

EXPERIMENT 3

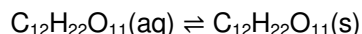
Equilibrium Constants

Introduction

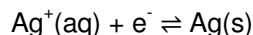
Most of Chemistry 101 was involved with limiting reagents, reactants turning into products and calculating the amount of products produced based on the stoichiometry of a reaction. This was possible because all of the reactions studied went to completion, that is, the reaction continued until one or more of the reactants were gone and no more products could be made.

However, not all reactions go to completion. In Chemistry 112, a great deal of emphasis is placed on the study of equilibrium reactions. An **equilibrium reaction** is a reaction that moves in the forward direction to produce products, but not all of the reactants will become products. It does not go to completion, so stoichiometric calculations cannot be used. In fact, to be a true equilibrium, there must be at least a small amount of every product and every reactant present in the solution. One example of an equilibrium reaction would be maintaining pH balance in a swimming pool by adding either baking soda to increase the pH or an acid to lower the pH. Other examples of manipulating equilibrium are shown below:

Making rock candy by dissolving sugar in hot water and allowing it to precipitate out when cooled. If more water is added, the sugar would dissolve back to its aqueous form. If water evaporates, more sugar precipitates.



Electroplating silver over a steel spoon. Adding electricity converts the silver ion in solution to a solid. Reversing the current would remove the solid silver coating and turn the silver atoms back into silver ions.

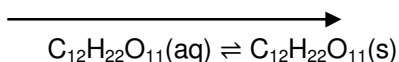


These reactions are very different from one another, but both involve reactions that move back and forth between reactants and products.

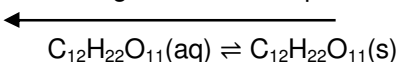
Equilibrium Reactions

Equilibrium reactions depend on a balance between products and reactants. Like a seesaw, if you add unequal weights to each side, the side with the heavier weight will go down while the other side rises to compensate for the extra weight. If left alone for a few minutes, the seesaw will stabilize and no longer move in either direction. In scientific terms, you can say it has reached a state of equilibrium, where no further movement is seen on either side.

Chemical reactions also reach a state of equilibrium where no further reaction appears to take place. For example, if you reach an equilibrium state with sugar in water, you will have some sugar in solution and some solid sugar on the bottom of the beaker. If you boil off some of the water, you increase the concentration of sugar in the solution. If the concentration becomes too high, the water must release some of the sugar as solid and it will precipitate out on the bottom of the beaker. The reaction goes forward and produces more of the product, $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$.



However, if you add more water, you dilute the sugar in the solution and more solid sugar will dissolve to compensate for the loss. The reaction will now go in reverse to produce more of your reactant, $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq})$.



This shifting back and forth is the hallmark of an equilibrium reaction and results in much more complicated calculations than those performed in Chemistry 191.

But before calculations can be performed with equilibrium reactions, a few new concepts must be introduced that relate the amount of products produced from a given amount of reactants. First, the reaction itself must be defined in such a way as to distinguish an equilibrium reaction from one that goes to completion. This is done by substituting a double arrow for the single forward arrow in the chemical equation to signify that the reaction can go in either direction; Reactants \rightleftharpoons Products. The ratio between the amounts of reactants and products will always be held constant, so that if you add additional product, the reaction will go backward to form more reactants. If product is removed from the solution, (for example, a gas escapes); the reactants will continue to form products to try to reestablish the reactant/product ratio. The second concept introduced is that numerical value for the ratio between products and reactants. This value is called the **equilibrium constant** and the symbol K_{eq} is used to represent this constant in mathematical formulas.

