

**Chemistry 1202 Review/Preview  
Chapter Five Review Guide**

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**Spring 2007**

**Thermochemistry**

**I. Definitions to Know and Understand**

System, Surroundings, Energy, Temperature, Heat, Potential Energy, Kinetic Energy, Work

**II. Heat and Energy**

- A. Potential Energy - the amount of stored energy.
- B. Kinetic Energy =  $1/2 mv^2$ , where m is in kilograms and v is in meters/second.  
The units of the energy value derived in this manner is the joule.

**III. Work**

- 1. **Work = Force x distance**, where force = pressure x area.  
The force due to gravity = mass x gravitational constant ( $9.8 \text{ m/s}^2$ ).

**IV. Internal Energy and the First Law of Thermodynamics**

- A. First Law of Thermodynamics - The total internal energy of an isolated system is constant.
- B. State Functions - Functions whose value depends only on the present state of the system, not on the path used to arrive at the state.

**V. Relating  $\Delta E$  to heat and work**

- A. The first law of thermodynamics is mathematically stated as follows:

$$\Delta E = q + w \quad \text{where } q \text{ is the heat exchanged and } w \text{ is the work done.}$$

The sign conventions for work and heat are as follows:

Work done on the system is positive; work done by the system is negative  
Heat entering the system is positive; heat going out of the system is negative

- B. Heat (q) and work (w) are two different forms of energy and are therefore equivalent and can be interconverted by using the appropriate conversion factor.

- C. Enthalpy - heat content (H)

Enthalpy, H, has the following relationship to the internal energy and the work:  $H = E + PV$ . The change in enthalpy,  $\Delta H$ , is related to the change in energy and the work by the following equation:

$$\Delta H = \Delta E + P\Delta V, \quad \text{and} \quad \Delta E = \Delta H - P\Delta V \quad (\Delta E = q + w)$$

$$\text{At constant pressure, } q_p = \Delta E + P\Delta V = \Delta H.$$

**VI. Enthalpies of Reaction**

The enthalpy of reaction ( $\Delta H$ ) is the difference in the enthalpies of the products and the enthalpies of the reactants in a given reaction.

$$\Delta H = H(\text{products}) - H(\text{reactants}) \quad \begin{array}{l} \Delta H < 0 \text{ for exothermic reactions} \\ \Delta H > 0 \text{ for endothermic reactions} \end{array}$$

Standard Enthalpy of Reaction ( $\Delta H_r$ )- the heat change associated with a reaction carried out under standard state conditions. See Section 5.4 of the text for general rules concerning enthalpies of reaction.)

**VII. Calorimetry and Heat Capacity**

- A. Calorimetry - a method of determining  $\Delta H$  for reactions by carrying out the reaction in an insulated vessel called a calorimeter. The fundamental calorimeter equation is

$$q = C \times \Delta T$$

where C = is the heat capacity (C = the heat absorbed/increase in temperature)

*Note: Heat capacity (C) is the heat required to raise the temperature of an entire sample by one degree celsius. The units are J/ C*

Specific heat capacity is the heat required to raise the temperature of one gram of a substance by one degree celsius.

$$\text{specific heat} = \frac{\text{heat transferred}}{\text{grams} \times \text{temperature change}}$$

The units of specific heat are J/gram x C

The molar heat capacity ( $C_m$ ) is the J/mole x C

To determine the heat change for a process: Joules = specific heat x grams x  $\Delta T$

B. Constant Pressure calorimetry is used to determine the heats of reactions for reactions occurring in solution.

C. Constant Volume calorimetry (bomb calorimetry)

The energy produced by a reaction is determined by the temperature increase (or decrease) of the water and other parts of the calorimeter. The calorimeter must be calibrated by determining the temperature increase associated with a reaction whose heat of reaction is known.

### VIII. Hess's Law:

Hess's Law states that when reactants are converted to products the enthalpy change is the same whether the process occurs in one step or more than one step. Hence, the heat of a particular reaction can be determined by adding the heats of reactions from two or more reactions whose combined net reaction equals the original reaction.

Notes about Hess's Law:

- i. When a reaction is written in reverse, the  $\Delta H$  of the reverse reaction is the negative of the  $\Delta H$  of the forward reaction.
- ii. If a reaction is multiplied by a certain factor, the new  $\Delta H$  is obtained by multiplying the original  $\Delta H$  by the same factor.

See examples 5.8 and 5.9 of the text.

### V. Standard Heats of Formation and Reaction

A. Standard Enthalpy of Formation ( $\Delta H_f$ )- the heat change associated with forming one mole of compound from its elements in their standard states.

B. Standard Enthalpy of Reaction ( $\Delta H_r$ )- the heat change associated with a reaction carried out under standard state conditions.

C. Calculating heats of reaction ( $\Delta H_r$ )

1. From the heats of formation of the components of the reaction:

$$\Delta H_r = \sum \Delta H_f \text{ products} - \sum \Delta H_f \text{ reactants}$$

(See  $\Delta H_f$  values in Appendix C of the text.)

2. By Using Hess's Law - See earlier discussion.

### VII. Foods and Fuels

See lecture notes from class for appropriate coverage

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**Chemistry 1202 Review/Preview**  
**Chapter Fourteen Review Guide**

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**Spring 2007**

**I. Factors that Affect Reaction Rates**

Many things determine how rapidly a chemical reaction will occur. The four factors that we can change that will speed up or slow down a reaction are the state of the reactants (solid, liquid, gas, powdered solid), the concentration of the reactants, the temperature at which the reaction is occurring, and whether or not there is a catalyst present. (Read section 14.1 for a description of how these factors influence the rates of chemical reactions.)

**II. Reaction Rates**

A. Reaction Rate is defined as the change in the concentration of a reactant or product with respect to time. For a general reaction  $A \rightarrow B$ ,

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

B. Reaction Rates and Stoichiometry

The coefficients in the balanced equation must be taken into consideration when discussing reaction rates. For example, in the reaction  $2A \rightarrow B$ ,

$$\text{rate} = -\frac{1}{2} \frac{\Delta[A]}{\Delta t} = \frac{\Delta[B]}{\Delta t}$$

This relationship could also be written

$$\text{rate} = -\frac{\Delta[A]}{\Delta t} = 2 \frac{\Delta[B]}{\Delta t}$$

**II. Concentration and Rate; Determination of a Rate Law**

A. The differential form of the rate law (usually just called the rate law) relates the rate of a reaction to the rate constant and the concentrations of the reactants.

