Le Chatelier's Principle

Introduction:

In this experiment you will observe shifts in equilibrium systems when conditions such as concentration and temperature are changed. You will explain the observed color changes of four reactions in terms of Le Chatelier's principle.

Le Chatelier predicted how an equilibrium system would respond to changes or stresses imposed on it. When a reversible reaction is at equilibrium, the rates of the forward and reverse reactions are equal, and the concentrations of all reactants and products remain constant with time. The resulting ratio of product concentrations over reactant concentrations is called the equilibrium constant. If a stress is added to an equilibrium system, it shifts in the direction that will relieve the stress until it reaches a new equilibrium state in which the concentration of each chemical will be different than the original yet the value of the equilibrium constant will be the same.

In this experiment you will investigate four different equilibrium systems. For the first two systems, colored substances are used to allow you to determine the direction of equilibrium shifts based on color changes. The substances involved in the last two systems are weak acid and weak base equilibrium reactions, so indicators are added to help detect equilibrium shifts. (An indicator is a substance that changes color in a specific pH range so you will also be able to determine whether the mixture is becoming more acidic or basic.)
Refer to Section 13.3 of Openstax Chemistry for information on Le Chatelier's Principle and determining directions of shift.

Explanations of Observations and Drawings of Equilibrium Solutions:

Explanations of Observations:

Your data tables will have three columns: **Change imposed** (what stress or chemical was added to the system?), **Observation** (what color change was observed?), and **Explanation**. For your explanation, answer the following questions: 1) What change was imposed on the system? (i.e. what chemical's concentration **in the reaction** was increased or decreased, or was heat added or removed?) 2) Does the reaction shift right (forward direction) or left (reverse direction)? 3) As a result of this shift, which chemical has increased or decreased in concentration causing the observed color change?

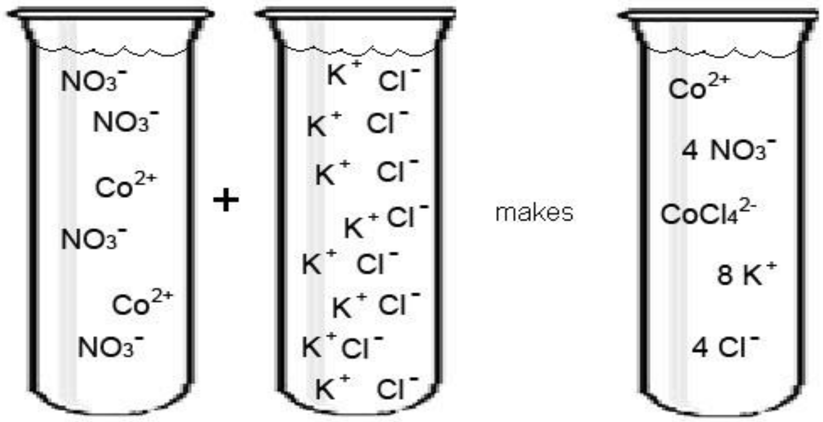
For example, when **cobalt(II) nitrate and potassium chloride** are added together at room temperature the resulting solution is purple once equilibrium is reached, indicating that both pink Co^{2+} ions and blue $CoCl_4^{2-}$ ions (both reactants and products) are present in appreciable amounts. This means this reaction is not heavily reactant or product favored at room temperature. If HCl(aq) were added to the system, the concentration of the Cl^- ions would increase (H^+ ions are spectators), which would cause the equilibrium to shift to the right. As a result of the shift, some of the pink Co^{2+} ions would react, while blue $CoCl_4^{2-}$ ions would form; thus, the solution turns more blue. An example of the data table is shown below.

