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 the limiting reactant was used up 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 system. 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.

$$C_{12}H_{22}O_{11}(aq) \rightleftharpoons C_{12}H_{22}O_{11}(s)$$

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.

$$Ag^{+}(aq) + e^{-} \rightleftharpoons Ag(s)$$

These reactions are very different from one another, but both involve reactions that move back and forth producing varying amounts of products and leaving some reactants behind.

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 dynamic equilibrium** where no further reaction appears to take place. What is really happening is that the forward reaction and reverse reactions are taking place at equal rates. 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, $C_{12}H_{22}O_{11}(s)$.

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$$\overrightarrow{C_{12}H_{22}O_{11}(aq) \rightleftharpoons C_{12}H_{22}O_{11}(s)}$$

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, $C_{12}H_{22}O_{11}(aq)$.

$$\overleftarrow{C_{12}H_{22}O_{11}(aq) \rightleftharpoons C_{12}H_{22}O_{11}(s)}$$

This shifting back and forth is the hallmark of an equilibrium reaction and results in much more complicated calculations than those performed in Chemistry 114.

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 $\mathbf{K_{eq}}$ is used to represent this constant in mathematical formulas.

The Equilibrium Constant 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 write the equilibrium constant expression. The format for an equilibrium constant expression is as follows:

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

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

The first step in determining the equilibrium constant expression for a reaction is to write the chemical equation associated with the equilibrium. The chemical equation MUST be balanced correctly or your error will be magnified exponentially, since 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 numerical value of K_{eq} is reported without units since the units would be variable since K_{eq} depends completely on which units are being used and the number of reactants and products in the chemical reaction. The most commonly used units when calculating K_{eq} are molarity for concentration and atmospheres for pressure. Pressure would be used if working with gas equilibriums and the equilibrium constant would be rewritten as $\mathbf{K_p}$ 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 $\mathbf{K_c}$, the concentration equilibrium constant.

Concentration Equilibrium Constant, K_c

While many units can be used when describing equilibrium, solution concentration (in molarity!) 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 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 ($FeSCN^{2+}$) is yellow. Like any colored species in an aqueous solution, the higher the concentration of $FeSCN^{2+}$ present in the solution, the darker the color. The chemical equation associated with this reaction is given below.

$$\underset{\text{Colorless}}{Fe^{3+}(aq) + SCN^-(aq)} \rightleftharpoons \underset{\text{Yellow}}{FeSCN^{2+}(aq)}$$

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 enter them into the equilibrium constant 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 of all species once the reaction has reached equilibrium. 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.

$$Fe^{3+}(aq) \quad + \quad SCN^-(aq) \quad \rightleftharpoons \quad FeSCN^{2+}(aq)$$

The first line in your ICE table will contain the initial concentrations of any of the species present. If you start with **1M** Fe^{3+} (aq) and **1M** SCN^-(aq), these are your initial concentrations before the reaction. No $FeSCN^{2+}$ (aq) has yet been formed, so the concentration of your product initially 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:

	Fe^{3+} (aq)	+	SCN^- (aq)	$\rightleftharpoons$	$FeSCN^{2+}$ (aq)
Initial	**1M**		**1M**		0M

Any chemical listed in a problem with a given amount should be used as the initial concentration of that species unless it is specifically designated as an equilibrium concentration. Also note that while the amount of initial product in the ICE table is zero for the example above, but 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, you can assume its 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:

$$\mathbf{1}Fe^{3+}(aq) + \mathbf{1}SCN^{-}(aq) \rightleftharpoons \mathbf{1}\ FeSCN^{2+}(aq)$$

Initial	1M	1M	0M
Change	-**1**x	-**1**x	+**1**x

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.

$$\mathbf{1}Fe^{3+}(aq) + \mathbf{1}SCN^{-}(aq) \rightleftharpoons \mathbf{1}\ FeSCN^{2+}(aq)$$

Initial	1M	1M	0M
Change	-**1**x	-**1**x	+**1**x
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.

