

General Chemistry Chapter 1 – Matter and Measurement

Matter

Elements (Na, Si, H₂, S₈) vs. compounds (CO₂, NaCl, NaHCO₃)

Pure substances vs. mixtures (more than one element and/or compound)

Homogeneous vs. heterogeneous mixtures

Chemical vs. physical properties and changes

Intensive vs. extensive properties

Precision vs. accuracy

Units of Measurement

Convert 1500g to kilograms.

Convert 3200m to miles (1in = 2.54cm, 1mile = 5280ft).

A hose fills your swimming pool at a rate of 1,000.0 gallons per hour. How much is this in mL per second? (1 gal=3.785L)

Density calculations (d=m/v)

A cube with 2.00cm edges weighs 10.0 grams. What is its density?

How much would a cube of the same substance with 5.00cm edges weigh?

Significant Figures

Zeros in between significant figures are significant

Zeros right of the decimal, at the end are significant

-Multiplication/division vs addition/subtraction

SI Units

Length	m
Mass	kg
Time	s
Temperature	K

Derived Units

Volume	m ³
Energy/Work	J
Pressure	Pa

Terra	10 ¹²
Giga	10 ⁹
Mega	10 ⁶
kilo	10 ³
centi	10 ⁻²
milli	10 ⁻³
micro	10 ⁻⁶
nano	10 ⁻⁹
pico	10 ⁻¹²
femto	10 ⁻¹⁵

General Chemistry Chapter 2 – Atoms, Molecules, and Ions

Atomic Structure - nucleus composed of protons and neutrons surrounded by an electron cloud

Dalton proposed his atomic theory.

Thompson discovered electrons.

Millikan determined the mass of an electron (oil drop experiment).

Rutherford discovered the nucleus (scattering of alpha particles in thin gold foil).

Isotope symbols (atomic number, mass number, isotopes)

Atomic Weights

Chlorine has two abundant isotopes; 75% have a mass of 35a.m.u, 25% have a mass of 37 a.m.u. What is the atomic mass (also called atomic weight) of chlorine?

Periodic Table

Periods vs. groups

metals, non-metals, and metalloids

Group names

Types of Compounds

Ionic compounds (metal w/nonmetal or involving polyatomic ions)

Molecular compounds (two nonmetals)

Acids (H with nonmetal or oxyanion)

Nomenclature

Ionic compounds

- 1) Name the metal (or polyatomic cation)
- 2) State its oxidation state as a roman numeral in parenthesis (except Group I/II, Al, Zn, Cd, Ag)
- 3) Name the non-metal with an -ide ending (or name the polyatomic anion)

Molecular compounds

- 1) Give the numerical prefix of the first element (omit if there is only one)
- 2) Name the first element
- 3) Give the numerical prefix for the second element
- 4) Name the second element with the -ide ending

Polyatomic Ions

SO₄²⁻ sulfate
SO₃²⁻ sulfite
NO₃⁻ nitrate
NO₂⁻ nitrite
PO₄³⁻ phosphate
CO₃²⁻ carbonate
HCO₃⁻ bicarbonate
OH⁻ hydroxide
MnO₄⁻ permanganate
CrO₄²⁻ chromate
Cr₂O₇²⁻ dichromate
NH₄⁺ ammonium
CN⁻ cyanide

ClO⁻ hypochlorite
ClO₂⁻ chlorite
ClO₃⁻ chlorate
ClO₄⁻ perchlorate
(for Br and I also)

Acids

1) Binary Acids (only hydrogen and one other element)

Hydro-*element*-ic acid

2) Oxyacids

HNO ₂	nitrous acid	HClO	hypochlorous acid
HNO ₃	nitric acid	HClO ₂	chlorous acid
		HClO ₃	chloric acid
		HClO ₄	perchloric acid

General Chemistry Chapter 3 – Stoichiometry

Common Types of Reactions

Combination,
Decomposition
Combustion

What are the products of the combustion of C₃H₈? Balance the reaction.

Write a balanced reaction for the combination reaction between elemental aluminum and fluorine.

Calculations with Moles

What is a mole?

You have 12g of carbon (diamond). How many moles of carbon do you have? How many atoms?

You have 196g of H₂SO₄.

How many moles of H₂SO₄ do you have?

How many oxygen atoms are in the sample?

How many grams of hydrogen atoms are in the sample?

Mass Percentages and Molecular/Empirical Formulas

What is the percentage of sulfur in your sample of H_2SO_4 ?

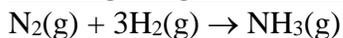
What is the %O?

What is the empirical formula of C_6H_6 ?

A certain compound is 80.0% carbon and 20.0% hydrogen by mass. What is its empirical formula?

If the compound above has a molecular weight of 30a.m.u., what is its molecular formula?

Limiting Reagent Calculations



How many moles of hydrogen gas are required for 6 moles of nitrogen gas to react completely?

How many moles of hydrogen gas are required for 56 grams of nitrogen gas to react completely?

If 128 grams of nitrogen and 30.0 grams of hydrogen react completely, which is the limiting reagent?

How much NH_3 is formed?

How much of the reagent in excess is left over?

General Chemistry Chapter 4: Aqueous Reactions and Solution Stoichiometry

A solution is composed of a solvent and one or more solutes.

The solvent is a liquid and is present in a higher concentration than any of the solutes.

The solute is dissolved in the solvent.

Electrolytes dissolve in water to form ions making the solution electrically conductive. They include ionic compounds, acids, and bases.

1) *Strong electrolytes* dissociate into ions completely.

a) soluble ionic compounds

b) strong acids (HCl, HBr, HI, HClO₄, HClO₃, H₂SO₄, HNO₃)

c) strong bases (Group I metal hydroxides, Ba(OH)₂, Sr(OH)₂, Ca(OH)₂.)

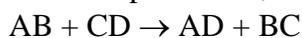
2) *Weak electrolytes* dissociate into ions only partially.

weak acids and bases (NH₃)

3) *Nonelectrolytes* can dissolve in water but don't form ions.

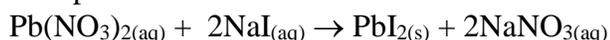
molecular compounds (such as C₆H₁₂O₆ and CH₃OH)

Double Displacement Reactions (also called double replacement, exchange, or metathesis reactions)



3 types

1) Precipitation reactions

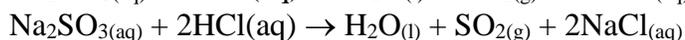
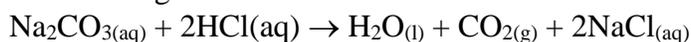


Complete Equation

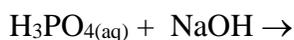
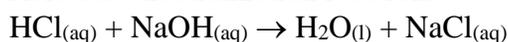
Ionic Equation

Net Ionic Equation

2) Gas-forming reactions



3) Acid base neutralization reactions



Concentration

Concentration is often measured in units of molarity (M)

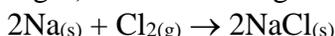
$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}} \quad M = \frac{\text{mol}}{V_L} \quad (M)(V_L) = \text{mol} \quad V_L = \frac{\text{mol}}{M}$$

What volume of 0.15 M H₂SO₄ is required to neutralize 0.50 L of 0.40 M NaOH?

