Experiment 4
CHEMICAL KINETICS AND CHEMICAL EQUILIBRIA

INTRODUCTION

I. Factors Affecting Rates of Chemical Reactions

Chemical kinetics is the study of the rates of chemical reactions and the mechanisms by which reactions occur. The rate of a reaction is defined as the change in concentration of a reactant or a product per unit of time. If the chemical equation for a reaction is known, its rate can be determined by following the change in concentration of any product or reactant that can be detected quantitatively. It is usually expressed in molar per unit of time. The main factors that influence the reaction rate include the nature of the reactants, the concentrations of the reactants, the temperature at which the reaction occurs, and whether or not catalysts are present in the reaction.

A. Nature of the reactants

The physical states of reacting substances are important in determining their reactivities. If particles are in the same phase (e.g., all reactants are in the liquid or gas phase), it is very easy for them to mix with each other. This gives particles the maximum opportunity to collide and react. If the reactants are immiscible, this will mean that the reaction can only occur at the interface of the fluids. If one of the reactants is a solid, the reaction can only take place on the surface of the solid. The smaller the size of the solid particles, the greater the surface area that the reaction can take place in. Therefore, finely divided powder reacts more quickly than the same substance in a big lump. Depending upon what substances are reacting, the time varies. Acid reactions, formation of salts, and ion exchange are fast reactions. When covalent bond formation takes place between the molecules and when large molecules are formed, the reactions tend to be very slow.

B. Concentration

According to the collision theory of chemical reactions, concentration plays a very important role in reactions because molecules must collide in order to react together. As the concentration of the reactants increases, the frequency of the molecules colliding increases, striking each other more frequently by being in closer contact at any given point in time. Consider two reactants in a closed container whose molecules are colliding constantly. Increasing the amount of one or more of the reactants causes these collisions to happen more often, increasing the reaction rate.

C. Temperature

An increase in temperature is accompanied by an increase in the reaction rate. Temperature is a measure of the kinetic energy of a system, so higher temperature implies higher average kinetic energy of molecules and more collisions per unit time. A general rule of thumb for chemical reactions is that the rate at which the reaction proceeds will approximately double for each 10°C increase in temperature. Once the temperature reaches a certain point, some of the chemical species may be altered (e.g., denaturing of proteins) and the chemical reaction will slow or stop.

D. Catalysts

A catalyst is a substance that accelerates the rate of a chemical reaction but remains chemically unchanged afterwards. The catalyst increases rate reaction by providing a different reaction mechanism to occur with a lower activation energy. In autocatalysis, a reaction product is itself a catalyst for that reaction leading to positive feedback. Proteins that act as catalysts in biochemical reactions are called enzymes. Michaelis-Menten kinetics describes the rate of enzyme mediated reactions. A catalyst does not affect the position of the equilibria, as the catalyst speeds up the backward and forward reactions equally.

II. Le Châtelier's Principle

An irreversible reaction is a reaction in which the reactants appear to be used up in forming the products. The reaction seems to go only in one direction or to go to completion. The products do not seem to change back and form the reactants. A reversible reaction is a reaction in which a dynamic equilibrium is maintained by the reactants and the products within a system. Dynamic equilibrium is a condition in which two opposing processes occur at equal rates and in which there is no net change with time. In a static equilibrium, there is no motion whatsoever. A double arrow in the equation indicates a reversible reaction. The reaction proceeding to the right is the forward reaction. The reaction proceeding to the left is the backward or reverse reaction.

Reversible reactions are governed by Le Châtelier's principle. Henri Louis Le Châtelier, a noted French chemist of the late 19th century, formulated this principle in 1884. According to this principle, if a stress is brought to bear upon a system in equilibrium, the equilibrium will be displaced in a direction tending to relieve the stress. The word stress denotes a change in concentration, temperature, or pressure since such changes affect the forward and backward reactions unequally.

A. Effect of Changing the Concentration of a Reactant or a Product

Increasing the concentration of one substance in an equilibrium mixture displaces the equilibrium in that direction which consumes some of the material added. Conversely, decreasing the concentration of a substance causes the production of more of that substance.

