EXPERIMENT

Equilibrium constant

Goal

The goal of this laboratory is to determine *the equilibrium constant* of a chemical reaction by using the *Lambert-Beer's law* to measure the concentration of the species in solution.

Background

Most chemical reactions do not proceed to completion, proceeding only to the point where both reactants and products have constant concentration. This is because most reactions are reversible and they can run in both forward and reverse directions. The concentrations of the species involved in the equilibrium will achieve an *equilibrium state* when the reaction rates in both directions equalize. For a certain chemical reaction:

$$\mathrm{aA} \quad + \quad \mathrm{bB} \quad \rightleftharpoons \quad \mathrm{cC} \quad + \quad \mathrm{dD}$$

The equilibrium can be characterized by the expression:

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

where K_c (capital letter) is the equilibrium constant in terms of molarity, the brackets represent the concentrations of the respective species and the exponents are their corresponding stoichiometric coefficients.

Spectrophotometry

Spectrophotometry is a technique that measures the amount of light absorbed by a chemical substance, typically in solution, using the use of a *spectrophotometer*. Spectrophotometers display the absorbance, the amount of light absorbed by a sample. To properly measure the absorbance of a chemical in a water-based solution you need to set the blank by using a sample containing only the solvent without a solute. The blank establishes the baseline of the measurement, eliminating any absorbance caused by the solvent.

The spectrometer allows the measurement of absorbance at a single wavelength. Interestingly, for each substance, there is a wavelength of *maximum absorbance*, which allows us to differentiate among chemical substances with different maximum absorbances. For example, in the current experiment, you will analyze mixtures containing $Fe^{3+}_{(aq)}$, $SCN^{-}_{(aq)}$, and $Fe(SCN)^{2+}_{(aq)}$. However, only $Fe(SCN)^{2+}_{(aq)}$ will absorb radiation at λ=450 nm, and hence we can track the absorbance of this chemical only in the mixture.

Lambert-Beer's law

Liquids *attenuated* the intensity of light passing through a colored solution. The effect can be comparable to a dirty window that attenuates the light passing through. The dirtier the window (higher concentration of dirt), the less one can see through it (because more light is absorbed, reflected, or refracted by the dirt). Lambert-Beer's law establishes the relationship between absorbance and concentration:

$$A = k_\lambda c$$

In this formula, A stands for the *absorbance* (light absorbed by the liquid), k_λ (lower case, not to be confused with the equilibrium constant, K) is a **proportionality constant**, which is specific to each compound and at the wavelength (λ) specified by the subindex, and c is the solution concentration. Lambert-Beer's law is a linear function of the form $y = mx + b$. When graphing the absorbance (y-axes) versus the concentration (x-axes), the slope of the line corresponds to Lambert-Beer's proportionality constant. This line should cross the origin as there should be no absorbance when the concentration is null.

The experiment

The experiment today is divided into 2 parts: determining Lambert-Beer's constant k and determining the equilibrium constant K.

Part A. Determining Lambert-Beer's constant k.

This experiment studies the reaction between iron(III) and thiocyanate to give an iron-thiocyanate complex:

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

Interestingly, Fe^{3+} in solution is yellow and becomes red when reacting with thiocyanate to produce the $(FeSCN)^{2+}$ complex. This intense red coloration allows us to use spectrophotometry to measure the $(FeSCN)^{2+}$ concentration and hence, of all the species in equilibrium.

The first part of this experiment requires the determination of Lambert-Beer's constant, k_λ. However, here we deal with a chemical reaction. According to the equilibrium, $Fe(SCN)^{2+}$ in solution will follow the reverse reaction forming iron(III) and thiocyanate ions. The shift due to the equilibrium makes the concentration of $Fe(SCN)^{2+}$ unknown.

To solve this problem Fe^{3+} ion will be added in excess (0.0025 M) while adding only small amounts of SCN^{-} (10^{-4} M). This trick will displace the equilibrium to the right: according to Le Châtelier's, the equilibrium shifts to the products when one reactant is added. Hence, if one of the reactants is added in overwhelming excess, the other reactant will be consumed almost to exhaustion forming the product. Therefore, the $FeSCN^{2+}$ concentration can be calculated using the concentration of SCN^{-}.

In this part, a set of solutions will be added to a clean cuvette, and absorbance at λ=450 nm will be measured using the spectrophotometer. With the 5 points obtained and the origin of the coordinates, a graph will be plotted to calculate the slope of the line. Notice again that in this linear representation, the line must pass through the origin by definition, since at concentration=0, absorbance must be 0.

