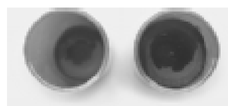
`DATE: NAME: CLASS:

CHAPTER 16		BLM 16.3.3
HANDOUT		

Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant

The colour intensity of a solution is related to the type of ions present, their concentration, and the depth of the solution (the linear measure of the solution through which you are looking). By adjusting the depth of a solution with an unknown concentration until it has the same intensity as a solution with known concentration, you can determine the concentration of the unknown solution. For example, if the concentration of a solution is lower than the standard, the depth of the solution has to be greater in order to have the same colour intensity. For this reason, the ratio of the concentrations of two solutions with the same colour intensity is in inverse ratio to their depths.



Which solution is the least concentrated? Why is the colour intensity the same when you look vertically through the solutions?

In this investigation, you will examine the homogeneous equilibrium between iron(III) ions and thiocyanate ions, and iron(III) thiocyanate ions, $Fe(SCN)^{2+}(aq)$:

$$Fe^{3+}(aq) + SCN^{-}(aq) \leftrightarrow Fe(SCN)^{2+}(aq)$$

You will prepare four equilibrium mixtures with different initial concentrations of $Fe^{3+}(aq)$ and $SCN^{-}(aq)$. You will calculate the initial concentrations of these reacting ions from the volumes and concentrations of the stock solution used and the *total* volumes of the equilibrium mixtures. Then you will determine the concentration of $Fe(SCN)^{2+}(aq)$ ions in each mixture by comparing the colour intensity of the mixture with the colour intensity of a solution with known concentration. After you find the concentration of $Fe(SCN)^{2+}(aq)$ ions, you will use it to calculate the concentrations of the other two ions at equilibrium. You will substitute the three concentrations for each mixture into the equilibrium expression to determine the equilibrium constant.

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Your teacher might choose to do this investigation as a demonstration. Alternatively, your teacher might have stations set up with solutions already prepared and you will make observations and record the data.

Question

What is the value of the equilibrium constant at room temperature for the following reaction?

$$Fe^{3+}(aq) + SCN^{-}(aq) \leftrightarrow Fe(SCN)^{2+}(aq)$$

Prediction

Write the equilibrium expression for this reaction.

Safety Precautions

- The $Fe(NO_3)_3(aq)$ solution is acidified with nitric acid. It should be handled with care. Nitric acid is corrosive and will damage eyes, skin, and mucous membranes. Immediately wash any spills on your skin or clothing with plenty of water and inform your teacher. Also inform your teacher if you spill any solutions on the lab bench or floor.
- All glassware used in this experiment must be clean and dry before using it.
- Wash your hands when you have completed the investigation.

Materials

- 30 mL 0.0020 mol/L KSCN (aq)
- 30 mL 0.0020 mol/L $Fe(NO_3)_3(aq)$ (acidified)
- 25 mL 0.200 mol/L $Fe(NO_3)_3(aq)$ (acidified)
- distilled water
- 5 test tubes (18 mm × 150 mm)
- 5 flat-bottom vials
- 3 beakers (100 mL)
- test-tube rack
- labels or grease pencil
- 1 pipette (20.0 mL)
- 3 pipettes (5.0 mL)
- pipette bulb
- stirring rod
- paper towel
- thermometer (alcohol or digital)
- strip of paper
- diffuse light source, such as a light box (used by doctors to look at x-rays)
- medicine dropper

Procedure

1. Give titles to the following tables. You will use them to record your measurements and calculations.

Test tube #	$Fe(NO_3)_3(aq)$ (mL)	$H_2O(\ell)$ (mL)	KSCN(aq) (mL)	Initial $[SCN^{-}(aq)]$ (mol/L)
2	5.0	3.0	2.0	
3	5.0	2.0	3.0	
4	5.0	1.0	4.0	
5	5.0	0	5.0	

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Vial #	Depth of solution in vial (mm)	Depth of standard solution (mm)	$\frac{\text{Depth of standard solution}}{\text{Depth of solution in vial}}$
2			
3			
4			
5			

2. Label five test tubes and five vials with the numbers 1 through 5. Label three beakers with the names and concentrations of the stock solutions: 2.00×10^{-3} mol/L KSCN(aq), 2.00×10^{-3} mol/L $Fe(NO_3)_3$(aq), and 0.200 mol/L $Fe(NO_3)_3$(aq). Pour about 30 mL of each stock solution into its labelled beaker. Be sure to distinguish between the different concentrations of the iron(III) nitrate solutions. Make sure you use the correct solution when needed in the investigation. Measure the volume of each solution as carefully as possible to ensure the accuracy of your results.

3. Prepare the standard solution of $Fe(SCN)^{2+}$(aq) in Test Tube 1. Use the 20 mL pipette to transfer 18.0 mL of 0.200 mol/L $Fe(NO_3)_3$(aq) into the test tube. Then use a 5 mL pipette to add 2.0 mL of 2.00×10^{-3} mol/L KSCN(aq). The large excess of Fe^{3+}(aq) is to help ensure that essentially all of the SCN^- (aq) will react to form $Fe(SCN)^{2+}$(aq).