The Equilibrium Expression and K_{eq}

The equilibrium constant in an equilibrium reaction is an experimentally determined number that uses the stoichiometry of the reaction to relate the products and reactants to each other. Therefore, if you know the chemical reaction governing the equilibrium, you can always write the equilibrium expression. The format for an equilibrium expression is as follows:

For a reaction; $aA + bB \rightleftharpoons cC + dD$

$$K_{eq} = \frac{(\text{Products})^{\text{coeff}}}{(\text{Reactants})^{\text{coeff}}} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

The first step in determining the equilibrium expression for a reaction is to be able to write the chemical equation associated with the equilibrium. Note that the equation **MUST** be balanced correctly or your error will be magnified exponentially, as your stoichiometric coefficients from the chemical equation become exponential variables in the equilibrium expression. If you are calculating the equilibrium constant from experimentally determined equilibrium concentrations, your value for K_{eq} will quickly become meaningless unless the coefficients in the chemical reaction have been balanced correctly.

K_{eq} can be measured in many units, but they must be defined from the chemical equation, so that a researcher can duplicate the same product to reactant ratio and obtain the same value for K_{eq} . The most commonly used units when calculating K_{eq} are concentration and pressure. Pressure would be used if working with gas equilibria and the equilibrium constant would be rewritten as K_p to define that all values are units of pressure. We will be working primarily with concentration, as most of the reactions we do are in aqueous solutions, and the equilibrium constant, K_{eq} will be redefined as K_c , the concentration equilibrium constant.

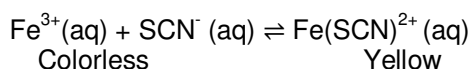
Concentration Equilibrium Constant, K_c

While many units can be used when describing equilibrium, solution concentration is one of the most common. To define your equation in terms of concentration, rather than pressure, mass, etc., a defined symbol signifying that the equilibrium constant is measured in concentration is used. This symbol, K_c , simply means that you must base your stoichiometry and units on molar concentration and write the K_{eq} as in the previous section. The new symbol, K_c , is then tabulated for many different reactions, allowing a scientist to simply look up the K_c for a reaction and be able to manipulate an experiment to obtain the equilibrium concentrations desired.

Calculating the Equilibrium Constant for the Reaction of Iron (III) Nitrate with Sodium Thiocyanate

The Reaction

In your experiment, you will be reacting a solution containing colorless iron (III) ions with a solution containing thiocyanate, SCN^- . The product produced, iron(III)thiocyanate ion ($Fe(SCN)^{2+}$) is yellow. Like any colored species in an aqueous solution, the higher the concentration of $Fe(SCN)^{2+}$ present in the solution, the darker the color. The chemical equation associated with this reaction is given below.



The amount of product produced from the reactants will be governed by the equilibrium expression. No matter how much reactant is present, you will generate a corresponding amount of product according to the following mathematical expression:

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

To find K_c , you must first know the equilibrium concentrations of the 2 reactants and the product. Once you know these values, you just plug them into the equilibrium expression and solve for K_c .

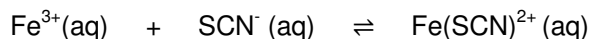
Equilibrium versus Non-Equilibrium Concentrations

The concentrations of reactants and products in the equilibrium constant expression must always be the concentrations at equilibrium. In general, the initial concentrations given in a problem will not be the concentration of a species once the reaction has reached equilibrium, thus making it very easy to make a mistake when performing calculations. While you can use these concentrations as a starting point, the concentrations will change as the equilibrium shifts to match the K_c . However, if you know the concentration of one of the species, either reactant or product, at equilibrium, you can calculate the concentrations of the others based on the stoichiometry of the reaction. One way of doing these calculations is to introduce the use of an ICE table.

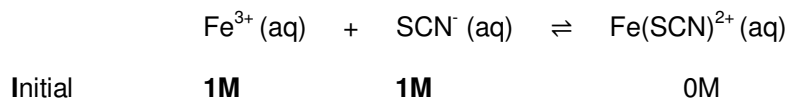
ICE Tables

An ICE table is a tabulated means of keeping track of concentrations at the start of a reaction and their relationship to the equilibrium concentrations that result when a reaction has reached equilibrium. It is based on the fact that while you may not know the equilibrium concentrations of your products and reactants, they can be calculated based on knowing the equilibrium concentrations of only one of the species in the reaction. The three letters, ICE stand for Initial, Change and Equilibrium.