Parts B. Determining the equilibrium constant K_c.

Using the absorbance measured and the value of k_{450nm}, in this second part of the experiment you will calculate the concentration in equilibrium of all species involved in the equilibrium. First, the concentration of the product will depend on the absorbance, and will be given by the formula below:

$$A = k_{450nm}[Fe(SCN)^{2+}] \qquad [Fe(SCN)^{2+}] = \frac{A}{k_{450nm}}$$

An ICE chart given below will be used to calculate the equilibrium concentration of reactants, where 3 is calculated from the absorbance, and 2 and 1 are recalculated from the initial concentrations:

	$[Fe(NO_3)_3]$	[KSCN]	$[Fe(SCN)^{2+}]$
Initial	1	2	0 M
Change	–3	–3	+3
Equilibrium	1 – 3	2 – 3	3

Procedure

Preparing a diluted KSCN solution

☐ *Step 1:* – Clean 5 large test tubes and the matching stoppers and set them upside down on a rack. Number the test tubes from 1 to 5.

☐ *Step 2:* – Obtain a 100.0 mL volumetric flask and a Mohr pipette.

☐ *Step 3:* – Obtain about 15 mL of 0.0025 M KSCN in a 50 mL beaker. Rinse the Mohr pipette a couple of times with a small amount of this solution. Always discard the rinsing liquid. Then, transfer 4.00 mL into the volumetric flask.

☐ *Step 4:* – In a clean 100 mL beaker, obtain about 100 mL of distilled water. Carefully add water to the volumetric flask directly from the beaker and stop adding before you reach the 100.0 mL mark. Use a plastic dropper to level the water to the mark, drop by drop. This is the diluted KSCN solution.

☐ *Step 5:* – Rinse the Mohr pipette with the diluted solution of KSCN again.

Good Lab Practice

- Pipettes are always used together with a suction bulb or a syringe. **Never** suck a chemical using the mouth!
- **Never** leave a pipette lying on the table **while connected to the syringe**. Remaining liquid can leak into the syringe. The syringe will be damaged and the liquid will return to the pipette contaminated.
- Always wash pipettes with distilled water and rinse them using the same product to be measured. Suck a small volume, disconnect the syringe and move the liquid about, displacing the distilled water. Discard the rinsing liquid as waste. Rinse twice.

Part A. Determining Lambert-Beer's constant k. Preparing the mixtures

☐ *Step 1:* – Following the Table in the description of Part A, add 1.00 mL of the diluted KSCN solution to jumbo test tube 1, 2.00 mL to test tube 2, 3.00 mL to test tube 3, 4.00 mL to test tube 4 and 5.00 mL to test tube 5. You will use jumbo test tubes throughout the experiment.

☐ *Step 2:* – Rinse the pipet with distilled water. Obtain about 50 mL of 0.25 M $Fe(NO_3)_3$ in a 100 mL beaker. Rinse the pipette with the solution of $Fe(NO_3)_3$ from the beaker.

☐ *Step 3:* – Add 5.0 mL of 0.25 M $Fe(NO_3)_3$ in each test tube.

☐ *Step 4:* – Rinse the pipette with distilled water. Obtain about 40 mL of 0.1 M HNO_3 in a 50 mL beaker. Rinse the pipette with the solution of HNO_3 from the beaker.

☐ *Step 5:* – Following the Table below, add 4.00 mL of the HNO_3 solution to test tube 1; 3.00 mL to test tube 2, 2.00 mL to test tube 3, and 1.00 mL to test tube 4.

☐ *Step 6:* – At this point, all your test tubes should have the same volume of liquid. If not, repeat the ones that diverge.

☐ *Step 7:* – Put the stoppers in the test tubes and mix the solutions.

Part A	$Fe(NO_3)_3$	KSCN	HNO_3	v_{total}
	0.25 M	0.0025 M	0.1 M	
Test Tube 1	5.0 mL	1.0 mL	4.0 mL	10.0 mL
Test Tube 2	5.0 mL	2.0 mL	3.0 mL	10.0 mL
Test Tube 3	5.0 mL	3.0 mL	2.0 mL	10.0 mL
Test Tube 4	5.0 mL	4.0 mL	1.0 mL	10.0 mL
Test Tube 5	5.0 mL	5.0 mL	0.0 mL	10.0 mL

Part A. Determining Lambert-Beer's constant k. Measuring absorbance

☐ *Step 1:* – Start the spectrophotometer. The bulb needs time to heat up to a stable temperature.

☐ *Step 2:* – Get a 250 mL beaker for waste. Refill the 100 mL beaker with distilled water.

