AP Chemistry - Le Chatelier's Principle
Student Guide
IS 8004

INTRODUCTION

Le Chatelier's principle is a simple guideline to help predict the effect of conditional changes on chemical equilibrium. Named for the French chemist Henri Le Chatelier, who in 1884 suggested that if a chemical system at equilibrium were to experience a change, the equilibrium of the system would adjust to minimize the effects of the change. Examples of conditional changes include pressure, temperature, and concentration.

It is important to note that Le Chatelier's principle is strictly qualitative. It is also important to note that Le Chatelier's principle does not explain anything. It merely provides an indication of which direction the equilibrium will shift under changing conditions. Remember that a system at dynamic equilibrium is not static. It is still reacting, it just reaches a point where it is reacting equally in both directions.

Consider the following fictional reaction:

$$A + B \longleftrightarrow C + D$$

It is a reversible reaction. In other words, it can go in either direction, depending on certain conditions. A and B can react to form C and D and C and D can react to form A and B. Left alone, the reaction will eventually reach dynamic equilibrium. It does not mean that it is not reacting, it is simply reacting equally in both directions, resulting in no further net gain on either side of the reaction.

Suppose you were to increase the concentration of the reactant A. You have now introduced a conditional change in the system. In an attempt to maintain equilibrium, the system will attempt to decrease the concentration of A by reacting A and B, in turn producing more C and D. It can be said that the equilibrium for the system has shifted to the right since it moved in the direction of C and D production.

The opposite applies as well. Suppose you were to decrease the concentration of B. To maintain equilibrium, the C and D would react to replace the B that has been removed from the system. In this case, equilibrium has shifted to the left.

In a gaseous system, pressure is the result of the number of molecules in a given volume. The Haber process is the reaction of nitrogen and hydrogen to form ammonia:

$$N_2\,(g) + 3\,H_2\,(g) \longleftrightarrow 2\,NH_3 + \text{heat}$$

Since pressure is the result of the number of molecules of gas, if pressure were to be increased in the above system, equilibrium would shift to the right. In other words, more nitrogen and hydrogen would combine to form ammonia, reducing the number of molecules. The above reaction can also be used to demonstrate the effect of temperature change on equilibrium. In the Haber process, when the reaction proceeds from left to right, it is exothermic. If temperature is increased, the equilibrium will shift to the left. Because the reverse reaction is endothermic, the shift to the left would occur as the system attempts to use the excess heat.

In this experiment, you will examine three different reactions. After dynamic equilibrium is reached, you will then change (stress) the system to determine the equilibrium shift.

In the first reaction, you will prepare a saturated solution of sodium chloride (NaCl). When placed in water, NaCl dissociates into Na^+ and Cl^- ions.

$$NaCl\,(s) \longleftrightarrow Na^+\,(aq) + Cl^-\,(aq)$$

You will then increase the concentration of one of the ions to determine how the system reacts to the stress.

In the second reaction, you will use a pH indicator, more specifically bromothymol blue (BTB). Bromothymol blue changes color at a specific pH (pH being an indication of the hydronium or hydroxyl ion concentration). Different indicators change color at different pH ranges. The pH range an indicator changes color in is called the visual transition interval. For bromothymol blue, the visual transition interval is pH 6.0 (yellow) to pH 7.6 (blue).

$$\underset{\text{(yellow)}}{BTBH\,(aq)} \longleftrightarrow \underset{\text{(blue)}}{BTB^-\,(aq)} + H^+$$

By altering the concentration of hydronium ions, you can examine the equilibrium shift.

The last reaction involves combining complex ions in solution. Two colorless components, Fe^{3+} and SCN^-, combine to form a molecule red in color.

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

The reaction is stressed by the addition and removal of ions. Since the $FeSCN^{2+}$ is red, if the reaction shifts to the right, the darker red the solution will be.

Objective

Demonstrate Le Chatelier's Principle by observing physical changes to these solutions as the systems are stressed.

Materials Included in the Kit

2 X 25 mL	Bromothymol blue indicator solution
1 X 25g	Potassium Thiocyanate
2 X 25mL	Ferric nitrate solution
2 X 200mL	Potassium Thiocyanate solution
1 X 25g	Sodium Phosphate, dibasic
6 X 25mL	Hydrochloric acid 36%
1 X 200 mL	Sodium hydroxide solution
1 X 100g	Sodium Chloride

Materials Needed but not Supplied

75 each	13 X 100mm Test tubes
15 each	Test tube racks
15 each	Funnels with filter paper

Safety

Rubber gloves
Apron
Safety goggles

Safety note: *Concentrated Hydrochloric acid is extremely corrosive and the liquid or vapors can cause severe burns. Sodium Hydroxide is also hazardous and can cause extreme irritation to skin and eyes. Wash hands after handling. Consult MSDS for first aid information.*

Procedure

Part I: Equilibrium in a Saturated Solution

1. Place a small amount of sodium chloride in a 13 X 100mm test tube and fill 3/4 with water. Mix well. If all of the sodium chloride goes into solution, add more and mix. Continue until there is a slight excess of solid in the bottom of the tube indicating that the solution is saturated. Filter into a second test tube.

2. Add a few drops of concentrated hydrochloric acid (a source of Cl^- ions) to the filtered sodium chloride solution and note the results.

Part II: Equilibrium of a pH Indicator in Solution

1. Bromothymol blue is an organic dye which changes color based on the presence or absence of hydrogen ions. Fill a 13 X 100mm test tube 1/2 full of water and add several drops of BTB and mix.

2. In a second test tube fill 3/4 with water and add one drop of concentrated hydrochloric acid. Mix well.

3. Add 5 drops of the diluted hydrochloric acid to the BTB solution and mix. This increases the H^+ concentration. Note the color of the indicator solution.

4. Add the sodium hydroxide solution one drop at a time while mixing until the color changes. Adding sodium hydroxide decreases the H^+ concentration and increases the pH. Note the color change.

Part III: Equilibrium of a Complex Ion in Solution

1. Place about 5mL of potassium thiocyanate solution in a 13 X 100mm test tube. Add 2-3 drops of the ferric nitrate solution and mix well. Note the color and write the ionic equation for the reaction.

2. Add a small crystal of KSCN to the test tube. Do not mix. Observe the color change and note the shift in the ionic equation.

3. Add a drop of the ferric nitrate solution. Observe the color change and note the shift in the ionic equation.

4. Add a crystal of sodium phosphate, dibasic. Observe the color change and note the shift in the ionic equation.

Chemical disposal: *These dilute solutions can all be safely flushed down the drain with copious amounts of water.*

Name:	Instructor:
Date:	Class/Lab Section:

DATA ANALYSIS

	Change	Observations
Part I **Saturated Solution**		
Part II **pH Indicator in Solution**		
Part III **Complex Ion in Solution**		

Name:	Instructor:
Date:	Class/Lab Section:

DATA ANALYSIS

1. Define Le Chatelier's principle and the role equilibrium plays in the reduction of stress.

2. Define the terms ion, anion, cation, dissociation and pH.

3. Describe what you observed when the saturated solution of sodium chloride was stressed by the addition of Cl^- ions, write the ionic equation and explain how Le Chatelier's principle is applied to this experiment.

Name:	Instructor:
Date:	Class/Lab Section:

DATA ANALYSIS

4. As in #3, describe your observations and apply Le Chatelier's principle to the pH indicator experiment.

5. As in #3, describe your observations and apply Le Chatelier's principle to the complex ion experiment.