To set up an ICE table for your lab, start with writing down your chemical reaction.

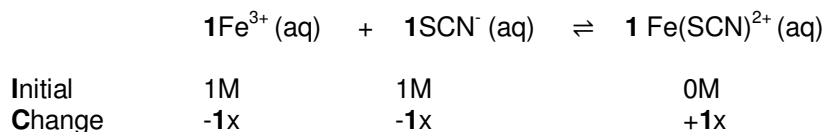


The first line in your ICE table will contain the initial concentrations of any of the species present. If you start with **1M** $\text{Fe}^{3+}(\text{aq})$ and **1M** $\text{SCN}^{-}(\text{aq})$, these are your initial concentrations before the reaction. No $\text{Fe}(\text{SCN})^{2+}(\text{aq})$ has yet been formed, so the concentration of your product is zero. It is convenient to write the concentrations in first line of the ICE table directly below the chemicals to which they correspond. The first line of your ICE table would be as follows:



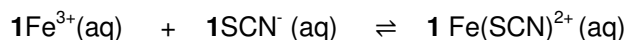
Any of the amounts listed in a problem should be considered to be initial concentrations unless they are specifically designated as equilibrium concentrations. Also note that while the amount of initial product in the ICE table is zero for this particular situation that is not always the case. It is possible to start with some product already present before equilibrium is established. However, if a chemical is not specifically listed in the problem, there is a good chance the initial concentration is zero.

The second line of the ICE table relates the stoichiometry between products and reactants. For this reaction, for every mole of product you produce, you use up 1 mole of reactant. Thus, if you produce x moles of products, you will have used up x moles of reactants. This change is represented in the ICE table below:



The sign in front of the x designates whether the concentration is increasing or decreasing. In the case of reactants going forward to products you decrease by the amount x (-) and increase the products also by x (+). The coefficient of 1 before each of the x values designates the mole ratio from the chemical equation. In your reaction, all coefficients are 1. You must always include the coefficient before the x on the Change line in the ICE table.

The Equilibrium concentrations are the result of how much of a reactant or product you start with in the reaction and how the concentration changes based on the equilibrium. Thus, you add the two values together to arrive at the value or equation found for each concentration in the equilibrium line of the ICE table.



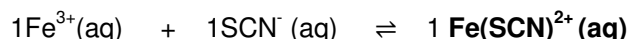
Initial	1M	1M	0M
Change	-1x	-1x	+1x
Equilibrium	(1-1x)	(1-1x)	(0+1x)

The values or formulas written on the equilibrium line are then substituted into the equilibrium constant expression. At this point, you can solve for the equilibrium constant K_c if you can solve for x in the equation, as you now have the algebraic relationship between all the species in the chemical equation.

$$K_c = \frac{[\text{Fe}(\text{SCN})^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^{-}]} = \frac{[(0+1x)]}{[(1-1x)][(1-1x)]} = \frac{[x]}{[(1-1x)][(1-1x)]}$$

Using ICE Table Data

At this point, you now have an algebraic equation that relates K_c and all of the equilibrium concentrations to a single variable, x . Unfortunately, at this time, you do not have a value for x . To find x , you first need to use the ICE table to see how it correlates to the reactants and products in your equilibrium. Looking in the products column, the equilibrium amount of $\text{Fe}(\text{SCN})^{2+}(\text{aq})$ corresponds to the x value in the equilibrium constant expression. Now you need to find a means of measuring the concentration of $\text{Fe}(\text{SCN})^{2+}(\text{aq})$ produced in your reaction.