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

Using ICE Table Data

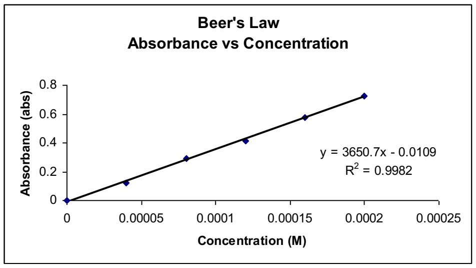
At this point, you now have an algebraic equation that relates K_c and all of the equilibrium concentrations to a single variable, x. Unfortunately, at this time, you do not have a value for x. To find x, you first need to use the ICE table to see how it correlates to the reactants and products in your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

$$1Fe^{3+}(aq) + 1SCN^{-}(aq) \rightleftharpoons 1\ \mathbf{FeSCN^{2+}(aq)}$$

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 spectrophotometer and this absorbance was generally related to the color of the solution at the time it was measured. In this experiment, you will use Beer's Law again to determine the concentration of an analyte in a colored solution.



A= εbc

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 the molar absorptivity
b is the pathlength of the cuvette. (1cm)
c is the concentration of the solution.

A plot of absorbance versus concentration for a series of diluted solutions is called a **calibration curve**. (see figure above) This method is a very effective means of relating absorbance to concentration. The correlation coefficient for the data is usually very high. In the data plotted above R^2 was 0.9982.

Determination of Equilibrium Concentrations from Graphed Data

By plotting the absorbance measured from the spectrophotometer against the concentration of the $FeSCN^{2+}$ in the solution, you can obtain a linear relationship that will allow you to calculate the concentration of $FeSCN^{2+}$ 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 $FeSCN^{2+}$.

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 use up as much reactant as possible, nearly all of the limiting reagent will be converted to product and instead of an equilibrium reaction, you can treat the chemical reaction as one that essentially goes to completion. The amount of the limiting reagent left is negligible compared to the amount of product that has been produced.

$$\longrightarrow$$

Fe^{3+}(aq) +	SCN^- (aq) ⇌	$FeSCN^{2+}$ (aq)
Colorless		Yellow
10^{-1}M	10^{-4}M	10^{-4}M (maximum amount that can be made)
Excess	Limiting	

To generate your calibration curve, you will make a series of solutions with differing concentrations of the limiting reagent and measure the absorbance of each solution. 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 ($FeSCN^{2+}$) and the absorbance will be seen.

Once you have your plot, you can use the equation for a line to determine the concentration of $FeSCN^{2+}$ 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 AU would give you a concentration of 9.9×10^{-5}M. This is the value for x which is also the value of the unknown concentration of $FeSCN^{2+}$.

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 equilibrium solution. In this example, both the Fe^{3+}(aq) and the SCN^- (aq) were initially $1x10^{-4}$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.**

	$1Fe^{3+}$(aq)	+	$1SCN^-$ (aq)	⇌	1 **$FeSCN^{2+}$ (aq)**
Initial	$1x10^{-4}$M		$1x10^{-4}$M		0M
Change	-1x		-1x		+1x
Equilibrium	($1x10^{-4}$ -$9.0x10^{-5}$)		($1x10^{-4}$ - $9.9x10^{-5}$)		**($9.9x10^{-5}$ from graph)**

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.

$$K_c = \frac{[Fe(SCN)^{2+}]}{[Fe^{3+}][SCN^-]} = \frac{[9.0x10^{-5}]}{[(1x10^{-4} - 9.0x10^{-5})][1x10^{-4} - 9.0x10^{-5}]}$$