For the general reaction  $aA + bB \rightarrow cC + dD$ ,

$$\text{rate} = k[A]^m[B]^n, \text{ where } m \text{ and } n \text{ are experimentally determined.}$$

To determine the value of  $m$  and  $n$ , consider the power to which the change in concentrations must be raised in order to obtain the rate change. The power to which the change in  $[A]$  must be raised is the value of  $m$ , and the power to which the change in  $[B]$  must be raised is the value of  $n$ . These values can be determined by "eyeballing" the table of initial rates versus concentration, as explained in Chem 1202 and the Chem 1202 review/preview session. See pages 582-587 of the text for examples of this.)

B. Order of a Reaction

The overall order of a chemical reaction is the sum of the powers to which all reactant concentrations appearing in the rate law are raised. The power to which a specific concentration is raised is the order of the reaction with respect to that reactant. The value of  $k$  can be determined by substituting experimental values of concentrations and rates into the rate law once it has been determined.

#### IV. The Change of Concentration with Time

The integrated rate law shows how the concentrations of species in the reaction depend on time.

##### A. Integrated Rate Law of First Order Reactions

The relationship between the concentrations and time for first order reaction is given by

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \ln[A]_t - \ln[A]_0 = -kt, \quad \text{or} \quad \ln \frac{[A]_t}{[A]_0} = -kt$$

The equation on the left has the form of the equation for a straight line ( $y = mx + b$ , where  $m$  is the slope and  $b$  is the y-intercept). *If we graph  $\ln[A]_t$  versus time, a first order reaction will yield a straight line where  $k = -\text{slope}$ .* See pages 587-588 of the text for more detail.

##### B. Half-Life of First Order Reactions

The half life,  $t_{1/2}$ , of a reaction is the amount of time required for the concentration of the reactant to decrease to half of its initial concentration. The expression relating half life of a first order reaction to the value of the rate constant for a first order reaction is

$$t_{1/2} = 0.693/k$$

##### C. Integrated Rate Law of Second Order Reactions

The relationship between the concentrations and time for second order reactions is given by:

$$1/[A]_t = kt + 1/[A]_0$$

Since this equation has the form of the equation for a straight line, *if we graph  $1/[A]$  versus time, a second order reaction will yield a straight line where  $k = \text{slope}$ .*

##### D. Half-Life of Second Order Reactions

The expression relating half life of a second order reaction to the value of the rate constant for a reaction is

$$t_{1/2} = 1/k[A]_0$$

(See pages 587 – 593 of the text for more information and examples.)

#### V. Temperature and Rate

##### A. The Collision Model of Chemical Kinetics

The collision model of chemical kinetics states that molecules must collide to react. Those collisions must have both the appropriate energy and the appropriate orientation for a reaction to occur.

##### B. Activation Energy and Temperature Dependence of Rate Constants

The activation energy is the minimum energy required for a chemical reaction to occur. It is the energy required to convert the reactants to the activated complex (or transition state) where bonds are in the process of forming and breaking. (See Figure 14.15 of the text.) The rate constant for a reaction can be related to a frequency factor, the activation energy, the temperature, and  $R$  as in the Arrhenius equation:

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

If we take the natural logarithm of each side, the Arrhenius equation becomes:

$$\ln k = (-E_a/RT) + \ln A$$

The value of  $E_a$  can be determined by measuring  $k$  at two different temperatures, and using the relationship:

$$\ln(k_2/k_1) = E_a/R(1/T_2 - 1/T_1)$$

## VI. Reaction Mechanisms

### A. Rate Laws and Reaction Mechanisms -

A reaction mechanism is the series of elementary reactions (one step reactions) that make up the pathway by which the reaction occurs. The overall reaction is the sum of the individual steps in the mechanism. The number of molecules that are reactants in an elementary reaction determines the molecularity of the reaction. One reactant molecule results in a unimolecular reaction; two reactant molecules in a bimolecular reaction, etc. The slowest step in a mechanism is rate determining, and the rate law can be written using the stoichiometry of this step. The reactants in the slow step of a mechanism are the ones that appear in the rate law for the reaction. If the slow step involves only one reactant molecule, the reaction is said to be unimolecular; if it involves two reactant molecules, it is said to be bimolecular; if it involves three reactant molecules, it is said to be termolecular. (See lecture notes from Chem 1202 and the Chem 1202 Review/Preview Session, as well as section 14.6 and Table 14.3 of the text for a thorough discussion of reaction mechanisms and rate laws.)

B. Reaction Intermediates - Species that are formed in one step of a reaction mechanism, and consumed in a subsequent step. Hence, in the series of steps that represents the mechanism, reaction intermediates show up first as products, and then as reactants.

## VII. Catalysts

Catalysts are substances that speed up the rates of reactions without being consumed. Hence, catalysts are used in one step of a mechanism, but regenerated in a later step. In the series of steps that represents the mechanism, catalysts show up first as reactants, and then as products.

There are many different types of catalysts, including homogeneous catalysts (same phase as the reactants), heterogeneous catalysts (different phase from the reactants), and enzymes (biological catalysts that are composed of proteins).

See pp. 610 – 614 of the text for a more complete discussion of catalysts.

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**Chemistry 1202 Review/Preview**  
**Chapter Fifteen Review Guide**

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**Spring 2007**

**I. The Concept of Equilibrium**

A reaction at equilibrium indicates that the rates of the forward and reverse reactions are equal. It does NOT indicate that the amounts of reactants and products are equal.

For the reaction  $A \rightleftharpoons B$

the rates of the forward reaction and reverse reactions are :

$$\text{rate}_f = k_f[A] \quad \text{and} \quad \text{rate}_r = k_r[B]$$

When the two rates are equal:  $k_f[A] = k_r[B]$

Dividing both sides of the equation above by  $k_r[A]$ , we obtain

$$\frac{k_f}{k_r} = \frac{[B]}{[A]} = \text{a constant}$$

**II. The Equilibrium Constant**

A. The Equilibrium Constant,  $K_c$ , represents the quantitative relationship between the concentrations of reactants and products in a chemical reaction.

The equilibrium constant expression for a general reaction is shown below.



$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Note that the brackets indicate moles/liter concentrations.

