Dry Lab
Determination of an Equilibrium Constant

Background

In the lecture we have examined the principles behind chemical equilibrium and the mathematics involved in calculating equilibrium constants. In this experiment we will use the same theory to find equilibrium constant. Since there is no instrument designed to do such a measurement we must explore indirect methods. Last semester we used spectrophotometric techniques to measure the concentration of a species in aqueous solution. Since the system we will be looking at in this experiment is colored a visible light spectrum can be of great value.

Data Analysis
The reaction:

$$Fe^{3+}(aq) + SCN^{-}(aq) \leftrightarrows Fe(SCN)^{2+}$$

has been widely used as a freshman chemistry experiment for years due to its brilliant colors and ease to prepare and work with. You will be supplied with a table of data points for the absorbance of $Fe(SCN)^{2+}$ for determining a molar extinction coefficient for the sample. These are data tubes #1–5. These points are a progressive increase in concentration for the sample. By numerical means, see Zumdahl's appendix or the handout from last semester for assistance, you will calculate the constant a in the equation A = abc where A is absorbance, a is the constant (different for all chemicals but a constant in our reaction), b is path length for the cell (1 cm for our instrument), and c is the concentration in molarity.

Next you will actually calculate the equilibrium constant for the above reaction. Ten different equilibrium mixtures (tubes 6–15) were prepared and their absorbance was measured with a visible light spectrophotometer. The equilibrium mixture was prepared by mixing exact volumes of aqueous $Fe(NO3)3$ (this is the source of Fe^{3+} ion) and KSCN (the source of the SCN^{-} ion) in known ratios. The concentrations of these are listed in the data section of this handout. Once mixed equilibrium is established according to the equation above. *This reaction does not go to completion implying that you will have left over starting materials and an unknown amount of product.* To measure the amount of product present is possible using absorbance data. The color of the complex can be measured using a visible light spectrophotometer. This can tell us the color of the product and hence its concentration. The absorbance for each of these samples is also provided.

The calculations for this lab are made much simpler by the use of a computer spreadsheet.

You will have to calculate four things for each of the samples in tubes six through fifteen.

1. The concentration of complex ion formed. The absorbance of the ion was measured so you can calculate the concentration using Beer's Law.

2. The remaining iron ion concentration. Since we know the concentration of the product we know how much iron was used. This means we can subtract that from how much we started with.

3. The remaining SCN^{-} concentration. Done the same way as the iron.

4. The equilibrium constant for this trial. Write the formula for Keq and substitute numbers. We will use the linear regression method discussed in the library handout.

Pre–Lab Questions:
1) Write the equilibrium constant for this reaction.
2) What is Beer's Law?
3) What do each of the letters in Beer's Law represent?
4) Why can we use visible light spectroscopy to study this product? How is it different from the reactants?
5) What is actually equal at equilibrium?

Data:
Absorbencies for the fifteen samples:

Tube	Absorbance	Tube	Absorbance	Tube	Absorbance
1.	0.080	6.	0.120	11.	0.243
2.	0.179	7.	0.198	12.	0.330
3.	0.253	8.	0.268	13.	0.461
4.	0.319	9.	0.324	14.	0.542
5.	0.400	10.	0.351	15.	0.695

Concentration Data:

Tube #	$[Fe(SCN)^{2+}]$
1.	1.0×10^{-5}
2.	2.0×10^{-5}
3.	3.0×10^{-5}
4.	4.0×10^{-5}
5.	5.0×10^{-5}

Calculate the value of the molar extinction coefficient before you continue.

Concentration data for calculating K.

Test Tube	Initial $[Fe^{3+}]$	Initial $[SCN^{-}]$	Equilibrium $[Fe(SCN^{2+}]$	Equilibrium $[Fe^{3+}]$	Equilibrium $[SCN^{-}]$	Keq
6	3.6×10^{-4}	3.6×10^{-4}				
7	3.6×10^{-4}	5.4×10^{-4}				
8	3.6×10^{-4}	7.1×10^{-4}				
9	3.6×10^{-4}	8.9×10^{-4}				
10	3.6×10^{-4}	1.1×10^{-3}				
11	7.1×10^{-4}	3.6×10^{-4}				
12	7.1×10^{-4}	5.4×10^{-4}				
13	7.1×10^{-4}	7.1×10^{-4}				
14	7.1×10^{-4}	8.9×10^{-4}				
15	7.1×10^{-4}	1.1×10^{-3}				

Post Lab Questions

1) What is the purpose of the first five test tubes where you already know the concentration of the complex ion?

2) What is the value of the molar extinction coefficient you calculated and its units?

3) What is the purpose of performing ten trials to get the equilibrium constant?

4) What is the average equilibrium constant for all of your trials?

5) Imagine that you touched the test tube containing the complex ion and left a finger print on it. How would this effect your values for the following:

    a) Absorbance
    b) Complex ion concentration
    c) Equilibrium constant