Initial	1M	1M	0M
Change	-1x	-1x	+1x
Equilibrium	(1-1x)	(1-1x)	(x)

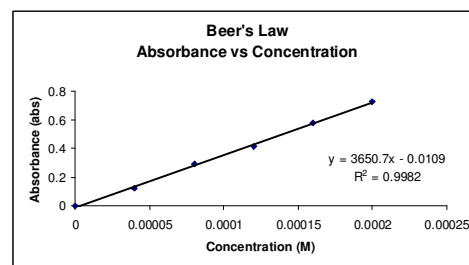
Beer's Law

In experiment 2, you used the fact that the crystal violet became gradually lighter as it reacted with sodium hydroxide. The degree of color was measured as absorbance in a colorimeter and this absorbance was generally related to the color of the solution at the time it was measured. In this experiment, you need a more exact measurement of concentration rather than just a relative correlation. In terms of spectroscopy, this relationship is called Beer's Law:

$$A = \epsilon bc$$

Beer's Law is an equation that matches the absorbance of a colored solution to the concentration of that solution. The variables are as follows:

- A is the absorbance of a solution being measured
- ϵ is a proportionality constant called the molar absorptivity
- b is the pathlength of the cuvette
- c is the concentration of the solution.

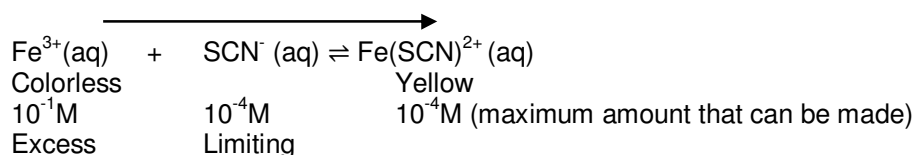


A plot of absorbance versus concentration will result in a graph similar to that shown on the right. This method is extremely accurate as can be seen by the high correlation coefficient, a very effective means of relating absorbance to concentration.

Determination of Equilibrium Concentrations from Graphed Data

By plotting the absorbance measured from the colorimeter against the concentration of the $\text{Fe}(\text{SCN})^{2+}$ in the solution, you can obtain a linear relationship that will allow you to calculate the concentration of $\text{Fe}(\text{SCN})^{2+}$ in an unknown solution. But first you need to be able to make the standard curve, which means having 6 standards, each of which has a known concentration of $\text{Fe}(\text{SCN})^{2+}$.

To do this, you must first force the equilibrium towards products as much as possible by using a large amount of one of your reactants. This forces your other reactant to act as a limiting reagent. Since the equilibrium will be trying to make as much product as possible to use up as much reactant as possible, nearly all of the limiting reagent will be exhausted and instead of equilibrium, you can treat the chemical reaction as one that essentially goes to completion. The amount of limiting reagent left is negligible compared to the amount of product that has been produced.



To generate your calibration curve, you regularly change the concentration of the limiting reagent measure the absorbance that is due to the yellow color of your product. A low concentration of the limiting reagent generates a lower concentration of your yellow product, which will appear lighter in color. The lighter the color, the lower the absorbance. When all points are plotted, the linear relationship between the concentration of the product ($\text{Fe}(\text{SCN})^{2+}$) will result.

Once you have your plot, you can use the equation for a line to determine the concentration of $\text{Fe}(\text{SCN})^{2+}$ in an equilibrium solution. To do this, solve for x in the equation for a line. The y value is the absorbance reading corresponding the the solution of unknown concentration. For the example above, the equation for the line was:
 $y = 3650.7x - 0.0109$

Solving for x if your absorbance was 0.35 abs would give you a concentration of $9.9 \times 10^{-5}\text{M}$. This is the value for x which is also the value of the unknown concentration of $\text{Fe}(\text{SCN})^{2+}$.

Going back to your ICE table, you can now fill in for x and solve for the remaining equilibrium concentrations of Fe^{3+} and SCN^{-} . Note the initial concentrations for both reactants must correspond to the initial concentrations used for your particular equilibrium solution. In this example, both the $\text{Fe}^{3+}(\text{aq})$ and the $\text{SCN}^{-}(\text{aq})$ were initially $1 \times 10^{-4}\text{M}$. Also note, these numbers are not representative of the data you will be obtaining, so be sure to use your own data in the calculations. **DO NOT USE THESE NUMBERS FOR YOUR CALCULATIONS FROM THE EXPERIMENT. THIS IS ONLY AN EXAMPLE, NOT REAL DATA.**

