### AP Lab #7 – Finding the Constant - Kc

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

**Fe3+(aq) + SCN–(aq) FeSCN2+(aq)**

When Fe3+ and SCN- are combined, equilibrium is established between these two ions and the FeSCN2+ ion. In order to calculate *Kc* for the reaction, it is necessary to know the concentrations of all ions at equilibrium: [FeSCN2+], [SCN–], and [Fe3+]. 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 *Kc* is indeed constant. You will use a spectrophotometer to determine [FeSCN2+]. The FeSCN2+ ion produces solutions with a red color. Because the red solutions absorb blue light very well, so spectrophotometer users will be instructed to set the wavelength reading to 447 nm (blue). This is the 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 to the absorbance of a *standard* solution, you can determine [FeSCN2+]. The standard solution has a known FeSCN2+concentration.

To prepare the standard solution, a very large concentration of Fe3+ will be added to a small initial concentration of SCN– (hereafter referred to as [SCN–]i. The [Fe3+] in the standard solution is 100 times larger than [Fe3+] in the equilibrium mixtures. According to Le Chatelier’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 FeSCN2+ is produced. Thus [FeSCN2+] is assumed to be equal to [SCN–]i. Assuming [FeSCN2+] and absorbance are related directly (Beer’s law).

After you have done your first set of experiments to determine the molar absorptivity of the FeSCN2+ ion, you will do a second set of experiments where an equilibrium exists between [Fe3+], [SCN-] and the [FeSCN2+]. Using the ICE box method, you will determine the equilibrium concentrations of all of the species in each of the reactions. Then you will determine the *Kc* for each experiment (as discussed in the first paragraph above).

# *Science Practices: [2.E, 3.A, 3.C, 4.D, 5.A, 5.C, 5.D, 6.D]*

# Objective

How can an equilibrium constant, *Kc,* be determined from absorbance data of a chemical species?

**Safety**

Wear protective glasses at all times.

Be careful with the equipment being used, it is extremely delicate.

0.5 M HNO3 and should be handled with care.

**Materials**

0.10 M KSCN 0.100 M Fe(NO3)3 0.250M HNO3

0.005 M Fe(NO3)3 disposable pipets four 10-mL volumetric flasks

 Spectrophotometer cuvettes labels

10-mL graduated cylinder 50 mL beakers Kimwipes,

two 2-mL graduated pipette pipette pumps

**Procedure**

**Preparation of the 0.00100 M HSCN solution *(done by teacher)***

1. Pipette 5.00 mL of 0.10 M KSCN solution into a clean 500 mL volumetric flask. Dilute

 to the mark with 0.250 M HNO3. Mix well. Transfer the solution to a bottle.

 Label the bottle **0.00100 M HSCN.**

**Preparation of the Beer’s Law Calibration Plot for the FeSCN+2 Solutions**

2. Label four clean 10.00 mL volumetric flasks #1-4. Using pipettes, carefully prepare four

 standard solutions as described in **Table I** below.

3. Transfer 4.00 mL of the **0.100M Fe(NO3)3** solution to each of the four 10 mL volumetric

 flask. Use a 10-mL graduated cylinder for this process.

4. Add the appropriate volume of the **0.00100 M HSCN** solution directly to the 10 mL flask as

 described in the table below. Use a graduated pipette for this process

5. Dilute each flask to the mark with **0.250 M HNO3**. Mix the solutions thoroughly.

**TABLE I: Standard Solutions of FeSCN+2**

**Solution #** **0.100M Fe(NO3)3 0.00100 M HSCN 0.250 M HNO3**

 1 4.00 mL 0.40 mL Dilute

 2 4.00 mL 0.80 mL to

 3 4.00 mL 1.00 mL the

 4 4.00 mL 2.00 mL mark

**Using the Spectrophotometer to make the Beer’s Law Plot**

6. Turn on the Spectrophotometer and let it warm up for 15 minutes.

7. Set the wavelength to **447 nm** using the *Wavelength Dial*.

8. Set the *Filter Position Lever* to **340-599 nm**.

9. Using the *0%T Dial*, set the % Transmittance to read **0.0**.

10. Fill a cuvette with **0.250 M HNO3**. You will use HNO3 as your blank instead of distilled

 water. Clean the outside of the cuvette with a *Kimwipe* to remove oil from fingerprints left

 on the cuvette. Insert the cuvette **correctly** into the cuvette slot.

11. Use the *100%T / 0A Dial* to set the **Transmittance** to read 100.0. Press the *Mode Button* on

 the Spectrophotometer to change it to read **Absorbance**. It should read .000. Record this

 number.