Laboratory Equipment Procedures

Using the Vernier workstation for spectrophotometry

1. Turn on the spectrometer by pressing the power button on the top of the LabQuest 2 unit.
2. Touch the red absorbance display box.
3. Touch “change wavelength”.
4. Type in 500 for the wavelength.
5. Touch the OK button.
6. Touch the “mode” button.
7. Select “Time based” from the menu.
8. Change “sample/min” to 10. Interval should automatically display 0.1min/sample
9. Touch absorbance display.
10. Fill a cuvette with water to use as a blank. Insert the cuvette into the spectrophotometer.
11. Touch “calibrate” and follow prompts to calibrate spectrophotometer.
12. Remove the blank from the spectrometer and refill the cuvette with your sample. Be sure all water has been removed so you do not dilute your sample.
13. Insert the cuvette containing your sample into the spectrophotometer
14. Once you have recorded the absorbance, you can remove the cuvette, refill it with your next sample and insert it back into the spectrophotometer. A second calibration step is not needed

Using Excel for Data Analysis

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

Chemical Hazards

0.20M Iron(III) Nitrate Solution and 2.0×10^{-3}M 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.

2.0×10^{-3}M 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.

Experiment 3: Prelab Worksheet

(Submit through Brightspace before coming to lab)

Name: ______________________________ Date: ________ Section: ________ Grade: ________

Record all values with the correct number of significant figures and units.
Place all answers on the line next to the question.
Show calculations for any numerical answers.
See any 114 TA in the help office before your prelab is due if you have any questions.
Your answer must be completely correct to get any credit for the answer, no partial credit.

1. Why do you use an excess of iron nitrate when making the standard curve for this experiment?

   ____

2. Which solutions do you use to make your Beer's Law plot, the standard solutions, the equilibrium solutions or both?

   ____

3. What is the final concentration of Fe $^{3+}$ in a solution if you add 2.5 ml of 3.0M $Fe(NO_3)_3$ and 3.2ml of water in a test tube?

   ____

Answer the next two questions based on the following reaction.

$$H_2(g) + 1/2S_2(g) \leftrightarrows H_2S(g)$$

4. Write K_c for the given reaction above.

   ____

5. Write the expression for K_c **for the production** of 4 moles of $H_2(g)$. (Hint: How does chemical equation change?)

   ____

Answer the remaining questions based on the following reaction.

$$Cl_2(g) + PCl_3(g) \leftrightarrows PCl_5(g)$$

6. If the equilibrium amounts of Cl_2, PCl_3, and PCl_5 are 2.1, 5.4 and 3.6 moles respectively in a 10.0 L flask, use the original given chemical equation to calculate K_c for the reaction.

____

7. Set up an ICE table for the given reaction with the initial amounts of the following chemicals: $Cl_2(g)$, $PCl_5(g)$: 4.0M each and $PCl_3(g)$: 3.0M (2pts)

8. Starting with only 3.0M PCl_5, calculate the concentration of Cl_2 at equilibrium if the K_c= 5.0 at a given temperature. This will require a quadratic equation. Do not reverse the chemical equation. (2pts)

Experiment 3: Experimental Procedures and Data Sheet

Submit as part of your informal report

Name: ____ Date: ________ Section: ________

TA Signature: ______________________________________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all measurements with the correct number of significant figures and units.
TA signature and TA initials on any changes made to the data must be present or your data is invalid.

1 person does part 1 and a second person does part 2. Be sure to record the data for both sections. **Make sure to maintain appropriate social distancing when exchanging data.**

Part 1: Preparation of standard solutions

1. 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.
2. Transfer 20 drops of **0.20M** $Fe(NO_3)_3$ solution into the test tubes labeled S1-S6
3. Transfer the exact amount of 2.0 x 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

4. Transfer the exact amount of distilled water listed table below into each of the corresponding test tubes. When finished all test tubes should have the same total volume.

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 **$2.0x10^{-3}$M** $Fe(NO_3)_3$ 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 **$2.0x10^{-3}$M** $Fe(NO_3)_3$ into the test tube labeled E1-E6.
4. Transfer the exact amount of 2.0 x 10^{-3} M solution NaSCN solution listed table below into each of the corresponding test tubes.

E1	5.0 drops	E4	20.0 drops
E2	10.0 drops	E5	30.0 drops
E3	15.0 drops	E6	40.0 drops

5. Transfer the exact amount distilled water listed table below into each of the corresponding test tubes. When finished all test tubes should have the same total volume.