	$1\text{Fe}^{3+}(\text{aq})$	+	$1\text{SCN}^{-}(\text{aq})$	\rightleftharpoons	$1\text{Fe}(\text{SCN})^{2+}(\text{aq})$
Initial	$1 \times 10^{-4}\text{M}$		$1 \times 10^{-4}\text{M}$		0M
Change	-1x		-1x		+1x
Equilibrium	$(1 \times 10^{-4} - 9.0 \times 10^{-5})$		$(1 \times 10^{-4} - 9.9 \times 10^{-5})$		(9.9×10^{-5})

Calculation of Equilibrium Constants

Once you have the equilibrium concentrations of each reactant and product, you substitute them into the equilibrium constant expression and solve for K_c . Note: The K_c from your experimental data will vary from the K_c that you would calculate from this data.

$$K_c = \frac{[\text{Fe}(\text{SCN})^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^{-}]} = \frac{[9.0 \times 10^{-5}]}{[(1 \times 10^{-4} - 9.0 \times 10^{-5})][1 \times 10^{-4} - 9.0 \times 10^{-5}]}$$

Laboratory Equipment Procedures

Using the MeasureNet System for Colorimetry

The colorimeter should already be connected to the MeasureNet system at the start of lab. Follow the instructions below to set up for absorbance measurements.

Colorimeter Calibration: Sets distilled water to read 0 abs units

1. Fill a wash bottle with distilled water.
2. Use the wash bottle to fill 2 cuvettes with distilled water.
 - a. Remember to avoid touching the sides of the cuvette near the midpoint.
3. Dry the sides of the cuvettes with a kimwipe.
4. Gently place the 2 cuvettes containing the water in the reference and sample cell holders in the colorimeter. See your TA if they will not go in easily, **DO NOT FORCE THEM INTO THE HOLDER.**
5. Close the lid of the colorimeter
6. Press the following keys on the workstation to start the kinetics program.
 - a. Press **ON/OFF** on the workstation.
 - b. Press **MAIN MENU**.
 - c. Press **F6**- Colorimetry/Fluor./Trb./Lum.
 - d. Press **F1**-Colorimetry.
 - e. Press **F2**-Green LED (wavelength = 515 nm)
 - f. Press **F1**- Kinetics
 - g. Press **ENTER** to calibrate your zero point on the Y axis (absorbance)
7. Wait until green light shows on colorimeter and you hear a beep acknowledging that the calibration is complete.
Press **DISPLAY** to accept the new value for the new settings.
8. Remove the sample cuvette (S) and dump the water into the sink. Do not remove the reference cuvette.
9. Dry the inside of the sample cuvette with a kimwipe before proceeding to the next section.

Sample Analysis

1. You will always be using the sample cuvette (S) when measuring any absorbance.
2. Make sure the cover is completely closed before recording absorbance values.
3. Remember to avoid touching the sides of the cuvet near the midpoint.
4. If making a series of absorbance values, you do not need to press any further buttons on the workstation, just insert the cuvette and record the values.

Recalibrating the Colorimeter

5. Fill the sample cuvette with distilled water.
6. Insert into colorimeter.
7. Close the cover.
8. Press **CALIBRATE**
9. Press **F2**
10. Press **F1**
11. Press **ENTER**
12. After beep, press **DISPLAY**
13. Remove sample cuvette from colorimeter and you are ready to collect data.

Using Excel for Data Analysis

See Experiment 2: Reaction Order for instructions for graphing with Excel.

Chemical Hazards

Iron(III) Nitrate Solution

NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 0

DERMAL EXPOSURE:

Wash off with soap and plenty of water.

EYE EXPOSURE:

Flush with water. Contact your TA immediately.

0.1M Nitric Acid Solution

NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 0

DERMAL EXPOSURE:

Wash off with soap and plenty of water.

EYE EXPOSURE:

Flush with water for at least 15 minutes. Contact your TA immediately.