B. Effect of Changing the Temperature

Raising the temperature of an equilibrium mixture causes the equilibrium condition to shift in the direction of the endothermic reaction. Lowering the temperature causes a shift in the direction of the exothermic reaction.

C. Effect of Changing the Pressure

When the pressure on an equilibrium mixture involving gases is increased, a net reaction occurs in the direction in which the number of moles of gases becomes smaller. If the pressure is decreased, a net reaction occurs in the direction producing a larger number of moles of gases. The effect of pressure on an equilibrium condition is not limited to reactions involving gases but because gases are so much compressible than liquids or solids, the effect is more pronounced with gases.

In this experiment, the shift in the direction of the equilibrium by changing the concentration of a substance in the equilibrium mixture and by changing the temperature can be readily observed by noting the color changes and by formation of precipitate. The following are the equilibrium systems to be studied.

1. Chromate – dichromate equilibrium

$$\underset{\text{(yellow)}}{2CrO_4^{2-}} + 2H^+ \leftrightarrow \underset{\text{(orange)}}{Cr_2O_7^{2-}} + H_2O$$

2. Thiocyanatoiron (III) complex ion equilibrium

$$Fe^{3+} + SCN^- \leftrightarrow \underset{\text{(red)}}{FeSCN^{2+}}$$

3. Weak acid equilibrium (ionization of acetic acid)

$$HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$$

4. Weak base equilibrium (ionization of ammonia)

$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

5. Saturated salt (sodium chloride) equilibrium

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

OBJECTIVES

1. To illustrate the effects of the various factors affecting the rates of chemical reactions
2. To demonstrate the effects of changing the concentration and temperature to the equilibrium of a system

LIST OF CHEMICALS

10 drops	1M potassium chromate	10drops	0.1M ammonium hydroxide
5 drops	6M hydrochloric acid	5 drops	1M ammonium chloride
2 mL	12M hydrochloric acid	4 mL	5.4M sodium chloride
3 mL	3M hydrochloric acid	6 mL	0.15M sodium thiosulfate
1 mL	0.5M hydrochloric acid	1 mL	1.5M hydrochloric acid
3 mL	1.0M hydrochloric acid	3 mL	1.0M sodium hydroxide
5 drops	6M sodium hydroxide	5 drops	1M sodium acetate
0.5 g	solid potassium nitrate	1 drop	methyl orange
3 mL	0.1M potassium thiocyanate	1 drop	phenolphthalein
10 drops	0.1M acetic acid	3 strips	1-cm aluminum
5 mL	0.1M ferric nitrate	1 strip	1-cm magnesium ribbon
2 mL	10% hydrogen peroxide	0.5 g	iron filings
3 drops	1M cupric sulfate solution		ice, distilled water

LIST OF APPARATUS

8 test tubes, 4 micro test tubes, test tube rack, 10-mL graduated cylinder, 50-mL graduated cylinder, two 150-mL beaker, 50-mL beaker, iron stand, iron ring, Bunsen burner, stirring rod, wire gauze, test tube rack

SAFETY PRECAUTIONS: Hydrochloric acid, acetic acid ammonium hydroxide, and sodium hydroxide are corrosive irritants. These reagents act by injuring the skin, eyes, and, if inhaled, the surface tissues of the respiratory tract. Potassium chromate is a toxic and carcinogenic reagent. Handle them with great care. If any of the solution gets in contact with the skin, immediately wash with abundant amount of water.

PROCEDURE

A. Effect of Nature of the Reactants to the Reaction Rate

1. Get 2 micro test tubes. Put 1 mL of 1M hydrochloric acid in each micro test tube.

2. In one of the micro test tubes, add 0.5 gram of aluminum turnings. To the other micro test tube, place 0.5 gram of iron filings. The two metals must be added to the hydrochloric acid simultaneously.

