

Equilibria: Measuring an Equilibrium Constant Using Spectrophotometry

Chemistry 102: *General Chemistry I*

Post-Lab & Lab Report #9



Sitthiphol Yuwanaboon

Professor Benny Ng

Lab Professor

April 8, 2020

Title: Equilibria: Measuring an Equilibrium Constant Using Spectrophotometry

Purpose:

Observing the equilibrium of Fe^{3+} and SCN^- , and measuring the equilibrium constant

Spectrophotometry is a method to measure how much a chemical substance absorbs light by measuring the intensity of light as a beam of light passes through sample solution in certain wavelength of the light for the known substance. "Spectrophotometry is one of the most useful methods of quantitative analysis in various fields such as chemistry, physics, biochemistry, material and chemical engineering and clinical applications." (libre Text)

Procedure

Part 1: Calibration curve using standard $\text{Fe}(\text{SCN})_2^{2+}$ solutions.

1. Copy the following table from the canvas into your notebook and then prepare the solutions A-1 through A-7 and B1-B5. Note that the concentration (we do online lab since we get the given number from the canvas)
2. Calculate the concentration of $\text{Fe}(\text{SCN})_2^{2+}$ in each solution and complete the values in the column labeled 'Calculated $[\text{Fe}(\text{SCN})_2^{2+}]$ ' of table 1. For example, for solution A-2, the concentration of the $\text{Fe}(\text{SCN})_2^{2+}$ is the same as the initial concentration of SCN^- (the limiting reagent). Therefore, the concentration is
3. Using Excel, plot a graph of $[\text{Fe}(\text{SCN})_2^{2+}]$ on the x axis versus absorbance on the y axis.
4. Add a straight line fit to your data and include the equation (5 sig. fig.) of the line on your graph.

Part II Measuring the equilibrium constant, K_c

1. By using an equation we got from part I we can find the values from an equation then solve for X which is the concentration of $\text{Fe}(\text{SCN})_2^{2+}$ as absorbances are given in the this experiment

Part III calculating K_c

1. To find the initial values of the K_c we have to use the formula from the lab manual. The new initial concentrations (after dilution) of the Fe^{3+} and SCN^- are calculated by taking the number of moles and dividing by the total volume of the solution. For example, the initial $[\text{SCN}^-]$ in solution B-1 is given by $[\text{SCN}^-]_{\text{init}} = \text{moles } \text{SCN}^- / \text{total volume} = 1.82 \times 10^{-4} \text{ M}$ And the initial Fe^{3+} concentration in solution B-1 is given by $[\text{Fe}^{3+}] = \text{moles } \text{Fe}^{3+} / \text{total volume} = 9.1 \times 10^{-4} \text{ M}$
2. Then we had initial values of the concentration of each species, so we subtract from the $\text{Fe}(\text{SCN})_2^{2+}$ eq so we find the different, since the initial values of concentration will decrease since we create more product in the experiment, $[\text{SCN}^-]_{\text{init}} - [\text{Fe}(\text{SCN})_2^{2+}]_{\text{eq}}$ then we will find the final equilibrium solution of each species.
3. Then calculate from K_c by, the equation for K_c is $K_c = [\text{Fe}(\text{SCN})_2^{2+}] / [\text{Fe}^{3+}] \cdot [\text{SCN}^-]$

4. Then we find the deviation from an avg of Kc
5. Then we find average deviation.

Data:

Solution	0.200 M $\text{Fe}(\text{NO}_3)_3$ (mL) Use a Volumetric Pipet or a Buret	0.00200 M KSCN (mL) Use a Buret	0.100 M HNO_3 (mL) Fill up the 100 mL volumetric flask	Calculated $[\text{Fe}(\text{SCN})^{2+}]$ (M) 3 sig. fig.	Measured Absorbance Should increase linearly
A-1	25.00	0.00	75.00	0.00E+00	0.00
A-2	25.00	1.00	74.00	2.00E-05	0.07
A-3	25.00	2.00	73.00	4.00E-05	0.20
A-4	25.00	4.00	71.00	8.00E-05	0.35
A-5	25.00	6.00	69.00	1.20E-04	0.47
A-6	25.00	8.00	67.00	1.60E-04	0.63
A-7	25.00	10.00	65.00	2.00E-04	0.93

