Le Chatelier's Principle

Introduction

Le Chatelier's principle is a qualitative rule, which allows the prediction of the effect of temperature, pressure and concentration changes on chemical reactions. The principle states: *A chemical system at equilibrium when stressed by external forces will adjust in such a way as to minimize that stress.* For example when a system is subjected to increased pressure it adjusts so that it will occupy less volume. This offsets the pressure increase. If ice is placed under an increased pressure, it melts because the water obtained from a given mass of ice occupies less volume. In the formation of ammonia (the Haber process) from hydrogen and nitrogen, the product of the reaction (NH_3) occupies less volume than the two uncombined gases. The increase in pressure favors the production of ammonia.

This experiment is divided into three separate reactions demonstrating how different types of stress effect equilibrium. Students are asked to predict the outcome of each situation and then prove or disprove their predictions.

The first experiment involves equilibrium in a saturated solution. Sodium chloride is dissolved in water disassociating into Na^+ and Cl^-. The system is then stressed by the addition of one of the ions.

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

The second experiment involves an equilibrium reaction using a pH indicator dye. This dye changes color at a specific pH (hydronium ion or hydroxyl ion concentration). The pH at which the indicators change color is referred to the visual transition interval. The indicator used in this experiment is bromothymol blue which has a visual transition interval of pH 6.0 (yellow) to pH 7.6 (blue). The system is stressed by adding H^+ or OH^- to the solution of indicator.

$$BTBH(aq) \Leftrightarrow BTB^{-}(aq) + H^{+}$$

yellow → blue

The final experiment involves an equilibrium reaction involving complex ions in solution. By adding and removing ions from the solution, the reaction can be stressed and shifted to the right or left. The darker the color of the solution the more the equilibrium of the reaction has shifted to the right.

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

Objective

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

Chemicals and Equipment

Materials included in this 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

1 set Student study and analysis copy masters
1 Teacher's guide

Materials needed but not supplied:

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

Safety equipment required:

- 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

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.

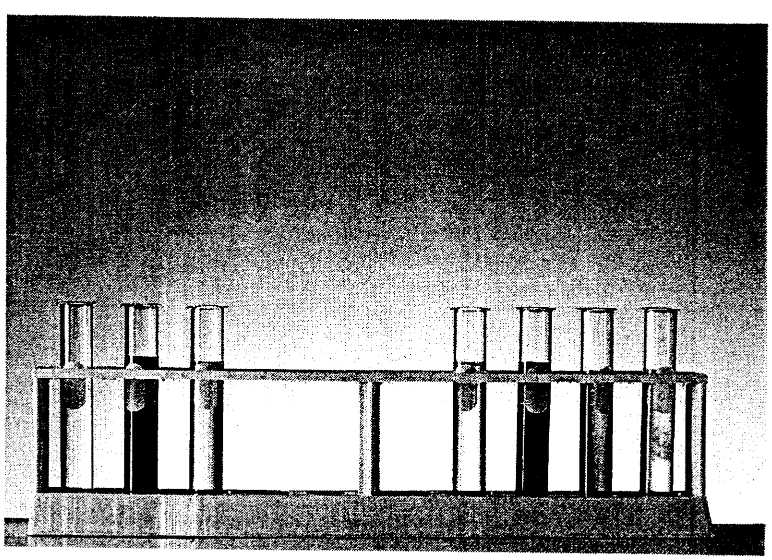
2.) Add a few drops of concentrated hydrochloric acid (a source of Cl^- ions) to the filtered sodium chloride solution and note the results.

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.



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.

Discussion and Laboratory Report

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.

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.

DISCUSSION: Give a thorough explanation of your results. Write the equation for each part and explain your results showing what is increasing or decreasing and how the reaction shifts. You can use diagrams if you want. Make sure you are clear and specific.

6. **For the following reaction, tell how the amount of glucose present at equilibrium would be affected by each of the following**

$$6CO_2(g) + 6H_2O(l) \rightleftarrows C_6H_{12}O_6(s) + 6O_2(g) \quad \Delta H = 2820$$

a) addition of ~~solid~~ carbon dioxide (g)

