SPECTROPHOTOMETRIC DETERMINATION OF AN EQUILIBRIUM CONSTANT Lab report

Name: ID #: TA name: Lab Section: Lab Date:

Looking at Table 5 (data sheet), how is K_c related to the initial concentrations ([Fes*]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 Kc is actually independent of the initial concentrations of the reactants since they are relatively consistent among the five test tubes. Furthermore, the Kc values are much more relative as opposed to the significant changes in Kc 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 Kc 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 Kc since equilibrium will shift to the right of the reaction to increase the concentration of [FeSCN²⁺] and balance out.

$$Kc = \frac{[FeSCN^{2+}]eq}{[Fe^{3+}] * [SGN^{-}]}$$

To obtain [FeSCN2+]eq in tube 5, you made the assumption that 100% of the ions SCN- had
reacted. Now that you know the value for Ke (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)

	[Fe-3]	+ [SCN-]	
Initial	0.180 M	0.000199 M	0 M
Change	-x	-x	+x
Equilibrium	(0.180 - x) M	(0.000199 - x) M	× M

Equilibrium Constant Lab Report

RC Schank

Equilibrium Constant Lab Report:

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