E1	35.0 drops	E4	20.0 drops
E2	30.0 drops	E5	10.0 drops
E3	25.0 drops	E6	0.0 drops

Part 3: Spectrophotometer Analysis

1. Calibrate the spectrophotometer 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.

Absorbance values for the standard solutions (used to create Beer's Law Plot)

S1 ________________

S2 ________________

S3 ________________

S4 ________________

S5 ________________

S6 ________________

Absorbance values for the equilibrium solutions (used to find concentration of $FeSCN^{2+}$ from Beer's Law plot) DO NOT MAKE A BEER'S LAW PLOT FROM THESE VALUES!!

E1 ________________

E2 ________________

E3 ________________

E4 ________________

E5 ________________

E6 ________________

Part 4: Cleanup

1. Remove both cuvettes from the spectrophotometer.
2. Pour all remaining solutions into waste beaker.
3. Empty waste beaker into waste container in hood.
4. Rinse all equipment, dry and return to lab drawer.

Grading

Points

Neatness and Clarity of Data	5pts	___________pts
Significant figures and units	5pts	___________pts
All data is present	10pts	___________pts

Deductions (sliding based on TA discretion)

Lab area left unclean	20pts	___________pts
Improper waste disposal	20pts	___________pts
Disruptive behavior	20pts	___________pts
Other: ____		___________pts

Grade for Experimental Procedures and Data ________pts

Experiment 3: Results Table and Calibration Curve

Submit as part of your informal report.

Name: ____ Date: ________ Section: ________

All results must be written in pen. **Pencil is not allowed!!**
Record all results with the correct number of significant figures and units
No stray marks or notes should be present on this page.
Only the tabulated results and your calibration curve are allowed on the concept review

Part 1: Creating a Beer's Law Plot

Concentrations of standards	S1	S2	S3	S4	S5	S6
Initial concentration of Fe^{3+} ion						
Initial concentration of SCN^- ion						
Final concentration of $FeSCN^{2+}$						

Part 2: Determining K_C

Concentrations of equilibrium solutions	E1	E2	E3	E4	E5	E6
Equilibrium concentration of $FeSCN^{2+}$						
Initial concentration of Fe^{3+} ion						
Initial concentration of SCN^- ion						
Equilibrium concentration of Fe^{3+} ion						
Equilibrium concentration of SCN^- ion						
Equilibrium constant, K_C						
Average value for K_C						

Equation for the line from the Beer's Law plot ______________________________

K_C expression

____

Include a copy of your calibration curve with your results table

Grading

Points

Significant figures and units	5pts	__________pts
Table is neat and legible	5pts	__________pts
Calibration curve is present	5pts	__________pts
All results are present	5pts	__________pts

Deductions (sliding based on TA discretion)

Results do not make sense	20pts	__________pts
Results do not match data	20pts	__________pts
Calibration curve incorrectly labeled	5pts	__________pts
Other: ____		__________pts

Plagiarism!!! Results are identical to another student 100pts __________pts

Grade for results table ________pts

Calculations

Submit as part of your informal report.

Your graph must be included in report.

Total Volumes of Standard and Equilibrium Solutions

Add the volumes of the $Fe(NO_3)_3$, NaSCN and distilled water used to make each solution. The total volume will come to the same value for all standards. Leave the volume in drops.

Total volume of each solution ____________________

Initial Concentrations of [Fe^{3+}] and [SCN^-] in Standard Solutions

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

$$M_2 = M_1V_1 / V2$$

M_1= concentration of stock solution

V_1= volume of stock solution in drops

M_2= concentration of final solution (standard)

V_2= volume of final solution in drops (standard)

To find the concentration of each standard solution, multiply the concentration of the stock solution by the volume of the stock solution (see Part 1) and divide by the total volume calculated for each standard solution. Find the initial concentrations for all standards for both the Fe^{3+} and SCN^-.

