

# *CH 223 Spring 2024:* **“Determination of an Equilibrium Constant (*online*)”** *Lab - Instructions*

**Note: This is the lab for section W1 of CH 223 only.**

- *If you are taking section 01 or section H1 of CH 223, please use this link:*  
<http://mhchem.org/q/2a.htm>
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*Step One:*

**Watch the lab video** for the “Equilibrium Constant” lab, found here:

<http://mhchem.org/v/a.htm>

**Record** the data found at the *end* of the lab video on page Ib-2-7.

*Step Two:*

**Complete pages Ib-2-7 through Ib-2-14** using the “Equilibrium Constant” video and the actual lab instructions on pages Ib-2-2 through Ib-2-6. Include your name on page Ib-2-7!

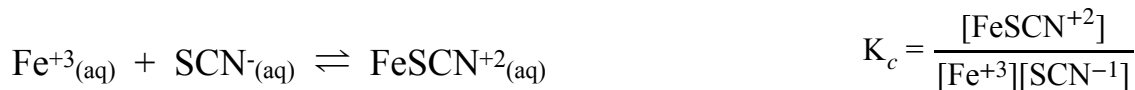
*Step Three:*

**Submit your lab** (pages Ib-2-7 through Ib-2-14 *only* to avoid a point penalty) **and computer generated graph as a single PDF file to the instructor via email** ([mike.russell@mhcc.edu](mailto:mike.russell@mhcc.edu)) **on Wednesday, April 10 by 11:59 PM.** I recommend a free program (ex: CamScanner, <https://camscanner.com>) or a website (ex: CombinePDF, <https://combinepdf.com>) to convert your work to a PDF file.

*If you have any questions regarding this assignment, please email ([mike.russell@mhcc.edu](mailto: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:



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

Before we can calculate the value of the equilibrium constant, we must be able to determine the concentration of  $\text{FeSCN}^{+2}$  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:

$$A = \epsilon bc = \log (100\% / \%T)$$

Here **A** is the measured absorbance of the colored solution,  $\epsilon$  is the molar absorptivity (with units of  $\text{M}^{-1} \text{cm}^{-1}$ ), **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  $\text{FeSCN}^{+2}$ . The value of the molar absorptivity,  $\epsilon$ , depends on the solute's overall energy stored within the molecule.  $\text{FeSCN}^{+2}$  has a deep red color while the reactants are virtually colorless, and only  $\text{FeSCN}^{+2}$  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  $\text{FeSCN}^{+2}$  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,  $\epsilon$ , for  $\text{FeSCN}^{+2}$  by measuring its absorbance at different known concentrations of  $\text{FeSCN}^{+2}$ . If absorbance is plotted versus concentration, the *slope* will give the molar absorptivity constant using Beer's Law.

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

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  $\text{FeSCN}^{2+}$  (or  $\text{SCN}^-$ ) 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  $\text{FeSCN}^{2+}$ , a linear relationship appears, and  $\epsilon$  can be calculated using linear regression (the slope equals  $\epsilon b$ ). Remember to report the value of the slope, y-intercept and correlation coefficient when using linear regression.  $\epsilon$  will allow you to calculate the concentration of  $\text{FeSCN}^{2+}$  for any combination of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$ . You will then prepare a series of solutions with varying amounts of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$  initially present, determine the absorbance for  $\text{FeSCN}^{2+}$ , and finally determine the value of  $K_c$  at room temperature.

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## 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.

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### Part A: Determining the Molar Absorptivity Constant ( $\epsilon$ )

Place approximately 30 mL of **0.100 M  $\text{Fe}(\text{NO}_3)_3$**  in 1 M  $\text{HNO}_3$  in a dry labeled 100 mL beaker. The  $\text{HNO}_3$  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  $\text{Fe}(\text{NO}_3)_3$  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  **$3.00 \times 10^{-4}$  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  $\text{Fe}(\text{NO}_3)_3$  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:

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### 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  $A = \log(100\%/ \%T)$ .

### SpectroVis Plus:

- 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  $\text{FeSCN}^{2+}$  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 ( $\lambda_{\text{max}}$ ). 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  $\lambda_{\text{max}}$  (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.