B. The Equilibrium Constant,  $K_p$

The quantitative relationship between the partial pressures (in atmospheres) of the reactants and products in a chemical reaction at equilibrium is described by the form of the equilibrium equation shown below:

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

Note that  $P_A$  represents the partial pressure of A.

The relationship between  $K_p$  and  $K_c$  is given by the following relationship:

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

where  $\Delta n$  is the number of moles of gaseous products minus the number of moles of gaseous reactants.  $T$  is in kelvins;  $R$  is the gas constant, 0.0821 L-atm/mol-K.

When the number of gaseous moles of reactants equals the number of gaseous moles of product,  $\Delta n = 0$ , and  $K_p = K_c$ .

### III. Interpreting and Working with Equilibrium Constants

#### A. Judging the Extent of Reaction from the Value of K

The size of the equilibrium constant K, indicates whether the equilibrium mixture is mostly products, mostly reactants, or an approximately equal mixture of the two.

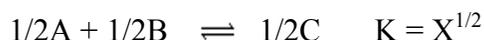
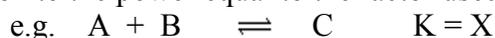
$K \ll 1$  is mostly reactants;                       $K \gg 1$  is mostly products;  
K between 0.01 and 100 is a mixture of significant amounts of both reactants and products.

#### B. Value of K for reactions written in reverse

When the equation for a reversible reaction is written in the opposite direction, the equilibrium constant becomes the inverse of the original equilibrium constant.

#### C. Value of K for reactions whose coefficients have been multiplied by a number.

The K of a reaction that is the multiple of another can be obtained by raising the K of the original reaction to the power equal to the factor used to obtain the new equation.



#### D. Value of K for net reactions

The K of the net reaction from two or more reactions is numerically equal to the product of the individual K's.

#### E. Dependence of K on Temperature

Note that the value of K is dependent on the temperature. If K decreases with a decrease in T, the reaction is endothermic. If K increases with a decrease in T, the reaction is exothermic.

### IV. Heterogeneous Equilibria

For a heterogeneous reaction (a reaction involving more than one phase) the concentrations of water, solids, and liquids are generally omitted from the equilibrium equation. Only the concentrations of gases and solutes in liquids are included. The concentration of water, pure solid substances and pure liquid substances are considered to be constant over the course of the reaction.

See pages 641 and 642 of the textbook for examples.

### V. Calculating Equilibrium Constants

A. From equilibrium concentrations: Substitute concentrations into the K expression.

B. From initial concentrations and one equilibrium concentration: Use the coefficients in the balanced equation to determine all other equilibrium concentrations, and substitute these values into the K expression.

(See pages 644 and 645 of the text for more explanation and examples.)

## VI. Applications of Equilibrium Constants

### A. Predicting the Direction of Reaction

Calculate  $Q$ , the reaction quotient. If  $Q < K$  the reaction will proceed to the right to reach equilibrium. If  $Q > K$ , the reaction will proceed to the left to reach equilibrium. If  $Q = K$ , the reaction is at equilibrium.

### B. Calculating one concentration from $K$ and other equilibrium values:

Substitute the known values into the  $K$  expression and solve for the unknown.

### C. Calculating equilibrium concentrations from $K$ and initial concentrations:

Set up an ICE (Initial, Change, Equilibrium) table to express equilibrium concentrations in terms of  $x$ . Solve for  $x$  and perform the relevant arithmetic operations to get the equilibrium concentration of each species.

NOTE: Before using the quadratic formula, determine if the equation can be simplified by taking the square root of both sides.

See page 647 of the text, and do sample exercises 15.10, 15.11, and 15.12 without looking at the solutions.

## VII. Le Chatelier's Principle

When an equilibrium system is stressed, reaction shifts toward the direction that relieves the stress. The stresses to the system can be in the form of changes in concentration, pressure, volume, temperature, or the addition of a catalyst. The generalizations about Le Chatelier's Principle are:

1. A system will always shift towards a decrease and away from an increase in concentration.
2. An increase in pressure will cause a shift towards the side with the fewer moles of gas. A decrease in pressure will shift towards the side with more moles of gas.
3. An increase in volume will cause a shift toward the side with more moles of gas. A decrease in volume will shift towards the side with fewer moles of gas. Note that an increase in volume results in a decrease in pressure, and vice versa.
4. Any change in temperature can be treated like a change in concentration, where heat is the reactant in an endothermic reaction and heat is the product in an exothermic reaction.
5. A catalyst will reduce the amount of time required for a system to reach equilibrium, but it will not affect the position of equilibrium. (The rates of the forward and reverse reactions are sped up by the same amount.)

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**Chapter Sixteen Review Guide**

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**Spring 2007**

**I. Acids and Bases: A Brief Review**

Acids were originally defined by Arrhenius as substances that produce  $H^+$  ions in water.

Bases were defined by Arrhenius as substances that produce  $OH^-$  ions in water.

**II. Bronsted-Lowry Acids and Bases**

Acids are proton donors; bases are proton acceptors.

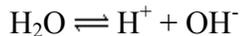
Conjugate acid-base pairs differ only in that the acid contains one more proton than the base. e.g. HCN (acid) and  $CN^-$  (base)

The stronger an acid, the weaker its conjugate base.

See section 16.2 and pages 670 to 676 of the text for more information and examples.

**III. The Autoionization of Water**

Water can act as an acid or a base, and one molecule of water can react as an acid with another molecule that acts as the base, according to the following reaction:



For aqueous solutions,  $[H^+][OH^-] = K_w = 1.0 \times 10^{-14}$  at  $25^\circ C$

In a neutral solution,  $[H^+] = [OH^-] = 1.0 \times 10^{-7}$

**IV. The pH Scale**

The concentration of  $H^+$  in a solution is indicated by the pH.

A. Calculating pH, pOH,  $[H^+]$ ,  $[OH^-]$

$pH = -\log [H^+]$ ;  $pOH = -\log [OH^-]$ ;

$[H^+] = 10^{-pH}$ ;  $[OH^-] = 10^{-pOH}$  (Note that  $10^{-x} = \text{antilog of } -x$ )

$pOH + pH = 14$

For acidic solutions,  $pH < 7$ ; for basic solutions,  $pH > 7$ ; for neutral solutions,  $pH = 7$

B. The pH can be measured using a pH meter (See Figure 16.6), acid-base indicators, or litmus paper.

**V. Strong Acids and Bases**

A. Strong acids and bases are almost completely dissociated in aqueous solutions. The strong acids that should be memorized and recognized are, HCl, HBr, HI,  $H_2SO_4$ ,  $HNO_3$ ,  $HClO_3$ , and  $HClO_4$ .

