CHEMISTRY

Lab Manual and Worksheet

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Name:………………………………….……………….

Troy ID: …………………………………………..

Experiment 1

Measurement: Identification of Unknown Liquids by Measuring Their Densities

INTRODUCTION

Measurement is the assignment of a number to a characteristic of an object or event which can be compared with other objects or events. In this experiment, you will work with two liquids and measure their densities. Liquid density is calculated by the following equation:

$$D = \frac{m}{V}$$

Where D is the density, m is the weight of the liquid, and V is the volume of the liquid.

OBJECTIVES

This experiment aims to identify two liquids by their density.

CONCEPTS

This experiment uses the concepts of weight, volume, and density.

TECHNIQUES

Correct use of scales and volumetric flasks are some of the techniques encountered in the experiment.

PROBLEM DESCRIPTION

Two research groups worked independently on two new compounds, a food preservative, and a neurotoxin. They all named their products “Compound A.” An intern reorganized the chemical storage stockroom and placed all the bottles labeled “Compound A” on the same shelf.

You would like to begin testing the new food preservative but do not know which bottle contains the food preservative and which bottle contains the neurotoxin. It would clearly not be a good idea to put neurotoxin into your food products. You have asked a theoretical chemist what to do, and he said that the preservative would have a higher density.[1]

You will need to perform an experiment to determine which bottle of Compound A contains the food preservative.

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[1] The problem description was copied from the Virtual Lab module written by Prof. Sophia Nussbaum, University of British Columbia, 2003. All else was written by the author of this book..

PROCEDURE

Determining the Density of Solution A-1:

1. Use a 100 mL volumetric flask labeled A-1-1. Weigh and record the mass of the empty flask **($g_{đ}$ =)**
2. Using beaker labeled A-1, pour a portion of A-1 solution into flask A-1-1. Then, add more A-1 solvent carefully, using dropper or pipette labeled A-1 to reach the calibration mark.
3. Weigh the volumetric flask containing the solution and record its mass **(gc =).**
4. Calculate the mass of the solution in the flask **g = gc – gđ** (grams).
5. Repeat steps 1 to 3 with two other 100 mL volumetric flasks labeled A-1-2 and A-1-3

Determining the Density of Solution A-2:

Perform the same steps as for determining the density of solution A-1

Pre-lab Discussion

1. Why are measurements important in every aspect of life?

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2. What is the difference between accuracy and precision? Please give two examples to clarify your point.

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3. Why is the last digit in a measurement called the uncertain digit?

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Experiment 1

Name:……………………………………….. ………..Troy ID:…………….……………………..

Partners' name (if appropriate):……………………………………………………………………….

Date (of Lab Meeting):…………………………………………………………………… ………..

Instructor:

TITLE OF EXPERIMENT:…………………………………………………………………………………….

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RESULTS

Table 1. Data obtained for Compound A-1

Trial	Label of the flask	Volume (mL)	Weight (g)	Density (g/mL)
1				
2				
3				
			Average	

Table 2. Data obtained for Compound A-2

Trial	Label of the flask	Volume (mL)	Weight (g)	Density (g/mL)
1				
2				
3				
			Average	

ANALYSIS

1. Which compound might be the new food preservative? Why?

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2. Why might you obtain different data in three trials?

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CONCLUSION

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Experiment 2

Concentration: Classification of Alcoholic Beverages by Measuring Their Densities

INTRODUCTION

The concentration of a solution is the quantity of a solute in a particular quantity of solvent or solution. Knowing the concentration of solutes is important in controlling the stoichiometry of reactants for solution reactions. It is also important in dosage calculation in medical practice.

There are different ways to express the relative amounts of solute and solvent in a solution, including percent composition (by mass or by volume), molarity, molality, and mole fraction.

Percent composition by mass, so-called mass fraction, is calculated by the following equation:

$$Percent\ by\ mass = \frac{mass\ of\ solute}{mass\ of\ solution} \times 100$$

Percent composition by volume is calculated by the following equation:

$$Percent\ by\ volume = \frac{volume\ of\ solute}{volume\ of\ solution} \times 100$$

Molarity is calculated by the following equation:

$$Molarity = \frac{moles\ of\ solute}{volume\ of\ solution\ in\ liters}$$

Molality is calculated by the following equation:

$$Molality = \frac{moles\ of\ solute}{mass\ of\ solvent\ in\ kilograms}$$

The mole fraction of A, X_A, in a solution consisting of A, B, C, ... is calculated using the equation:

$$X_A = \frac{moles\ of\ A}{moles\ of\ A + moles\ of\ B + moles\ of\ C + \cdots}$$

In many cases, the density of solutions relates to their solute concentration. Hence, measuring the density of a solution is sometimes a quick and low-cost method to approximate its solute concentration.

OBJECTIVES :This experiment aims to determine the concentration of two alcohol solutions by their density.