$Co^{2+}(aq) + 4\,Cl^-(aq) \leftrightarrows CoCl_4^{2-}(aq)$ (original mixture: purple color)
pink *colorless* *blue*

Change Imposed	Observation	Explanation
Add HCl(aq)	The solution turns blue.	1. $[Cl^-] \uparrow$
		2. Reaction $\rightarrow$
		3. $[CoCl_4^{2-}] \uparrow$, solution turns blue

Drawing Example: When we draw an equilibrium reaction, we must draw all the chemicals present in solution, not just the substances in the net ionic equation. Thus, spectator ions are included in equilibrium drawings. Because this reaction reaches equilibrium, the resulting solution is not just products, but a mixture of both reactants and products.

In the example below, two units of cobalt(II) nitrate are combined with a stoichiometric amount of potassium chloride (8 units), which forms the resulting equilibrium solution. (The net ionic equation shown on page 1 cannot be written from a double replacement reaction because it contains the complex ion, $CoCl_4^{2-}$). Note that one cobalt ion has reacted, while the other one has not so the equilibrium solution contains both reactants and products!



Co^{2+} (aq)	+	$4\ Cl^-$ (aq)	$\leftrightarrows$	$CoCl_4^{2-}$ (aq)
2 units $Co(NO_3)_2$		8 units KCl (from 4:1 ratio of Cl^- to Co^{2+} in the net ionic reaction!)		equilibrium mixture: 1 Co^{2+} combined with 4 Cl^- to make $CoCl_4^{2-}$; 1 Co^{2+} and 4 Cl^- are left unreacted; spectator ions: 8 K^+ and 4 NO_3^-

Materials:

Hot plate	0.20 M $CuSO_4$
250 mL beaker	saturated KBr
2 - 150 mL beakers	0.10 M $FeCl_3$
3 - 30 mL beakers	0.10 M NH_4SCN
Test tubes and rack	0.10 M $HgCl_2$
Tap water	15 M NH_3
DI water	Phenolphthalein indicator
Ice	2 M NH_4Cl
Rock salt	2 M CH_3COOH
stirring rod	Methyl orange indicator
	2 M $NaCH_3COO$

Procedure:

Part I. Formation of $CuBr_4^{2-}$ and Temperature Effects

Cu^{2+} ions from copper(II) sulfate combine with Br^- ions from potassium promide to form the yellow $CuBr_4^{2-}$ complex ion in the equilibrium shown in the **net ionic** equation below:

$$Cu^{2+}(aq) + 4\,Br^-(aq) \leftrightarrows CuBr_4^{2-}(aq)$$

blue *colorless* *yellow*

1. Prepare a hot water bath by filling a 250 mL beaker about half full with tap water and setting it on a hot plate (for use in Step 6 below). Turn the heat setting to 5 (or a temperature of 85 – 90 degrees). While the water bath is heating, prepare the solutions below.
2. Prepare an ice-water/salt bath by adding 1 scoop of rock salt to about 75 mL of an ice-water mixture in a 150 mL beaker.
3. Place about 15 mL of 0.20 M $CuSO_4$ solution in a 30 mL beaker. Add about 5 mL of saturated KBr solution and stir to mix. The equilibrium between the two reactants will set up very quickly.
4. Divide the solution into three roughly equal parts in three test tubes. Keep one tube at room temperature for comparison. Use the other tubes for steps 5 and 6 below.

Note: While waiting for steps 5 and 6 below, you can continue on to the other reactions.

5. Place the second test tube in the hot water bath for **at least** 15 minutes or until a color change is observed. Turn off the heat if the solution begins to boil or turns brown because the solution may decompose at higher temperatures. Compare the color of the heated sample to the room temperature sample and record your observations.
6. Place the third test tube in the ice-water/salt bath to chill for **at least** 15 minutes. Compare to the room temperature sample and record your observations.

Waste Disposal: Discard your solutions in the waste container in the fume hood.

Safety: When you have completed each experiment, place rinsed, wet test tubes upside down in the test tube rack to dry. DO NOT attempt to dry the test tubes with a paper towel and glass stir rod since the pressure can break the test tubes.