B. The strong bases are the hydroxides of the Group IA Group IIA metals (except Be). Note that there are relatively few strong acids and strong bases, but many weak acids and weak bases.

Note that the conjugate bases of strong acids are very weak bases; i.e. they have a negligible tendency to be protonated.

**VI. Weak Acids**

Since weak acids dissociate only partially and establish an equilibrium condition.



Therefore an equilibrium equation must be set up to determine equilibrium concentrations of species in solution. The equilibrium constant is referred to as the acid-dissociation constant, and is specified as  $K_a$ . See Table 16.2 for the values of  $K_a$  for selected acids.

A. Calculating  $K_a$  from pH

If we have the pH, we can find the  $[H^+]$  concentration at equilibrium. (Take the antilog of  $-pH$ .) We can then set up an ICE table (like we did in chapter 15) to calculate the other equilibrium concentrations and the value of the  $K_a$ . (See pages 685 – 686 of the textbook for more examples.)

B. Using  $K_a$  to calculate pH.

If we have the  $K_a$  and the initial concentration of the weak acid, we can calculate the equilibrium concentrations using the ICE table, letting  $x$  represent the equilibrium concentrations of the unknown species. (See page 687 and Sample Exercise 16.11 of the text.)

C. Polyprotic Acids - acids that have more than one dissociable hydrogen. In the case of a diprotic acid in which both acids are weak, there will be two  $K_a$  values.  $K_{a1}$  is the constant for the first hydrogen dissociation and  $K_{a2}$  is the constant for the second hydrogen dissociation. Because it is always harder to remove the second  $H^+$  than the first one, and easier to remove the second one than the third one,  $K_a$  values decrease, such that  $K_{a1} > K_{a2} > K_{a3}$ , etc. See section 16.6 and lecture notes for a more complete discussion of solutions of polyprotic acids.

## VII. Weak Bases

Weak bases produce  $OH^-$  ions in an equilibrium reaction. Most weak bases are of two types: molecules containing atoms with a lone pair of electrons (such as N as is the case in ammonia derivatives containing an NH bond, and the anions of weak acids (such as  $F^-$ ).

See lecture notes and Sample Exercise 16.15 to learn the procedure for calculating the concentration of a salt from the pH.)

### Calculating the pH of Solutions of Weak Bases

Since weak bases establish an equilibrium condition in aqueous solution, an equilibrium equation must be set up to determine equilibrium concentrations of species in solution. The equilibrium constant is referred to as the base-dissociation constant, and is specified  $K_b$ . See Table 16.4 for the values of  $K_b$  for selected bases. (See section 16.7 and Sample Exercise 16.4 of the text and lecture notes for examples of calculations involved solutions of weak bases.)

## VIII. Relationship Between $K_a$ and $K_b$

For conjugate acid-base pairs, the conjugate acid contains one more  $H^+$  than the conjugate base of the pair. Some examples are:  $H_2S$  and  $HS^-$ ;  $HCN$  and  $CN^-$ ;  $HSO_4^-$  and  $SO_4^{2-}$ , etc.

For conjugate acid-base pairs,  $K_a \times K_b = K_w$  and  $pK_a + pK_b = 14$

## IX. Acid-Base Properties of Salts Solutions

Salts in which the anion is derived from a weak acid or the cation is derived from a weak base undergo hydrolysis when placed in water. Hydrolysis is the process by which the anions of weak acids and the cations of weak bases react with water to produce a weak acid and  $OH^-$  or a weak base and  $H^+$ , respectively.

See pages 700 – 701 of the text for the mechanism for determining whether a salt solution will be acidic, basic, or neutral.

## X. Acid Base Behavior and Chemical Structure

A. Binary Acids (HX) - Both electronegativity and bond strength determine acid strength, with the bond strength being the more important factor. The bond strength decreases from top to bottom in the periodic table, and the acid strength increases (e.g.  $\text{HI} \gg \text{HF}$ ). When moving across a period, the acid strength increases with increasing electronegativity of the element bonded to the hydrogen. (e.g.  $\text{HF} > \text{CH}_4$ .) See section 16.10 of the text and lecture notes for more discussion of these trends.

B. Oxoacids - Both the electronegativity of the central atom and the number of OH groups in the acid contribute to determining the acid strength. If the number of OH groups is the same in two oxoacids (such as  $\text{HClO}_4$  and  $\text{HIO}_4$  the stronger acid is the one with the more electronegative central atom ( $\text{HClO}_4$  in this case). If the number of OH groups is not the same, the stronger acid is the one with more OH groups.

C. Carboxylic acids - Carboxylic acids have the general formula  $\text{R-COOH}$ . The more electronegative atoms in the R group, the stronger the acid.

## XI. Lewis Acids and Bases

Lewis acids - electron pair acceptors; Lewis bases - electron pair donors  
Species with lone pairs of electrons, such as  $\text{NH}_3$  can act as Lewis bases. Species with empty orbitals, such as  $\text{H}^+$ , elements in Group IIIA or metals can act as Lewis acids. Hydrated metal cations often act as Lewis acids. (See section 16.11. and lecture notes.) In general the higher the charge on the metal ion the stronger the acid, and the smaller the size of the ion, the stronger the acid.

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**Spring 2007**

**Additional Aspects of Aqueous Equilibria**

**I. The Common Ion Effect**

The addition of a common ion to a weak acid or base solution shifts the equilibrium in the direction of the undissociated species and changes the pH of the solution. This is the same result as would be predicted by LeChatelier's principle. This phenomenon is generally known as the common ion effect, and is defined as the shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substances. (See pages 722 – 724 of the text.)

**II. Buffered Solutions**

Buffer solutions are solutions that resist a change in pH when acid or base is added.

A. Buffer solutions are made from a weak acid and its salt or a weak base and its salt

B. The Henderson-Hasselbach Equation can be used to calculate the pH of buffered solutions:  $\text{pH} = \text{pK}_a + \log [\text{base}]/[\text{acid}]$

C. Buffer Capacity

Buffer solutions can resist a change in pH after the addition of base or acid. The effective buffer region is one in which the pH of the solution is within one unit of the  $\text{pK}_a$  of the weak acid to be used. Effective buffer region:  $\text{pK}_a \pm 1$ . The optimal pH of a buffer solution is equal to the  $\text{pK}_a$  of the acid.

