CH 223 Spring 2024: **"Determination of an Equilibrium Constant** *(in class)*" ***Lab** - Instructions*

Note: **This is the lab for section 01 and H1 of CH 223 only.**

- *If you are taking section W1 of CH 223*, please use this link:
  http://mhchem.org/q/2b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-2-2 through Ia-2-15 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, April 1 (section 01) *or* **Wednesday, April 3 (section H1.)** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in (**pages Ia-2-7 through Ia-2-15 *only* to avoid a point penalty**) and your computer generated graph at the beginning of recitation to the instructor on Monday, April 8 (section 01)** *or* **Wednesday, April 10 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Determination Of An Equilibrium Constant

In this experiment the equilibrium properties of the reaction between the iron(III) ion and the thiocyanate ion will be studied. The relevant chemical equation for this lab is:

$$Fe^{+3}{}_{(aq)} + SCN^{-}{}_{(aq)} \rightleftharpoons FeSCN^{+2}{}_{(aq)} \qquad\qquad K_c = \frac{[FeSCN^{+2}]}{[Fe^{+3}][SCN^{-1}]}$$

When solutions of Fe^{+3} and SCN^{-} are mixed, the above reaction occurs to some extent, forming the *complex ion* $FeSCN^{+2}$. The concentrations of Fe^{+3} and SCN^{-} will decrease by one mole for every mole of $FeSCN^{+2}$ that forms, but not all of the Fe^{3+} and SCN^{-} ions will be converted to the complex ion under normal circumstances.

The objective of this experiment is to determine the **equilibrium constant, K_c**, for this reaction. The value of K_c is constant at a given temperature. Any mixture of Fe^{+3} and SCN^{-} will react until the same value of K_c is obtained. In this experiment, we will determine K_c for this reaction using several different mixtures of Fe^{+3} and SCN^{-}.

Before we can calculate the value of the equilibrium constant, we must be able to determine the concentration of $FeSCN^{+2}$ in solution. The Beer-Lambert Law, which is commonly referred to as simply **Beer's Law**, relates the absorption of light in a colored sample to its concentration in solution:

$$\mathbf{A = \varepsilon bc = \log (100\% / \%T)}$$

Here **A** is the measured absorbance of the colored solution, ε is the molar absorptivity (with units of M^{-1} cm^{-1}), **b** is the path length (in cm), and **c** is the concentration (molarity, or moles per Liter) of the species being studied, in this case $FeSCN^{+2}$. The value of the molar absorptivity, ε, depends on the solute's overall energy stored within the molecule. $FeSCN^{2+}$ has a deep red color while the reactants are virtually colorless, and only $FeSCN^{2+}$ will absorb light in the visible region.

Absorbance is a unitless quantity that corresponds with the amount of light removed by a colored system. Depending on if you use an analog Spectronic 21 or a digital Vernier SpectroVis, you may need to measure the **percent transmittance (%T)**, or how much light gets through colored solutions. Calculating absorbance from %T is straightforward (A = log (100%/%T).

Because $FeSCN^{+2}$ has a red color, we will use a wavelength appropriate for measuring red light – the SpectroVis will tell us what the exact wavelength will be, but it should be about 450 nm (which is what you'll use when utilizing the Spectronic 21.) Using the Beer's Law equation, we can determine the molar absorptivity constant, ε, for $FeSCN^{+2}$ by measuring its absorbance at different known concentrations of $FeSCN^{2+}$. If absorbance is plotted versus concentration, the *slope* will give the molar absorptivity constant using Beer's Law.

Finding the concentration of $FeSCN^{+2}$ can be achieved using Le Chatelier's Principle. If an excess of Fe^{3+} is added to a small amount of SCN^{-}, one can assume that the amount of SCN^{-} present in solution equals the amount of $FeSCN^{2+}$ formed at equilibrium. In other words, we know the final concentration of $FeSCN^{+2}$ in the solution by creating a solution that is not in equilibrium but goes to *completion* using the principle of *limiting reagent*. The SCN^{-} will be completely converted to $FeSCN^{+2}$, such that the final concentration of $FeSCN^{+2}$ is equal to the initial concentration of SCN^{-}.

