

■ Determination of an Equilibrium Constant

Step 2

Enter the equilibrium value of $[\text{FeNCS}^{2+}]$ from Table 1.1. Use Equation (12) to find the number of moles of FeNCS^{2+} in each of the mixtures. Enter your results in the fifth column of Table 1.2. (Note that this is also the number of moles of Fe^{3+} and SCN^- consumed in the reaction.)

Step 3

Calculate the number of moles of Fe^{3+} and SCN^- that remain at equilibrium for each of the mixtures. Use Equation (13) and enter the results in columns 3 and 4 in Table 1.2.

Step 4

Calculate the concentrations of all of the species at equilibrium using Equation (14) and the results of Step 3. The volume of the mixture is 10.00 mL (0.0100 L) for all mixtures. Enter your results into the appropriate columns in Table 1.2.

Step 5

Calculate the equilibrium constant, K , for the reaction by substituting the values obtained for $[\text{Fe}^{3+}]$, $[\text{SCN}^-]$, and $[\text{FeNCS}^{2+}]$ into Equation (10). Calculate the K_{avg} from the three values that are in closest agreement.

Table 1.2

Equilibrium Mixture	Initial Moles ($\times 10^{-3}$)		Equil. Number of Moles ($\times 10^{-3}$)			Equilibrium Concentrations ($\times 10^{-3}$)			K
	Fe^{3+}	SCN^-	Fe^{3+}	SCN^-	FeNCS^{2+}	Fe^{3+}	SCN^-	FeNCS^{2+}	
1	20.0	20.0	0.6	0.6	0.6				
2	20.0	20.0	0.8	0.8	0.8				
3	20.0	20.0	1.6	1.6	1.6				
4	20.0	20.0	0.2	0.2	0.2				
5									
						K_{avg}			

3 4 5

Lab 11 Rate And Equilibrium Report Answer

B Lingard



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