Dilutions ($M_1V_1=M_2V_2$ or $C_1V_1=C_2V_2$)

Oxidation-Reduction Reactions (Redox Reactions)

A reaction in which the oxidation states (charges) of atoms change from reactants to products



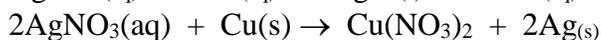
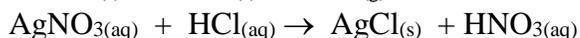
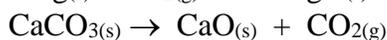
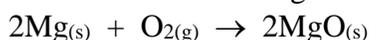
oxidation is the loss of electrons (OIL) (LEO)

reduction is the gain of electrons (RIG) (GER)

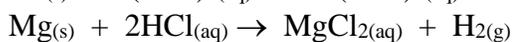
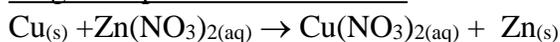
Determining Oxidation States of atoms

1. Atoms in elemental form: 0
2. Monatomic Ions in ionic compounds: ion charge
3. Atoms in Molecular Compounds:
 - Oxygen -2 except in peroxide O₂²⁻
 - Hydrogen +1 (when bonded to nonmetals)
 - Fluorine -1 always
 - Halogens -1 except when bonded to Oxygen
4. The oxidation numbers sum to zero for a compound or to the charge of an ion.

Which of the following are oxidation-reduction reactions?



Single Displacement Reactions



Which of the two reactions above are spontaneous?

<u>Activity Series</u>	
Li	↑ More Easily Oxidized
K	
Ba	
Ca	
Na	
Mg	
Al	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
H	
Cu	
Ag	
Hg	
Pt	
Au	

General Chemistry Chapter 5 – Thermochemistry

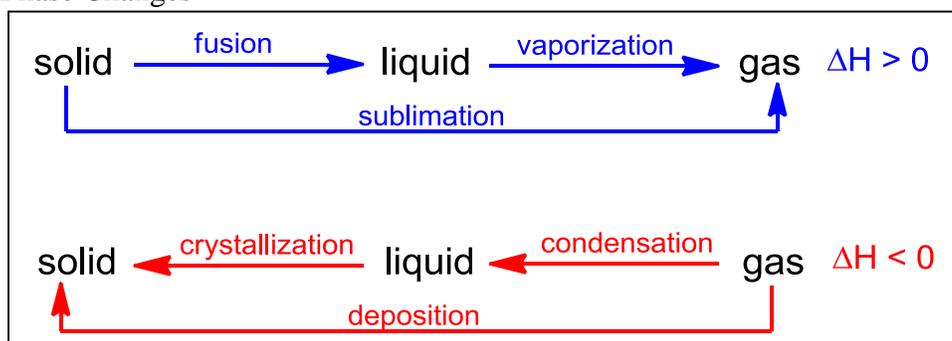
1st Law of Thermodynamics

$$\Delta E = q + w \quad \text{PV work (} w = -P\Delta V \text{)}$$

Enthalpy (H; a state function)

Endothermic ($\Delta H > 0$) vs. Exothermic reactions ($\Delta H < 0$)

Phase Changes



Calorimetry ($q = sm\Delta T$)

$$q = C\Delta T$$

$$q = mC_s\Delta T \quad m = \text{mass}$$

C = heat capacity

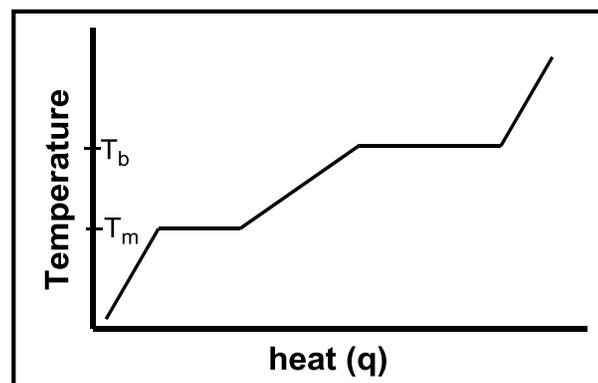
$$\Delta T = T_f - T_i$$

C_s = specific heat

$$\Delta T = T_f - T_i$$

-heat of fusion

-heat of vaporization



If an object absorbs 1000J of heat and its temperature increases 5 degrees, what is its heat capacity?

How much heat would be required to raise 90g of ice at -20°C to 30°C?

$$C_{s,\text{ice}} = 2.03\text{J/g}\cdot^\circ\text{C}$$

$$C_{s,\text{liquid water}} = 4.18\text{J/g}\cdot^\circ\text{C}$$

$$\Delta H^\circ_{\text{fus}} = 6.01\text{kJ/mol}$$

$$C_{s,\text{ice}} = 2.03\text{J/g}\cdot^\circ\text{C}$$

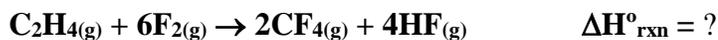
$$C_{s,\text{liquid water}} = 4.18\text{J/g}\cdot^\circ\text{C}$$

$$\Delta H^\circ_{\text{fus}} = 6.01\text{kJ/mol}$$

Hess's Law



$2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$	$\Delta H^\circ = -113\text{kJ}$
$2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{N}_2\text{O}(\text{g})$	$\Delta H^\circ = +163\text{kJ}$
$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$	$\Delta H^\circ = +181\text{kJ}$



$\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightarrow 2\text{HF}(\text{g})$	$\Delta H = -537\text{kJ}$
$\text{C}(\text{s}) + 2\text{F}_2(\text{g}) \rightarrow \text{CF}_4(\text{g})$	$\Delta H = -680\text{kJ}$
$2\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g})$	$\Delta H = +52.3\text{kJ}$

Enthalpy of Formation

$$\Delta H^\circ_{\text{rxn}} = \sum n\Delta H^\circ_{\text{products}} - \sum m\Delta H^\circ_{\text{reactants}}$$

Given the following standard enthalpy of formation values, calculate the $\Delta H^\circ_{\text{rxn}}$ of the following reaction:



ΔH_f° (kJ/mol)	
$\text{N}_2\text{H}_4(\text{g})$	+95.4kJ
$\text{H}_2\text{O}(\text{g})$	-242kJ

Formation Reactions

- 1) Form 1 mole of a single product
- 2) Reactants are elements in their standard states

STANDARD STATES

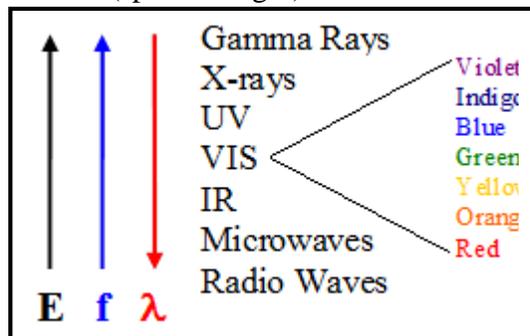
Gases	He, Ne, Ar, Kr, Xe, Rn, H ₂ , N ₂ , O ₂ , F ₂ , Cl ₂
Liquids	Br ₂ , Hg
Solids	Everything Else
Special	C(graphite), S ₈

General Chemistry Chapter 6 – Electronic Structure

Electromagnetic Radiation (spectrum)

$$\lambda \nu = c \quad \lambda = \text{wavelength} \quad \nu = \text{frequency} \quad c = 3.0 \times 10^8 \text{ m/s (speed of light)}$$

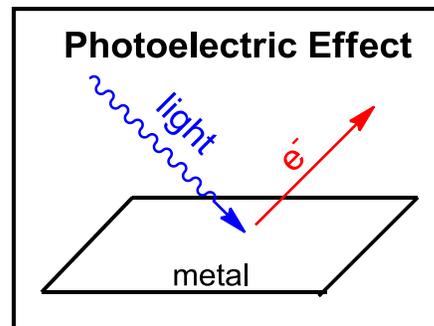
$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda} \quad h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s (Planck's constant)}$$



Photoelectric Effect

$$K.E._{\text{electron}} = E_{\text{photon}} - \phi$$

ϕ = work function i.e. the minimum energy to ionize an electron
1 electron per photon (if it is above the threshold)



Bohr Model of the Atom

-Quantized

Electronic Transitions

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$$R_H = 1.096776 \times 10^7 \text{ m}^{-1}$$

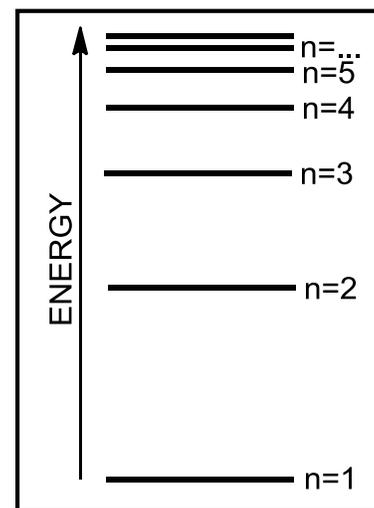
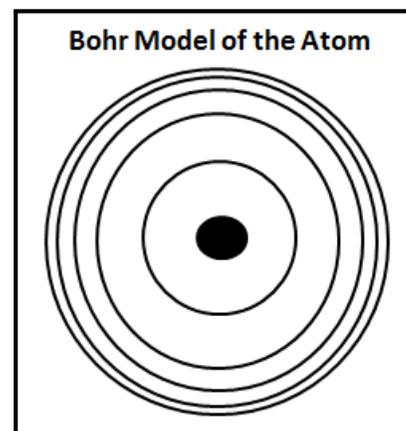
$$E_n = \left(\frac{-2.18 \times 10^{-18} \text{ J}}{n^2} \right)$$

$$n = 1, 2, 3, \text{ etc.}$$

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$\Delta E = E_{\text{photon}}$$

Lyman ($n_f = 1$), **Balmer** ($n_f = 2$), Paschen ($n_f = 3$), Bracket ($n_f = 4$) series



Wave Behavior of Matter (de Broglie Relation)

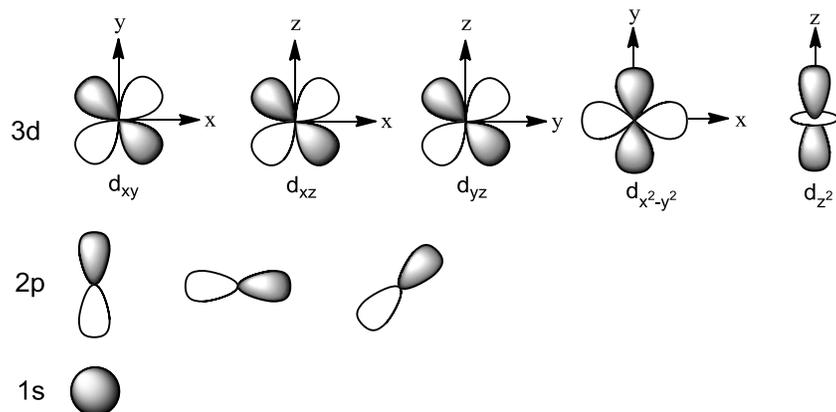
$$\lambda = \frac{h}{mv} \quad m = \text{mass} \quad v = \text{velocity} \quad \lambda = \text{wavelength} \quad h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$$

Uncertainty Principle

Impossible to simultaneously know an electron's *position* (x) and *momentum* (mv) with infinite precision.

$$\Delta x \Delta p \geq \text{constant}$$

Shapes of s, p, and d orbitals



Quantum Numbers

Pauli Exclusion Principle – no two electrons in an atom can have the same set of four quantum numbers

	Name	What it means	Range
n	Principal	Shell #	$[1 \dots \infty]$
l	Azimuthal	Subshell (type of orbital) $0=s; 1=p; 2=d; 3=f$	$[0 \dots (n-1)]$
m_l	Magnetic	Orientation (specifies a single orbital)	$[-l \dots +l]$
m_s	Spin	Spin up or spin down	$+1/2$ or $-1/2$

Ground State Electron Configurations

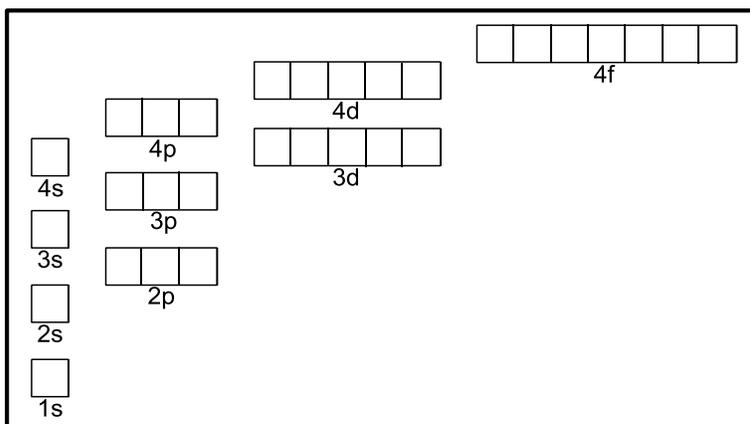
Standard and Noble Gas configurations

-Exceptions (Cr, Mo and Cu, Ag, Au)