CONCEPTS :This experiment uses the concepts of volume percent concentration, molarity, and density.

TECHNIQUES : Correct use of scales, volumetric flasks, and graduated cylinders are some of the techniques encountered in the experiment.

PROBLEM DESCRIPTION

A traditional way of determining the concentration of alcohol in alcoholic beverages is by measuring the density of the solution since these two characteristics of the solution (density and percent alcohol) are related. This is a convenient determination method because it is low cost and can be done quickly. This is why customs officials sometimes use density measurements to classify alcoholic beverages and spirits. In the stockroom, you will find a variety of solutions, including two alcoholic beverages. You will need to perform an experiment to classify the two alcoholic beverages into one of the following five categories of alcohol strength defined by Canadian importation laws:[2]

1. Less than 0.5 vol.%
2. 0.5 – 1.2 vol.%
3. 1.2 – 7.0 vol. %
4. 7.0 – 22.9 vol. %
5. More than 22.9 vol. %

PROCEDURE

1. Use three **25-mL** volumetric flasks and a scale to measure the density of distilled water, Alcoholic **Beverage A, 1M alcohol, and alcohol 70%**. Be sure to label the containers appropriately. Note that the weight of the volumetric flasks might be different.
2. Prepare five water-alcohol solutions as suggested in the following table.

Table 1. Alcohol-water solution preparation

Beaker label	20/100	40/100	60/100	80/100	100/100
Alcohol 70% (mL)	5	10	15	20	25
Water (mL)	25	25	25	25	25

3. Measure the density of each solution using a volumetric flasks and a scale.
4. Use a MS Excel worksheet to build a graph of density versus concentration.
5. Add a trendline on the graph, and be sure that you obtain a slope value of at least 4 significant figures.
6. Use the trendline equation to calculate the concentrations of the **1M Alcohol**, Alcoholic **Beverage A**

Pre-lab Discussion

1. Which formula will you use to calculate the volume concentration of the prepared solutions?

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Present the data that you have calculated in the following table:

Table 2. Volume percent concentration of the prepared solutions

Beaker label	20/100	40/100	60/100	80/100	100/100
Alcohol 70% (mL)	5	8	15	20	25
Water (mL)	25	25	25	25	25
Alcohol concentration (vol.%)					

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[2] The problem description was copied from the Virtual Lab module written by Mr. Tim Palucka and Dr. Jordi Cuadros, Instituto Químico de Sarriá, 2011. All else was written by the author of this book.

Experiment 2

Name:……………………………………….. ……….Troy ID:……………..…………………….

Partners' name (if appropriate):…………………………………………………………………….

Date (of Lab Meeting):…………………………………………………………………… ………..

Instructor:

TITLE OF EXPERIMENT:………………………………………………………………………………………….

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RESULTS

Table 3. The density dependence of the alcohol solution on the concentration

Beaker label	water	20/100	40/100	60/100	80/100	100/100	1M	A	70
Concentration (vol.%)	0								70
Density (g/mL)									

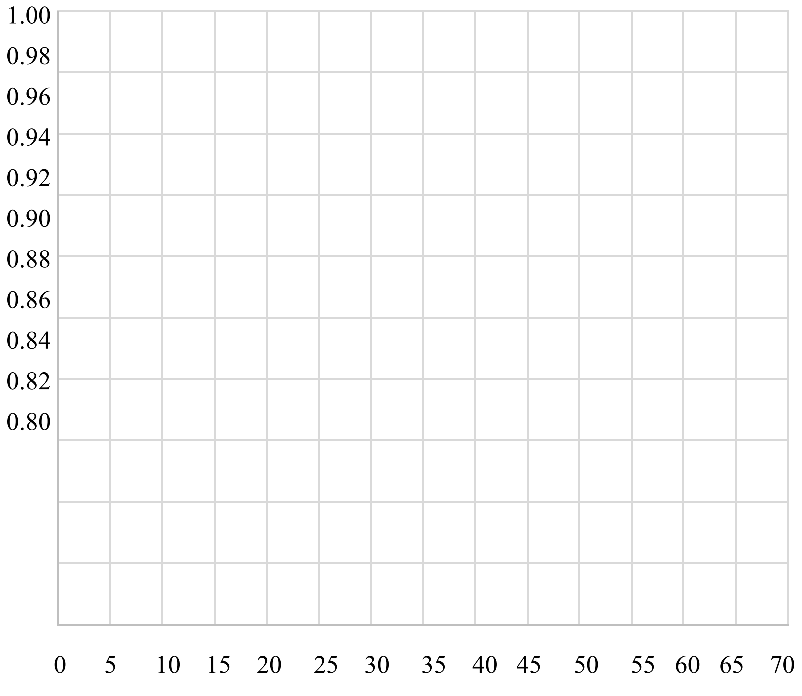

Figure 1. The density dependence of the alcohol solution on the concentration

ANALYSIS

1. How are densities and concentrations of the prepared alcohol solutions dependent on each other?

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2. What are the volume percent concentrations of the 1M Alcohol, Alcoholic Beverage A

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3. What categories of alcohol strength do the 1M Alcohol, Alcoholic Beverage A belong to?

