

LE CHATELIER'S PRINCIPLE AND BUFFERS

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INTRODUCTION

Chemical equilibrium in a reaction mixture occurs when the rate of the forward reaction is the same as the rate of the reverse reaction.

Once a reaction mixture has reached equilibrium, changes in concentration, temperature, or pressure that are made to the system can cause the system to no longer be at equilibrium. The composition of the reaction mixture will then shift until equilibrium has been reestablished. This is known as Le Chatelier's principle.

More formally, Le Chatelier's principle states that if a reaction mixture at equilibrium is disturbed by a change, the system will shift its equilibrium composition in a way that tends to counteract the change in an attempt to reestablish equilibrium.

In the first part of the experiment, you will study how changes in concentration and temperature affect an equilibrium reaction mixture.

A buffer is a solution that undergoes a minimal change in pH when small amounts of acid or base are added to it. In the second part of the experiment, the pH of various solutions will be determined both before and after small amounts of acid and base are added to them.

PROCEDURE and DATA REPORT

Dispose of all chemicals according to your teaching assistant's instructions.

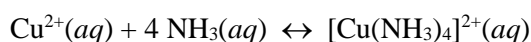
Prepare a hot water bath (250 mL beaker) for use later on in the experiment.

You should write down all observations for each experiment.

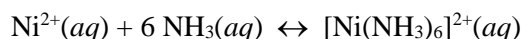
Part I.A

- 1) Place 20 drops of 0.1 M CuSO_4 and 20 drops of 0.1 M NiCl_2 into separate small test tubes. You should note the color of the solutions for each step.
- 2) Slowly add drops of concentrated NH_3 to each solution until a color change occurs. Do this step in the hood since NH_3 has a strong odor and you should avoid breathing it. Aga

When NH_3 is added to an aqueous solution containing Cu^{2+} , the complex ion $[\text{Cu}(\text{NH}_3)_4]^{2+}$ forms.



When NH_3 is added to an aqueous solution containing Ni^{2+} , the complex ion $[\text{Ni}(\text{NH}_3)_6]^{2+}$ forms.



- 3) Slowly add drops of 1 M HCl to each solution until the color changes again.



For your lab report you will need to explain your observations using Le Chatelier's principle.

Part I.B

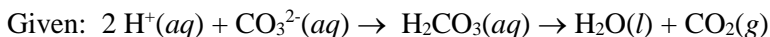
When a saturated solution of a salt is prepared, an equilibrium is established between the cations and anions that make up the salt and the salt in the solid phase. According to the solubility rules given in your textbook, Ag_2CO_3 , AgCl , AgI , and Ag_2S are all insoluble salts containing Ag^+ . In this part of the experiment, the solubility equilibria of these four salts will be investigated. Be sure to write down your observations and be able to explain your observations for each step. For your lab report you will need to explain your observations using Le Chatelier's principle.

****USE THE SAME TEST TUBE WITHOUT POURING OUT THE TEST TUBE CONTENTS IN ALL SIX PROCEDURES IN PART I.B!!!****

- 1) Ag_2CO_3 is an insoluble salt. In aqueous solution: $2 \text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \leftrightarrow \text{Ag}_2\text{CO}_3(s)$

In a small test tube add 10 drops of 0.01 M AgNO_3 to 10 drops of 0.1 M Na_2CO_3 .

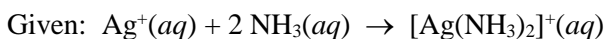
- 2) Slowly add drops of 6 M HNO_3 until the chemical change that you observe is complete.



- 3) AgCl is an insoluble salt. In aqueous solution: $\text{Ag}^+(aq) + \text{Cl}^-(aq) \leftrightarrow \text{AgCl}(s)$

Slowly add drops of 0.1 M HCl until you see evidence of a chemical change.

- 4) Slowly add drops of concentrated NH_3 to your assigned solution until the chemical change that you observe is complete. Do this step in the hood since NH_3 has a strong odor and you should avoid breathing it.



- 5) AgI is an insoluble salt. In aqueous solution: $\text{Ag}^+(aq) + \text{I}^-(aq) \leftrightarrow \text{AgI}(s)$

Add drops of 0.1 M KI until you see evidence of a chemical change.

- 6) Ag_2S is an insoluble salt. In aqueous solution: $2 \text{Ag}^+(aq) + \text{S}^{2-}(aq) \leftrightarrow \text{Ag}_2\text{S}(s)$

Add drops of 0.1 M Na_2S until evidence of a chemical change has occurred.

Part I.C

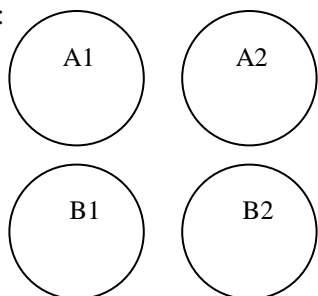
Be sure to write down your observations and be able to explain your observations for each step.

- 1) Place a small amount (pea sized) of CuBr_2 in a small test tube. Add 5 drops of water.
- 2) Add additional water, 5 drops at a time, until a total of about 2 mL of solution is in the test tube. Note the solution's **color** after each addition.
- 3) Put the test tube in the hot water bath for several minutes.
- 4) Is the reaction, $[\text{CuBr}_4]^{2-}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l}) \leftrightarrow 4 \text{Br}^{-}(\text{aq}) + [\text{Cu}(\text{H}_2\text{O})_4]^{2+}(\text{aq})$, endothermic or exothermic as written? Explain your answer using Le Chatelier's principle.

Part II.

The pH of the solutions in this part of the lab is determined by comparing the color of the solution with a universal indicator color chart.

Well plate:



- 1) Place 20 drops of deionized water in wells A1 and A2 of the well plate. Add 3 drops of universal indicator.

Well A1: color: _____ pH: _____

Well A2: color: _____ pH: _____

- 2) Place 10 drops of 0.1 M $\text{HC}_2\text{H}_3\text{O}_2$ in wells B1 and B2 of the well plate. Add 3 drops of universal indicator.

Well B1: color: _____ pH: _____

Well B2: color: _____ pH: _____

- 3) Add 10 drops of 0.1 M $\text{NaC}_2\text{H}_3\text{O}_2$ to wells B1 and B2. Use $\Delta\text{pH} = \text{pH}_{\text{final}} - \text{pH}_{\text{initial}}$

Well B1: color: _____ pH: _____ ΔpH : _____

Well B2: color: _____ pH: _____ ΔpH : _____

- 4) Add 5 drops of 0.10 M HCl to wells A1 and B1.

Well A1: color: _____ pH: _____ ΔpH : _____

Well B1: color: _____ pH: _____ ΔpH : _____

- 5) Add 5 drops of 0.10 M NaOH to wells A2 and B2.

Well A2: color: _____ pH: _____ ΔpH : _____

Well B2: color: _____ pH: _____ ΔpH : _____

- 6) Explain all of your observations and results for this part of the experiment in your lab report. Pay particular attention to the magnitude of ΔpH in parts 3 through 5.