Spectrophotometric determination of the $Fe(SCN)^{2+}$ equilibrium constant

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Introduction

This research poster focuses on exploring the equilibrium constant in the reaction between thiocyanate anion (SCN–) and iron(III) cation (Fe3+) in water. Understanding this constant is vital in grasping how conditions influence equilibrium. The experiment investigates how changes in reactant concentrations impact the equilibrium position, using **Le Chatelier's Principle** as a guide. This principle explains how systems at equilibrium respond to changes, adjusting to maintain balance. For example, if there's more of a substance, the system tries to restore balance by favoring either the forward or reverse reaction.

By precisely manipulating reactant concentrations, the experiment aims to uncover how changes affect the equilibrium position of the iron(III) thiocyanate reaction. The main question driving this investigation is: *How does altering each reactant's concentration shift the equilibrium position of the iron(III) thiocyanate reaction?* Understanding this dynamic equilibrium not only sheds light on this specific reaction but also helps elucidate broader principles governing chemical equilibria. This research aims to reveal the complexities of the equilibrium constant in the interaction between thiocyanate anion and iron(III) cation, providing insights into their equilibrium positions.

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftarrows Fe(SCN)^{2+}(aq)$$

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ system at equilibrium}$$

Materials Utilized

Chemicals & Aqueous Solutions

- .0025M Iron(III) Nitrate
- .5M Potassium Thiocyanate
- .1M Nitric Acid
- .0025M Potassium Thiocyanate

Part A & Part B: Equipment

- Spectrophotometer for observing absorbance with respect to concentration

Methods: Determining Beer-Lambert Law

In order to determine the concentration of products for the reaction in a state of equilibrium, we utilized Iron Thiocyanate's properties of light and color. With help from a spectrophotometer, we measured the absorbance of impinging light on a sample of Iron Thiocyanate at different concentrations.

We plotted such values on the graph of Absorbance vs Concentration (Plot 1) and observed a linear relationship between the two data values.

Table 1 illustrates a list of different concentrations for Iron Thiocyanate and its associated absorbance values. The absorbance, A, is measured on a logarithmic scale of the ratio of the initial intensity of light to the absorbed intensity of light.

$$A = \log_{10}\left(\frac{I_0}{I}\right)$$

Methods: Measuring K for Equilibrium Reaction

In actually determining K, we must setup an equilibrium reaction for Iron(III) nitrate and Thiocyanate. This differs from the reaction done in Part A because we will NOT be using excess reactants to create 100% products; instead we will be changing the amounts/concentrations of our reactants to have a reaction that reaches equilibrium.

Table 2 lists the amounts and concentrations of each reactant for each trial along with the absorbances for each of those trails.

From the absorbance values, we can determine the concentration of the product - by way of the Beer-Lambert plot made in Part A. Specifically, the equation below specifies how we do so.

$$\left[Fe(SCN)^{2+}\right]_j = \frac{(A_j - b)}{m}$$

After combining each reactant in a test tube, we wait for roughly 1 minute till the reaction reaches equilibrium. In essence, we wait until we stop observing any macroscopic color changes to the solution. We then sample about 1mL of each solution into a cuvette to be measured for its absorbance at a wavelength of around 460.1nm.

Discussion

Part A: Analyzing trends for Beer-Lambert Law

- As the concentration of the reactants increases, the concentration of the products follows suite
- This in turn increases the pigmentation for the solutions as higher concentrations are darker
- Referring to the equation in obtaining absorbance, as the solution becomes darker, the absorbance increases since the impinging light is being absorbed more
- This is as to why there is a linear relationship between the concentration of the solution and its absorbance. The slope of the line is a representation for the Beer-Lambert Law

Part B: Analyzing trends for adjusting reactant concentration on K

- Table 2 illustrates explicit trends for how the absorbance is directly correlated with the concentration for the Thiocyanate Anion
- The increasing absorbance indicates increased concentration for Iron Thiocyanate, evident in Table 3