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CONCLUSION

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Experiment 3:

Thermochemistry: Heat of Neutralization

In this experiment, you will use calorimetry to experimentally determine the heat of neutralization of NaOH-HCl, or the enthalpy of the acid-base reaction between NaOH and HCl. The calorimeter is a thermos container equipped by a thermometer to monitor temperature and magnetic stirrer to ensure the contents are at uniform temperature. The thermos is assumed to be a perfect insulator such that no heat comes in or out of it. When a reaction is allowed to occur in it, the amount of heat absorbed or released by the reaction can be determined by monitoring the temperature change. If the reaction releases heat (exothermic), the solution in the thermos absorbs the heat released, thus showing a temperature increase. On the other hand, the reaction may absorb heat (endothermic) from the solution, which in turn shows a temperature decrease.

When aqueous solutions of NaOH and are mixed in the calorimeter, the chemical equation for the reaction that occurs is:

$$\mathrm{NaOH\ (aq) + HCl\ (aq) \rightarrow H_2O\ (l) + NaCl\ (aq)}$$

Assuming NaOH and HCl fully react, the solution in the calorimeter is composed of salt and water. Because the thermos is assumed to be a perfect insulator, the amount of heat absorbed or released by the reaction (q_{rxn}) is equal in magnitude, but opposite in direction, to the amount of heat released or absorbed by the solution (q_{soln}), such that the q_{rxn} is just the negative of the q_{soln} as shown in Equation 1.

$$q_{rxn} = -q_{soln} \qquad \text{Equation 1}$$

The q_{soln} can be calculated from experimental data using Equation 2, where m_{soln}, $C_{s,soln}$, and ΔT_{soln} are the mass, specific heat and temperature change of the solution, respectively:

$$q_{soln} = m_{soln}C_{s,soln}\Delta T_{soln} \qquad \text{Equation 2}$$

As an example, let us calculate the enthalpy of the reaction below in kJ per mole of AgCl produced.

$$\mathrm{1\ AgNO_3\ (aq) + 1\ HCl\ (aq) \rightarrow 1\ AgCl\ (s) + 1\ HNO_3\ (aq)}$$

In a calorimetry experiment, when 25.0 mL of 0.100 M $AgNO_3$ and 25.0 mL of 0.100 M HCl are mixed in a thermos container, an increase in temperature from 23.42°C to 24.27 °C.

Since the solution warmed up, this means that the reaction released heat, and is thus exothermic. The q_{soln} is calculated from Equation 2, and q_{rxn} is the negative of q_{soln} by Equation 1. First, we will assume that volume is additive and that the density of the solution is the same as that of water at 1.00 g/mL, allowing us to calculate the m_{soln}. The $C_{s,soln}$ is also assumed to be the same as that of water which is 4.184 J/g. °C.

$$V_{soln} = 25.0\ \text{mL AgNO}_3\ \text{soln} + 25.0\ \text{mL HCl soln} = 50.0\ \text{mL solution}$$
$$m_{soln} = 50.0\ \text{mL}\ (1.00\ \text{g/mL}) = 50.0\ \text{g}$$

$$q_{soln} = (50.0\ \text{g})(4.184\ \text{J/g}\cdot{}^{\circ}\text{C})(24.27\ {}^{\circ}\text{C} - 23.42\ {}^{\circ}\text{C}) = 178\text{J}$$
$$q_{rxn} = -178\ \text{J}$$

The reaction released 178 J of heat. However, we need to calculate the heat of reaction in kJ per mole of AgCl produced. For the reaction that released 178 J of heat, the number of moles of $AgNO_3$ and HCl that reacted are calculated from the molarities and volumes of the solutions before they were mixed:

$$\text{moles AgNO}_3 = (0.100\text{ M})(0.0250\text{ L}) = 2.50 \times 10^{-3}\text{ moles AgNO}_3$$

$$\text{moles HCl} = (0.100\text{ M})(0.0250\text{ L}) = 2.50 \times 10^{-3}\text{ moles HCl}$$

Using the moles and molar relationships from the balanced chemical equation, the number of moles of AgCl can be calculated. But first, you must check if there is a limiting reagent in the reaction.

$$2.50 \times 10^{-3}\text{moles AgNO}_3 \left(\frac{1\, mol\, HCl}{1\, mol\, AgNO_3}\right) = 2.50 \times 10^{-3}\text{moles HCl}$$

This is equal to the amount of HCl used; therefore, there is no limiting reagent. This means that the number of moles of either $AgNO_3$ or HCl can be used to calculate the number of moles of AgCl.