b) increase in T

c) decrease in V

d) removal of some oxygen gas

e) removal of some glucose

f) addition of a catalyst

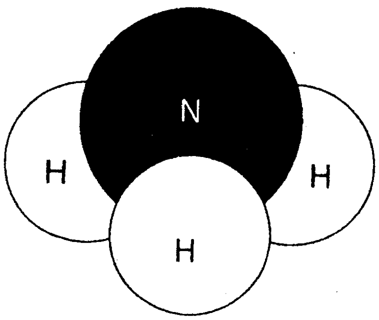
g) removal of some water

Name ______________________ Date ______________ Class ______________

Chapter 22

CHEMISTRY AND INDUSTRY

AMMONIA: THE MAKING OF A SOLUBLE BASE



The pungent gas ammonia takes its name from the Ammonians. These ancient worshippers of the Egyptian god, Amun, used *sal volatile* (ammonium chloride) in their religious rites. During the Middle Ages, ammonia was one of the few known soluble bases and was used in the dyeing of woolen goods and in tanning. This colorless gas is generally thought of as nonflammable but will burn in air under some conditions. It dissolves readily in water because it forms hydrogen bonds with water molecules. The fact that ammonia is extremely soluble in water is what makes it so obvious to your nose. It dissolves in the aqueous mucus that coats the olfactory tissue of the nose. Water might smell just as pungent if our nasal sensors were not constantly saturated with it!

At the beginning of the twentieth century, prominent scientists were warning of approaching world famine because of a scarcity of fertilizer containing nitrogen. The nitrogen is needed by plants to make proteins. Some types of bacteria "fix" nitrogen from the air, forming nitrates, which plants can use. But large-scale farming requires more abundant nitrates, which can be made from ammonia. However, because ammonia could not be produced in abundance, most nitrates had to be imported from mined deposits, mainly from Chile. Nitrates are also used in the manufacturing of explosives. In 1913, as World War I was approaching, Germany was under pressure to obtain ammonia needed in the manufacturing of explosives. These factors led researchers to investigate methods of producing ammonia on an industrial scale. Fritz Haber, a German chemist, learned that ammonia could be produced by the direct combination of nitrogen from the air and hydrogen through the following reaction:

$$N_2 + 3H_2 \rightleftarrows 2NH_3$$

The reaction requires the presence of a catalyst and high temperature and pressure. Because the reaction is reversible, the ammonia must be removed as it is produced to keep the reaction moving to the right, or in favor of the product. Karl Bosch, an engineer from a German company interested in Haber's work, designed equipment that could operate at temperatures up to 550°C and up to 200 atmospheres, making the large scale production of ammonia possible. The process developed to produce ammonia is known as the Haber-Bosch process.

Today, ammonia ranks as one of the most important industrial substances. Modern chemical plants that produce ammonia manufacture thousands of tons per day. The ammonia is then used in the manufacturing of explosives, plastics, soap, and many other common products. However, the greatest percentage of ammonia is used in the production of fertilizers. The nitrogen in fertilizers is supplied directly or indirectly by ammonia. Fertilizers replenish nitrogen and other substances, particularly potassium and phosphorus, that have been reduced or exhausted in soils.

Ammonia is used as a fertilizer in both gaseous and liquid form. The gas is pumped directly into the soil. The liquid form, called anhydrous ammonia, is also added directly to the soil. Plants are able to absorb some of the ammonia, using it to make proteins. Bacteria in the soil convert much of the ammonia to nitrites (NO_2^-) and then to nitrates (NO_3^-). Plants absorb the nitrates and utilize the nitrogen in making proteins.

Many commercial fertilizers supply nitrogen in the form of nitrate salts that are manufactured from nitric acid. Here again, ammonia plays a role because nitric acid is produced from ammonia from a method called the Ostwald process. In the Ostwald process, ammonia is oxidized in the presence of a platinum-rhodium catalyst to yield nitrogen(II) oxide.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

This product oxidizes to nitrogen(IV) oxide.

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

When NO_2 is combined with water, a mixture of nitrous and nitric acids forms.

$$2NO_2(g) + 2H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

Further processing converts HNO_2 to HNO_3 to achieve nearly pure nitric acid. To produce high-nitrogen fertilizer, nitric acid can be reacted with ammonia to produce ammonium nitrate, NH_4NO_3, an important fertilizer. Nitric acid is also used to produce metallic salts, such as potassium nitrate (KNO_3). Such salts are important ingredients in many fertilizers.

Fertilizers can be used to tailor soils for specific crops, to enrich poor soils, and to increase crop yield. However, runoff is a problem with these fertilizers, as they are highly soluble in water. Contamination of bodies of water by nitrogen fertilizers is a serious form of pollution. Increased plant growth in the affected body of water can lead to oxygen depletion and the "death" of the body of water. Controlled release fertilizers are under development, but are currently used only in nonfarming applications due to their cost.

1. What factors led to investigation of methods for producing ammonia on a large scale?

2. Describe the Haber-Bosch process for producing ammonia.

3. Explain why ammonia might be said to be at the beginning of the food chain in industrial-scale farming for food production.

4. Why is ammonia such an important chemical material?

5. What effect does fertilizer runoff have on bodies of water?

6. How does the concept of equilibrium relate to the fact that ammonia must be removed during the Haber-Bosch process if the reaction is to continue?