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### **Part B: Determining the Equilibrium Constant (K)**

Place approximately 30 mL of  $2.00 \times 10^{-3} \text{ M Fe(NO}_3)_3$  in 1 M  $\text{HNO}_3$  in a dry labeled 100 mL beaker. Place approximately 20 mL of  $2.00 \times 10^{-3} \text{ M KSCN}$  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  $\text{Fe(NO}_3)_3$  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  $b = 1.00 \text{ cm}$ . 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.

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### **Calculations for Part A: *Finding the value of $\epsilon$***

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

*Example:* Find the concentration of  $\text{SCN}^-$  when 4.00 mL of a  $3.00 \times 10^{-4} \text{ M KSCN}$  is diluted to a total of 10.00 mL.

*Solution:* The initial concentration of KSCN,  $3.00 \times 10^{-4} \text{ M}$ , 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  $M_2$ .

$$M_2 = 3.00 \times 10^{-4} \text{ M} \times 4.00 \text{ mL} / 10.00 \text{ mL} = \mathbf{1.20 \times 10^{-4} \text{ M}}$$

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

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

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

Construct a **graph** of absorbance versus the concentration of  $\text{FeSCN}^{2+}$  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  $\epsilon$**  from the slope (remember that the path length, b, equals 1.00 cm.) **Print this graph** and include it with your final lab report.

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

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

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

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

*Solution:* First convert  $\%T$  to absorbance.  $A = \log(100\%/45.2\%) = 0.345$

Now solve for the concentration, c.  $c = A / b \epsilon = 0.345 / 1.00 \text{ cm} \times 3420 \text{ cm}^{-1} \text{ M}^{-1} = 1.00 \times 10^{-4} \text{ M}$

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

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

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

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

	$[\text{Fe}^{3+}]$	$[\text{SCN}^-]$	$[\text{FeSCN}^{2+}]$
<i>initial</i>	$[\text{Fe}^{3+}]_0$	$[\text{SCN}^-]_0$	0
<i>change</i>	- x	- x	+ x
<i>equilibrium</i>	$[\text{Fe}^{3+}]_0 - x$	$[\text{SCN}^-]_0 - x$	x

Finally, calculate the equilibrium constant,  $K_c$ , for each of the five solutions:

$$K_c = \frac{[\text{FeSCN}^{2+}]_{\text{eq}}}{\{[\text{Fe}^{3+}]_{\text{eq}} * [\text{SCN}^-]_{\text{eq}}\}} \quad \text{or:}$$
$$K_c = \frac{(x)}{([Fe^{+3}]_0 - x)([SCN^-]_0 - x)}$$

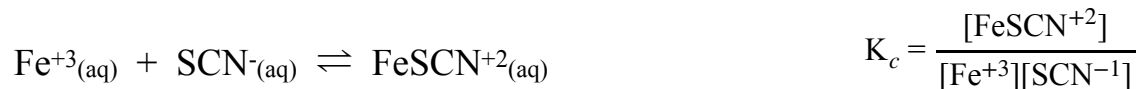
You should have five values of  $K_c$  that are all relatively close to each other.

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 - *Worksheet*

**Your name:**

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



**Goal #1:** Watch the following video related to this lab: <http://mhchem.org/v/a.htm>

Write down the following data *from the video* which relates to this lab:

*Part A:*

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

*Part B:*

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

*Notes: (optional)*

**Goal #2: Determine the value of  $\epsilon$**  We will convert the mL of KSCN into a concentration of KSCN. In Part A, the  $[KSCN] \ll [Fe^{3+}]$ , 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  $\epsilon$ , 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  $M_1 = 3.00 \times 10^{-4} \text{ M}$  for the initial molarity of KSCN and use a final volume  $V_2 = 10.00 \text{ mL}$ . (see page I-2-4 for an example.)

b. Since  $[KSCN] \ll [Fe^{3+}]$  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 I-2-5 for an example.)