$$2.50 \times 10^{-3}\text{moles AgNO}_3 \left(\frac{1\, mol\, AgCl}{1\, mol\, AgNO_3}\right) = 2.50 \times 10^{-3}\text{moles AgCl}$$

When 2.5×10^{-3} moles of both $AgNO_3$ and HCl react, 178 J of heat is released, which must be converted to kJ to calculate ΔH in kJ.

$$-178\text{ J}\left(\frac{1\, kJ}{1000\, J}\right) = -0.178\text{ kJ}$$

The enthalpy of reaction in terms of kJ per mole of AgCl produced can then be calculated as:

$$\Delta H = \frac{-0.178\text{ kJ}}{2.50 \times 10^{-3}\text{ mol AgCl}} = -71.2\ \frac{\text{kJ}}{\text{mol AgCl}}$$

PROCEDURE

1. Dry the thermos and the thermometer with a paper towel. Using a dry, clean graduated cylinder, measure **50.0 mL of 0.5 M NaOH** and pour it in the thermos. Put a dry, clean magnetic stirring bar into the thermos and put the thermos on top of a stirring plate.

> CAUTION: You are provided with a combination heat/stir plate but will be using only the stirring function. **Do not turn the heat on as this will ruin the thermos.**

2. Rinse the graduated cylinder and wipe dry the inside. Measure **50.0 mL of 0.5 M HCl** and pour it into a dry, clean beaker.

3. Measure the temperature of the acid. Rinse and wipe the thermometer, then measure the temperature of the base. The temperatures of the two solutions must be within 0.5°C of each other. If not, warm the HCl by holding the beaker in your hands or cool by immersing the beaker in tap water.

4. Record the temperature of the NaOH solution; this will be your initial temperature and enter it as the temperature at 0 s.

5. Insert the thermometer into the thermos lid, making sure it passes through the top side of the lid then through the bottom side.

6. Add the HCl solution from the beaker to the NaOH solution in the thermos quickly, but taking care not to splash any solution to the sides of the thermos. Stir gently, close the lid, and lower the

thermometer to the solution. Record the temperature 15 s after the addition of HCl. Continue recording the temperature every 15 s for 3 min.

The temperature rises at the beginning then stabilizes. Note the highest temperature that was achieved by the system; this will be the final temperature. If your readings show □T less than 4 □C, your thermometer is likely defective, and you need to repeat the experiment using another thermometer. Rinse and wipe the thermometer then turn it off when finished with the experiment.

7. Empty the thermos into the appropriate waste container and thoroughly rinse it and the stir bar.

8. Repeat Steps 1-7 two more times. Average the ΔT from all three trials.

CLEAN-UP

- Dispose of wastes in the container in the hood.
- Wash all glassware and the thermos container. Return all materials where they belong. Return the stir bar on top of the heat/stir plate.
- Make sure the thermometer is rinsed, wiped and turned off.

Experiment 3

Name:……………………………………….. ………..Troy ID:……………..…………………….

Partners' name (if appropriate):……………………………………………………………………….

Date (of Lab Meeting):…………………………………………………………………… ………..

Instructor:

TITLE OF EXPERIMENT:……………………………………………………………………………

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DATA

Trial 1 : T_{initial} (temperature at 0 seconds): ____________________

Time (s)	Temp (°C)	Time (s)	Temp (°C)
15	____________	105	____________
30	____________	120	____________
45	____________	135	____________
60	____________	150	____________
75	____________	165	____________
90	____________	180	____________

T_{final} (highest temp): ____________________ ΔT ($T_{\text{final}} - T_{\text{initial}}$): ____________________

Trial 2: T_{initial} (temperature at 0 seconds): ____________________

Time (s)	Temp (°C)	Time (s)	Temp (°C)
15	____________	105	____________
30	____________	120	____________
45	____________	135	____________
60	____________	150	____________
75	____________	165	____________
90	____________	180	____________

T_{final} (highest temp): ____________________ ΔT ($T_{\text{final}} - T_{\text{initial}}$): ____________________

Trial 3: T_{initial} (temperature at 0 seconds): ____________________

Time (s)	Temp (°C)	Time (s)	Temp (°C)
15	____________	105	____________
30	____________	120	____________
45	____________	135	____________
60	____________	150	____________
75	____________	165	____________
90	____________	180	____________

T_{final} (highest temp): ____________________ ΔT ($T_{final} - T_{initial}$): ____________________

Average ΔT: ___________________

The specific heat (C_s) of water is 4.184 J/g· °C

CALCULATIONS

Show clearly the complete calculations with correct number of significant figures and units.

$$\_\_\ NaOH\ (aq) + \_\_\ HCl\ (aq) \rightarrow \_\_\ H_2O\ (l) + \_\_\ NaCl\ (aq)$$

1. Calculate the mass of the solution in g.

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2. Calculate the q_{soln} in J using Equation 2. Use the average ΔT.

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3. Determine q_{rxn} in J using Equation 1.