Slope (5 sig. fig. --> Adjust the number of sig. fig on the Format Trendline Label)
$y = 4356.7x - 0.0073$
y-intercept (5 sig. fig.)
-0.0073
r^2 5 sig. fig.
0.98250

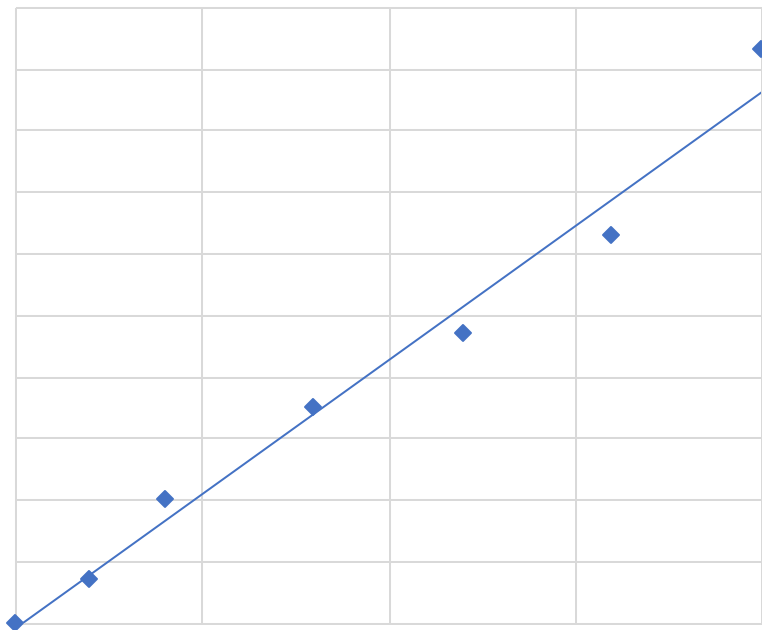
Solution	0.00200 M $\text{Fe}(\text{NO}_3)_3$ (mL) Use a Buret	0.00200 M KSCN (mL) Use a Buret	0.100 M HNO_3 (mL) Use a Buret	Measured Absorbance Should increase linearly	Calculated $[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}$ (M) 3 sig. fig. & See pg. 8 of the lab handout
B-1	5.00	1.00	5.00	0.2	4.76E-05
B-2	5.00	2.00	4.00	0.31	7.28E-05
B-3	5.00	3.00	3.00	0.41	9.58E-05
B-4	5.00	4.00	2.00	0.54	1.26E-04
B-5	5.00	5.00	1.00	0.64	1.49E-04

Table 3. Calculated Concentrations needed to find Kc

Solution	Initial $[\text{Fe}^{3+}]$ (M) 3 sig. fig.	Initial $[\text{SCN}^-]$ (M) 3 sig. fig.	$[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}$ (M)	$[\text{Fe}^{3+}]_{\text{eq}}$ (M) 3 sig. fig. + 3 extra digits	$[\text{SCN}^-]_{\text{eq}}$ (M) 3 sig. fig. + 3 extra digits
B-1	9.09E-04	1.82E-04	4.76E-05	8.61E-04	1.34E-04
B-2	9.09E-04	3.64E-04	7.28E-05	8.36E-04	2.91E-04
B-3	9.09E-04	5.45E-04	9.58E-05	8.13E-04	4.49E-04
B-4	9.09E-04	7.27E-04	1.26E-04	7.83E-04	6.01E-04
B-5	9.09E-04	9.09E-04	1.49E-04	7.60E-04	7.60E-04

Table 4. Summary of Equilibrium constants for solutions B1 to B5

Solution	K_c 3 sig. fig. + 3 extra digits	pt	Solution	Deviation = $ K_c$ - Avg. $K_c $ 3 sig. fig. + 3 extra digits	pt
B-1	4.13E+02	0.1	B-1	1.13E+02	0.2
B-2	2.99E+02	0.1	B-2	7.51E-01	0.2
B-3	2.62E+02	0.1	B-3	3.80E+01	0.2
B-4	2.68E+02	0.1	B-4	3.20E+01	0.2
B-5	2.58E+02	0.1	B-5	4.20E+01	0.2
Avg. K_c	3.00E+02	0.50	Avg. deviation of K_c	4.52E+01	1.00



Calculations

Part I:

1. we given the absorbance and we have to find the graph.
2. First of all we have to the limiting reagent, which SCN- by moles we got is the smallest in the solution.
3. We can find the total volumes of all species which is $75\text{ml} + 24\text{ml} + 1\text{ml} = 100\text{m}$ then we find the mole from SCN- By using the $.002\text{ M SCN-} \times 1\text{ml} / \text{total volume}$ then we get the first portion equal to $2.00\text{E-}05$
4. Then we can find all the rest.