Part II. Formation of $FeSCN^{2+}$ and Concentration Effects

Iron(III) ions, Fe^{3+}, and thiocyanate ions, SCN^-, combine to form the iron(III) thiocyanate complex ion, $FeSCN^{2+}$, in the equilibrium shown in the **net ionic** equation below:

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

yellow *colorless* *red*

1. Place 50 mL of deionized water in a 150 mL beaker, then add 20 drops of 0.10 M $FeCl_3$ and 20 drops of 0.10 M NH_4SCN. Stir the solution until it is thoroughly mixed.
2. Half-fill four test tubes with the equilibrium mixture prepared in step 1. Keep one tube as a color control for comparison. Add 20 drops of the 0.10 M $FeCl_3$ solution to the second tube, 20 drops of the 0.10 M NH_4SCN solution to the third tube, and 20 drops of 0.10 M $HgCl_2$ solution to the fourth tube. Compare the colors of each tube to the control tube and record your observations.

Note: Hg^{2+} ions will react with SCN^- ions to form a colorless complex ion, $Hg(SCN)_4^{2-}$.

Waste Disposal: Discard your solutions in the waste container in the fume hood.

Part III. Ionization of a Weak Base and Concentration Effects

The weak base ammonia, NH_3, reacts with water to produce hydroxide ions, OH^-, according to the following equilibrium reaction:

$$NH_3\ (aq) + H_2O\ (l) \leftrightarrows NH_4^+\ (aq) + OH^-\ (aq)$$

You will use phenolphthalein as the indicator to study this reaction. Note that phenolphthalein is colorless in acidic solutions and bright pink in basic solutions.

1. Place 15 mL of deionized water in a 30 mL beaker, then add two drops of 15 M NH_3 and two drops of phenolphthalein indicator. Mix the solution, then divide the solution into two portions by pouring it into two separate test tubes. Keep one tube as a color control for comparison.
2. Add 15 drops of 2 M NH_4Cl solution to the other tube and gently shake the test tube to mix the contents. Record your observations.

Waste Disposal: Discard your solutions in the waste container in the fume hood.

Part IV. Ionization of a Weak Acid and Concentration Effects

Acetic acid, CH_3COOH*(aq),* is a weak acid that undergoes the following ionization reaction with water:

$$CH_3COOH\ (aq) + H_2O\ (l) \leftrightarrows H_3O^+\ (aq) + CH_3COO^-\ (aq)$$

For this system, you will use the methyl orange indicator, which is red in very acidic solutions having pH < 3.2, yellow in solutions having pH > 4.4, and orange in between.

1. Place 15 mL of deionized water in a 30 mL beaker, then add two drops of 2 M CH_3COOH and two drops of methyl orange indicator. Mix the solution, then divide the solution into two portions by pouring it into two separate test tubes. Keep one tube as a color control for comparison.
2. Add 5 drops of 2 M $NaCH_3COO$ solution to the other tube, and gently shake the test tube to mix the contents. Record your observations.

Waste Disposal: Discard your solutions in the waste container in the fume hood.

Le Chatelier's Principle Pre-Lab Questions and Calculations

You will not turn this assignment in; it will completed in Canvas 1 hour before your lab period.

1. State Le Chatelier's Principle.

2. What is/are the stress(es) that will be applied to the system in Part I (Formation of $CuBr_4^{2-}$)?

3. What is/are the stress(es) that will be applied to the system in Part II (Formation of $FeSCN^{2+}$)?

4. What indicator is used in Part III (Ionization of a Weak Base)? Spelling counts!

5. What indicator is used in Part IV (Ionization of a Weak Acid)? Spelling counts!

6. What safety clothing/equipment are needed for this experiment?

7. How should test tubes be handled at the end of the experiment?

8. How should chemical waste be handled at the end of the experiment?

Name: ______________________________ Partners: _________________________

Le Chatelier's Principle Lab Report, pages 6 – 7

Data

Table 1: Part I. Net ionic equation: **Cu^{2+} (aq) + $4Br^{-}$ (aq) ⇌ $CuBr_4^{2-}$ (aq)**

colors: _____________ _____________ _____________

original equilibrium color: ______________________________

Change Imposed	**Observation**	**Explanation**
		1.
		2.
		3.
		1.
		2.
		3.

