I. The Concept of Equilibrium
   A. When a system (a chemical reaction) is at equilibrium, the rate at which products are produced from reactants equals the rate at which reactants are produced from products.
   B. At equilibrium, the ratio of the concentrations of the reactants to the products is a constant. This is true no matter which direction the reaction is performed in, or even if you start with a mixture of reactants and products.
   C. Once at equilibrium, reactants and products continue to react, but the concentrations of each do not change
      1. this is called dynamic equilibrium
      2. it is possible since the forward and reverse reactions are occurring at the same rate
   D. Example of how an equilibrium reaction is written:
      \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \]

II. The Equilibrium Constant
   A. The Law of Mass Action
      1. Expresses the relationship between the concentrations of the reactants and products at equilibrium in any reaction
      2. If you have the general equation \[aA + bB \rightleftharpoons pP + qQ\]
         then the law of mass action gives the following relationship at equilibrium:
         \[ K_c = \frac{[P]^p[Q]^q}{[A]^a[B]^b} \]
         [ ] means molar concentration(*use molarity!!)
      3. This relationship is called the equilibrium expression, and can be written for any system that reaches equilibrium
      4. \(K_c\) is called the equilibrium constant; it varies with temperature but is not affected by catalysts or initial concentrations
      5. Law of Mass Action and values of \(K_c\) have been experimentally determined and proven
   B. Expressing Equilibrium Constant in Terms of Pressure
      1. When reactants and products in a equilibrium are gases, the equilibrium expression is written in terms of partial pressures (use atm)
      2. Use \(K_p\) (p for pressures) instead of \(K_c\) (c for concentrations)
      3. From the general equation \[aA + bB \rightleftharpoons pP + qQ\]
         \[ K_p = \frac{(P_P)^p(P_Q)^q}{(P_A)^a(P_B)^b} \]
   C. The Magnitude of Equilibrium Constants
      1. If \(K >> 1\), the equilibrium "lies to the right", the products are favored
      2. If \(K << 1\), the equilibrium "lies to the left", the reactants are favored
      3. Often the double arrows are written adjusted to illustrate the above
   D. The Direction of the Chemical Equation and \(K\)
1. The equilibrium expression for a reaction written in one direction is the reciprocal of the one for the reaction written in the reverse direction.
2. It is necessary then to know in which direction the original equation was written, since the values of the equilibrium constants will be different in each case.

III. Heterogeneous Equilibria

A. When all substances in an equilibrium are all in the same phase, it is called a homogeneous equilibrium.
B. If substances are in different phases, it is called a heterogeneous equilibria.
   1. If a pure solid or pure liquid is involved in a heterogeneous equilibrium, its concentration is not included in the equilibrium expression.
      □ Why???
   2. When you have an equilibrium that involves solutions and gases, the solution concentrations are expressed in molarity and the gas pressures are expressed in atmospheres, both are written in the same expression.

IV. Calculating Equilibrium Constants

A. Steps for creating an ICE diagram
   1. Tabulate the known initial and equilibrium concentration of all species involved in the equilibrium.
   2. For those species for which both the initial and equilibrium concentrations are known, calculate the change in concentration that occurs as the system reached equilibrium.
   3. Use the stoichiometry of the reaction to calculate the changes in concentration for all the other species in the equilibrium.
   4. From the initial concentrations and the changes in concentration, calculate the equilibrium concentrations--these are used to evaluate the equilibrium constant.

B. Relating $K_c$ and $K_p$
   1. Since $PV = nRT$, then $P = (n/V)RT$, or $P = [X]RT$ where $[X]$ is the concentration in molarity of a particular substance in an equilibrium.
   2. So, $K_p = K_c(RT)^\Delta n$ where $\Delta n$ = the change in the number of moles of gas upon going from reactants to products in an equilibrium reaction (the # of moles of gaseous products — the number of moles of gaseous reactants).
      □ Why does this work??

V. Applications of Equilibrium Constants

A. Predicting the Direction of Reaction
   1. Determination of the reaction quotient, $Q$, allows you to predict the direction an equilibrium reaction will proceed.
   2. $Q$ is found by substituting given concentrations or pressures into an equilibrium expression--note $Q$ is not $K$, but $Q = K$ only at equilibrium.
      a. If $Q > K$, substances on the right side of the chemical equation will react to form substances on the left; the reaction moves from right to left in approaching equilibrium.
      b. If $Q < K$, the reaction will achieve equilibrium by forming more products; it moves from left to right.

B. Calculation of Equilibrium Concentrations
1. Write the equilibrium expression, then substitute in the values known
2. You should be given $K$, so one (or more!!) of the concentrations becomes the unknown
3. Solve for the concentration
4. In some problems, an ICE diagram will again be helpful

VI. Le Châtelier's Principle
A. Le Châtelier's Principle--If a system at equilibrium is disturbed by a change in temperature, pressure, or the concentration of one of the components, the system will shift its equilibrium position so as to counteract the effect of the disturbance.

B. Change in Reactant or Product Concentration--If a chemical system is at equilibrium and we add a substance (either a reactant or a product), the reaction will shift so as to reestablish equilibrium by consuming part of the added substance. Conversely, removal of a substance will result in the reaction moving in the direction that forms more of the substance.

\[ \text{What about solids and liquids?} \]

C. Effects of Volume and Pressure Changes--At a constant temperature, reducing the volume of a gaseous equilibrium mixture causes the system to shift in the direction that reduces the number of moles of gas. Conversely, increasing the volume causes a shift in the direction that produces more gas molecules.

\[ \text{How does this involve pressure?} \]
\[ \text{Why is the above true?} \]
\[ \text{What about solids and liquids?} \]
\[ \text{What about addition of inert gas?} \]

D. Effect of Temperature Changes
1. When the temperature is increased, the equilibrium shifts in the direction that absorbs heat
2. In an endothermic reaction: reactants + heat $\rightleftharpoons$ products
   
   Heat is absorbed as reactants are converted to products, so an increase in temperature (heat) causes the equilibrium to shift to the right (toward products), and $K$ increases
3. In an exothermic reaction: reactants $\rightleftharpoons$ products + heat
   
   Heat is absorbed as products are converted to reactants, so the equilibrium shifts to the left and $K$ decreases

E. Effect of Catalysts--A catalyst increase the rate at which equilibrium is achieved, but it does not change the composition of the equilibrium mixture

F. A summary:
1. Concentration-shift in eq. position
2. Volume and pressure-shift in eq. position
3. Temperature-shift in eq. position and $K$ value changes (need to know if endo or exo)
4. Catalyst-neither eq. position nor $K$ changes