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4. Calculate the number of moles of NaCl produced.

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5. Calculate the enthalpy of the reaction in kJ per mole NaCl produced.

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6. The combustion of methylhydrazine (CH_6N_2), a liquid rocket fuel produces N_2 (g), CO_2 (g), and H_2O (l):

   $$\_\_\ CH_6N_2(l) + \_\_\ O_2(g) \rightarrow \_\_\ N_2(g) + \_\_\ CO_2(g) + \_\_\ H_2O(l)$$

   When 4.00 g of methylhydrazine is combusted in a bomb calorimeter, the temperature of the calorimeter increases from 25.00 °C to 39.50°C. In a separate experiment, the heat capacity of the calorimeter is measured to be 7.794 kJ/ °C. Calculate the heat of reaction for the combustion of a mole of CH_6N_2.

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Experiment 4

Le Châtelier's Principle

Introduction

Le Châtelier's principle applies to a chemical system that is initially at equilibrium. W*hen a change is made to such a system, the equilibrium position will shift so as to counteract the change*. Suppose a solution contains the species chromate, CrO_4^{2-}, and dichromate, $Cr_2O_7^{2-}$, in equilibrium, with some concentration of reactants and some concentration of products present.

$$\underset{\text{yellow}}{2CrO_4^{2-}(aq)} + 2H_3O^+(aq) \rightleftharpoons \underset{\text{orange}}{Cr_2O_7^{2-}(aq)} + 3H_2O(l)(aq) \qquad \text{Eq (1)}$$

To this solution is added a few drops of a concentrated solution of $Na_2CrO_4(aq)$. The task is to predict whether adding the $Na_2CrO_4(aq)$ to the mixture will result in the reaction shifting to the left, to the right, or staying the same (that is, to predict how the equilibrium position will shift). To understand how $Na_2CrO_4(aq)$ is related to this equilibrium requires recognizing that sodium salts **dissociate**: $Na_2CrO_4(aq) \rightarrow 2Na^+(aq) + CrO_4^{2-}(aq)$. So, adding $Na_2CrO_4(aq)$ increases the concentration of CrO_4^{2-} in the solution. According to Le Châtelier's principle, the reaction will shift to counteract what was done to it, so the reaction will shift to the right, which will lower the concentration of CrO_4^{2-}.

It is not easy to predict the color change in this case. Adding yellow $Na_2CrO_4(aq)$ to the beaker would, of course, make the solution more yellow. But, the shift in the reaction to the right would make the solution more orange.

Suppose that $AgNO_3(aq)$ was added to the original equilibrium mixture. Neither Ag^+ nor NO_3^- is directly involved in the equilibrium. However, in this case the silver ion will react with the chromate ion (CrO_4^{2-}): $2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)$. So, adding $AgNO_3(aq)$ to the equilibrium mixture will precipitate the chromate ion, making its concentration in solution less. The reaction will shift to counteract that change, which corresponds to shifting to the left. Shifting to the left corresponds to a decrease in the concentration of $Cr_2O_7^{2-}$, so less orange would be present after adding $AgNO_3(aq)$.

Suppose that a few drops of concentrated phosphoric acid, $H_3PO_4(aq)$, were added to the original equilibrium mixture. The problem is to figure out what H_3PO_4 has to do with the equilibrium being studied. There is no phosphate in that equilibrium. However, in solution H_3PO_4 dissociates: $H_3PO_4 \rightarrow H^+(aq) + H_2PO_4^-(aq)$. And, $H^+(aq)$ is the same thing as $H_3O^+(aq)$, so adding H_3PO_4 to a solution increases the concentration of hydronium ion, H_3O^+. Consequently, the reaction would be expected to shift to the right, consuming some of the hydronium ion that was added; the solution would become less yellow, since CrO_4^{2-} decreases on shifting to the right, and more orange, since $Cr_2O_7^{2-}$ increases.

Suppose a few drops of concentrated NaOH(aq) are added to the equilibrium. The equilibrium does not show OH^- ion. However, hydroxide reacts with hydronium ion (that is, base neutralized acid): $OH^-(aq) + H_3O^+(aq) \rightarrow 2H_2O(l)$. So, adding hydroxide ion will decrease the concentration of hydronium ion, H_3O^+. The reaction will shift to counteract what was done to it; that is, it will shift to the left, becoming more yellow and less orange.

For some equilibria, all of the components may be colorless, so a shift in the position of the reaction cannot be seen. However, if the reaction is an acid-base reaction, the shift may be made visible by adding an acid-base indicator. Suppose an indicator is used that is green in acid and yellow in base (this is **not** the indicator used in this lab). Also, suppose the following equilibrium is being examined with that indicator:

$$\underset{\text{yellower}}{H_3PO_4(aq) + H_2O(aq)} \rightleftharpoons \underset{\text{greener}}{H_3O^+(aq) + H_2PO_4^-(aq)} \qquad \text{Eq (2)}$$

This reaction produces hydronium ion, so if the reaction shifts to the right, the solution will become more acidic and greener. If it shifts to the left, some of the hydronium ion would be consumed, making the solution more basic, and so yellower.