-Ions (transition metals too)

-Number of valence electrons

-Ground state vs. excited state



Aufbau Principle – electrons fill lowest energy orbitals first

Hunds Rule – degenerate orbitals each get an electron before pairing

Pauli Exclusion Principle – no 2 electrons in an atom have the same 4 quantum numbers

General Chemistry Chapter 7 – Periodic Properties of the Elements

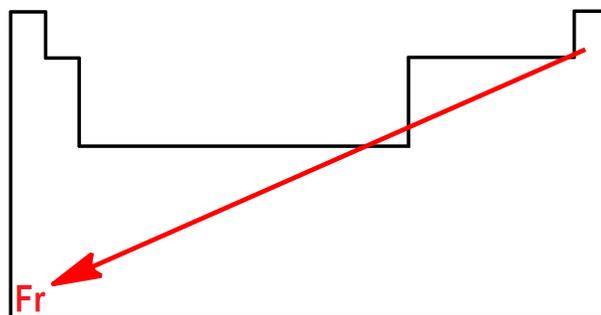
Atomic Size (Radius)

Effective Nuclear Charge

Bond length

Ionic radii

Isoelectronic series

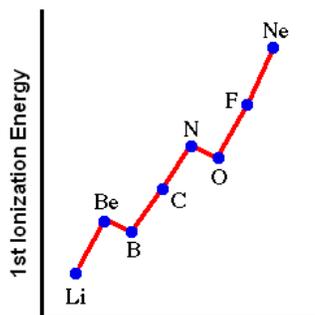
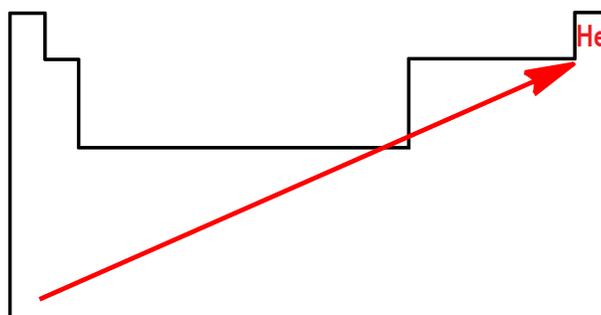


Ionization Energy

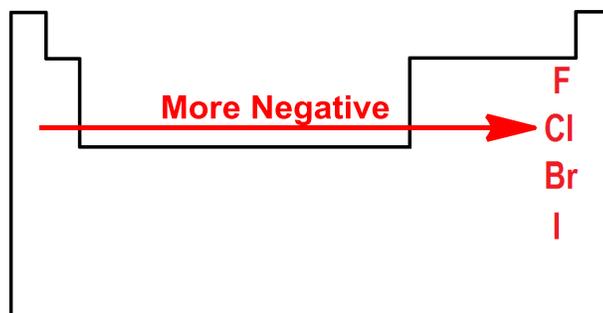
-Energy required to remove an electron

2nd, 3rd, 4th, etc. ionization energies

Irregularities in the trend



Electron Affinity – Energy change of gaining an electron



Alkali metals are highly reactive with water (due to low ionization energies).

Alkaline Earth metals are not nearly as reactive with water (increasing reactivity going down).

Halogens have high (very negative) electron affinities.

Metal oxides are basic (examples: Na_2O , BaO)

Non-metal oxides are acidic (examples: NO_2 , SO_3 , CO_2)

General Chemistry Chapter 8 – Bonding and Lewis Structures

Bonding

Ionic – metal with nonmetal or involving polyatomic ions (electrons are transferred)
forms crystalline solids

Covalent – non-metal with non-metal (electrons are shared)
forms molecular compounds and network covalent solids

Metallic – metal with metal

Octet Rule

Octet Rule exceptions

- 1) Third row and lower can exceed the octet rule
- 2) H only 2, Be only 4, B and Al only 6 (usually)
- 3) Molecules with odd numbers of electrons

Ionic Compounds (high melting and boiling pts, brittle)
lattice energy – increases with higher charges and smaller sizes

Molecular (covalent) Compounds

Bond polarity and electronegativity

Lewis structures

Drawing Lewis Structures

- 1) Sum up the valence electrons.
- 2) The atom that can make the most bonds is the central atom.
- 3) Fill up the outer atoms first.
- 4) Place any remaining electrons on the central atom.
- 5) If the central atom isn't satisfied then use the outer atoms electrons to make double/triple bonds to the central atom.

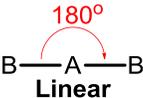
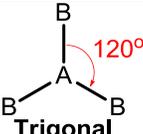
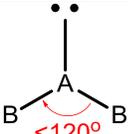
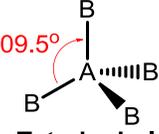
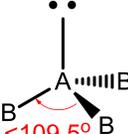
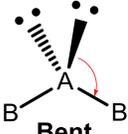
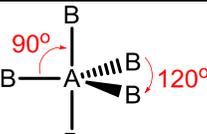
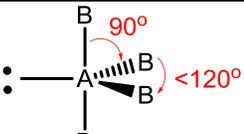
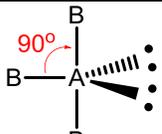
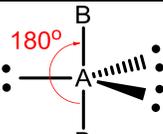
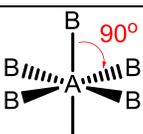
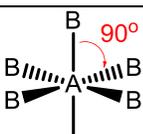
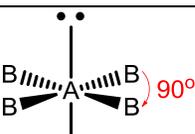
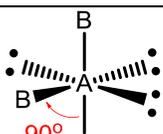
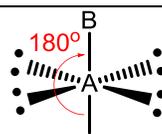
Resonance (electron delocalization)

Effect on bond length

General Chemistry Chapter 9 – Molecular Geometry and Bonding

VSEPR Theory (Valence Shell Electron Pair Repulsion)

Electron Domains	Electron Domain Geometry	Bond Angles	Hybridization of Central Atom
2	Linear	180°	sp
3	Trigonal Planar	120°	sp ²
4	Tetrahedral	109.5°	sp ³
5	Trigonal Bipyramidal	90°, 120°, 180°	sp ³ d
6	Octahedral	90°, 180°	sp ³ d ²