See section 17.2 of the text for a more detailed discussion of buffer solutions.

D. Calculating the pH of a buffer solution after addition of acid or base. To calculate the pH of a buffer solution after the addition of acid or base, calculate new concentrations of the major species in the buffer solution and use the Henderson Hasselbach equation to calculate the new pH. See pages 729 and work through Sample Exercise 17.5 for practice with these types of problems.

**III. Acid-Base Titrations**

A. Strong Acid/Strong Base:  $\text{pH} = 7$  at the equivalence point (neutral salt produced).

B. Weak Acid/Strong Base:  $\text{pH} > 7$  at the equivalence point (basic salt produced).

C. Regions in Titration Calculations

1. The initial pH -- before the addition of any base (or acid)

Calculate the pH of the pure acid (or base)

2. Between the initial pH and the equivalence point -- after the addition of some base (or acid) but before reaching the equivalence point.

The calculations over this region differ depending on whether the titration involves only strong species or if a weak species is involved. See Section 17.3 for the correct procedures in each case.

3. At the equivalence point.

The salt and water are present. Calculate the pH of the salt solution.

4. After the equivalence point.

There will be excess base (or acid) present. Calculate the pH of the base (or acid) solution.

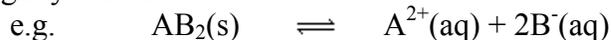
Always remember to use the new volume to calculate the new concentrations of species in solution for titration problems.

D. Titrations of Polyprotic Acid

These titrations have multiple equivalence points. (See page 739.)

#### IV. Solubility Equilibria

A.  $K_{sp}$  expressions - Equilibrium constant expressions for the dissolution of very slightly soluble salts.



$$K_{sp} = [A^{2+}][B^-]^2$$

B. Calculating  $K_{sp}$  from solubility data

Convert the solubility in g/L (if that is what is given in the problem) to moles/L. Use stoichiometry to calculate the molar concentrations of the ions. Substitute these values into the  $K_{sp}$  expression. Work through sample exercise 17.10 for practice.

C. Calculating solubility from  $K_{sp}$ .

Let  $x$  represent the moles of solid that react. Determine the concentrations of aqueous ions stoichiometrically. e.g.  $2x$ ,  $3x$ , etc. Substitute into the  $K_{sp}$  expression and solve for  $x$ . Convert the molar solubility to grams to get the solubility in g/L. Work through sample exercise 17.11.

#### V. Factors that Affect Solubility

A. The Common Ion Effect and Solubility

The presence of a common ion always decreases the solubility of a compound in water. Work through sample exercise 17.12.

B. pH and Solubility

Use LeChatlier's Principle to determine the position of equilibrium shift. See page 745 in the text and lecture notes for a discussion of this concept.

C. Formation of Complex Ions and Amphoterism

See section 17.5 and lecture notes for a discussion of this.

#### VI. Precipitation and Separation of Ions

A. To determine whether a precipitate will form:

Calculate  $Q$  from the information given.

If  $Q > K_{sp}$  a precipitate will form. If  $Q < K_{sp}$  a precipitate will not form.

If  $Q = K_{sp}$  the solution is saturated.

B. Selective precipitation of ions

A mixture of ions in solution can be separated by using a reagent with an ion that forms a precipitate with only one of the ions in the mixture. The two major types of steps involved in calculations involving selective precipitation are:

1. Calculation of the amount of added ion needed to initiate precipitation of the least soluble ion by using the  $K_{sp}$  expression of this salt.
2. Calculation of the amount of the initial ion remaining just before the second ion begins to precipitate by using the  $K_{sp}$  values of both salts.

Work through sample exercise 17.16 for more practice in solving these types of problems.

#### XI. Qualitative Analysis for Metallic Elements

See lecture notes from Chem 1202 to see coverage (if any) of this topic.

**NOTE: This review sheet is for Chem 1202 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1202.**



**Chemistry 1202 Review/Preview  
Chapter Nineteen Review Guide**

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**Spring 2007**

**Thermodynamics**

IMPORTANT NOTE: Chapter 19 assumes mastery of the material in Chapter 5. The Chapter 19 problems require the application of the principles learned in Chapter 5. The Chapter 5 review guide provided in the Chem 1202 Review/Preview session should be thoroughly reviewed before the study of Chapter 19.

**I. Spontaneous Processes**

Spontaneous processes proceed on their own without any external assistance. The process that is the reverse of a spontaneous process is always non-spontaneous. The spontaneity of a chemical reaction generally depends on the temperature, pressure, and composition of the reaction mixture. A spontaneous process can occur at a rapid rate or at a rate that is so slow it is impossible to tell that a change is occurring. A process is at equilibrium if the forward and reverse processes are occurring at the same rate and there is not preference for either the forward or the reverse process. Thermodynamics tells us the direction and extent of a chemical process, but does not indicate how fast the reaction will occur. (See pages 803-808 of the text for an in depth discussion of this concept.)

**II. Entropy and the Second Law of Thermodynamics**

**A. Entropy change**

The entropy is the amount of randomness or disorder in a system, and is a characteristic of the state of the system. The entropy of a gas is greater than that of a liquid, which is greater than that of a solid. The entropy change of a process,  $\Delta S$  can be determined from the initial and final entropy values as follows:

$$\Delta S = S_{\text{final}} - S_{\text{initial}}$$

When a process occurs at constant temperature (an isothermal process),  $\Delta S$  is related to the heat that would be transferred in the reversible process divided by the temperature (in kelvins), as follows:  $\Delta S = q_{\text{rev}} / T$

Spontaneous reactions are often ones in which there is an increase in the entropy of the system (e.g. ice melting at 273 K), but not always. Systems tend to move toward an increase in disorder because a random state is more probable than an ordered one.

**B.  $\Delta S$  for Phase Changes**

Because phase changes are isothermal processes (no temperature change is occurring during the phase change),  $q = \Delta H$ , and the entropy change can be calculated from the relationship:

$$\Delta S = \Delta H / T$$

(Notice that the units on  $\Delta S$  will be J/K (joules per kelvin))

Do Sample Exercise 19.2 and the practice exercise for practice calculating  $\Delta S$  for phase changes.)