Would this reaction mixture get yellower or greener if a few drops of concentrated $H_3PO_4(aq)$ were added to it? Well, the $H_3PO_4(aq)$ is on the reactants side, so the equilibrium would shift to the right to consume some of the added material. Shifting to the right produces more $H_3O^+(aq)$, making the solution more acidic, so the indicator would become greener.

Le Châtelier's principle also applies to chemical reactions involving heat. Heat can be added or removed from a reaction by heating or cooling the reaction mixture. If heat is added to the mixture by raising the temperature, the equilibrium will shift to produce less heat; if heat is removed from the mixture by lowering the temperature, the reaction will shift to produce more heat. This will be used to determine if a reaction is exothermic or endothermic. To do this, heat is included in the previous reaction. It is placed on one side of the reaction, then the evidence is examined to see if that is the correct side. In the following reaction, it is placed on the reactants side (that corresponds to the reaction being endothermic).

$$\text{heat} + H_3PO_4(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + HPO_4^-(aq) \qquad \text{Eq (3)}$$

yellower greener

Suppose that on heating this system the indicator became more yellow. Adding more heat should have shifted the reaction to the right to consume some of the added heat, resulting in the solution becoming greener, not yellower. Therefore, the heat was placed on the wrong side: heat must be a product, not a reactant, and the reaction must be exothermic. The correct equation is the following:

$$H_3PO_4(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + H_2PO_4^-(aq) + \text{heat} \qquad \text{Eq (4)}$$

yellower greener

Procedure

Caution: As always, wear safety glasses while performing this experiment. Concentrated acid, which causes burns, will be used.

Preparation

Hot and cold water baths:

Start heating a 100 mL beaker about half-full of tap water to near-boiling on a hot plate. (Do not boil the water so that it won't splatter on you.)

Fill another 100 mL beaker with about half full with ice. Add water to make a slush (having water present improves cooling, because an object will have a greater contact area with a liquid than with ice).

Label four 15 × 125 mm test tubes as 1, 2. Place the tubes in a test tube rack.

Equilibrium of a Weak Acid, HSO_4^-

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq) \qquad \text{Eq (5)}$$

Eq (5) is the net ionic equation for the next equilibrium to be examined. The acid-base indicator used in this part is thymol blue, which becomes yellow-orange as a solution becomes more basic, and red as a solution becomes more acidic. Copy the above reaction to the data sheet. Below that reaction **write the indicator color change** expected after the reaction has shifted to the left and to the right, as in $H_3PO_4(aq) + H_2O(aq) \rightleftharpoons H_3O+(aq) + H_2PO_4^-(aq)$ Eq (2).

1. **Prepare storage solutions.** Transfer about *5 mL* of the provided $NaHSO_4$ solution to *the test tube labeled 1.* Place a disposable plastic transfer pipette into this tube.
2. **Add indicator.** Add 1-2 drops of the thymol blue indicator from the dropper bottle to the test tube cotained $NaHSO_4$. Swirl the mixture in the test tubes. Record the color of these solutions, which are at equilibrium.

3. **Transfer the mixture to the 24-well plate.** Using the disposable pipette, transfer about 10 drops of the mixture into wells *A1, A2, A3,* and the other *test tube labeled 2*.

Figure 1. 24-well plate. Notice the letters along the side and numbers along the top, which allow each well to be identified. For example, the bottom right well is D6. Picture by Joseph Elsbernd, on Flickr; creative commons SA-BY

4. **Add Na_2SO_4 to a well.** With the pipette, add one drop of the Na_2SO_4 solution to well A1. Mix the solution with a toothpick. Probably not much change in color is observed, so add a drop more with mixing until a color change is observed. Record the color change.
   For example, if the indicator had originally been pink, the color change might be "less pink" or "more pink". If the color had been purple, the color change might be "redder" or "bluer", or "more blue" or "less red".

5. **Analyze.** Based on the color change, record on the data sheet the direction the reaction shifted on adding Na_2SO_4 solution to well A1. Explain the direction of the shift on the data sheet using Le Châtelier's principle.

   Here is a template for creating the explanation. "Adding/removing ________________ increases/decreases the concentration of ____________. To counteract this change, the reaction will shift to the left/right, which will increase/decrease the concentration of _____________, resulting in the indicator color becoming __________________________."

   "First, what was done: adding/removing ____________ (the reagent) increases/decreases the concentration of __________________ (one of the species involved in the reaction)
   Second, how the system responded: The system will counteract this disturbance by shifting to the ________ (left or right); as a result, the concentration of (one of the species involved in the reaction) will decrease/increase, resulting in the indicator color becoming __________________________."