- [ ] *Step 3:* – Using the plastic dropper, fill the cuvette 3/4 full with distilled water.
- [ ] *Step 4:* – Insert the cuvette into the spectrophotometer, set the wavelength to 450 nm, and press "blank" or "zero".
- [ ] *Step 5:* – Use a new cuvette. Fill the cuvette 3/4 full with the solution from test tube 1.
- [ ] *Step 6:* – Insert the cuvette into the spectrophotometer to measure the absorbance. Record the result.
- [ ] *Step 7:* – Repeat for test tubes 2, 3, 4, and 5, always using new cuvettes for each measurement.
- [ ] *Step 8:* – Plot absorbance 7 vs. concentration 6 in order to obtain the Lambert-Beer's constant k.
- [ ] *Step 9:* – Measure the absorbance of an unknown solution and use the linear regression (A/k) to obtain its concentration.
- [ ] *Step 10:* – Do not proceed unless you show the instructor your graph.

Good Lab Practice

✍ For best results, use the same cuvette, with the same orientation, for all the measurements.

✍ Keep the outer walls of the cuvette clean. Do not touch them. Wipe them with a piece of optical paper.

✍ Rinse the cuvette at the same time as you rinse the pipette.

Part B. Determining the equilibrium constant K. Preparing a different set of mixtures

- [ ] *Step 1:* – Prepare a new set of mixtures based on the proportions reported in the Table below.
- [ ] *Step 2:* – Make sure you use the 0.0025 M $Fe(NO_3)_3$ solution and not the 0.25M.
- [ ] *Step 3:* – The 0.0025 M KSCN solution to be used has already been prepared at the lab.
- [ ] *Step 4:* – Measure the absorbance for the new set of mixtures.

Part B	$Fe(NO_3)_3$ 0.0025 M	KSCN 0.0025 M	HNO_3 0.1 M	v_{total}
Test Tube 6	1.0 mL	5.0 mL	1.0 mL	7mL
Test Tube 7	1.0 mL	4.5 mL	1.5 mL	7mL
Test Tube 8	1.0 mL	4.0 mL	2.0 mL	7mL
Test Tube 9	1.0 mL	3.5 mL	2.5 mL	7mL
Test Tube 10	1.0 mL	3.0 mL	3.0 mL	7mL
Test Tube 11	2.0 mL	4.0 mL	1.0 mL	7mL
Test Tube 12	2.0 mL	3.5 mL	1.5 mL	7mL
Test Tube 13	2.0 mL	3.0 mL	2.0 mL	7mL
Test Tube 14	2.0 mL	2.5 mL	2.5 mL	7mL
Test Tube 15	2.0 mL	2.0 mL	3.0 mL	7mL

Calculations

1 This is the volume of $Fe(NO_3)_3$ added.

2 This is the volume of KSCN added.

3 This is the total volume of the mixtures.

4 This is the initial Iron(III) concentration in the mixtures:

$$[Fe^{+3}]_0 = \frac{v_{Fe^{+3}} \cdot c_{Fe^{+3}}}{v_{total}} = \frac{1 \cdot 0.25}{3}$$

5 This is the initial thiocyanide concentration in the mixtures:

$$[SCN^-]_0 = \frac{v_{SCN^-} \cdot c_{SCN^-}}{v_{total}} = \frac{2 \cdot 0.0025}{3}$$

6 This is the concentration of $Fe(SCN)^{+2}$ in the mixture:

$$[Fe(SCN)^{+2}] = [SCN^-] = 5$$

7 This is the measured absorbance of the mixture.

8 This step is needed to calculate numerically the Lambert-Beer's constant k:

$$[Fe(SCN)^{+2}] \cdot A = 6 \cdot 7$$

9 This step is needed to calculate numerically the Lambert-Beer's constant k:

$$[Fe(SCN)^{+2}]^2 = 6^2$$

10 This step is needed to calculate numerically Lambert-Beer's constant k, and results from adding all values of 8 .

11 This step is needed to calculate numerically Lambert-Beer's constant k, and results from adding all values of 9 .

12 This step the Lambert-Beer's constant k:

$$k = \frac{\sum [Fe(SCN)^{+2}] \cdot A}{\sum [Fe(SCN)^{+2}]^2} = \frac{10}{11}$$

13 This is the volume of $Fe(NO_3)_3$ added.

14 This is the volume of KSCN added.

15 This is the total volume of the mixtures.

16 This is the initial Iron(III) concentration in the mixtures:

$$[\mathrm{Fe}^{+3}]_0 = \frac{v_{\mathrm{Fe}^{+3}} \cdot c_{\mathrm{Fe}^{+3}}}{v_{total}} = \frac{13 \cdot 0.0025}{15}$$

