Chemical Equilibrium:  
Finding a Constant, Kc

The purpose of this lab is to experimentally determine the equilibrium constant, Kc, for the following chemical reaction:

Fe3+(aq) + SCN–(aq) **** FeSCN2+(aq)

iron(III) thiocyanate thiocyanoiron(III)

When Fe3+ and SCN- are combined, a dynamic equilibrium is established between these two ions and the FeSCN2+ ion. In order to calculate the equilibrium constant, Keq, for the reaction, it is necessary to know the concentrations of all the ions at equilibrium. In this experiment four separate equilibrium systems, or trials, containing different concentrations of these three ions (Fe3+, SCN-, and FeSCN2+) will be determined experimentally. The values for these equilibrium concentrations will be substituted into the equilibrium constant expression to see if Keq is indeed constant despite varied initial concentrations for the reactants. The Keq, is determined by using the Law of Mass Action.

aA + bB ⇔ cC + dD

This equation gives the equilibrium constant expression of:

Keq = [C]c[D]d/[A]a[B]b

In order to determine the equilibrium concentrations for the three ions a *standard solution* needs to be prepared. To prepare the standard solution, a very large concentration of Fe3+ will be added to a small initial concentration of SCN– (hereafter referred to as [SCN-]i. The initial [Fe3+] in the standard solution is 900 times larger than [SCN-]i. According to LeChatelier's principle, which states that when a system in dynamic equilibrium is disturbed, the system responds so as to minimize the disturbance and return the system to a state of equilibrium. This high initial concentration Fe3+ ions on the left side of the equation forces the reaction far to the right, using up nearly 100% of the initial SCN– ions. Using stoichiometry and the balanced equation, it is assumed that for every mole of FeSCN2+ produced, one mole of SCN– is used up. Thus since nearly all of the SCN- ions are consumed in order to minimize the disturbance, the product’s concentration, [FeSCN2+]std, at equilibrium is assumed to be equal to the [SCN–]i.

Since the reaction produces the FeSCN2+ ions and this ion transmits the color red, the solution’s absorbance of blue light can be measured through the use of a colorimeter (see Figure 1). Because the red solutions absorb blue light very well, the blue LED setting on the Colorimeter is used. The computer-interfaced Colorimeter measures the amount of blue light absorbed by the colored solutions (absorbance, A).





*Figure 1* *Figure 2*

According to Beer’s Law, there is a direct relationship between a solution’s concentration and its absorbance. In this case the concentration is the FeSCN2+ ion and the absorbance is blue light (470 nm). In other words, as the concentration of FeSCN2+ increases so will the absorbance of blue light (see Figure 2). The concentration of FeSCN2+ for any of the equilibrium systems, trials 1-4, can be found by comparing the absorbance of each equilibrium system, *Aeq*, to the absorbance of the *standard* solution, *Astd*, according to the following equation:

[FeSCN2+]*std*/A*std*= [FeSCN2+]*eq*/ A*eq*

Since the concentration of [FeSCN2+]*std* is known and the all of the absorbances for the equilibrium solutions and the standard are measured and recorded all that needs to be done is to solve for the unknown.

[FeSCN2+]eq = X [FeSCN2+]std

Knowing the [FeSCN2+]*eq* allows you to determine the concentrations of the other two ions at equilibrium. For each mole of FeSCN2+ ions produced, one less mole of Fe3+ and SCN- ions will be found in the solution (see the 1:1 ratio of coefficients in the equation on the previous page). At equilibrium the [Fe3+] and [SCN-] can be determined according to the following equations:

[Fe3+]eq = [Fe3+]i – [FeSCN2+]eq

[SCN–]eq = [SCN–]i – [FeSCN2+]eq

Knowing the values of [Fe3+]eq, [SCN–]eq, and [FeSCN2+]eq, you can now calculate the value of *Kc*, the equilibrium constant.

