

Chemical Equilibrium: Finding a Constant, K_c

The purpose of this lab is to experimentally determine the equilibrium constant, K_c , for the following chemical reaction:



When Fe^{3+} and SCN^{-} are combined, equilibrium is established between these two ions and the FeSCN^{2+} ion. In order to calculate K_c for the reaction, it is necessary to know the concentrations of all ions at equilibrium: $[\text{FeSCN}^{2+}]_{\text{eq}}$, $[\text{SCN}^{-}]_{\text{eq}}$, and $[\text{Fe}^{3+}]_{\text{eq}}$. You will prepare four equilibrium systems containing different concentrations of these three ions. The equilibrium concentrations of the three ions will then be experimentally determined. These values will be substituted into the equilibrium constant expression to see if K_c is indeed constant.

You will use a Colorimeter or a Spectrometer to determine $[\text{FeSCN}^{2+}]_{\text{eq}}$. The FeSCN^{2+} ion produces solutions with a red color. Because the red solutions absorb blue light very well, so Colorimeter users will be instructed to use the 470 nm (blue) LED. Spectrometer users will determine an appropriate wavelength based on the absorbance spectrum of the solution. The light striking the detector is reported as *absorbance* or *percent transmittance*. By comparing the absorbance of each equilibrium system, A_{eq} , to the absorbance of a standard solution, A_{std} , you can determine $[\text{FeSCN}^{2+}]_{\text{eq}}$. The standard solution has a known FeSCN^{2+} concentration.

To prepare the standard solution, a very large concentration of Fe^{3+} will be added to a small initial concentration of SCN^{-} (hereafter referred to as $[\text{SCN}^{-}]_i$). The $[\text{Fe}^{3+}]$ in the standard solution is 100 times larger than $[\text{Fe}^{3+}]$ in the equilibrium mixtures. According to LeChatelier's principle, this high concentration forces the reaction far to the right, using up nearly 100% of the SCN^{-} ions. According to the balanced equation, for every one mole of SCN^{-} reacted, one mole of FeSCN^{2+} is produced. Thus $[\text{FeSCN}^{2+}]_{\text{std}}$ is assumed to be equal to $[\text{SCN}^{-}]_i$.

Assuming $[\text{FeSCN}^{2+}]$ and absorbance are related directly (Beer's law), the concentration of FeSCN^{2+} for any of the equilibrium systems can be found by:

$$[\text{FeSCN}^{2+}]_{\text{eq}} = x [\text{FeSCN}^{2+}]_{\text{std}}$$

Knowing the $[\text{FeSCN}^{2+}]_{\text{eq}}$ allows you to determine the concentrations of the other two ions at equilibrium. For each mole of FeSCN^{2+} ions produced, one less mole of Fe^{3+} ions will be found in the solution (see the 1:1 ratio of coefficients in the equation on the previous page). The $[\text{Fe}^{3+}]$ can be determined by:

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

Because one mole of SCN^{-} is used up for each mole of FeSCN^{2+} ions produced, $[\text{SCN}^{-}]_{\text{eq}}$ can be determined by:

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

Knowing the values of $[\text{Fe}^{3+}]_{\text{eq}}$, $[\text{SCN}^{-}]_{\text{eq}}$, and $[\text{FeSCN}^{2+}]_{\text{eq}}$, you can now calculate the value of K_c , the equilibrium constant.

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