17 This is the initial thiocyanide concentration in the mixtures:

$$[\mathrm{SCN}^-]_0 = \frac{v_{\mathrm{SCN}^-} \cdot c_{\mathrm{SCN}^-}}{v_{total}} = \frac{14 \cdot 0.0025}{15}$$

18 This is the measured absorbance of the mixture.

19 This is the concentration of $\mathrm{Fe(SCN)}^{+2}$ in equilibrium:

$$[\mathrm{Fe(SCN)}^{+2}]_{eq} = \frac{A}{k} = \frac{18}{12}$$

20 This is the concentration of Fe^{+3} in equilibrium:

$$[\mathrm{Fe}^{+3}]_{eq} = [\mathrm{Fe}^{+3}]_0 - [\mathrm{Fe(SCN)}^{+2}]_{eq} = 16 - 19$$

21 This is the concentration of SCN^- in equilibrium:

$$[\mathrm{SCN}^-]_{eq} = [\mathrm{SCN}^-]_0 - [\mathrm{Fe(SCN)}^{+2}]_{eq} = 17 - 19$$

22 This is the equilibrium constant:

$$K = \frac{[\mathrm{Fe(SCN)}^{+2}]_{eq}}{[\mathrm{Fe}^{+3}]_{eq} \cdot [\mathrm{SCN}^-]_{eq}} = \frac{19}{20 \cdot 21}$$

STUDENT INFO

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Pre-lab Questions

Equilibrium constant

1. You prepare a solution by following the procedure described next. You first obtain about 15 mL of 0.0025 M KSCN in a 50 mL beaker. Then you transfer 4.00 mL into the volumetric flask. In a clean 100 mL beaker, you obtain about 100 mL of distilled water, and carefully add water to the volumetric flask directly from the beaker stopping before reaching the 100.0 mL mark. Finally, you use a plastic dropper to level the water to the mark, drop by drop. Calculate the molarity of the solution you prepared.

2. You mix 1mL of $Fe(NO_3)_3$ 0.25 M with 5mL of KSCN 10^{-4} M and 1mL of HNO_3 0.1 M. Calculate the initial concentration of Fe^{3+} in the mixture.

3. Write down the formula for the equilibrium constant in terms of concentration for the equilibrium below:

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

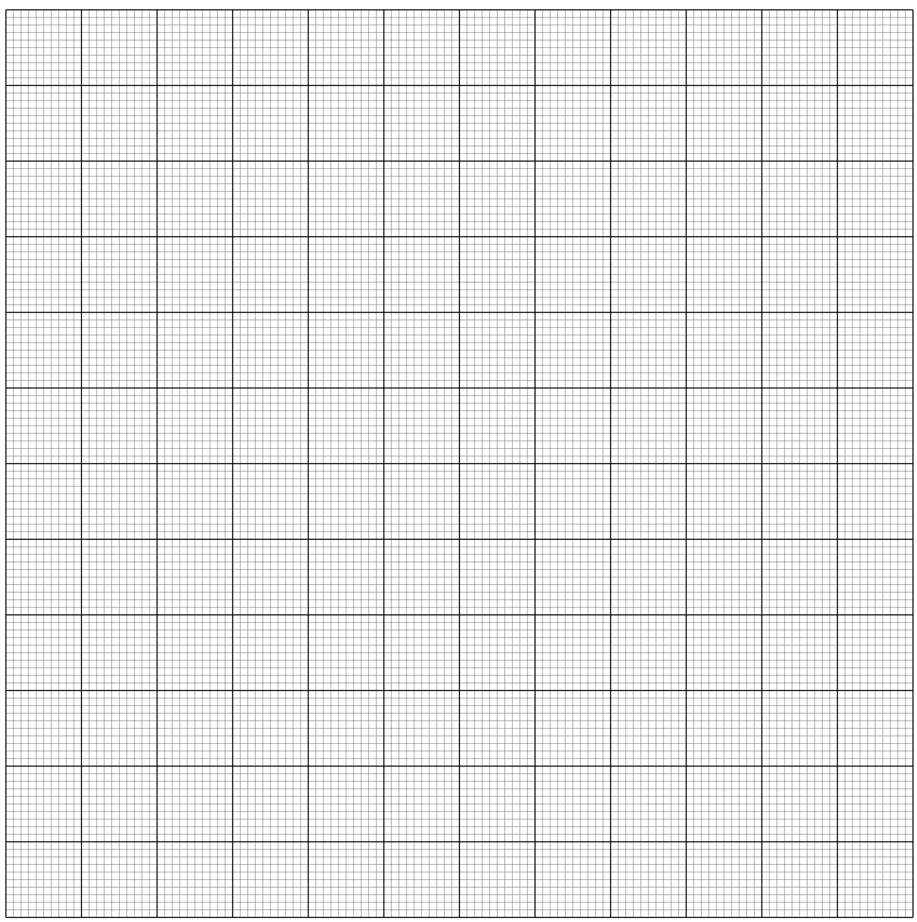
4. A set of absorbances, A, for different concentrations, c, are given below.

