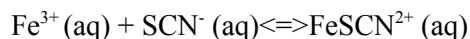


## Chemical Equilibrium: Finding a Constant, $K_c$

### **Introduction:**

Every chemical reaction that takes place within a closed system, with no external disturbances, will eventually reach a state of equilibrium. At this state, the rate of the forward reaction equates the rate of the reverse reaction, but the concentrations of the entities on both sides are not necessarily equal. The direction in which the reaction proceeds to achieve equilibrium can be determined by finding the equilibrium constant  $K_c$ . The value of  $K_c$  is calculated by dividing the concentrations of the products, the entities on the right, by the reactants, the entities on the left. This value is specific to one temperature, and thus will change as temperature changes in accordance to the reaction whether it's exothermic or endothermic.

In this investigation we will experimentally determine the equilibrium constant for the reaction between Iron (III) and thiocyanate at a constant temperature:



The  $\text{Fe}^{3+}$  ion will be obtained from  $\text{Fe}(\text{NO}_3)_3$  solution and the  $\text{SCN}^{-}$  ion will be obtained from KSCN solution.

When the two solutions are mixed, the two ions will react to form a more complex ion of  $\text{FeSCN}^{2+}$  and will emit a red color. In this reaction for every mole of  $\text{FeSCN}^{2+}$  formed, one mole of  $\text{Fe}^{3+}$  and  $\text{SCN}^{-}$  will be formed thus the expression for the equilibrium constant will be:

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

In this investigation, there will be 4 solutions prepared with different concentrations of KSCN and  $\text{H}_2\text{O}$  but will all have the same concentrations of  $\text{Fe}(\text{NO}_3)_3$ . A 5th solution will remain as the standard solution where large amounts  $\text{Fe}(\text{NO}_3)_3$  with a 0.200 molarity and small amount of KSCN will be added together. This so  $[\text{Fe}^{3+}]$  is as much as a 100 times more than the previous 4 trials, so that according to Le Chatelier's principle: the reaction will proceed completely to the right so that it can be considered qualitative or complete and all of the  $[\text{SCN}^{-}]$  has been 100% used up, and thus we can assume that the  $[\text{FeSCN}^{2+}]_{\text{eq}}$  is equal to  $[\text{SCN}^{-}]_i$ . Le Chatelier's principle can also be applied in the first 4 trials, when as we change the initial concentrations of  $[\text{SCN}^{-}]$  the equilibrium constant for all trials must be quite similar so long that the temperature remains controlled. This can be explained as the following: when more of  $\text{SCN}^{-}$  is added to the solution, it is considered a disturbance to the system or a stress, and the system will adjust itself to counteract this disturbance by increasing the forwards rate so that excess  $\text{SCN}^{-}$  is consumed thus equilibrium can be reached again but it will be shifted slightly to the right more.

Therefore, our purpose of this investigation is to determine the equilibrium constant for iron (III) reacting with thiocyanate to form iron (III) thiocyanate using a colorimeter. This is based on the relation between absorbance and the concentration of  $\text{FeSCN}^{2+}$  where they are directly related, this is known as Beer's law. Essentially, a colorimeter is a light-sensitive instrument that can determine the intensity of the color of a substance by measuring the amount of color it absorbed from different parts of the visible-light spectrum. Although we are using this device in our investigation, colorimetry is commonly used in the field of biology. For instance, the growth of yeast and bacterial cultures can be determined using colorimetry, so that as the culture grows it'll be seen to absorb more light as it becomes more opaque. In addition to that, a colorimeter can be used to determine the haemoglobin content of the blood as a measure of health and wellness of a patient. This concept is applied in our investigation when we use a colorimeter to determine the absorbance of each trial of blue light, as the solutions obtained from the experiments will have a dark red color. However, before using it, it must be calibrated to ensure that we get consistent readings, and that can be done by filling the cuvette with distilled water to use it as the control solution, and the setting the wavelength of the colorimeter to 470 nm so that it's blue.

Thus based on the colorimetric measures and Le Chatelier's principle we are predicting that for the 4 trials, the  $K_c$  value will be very similar despite the change in the initial concentration of  $\text{SCN}^{-}$ .

### **An overview of calculations:**

To experimentally determine the equilibrium constant for the reaction of iron (III) and thiocyanate we must first calculate the initial concentrations of the two ions  $\text{Fe}^{3+}$  and  $\text{SCN}^-$ . This can be done by determining the new, diluted concentrations using the formula  $C_1V_1=C_2V_2$ . Where we already know that  $\text{Fe}^{3+}$  and  $\text{SCN}^-$  have a molarity of 0.0020 M and the total mL is 10 mL. So for  $\text{Fe}^{3+}$  it'll be:  $\text{Fe}(\text{NO}_3)_3 / \text{total mL} \times (0.0020 \text{ M})$ . Similarly for  $\text{SCN}^-$  it'll be  $\text{KSCN} / \text{total mL} \times (0.0020 \text{ M})$

Then to determine  $[\text{FeSCN}^{2+}]$  at equilibrium, we must rely on colorimetry, which will measure the absorbance of each solution of light with a wavelength of 470 nm- blue light . Once absorbance of each trial is determined, the  $[\text{FeSCN}^{2+}]_{\text{eq}}$  using the formula:

$$\frac{A_{\text{eq}}}{A_{\text{std}}} \times [\text{FeSCN}^{2+}]_{\text{std}}$$

When the equilibrium concentration of  $\text{FeSCN}^{2+}$  has been determined we can thus calculate the concentration of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$ , where we rely on the fact that the molar ratio in this chemical reaction 1:1:1 and because  $\text{Fe}^{3+}$  and  $\text{SCN}^-$  were firstly added, their equilibrium concentrations will be lower than the initial. Therefore,  $[\text{Fe}^{3+}]_{\text{eq}}$  can be determined by:

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

And  $[\text{SCN}^-]_{\text{eq}}$  can be determine by:

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

Once the equilibrium concentrations have been all determined, they can be substituted into the expression of  $K_c$ .