Conclusion

In changing the concentrations of the reactants, we observe an increase in the concentrations in the products. However, on average, the data does not indicate significant correlation for the equilibrium constant K. **In other words, changing the concentration of the reactants does not significantly affect the equilibrium constant K.**

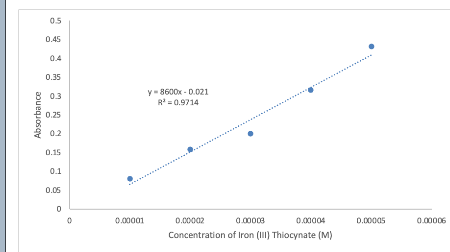
Consider the following to support the claim above:

- Table 3 indicates for trials 6-10 how increasing the concentration of the thiocyanate anion does not increase the equilibrium constant K. Rather, the constant shifts from 134.65 to 144.35 sporadically, and not consistently.
  - The same can be said for trials 10-15 where the iron(III) cation is doubled in concentration and the thiocyanate anion is varying.
- Le Chatlier's Principle explains how equilibrium systems that exhibit a change will alter the state of their system to return back to equilibrium.
  - In changing the concentration of the reactants, the reaction produces more product - Iron Thiocyanate - in order to maintain a consistent equilibrium constant

Beer-Lambert Plot for Determining Concentration

Test Tube	Fe^{3+} V (ml)	SCN^- V (ml)	HNO_3 V (ml)	Total V (ml)	$[Fe(SCN)^{2+}]$ (mol/L)	Absorbance
1	1	5	4	10	0.00001	0.08
2	2	5	3	10	0.00002	0.158
3	3	5	2	10	0.00003	0.2
4	4	5	1	10	0.00004	0.316
5	5	5	0	10	0.00005	0.431

Table 1: Data for Beer's Law



Plot 1: Beer-Lambert Law Plot

Data for Determining K

Test Tube	Fe^{3+} V (ml)	SCN^- V (ml)	HNO_3 V (ml)	Total V (ml)	Absorbance
6	1	1	5	7	0.114
7	1	1.5	4.5	7	0.222
8	1	2	4	7	0.279
9	1	2.5	3.5	7	0.305
10	1	3	3	7	0.375
11	2	1	4	7	0.295
12	2	1.5	3.5	7	0.411
13	2	2	3	7	0.497
14	2	2.5	2.5	7	0.629
15	2	3	2	7	0.733

Table 2: Concentrations of reactants & measured Absorbances

Test Tube	$[Fe^{3+}]_i$ (mol/L)	$[SCN^-]_i$ (mol/L)	$[Fe(SCN)^{2+}]_{eq}$ (mol/L)	$[Fe^{3+}]_{eq}$ (mol/L)	$[SCN^-]_{eq}$ (mol/L)	K
6	0.0003571	0.000357	1.5698E-05	0.00034145	0.00034145	134.645963
7	0.0003571	0.000536	2.8256E-05	0.00032889	0.00050746	169.301401
8	0.0003571	0.000714	3.4884E-05	0.00032226	0.0006794	159.327502
9	0.0003571	0.000893	3.7907E-05	0.00031924	0.00085495	138.888616
10	0.0003571	0.001071	4.6047E-05	0.0003111	0.00102538	144.349775
11	0.0007143	0.000357	3.6744E-05	0.00067754	0.0003204	169.262989
12	0.0007143	0.000536	5.0233E-05	0.00066405	0.00048548	155.815107
13	0.0007143	0.000714	6.0233E-05	0.00065405	0.00065405	140.800831
14	0.0007143	0.000893	7.5581E-05	0.0006387	0.00081728	144.792625
15	0.0007143	0.001071	8.7674E-05	0.00062661	0.00098375	142.22898
					Mean	149.941379
					St. Dev.	12.5947731

Table 3: Calculating K for each trial, determining Average K