CHM 152LL, Determination of an Equilibrium Constant — K_f for $Fe(SCN)^{2+}$

Introduction:

Equilibrium is a dynamic state in which the *rate of formation of the products* **is equal to** the *rate of formation of the reactants* (also called the Law of Mass Action). Reactions in chemical equilibrium will remain so until the system is altered by some outside factor, such as removing some material from the product side. The reaction will then counter-balance that stress and restore to the equilibrium concentration constant, $\mathbf{K_c}$.

For the general reaction, a A + b B $\rightleftarrows$ c C + d D, the equilibrium constant is described by:

$$K_c = [C]^c[D]^d/([A]^a[B]^b)$$

Note: Square brackets ([ ]) indicate concentration in units of molarity (*M* or mol/L).

For a given reaction at a given temperature, no matter how a reaction is performed (i.e. A, B, C and D are mixed in arbitrary amounts), the reaction will always reach a state of equilibrium according to its given equilibrium constant, $\mathbf{K_c}$.

To demonstrate this phenomenon, we will study the equilibrium system of iron(III) and thiocyanate ion and forming thiocyanatoiron (III), a complex ion.

$$\underset{\text{iron(III)}}{Fe^{3+}} \quad + \quad \underset{\text{thiocyanate}}{SCN^-} \quad \rightleftarrows \quad \underset{\text{thiocyanatoiron (III)}}{Fe(SCN)^{2+}}$$

When solutions containing iron(III) ion and thiocyanate ion are mixed, the **deep red** thiocyanatoiron(III) complex ion is formed. The starting concentrations of iron(III) ion and thiocyanate ion (SCN^-) decrease. By determining the concentrations of these three chemical species in several solutions, we can calculate the equilibrium constant within experimental tolerance. Since this equilibrium is a complex ion formation equilibrium, the equilibrium constant is given a special name and symbol — formation constant, K_f:

$$K_f = [Fe(SCN)^{2+}] /([\ Fe^{3+}][\ SCN^-])$$

Due to the deep red color of the complex ion $Fe(SCN)^{2+}$ formed, its concentration will be measured by spectrophotometry. Using that information, calculating the concentrations of all of the other species for the equilibrium expression can be accomplished.

Material:

Due to Fe^{3+} ion's instability in neutral water (possibly forming $Fe(OH)^{2+}$ and etc.), all solutions are prepared in 0.1 M HNO_3. Thus, when some solution needed to be diluted, 0.1 M HNO_3 is to be used in place of deionized water, so that the acidic condition is kept uniform throughout the experiment.

0.00200 M $Fe(NO_3)_3$
0.00200 M KSCN
0.10 M HNO_3
Standard $Fe(SCN)^{2+}$ solution [concentration = ________________ (given at the beginning of the lab)]

Procedure:

Mark 4 disposable pipets with permanent markers so that the volume delivered are equal and around 1 mL.

Label 1 as standard $Fe(SCN)^{2+}$, 1 as HNO_3, 1 as $Fe(NO_3)_3$ and 1 as KSCN.

Part I. Calibration

Since **$[Fe(SCN)^{2+}]$** is measured via spectrophotometry, a calibration curve need to be obtained with solutions of know concentration (Please refer to Beer's Law experiment).

Prepare five standard $FeSCN^{2+}$ solutions in 5 clean, dry & labeled test tubes for calibrating the spectrophotometer as indicated in the following table:

Tube #		2	3	4	5	blank
droppers of **standard** $FeSCN^{2+}$ solution	1.0	2.0	3.0	4.0	5.0	0
droppers of HNO_3 **?**						
Total volume in droppers	5.0	5.0	5.0	5.0	5.0	5.0
$[Fe(SCN)^{2+}]$?						

Rinse 6 empty cuvettes with deionized water and tap as much water out on a paper towel. Then rinse each of the cuvettes with a solution above and then fill it ¾ full with the solution.

Open the *Logger Pro* file titled as *11 Beer's Law* inside the *Chemistry with Vernier* folder. The wavelengths selections for the Colorimeter are 430 nm (violet), 470 nm (blue), 565 nm (green) and 635 nm (red). (What is the right wavelength for the **deep red** thiocyanatoiron(III) complex ion? How can you tell?) Obtain the calibration curve. Print out the graph with the linear regression equation (or save the image file and insert it into a word file and email it to yourselves to be included in the lab report.). **Don't close the program**.

Discard all waste solution in the heavy metal waste bottle, rinse test tubes and the cuvettes clean.

Part II. Obtaining Equilibrium Constant

Label four clean, dry test tubes with identification numbers 1 through 4 and prepare equilibrium systems according to the table below:

Tube #	**1**	**2**	**3**	**4**
droppers of 0.00200M $FeCl_3$	1	2	3	4
droppers of 0.00200M KSCN	4	3	2	1
Total Volume				

Rinse 4 empty cuvettes with deionized water and tap as much water out on a paper towel. Rinse the each of the cuvette with a solution above and fill it ¾ full with the solution.

Measure the absorbance of each of the solutions above. Use calibration curve with calibration equation to figure out the **[$Fe(SCN)^{2+}$]** for each of the reaction mixtures.

The initial concentrations of Fe^{3+} and SCN^- are determined by how the solutions are mixed together. The equilibrium concentration of $Fe(SCN)^{2+}$was obtained via the absorbance data. ICE table should be setup for the calculation of the measure K_f for each of the mixtures above. Do they agree with each other? What is the average? What is the Range of your findings? Your experiments are disagreeing by what percentage?

Again, discard all waste solution in the heavy metal waste bottle, rinse test tubes and the cuvettes clean. Return all glassware and equipment back in their drawers.