Sodium Thiocyanate Solution

NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 1

ORAL EXPOSURE

Poisonous. If swallowed, wash out mouth with water provided person is conscious. Do not induce vomiting. Contact your TA immediately.

DERMAL EXPOSURE

Slightly harmful. In case of extensive skin contact, flush with water for at least 15 minutes. Remove contaminated clothing and shoes. Contact your TA immediately.

EYE EXPOSURE

Poisonous. In case of contact with eyes, flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately.

REACTIVITY

At concentrations above 1M, contact with strong acids, strong bases and oxidizing agents may liberate toxic gas.

Experimental Procedures

Part 1: Preparation of standard solutions

1. Record the concentrations of 0.20M $\text{Fe}(\text{NO}_3)_3$ and 2.00×10^{-3} M NaSCN solutions on your data sheet.
2. Label six test tubes "S1" through "S6" and place in test tube rack. You can use a paper labeled S1-S6 under the test tube rack if necessary.
3. Transfer 20 drops of 0.20M $\text{Fe}(\text{NO}_3)_3$ solution into the test tubes labeled S1-S6
4. Transfer the exact amount of 2.00×10^{-3} M solution NaSCN solution listed table below into each of the corresponding test tubes.

S1	0.0 drops	S4	6.0 drops
S2	2.0 drops	S5	8.0 drops
S3	4.0 drops	S6	10.0 drops

5. Transfer the exact amount of distilled water listed table below into each of the corresponding test tubes.

S1	60.0 drops	S4	54.0 drops
S2	58.0 drops	S5	52.0 drops
S3	56.0 drops	S6	50.0 drops

Part 2: Preparation of equilibrium mixtures

1. Record the exact concentration of the 2.0×10^{-3} M $\text{Fe}(\text{NO}_3)_3$ solution on your data sheet.
2. Label six test tubes "E1" through "E6" and place in test tube rack. You can use a paper labeled E1-E6 under the test tube rack if necessary.
3. Transfer exactly 40 drops of 2.0×10^{-3} M $\text{Fe}(\text{NO}_3)_3$ into the test tube labeled E1-E6.
4. Transfer the exact amount of 2.00×10^{-3} M solution NaSCN solution listed table below into each of the corresponding test tubes.

E1	0.0 drops	E4	20.0 drops
E2	5.0 drops	E5	30.0 drops
E3	10.0 drops	E6	40.0 drops

5. Transfer the exact amount distilled water listed table below into each of the corresponding test tubes.

E1	40.0 drops	E4	20.0 drops
E2	35.0 drops	E5	10.0 drops
E3	30.0 drops	E6	0.0 drops

Colorimeter Analysis

1. Calibrate the colorimeter according to the instructions given in the Equipment Procedures.
2. Analyze all your standards starting from the least concentrated standard and moving to the most concentrated using the instructions given for sample analysis in the Equipment Procedures.
3. Pour your standard back into the appropriate test tube when finished recording the absorbance.
4. Repeat for each of the equilibrium mixtures.

Cleanup

1. Pour all remaining solutions into waste beaker.
2. Empty waste beaker into waste container in hood.
3. Rinse all equipment, dry and return to lab drawer.

Laboratory Data

Create the following data tables in your lab notebook before coming to class. Include the correct number of significant figures for each measurement.

Concentration of stock solution of $\text{Fe}(\text{NO}_3)_3$ $[\text{Fe}^{3+}]$ _____ M
Concentration of stock solution of NaSCN $[\text{SCN}^-]$ _____ M

Table 1: Standard mixtures: Record the absorbance values of solutions S1 through S6

Table 2: Equilibrium mixtures: Record the absorbance values of solutions E1 through E6

Calculations

Volumes of Standard and Equilibrium Solutions

Copy values from lab instructions for the volume of each of the three chemicals in each standard.

Total Volumes of Standard and Equilibrium Solutions

Add the volumes of the $\text{Fe}(\text{NO}_3)_3$, NaSCN and distilled water used to make each sample. The total volume will come to the same value for all standards.