(a) Plot A vs. c in the graph below.

(b) Compute the slope of the graph by using the formula:

$$k = \frac{\sum c \cdot A}{\sum c^2}$$

c (M)	A	$c \cdot A$ (M)	c^2 (M^2)
0.0120	0.681		
0.00960	0.540		
0.00720	0.389		
0.00480	0.270		
0.00240	0.133		
Sum			
k (M^{-1})			



(c) After calculating k, now write down the formula for absorbance in the form: $A = k \cdot c$

Results
EXPERIMENT

Equilibrium constant

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Parts A. Determining Lambert-Beer's constant k.

	$Fe(NO_3)_3$ 0.25M (1)	KSCN 0.0025M (2)	HNO_3 0.1M	v_{tot} (3)	$[Fe^{+3}]_0$ (4)	$[SCN^-]_0$ (5)	$[FeSCN^{+2}]$ (6)	A (7)	$[FeSCN^{+2}]\cdot A$ (8)	$[FeSCN^{+2}]^2$ (9)
Test Tube 1	5mL	1mL	4mL	10mL	______	______	______	______	______	______
Test Tube 2	5mL	2mL	3mL	10mL	______	______	______	______	______	______
Test Tube 3	5mL	3mL	2mL	10mL	______	______	______	______	______	______
Test Tube 4	5mL	4mL	1mL	10mL	______	______	______	______	______	______
Test Tube 5	5mL	5mL	0mL	10mL	______	______	______	______	______	______
Origin							0	0	0	0
								Sum=	(10) ______	(11) ______

(12) k = ____________

Unknown							______	______		

Parts B. Determining the equilibrium constant K.

Test Tube	$Fe(NO_3)_3$ 0.0025M (13)	KSCN 0.0025M (14)	HNO_3 0.1M	v_{tot} (15)	$[Fe^{3+}]_0$ (16)	$[SCN^-]_0$ (17)	A (18)	$[Fe(SCN)^{2+}]_{eq}$ (19)	$[Fe^{3+}]_{eq}$ (20)	$[SCN^-]_{eq}$ (21)	K (22)
6	1mL	5.0mL	1.0mL	7.0mL	____	____	____	____	____	____	____
7	1mL	4.5mL	1.5mL	7.0mL	____	____	____	____	____	____	____
8	1mL	4.0mL	2.0mL	7.0mL	____	____	____	____	____	____	____
9	1mL	3.5mL	2.5mL	7.0mL	____	____	____	____	____	____	____
10	1mL	3.0mL	3.0mL	7.0mL	____	____	____	____	____	____	____
11	2mL	4.0mL	1.0mL	7.0mL	____	____	____	____	____	____	____
12	2mL	3.5mL	1.5mL	7.0mL	____	____	____	____	____	____	____
13	2mL	3.0mL	2.0mL	7.0mL	____	____	____	____	____	____	____
14	2mL	2.5mL	2.5mL	7.0mL	____	____	____	____	____	____	____
15	2mL	2.0mL	2.0mL	7.0mL	____	____	____	____	____	____	____

Average K= ____

Figure 1: Column 6 [$Fe(SCN)^{+2}$] (X axis) vs. A Column 7 (Y axis)

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Post-lab Questions

Equilibrium constant

1. The absorbance of a colored substance in a colorless liquid is measured at each of a series of wavelengths, and the data is given below:

λ(nm)	325	350	375	400	425	450	475	500	525
A	0.016	0.144	0.341	0.578	0.681	0.558	0.281	0.092	0.031

(a) Plot A vs. λ in the graph below.

(b) Calculate the λ value that gives a maximum A.

$\lambda(A_{max})$ = ______________________

2. In parts A and B you use different iron and thiocyanate solutions: in part, A you use a 0.25M-$Fe(NO_3)_2$ solution and a 0.0025M-KSCN solution, whereas on part B you use a 0.0025M-$Fe(NO_3)_2$ solution and a 0.0025M-KSCN solution. Explain why you do this.

