Prelab #9 Week 2—Equilibrium Name ______________________

Lab day: M T W Th F

From Week 1:

1. Attach Beer's Law Standard Curve
2. Determine the extinction coefficient, ε, of $FeSCN^{+2}$ at 455 nm. *Include the correct units for* ε. Absorption has no units.

Equation from Beer's Law plot	
Extinction coefficient	

Experimental plan for Week 2: (attach calculations)

Our goal in Week 2 is to determine the equilibrium constant for the reaction:

$$Fe^{3+}_{(aq)} + SCN^{-}_{(aq)} \rightleftharpoons FeSCN^{2+}_{(aq)}$$

In order to do this, we will be making three solutions that are at equilibrium. If we know the initial concentrations of reactants and the equilibrium concentration of product, we can calculate the equilibrium constant. Unlike Week 1, we will be using similar concentrations of both reactants. We will know the initial concentrations of reactants based on the volumes of stock solution that we use to make our solutions. We will be measuring the concentration of $FeSCN^{+2}$ in solution based on the extinction coefficient you determined in Week 1. In order to do this, we need to produce measurable amounts of $FeSCN^{+2}$. Measurable means that the absorbance will be between 0.1 and 1 on the scale of the Spec 20.

This Prelab will walk you through the steps you need to take in order to design your three solutions for Week 2. First, we will use our Beer's Law relationship to help us decide how much product we need to form to have a measurable absorbance. Then we can use the equilibrium constant expression to decide how much reactants we need.

1. First examine your Beer's law plot. What is the concentration of $FeSCN^{2+}$ that gives a midpoint absorbance value of 0.500? ____________________ . This is a good target absorbance for the equilibrium reaction you will set up and will serve as the equilibrium concentration of our product.
2. To simplify the calculation, assume that K is approximately 100, $[FeSCN^{2+}]_{eq}$ is the concentration determined in the previous question, and that $[Fe^{3+}]_{eq}$ and $[SCN^{-}]_{eq}$ are equal ($x = [Fe^{3+}]_{eq} = [SCN^{-}]_{eq}$). The subscript "eq" means equilibrium concentration. We want to know "x" because it is the concentration of reactants that will give us a measurable amount of product. Then we can make solutions with $[Fe^{3+}]_{eq} = [SCN^{-}]_{eq} =$

x, measure $[FeSCN^{2+}]_{eq}$, and calculate the actual value of K (we're just assuming K $\approx$ 100 in order to design the experiment). To solve for "x", use the equilibrium expression:

$$K = \frac{\left[FeSCN^{2+}\right]_{eq}}{\left[Fe^{3+}\right]_{eq}\left[SCN^{-}\right]_{eq}} = \frac{\left[FeSCN^{2+}\right]_{eq}}{x^2} \approx 100$$

That will give us our target equilibrium concentrations of the reactants: x $\approx$ $[Fe^{3+}]_{eq}$ $\approx$ $[SCN^{-}]_{eq}$ $\approx$ _________________. It turns out that it will work fine if we make solutions with x $\approx$ $[Fe^{3+}]_o$ $\approx$ $[SCN^{-}]_o$, where the subscript "o" signifies initial concentrations.

3. You will be given one stock solution with $[Fe^{3+}]$ = 0.2M and one stock solution with $[SCN^{-}]$ = 0.002M. How much of each solution will you need to put into a 50.00 mL flask in order to make a solution with x = $[Fe^{3+}]_o$ = $[SCN^{-}]_o$, the concentration you determined above?

   mL of 0.2M $[Fe^{3+}]$: _________________

   mL of 0.002M $[SCN^{-}]$: _________________

   Remember from last week that we need our solution to be 0.5M HNO_3, so both of our stock solutions also have 0.5M HNO_3 and you will be diluting with 0.5M HNO_3.

4. You should have found that you need an amount of Fe^{3+} stock solution that is too small to measure. Therefore, you will have to dilute your Fe^{3+} stock solution before using it to make solutions. Decide what factor you want to dilution your stock solution by, and make sure that the glassware you will need is available in the Materials list in the lab manual.
   a) What factor will you dilute the stock iron solution by? _________
   b) What will be the concentration of your new iron stock solution? _________
   c) What glassware (exact sizes) will you use for diluting the stock solution?

      ____

   d) What solution will you use to dilute your iron solution? ___________________

5. Now that you have a new, diluted stock solution, go back and determine how to make your target solution. How much of each solution will you need to put into a 50.00 mL flask in order to make a solution with x = $[Fe^{3+}]_o$ = $[SCN^{-}]_o$, the concentration you determined above?

   mL of new, diluted $[Fe^{3+}]$ stock: _________________

   mL of 0.002M $[SCN^{-}]$: _________________ (same as before)

6. You are now ready to design your experiment. You want to make three different solutions that have reactant concentrations similar to the target concentrations you determined above. You probably cannot measure exactly the volumes you determined in the previous question. However, those amounts are only approximate. Using glassware available (refer to the lab manual), choose volumes that are within a factor of two of the amounts determined above to make your three solutions. Fill in the table below for the three solutions that you will make:

	Stock solution: _____M Fe^{3+} in 0.5 HNO_3 (from question 4.b)		**Stock solution: 0.002M SCN^- in 0.5 HNO_3**		**0.5 M HNO_3****			
	Fe^{3+} Vol. (mL)	**Fe^{3+} Conc.* (M)**	**SCN^- Vol. (mL)**	**SCN^- Conc.* (M)**	**HNO_3 Vol. (mL)**	**TOTAL Vol. (mL)**	**Molar Ratio $[Fe^{3+}]$:$[SCN^-]$**	**Glassware used**
1						50.00		
2						50.00		
3						50.00		

* This is the concentration immediately after the reactants are mixed but before any reaction occurs to bring the system to equilibrium

** The concentration of 0.5M HNO_3 does not change since all the other solutions also contain 0.5M HNO_3 as the diluent.