Chemical Equilibrium

Key Concepts:

- How do we measure the amounts of matters in Chemistry?
  - Concentrations
    - Molarity (M)
- Brief Intro. to Writing Chemical Equations/Stoichiometry
- What is Dynamic Equilibrium?
- What is Equilibrium Constant?
- Le Chatelier's Principle

- Concentrations
  - homogeneous vs. heterogeneous
    - homogeneous: appears the same throughout the liquid. Nothing settles at the bottom or floats at the top—dissolving sugar or salt. This resulting sugar/salt water is solution
    - heterogeneous—if something settles at the bottom or floats at the top. Sand or flour in water. NOT a solution
  - solute: something you dissolve. Ex: salt, sugar
  - solvent: the medium that you dissolve the solutes in. Ex: water
  - Expressing Concentrations of solutions (or gases)
    - Molarity and Normality
      - Molarity

$$M = molarity = \frac{moles\ \ of\ \ solute}{liters\ \ of\ \ solution}$$

  - [A]= concentration (molarity) of A
  - another exercise calculating molarity

- Chemical Equations
  - $aA + bB \Leftrightarrow cC + dD$
  - reactants: A and B
  - products C and D
  - a, b, c, d are constants (coefficients), to note the relative amounts of one another
  - double arrow means it's in equilibrium. If a reaction is written with a single-headed arrow, that means the reaction is spontaneous (once the product is formed, it never comes back to form the reactants)
    - example: $HCl \rightarrow H^+ + Cl^-$

- Concept of Dynamic Equilibrium
  - Equilibrium: if the amount of the reactants and the products stays the same, we say that the system has reached an equilibrium. However, this doesn't mean that the reaction stops; the reaction still occurs, but the rate of product formation is the same as the rate of product breakdown (and formation of the reactants). Thus the amount stays the same, but if we were to track the individual atoms, we would see that the product is formed and broken down constantly.

- Equilibrium Constant (K_{eq}) specifies relative amounts of each chemical at equilibrium
    - Law of mass action: for a reaction $aA + bB \Leftrightarrow cC + dD$, the law of mass action is represented by the following equilibrium expression
      $K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$, where [A], [B], [C], [D] are the concentrations of each at equilibrium
        - http://www.kentchemistry.com/links/Kinetics/Equilibrium/equilibrium.swf

    - Reaction Quotient
      $Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$, which looks the same as K, but here we use the concentrations that are NOT at equilibrium (often called the "initial concentrations")
    - Different reactions have different K values, and K values differ at different temperatures (and for a given system, K values differ ONLY with temperatures)
    - Value of K and the reaction
        - If K>1, we say the forward reaction (forming the products) is favored
        - If K=1 neither one is favored
        - If K<1, then the reverse reaction is favored
    - Q vs. K
        - If Q<K, then the reaction will go forward to achieve equilibrium (and form more of the product than before)
        - If Q>K, then the reaction will go reverse to achieve equilibrium (and form more of the reactant than before)
    - We do not include pure liquids or pure solids in the calculation of K or Q
    - Go through some problems together as an example—use the reaction from the flash (NO_2 and N_2O_4), with different initial concentrations
    - 1^{st} page of the Handout

- Le Chatelier's Principle
    - If a change is imposed on a system at equilibrium, the position of the equilibrium will shift in a direction that tends to reduce that change
        - $aA + bB \Leftrightarrow cC + dD$
          If we put in more of the A, then the system will try to reduce the amount of A (since it disturbed the equilibrium), which means we make more of C and D. Once a new equilibrium has reached, the amount of A is likely to be more than the amount before the addition, but still less than the amount after the addition. The reverse situation works the same way when we add one of the products.
    - Examples: coke bottle opening
        - Before opening the bottle, not much bubbles. There are some carbon dioxide molecules in the gas phase and lots are dissolved in the coke. The system is already at equilibrium, so nothing profound happens.
        - As you open the bottle, you get rid of the gas that was charged in the bottle. You got rid of the gases, so according to the Le Chatelier's

Principle, the system will try to minimize this change by bubbling out more carbon dioxide from the coke

- Other real-life examples include blood CO_2 level and acidity, how hemoglobin transports oxygen and CO_2 back and forth
  - http://www.scienceclarified.com/everyday/Real-Life-Chemistry-Vol-2/Chemical-Equilibrium-Real-life-applications.html
- 2nd page of the Handout

- Extension
  - Thermodynamics
  - Solubility constant
  - Acid-base equilibria
  - chemical equilibrium in human body