3. You mix 5mL of $Fe(NO_3)_3$ 0.25 M with 2mL of KSCN 0.0025 M and 3mL of HNO_3 0.1 M. Calculate the initial concentration of SCN^- in the mixture.

4. You mix 5mL of $Fe(NO_3)_3$ 0.25 M with 4mL of KSCN 0.0025 M and 1mL of HNO_3 0.1 M. Calculate the concentration of $FeSCN^{2+}$ in the mixture.

EXPERIMENT

Le Châtelier’s Principle

Goal

The goal of this laboratory is to see the shift in the **equilibrium position** and to connect the observations to **Le Châtelier's Principle**, which states that:

Any change in status quo prompts an opposing reaction in the responding system.

Background

A double arrow in a chemical reaction indicates that the reaction can proceed in both directions. Products can react with each other to generate the original reactant.

$$A + B \rightleftharpoons C + D$$

Since the speed of most reactions depends on the concentration of the combining species, the forward reaction rate will decrease when products are formed, and the reverse reaction rate will increase. The latter will replenish the concentration of reactants which in turn will accelerate the forward reaction. Eventually, forward and backward reaction rates will equal each other and the concentrations of reactants and product will remain constant; at this point, the equilibrium has been reached.

The experiment

The experiment today is divided into the following parts: colored complexed ions ($FeSCN^{2+}$), colored complexed ions ($Ni(NH_3)_6{}^{2+}$), the effect of pH on an indicator, effect of pH on solubility, and heat as a product.

Part A. Colored complexed ions; $FeSCN^{2+}$.

Many metals form colored complex ions with several ligands, such as iron (III) ion Fe^{3+} with thiocyanate SCN:

$$\underset{\text{yellow}}{Fe^{3+}\,(aq)} + SCN^{-}\,(aq) \rightleftharpoons \underset{\text{red}}{FeSCN^{2+}\,(aq)}$$

The Iron (III) solution is yellow, the thiocyanate is transparent and the Iron (III) thiocyanate is red. The final color of the solution will be determined by the concentrations in the final equilibrium position.

Part B. Colored complexed ions; $Ni(NH_3)_6{}^{2+}$

.

In the case of Hexaaminenickel (II), you will also study the change in equilibrium position when adding an acid. The ligand will act as a base reacting with the acid and detaching from the metal.

$$\overset{\text{green}}{Ni^{2+}\,(aq)} + 6\,NH_3\,(aq) \rightleftharpoons \overset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}}$$

$$\Downarrow$$

$$\underset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}} + 6\,HCl\,(aq) \rightleftharpoons \underset{\text{green}}{NiCl_2\,(aq)} + 6\,NH_4Cl\,(aq)$$

Part C. Effect of pH on an indicator.

pH indicators are substances that change color depending on the medium pH. Typically they are weak bases or acids. As such, they dissociate slightly in water forming ions. Those ions might be colored species. The general expression for the dissociation of a divalent weak acid indicator (H2In) is:

$$\underset{\text{yellow}}{H_2In_{(aq)}} \quad + \quad H_2O \rightleftharpoons \quad \underset{\text{red}}{HIn^{-}_{(aq)}} \quad + \quad H_3O^{+}$$

In the case of Methyl orange, the protonated form is red, while the conjugated base is yellow. Notice that hydronium is a product in this equilibrium. How will adding an acid affect this equilibrium?

Part D. Effect of pH on solubility.

Hydroxide salts of Group II elements (Ca, Sr, and Ba) are slightly soluble. Adding acid will neutralize the hydroxides in the solution while adding more hydroxide will increase the ion product. In both cases, the amount of precipitate will be affected.

$$\mathrm{Ca(OH)_2}\,(s) \rightleftharpoons \mathrm{Ca^{2+}}\,(aq) + \mathrm{OH^-}\,(aq)$$
$$\Downarrow$$
$$\mathrm{OH^-}\,(aq) + \mathrm{H^+}\,(aq) \rightleftharpoons \mathrm{H_2O}\,(aq)$$

Part E. Heat as a product

Lastly, you will experience Le Châtelier's Principle in an exothermic reaction, where the heat can be viewed as a product of the reaction.

$$\underset{\text{violet}}{\mathrm{CoCl_4{}^{2-}}\,(aq)} + 6\,\mathrm{H_2O}\,(l) \rightleftharpoons \underset{\text{pink}}{\mathrm{Co(H_2O)_6{}^{2+}}\,(aq)} + 4\,\mathrm{Cl^-}\,(aq) + \textbf{HEAT}$$

Example

The following endothermic reaction is allowed to reach equilibrium:

$$\underset{blue}{A(aq)} + B(aq) + HEAT \rightleftharpoons C(aq) + D(aq)$$

Where A, the only colored compound, is blue. How will the following changes affect the color/equilibrium: (a) Adding more B reactant, (b) Adding more C product, (c) Heating the mixture.