OBJECTIVE

In this experiment, you will determine the equilibrium constant, *Kc*, for the following chemical reaction:

Fe3+(aq) + SCN–(aq) **** FeSCN2+(aq)

iron(III) thiocyanate thiocyanoiron(III)

MATERIALS

|  |
| --- |
| Vernier Colorimeter |
| 1 plastic cuvette |
| five 20 🞨 150 mm test tubes |
| pipet bulb or pipet pump |
| Serological pipet |

PROCEDURE

1. Obtain and wear goggles.

2. Label four 20 🞨 150 mm test tubes 1-4. Pour about 30 mL of 0.0020 M Fe(NO3)3 into a clean, dry 100 mL beaker. Pipet 5.0 mL of this solution into each of the four labeled test tubes. Use a pipet pump or bulb to pipet all solutions. **CAUTION:** *Fe(NO*3*)*3 *solutions in this experiment are prepared in 1.0 M HNO*3 *and should be handled with care.* Pour about 25 mL of the 0.0020 M KSCN into another clean, dry 100 mL beaker. Pipet 2, 3, 4 and 5 mL of this solution into Test Tubes 1-4, respectively. Obtain about 25 mL of distilled water in a 100 mL beaker. Then pipet 3, 2, 1 and 0 mL of distilled water into Test Tubes 1-4, respectively, to bring the total volume of each test tube to 10 mL. Mix each solution *thoroughly* with a stirring rod. Be sure to clean and dry the stirring rod after each mixing. Measure and record the temperature of one of the above solutions to use as the temperature for the equilibrium constant, *Kc*. Volumes added to each test tube are summarized below:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Test Tube Number | Fe(NO3)3 (mL) | KSCN (mL) | H2O (mL) |
|  | 1 | 5 | 2 | 3 |
|  | 2 | 5 | 3 | 2 |
|  | 3 | 5 | 4 | 1 |
|  | 4 | 5 | 5 | 0 |

3. Prepare a standard solution of FeSCN2+ by pipetting 18 mL of 0.200 M Fe(NO3)3 into a 20 🞨 150 mm test tube labeled “5”. Pipet 2 mL of 0.0020 M KSCN into the same test tube. Stir thoroughly.

4. Connect the Colorimeter to the computer interface. Prepare the computer for data collection by opening the file “Lab 8 Equilibrium” from the *Chemistry 228* folder of Logger *Pro*

5. Prepare a blank by filling a cuvette 3/4 full with distilled water. To correctly use a Colorimeter cuvette, remember:

* All cuvettes should be wiped clean and dry on the outside with a tissue.
* Handle cuvettes only by the top edge of the ribbed sides.
* All solutions should be free of bubbles.
* Always position the cuvette with its reference mark facing toward the white reference mark at the top of the cuvette slot on the Colorimeter.

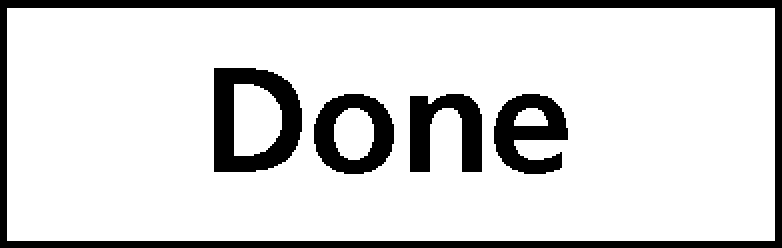
6. Calibrate the Colorimeter.

1. Open the Colorimeter lid.
2. Holding the cuvette by the upper edges, place it in the cuvette slot of the Colorimeter. Close the lid.
3. If your Colorimeter has a CAL button, Press the < or > button on the Colorimeter to select a wavelength of 470 nm (Blue) for this experiment. Press the CAL button until the red LED begins to flash. Then release the CAL button. When the LED stops flashing, the calibration is complete. Proceed directly to Step 7. If your Colorimeter does not have a CAL button, continue with this step to calibrate your Colorimeter.