Table 2: Part II. Net ionic equation: **Fe^{3+} (aq) + SCN^{-} (aq) ⇌ $FeSCN^{2+}$ (aq)**

colors: _____________ _____________ _____________

original equilibrium color: ______________________________

Change Imposed	**Observation**	**Explanation**
		1.
		2.
		3.
		1.
		2.
		3.
		1.
		2.
		3.

Table 3: Part III. **NH_3 (aq) + H_2O (l) ⇌ NH_4^+ (aq) + OH^- (aq)**

original equilibrium color with indicator: ______________________________

Change Imposed	**Observation**	**Explanation**
		1.
		2.
		3.

Table 4: Part IV. **CH_3COOH (aq) + H_2O (l) ⇌ CH_3COO^- (aq) + H_3O^+ (aq)**

original equilibrium color with indicator: ______________________________

Change Imposed	**Observation**	**Explanation**
		1.
		2.
		3.

Instructor Initials: _______

Conclusion: In the space below, **summarize** the shifts in equilibrium that occur as changes in reactant and product concentrations are made. Also summarize the shifts in equilibrium that occur with temperature changes; how can a reaction be determined to be endothermic or exothermic based on equilibrium shifts? Be sure to include examples from the four equilibrium systems to support your conclusions.

Post-Lab Questions – These questions will not be graded as part of your lab report grade. You will be responsible for the information in these questions and able to answer these or similar questions on the post-lab quiz at the start of next week's lab period.

1. Copper(II) sulfate solution is blue and potassium bromide solution is colorless. Why does the solution in Part I turn green when they are added together? What substances are in the resulting solution that make it green?

2. Is the reaction in Part I *endothermic or exothermic* in the forward direction? (It can't be both!) Explain based on your observations for the hot and cold baths.

3. The chemical $HgCl_2$ is not part of the reaction in Part II, so explain how its addition caused the equilibrium to shift.

4. Is the solution in Part III becoming more acidic or basic upon addition of NH_4Cl? Explain. (Hint: what is happening to the hydroxide ion concentration, $[OH^-]$?)

5. Is the solution in Part IV becoming more acidic or basic upon addition of $NaCH_3COO$? Explain. (Hint: what is happening to the hydronium ion concentration, $[H_3O^+]$?)

6. Consider the following equilibrium reaction:

$$2\ SO_2\,(g) + 2\ H_2O\,(g) \rightleftharpoons 2\ H_2S\,(g) + 3\ O_2\,(g) \qquad \Delta H = +\ 1037\ kJ$$

   For a-d, predict the effect of the following changes on the position of equilibrium (left, right, no change), when each of the following changes is made:

   a. Adding H_2S*(g)* _________________   c. Removing SO_2*(g)* _________________

   b. Increasing the temperature _________________   d. Adding a catalyst _________________

7. Discuss at least 2 sources of error. How did they affect your results and how would you correct them if you were to repeat the experiment?

Molecular-level Drawings of Equilibrium Systems – next page

I. Formation of $CuBr_4^{2-}$ (Hint: See pages 1 and 2 for a similar example!)

$$\underset{\textit{blue}}{Cu^{2+}\,(aq)} \quad + \quad \underset{\textit{colorless}}{4\ Br^-\,(aq)} \quad \rightleftharpoons \quad \underset{\textit{yellow}}{CuBr_4^{2-}\,(aq)}$$

The reaction above is a net ionic reaction which means the spectator ions are not shown. But it is important to remember that the spectator ions are still present in the solution so we will include them in our drawings of reactions in solution. However, we will not draw the solvent water molecules, even though they are also present in all aqueous solutions.