Answer: (a) less blue/equilibrium shifts to the right. (b) more blue/ equilibrium shifts left. (c) Heat is a reactant; less blue/equilibrium shifts to the right.

Procedure

Part A. Colored complexed ions: $FeSCN^{2+}$

- [ ] *Step 1:* – Find 3 test tubes and a 100 mL beaker, clean them, and mark the test tubes with letters A, B, and R.
- [ ] *Step 2:* – Read the Good Lab Practice box

Good Lab Practice

✍ Concentrated solutions of ammonia, sodium hydroxide and hydrochloric acid should be handled with care.

✍ Do not shake open test tubes. Do not use your fingers as stoppers, even if you wear gloves.

- [ ] *Step 3:* – Add about 20 mL of distilled water to the beaker, 20 drops of 0.1 M $Fe(NO_3)_3$ and 20 drops of 0.1 M KSCN. Mixed the solution until the color is homogeneous.
- [ ] *Step 4:* – Use a 10 mL graduated cylinder to add 3 mL of the mixture to each of the test tubes.

- [ ] *Step 5:* – Add 20 drops of 0.1 M $Fe(NO_3)_3$ to test tube A. Put a stopper and mix the solution.
- [ ] *Step 6:* – Add 20 drops of 0.1 M KSCN to test tube B. Put a stopper and mix the solution.
- [ ] *Step 7:* – Add 20 drops of distilled water to test tube R. Put a stopper and mix the solution.
- [ ] *Step 8:* – Compare the color of the test tubes A and B to R and write down your observations.

Part B. Colored complexed ions: $Ni(NH_3)_6^{2+}$.

- [ ] *Step 1:* – Obtain 1 test tube and clean it.
- [ ] *Step 2:* – Add 10 drops of 0.1 M $Ni(NO_3)_2$. Indicate the color on the results page.
- [ ] *Step 3:* – Add drops of 6 M NH_3 until the color changes.
- [ ] *Step 4:* – Add drops of 6 M HCl until the color changes.

Part C. Effect of pH on an indicator.

- [ ] *Step 1:* – Obtain two 50 mL beakers, clean them, and mark them with letters A (for Acid) and B (for Base). Find and clean a test tube.
- [ ] *Step 2:* – Add 10 mL of distilled water to each beaker and 1 mL of distilled water to the test tube.
- [ ] *Step 3:* – Add 4 drops of 6 M HCl to beaker A and stir it (this is the diluted acid).
- [ ] *Step 4:* – Add 4 drops of 6 M NH_3 to beaker B and stir it (this is the diluted base).
- [ ] *Step 5:* – Add 4 drops of the indicator (methyl orange) to the test tube.
- [ ] *Step 6:* – Add 2 drops from the diluted acid in beaker A to the test tube. Mix gently the solution. Record the color on the results page.
- [ ] *Step 7:* – Drop by drop, add the diluted ammonia solution from beaker B until the color changes. Homogenize the solution from time to time. Write down the number of drops.
- [ ] *Step 8:* – Attempt one more time for another color change by adding drop-by-drop diluted acid solution. Write down the number of drops.

Part D. Effect of pH on solubility.

- [ ] *Step 1:* – Add 5 mL of 6 M NaOH to a 50mL beaker using the graduated cylinder.
- [ ] *Step 2:* – Rinse a 10mL cylinder with water 3 times. Then, use it to add 5 mL of 1 M $Ca(NO_3)_2$ to the beaker.
- [ ] *Step 3:* – Stir the mixture.
- [ ] *Step 4:* – Make a cone with filter paper and place it in afunnel, on top of the Erlenmeyer. Filter the solution with the precipitate. Carry out a couple of washings with distilled water to make sure you collect all solids. Transfer with a spatula the white solid to a small clean beaker.
- [ ] *Step 5:* – Add 10 mL of distilled water to the beaker with the white solid. Stir the mixture. Do not expect the solid to completely dissolve. The solution is saturated.
- [ ] *Step 6:* – Add drops of 6 M HCl until a change is observed. Record your results.
- [ ] *Step 7:* – Add drops of 6 M NaOH until a change is observed. Record your results.

Part E. Heat as a product.