6. **Add $NaHSO_4$ to a well.** Since $NaHSO_4$ is already in the well, to increase its concentration, solid $NaHSO_4$ will be added. With a microspatula, add a crystal of $NaHSO_4$ to well A3 (well A2 is the color with nothing added, for comparison). Mix the solution with a toothpick. Add crystals with mixing until a color change is observed. Record the color change.
7. **Analyze.** Based on the color change, record on the data sheet the direction the reaction shifted on adding $NaHSO_4$ solution to well A3. As before, explain the direction of the shift on the data sheet using Le Châtelier's principle.
8. **Clean the pipets.** Fill a beaker with tap water, and use it to rinse out the two pipets a couple of times, then rinse them with deionized water. Get as much of the liquid out of the pipets as possible.
9. **Place a mixture in the hot-water bath.** Place the mixture in the test tube labeled 1 into the hot-water bath.
10. **Place a mixture in the cold-water bath.** Place the mixture in the test tube labeled 2 into the cold-water bath.
11. **Equilibrate.** Swirl the mixtures in the test tubes for **a minute or two** to speed up the heat transfer. When the color has largely stopped changing, record the changes in color of the two solutions.

12. **Reverse the process.** Exchange the test tubes in the two baths to see whether the color changes are permanent, or whether the process can be reversed.
13. **Analyze.** From the color change, determine whether the reaction shifted to the left or to the right on heating. Based on the direction the reaction shifted on heating, determine whether heat is a reactant or product in this reaction. Rewrite the net ionic equation, including heat in the reaction, as in Eq (3) and Eq (4). State whether this is an exothermic or endothermic reaction.
14. **Clean up.** Take the 24-well plate and a bottle of deionized water over to the hood containing the waste containers. Dump the contents of the plate into the "Discarded HSO_4- Reaction Mixtures" container. Rinse the plates with the deionized water, letting the water run into the waste container. Be careful to completely rinse out any crystals of $NaHSO_4$ that may remain. Dry the plate with a "Lab Wipe" tissue paper.

    Empty the two test tubes into the waste container, and rinse with deionized water from a squirt bottle.

Equilibrium of a Slightly Soluble Salt, $Mg(OH)_2$

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^{-}(aq) \qquad \text{Eq (6)}$$

Eq (6) is the net ionic equation for this equilibrium.

1. **Prepare storage solutions.** Transfer about *10 mL* of the provided *$MgCl_2$* solution to *the test tube labeled "1"*. Place the disposable plastic transfer pipet into the test tube.
2. **Transfer $MgCl_2$ to the 24-well plate.** Transfer 20 drops of the $MgCl_2$ solution into wells B1, B2 and *the test tube labeled "2"*. Place a toothpick in each well for future use.

Caution: As always, wear safety glasses while performing this experiment. Sodium hydroxide, NaOH is very hazardous. If you get NaOH on your skin, rinse with water until it stops feeling soapy. Likewise, HCl can cause burns. If you get HCl on your skin, rinse it off with water.

3. **Transfer NaOH to the 24-well plate.** Transfer 10 drops of the NaOH solution into wells B1, B2 and the *the test tube labeled "1" and "2"*. Mix with the toothpicks. Write a complete chemical reaction for what was observed in cells B1 and B2.
4. **Net ionic equation for the equilibrium.** In the data sheet, record the net ionic equation, Eq (6), for the equilibrium involving $Mg(OH)_2$.
5. **Add concentrated HCl to a well.** The concentrated HCl is stored in the hood. Take the plate to the hood and add 1 drop of concentrated HCl to well B1. Mix the solution with a toothpick. If not much change is observed, add a drop more with mixing until a change is observed. Describe this change in the data sheet.
6. **Analyze.** State on the data sheet which component of the equilibrium reaction has its concentration directly changed by adding HCl(aq). Based on the observed change, record on the data sheet the direction the reaction shifted on adding conc. HCl solution to well B1. Explain the direction of the shift on the data sheet using Le Châtelier's principle.
7. **Add Na_4EDTA to a well.** Add 1 drop of Na_4EDTA to well B2. Mix the solution with a toothpick. Probably not much change will be observed, so continue adding drops, mixing after each drop, until a distinct change is observed (this reaction may take a couple of minutes of stirring). Record the change on the data sheet. EDTA stands for ethylene diamine tetraacetic acid, Fig. 7. It is available as a sodium salt, Na_4EDTA. EDTA wraps around metal ions, like Ca^{2+}, Mg^{2+}, and Co^{2+}, as shown in the structure on the right. Although the ion stays in solution, the EDTA prevents it from reacting with OH_{-}, so in effect, EDTA lowers the concentration of Mg^{2+} ion.