Electron Domains	No lone pairs	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 <p style="text-align: center;">Linear</p>				
3	 <p style="text-align: center;">Trigonal Planar</p>	 <p style="text-align: center;">Bent</p>			
4	 <p style="text-align: center;">Tetrahedral</p>	 <p style="text-align: center;">Trigonal Pyramidal</p>	 <p style="text-align: center;">Bent</p>		
5	 <p style="text-align: center;">Trigonal Bipyramidal</p>	 <p style="text-align: center;">See-saw (Sawhorse)</p>	 <p style="text-align: center;">T-shaped</p>	 <p style="text-align: center;">Linear</p>	
6	 <p style="text-align: center;">Octahedral</p>	 <p style="text-align: center;">Square Pyramidal</p>	 <p style="text-align: center;">Square Planar</p>	 <p style="text-align: center;">T-shaped</p>	 <p style="text-align: center;">Linear</p>

How do lone pairs affect bond angles?

Polarity (Molecular)

Valence-Bond Theory – electrons are shared between two atoms where their orbitals overlap

Sigma and Pi bonds

Hybridization

General Chemistry Chapter 10 – Gases

Pressure

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}} \quad \left(P = \frac{F}{A} \right)$$

$$1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg}$$

Ideal Gas Theory (Kinetic Molecular Theory)

- Assumptions: 1) Gas molecules don't have volume.
2) Gas molecules don't have any attractive forces between themselves.
(i.e. all collisions are elastic.)

Boyle's Law $PV = \text{constant}$ or $P_1V_1 = P_2V_2$

Charles' Law $\frac{V}{T} = \text{constant}$

Avogadro's Law $\frac{V}{n} = \text{constant}$

Combined Gas Law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ Better Combined Gas Law $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$

Ideal Gas Law

$$PV = nRT \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

Dalton's Law of Partial Pressures

$$P_{\text{tot}} = P_A + P_B + P_C + \dots$$

$$P_A = \chi_A P_{\text{tot}} \quad \chi_A = \text{mole fraction A}$$

Gas density and molar mass $d = \frac{m}{V} = \frac{PM}{RT}$ $M = \text{molar mass}$

Molar volume at STP (1 atm and 273K) is 22.4L / mole gas

Graham's Law (Effusion, Diffusion) $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

Ideal Behavior

- 1) Gases behave more ideally at low Pressure and high Temperature.
2) Gases with less Intermolecular Forces tend to behave more ideally.

Real Gases

Van der Waals equation: $\left(P + \frac{an^2}{V^2} \right) (V - nb) = nRT$

General Chemistry Chapter 11 – Intermolecular Forces and Phase Diagrams

Intermolecular Forces

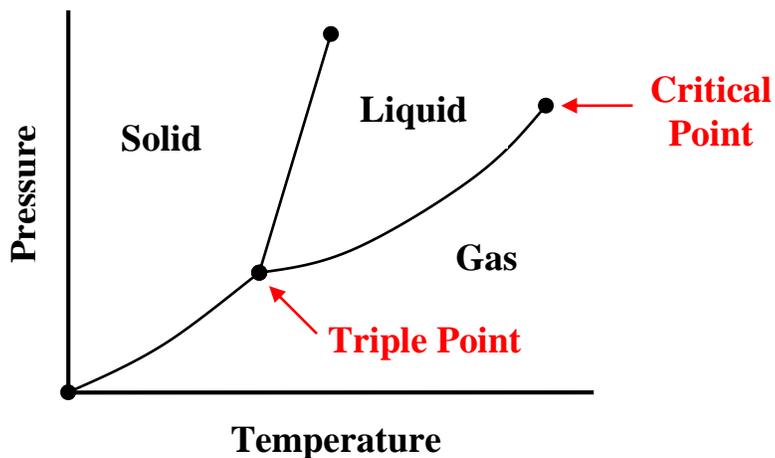
- 1) Hydrogen Bonding – a super strong dipole-dipole force
 - must have hydrogen bound to F, O, N (F-H, O-H or N-H bond) to Hydrogen Bond as a pure liquid
 - must only have F, O, N to hydrogen bond with water
- 2) Dipole-Dipole Forces – interaction between molecules having permanent dipole moments
 - the larger the dipole moment, the larger the force
- 3) London Dispersion Forces – weak interactions due to a transient (temporary) dipole
 - all molecules have these
 - the larger the molecule, the larger the force
- 4) Ion-Dipole Forces--forces between ions and polar liquids

Higher IM Forces → Higher Boiling Pt, Higher Viscosity, Higher Surface Tension, Lower Vapor Pressure

“Like dissolves like.”

Phase Changes and Phase Diagrams

- lines of equilibrium, triple point, critical point, normal m.p. and b.p.
- phase diagrams for CO₂ and H₂O



General Chemistry Chapter 12 – Colligative Properties

$$\text{molality} = \frac{\text{mol}_{\text{ solute}}}{\text{kg}_{\text{ solvent}}}$$

van't Hoff factor (i) – number of ions a compound dissociates into per formula unit

Colligative Properties

1) Freezing Pt. Depression

$$\Delta T_F = -iK_F m$$

2) Boiling Pt. Elevation

$$\Delta T_B = iK_B m$$

3) Vapor Pressure Depression (Raoult's Law)

$$P_A = \chi_A P_{\text{Total}}$$

4) Osmotic Pressure

$$\pi = iMRT$$

General Chemistry Chapter 13 – Kinetics

Rate expressions



$$\frac{-\Delta[\text{NH}_3]}{2\Delta t} = \frac{\Delta[\text{N}_2]}{\Delta t} = \frac{\Delta[\text{H}_2]}{3\Delta t}$$

Rate laws

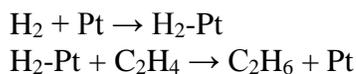
$$\text{rate} = k[\text{A}]^x[\text{B}]^y$$

For the reaction $2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g})$
the following results were obtained:

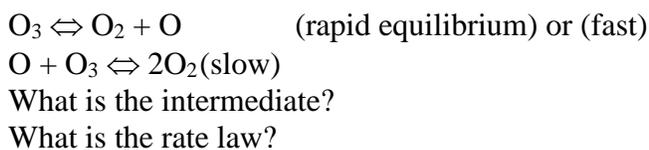
[NO] ₀ (mol/L)	[Cl ₂] ₀ (mol/L)	Initial Rate (mol/L·s)
0.10	0.10	0.18
0.10	0.20	0.36
0.20	0.20	1.44

What is the rate law for this reaction?

What is the rate constant?