The purpose of this lab is to find the value of the equilibrium constant, K_c. First, you will prepare a series of solutions with known concentrations of $FeSCN^{2+}$ (or SCN^-) and measure the absorbance (or percent transmittance) values at a wavelength appropriate for a red solution around 450 nm. When the absorbance values are plotted versus the concentration of $FeSCN^{2+}$, a linear relationship appears, and ε can be calculated using linear regression (the slope equals εb.). Remember to report the value of the slope, y-intercept and correlation coefficient when using linear regression. ε will allow you to calculate the concentration of $FeSCN^{2+}$ for any combination of Fe^{3+} and SCN^-. You will then prepare a series of solutions with varying amounts of Fe^{3+} and SCN^- initially present, determine the absorbance for $FeSCN^{2+}$, and finally determine the value of K_c at room temperature.

PROCEDURE:

In **Part A**, the goal is to find the molar absorptivity constant, and in **Part B**, you will find the value of the equilibrium constant. Both parts use similar techniques, but **make sure you use the correct reactant concentrations in each section.** Check and double-check the concentrations before analyzing your solutions!

Before you leave lab, it is important that you share data with the other group members (essentially mL of KSCN and absorbance (or %Transmittance) values). Also make sure to get all of the names of your lab partners for your final lab report.

Part A: Determining the Molar Absorptivity Constant (ε)

Place approximately 30 mL of **0.100 M $Fe(NO_3)_3$** in 1 M HNO_3 in a dry labeled 100 mL beaker. The HNO_3 allows the solute components to remain in solution, but it does not affect our calculations in this lab. Note also that there are *two different stock solutions* of $Fe(NO_3)_3$ and KSCN at different concentrations. Make sure you choose the correct solution for part A or your experiment will not work! ***NEVER*** **place solution back into the reagent bottle** - once the solution is removed, the solution has to be placed in the waste bottle.

Place approximately 20 mL of **3.00×10^{-4} M KSCN** into a second dry labeled 100 mL beaker. Clean and dry six 18 x 150 mm test tubes and label 1-5 and B (blank).

Pipet 5.00 mL of the $Fe(NO_3)_3$ solution into tubes 1-5 as demonstrated by the instructor. ***NEVER*** **pipet directly from the bottle, only pipet from a beaker**. Pipet 1, 2, 3, 4, or 5 mL of KSCN into the corresponding labeled test tube. Then add the correct number of mL of water to each test tube so that the *total volume* is 10.00 mL. The sixth test tube is the **blank** (or "dark tube') and will contain only water.

You will need to transfer a small quantity of each solution into a **cuvette** for analysis on your instrument (see below); ideally you will have six cuvettes, one for each of the test tubes from above. Make sure the cuvette is filled to at least 80% of the maximum volume to ensure that light flows through the sample.

You will be using either a Spectronic 21 or a SpectroVis Plus. Follow the appropriate instructions:

Spectronic 21:

Determine the %T for each of the five solutions using a Spec 21 at 450 nm. Assume the path length, b, equals 1.00 cm for these trials. Convert the %T readings into absorbance. Recall that $A = \log(100\%/\%T)$.

<u>SpectroVis Plus</u>:

- Connect the SpectroVis Plus to the LabQuest II. If everything is correct, you should see “USB: Abs” on the screen.
- Place the blank cuvette (with water only) in the SpectroVis, noting the arrow (and direction of light - do not place the "ribbed" walls of the cuvette face the arrow or source of light symbol.)
- Select “Sensors -> Calibrate -> USB: Spectrometer”, then press OK if necessary. The following message is displayed: “Waiting 90 seconds for lamp to warm up…” After 90 seconds, the message will change to “Warmup complete” then select “Finish Calibration”. When the message “Calibration completed” appears, select OK.
- Place one of the $FeSCN^{2+}$ samples in the SpectroVis. Start data collection by pushing the green “start” button in the lower left corner of the LabQuest 2. A full spectrum graph of the solution will be displayed. Stop data collection (by pressing the red button in the lower left corner) and move the cursor to the maximum wavelength at about 450 nm; this is the wavelength of maximum absorbance (λ max). Record both the wavelength of maximum absorbance and the actual absorbance value in your lab notebook.
- Remove the cuvette and place a new cuvette (with another solution) inside. Repeat the above process until all solutions have recorded values of absorbance and the corresponding λ max (and they should all be about 450 nm.)

You should see a linear relationship between mL of KSCN added and Absorbance; if not, you might want to re-do the measurements. Dispose of the solutions in the waste bottle and clean your beakers and test tubes.

Part B: Determining the Equilibrium Constant (K)

Place approximately 30 mL of **2.00×10^{-3} M $Fe(NO_3)_3$** in 1 M HNO_3 in a dry labeled 100 mL beaker. Place approximately 20 mL of **2.00×10^{-3} M KSCN** in a second dry labeled 100 mL beaker. Clean and dry five 18 x 150 mm test tubes and label 1-5.

