

SPECTROPHOTOMETRIC DETERMINATION OF AN EQUILIBRIUM CONSTANT

Lab report

Name:

ID #:

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Lab Date:

1. Looking at Table 5 (data sheet), how is K_c related to the initial concentrations ($[Fe^{3+}]_0$ and $[SCN^-]_0$)? Is it what you expected? (Explain briefly) (5 marks)

Looking at Table 5 in the results, it is somewhat evident that K_c is actually independent of the initial concentrations of the reactants since they are relatively consistent among the five test tubes. Furthermore, the K_c values are much more relative as opposed to the significant changes in K_c due to temperature variation which holds true because the equilibrium constant is affected by temperature.

However, looking at the data for any trends, it appears that K_c drops as the concentration of $[SCN^-]$ increases, but this should not be true. So, accounting for relative human error and limited machine capabilities, the value only fluctuated minimally from the average which was 204 (M⁻¹). Thus the results are not exactly as I had expected, but ignoring the small deviations, the results are based on Le Chatelier's Principle which states that when a chemical system in equilibrium is disturbed there will be a shift in the system to reduce the effect of that disturbance. So, by increasing $[SCN^-]$ from tube 1 to 5, there should be no change of K_c since equilibrium will shift to the right of the reaction to increase the concentration of $[FeSCN^{2+}]$ and balance out.

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

2. To obtain $[FeSCN_{2-}]_{eq}$ in tube 5, you made the assumption that 100% of the ions SCN^- had reacted. Now that you know the value for K_c (average in table 5), build an ICE diagram to calculate the true percentage of SCN^- reacted in tube 5. Comment on the validity of the assumption. (Make sure you include the ICE diagram in your answer) (10 marks)

ICE Table

	$[Fe^{3+}]$	$+ [SCN^-]$	$\rightleftharpoons [FeSCN_{2-}]$
Initial	0.180 M	0.000199 M	0 M
Change	-x	-x	+x
Equilibrium	$(0.180 - x)$ M	$(0.000199 - x)$ M	x M

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