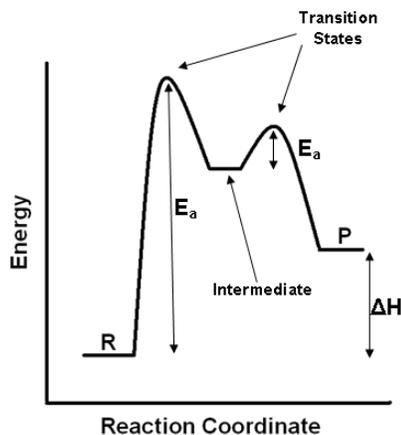
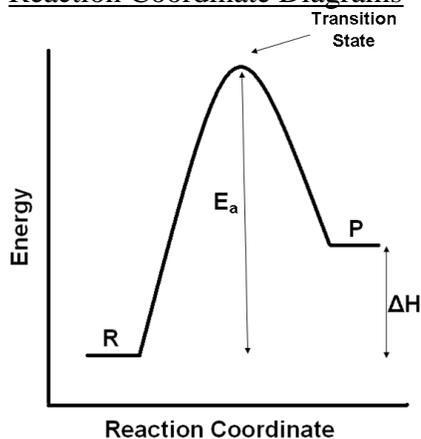
Identify the intermediate and the catalyst.



Catalyst – speeds up a reaction by lowering the activation energy by providing an alternate mechanism (pathway) for the reaction to occur

A catalyst lowers the E_a in both directions, is not consumed in a rxn, and does not shift the equilibrium.

Reaction Coordinate Diagrams



Collision Theory (Requirements for a chemical reaction)

- 1) Collision
- 2) Proper orientation of reactant molecules
- 3) Sufficient energy for a reaction to occur

An increase in temperature increases collision frequency and the percentage of high energy collisions.

Integrated Rate Laws

Zero Order $[A] = [A]_0 - kt$

1st Order $\ln[A] = \ln[A]_0 - kt$

2nd Order $1/[A] = 1/[A]_0 + kt$

Half-Life

1st order $t_{\frac{1}{2}} = \frac{0.693}{k}$

The rate constant (k) increases with increasing temperature and decreasing activation energy (by adding a catalyst).

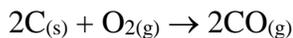
$$k = Ae^{-E_a/RT}$$

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

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

General Chemistry Chapter 14--Equilibrium

Equilibrium constants (Law of Mass Action)



What is the equilibrium constant expression for the reaction above?

$K \gg 1$	Favors products
$K \ll 1$	Favors reactants

Equilibrium constants are CONSTANTS. The constant will only change with a change in temperature.

$$K_p = K_c (RT)^{\Delta n} \quad \Delta n = \text{change in moles gas} \quad R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$$



Based on the reaction above, what are the equilibrium constants for the following reactions?



Le Chatelier's Principle

If a stress is placed on a system at equilibrium, the system will respond to counteract the stress.

Q	Result
$Q < K$	shift right
$Q > K$	shift left
$Q = K$	at equilibrium



If 2 moles of both H_2 and Cl_2 are each placed in a 2L vessel, what will be the equilibrium concentrations of H_2 , Cl_2 , and HCl ?



0.1 moles of HCN are added to a liter of solution. What is the equilibrium concentration of H^+ in the solution?

General Chemistry Chapter 15--Acids and Bases

Definitions of acids and bases

	Acid	Base
Arrhenius	H ⁺ donor in H ₂ O	OH ⁻ donor in H ₂ O
Bronsted-Lowry	H ⁺ donor	H ⁺ acceptor
Lewis	Electron acceptor	Electron donor

Strong acid or base – one that dissociates or ionizes or hydrolyzes completely in water

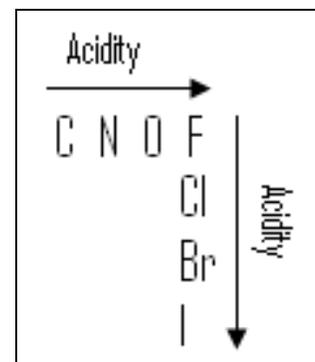
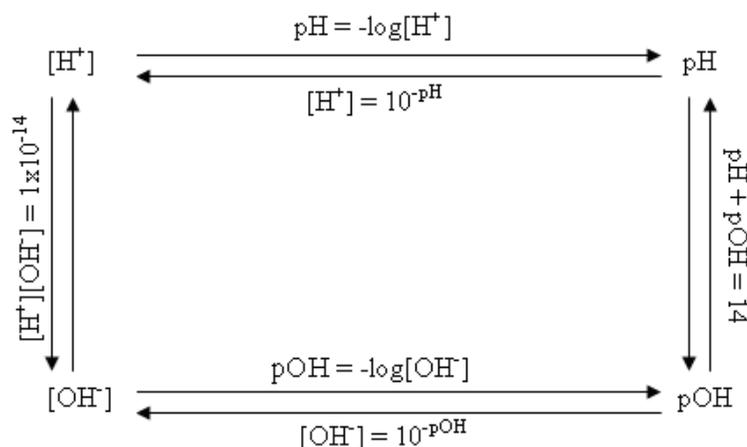
Strong acids	Strong Bases
HClO ₄	Group 1 metal hydroxides
H ₂ SO ₄	Ba(OH) ₂
HI	Sr(OH) ₂
HBr	Ca(OH) ₂ to solubility limit
HCl	
HNO ₃	
HClO ₃	

Binary acid trend

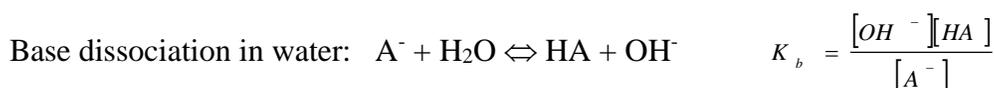
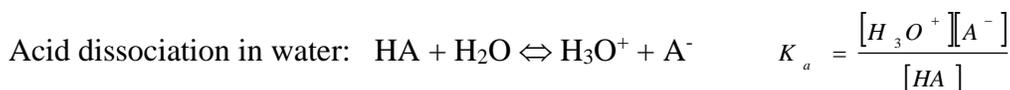
Oxoacid trends

- 1) More oxygens, more acidic
- 2) More electronegative heteroatom, more acidic

Polyprotic Acids (like H₂SO₄ and H₃PO₄)



pH, pOH, [H⁺], [OH⁻];
if you know 1 of the 4, you
can calculate the other 3.