Initial Concentrations of [Fe^{3+}] in Standard Solutions

Initial [Fe^{3+}] S1____________ S2_____________ S3____________

Initial [Fe^{3+}] S4____________ S5_____________ S6____________

Initial Concentrations of [SCN^-] in Standard Solutions

Initial [SCN^-] S1____________ S2_____________ S3____________

Initial [SCN^-] S4____________ S5_____________ S6____________

Concentration of $FeSCN^{2+}$ in Standard Solutions

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

Fe^{3+}(aq) + SCN^- (aq) → $FeSCN^{2+}$ (aq)
(excess) (limiting) (same amount as limiting reagent)

Concentration of $FeSCN^{2+}$ in Standard Solutions

Final [$FeSCN^{2+}$] S1____________ S2_____________ S3____________

Final [$FeSCN^{2+}$] S4____________ S5_____________ S6____________

Creating a Beer's Law Plot

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

Standard	[$FeSCN^{2+}$]	Absorbance
S1	__________	__________
S2	__________	__________
S3	__________	__________
S4	__________	__________
S5	__________	__________
S6	__________	__________

Plot the [$FeSCN^{2+}$] data on the x-axis and the absorbance data on the y-axis on graph paper. Estimate the equation for the line based on your data (y=mx+b). Make sure you can reproduce the equation based on your calculations, since you must be able to do this on your concept review.

Concentration of $FeSCN^{2+}$ in Equilibrium Solutions

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

Calculated linear equation ____

Concentration of $FeSCN^{2+}$ in Equilibrium Solutions

Equilibrium $[FeSCN^{2+}]$ E1____________ E2_____________ E3____________

Equilibrium $[FeSCN^{2+}]$ E4____________ E5_____________ E6____________

Initial Concentrations of $[Fe^{3+}]$ and $[SCN^{-}]$ in Equilibrium Solutions

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

$$M_2= M_1V_1/ V2$$

M_1= concentration of stock solution
V_1= volume of stock solution
M_2= concentration of final solution (equilibrium)
V_2= volume of final solution (equilibrium)

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 equilibrium solution. Find the initial concentrations for both the Fe^{3+} and SCN^{-} in the equilibrium solutions.

Initial Concentrations of $[Fe^{3+}]$ in Equilibrium Solutions

Initial $[Fe^{3+}]$ E1____________ E2_____________ E3____________

Initial $[Fe^{3+}]$ E4____________ E5_____________ E6____________

Initial Concentrations of [SCN^-] in Equilibrium Solutions

Initial [SCN^-] E1____________ E2_____________ E3____________

Initial [SCN^-] E4____________ E5_____________ E6____________

Concentrations of [Fe^{3+}] and [SCN^-] in Equilibrium Solutions

For every mole of $FeSCN^{2+}$ (aq) produced, you used up 1 mole of Fe^{3+}(aq). To account for this loss, subtract the concentration of the $FeSCN^{2+}$ at equilibrium for E1 from the initial concentration of Fe^{3+}(aq) in E1 to find the equilibrium concentration of Fe^{3+}(aq) for solution E1. Do the same for the SCN^-(aq) and then repeat the calculations for E2-E6.

Concentrations of [Fe^{3+}] in Equilibrium Solutions

Equilibrium [Fe^{3+}] E1____________ E2_____________ E3____________

Equilibrium [Fe^{3+}] E4____________ E5_____________ E6____________

Concentrations of [SCN^-] in Equilibrium Solutions

Equilibrium [SCN^-] E1____________ E2_____________ E3____________

Equilibrium [SCN^-] E4____________ E5_____________ E6____________

K_c Expression

Write the K_c expression for the following reaction: $Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$. Remember, products over reactants with each concentration raised to its coefficient.

K_c=

____

Value of K_c and Average K_c

Fill in the equilibrium concentrations you calculated for $Fe^{3+}(aq)$, $SCN^{-}(aq)$ and $FeSCN^{2+}(aq)$ 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.

Equilibrium constants for equilibrium solutions

K_c E1____________ E2____________ E3____________

K_c E4____________ E5____________ E6____________

Average K_c ___________

There are no additional questions for this experiment. Make sure all graphs are correctly formatted and all work is clearly shown.