To create this equilibrium system we began with the first reactant: the copper(II) sulfate solution. Draw a representation of **two units of copper(II) sulfate** in the first test tube. Draw potassium bromide in the second test tube in the correct stoichiometric amount required to be added to the two units of copper(II) sulfate. **Refer to Equation 1 above for the correct mol-mol ratio for the reactants!** Then draw the resulting equilibrium mixture in the third test tube; don't forget to include the spectator ions. Remember reactions that set up equilibrium don't go to completion, there are always some reactants and some products in the mixture, so your drawing should represent this.

I. Formation of $CuBr_4^{2-}$ (Hint: See pages 1 and 2 for a similar example!)

$Cu^{2+}(aq) + 4\,Br^{-}(aq) \rightleftharpoons CuBr_4^{2-}(aq)$

blue *colorless* *yellow*

2 units
copper(II) sulfate

___ units
potassium bromide

equilibrium mixture

II. Formation of $FeSCN^{2+}$

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

yellow *colorless* *red*

2 units
iron(III) chloride

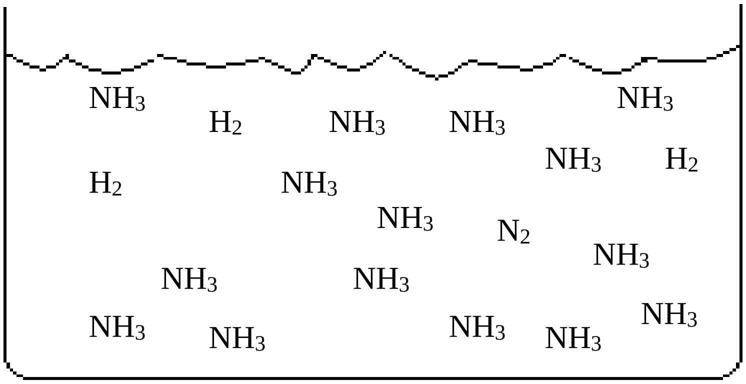
___ units
ammonium thiocyanate

equilibrium mixture

III. Calculating K_c from a drawing: For the two examples below, balance the equation, write the K_c expression, predict whether K_c is large or small based on the drawing shown, and calculate the value of K_c (to 3 sig figs) based on the drawing shown in the beaker.

a. ____ N_2 (g) + ____ H_2 (g) $\leftrightarrows$ ____ NH_3 (g)

K_c expression = ________________

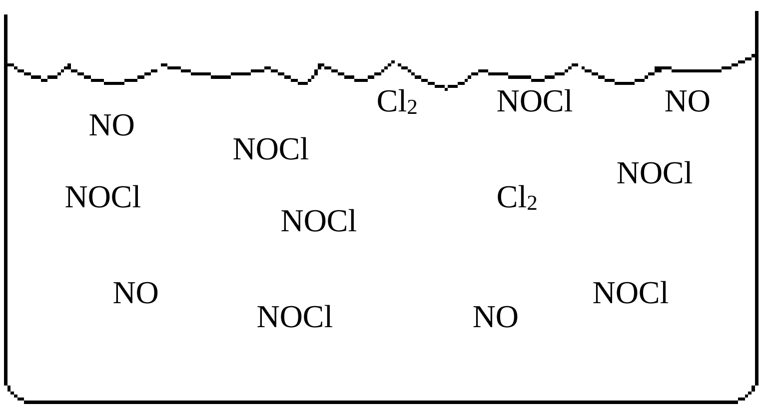


K value prediction: large or small

K_c value: ______________________

b. ____ NOCl (g) $\leftrightarrows$ ____ NO (g) + ____ Cl_2 (g)

K_c expression = ___________________



K value prediction: large or small

K_c value: _______________________