3. Observe the appearance of bubbles and compare the rates of the chemical reactions.

4. **Waste Disposal:** Place the solutions in a test tube and add 2 drops of phenolphthalein. Add dropwise 1M sodium hydroxide until the color becomes light pink. Dispose the neutralized solution to the sink with plenty of water.

B. Effect of Temperature to the Reaction Rate

1. Place the following reagents in separate micro test tubes.

Test Tube	Reagents
1	3 drops of 0.15M sodium thiosulfate
2	3 drops of 3M hydrochloric acid and 5 drops of distilled water

2. Immerse the 2 micro test tubes in a water bath for 3 minutes. Use a 150-mL beaker with 100 mL tap water as the water bath.

3. Measure the temperature of the tap water (room temperature).

4. Get a piece of white paper marked with an X. Hold the micro test tube 2 over the mark X and pour the content of micro test tube 1 to micro test tube 2. Take the time in seconds to complete the reaction. The reaction is complete when the mark X is no longer visible.

5. Repeat steps 1 to 4 for the other two temperatures: one temperature at 20°C higher than the room temperature and the other temperature at 40°C higher than the room temperature.

6. **Waste Disposal:** Place all solutions in a labeled container.

C. Effect of Concentration to the Reaction Rate

1. Get 3 micro test tubes and place them in a test tube rack.

2. Into each micro test tube, put 1 mL of hydrochloric acid of different concentrations: 0.5M, 1M, and 1.5M.
3. Drop a 1-cm strip of magnesium ribbon in the micro test tube with the 0.5M hydrochloric acid. Record the time in seconds to complete the reaction.

4. Repeat step 3 for the remaining acids.

5. **Waste Disposal:** Place the solutions in a test tube. Add 1 drop of phenolphthalein. Put dropwise 1M sodium hydroxide until the mixture becomes light pink. Dispose the neutralized mixture into the sink with plenty of water.

D. Effect of Catalyst to the Reaction Rate

1. Get two micro test tubes. Into each micro test tube, place 15 drops of 10% hydrogen peroxide.

2. Into one of the micro test tubes, add 3 drops of 1M cupric sulfate.

3. Observe for the appearance of bubbles in the two micro test tubes and compare the rates of the reactions.

4. **Waste Disposal:** Dispose the solutions into the sink with plenty of water.

E. Chromate – Dichromate Equilibrium

1. Place 10 drops of 1M potassium chromate in a micro test tube.

2. Add dropwise 6M hydrochloric acid until a color change is observed.
3. Put drop by drop 6M sodium hydroxide until a color change occurs.
4. Again, add dropwise 6M hydrochloric acid until a color change is observed.
5. **Waste Disposal:** Place the solution in a properly labeled container.

F. Thiocyanatoiron(III) Complex Ion Equilibrium

1. Using a 10-mL graduated cylinder, measure 3 mL of 0.1M ferric nitrate and place the solution in a 150-mL beaker.
2. Add 3 mL of 0.1M potassium thiocyanate.
3. Dilute the solution by adding 50 mL of distilled water. Measure the distilled water using a 50-mL graduated cylinder.
4. Label 6 test tubes with A, B, C, D, E, and F.
5. Into each test tube, put 5 mL of the solution.
6. To test tube A, add 1 mL of 0.1M ferric nitrate.
7. To test tube B, add 1 mL of potassium thiocyanate.
8. To test tube C, add 0.5g of potassium nitrate.
9. Place test tube D in a boiling water bath for 2 minutes.
10. Place test tube E in an ice bath for 2 minutes.
11. To test tube F, add 1 mL of distilled water. This will be the reference solution.
12. Compare the relative intensity of the red color of the thiocyanatoiron(III) complex ion in each of the first five test tubes to that of test tube F.
13. **Waste Disposal:** Dispose the solutions to the sink with plenty of water.

G. Weak Acid Equilibrium (Ionization of Acetic Acid)

1. Place 10 drops of 0.1M acetic acid in a micro test tube.
2. Add 1 drop of methyl orange and observe the color of the solution.
3. Put dropwise 1.0M sodium acetate until a color change occurs.