$K_a K_b = K_w = 1 \times 10^{-14}$ (If you know K_a , you know K_b and vice versa)
The stronger the acid, the weaker the conjugate base and vice versa

Calculations involving K_a and K_b

a) What is the pH of a 0.1M solution of acetic acid ($K_a = 1.76 \times 10^{-5}$)?

b) What is the K_a of a weak acid, HA, if the pH of a 0.1M solution is 3?

c) What is the percent dissociation of the acid in part b?

d) What is the pH of a 0.1M solution of ammonia ($K_b = 1.75 \times 10^{-5}$)?

e) What is the K_b of a weak base, A^- , if the pH of a 0.1M solution is 10.2?

pH of Salts

<u>Neutral Cations</u>	<u>Neutral Conjugate Bases</u>
Li^+	Cl^-
Na^+	Br^-
K^+	I^-
Rb^+	NO_3^-
Cs^+	ClO_4^-
Ca^{2+}	ClO_3^-
Sr^{2+}	
Ba^{2+}	
All other cations are acidic; the more positively charged, the more acidic.	Acidic Conjugate <u>Base</u> HSO_4^-
	Almost all other anions are bases.

Other Acid/Base Equilibria

Buffer – A solution that resists changes in pH. It is composed of a weak acid and its conjugate base or a weak base and its conjugate acid.

Calculations with buffers

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

What is the pH of a solution that is 0.1M in HA and 0.075M in NaA ($K_a = 2.5 \times 10^{-6}$)?

Problem Type	
Strong Acid	$pH = -\log[H^+]$
Strong Base	$pOH = -\log[OH^-]$
Weak Acid	Ka / ICE chart
Weak Base	Kb / ICE chart
Buffer	$pH = pK_a + \log \frac{[A^-]}{[HA]}$

Titration

100mL of 0.125M NaOH is added to 150mL of 0.10M HCl. What is the pH of the resulting solution?

100mL 0.10M HF ($K_a = 6.8 \times 10^{-4}$) is titrated with 0.1M NaOH. Calculate the pH after the addition of 25, 50 and 100ml NaOH.

General Chemistry Chapter 16--Solubility Equilibria

General Solubility

a) What is the molar solubility of AgCl ($K_{sp} = 1.8 \times 10^{-10}$)?

b) What is the molar solubility of $Mg(OH)_2$ ($K_{sp} = 1.6 \times 10^{-12}$)?

c) The molar solubility of BiI_3 is 1.32×10^{-5} . Calculate K_{sp} .

Common Ion Effect

What is the molar solubility of AgCl in 0.1M HCl? ($K_{sp} = 1.8 \times 10^{-10}$)

Precipitation

To a 0.0001M solution of $Mg(NO_3)_2$, NaOH was added to a final concentration of 0.001M. Did a precipitate form? ($K_{sp} Mg(OH)_2 = 1.6 \times 10^{-12}$)

Increased solubility in acids/bases or other reagents

Acidic salts are more soluble in base, less soluble in acid.

Basic salts are more soluble in acid, less soluble in base.

Neutral salts' solubility is largely unaffected by a change in pH.

General Chemistry Chapter 17--Thermodynamics

Three laws of thermodynamics

- 1) Conservation of energy (Energy can't be created or destroyed).
- 2) For a spontaneous process, the entropy of the universe increases.
- 3) A perfect crystal at 0K has zero entropy.

Entropy (disorder or randomness)

For what processes are there significant increases or decreases in entropy?

Gibbs Free Energy – the energy available to do work

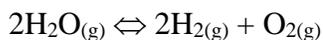
$\Delta G < 0$	Spontaneous
$\Delta G > 0$	Nonspontaneous
$\Delta G = 0$	At equilibrium

$$\Delta G = \Delta H - T\Delta S$$

ΔH	ΔS	$-T\Delta S$	
-	+	-	Spontaneous at all temperatures
+	-	+	Non-spontaneous at all temperatures
-	-	+	Spontaneous at low temperatures
+	+	-	Spontaneous at high temperatures



Is this reaction spontaneous at all temps, no temp, low temp, or high temps?



The reaction above is only spontaneous at temperatures above 2000K. Is the reaction endothermic (+ ΔH) or exothermic (- ΔH)?

Calculating ΔG° , ΔH° , and ΔS°

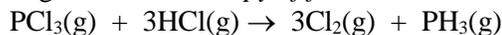
$$\Delta G^\circ = \sum n\Delta G^\circ_{f, \text{products}} - \sum n\Delta G^\circ_{f, \text{reactants}}$$

$$\Delta H^\circ = \sum n\Delta H^\circ_{f, \text{products}} - \sum n\Delta H^\circ_{f, \text{reactants}}$$

$$\Delta S^\circ = \sum nS^\circ_{\text{products}} - \sum nS^\circ_{\text{reactants}}$$

Enthalpies of Formation

Given the following *standard enthalpy of formation* values, calculate the ΔH°_{rxn} of the following reaction:



<u>Compound</u>	<u>ΔH_f (kJ/mol)</u>
PCl ₃ (g)	-288.07
HCl(g)	-92.30
PH ₃ (g)	5.40

Miscellaneous Mathematical Relationships

$$\Delta G = \Delta G^\circ + RT \ln Q \quad (\text{Free energy under non-standard conditions})$$

$$\Delta G^\circ = -RT \ln K_{eq}$$

$\Delta G^\circ < 0$	$K_{eq} > 1$
$\Delta G^\circ > 0$	$K_{eq} < 1$
$\Delta G^\circ = 0$	$K_{eq} = 1$

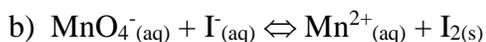
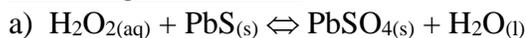
General Chemistry Chapter 18--Electrochemistry

<u>Oxidation</u> – loss of electrons	LEO	OIL
<u>Reduction</u> – gain of electrons	GER	RIG

Oxidation States

- 1) Elements in their elemental form are in the zero oxidation state.
- 2) Group 1 metals are +1 and Group 2 metals are +2.
- 3) Hydrogen is +1 except when bonded to metals (when it's -1).
- 4) Transition elements must be determined by context (except Al=+3, Zn=+2, Cd=+2, Ag=+1)
- 5) The most electronegative elements get their typical oxidation state.
- 6) The last element not assigned balances the charge of the compound/ion.

Balancing Redox Rxns



The two ways these questions are typically asked:

- 1) What is the coefficient for a specific compound or what is the sum of the coefficients? (Don't forget to include the implied coefficient of 1 in the sum)
- 2) Which of the following is/isn't balanced correctly? (First check to make sure charge is balanced as it is in all properly balanced reactions)

To find what's changing oxidation states narrow the search by first looking for elements in their zero oxidation state. This is helpful for identifying redox reactions. Once you find an element that has been oxidized or reduced, you know the rxn is a redox rxn.