<u>mL KSCN (<math>V_1</math>)</u>	<u>diluted [KSCN] (<math>M_2</math>)</u>	<u>[FeSCN<sup>2+</sup>]</u>
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. Convert your %Transmittance values *from part A* into **absorbance** using:

$$\text{Absorbance} = \log(100\%/\% \text{Transmittance})$$

Absorbance values are unitless and will constitute the y-axis in the graph. See page I-2-5 for an example.

mL KSCN	%Transmittance	Absorbance
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** (Goal #2, part c) **versus** **[FeSCN<sup>2+</sup>]** (Goal #2, part b). **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.

e. **Perform a linear regression of the values used in Goal #2, part d** above (Absorbance versus [FeSCN<sup>2+</sup>]). 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  $\epsilon$ , the **molar absorptivity coefficient**, which we will use in Goal #3 next.

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

slope =  $\epsilon$  = \_\_\_\_\_

y-intercept = \_\_\_\_\_

correlation coefficient ( $r$ ) = \_\_\_\_\_ (*r is not the same as R<sup>2</sup>!*)

*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  $\epsilon$ , 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  $M_1 = 2.00 \times 10^{-3} \text{ M}$  for the initial molarity of KSCN and use a final volume  $V_2 = 10.00 \text{ mL}$ . (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  $M_1 = 2.00 \times 10^{-3} \text{ M}$  for the initial molarity of  $Fe^{3+}$  and use a final volume  $V_2 = 10.00 \text{ mL}$ . 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.)

<u>mL KSCN (<math>V_1</math>)</u>	<u>diluted [KSCN] (<math>M_2</math>)</u>	<u>mL <math>Fe^{3+}</math> (<math>V_1</math>)</u>	<u>diluted <math>[Fe^{3+}]</math> (<math>M_2</math>)</u>
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. Convert your %Transmittance values *from part B* into **absorbance** using:

$$\text{Absorbance} = \log(100\%/\% \text{Transmittance})$$

Reminder: these %Transmittance values came from Goal #1 Part B (the Lab Video). Absorbance values are unitless and will constitute the y-axis in the graph. (see page I-2-5 for an example of conversion.)

mL KSCN	%Transmittance (Part B)	Absorbance
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  $[\text{FeSCN}^{2+}]$**  using:

$$\text{Absorbance} = A = \epsilon bc$$

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

value of  $\epsilon$  = \_\_\_\_\_ (from Goal #2, part e)

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

**Goal #3 (continued)**

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

$$[\text{SCN}^{-1}]_{\text{eq}} = [\text{SCN}^{-1}]_{\text{i}} - [\text{FeSCN}^{2+}]_{\text{eq}}$$

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

mL KSCN	$[\text{SCN}^{-1}]_{\text{i}}$	$[\text{FeSCN}^{2+}]_{\text{eq}}$	$[\text{SCN}^{-1}]_{\text{eq}}$
1.00	_____	_____	_____
2.00	_____	_____	_____
3.00	_____	_____	_____
4.00	_____	_____	_____
5.00	_____	_____	_____

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

$$[\text{Fe}^{3+}]_{\text{eq}} = [\text{Fe}^{3+}]_{\text{i}} - [\text{FeSCN}^{2+}]_{\text{eq}}$$

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

mL $\text{Fe}^{3+}$	$[\text{Fe}^{3+}]_{\text{i}}$	$[\text{FeSCN}^{2+}]_{\text{eq}}$	$[\text{Fe}^{3+}]_{\text{eq}} (\text{M})$
5.00	_____	_____	_____
5.00	_____	_____	_____
5.00	_____	_____	_____
5.00	_____	_____	_____
5.00	_____	_____	_____

**Goal #3 (continued)**

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

$$K_c = \frac{[\text{FeSCN}^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^{-1}]}$$

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

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

**Average of five  $K_c$  values =** \_\_\_\_\_

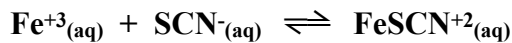
**Parts per thousand =** \_\_\_\_\_

*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  $2.00 \times 10^{-3}$  M  $\text{Fe}(\text{NO}_3)_3$  with 5.00 mL  $2.00 \times 10^{-3}$  M KSCN. She finds that in the equilibrium mixture the concentration of  $\text{FeSCN}^{2+}$  is  $1.40 \times 10^{-4}$  M.  $\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightleftharpoons \text{FeSCN}^{2+}(\text{aq})$



- a. What is the initial concentration in solution of the  $\text{Fe}^{+3}$  and  $\text{SCN}^-$ ?
- b. Create an ICE table for this problem. What is the equilibrium constant value for this reaction?