Initial Concentrations of Ions ($[\text{Fe}^{3+}]$ and $[\text{SCN}^-]$) in Standard and Equilibrium Solutions

This is a dilution using the equation: $M_1V_1 = M_2V_2$ and solve for M_2 .

$$M_2 = M_1V_1 / V_2$$

M_1 = concentration of stock solution

M_2 = concentration of final solution (standard)

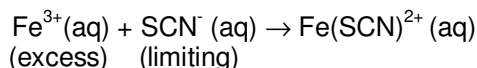
V_1 = volume of stock solution

V_2 = volume of final solution (standard)

To find the concentration of each standard solution, multiply the concentration of the stock solution by the volume of the stock solution and divide by the total volume calculated for the standard solution. Find the initial concentrations for all standards for both the Fe^{3+} and SCN^- .

Equilibrium Concentration of $\text{Fe}(\text{SCN})^{2+}$ in Standard Solutions

With such an excess of $\text{Fe}^{3+}(\text{aq})$ in the standard solutions, essentially all of the $\text{SCN}^-(\text{aq})$ is converted to $\text{Fe}(\text{SCN})^{2+}(\text{aq})$. Thus, this reaction would be considered to go to completion. As a result, the concentration of $\text{SCN}^-(\text{aq})$ in the standard solution would now be equal to the concentration of the product, $\text{Fe}(\text{SCN})^{2+}(\text{aq})$, as the $\text{SCN}^-(\text{aq})$ ion would be the limiting reagent.



Creating a Beer's Law Plot

Create a table containing the following information for each of your standards.

Standard	Equilibrium $[\text{Fe}(\text{SCN})^{2+}]$	Absorbance
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Graph this data in Excel (Instructions on graphing in Excel given in Experiment 2) Equilibrium $[\text{Fe}(\text{SCN})^{2+}]$ data on the x-axis and the absorbance data on the y-axis. Include the correlation coefficient and the equation for the line with your plot.

Equilibrium Concentration of $\text{Fe}(\text{SCN})^{2+}$ in Equilibrium Solutions

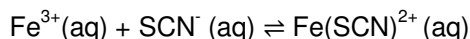
Write down the equation for a line from your Beer's Law Plot. Solve for the equilibrium concentrations of $\text{Fe}(\text{SCN})^{2+}$ for your equilibrium solutions by substituting the absorbance values you recorded for each equilibrium solution as the y values in the equation and solving for x. These numbers correspond to the concentration of $\text{Fe}(\text{SCN})^{2+}$ for each equilibrium solution.

Equilibrium Concentrations of Reactant Ions ($[\text{Fe}^{3+}]$ and $[\text{SCN}^-]$) in Equilibrium Solutions

For every mole of $\text{Fe}(\text{SCN})^{2+}(\text{aq})$ produced, you used up 1 mole of $\text{Fe}^{3+}(\text{aq})$. To account for this loss, subtract the concentration of $\text{Fe}(\text{SCN})^{2+}$ from the initial concentration of $\text{Fe}^{3+}(\text{aq})$ to find the equilibrium concentration of $\text{Fe}^{3+}(\text{aq})$. Do the same for the $\text{SCN}^-(\text{aq})$.

K_c Expression

Write the K_c expression for the following reaction. Remember, products over reactants with each concentration raised to its coefficient.



Value of K_c and Mean K_c

Fill in the equilibrium concentrations you calculated for $\text{Fe}^{3+}(\text{aq})$, $\text{SCN}^-(\text{aq})$ and $\text{Fe}(\text{SCN})^{2+}(\text{aq})$ into the K_c expression. Find the K_c for all 5 of the equilibrium solutions. Ideally, they will be identical. Average the 5 values when you are finished.

Results

Remember to show 1 example of each calculation either handwritten or typed on a separate sheet.

Data determined from Excel does not need to be hand calculated.

Include correct significant figures and units

Units can be included in column or row title as demonstrated in tables below

Brackets, [xx], can be used to designate concentration in molarity.