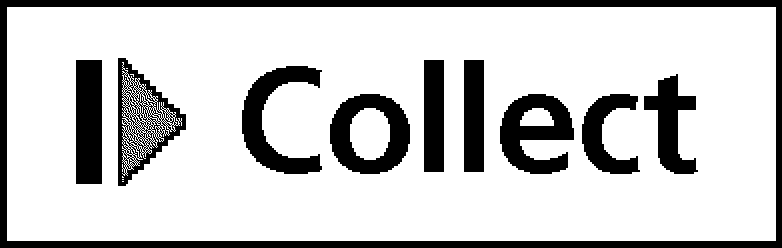
First Calibration Point

1. Choose Calibrate 🞂 CH1: Colorimeter (%T) from the Experiment menu and then click .
2. Turn the wavelength knob on the Colorimeter to the “0% T” position.
3. Type “0” in the edit box.
4. When the displayed voltage reading for Reading 1 stabilizes, click .

Second Calibration Point

1. Turn the knob of the Colorimeter to the Blue LED position (470 nm).
2. Type “100” in the edit box.
3. When the displayed voltage reading for Reading 2 stabilizes, click , then click .

7. You are now ready to collect absorbance data for the four equilibrium systems and the standard solution.

1. Click  to begin data collection.
2. Empty the water from the cuvette. Rinse it twice with ~1 mL portions of the Test Tube 1 solution.
3. Wipe the outside of the cuvette with a tissue and then place the cuvette in the Colorimeter. After closing the lid, wait for the absorbance value displayed in the meter to stabilize. Then click , type “1” (the trial number) in edit box, and press the ENTER key.
4. Discard the cuvette contents as directed by your teacher. Rinse the cuvette twice with the Test Tube 2 solution and fill the cuvette 3/4 full. Follow the Step-c procedure to find the absorbance of this solution. Type “2” in the edit box and press ENTER.
5. Repeat the Step-d procedure to find the absorbance of the solutions in Test Tubes 3, 4, and 5 (the standard solution).
6. From the table, record the absorbance values for each of the five trials in your data table.
7. Dispose of all solutions as directed by your instructor.

PROCESSING THE DATA

1. Write the *Kc* expression for the reaction in the Data and Calculation table.

2. Calculate the initial concentration of Fe3+, based on the dilution that results from adding KSCN solution and water to the original 0.0020 M Fe(NO3)3 solution. See Step 2 of the procedure for the volume of each substance used in Trials 1-4. Calculate [Fe3+]i using the equation:

[Fe3+]i = 🞨 (0.0020 M)

This should be the same for all four test tubes.

3. Calculate the initial concentration of SCN–, based on its dilution by Fe(NO3)3 and water:

[SCN–]i = 🞨 (0.0020 M)

In Test Tube 1, [SCN–]i = (2 mL / 10 mL)(0.0020 M) = 0.00040 M. Calculate this for the other three test tubes.

4. [FeSCN2+]eq is calculated using the formula:

[FeSCN2+]eq = 🞨 [FeSCN2+]std

where *Aeq* and *Astd* are the absorbance values for the equilibrium and standard test tubes, respectively, and [FeSCN2+]std = (1/10)(0.0020) = 0.00020 M. Calculate [FeSCN2+]eq for each of the four trials.

5. [Fe3+]eq: Calculate the concentration of Fe3+ at equilibrium for Trials 1-4 using the equation:

[Fe3+]eq = [Fe3+]i – [FeSCN2+]eq

6. [SCN–]eq: Calculate the concentration of SCN- at equilibrium for Trials 1-4 using the equation:

[SCN–]eq = [SCN–]i – [FeSCN2+]eq

7. Calculate *Kc* for Trials 1-4. Be sure to show the *Kc* expression and the values substituted in for each of these calculations.

8. Using your four calculated *Kc* values, determine an average value for *Kc.* How constant were your *Kc* values?

**Equilibrium Lab Report:**

Your report for this lab should include the following sections:

Abstract:

Your abstract should be written individually

Data:

Include your data table with initial concentrations of each reactant for each trial

Results:

Include a results table with calculated Keq values for each trial and an average value for Keq

Be sure to attach hand written sample calculations to the back of your report

Discussion:

Discuss **the experiment** and any possible sources of error

In addition, answer the following question as part of your report:

1. How are you Keq values to each other? Are they close enough to justify the assertion that an equilibrium constant is constant?
2. What factors could have led to variations in Keq between trials.