Then we plot the graph Absorbance (y) vs. Molarity

5. We an equation by using linear fit which is equal to ($y = 4356.7x - 0.0073$)

Part B:

1. We use the formula then we substitute the absorbance then solve for x Example: .2
 $= 4356.7x - 0.0073$ then we solve we got the $x = 4.76E-05$
2. Then solve all the rest.

Part 3

1. Then we calculating for initial concentration of each species, by using $M \cdot V / \text{total volume}$:
 $.002M \cdot 5\text{ml} / 11\text{ml} = 9.09E-04$
2. Then we solve like this for each species such SCN^- and Fe^{3+}
3. For SCN^- $1\text{ml} \cdot 0.002\text{ M} / 11\text{ ml} = 1.82E-04$
4. We got the concentration for X which is $[\text{FeSCN}^{2+}]$ then we can using subtraction since the molarity of initial will decrease since we are making more product
5. Then we find $[\text{Fe}^{3+}] - [\text{FeSCN}^{2+}]$ as same as $[\text{SCN}^-] - [\text{FeSCN}^{2+}]$, then got the value of new equilibrium constant: which are

8.61E-04	1.34E-04
----------	----------

Then we find all the rest

6. To calculating K_c we use EQ expression : $[\text{Fe}(\text{SCN})^{2+}] / [\text{Fe}^{3+}] \cdot [\text{SCN}^-]$

Then we find first $4.76E-05 / (8.61E-04 \cdot 1.34E-04) = 413\text{ M}$

7. So we find all the rest then sum it up we got AVG M $K_c = 300\text{ M}$
8. To find deviation $|\text{Avg} - \text{cal}| =$

Deviation = $\text{Kc} - \text{Avg. Kc}$ 3 sig. fig. + 3 extra digits
1.13E+02
7.51E-01
3.80E+01
3.20E+01
4.20E+01

Then we are averaging them $(113 + 75.1 + 38 + 32 + 42) / 5 = 45.2$

9. Avg Deviation = $4.52E+01$

Post-lab Questions.

Why do we not have to include water in the equilibrium expression for this experiment? Why can we leave pure liquids (like the solvent water in this reaction) and pure solids out of equilibrium expressions?

Pure liquid like water and solid are left of the equilibrium expression because it does not affect the amount of the reactant to have in chemical reaction and we determine the quantities as 1 and it does affect the concentration of the equilibrium solution.

What do your values for K_c tell you about where this particular equilibrium lays (does it favor products or reactants)?

Since K_c can tell us about what we are making it, if K_c is greater than 1, it is the product favor. In contrast if it is less than 1, it is reactants favored. Since we have more product, the majority of the substance are product by equilibrium expression $[Products]/[Reactants]$. In my case the K_c is greater than 1 (300 M) so this means it was making more product and it is shift to the right.

Should the values of K_c for solutions B-1 through B-5 be the same or different? Explain. How close are your values?

From B-1 to B-5 we have to look at the limiting reagent which can determine the solution we want since all reactant can not all react thoroughly in chemical reactions, so the product that form is by smaller number of mole that we have by using calibrating curve to determine the solution.

Can we calculate values for K_c for solutions A-1 through A-7? Why or why not?

Since everything is given as same as part B, I can calculate the amount of initial solution and K_c from part A, and SCN is also the limiting reactant in this chemical reaction, so we find equilibrium constant in the reaction then subtract from the initial value of concentration, then we can calculate for K_c of all in part A.

If this reaction was exothermic as written, would you expect the value for K_c to increase or decrease at a higher temperature according to Le Chatelier's principle? Briefly explain.

By increasing the temperature in exothermic, the reaction will make more reactant and shift to the left. So the K_c would be smaller than the K_c at the certain temperature since it makes more reactant.

Does the value of K_c tell us anything about how fast equilibrium is attained? Briefly explain.

By using unit of Molarity (Mol/Liter) does not have unit of time to determine the rate of change of chemical equilibrium but it tells us about how what is dominant and minor in the solution that is concentrate and chemical equilibrium has no unit such S^{-1} .