12. Remove the cuvette and insert another cuvette with one of your **FeSCN+2** samples. Make

 sure to clean the outside of the cuvette with a *Kimwipe* first. Record the Absorbance of your solution.

13. Repeat this process with the rest of your solutions. Record the Absorbance of each solution.

**Determination of FeSCN+2 at Equilibrium**

14. Clean and label four 10.00 mL volumetric flasks #1-4. Using pipets, carefully prepare

 four new solutions as described in **Table II** below.

15. Transfer 2.00 mL of the **0.0050 M Fe(NO3)3** solution to each of the four 10 mL volumetric

 flask. Use a graduated pipette for this process.

16. Add the appropriate volume of the **0.00100 M HSCN** solution directly to the 10 mL flask as

 described in the table below. Use a10-mL graduated cylinder for this process.

17. Dilute each flask to the mark with **0.250 M HNO3**. Mix the solutions thoroughly.

18. Find the absorbances of these four new solutions just as you did for Table I. Simply repeat

steps 10 -13 above, except use your new solutions and record the absorbances).

**TABLE II: Equilibrium Solutions of FeSCN+2**

**Solution #** **0.0050M Fe(NO3)3 0.00100 M HSCN 0.250 M HNO3**

 1 2.00 mL 2.00 mL Dilute

 2 2.00 mL 4.00 mL to

 3 2.00 mL 6.00 mL the

 4 2.00 mL 7.00 mL mark

**Data**

Set up two data tables like the ones seen in the procedure above, but include the absorbance values. (Note each table requires its own title).

**Calculations - from Table I**

1. Calculate the Molarity of each of the standard HSCN solutions you made in Table I.

Use M1V1=M2V2 to calculate the new molarities of the HSCN.

2. Since the initial concentration of HSCN is stoichiometrically 1:1 with the concentration of the

 FeSCN+2 that was produced, write down the FeSCN+2 concentrations the same as calculation #1.

3. Plot the absorbance versus [FeSCN+2]. Calculate the slope of the linear regression line (the

molar absorptivity).

**Calculations - from Table II**

4. Using the absorbance values from Table II and your Beer’s Law plot, determine the

EQUILIBRIUM concentration of FeSCN+2. NOTE: You can do this by using the slope and the y-intercept values from your plot.

5. Calculate the INITIAL concentrations of Fe3+ (from Table II) using M1V1=M2V2 to calculate the new molarities of the Fe+3 in each solution.

6. Calculate the INITIAL concentrations of HSCN (from Table II) using M1V1=M2V2 to calculate the new molarities of the HSCN in each solution.

7. Use the ICE box table provided to calculate the EQUILIBRIUM concentrations of Fe+3 AND

HSCN in each solution. NOTE: The [H+] (or HNO3) is 0.250 M throughout the experiment. Any additional changes in[H+] concentration would be insignificant, so keep the [H+] at 0.250 M.

8. Calculate *Kc* for each experiment and determine if the results are similar.

9. Calculate the average *Kc* value for this reaction.

**Equilibrium Constant Value Calculations**

Solution # Fe+3 + HSCN FeSCN+2 + H+

1 Initial M 0 0.250

 Change M

 Equilibrium M 0.250

2 Initial M 0 0.250

 Change M

 Equilibrium M 0.250

3 Initial M 0 0.250

 Change M

 Equilibrium M 0.250

4 Initial M 0 0.250

 Change M

 Equilibrium M 0.250

**Discussion**

* Answer the objective. Be clear and specific.
* What was the *Kc* for this experiment? Was it reproducible?
* Explain why your *Kc* was either a large or a small number.

 **Additional Discussion Questions**

1. Write the equilibrium constant expression, *Kc,* for the reaction done in this lab.

2. What does the value of than equilibrium constant tell you?

3. Explain why FeSCN+2 was yellowish orange in color.

4. Why was the wavelength of the spectrophotometer set to 447 nm?

5. What would have happened to the equilibrium if after the reaction reached equilibrium, an

 additional amount of 1.0-molar acid was added to the system? Would the FeSCN+2 concentration increase or decrease in this situation? Explain.

6. What would have happened to the equilibrium constant if after the reaction reached

 equilibrium, an additional amount of 1.0-molar acid was added to the system? Explain.

7. It has been suggested that an alternate reaction might occur in this experiment. What would

 the equilibrium expression look like if the reaction below had occurred instead of the one mentioned in the introduction?

 Fe+3 + 2HSCN Fe(SCN)2+1 + 2H+

8. What changes would you have to make in calculating the equilibrium concentrations of the reactants and products in your ICE box if the reaction above had occurred?