### C. The Second Law of Thermodynamics

Any spontaneous (irreversible) process occurs with an increase in entropy. Any reversible process occurs with no change in entropy.

$$\text{Reversible process: } \Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} = 0$$

$$\text{Irreversible process: } \Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$$

## III. The Molecular Interpretation of Entropy

### A. Molecular Motions and Entropy

Molecules can undergo three different types of motion, that can increase or decrease with changes in temperature. These are translational motion (movement in one direction), vibrational motion (movement of the atoms in a molecule toward or away from each other), and rotational motion (spinning around an axis).

### B. Boltzmann's Equations and Microstates

A microstate is a particular combination of position and kinetic energy of the molecules in a system. Boltzmann's Formula predicts the entropy of a state, based on the number of microstates possible. Boltzmann's formula is:

$$S = k \ln W$$

where  $k$  is Boltzmann's constant, ( $k = 1.38 \times 10^{-23} \text{ J/K}$ ), and  $W$  is the number of microstates associated with a particular macrostate. For a perfectly ordered state that can be achieved in only one way ( $W = 1$ ),  $S = 0$  because  $\ln 1 = 0$ .

$\Delta S$  is positive if the number of microstates increases during a process, and negative if the number of microstates decreases. Generally, the number of microstates increases when the volume, temperature, or number of particles increases during a process. (See pages 812 – 815 for a more detailed discussion of this concept.)

### C. Making Qualitative Predictions about $\Delta S$

The entropy generally increases when the following occurs:

1. Gases are produced from solids or liquids
2. Liquids or solutions are produced from solids
3. The number of gaseous molecules in the products is greater than the number of gaseous molecules in the reactants.

text Work through the Sample exercises and practice exercises on pages 817 of the text for practice doing these kinds of problems.

### D. The Third Law of Thermodynamics

The entropy of a pure crystalline substance at 0 kelvins is 0,  $S(0\text{K}) = 0$  is the third law of thermodynamics. The entropy of a substance increases as the temperature increases. (See Figures 19.13 and 19.14 on page 819 of the textbook.)

## IV. Entropy Changes in Chemical Reactions

### A. Standard Molar Entropy of a Substance

The standard molar entropy of a substance, denoted by  $S^\circ$ , is the entropy of one mole of the pure substance at 1 atmosphere pressure and a specified temperature, usually  $25^\circ\text{C}$  (298 K).

Standard entropy values are given in Appendix C of the text. Note that the units are in joules per kelvin mole (not kilojoules/mole, like enthalpy and free energy). Study the general observations on page 820 of the text!

The entropy change of a reaction, denoted by  $\Delta S^\circ$ , is obtained by subtracting the sum of the standard molar entropies of the reactants from the sum of the standard molar entropies of the products.

The entropy change in a chemical reaction is calculated by using the formula:

$$\Delta S^\circ = \sum nS^\circ(\text{products}) - \sum mS^\circ(\text{reactants})$$

where n and m are the coefficients in the balanced chemical equation.

(See exercises on page 821 for practice.)

#### B. Entropy Changes in the Surroundings

$$\Delta S_{\text{surroundings}} = -q_{\text{system}}/T = -\Delta H/T$$

Therefore if a process is exothermic (q is negative) the entropy of the surroundings always increases.  $-(-\text{value}) = + \text{value}$

### V. Gibbs Free Energy

Gibbs Free Energy (G) (also referred to as free energy) is the usable energy produced (or consumed) by a chemical reaction. The free energy is related to the absolute temperature, enthalpy, and entropy of a system according to the relationship:  $G = H - TS$ .

#### A. Standard Free Energy Changes

The change in free energy is given by the relationship:  $\Delta G = \Delta H - T \Delta S$ , where T must be in kelvins. The spontaneity of a reaction can be determined by the value of  $\Delta G$ . When  $\Delta G < 0$  (negative), the reaction is spontaneous, when  $\Delta G > 0$  (positive), the reaction is non-spontaneous, and when  $\Delta G = 0$ , the reaction is at equilibrium. Table 19.4 provides valuable information about the relationship between  $\Delta H$ ,  $\Delta S$ , the absolute temperature (T), and  $\Delta G$ . You should know the information in this table very well. The  $\Delta G$  of a reaction is related to the change in the entropy of the universe (system + surroundings) by the following relationship:  $\Delta G = -T\Delta S_{\text{universe}}$

As per equation 19.12,  $-T\Delta S_{\text{universe}} = \Delta H_{\text{system}} - T\Delta S_{\text{universe}}$

Substituting  $\Delta G$  for  $-T\Delta S_{\text{universe}}$ , we get the equation for free energy presented above:  $\Delta G = \Delta H - T \Delta S$

The temperature at which a non-spontaneous reaction becomes spontaneous can be determined by setting  $\Delta G$  equal to 0 and solving for T.

$$0 = \Delta H - T \Delta S$$

$$\text{Therefore } T = \frac{\Delta H}{\Delta S}$$

The equation above can always be used to calculate the temperature at which phase changes occur because phase changes are equilibrium processes at which  $\Delta G = 0$ .

#### B. Standard Free Energy Changes for Reactions ( $\Delta G^\circ$ ) - the free energy change that occurs when reactants in their standard states are converted to products in their standard states.

##### 1. Calculating $\Delta G^\circ$ from standard enthalpy and standard entropy changes

The  $\Delta G^\circ$  of a reaction can be determined from the  $\Delta H^\circ$  and  $\Delta S^\circ$  of a reaction by the following relationship:

$$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ, \text{ where T must be in kelvins}$$

## 2. Standard Free Energies of Formation

The  $\Delta G^\circ$  of a reaction can also be calculated from the  $\Delta G^\circ_{\text{formation}}$  of the reactants and products by the following relationship:

$$\Delta G^\circ = \sum nG^\circ(\text{products}) - \sum mG^\circ(\text{reactants})$$

## VI. Free Energy and Temperature

Because  $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$ , the free energy change depends on the temperature at which the reaction is occurring. Notice that the temperature affects the entropy term, and that T is always a positive value. See table 19.4 for a representation of the effect of T on the spontaneity of reactions. Do the exercises on page 829 for more practice.

## VII. Free Energy and the Equilibrium Constant

A. Relationship of free energy to the reaction quotient, Q:

$$\Delta G = \Delta G^\circ + RT \ln Q$$

where R is the gas constant (8.314 J/mol-K), T is the absolute temperature in kelvins, and Q is the reaction quotient (See handout of chapter 15).

