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CHMG 146 Section 12

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Lab 3: Keq of an Equilibrium

Abstract

In this experiment, a real-world application was used to determine the equilibrium constant using Beer's Law. It is noted that the K_{eq} value is different for each reaction due to it signifying how much chemical reactant to use to avoid unnecessary waste. The molarities of the two solutions (Thiocyanate and Ferric Nitric Acid) were determined to find this value. Then they were placed in a spectrophotometer to find their absorbances at different levels of concentration. Then, the Equilibrium Expression was used to attain graphable points. These points are then graphed, and a line-of-best-fit was found. From this plotted graph, it was determined that the K_{eq} of this reaction was 140.

Experimental

In this experiment, Beer's Law is applied to find the equilibrium constant for a given reaction. To begin, the Thiocyanate solution was prepared by mixing 10mL of 0.002M Potassium Thiocyanate, 25mL of 2M Nitric Acid, and 65 mL of distilled water in a flask. Next, the Ferrothiocyate solution was prepared by adding 5mL of 0.1M Fe^{3+} to the beaker containing the Thiocyanate solution. Then, Equation 1 was used to calculate the molarity of Fe^{3+} and SCN^{-}.

Equation 1: Used to find the concentration of a given solution

$$\frac{mL\ of\ solute * Molarity\ of\ Solute}{total\ mL\ of\ solution}$$

Next, a spectrophotometer was used to determine the absorbance of the solution in a cuvette. After, the cuvette was emptied back into the solution flask, and 1 more mL of Fe^{3+} was added, and the absorbance of the new solution was determined. This was repeated in order to get a total of 10 different readings. This was placed in *Results Table 1.*

Results and Discussion

Table 1: Shows the molarity of both solutions and their absorbances

Trial No.	**$[SCN^-]$ M**	**mL of Fe^{3+}**	**$[Fe^{3+}]$ M**	**Abs**	**λ (nm)**
1	1.90×10^{-4}	5	4.76×10^{-3}	0.41	458
2	1.89×10^{-4}	6	5.66×10^{-3}	0.45	457
3	1.87×10^{-4}	7	6.54×10^{-3}	0.48	463
4	1.85×10^{-4}	8	7.41×10^{-3}	0.51	455
5	1.83×10^{-4}	9	8.26×10^{-3}	0.53	456
6	1.82×10^{-4}	10	9.09×10^{-3}	0.55	453
7	1.80×10^{-4}	11	9.91×10^{-3}	0.57	460
8	1.79×10^{-4}	12	1.07×10^{-2}	0.58	467
9	1.77×10^{-4}	13	1.15×10^{-2}	0.59	456
10	1.75×10^{-4}	14	1.23×10^{-2}	0.60	458

Next, to get the plotted values, *Equation 2* was used, to which *Equations 3 and 4* were derived to create graphable points. Using these equations, *Table 2* and *Chart 1* was created.

Equation 2: The Equilibrium Expression by H.A. Frank and R.L. Oswald

$$K_{eq} = \frac{[FeSCN^{2+}_{(aq)}]}{[Fe^{3+}_{(aq)}] * [SCN^{-}_{(aq)}]}$$

Equation 3: Used to represent the X values in Equation 2

$$X = \frac{Abs * ([Fe^{3+}_{(aq)}] + [Fe^{-}_{(aq)}])}{[Fe^{3+}_{(aq)}] * [Fe^{-}_{(aq)}]}$$

Equation 4: Used to represent the Y values in Equation 2

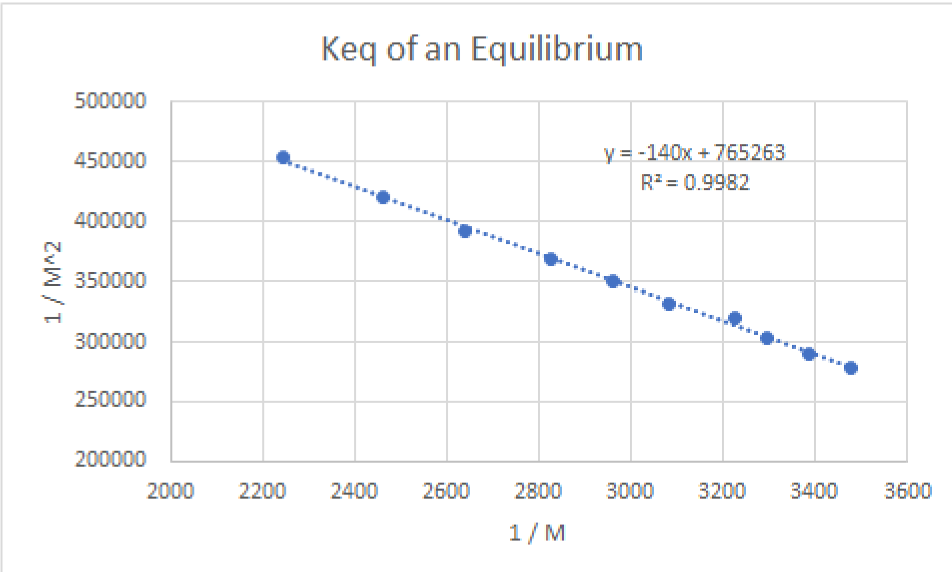
$$Y = \frac{Abs}{[Fe^{3+}_{(aq)}] * [Fe^{-}_{(aq)}]}$$

Table 2: The X and Y values found using Equation 3 and 4

Trial No.	X	Y
1	2244.03	453339.23
2	2460.46	420662.96
3	2640.24	392483.93
4	2825.03	369043.74
5	2960.34	350626.5
6	3082.48	332450.83
7	3224.18	319542.55
8	3294.43	302824.62
9	3384.64	289855.07

10	3477.35	278745.64

Chart 1: Shows the K_{eq} of equilibrium from determined values.



From this graph, it was derived that R^2 is 0.9982, and the equation of the line of best fit is y = -140x + 765263. Thus, K_{eq} is 140 (unitless), and the y-intercept is 765263. Upon searching for the reaction, the K_{eq} of this specific reaction is 126. Sources of error would include the fact that the solution might not have been entirely homogeneous due to a lack of a stir plate and insufficient trials to validate the solution. These error sources can be cleared up by making sure that the solution remained homogeneous and more trials.