4. Use the pipette to transfer 5.0 mL of 2.0×10^{-3} mol/L $Fe(NO_3)_3$(aq) into each of the other four test tubes (labelled 2 to 5).

5. Use the pipette to transfer 3.0, 2.0, 1.0, and 0 mL of distilled water into Test Tubes 2, 3, 4, and 5, respectively.

6. Use the pipette to transfer 2.0, 3.0, 4.0, and 5.0 mL of 2.0×10^{-3} mol/L KSCN(aq) into Test Tubes 2, 3, 4, and 5, respectively. Each of these test tubes should now contain 10.0 mL of solution. Notice that the first table you prepared (in Step 1) shows the volumes of the liquids you added to the test tubes. Use a stirring rod to mix each solution, being careful to rinse the rod with water and then dry it with paper towel before stirring the next solution. Measure and record the temperature of one of the solutions. Assume that all the solutions are at the same temperature.

7. Pour about 5 mL of the standard solution from Test Tube 1 into Vial 1.

8. Pour some of the solution from Test Tube 2 into Vial 2. Look down through Vials 1 and 2. Add enough solution to Vial 2 to make its colour intensity appear about the same as the colour intensity in Vial 1. Use a sheet of white paper as background to make your rough colour intensity comparison.

9. Wrap a sheet of paper around Vials 1 and 2 to prevent light from entering the sides of the solutions. Looking down through the vials over a diffuse light source, adjust the volume of the

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standard solution in Vial 1 until the colour intensity in the vials is the same. Use a medicine dropper to remove or add standard solution. Be careful not to add standard solution to Vial 2.

10. When the colour intensity is the same in both vials, measure and record the depth of solution in each vial as carefully as possible.

11. Repeat Procedure Steps 9 and 10 using Vials 3, 4, and 5.

12. Discard the solutions into the container supplied by your teacher. Rinse the test tubes and vials with distilled water, then return all the equipment. Remember to wash your hands when you have finished.

13. Summarize the results of your calculations in Table 1 on the following page.

(a) Calculate the equilibrium concentration of $Fe(SCN)^{2+}(aq)$ in the standard solution you prepared in Test Tube 1. The $[Fe(SCN)^{2+}(aq)]_{standard}$ is essentially the same as the starting concentration of $SCN^{-}(aq)$ in Tube 1. The large excess of $Fe^{3+}(aq)$ ensured that the reaction of $SCN^{-}(aq)$ was almost complete. However, remember to include the volume of $Fe(NO_3)_3(aq)$ in the total volume of the solution for your calculation.

(b) Calculate the initial concentration of $Fe^{3+}(aq)$ in Test Tubes 2 to 5. $[Fe^{3+}(aq)]_i$ is the same in these four test tubes. They all contained the same volume of $Fe(NO_3)_3(aq)$, and the total final volume was the same. Remember to use the total volume of the solution in your calculation.

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(c) Calculate the initial concentration of $SCN^-(aq)$ in Tubes 2 to 5. $[SCN^-(aq)]_i$ is different in each test tube.

(d) Calculate the equilibrium concentration of $Fe(SCN)^{2+}(aq)$ in Test Tubes 2 to 5. Use the following equation:

$$[FeSCN^{2+}(aq)]_{eq} = \frac{\text{Depth of standard solution}}{\text{Depth of solution in vial}} \times [FeSCN^{2+}(aq)]_{standard}$$

Table 1

Test tube #	Initial concentration (mol/L)		Equilibrium concentration (mol/L)			Equilibrium constant, K_c
	$[Fe^{3+}(aq)]_i$	$[SCN^-(aq)]_i$	$[Fe^{3+}(aq)]_{eq}$	$[SCN^-(aq)]_{eq}$	$[Fe(SCN)^{2+}(aq)]_{eq}$	
1						
2						
3						
4						
5						

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(e) Based on the stoichiometry of the reaction, each mole of $Fe(SCN)^{2+}(aq)$ is formed by the reaction of one mole of $Fe^{3+}(aq)$ with one mole of $SCN^{-}(aq)$. Thus, you can find the equilibrium concentrations of these ions by using the equations below:

$[Fe^{3+}(aq)]_{eq} = [Fe^{3+}(aq)]_{i} - [Fe(SCN)^{2+}(aq)]_{eq}$

$[SCN^{-}(aq)]_{eq} = [SCN^{-}(aq)]_{i} - [Fe(SCN)^{2+}(aq)]_{eq}$

(f) Calculate four values for the equilibrium constant, K_c, by substituting the equilibrium concentrations into the equilibrium expression. Find the average of your four values for K_c.

Analysis

1. How did the colour intensity of the solutions in Test Tubes 2 to 5 vary at equilibrium? Explain your observation.

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2. How consistent are the four values you calculated for K_c? Suggest reasons that could account for any differences.

3. Suppose the following reaction were the equilibrium reaction:

$$Fe^{3+}(aq) + 2SCN^{-}(aq) \leftrightarrow Fe(SCN)^{2+}(aq)$$

(a) Would the equilibrium concentration of the product be different from the concentration of the product in the actual reaction? Explain.

(b) Would the value of K_c be different from the value you calculated earlier? Explain.

Conclusion

4. Write a short conclusion, summarizing your results for this investigation.