Note that when mostly reactants are present,  $RT \ln Q$  will be a large negative number that will dominate the expression, and the reaction will proceed spontaneously. When mostly products are present,  $RT \ln Q$  will be a large positive number that will dominate, and the reverse reaction will proceed spontaneously. Read Section 19.7 very carefully, and do the exercises on pages 830 and 831 of the text.

B. Relationship of Free Energy to K: At equilibrium,  $\Delta G = 0$  and  $Q = K$ .

Therefore  $\Delta G^\circ = -RT \ln K$

The relationship between  $\Delta G^\circ$  and K is given in table 19.5 of the text. You should do the calculations to get these values so that you could recreate the table from scratch!

**NOTE: This review sheet is for Chem 1202 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1202.**

**Chemistry 1202 Review/Preview  
Chapter Twenty Review Guide**

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**Spring 2007**

**Electrochemistry**

IMPORTANT NOTE: Chapter 20 assumes mastery of the material on oxidation-reduction (redox) reactions in Chapter 4, especially section 4.4 (pages 137-139). The section on oxidation-reduction reactions in the Chapter 4 review guide should be thoroughly reviewed before beginning the study of Chapter 20.

**I. Oxidation States**

When electrons are transferred between the reactants in a chemical reaction, the element that loses electrons is said to be oxidized and the element that gains electrons is said to be reduced. Oxidation states are compared to tell if a reaction is a redox reaction. The substance containing the element that is oxidized is called the reducing agent (or reductant). The substance containing the element being reduced is called the oxidizing agent (or oxidant).

**A. Assigning Oxidation States (Numbers)**

Refer to the rules presented in Section 4.4 of the textbook, and memorize these rules if you have forgotten any of them! (Work through exercise 20.1 for practice.)

**II. Balancing Redox Equations in Acidic or Basic Solution**

**A. Half-Reactions**

An oxidation-reduction reaction can be broken into two half-reactions, one showing the oxidation process, and the other showing the reduction process. In the overall reaction, the number of electrons lost during oxidation must equal the number of electrons gained during reduction. (We can't have extra electrons floating out in space, and we can't grab electrons from space!) The electrons gained can only be produced by the substance that is losing electrons. When each half-reaction has the same number of electrons involved, we can add them together to get the overall oxidation-reduction reaction.

**B. Balancing Equations by the Method of Half-Reactions**

Important note: I do not recommend assigning oxidation numbers first (as the textbook does). It is not generally necessary to determine what is being oxidized and what is being reduced before you can write the half-reactions. It does not hurt to assign the oxidation numbers, but it is usually not necessary.

Dr. McGuire's recommended steps for balancing by the half-reaction method

1. Write the unbalanced ionic equation.
2. Write the individual half-reactions by doing the following:
  - a) Balance all atoms other than H and O
  - b) Balance the oxygens by adding water to the side deficient in oxygens
  - c) Balance the hydrogens by adding H<sup>+</sup> to the side deficient in hydrogens
  - d) Balance the charges by adding electrons to the more positive side of the equation
3. Multiply one or both half-reactions by the appropriate number(s) to make the number of electrons transferred equal.

4. Add the two half-reactions, canceling species that appear on both sides of the equation), and check that the overall equation is balanced.
  5. Check that all of the elements and the charges are balanced.
- C. Balancing Equations for Reactions Occurring in Basic Solution
- If an oxidation-reduction reaction occurs in basic solution, there will not be extra  $H^+$  ions in solution, so we cannot have  $H^+$  in the balanced equation. To remove the  $H^+$  we do add following step to the procedure outlined above for balancing redox reactions:
6. Add enough  $OH^-$  to both sides of the equation to neutralize the  $H^+$ . Combine water molecules on both sides of the equation, canceling as necessary.

VERY IMPORTANT: Practice this procedure by doing all of the examples and the odd numbered problems (problems in red) in the text that involve balancing redox equations. Check your answers! The more you practice, the less likely you will make a mistake on the test or quiz!

### III. Voltaic (or Galvanic) Cells

*Electrochemical cells* are devices used for interconverting chemical and electrical energy. *Voltaic* (or galvanic) cells produce electricity from a spontaneous chemical reaction, whereas *electrolytic* cells utilize an electric current to drive a non-spontaneous chemical reaction. Electrochemical cells involve oxidation-reduction reactions in which there is a transfer of electrons.

A. Voltaic cells - consist of two half-cells, which contain electrodes. The half-cells are connected by a salt-bridge. (Study Figures 20.5 and 20.6 *very* carefully, and be able to recreate them.) The various components of a voltaic cell are listed below:

1. Anode - The electrode at which oxidation occurs
2. Cathode - The electrode at which reduction occurs
3. Salt Bridge - a tube containing a salt solution that completes the electrical circuit and allows the electrical neutrality of the cell to be maintained. Note that the anions in the salt bridge flow toward the anode, while the cations flow toward the cathode. The electrons in the circuit flow toward the cathode and away from the anode.

D. Anode and Cathode Half-reactions - the oxidation and reduction half-reactions that occur at the anode and the cathode, respectively.

### IV. Cell EMF (Potentials) Under Standard Conditions

A. Electromotive force (or cell voltage or cell potential) is the voltage difference between the electrodes; it represents the driving force that pushes the negatively charged electrons away from the anode and toward the cathode, and has the unit of volts. The relationship between the volt (potential), the joule (energy), and the coulomb (electrical charge) is as follows:

$$1 \text{ Volt} = 1 \frac{\text{Joule}}{\text{Coulomb}}$$

or alternatively,  $1 \text{ Joule} = 1 \text{ Volt} \times 1 \text{ Coulomb}$

B. Standard Reduction (Half-Cell) Potentials

The standard reduction potential of a galvanic cell is the standard half-cell potential for the cathode (reduction) half-reaction *minus* the standard half-cell potential for the anode (oxidation) half-reaction, as shown below:

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}}(\text{cathode}) - E^\circ_{\text{red}}(\text{anode})$$

The half-reaction that occurs in the standard hydrogen electrode (SHE) is assigned a reduction potential of 0 V, and other potentials are determined relative to this. (See page 861 of the text.)

C. Table of Standard Reduction Potentials (Table 20.1) - lists the potentials of reduction half reactions that are used in the calculation of cell potentials. The oxidation reaction potentials are equal in magnitude but opposite in sign from the reduction potentials. This is the reason that when calculating the  $E^\circ_{\text{cell}}$  the oxidation half-cell potential must be subtracted from that of the reduction half-cell potential.