Pipet 5.00 mL of the $Fe(NO_3)_3$ solution into tubes 1-5. Pipet 1, 2, 3, 4, or 5 mL of KSCN into the corresponding test tube. Then add the correct number of mL of water to each test tube so that the total volume is 10.00 mL.

Determine the absorbance or %T for each of the five solutions using the Spec 21 (at 450 nm) or the SpectroVis Plus as before. Assume b = 1.00 cm. Convert the %T readings into absorbance values if using the Spec 21 (the SpectroVis does this automatically for you.) You should see a linear relationship between the concentration of KSCN added and Absorbance; if not, you might want to re-do the measurements. Dispose of the solutions in the waste bottle when complete.

Calculations for Part A: *Finding the value of* ε

Determine the diluted molarity of Fe^{+3} and of SCN^- present in each solution using the dilution equation ($M_1 \times V_1 = M_2 \times V_2$).

> *Example:* Find the concentration of SCN^- when 4.00 mL of a $3.00 * 10^{-4}$ M KSCN is diluted to a total of 10.00 mL.
>
> *Solution:* The initial concentration of KSCN, $3.00 * 10^{-4}$ M, is being diluted to a new solution volume of 10.00 mL. We can use $M_1 \times V_1 = M_2 \times V_2$ equation and solve for M_2.
> $M_2 = 3.00 * 10^{-4}$ M * 4.00 mL / 10.00 mL = **$1.20 * 10^{-4}$ M**

Find the concentration of $FeSCN^{2+}$ using the law of limiting reactants.

Example: Find the concentration of $FeSCN^{2+}$ when $[SCN^-] = 1.20 * 10^{-4}$ M and $[Fe^{3+}] = 0.100$ M.

Solution: Since $[SCN^-] << [Fe^{3+}]$, it can be assumed that all of the SCN^{-1} has been converted to $FeSCN^{2+}$ using Le Chatelier's Principle. Hence, the $[FeSCN^{2+}]$ at equilibrium equals **$1.20 * 10^{-4}$ M** in this example. *Note:* in part B you will use the value of ε and the absorbance to calculate $[FeSCN^{2+}]$.

Construct a **graph** of absorbance versus the concentration of $FeSCN^{2+}$ on a computer or tablet (no hand drawn graphs.) Perform a **linear regression** on the data, and record your values of the slope, the y-intercept and the correlation coefficient (r). **Determine the value of** ε from the slope (remember that the path length, b, equals 1.00 cm.) **Print this graph** and include it with your final lab report.

Calculations for Part B: *Finding the value of K*

To analyze this set of data, determine the initial molarity of Fe^{+3} and of SCN^- present in each of your five solutions. See the section in Part A for assistance on this procedure.

Calculate the equilibrium concentration of $[FeSCN^{+2}]_{eq}$ in each sample. This can be done using Beer's Law and the molar absorptivity constant determined in part A. $\mathbf{c_{FeSCN} = A / b\ \varepsilon}$

Example: Calculate the equilibrium concentration of $FeSCN^{2+}$ when $\varepsilon = 3420\ cm^{-1}\ M^{-1}$, %T = 45.2%, and b = 1.00 cm.

Solution: First convert %T to absorbance. A = log(100%/45.2%) = 0.345
Now solve for the concentration, c. $c = A / b\ \varepsilon = 0.345 / 1.00\ cm * 3420\ cm^{-1}\ M^{-1} =$ **$1.00 * 10^{-4}$ M**

Next, find the equilibrium concentrations of $[Fe^{+3}]$ and $[SCN^-]$ using the following equations:

$$\mathbf{[Fe^{+3}]_{eq}} = [Fe^{+3}]_{ini} - [FeSCN^{+2}]_{eq} \quad \textit{and} \quad \mathbf{[SCN^-]_{eq}} = [SCN^-]_{ini} - [FeSCN^{+2}]_{eq}$$

where **eq** = equilibrium concentration and **ini** = initial concentration

An **ICE table** may also be helpful (note that $x = [FeSCN^{+2}]_{eq}$):

	$[Fe^{3+}]$	$[SCN^-]$	$[FeSCN^{2+}]$
initial	$[Fe^{3+}]_0$	$[SCN^-]_0$	0
change	- x	- x	+ x
equilibrium	$[Fe^{3+}]_0$ - x	$[SCN^-]_0$ - x	x

Finally, calculate the equilibrium constant, $\mathbf{K_c}$, for each of the five solutions:

$$\mathbf{K_c = [FeSCN^{+2}]_{eq} / \{[Fe^{+3}]_{eq} * [SCN^-]_{eq}\}} \quad \textit{or:}$$

$$K_c = \frac{(x)}{\left([Fe^{+3}]_0 - x\right)\left([SCN^-]_0 - x\right)}$$

You should have five values of K_c that are all relatively close to each other.