- ☐ *Step 1:* – Obtain a Bunsen burner, a stand, two iron rings, a wire gauze, two 250 mL beakers, and a test tube.
- ☐ *Step 2:* – Prepare the setup for the bunsen burner, using the second iron ring to protect the beaker from falling. Put about 100 mL of distilled water in the beaker and bring the water to a boil.
- ☐ *Step 3:* – Add 5 drops of 0.1 M $Co(NO_3)_2$ to the test tube.
- ☐ *Step 4:* – Record the color of the liquid solution after each step.
- ☐ *Step 5:* – Add drops of 12 M HCl until the color of the solution changes. You might need to stir the test tube to help mix the reagents.
- ☐ *Step 6:* – Add 5 drops of distilled water and mix.
- ☐ *Step 7:* – Place the test tube in the boiling water and wait for another color change.
- ☐ *Step 8:* – Attempt to reverse the reaction by placing the test tube in a beaker with ice or cold water.

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Le Châtelier's Principle

1. Write the equilibrium constant expressions for all reactions involved in this experiment.

2. How does pH affect the solubility of $Ca(OH)_2$?

3. Is heat being consumed (reactants) or produced (product) in an endothermic reaction?

4. Is heat being consumed (reactants) or produced (product) in an exothermic reaction?

STUDENT INFO	
Name: ________________	Date: ________

Results
EXPERIMENT

Le Châtelier's Principle

Part A. Colored complexed ions: $FeSCN^{2+}$

$$\underset{\text{yellow}}{Fe^{3+}\,(aq)} + SCN^{-}\,(aq) \rightleftharpoons \underset{\text{red}}{FeSCN^{2+}\,(aq)}$$

After step #			Before color	After color
5	Add Fe^{3+}	Solution in test tube A	________	________
6	Add SCN^{-}	Solution in test tube B	________	________
7	Add H_2O	Solution in test tube R	________	________

Part B. Colored complexed ions: $Ni(NH_3)_6{}^{2+}$.

$$\overset{\text{green}}{Ni^{2+}\,(aq)} + 6\,NH_3\,(aq) \rightleftharpoons \overset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}}$$

$$\Downarrow$$

$$\underset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}} + 6\,HCl\,(aq) \rightleftharpoons \underset{\text{green}}{NiCl_2\,(aq)} + 6\,NH_4Cl\,(aq)$$

After step #	Color		
2	Add Ni^{2+}	________	
3	Add NH_3	________	# drops added _____
4	Add HCl	________	# drops added _____

Part C. Effect of pH on an indicator.

$$\underset{\text{yellow}}{H_2In_{(aq)}} \quad + \quad H_2O \longrightarrow \underset{\text{red}}{HIn^-_{(aq)}} \quad + \quad H_3O^+$$

After step #		Before color	After color	
5	Add indicator	________	________	
6	Add diluted acid	________	________	
7	Add diluted base	________	________	# drops added ___
8	Add diluted acid	________	________	# drops added ___

Part D. Effect of pH on solubility.

$$Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + OH^-\,(aq)$$
$$\Downarrow$$
$$OH^-\,(aq) + H^+\,(aq) \rightleftharpoons H_2O\,(aq)$$

After step #		Indicate # drops added
6	Add acid	___
7	Add base	___

Part E. Heat as a product.

$$\underset{\text{violet}}{CoCl_4{}^{2-}\,(aq)} + 6\,H_2O\,(l) \rightleftharpoons \underset{\text{pink}}{Co(H_2O)_6{}^{2+}\,(aq)} + 4\,Cl^-\,(aq) + \textbf{HEAT}$$

After step #		color	
3	Add Co^{2+}	________	
5	Add acid	________	# drops added ___
6	Add H_2O	________	
7	Heat	________	
8	Cool	________	

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Le Châtelier's Principle

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1. In part A. Explain the different colors based on the equilibrium reaction and Le Châtelier's Principle.

2. In part A. Given that water is not involved in the equilibrium, explain the color change you observed after adding water to the mixture.

3. In part B. What did you observe after adding the acid and after adding the base? Explain the different colors based on the equilibrium reaction and Le Châtelier's Principle.

4. In Part C. What did you observe after adding the acid and after adding the base? Explain the changes based on the equilibrium reaction and Le Châtelier's Principle.

5. In Part D. What did you observe after adding the acid and after adding the base? Explain the changes based on the equilibrium reaction and Le Châtelier's Principle?

6. In Part E. What did you observe after heating and after cooling? Explain the changes based on the equilibrium reaction and Le Châtelier's Principle.