Include a typed copy of this table with your lab report.

Table 3: Concentration of Ions in Standard Solutions

Calculation	unit	S1	S2	S3	S4	S5	S6
Volume of $\text{Fe}(\text{NO}_3)_3$	(drops)	_____	_____	_____	_____	_____	_____
Volume of NaSCN	(drops)	_____	_____	_____	_____	_____	_____
Volume of distilled water	(drops)	_____	_____	_____	_____	_____	_____
Total volume of solution	(drops)	_____	_____	_____	_____	_____	_____
Initial Concentration of Fe^{3+} ion	$[\text{Fe}^{3+}]$	_____	_____	_____	_____	_____	_____
Initial Concentration of SCN^- ion	$[\text{SCN}^-]$	_____	_____	_____	_____	_____	_____
Equilibrium Concentration of $\text{Fe}(\text{SCN})^{2+}$	$[\text{Fe}(\text{SCN})^{2+}]$	_____	_____	_____	_____	_____	_____

Table 4: Data for Beer's Law Plot

Standard	$[\text{Fe}(\text{SCN})^{2+}]$	Absorbance (au)
S1	_____	_____
S2	_____	_____
S3	_____	_____
S4	_____	_____
S5	_____	_____
S6	_____	_____

This table can be either typed into Excel or into Word, but must be typewritten in your lab report.

Table 5: Equilibrium Calculations

Calculation	unit	E2	E3	E4	E5	E6
Volume of $\text{Fe}(\text{NO}_3)_3$	(drops)	_____	_____	_____	_____	_____
Volume of NaSCN	(drops)	_____	_____	_____	_____	_____
Volume of distilled water	(drops)	_____	_____	_____	_____	_____
Total volume of solution	(drops)	_____	_____	_____	_____	_____
Initial Concentration of Fe^{3+} ion	$[\text{Fe}^{3+}]$	_____	_____	_____	_____	_____
Initial Concentration of SCN^- ion	$[\text{SCN}^-]$	_____	_____	_____	_____	_____
Equation of a line from Beer's Law Plot		_____				
Equilibrium Concentration of $\text{Fe}(\text{SCN})^{2+}$	$[\text{Fe}(\text{SCN})^{2+}]$	_____	_____	_____	_____	_____
Equilibrium Concentration of Fe^{3+} ion	$[\text{Fe}^{3+}]$	_____	_____	_____	_____	_____
Equilibrium Concentration of SCN^- ion	$[\text{SCN}^-]$	_____	_____	_____	_____	_____
K_c expression		_____				
K_c value		_____	_____	_____	_____	_____
Mean K_c value		_____				

Discussion Questions

You must use your data to answer these questions. If the answer you give does not reflect YOUR data and results, you will receive no credit for the question.

- 1) Give two specific sources of error in your experiment. Explain the impact of these errors on your calculation of K_c . Explain how they could be minimized in the future. (If you use the phrase "human error" in your experiment, you lose all credit!)

- 2) Calculate the pK_c using the following formula for each of your equilibrium concentrations.

$$pK_c = -\log K_c$$

- 3) Calculate the percent difference between the average pK_c and each pK_c that you calculated for the different solutions. Use the formula below.

$$\% \text{Difference} = \frac{|\text{Experimental } pK_c - \text{Average } pK_c|}{\text{Average } pK_c} \times 100$$

Full credit for this question if your values fall within 10% difference, 1/2 credit if within 20%, no credit if percent difference is greater than 20%, as this demonstrates that the error in your experimental technique was extremely high.

- 4) Will the reaction shift toward reactants, toward products or not at all, if you start with a solution that only contains 0.020M $\text{Fe}(\text{SCN})^{2+}$ ion? Briefly explain your answer.
- 5) Calculate the equilibrium concentration of iron (III) ion if your original solution only consists of a solution of 0.020M $\text{Fe}(\text{SCN})^{2+}$ ion. Show all calculations and you must use the value of K_c that you calculated from your data. (This problem may be handwritten)