Determine the average K_c value and deviation in parts per thousand (ppt) for your calculations. *See the ppt handout on the CH 223 website or here:* https://mhchem.org/ppt

Determination of an Equilibrium Constant

YOUR NAME: ______________________ ***LAB PARTNER(s):*** ________________

Purpose: To determine the value of an equilibrium constant (K_c) for the following reaction:

$$Fe^{+3}{}_{(aq)} + SCN^{-}{}_{(aq)} \rightleftharpoons FeSCN^{+2}{}_{(aq)} \qquad K_c = \frac{[FeSCN^{+2}]}{[Fe^{+3}][SCN^{-1}]}$$

Goal #1: **Obtain the Data**

a. **Create the five solutions** as discussed in Part A on page I-2-3**.** Use the Spec-21 *or* SpectroVis to **determine the absorbance (and %Transmittance, if necessary) for each of the five solutions.**

If converting from %T, **Absorbance = log(100%/%Transmittance)** *use base 10 log, not natural log (ln)!*

Data check: The solutions should get progressively more orange/red as the mL of KSCN increase (1.00 mL will be the most dilute, 5.00 mL the most intense.) Also, as mL KSCN increase, the %Transmittance will decrease and the Absorbance increase. If these are not observed, re-create your solutions and try again.

Part A:

mL KSCN	%Transmittance *(if necessary)*	mL KSCN	Absorbance
1.00	_______________	1.00	_______________
2.00	_______________	2.00	_______________
3.00	_______________	3.00	_______________
4.00	_______________	4.00	_______________
5.00	_______________	5.00	_______________

Notes: (optional)

Goal #1 (continued)

b. **Create the five solutions** as discussed in Part B on page I-2-4**.** Use the Spec-21 *or* SpectroVis to **determine the absorbance (and %Transmittance, if necessary) for each of the five solutions.**

If converting from %T, **Absorbance = log(100%/%Transmittance)** *use base 10 log, not natural log (ln)!*

Data check: As before, the solutions should get progressively more orange/red as the mL of KSCN increase (1.00 mL will be the most dilute, 5.00 mL the most intense.) Also, as mL KSCN increase, the %Transmittance will decrease and the Absorbance increase. If these are not observed, re-create your solutions and try again.

Part B:

mL KSCN	%Transmittance
1.00	_______________
2.00	_______________
3.00	_______________
4.00	_______________
5.00	_______________

mL KSCN	Absorbance
1.00	_______________
2.00	_______________
3.00	_______________
4.00	_______________
5.00	_______________

Notes: (optional)

Goal #2: **Determine the value of** ε *We will convert the mL of KSCN into a concentration of KSCN. In Part A, the [KSCN] << [Fe^{3+}], so we can assume that [KSCN] = [$FeSCN^{2+}$]. Next we will convert the %Transmittance values into Absorbance, then make a plot of Absorbance versus [$FeSCN^{2+}$]. Finally, a linear regression of the Absorbance versus [$FeSCN^{2+}$] data will lead us to the value of* ε*, the molar absorptivity coefficient, which will be used in the next section.*

a. **Convert mL of KSCN into the diluted concentration of [KSCN] (molarity) using $M_1V_1 = M_2V_2$.** Use **M_1 = 3.00 x 10^{-4} M** for the initial molarity of KSCN and use a final volume V_2 = 10.00 mL. *(see page I-2-4 for an example.)*

b. Since [KSCN] << [Fe^{3+}] in part A, so we can assume that **[KSCN] = [$FeSCN^{2+}$]**. Write the values for **[$FeSCN^{2+}$]** in the next column… these will be the x-axis values for your graph. *(see page I-2-4 and 1-2-5 for an example.)*

mL KSCN (V_1)	diluted [KSCN] (M_2)	**[$FeSCN^{2+}$] (M)**
1.00	_______________	_______________
2.00	_______________	_______________
3.00	_______________	_______________
4.00	_______________	_______________
5.00	_______________	_______________

Show an example of your work by converting 3.00 mL KSCN into a diluted [KSCN] and then into [$FeSCN^{2+}$] in the space below.