### Materials and Methods:

(Refer to lab handout)

### Results and analysis:

(Refer to lab handout)

### Conclusion and Discussion:

The results obtained from this investigation demonstrate the idea that the initial concentration of the entities on the right side of the equation have minimum effect on the equilibrium constant for the reaction, and thus our hypothesis is accepted and true, the average of  $K_c$  value was calculated at 54.01 at 20°C. The values of  $K_c$  for the 4 trials ranges from 56.60 to 49.70 (please refer to Appendix A for sample calculations). However there's no extreme change between the different  $K_c$  values as they were fairly consistent

This investigation was largely based on calculations and substituting numbers already provided along with the lab handout into expressions, hence many sources of error have already been eliminated since a calculator was used.

The few sources of error that can be found can only originate from the step of determining the absorbance of the solutions of each trial. One noticeable error was that the readings on the screen of the colorimeter continued to fluctuate as we observed, and it didn't seem to stop at a certain value so we had to record the number on the screen 10 seconds after the curveter has been placed into the colorimeter, even if that meant it was different afterwards. But, those differences were a matter of  $\pm 0.001$ , so it wouldn't have had a massive impact on our calculations. Nonetheless, what would have improved this investigation would that if more than one reading of absorbance was done for each of the solutions, thus more reliable results would have been obtained. Or perhaps if the solutions have tried into two different calorimeters to improve the reliability of the results.

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### Citations:

Nelson Chemistry Textbook Grade 12.

-The Equilibrium Constant  $K_c$  (16 Jan 2014)

<https://www.khanacademy.org/science/chemistry/chemical-equilibrium/equilibrium-constant/a/the-equilibrium-constant-k>

**Appendix A**Sample calculations using Trial 1**[Fe<sup>3+</sup>]<sub>i</sub>:**Fe(NO<sub>3</sub>)<sub>3</sub> mL = 5 mL

Total mL = 10 mL

$$[\text{Fe}^{3+}]_i = \frac{\text{Fe(NO}_3)_3 \text{ mL}}{\text{total mL}} \times 0.0020 \text{ M}$$

$$[\text{Fe}^{3+}]_i = \frac{5 \text{ mL}}{10 \text{ mL}} \times 0.0020 \text{ M}$$

$$[\text{Fe}^{3+}]_i = 1.0 \times 10^{-3} \text{ mol/L}$$

**[SCN<sup>-</sup>]:**

KSCN mL = 2 mL

Total mL = 10 mL

$$[\text{SCN}^-] = \frac{\text{KSCN mL}}{\text{total mL}} \times 0.0020 \text{ M}$$

$$[\text{SCN}^-] = \frac{2 \text{ mL}}{10 \text{ mL}} \times 0.0020 \text{ M}$$

$$[\text{SCN}^-] = 4.0 \times 10^{-4}$$

**[FeSCN<sup>2+</sup>]<sub>eq</sub>:**

$$A_{\text{eq}} = 0.116$$

$$A_{\text{std}} = 1.247$$

$$[\text{FeSCN}^{2+}]_{\text{std}} = 0.00020 \text{ mol}$$

$$[\text{FeSCN}^{2+}]_{\text{eq}} = A_{\text{eq}} \times \frac{0.00020 \text{ mol}}{A_{\text{std}}}$$

$$[\text{FeSCN}^{2+}]_{\text{eq}} = 1.86 \times 10^{-5} \text{ mol/L}$$

**[Fe<sup>3+</sup>]<sub>eq</sub>:**

$$[\text{Fe}^{3+}]_i = 1.0 \times 10^{-3} \text{ mol/L}$$

$$[\text{FeSCN}^{2+}]_{\text{eq}} = 1.86 \times 10^{-5} \text{ mol/L}$$

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

$$[\text{Fe}^{3+}]_{\text{eq}} = 1.0 \times 10^{-3} - 1.86 \times 10^{-5} \\ = 9.8 \times 10^{-4} \text{ mol/L}$$

**[SCN<sup>-</sup>]<sub>eq</sub>:**

$$[\text{SCN}^-]_i = 4.0 \times 10^{-4} \text{ mol/L}$$

$$[\text{FeSCN}^{2+}]_{\text{eq}} = 1.86 \times 10^{-5} \text{ mol/L}$$

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

$$[\text{SCN}^-]_{\text{eq}} = 4.0 \times 10^{-4} - 1.86 \times 10^{-5} \\ = 3.8 \times 10^{-4} \text{ mol/L}$$

**K<sub>c</sub> value:**

$$[\text{FeSCN}^{2+}]_{\text{eq}} = 1.86 \times 10^{-5} \text{ mol/L}$$

$$[\text{Fe}^{3+}]_{\text{eq}} = 9.8 \times 10^{-4} \text{ mol/L}$$

$$[\text{SCN}^-]_{\text{eq}} = 3.8 \times 10^{-4} \text{ mol/L}$$

$$K_c = 49.70$$