D. Strengths of Oxidizing and Reducing Agents

The order of the half-reactions in the Table of Standard Reduction Potentials is an indication of the strengths of the oxidizing and reducing agents in the table.

The half-reactions are listed in order of decreasing standard reduction potential.

Therefore, the strongest oxidizing agents are listed in the upper left of the table, and the strongest reducing agents are listed in the lower right of the table.

(Convince yourself of this by carefully looking at the numbers and the species involved in the half-reactions. Hint: Why would metals be good reducing agents? Why would fluorine be a good oxidizing agent?)

## V. Free Energy and Redox Reactions

The Table of Standard Reduction Potentials can be used to calculate cell potentials as well as to determine which reactions will occur spontaneously (have a negative  $\Delta G$ , and can therefore be utilized in a voltaic cell.)

A. Relationship Between the Value of the cell potential and spontaneity

Those reactions with a positive standard cell potential will occur spontaneously, and those with a negative standard cell potential will not occur spontaneously. In general, an oxidizing agent will oxidize any reducing agent that lies below it in the table, but not one that lies above it in the table. (Reactions will occur spontaneously between species that lie in the relationship upper left to lower right in the table of standard reduction potentials, but not in the opposite direction.)

e.g.  $\text{Cu}^{2+}(\text{aq}) + \text{H}_2(\text{g}) \rightarrow \text{Cu}(\text{s}) + 2 \text{H}^+(\text{aq})$  is spontaneous

$\text{Ni}^{2+}(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{Ni}(\text{s}) + \text{Cu}^{2+}(\text{aq})$  is not spontaneous

Note: Calculate the cell potentials of the two reactions above to confirm this!

B. EMF (Cell Potential) and  $\Delta G$  Free Energy Changes for Reactions

Since the cell potential and the change in free energy are both measures of the spontaneity of chemical reactions, they are related to each other by the following equation:

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

where  $n$  is the number of electrons transferred during the process,  $F$  is the Faraday constant which is equal to 96,485 Coulombs/mol  $e^-$ , and  $E$  is the calculated cell potential.

C. Standard Cell Potential and Standard Free Energy Change

The standard cell potential,  $E^\circ$ , is related to the standard free energy change,  $\Delta G^\circ$ , by the following relationship:  $\Delta G^\circ = -n F E^\circ$

Because  $\Delta G^\circ$  and  $E^\circ$  are directly proportional to each other; the higher the standard cell potential, the more negative the standard free energy change.

D. Standard Cell Potential and the Equilibrium Constant

The cell standard potential can also be used to calculate the equilibrium constant because of the relationship between  $\Delta G^\circ$  and  $K$ ,  $\Delta G^\circ = -RT \ln K$ . Combining both  $\Delta G^\circ$  expressions, we get

$$-n F E^\circ = -RT \ln K, \text{ or simply } n F E^\circ = RT \ln K$$

Note that when  $E^\circ_{\text{cell}} > 0$ ,  $K > 1$ , and when  $E^\circ_{\text{cell}} < 0$ ,  $K < 1$ .

(Work through Sample Exercise 20.10 for practice with these types of problems.)

## VI. Cell EMF under Nonstandard Conditions

A. The Nernst Equation - allows the calculation of the cell potential for reactions that are not at standard conditions, i.e. all of the concentrations are not 1M and/or not all of the pressures are 1 atm.

To calculate  $E_{\text{cell}}$  at these conditions we use the Nernst equation, shown below:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q \quad \text{or} \quad E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{2.303RT}{nF} \log Q,$$

where Q involves the concentrations and/or pressures of species present.

Since  $2.303RT/F$  is equal to 0.0592 at 25°C, the Nernst equation reduces to:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q \quad (T = 298\text{K})$$

(Work through Exercise 20.11 for practice in using the Nernst equation.)

B. Concentration cells (discussed on page 874 of the text) are those cells in which the cell potential is generated from differences in concentration of the same species, and not from using species with different reduction potentials. (See pages 874 and 875 for more on this topic.)

## VII. Batteries and Fuel Cells

A. Batteries are units that consist of one or more voltaic cells that can be used as a source of direct electric current at a constant voltage. See section 20.7 and Chem 1202 lecture notes for a complete discussion of batteries. You should probably know how to write the half-reactions involved in any batteries discussed in lecture. Check with your instructor to find out which specific reactions you should know if this was not made clear in lecture.

## VIII. Corrosion

Corrosion is the term used to describe the oxidative deterioration of metals to the ores from which they were originally obtained. See section 20.8 and Chem 1202 lecture notes for a discussion of corrosion.

A. Corrosion of Iron - See pages 881 – 882 of the text.

B. Prevention of Corrosion

1. Galvanizing - coating the metal with a layer of zinc
2. Cathodic Protection - connecting the metal to a second metal that is more easily oxidized than the metal being protected.

## IX. Electrolysis

A. Electrolysis reactions are reactions in which an electric current is used to bring about a non-spontaneous chemical change.

B. Electrolytic cells are the cells in which electrolysis reactions occur.

(See pages 883 - 885 of the text and Chem 1202 lecture notes for a thorough discussion of this topic.)

C. Quantitative Aspects of Electrolysis

The amount of product obtained in an electrochemical cell is related to the number of moles of electrons passed through the cell. The moles of electrons depend on the current, measured in amperes, and the time of the current flow, measured in seconds. The relationships are summarized below:

$$\text{Coulombs (charge)} = \text{amperes (current)} \times \text{seconds (time)}$$

$$1 \text{ Faraday} = 1F = 1 \text{ mole of electrons} = 96,485 \text{ Coulombs}$$

All electrochemical reactions will involve a certain number of moles of electrons, which can be related stoichiometrically to the other species in the balanced equation. Therefore, once the coulombs are determined from the amperes and the seconds, the faradays (moles of electrons) can be determined by dividing the coulombs by 96,485. The faradays can then be used to determine the number of moles of any other species in the balanced equation, and the moles can be converted to grams. (Work through Exercise 20.14 for practice with these types of problems.)

#### X. Review of Measures of Spontaneity

$\Delta G$	K	$E_{\text{cell}}$	Reaction Status
Negative	>1	Positive	Spontaneous
0	=1	0	At Equilibrium
Positive	<1	Negative	Non-spontaneous

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