Goal #2 (continued)

c. Rewrite your Absorbance values ***for Part A*** into the following table (Goal #1, part a) as well as your values for [$FeSCN^{2+}$] (Goal #2, part b).

mL KSCN	Absorbance (Part A)	[$FeSCN^{2+}$] (M)
1.00	____________________	_______________
2.00	____________________	_______________
3.00	____________________	_______________
4.00	____________________	_______________
5.00	____________________	_______________

d. **Create a computer-generated graph** in a program like Excel, Sheets, Numbers, etc. with **Absorbance versus [$FeSCN^{2+}$]** (Goal #2, part c). **Include the graph with this lab report**. Graphs should be labeled correctly with appropriate axes, and the data should follow a straight line with a positive slope.

Absorbance values are unitless and will constitute the y-axis in the graph. The values for [$FeSCN^{2+}$] will be the x-axis in the graph.

e. **Perform a linear regression of the values used in Goal #2, part d** above (Absorbance versus [$FeSCN^{2+}$]). You can do this in the graphing program you used in Goal #2 part d above, or on your calculator, or using an online linear regression calculator (for example: http://www.alcula.com/calculators/statistics/linear-regression/) The **slope** of the line equals ε**,** the **molar absorptivity coefficient,** which we will use in Goal #3 next.

The slope should have a value > 1000.

Report the values for slope (ε), y-intercept and the correlation coefficient (r) obtained from the linear regression below:

slope = ε = ____________________________

y-intercept = ___________________________

correlation coefficient (r) = _________________________ *(r is not the same as R^2!)*

Make sure you include a computer-generated graph with your lab report!

Goal #3: **Determine the value of K_c.** *We will convert the mL of KSCN and Fe^{3+} into diluted concentrations of KSCN and Fe^{3+} in a method similar to that of Goal #2. We will convert the Part B %Transmittance values into absorbance, then we can use our* **ε,** the **molar absorptivity coefficient**, *to calculate the equilibrium concentrations of $FeSCN^{2+}$ using the Beer-Lambert Law. Finally, using ICE tables, we will calculate the equilibrium concentrations of Fe^{3+}, SCN^{-1} and $FeSCN^{2+}$ in order to calculate K_c, the equilibrium constant for this lab.*

a. **Convert mL of KSCN into the diluted concentration of [KSCN] (molarity) using $M_1V_1 = M_2V_2$.** Use **$M_1 = 2.00 \times 10^{-3}$ M** for the initial molarity of KSCN and use a final volume $V_2 = 10.00$ mL. *(see page I-2-4 and Goal #2 section a for examples.)*

b. **Convert mL of Fe^{3+} into the diluted concentration of $[Fe^{3+}]$ (molarity) using $M_1V_1 = M_2V_2$.** Use **$M_1 = 2.00 \times 10^{-3}$ M** for the initial molarity of Fe^{3+} and use a final volume $V_2 = 10.00$ mL. *These calculated M_2 values should all be the same since the initial volume (V_1) are identical! (see page I-2-4 and Goal #2 section a for examples.)*

mL KSCN (V_1)	**diluted [KSCN] (M_2)**	**mL Fe^{3+} (V_1)**	**diluted $[Fe^{3+}]$ (M_2)**
1.00	______________	5.00	______________
2.00	______________	5.00	______________
3.00	______________	5.00	______________
4.00	______________	5.00	______________
5.00	______________	5.00	______________

Note:

- We will refer to "diluted [KSCN]" values as **$[SCN^{-1}]_i$ in Goal #3, part e, below**
- We will refer to "diluted [**Fe^{3+}**]" values as **$[Fe^{3+}]_i$ in Goal #3, part f, below**

Show an example of your work by showing how to convert 3.00 mL KSCN into a diluted [KSCN]:

Show an example of your work by showing how to convert 5.00 mL Fe^{3+} into a diluted $[Fe^{3+}]$:

Goal #3 (continued)

c. Rewrite your Absorbance (and %Transmittance values, if used) **for Part B** into the following table (Goal #1, part b).

