The Equilibrium Constant of FeSCN

# Introduction

For the reaction αA + βB <==> γC + δD, equilibrium law states

where Keq is the equilibrium constant for the reaction. The value of the equilibrium constant may be calculated from experimental data if the concentrations of both the reactants and products at equilibrium are known. The general experimental method is to combine different amounts of reactants with known concentrations and then allow the system to come to equilibrium. Although a single trial is all that is necessary to calculate Keq, a more accurate method is to prepare a number of experimental trials with different initial reactant concentrations and report the average value of Keq.

Since the concentrations of the products are initially zero and since the moles of reactants are known from volume and concentration data, it is possible to calculate the final concentrations of both products and reactants from the coefficients of the balanced reaction as long as the final concentration of at least one of the species at equilibrium is known. In this experiment, a spectrophotometer will be used to experimentally measure the concentration of one of the product at equilibrium.

The reaction studied in this experiment is formation of the iron(III) thiocyanate complex from the iron(III) and thiocyanate ions:

Fe+3 + SCN‑ ⬄ Fe(SCN)+2

with the corresponding equilibrium expression:

The thiocyanate ion is colorless and the iron cation is slightly yellowish, but the iron thiocyanate complex is a deep red‑orange. Most colored species will follow the Beer-Lambert law:

A = εbc

where A is the absorbance measured by a spectrophotometer, ε is the absorptivity at a given wavelength of light, b is the path length and c is the concentration of the absorbing species. Since the wavelength and path length are held constant, the relationship between absorbance and concentration should be linear.

From the absorbance of the iron thiocyanate complex, the concentration at equilibrium may be determined from a Beer's law plot. The definition of molarity and reaction stoichiometry may then be used to determine concentrations of the iron ion and the thiocyanate ion, and from these three concentrations the equilibrium constant may be calculated.

# Materials and Methods

Spectrophotometer and cuvettes

0.005 M Iron(III) nitrate, Fe(NO3)3

0.002 M Sodium thiocyanate, NaSCN

0.1 M Nitric acid, HNO3

A 5.0 mL and a graduated 5.0 pipet

Five 50-ml beakers

# Procedure

1. Pipet 5 ml of 0.005 M Fe(NO3)3 into five clean 50 mL beakers.
2. Pipet 1.00, 2.00, 3.00, 4.00, and 5.00 ml of 0.002 M NaSCN, respectively into beakers 1-5.
3. Pipet 4.00, 3.00, 2.00, 1.00 and no 0.1 M HNO3, respectively, into each beaker so that all solutions have a total volume of 10.0 ml.
4. Measure the absorbance of each solution at 447 nm.

# Data and Results

1. Determine the Fe[SCN]2+ concentration from the relationship:

A = 7000 [Fe(SCN)2+]

1. Determine the initial concentrations of the iron and thiocyanate ions for each trial before reaction by multiplying their concentrations by their volumes (in liters).
2. Determine the moles of the FeSCN complex formed for each trial by multiplying the concentration of the complex by the total volume (in liters).
3. Determine the equilibrium moles of the two reactants in each trial by subtracting the moles of complex from initial moles of each.
4. Determine the equilibrium concentrations of the two reactants in each trial by dividing the equilibrium moles by the total volume of solution (in liters).
5. Calculate the equilibrium constant for each trial and report each value, as well as the average value of Keq.