4. **Waste Disposal:** Dispose the solution to the sink with plenty of water.

H. Weak Base equilibrium (Ionization of Ammonia)

1. Place 10 drops of 0.1M ammonium hydroxide in a micro test tube.
2. Add 1 drop of phenolphthalein and observe the color of the solution.
3. Put dropwise 1.0M ammonium chloride until a color change occurs.
4. **Waste Disposal:** Dispose the solution to the sink with plenty of water.

I. Saturated Salt (Sodium Chloride) Equilibrium

1. In a test tube, place 4 mL of 5.4M sodium chloride solution.
2. Add 2 mL of 12.0M hydrochloric acid. Observe.
3. **Waste Disposal:** Dispose the solution to the sink with plenty of water.

REFERENCES

Chang, R. (2008). General Chemistry: The Essential Concepts, 5th Edition. McGraw-Hill, New York.

Nelson, J.H. and Kemp, K.C. (2003). Chemistry: The Central Science (Laboratory Experiments), 9th Edition. Pearson Education, Inc., USA.

Roberts, Jr.,J.L., Hollenberg, J.L., and Postma, J.M. (1991). General Chemistry in the Laboratory, 3rd Edition. W.H. Freeman and Co., USA

Group No.______

Name:	**Date Performed:**
Course & Section:	**Date Submitted:**
Program & Year:	**Professor:**

Experiment 4
CHEMICAL KINETICS AND CHEMICAL EQUILIBRIA

A. Effect of Nature of Reactants to the Reaction Rate

Which one of aluminum and iron had the faster rate of reaction with hydrochloric acid?

Write the chemical equations for the reactions of aluminum and iron with hydrochloric acid.

Explain your observation with the use of the activity series of the metals.

B. Effect of Temperature to the Reaction Rate

Temperature (^{o}C)	Time (seconds) to Complete the Reaction

Which of the reactions had the fastest rate?

What is the effect of temperature to the rate of a chemical reaction?

C. Effect of Concentration to the Reaction Rate

HCl Concentration	Time (seconds) to Complete the Reaction
0.5M	
1.0M	
1.5M	

Which of the reactions was the fastest?

What is the effect of concentration to the rate of a chemical reaction?

D. Effect of Catalyst to the Reaction Rate

Write the chemical equation for the decomposition of hydrogen peroxide.

What is the effect of cupric sulfate solution to the rate of decomposition of hydrogen peroxide?

E. Chromate – Dichromate Equilibrium

	Color
K_2CrO_4	
K_2CrO_4 + HCl	
K_2CrO_4 + HCl + NaOH	
K_2CrO_4 + HCl + NaOH + HCl	

Explain how HCl exerts an effect to the equilibrium system.

Explain how NaOH exerts an effect to the equilibrium system.

F. Thiocyanatoiron(III) Complex Ion Equilibrium

	Darker or Lighter than the Standard Solution (Test Tube F)
Test tube A + $Fe(NO_3)_3$	
Test tube B + KSCN	
Test tube C + KNO_3	
Test tube D in boiling water bath	
Test tube E in ice bath	

In the equilibrium system Fe^{3+} + SCN^- $\leftrightarrow$ $FeSCN^{2+}$, which reaction is endothermic? Which is exothermic?

G. Weak Acid Equilibrium (Ionization of Acetic Acid)

	Color
$HC_2H_3O_2$ + methyl orange	
$HC_2H_3O_2$ + methyl orange + $NaC_2H_3O_2$	

Explain how $NaC_2H_3O_2$ exerts an effect to the equilibrium system.

H. Weak Base equilibrium (Ionization of Ammonia)

	Color
NH_4OH + phenolphthalein	
NH_4OH + phenolphthalein + NH_4Cl	

Explain how NH_4Cl exerts an effect to the equilibrium system.

I. Saturated Salt (Sodium Chloride) Equilibrium

What happened to the saturated NaCl when HCl was added?

Explain how the addition of HCl shifts the equilibrium of the system.