Absorbance values are unitless and will constitute the y-axis in the graph.

mL KSCN	%Transmittance (*if used*)	Absorbance (Part B)
1.00	____________	____________
2.00	____________	____________
3.00	____________	____________
4.00	____________	____________
5.00	____________	____________

d. Using Beer-Lambert Law, we can convert your Absorbance (A) values (Goal #3, part c) into **concentration of [$FeSCN^{2+}$]** using:

$$\textbf{Absorbance} = \mathbf{A} = \varepsilon \mathbf{bc}$$

In this equation, **A** = Absorbance (Goal #3, part c), ε = the molar absorptivity coefficient (from Goal #2, part e), **b** = path length (for this lab, b = 1.00 cm) and **c =** concentration of [$FeSCN^{2+}$] (in mol/L, or just M). You want to solve for the **equilibrium concentration of [$FeSCN^{2+}$]**, so **c = A/εb**

value of ε = ____________________ *(from Goal #2, part e)*

mL KSCN	**Absorbance**	**[$FeSCN^{2+}$]$_{eq}$ (M)**
1.00	____________	____________
2.00	____________	____________
3.00	____________	____________
4.00	____________	____________
5.00	____________	____________

Goal #3 (continued)

e. We can use the values of **$[FeSCN^{2+}]_{eq}$** from Goal #3, part d (above) to find the equilibrium values of $[SCN^{-1}]$ or $[SCN^{-1}]_{eq}$. To find the equilibrium values of $[SCN^{-1}]_{eq}$ do the following:

$$[SCN^{-1}]_{eq} = [SCN^{-1}]_i - [FeSCN^{2+}]_{eq}$$

- **$[SCN^{-1}]_i$** values can be found in Goal #3, part b above
- **$[FeSCN^{2+}]_{eq}$** values can be found in Goal #3, part d above

mL KSCN	$[SCN^{-1}]_i$	$[FeSCN^{2+}]_{eq}$	$[SCN^{-1}]_{eq}$
1.00	________	________	________
2.00	________	________	________
3.00	________	________	________
4.00	________	________	________
5.00	________	________	________

f. In a similar fashion, we can use the values of $[FeSCN^{2+}]_{eq}$ from Goal #3, part d to find the equilibrium values of $[Fe^{3+}]$ or $[Fe^{3+}]_{eq}$. To find the equilibrium values of $[Fe^{3+}]_{eq}$ do the following:

$$[Fe^{3+}]_{eq} = [Fe^{3+}]_i - [FeSCN^{2+}]_{eq}$$

- **$[Fe^{3+}]_i$** values can be found in Goal #3, part b above - *they should be all identical!*
- **$[FeSCN^{2+}]_{eq}$** values can be found in Goal #3, part d above

mL Fe^{3+}	$[Fe^{3+}]_i$	$[FeSCN^{2+}]_{eq}$	$[Fe^{3+}]_{eq}$ (M)
5.00	________	________	________
5.00	________	________	________
5.00	________	________	________
5.00	________	________	________
5.00	________	________	________

Goal #3 (continued)

g. We can use the equilibrium values of $[Fe^{3+}]_{eq}$, $[SCN^{-1}]_{eq}$ and $[FeSCN^{2+}]_{eq}$ to calculate the value of $\mathbf{K_c}$**.**

$$\mathbf{K}_c = \frac{\mathbf{[FeSCN^{+2}]}}{\mathbf{[Fe^{+3}][SCN^{-1}]}}$$

- $\mathbf{[FeSCN^{2+}]_{eq}}$ values can be found in Goal #3, part d above
- $\mathbf{[SCN^{-1}]_{eq}}$ values can be found in Goal #3, part e above
- $\mathbf{[Fe^{3+}]_{eq}}$ values can be found in Goal #3, part f above

$\mathbf{[FeSCN^{2+}]_{eq}}$	$\mathbf{[SCN^{-1}]_{eq}}$	$\mathbf{[Fe^{3+}]_{eq}}$	$\mathbf{K_c}$

Average of five $\mathbf{K_c}$ values = ___________________

Parts per thousand = _______________________

Info on parts per thousand: **https://mhchem.org/ppt**

In the space below, show *how* you calculated the first K_c value, the average of the five K_c values, and the parts per thousand.

Postlab Questions:

A student mixes 5.00 mL of 2.00×10^{-3} M $Fe(NO_3)_3$ with 5.00 mL 2.00×10^{-3} M KSCN. She finds that in the equilibrium mixture the concentration of $FeSCN^{+2}$ is 1.40×10^{-4} M. $\mathbf{Fe^{+3}_{(aq)} + SCN^{-}_{(aq)} \rightleftharpoons FeSCN^{+2}_{(aq)}}$

a. What is the initial concentration in solution of the Fe^{+3} and SCN^{-}?

b. Create an ICE table for this problem. What is the equilibrium constant value for this reaction?

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