Oxidizing Agent (Oxidant) – species that is reduced

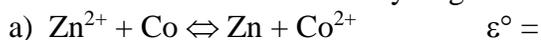
Reducing Agent (Reductant) – species that is oxidized

Electrochemical Cells

- a) The anode is always the site of oxidation (an ox)
- b) The cathode is always the site of reduction (red cat)
- c) Electrons always flow from anode to cathode
- d) Anions flow to the anode and cations to the cathode through the salt bridge
- e) In the galvanic/voltaic (spontaneous) cells, the cathode is + and the anode is –
The signs are the opposite in electrolytic (non-spontaneous) cells.
- f) For metal/metal salt solution cells, the cathode gains mass, while the anode loses mass.

Standard Cell Potentials ($\epsilon^\circ = \epsilon^\circ_{\text{cat}} - \epsilon^\circ_{\text{an}}$ $\epsilon^\circ = \epsilon^\circ_{\text{red}} + \epsilon^\circ_{\text{ox}}$)

Remember that the standard hydrogen electrode (SHE) has a potential of 0.00V



- c) What is the balanced rxn for the galvanic cell composed of Mn/Mn²⁺ and Fe/Fe²⁺ half cells?

Non-standard Cell Potentials $E = E^\circ - \frac{0.0592}{n} \log Q$ $E = E^\circ - \frac{0.0257}{n} \ln Q$ (Nernst Equation)

- a) Calculate the cell potential for the cell in part a) (last problem) with $[Zn^{2+}] = 0.001M$ and $[Co^{2+}] = 0.1M$
- b) Know the qualitative effects that will result in a higher/low potential.
- 1) Shift to the right – higher potential
 - 2) Shift to the left – lower potential

Reduction potentials

- a) What is the strongest oxidizing agent?
- b) What is the strongest reducing agent?
- c) Which pairs will react spontaneously?

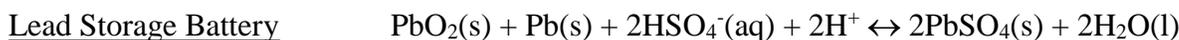
$Al^{3+} + 3e^- \rightarrow Al$	-1.66V
$Mn^{2+} + 2e^- \rightarrow Mn$	-1.18V
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76V
$Cr^{3+} + 3e^- \rightarrow Cr$	-0.74V
$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44V
$Co^{2+} + 2e^- \rightarrow Co$	-0.28V
$Ni^{2+} + 2e^- \rightarrow Ni$	-0.25V
$2H^+ + 2e^- \rightarrow H_2$	0V
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34V
$Ag^+ + 1e^- \rightarrow Ag$	+0.80V
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.07V

Spontaneous Rxns

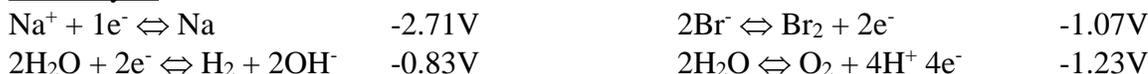
$$\Delta G < 0 \qquad \Delta G = -nFE$$

$$E_{cell} > 0 \qquad \Delta G^\circ = -RT \ln K$$

$$Q < K \qquad E^\circ = \frac{0.0592}{n} \log K$$



Electrolysis



- a) What are the products of electrolysis of $NaBr(l)$?
- b) What are the products of electrolysis of a solution of $NaBr$?

Electrolysis Calculations

$$\frac{(Amps)(T_s)(MW)}{(n)(F)} = \text{grams product}$$

$$\frac{(Amps)(T_s)}{(n)(F)} = \text{moles product}$$

- a) How many moles of Mn will be formed if $MnCl_2$ is electrolysed for 2 hours using 10 Amps?
- b) How long will it take to plate out 1.0kg of Al from molten $AlCl_3$ with a current of 100 A?

General Chemistry Chapter 19--Nuclear Chemistry

Understand the symbols used for nuclei. ${}_{92}^{235}\text{U}$

Determining which nuclei are radioactive (unstable):

- 1) Elements with $Z > 83$ are radioactive
- 2) Nuclei having even numbers of protons and/or neutrons are more stable
- 3) N/Z ratio – ratio of neutrons to protons ($N = \# \text{ neutrons}$, $Z = \# \text{ protons}$)
 - the lighter elements ($Z=1-20$) are stable with $N/Z = 1$
 - the heavier elements are stable with $N/Z \leq 1.6$ (a slight excess of neutrons)
- 4) Atoms with magic numbers of protons and/or neutrons are more stable (2,8,20,28,50,82,126)

Balancing Nuclear Reactions

Total mass and charge must be equal on both sides of the rxn arrow



Nuclear Reactions

Routes of Decay	Net Result	Who?
α-decay (${}^4_2\alpha$ is a product)	Reduces Mass	Large nuclei ($Z > 83$)
β-decay		
a) β^- decay (${}^0_{-1}\beta$ is a product)	n \rightarrow p	N/Z is too high (above belt)
b) Positron Emission (${}^0_{+1}\beta$ is a product)	p \rightarrow n	N/Z is too low (below belt)
c) Electron Capture (${}^0_{-1}\beta$ is a reactant)	p \rightarrow n	N/Z is too low (below belt)
γ-decay (${}^0_0\gamma$ is a product)		

Fission – the splitting of a heavy nucleus into lighter nuclei (and possibly nuclear particles)
 induced by neutron bombardment
 this is the rxn that takes place in nuclear reactors

Fusion – the combination of two or more lighter nuclei to form a heavier nucleus
 this occurs in H-bombs and in the sun's core and requires extremely high temps.

Kinetics of Nuclear Decay (First Order)

Radiocarbon Dating - Used to date objects that were once living and are less than ~50,000 years old. It is based on the decay of ${}^{14}\text{C}$ by β -emission which has a half life of 5730 years.

Nuclear Binding Energy – energy that holds the nucleus together – the nucleus always weighs less than the sum of its nucleons because some of the mass is converted into this nuclear binding energy

Mass defect – Difference between the actual mass of the nucleus and the mass of the constituent nucleons

Also is the difference in mass between the products and reactants of a nuclear rxn
 (the products will always weigh less—the mass lost is converted to energy)

To calculate nuclear binding energy: (1) calculate the mass defect, (2) convert to kg, (3) multiply by c^2 , and (4) divide by the number of nucleons (protons + neutrons)

mass defect (Δm) $E = \Delta mc^2$ Δm must be in kg 1 a.m.u. = 1.66×10^{-27} kg

${}^{56}\text{Fe}$ has the highest nuclear binding energy per nucleon and typically the closer in mass you are to 56, the higher the nuclear binding energy per nucleon.