8. **Analyze.** State on the data sheet which component of the equilibrium reaction has its concentration changed by adding $Na_4EDTA(aq)$ to the mixture. Based on the observed change, record on the data sheet the direction the reaction shifted on adding Na_4EDTA. Explain the direction of the shift on the data sheet using Le Châtelier's principle.
9. **Add indicator.** Add 1 drop of phenolphthalein indicator from the dropper bottle to the solutions in *the test tube labeled "1" and "2"*. Swirl the mixtures in the test tubes. Record the color if this indicator solution, which is at room temperature, on the data sheet.

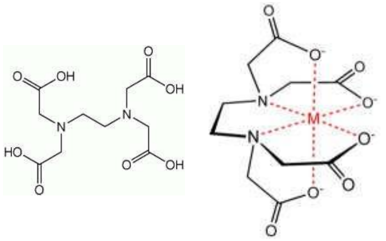

Figure Structure of EDTA (left), and a metal atom surrounded by an EDTA ion (right).

   The acid-base indicator used in this part, phenolphthalein, becomes pink as a solution becomes more basic, and colorless as a solution becomes more acidic.

   Also, below the net ionic equation for this equilibrium (step 4) write the indicator color change expected after the reaction has shifted to the left and to the right, as in $H_3PO_4(aq) + H_2O(aq) \rightleftharpoons H_3O+(aq) + H_2PO_4–(aq)$ Eq (2).
10. **Clean the pipets.** Fill a beaker with tap water, and use it to rinse out the two pipets a couple of times, then rinse them with deionized water. The rinse water can be placed in the waste container in the hood that is labelled "Discarded $Mg(OH)_2$ Reaction Mixtures". Get as much of the liquid out of the pipets as possible.
11. **Place a mixture in the hot-water bath.** Place the test tube labeled 1 into the hot-water bath.
12. **Place a mixture in the cold-water bath.** Place the test tube labeled 2 into the cold-water bath.
13. **Equilibrate.** Swirl the mixtures in the test tubes occasionally to speed up the heat transfer. After three or four minutes, record the changes in color of the two solutions.
14. **Reverse the process.** Exchange the pipets in the two baths to see whether the color changes are permanent, or whether the process can be reversed.
15. **Analyze.** From the color change, determine whether the reaction shifted to the left or to the right on heating. Based on the direction the reaction shifted on heating, determine whether heat is a reactant or product in this reaction. Rewrite the net ionic equation, including heat in the reaction, as in Eq (3) and Eq (4). State whether this is an exothermic or endothermic reaction.
16. **Clean up.** Take the 24-well plate and a bottle of deionized water over to the hood containing the waste containers. Dump the contents of the plate into the "Discarded $Mg(OH)_2$ Reaction Mixtures" container. Rinse the plates with the deionized water, letting the water run into the waste container. Be careful to completely rinse out any solid $Mg(OH)_2$ that may remain. Dry the plate with a "Lab Wipe" tissue paper.

   Empty the two test tubes into the waste container, and rinse with deionized water from a squirt bottle.

   Empty the two pipets into the waste container, and rinse them two or three times with water, and put them in the trash.

   Throw the toothpicks into the "Discarded Toothpicks" container.

Experiment 4

Name:………………………………………….. ………..Troy ID:……………..……………………..

Partners' name (if appropriate):………………………………………………………………………

Date (of Lab Meeting):…………………………………………………………………… ………..

Instructor:

TITLE OF EXPERIMENT:……………………………………………………………………………………

………

Data Sheet

Equilibrium of a Weak Acid, HSO_4^-

Net ionic equation for this equilibrium, with change in indicator colors: ………………………………...

……………………………………………………………………………………………………..

……………………………………………………………………………………………………..

Color of the initial solution after adding the indicator: ____

The solution should be more orange if the equilibrium shifts to the _______(left or right).

The solution should be redder if the equilibrium shifts to the _______ (left or right).

Adding Na_2SO_4

Color change after adding Na_2SO_4 solution: ____

Based on the color change, the reaction shifted to the (circle one): left right Explain the direction of the shift using Le Châtelier's principle:……………………………………………………………..

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

Adding $NaHSO_4$

Color change after adding $NaHSO_4$ solution: ____

Based on the color change, the reaction shifted to the (circle one): left right Explain the direction of the shift using Le Châtelier's principle:……………………………………………………………...

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

Changing the Temperature

Color change on heating: ____

Color change on cooling: ____

On heating, the reaction shifted to the (circle one): left right

Net ionic equation for this equilibrium, with change in indicator colors, and with heat as a reactant or product: ……………………………………………………………………………………………

………

………

According to the reaction above, this process is (circle one): exothermic endothermic

Equilibrium of a Slightly Soluble Salt, $Mg(OH)_2$

Adding HCl

Chemical reaction for formation of a precipitate on adding NaOH(aq) to $MgCl_2$(aq). …………………

………

………

Net ionic equation for the $Mg(OH)_2$ equilibrium. …………………………………………………

………

………

………

Change on adding conc. HCl to a mixture containing $Mg(OH)_2$(s): ________________________

Component of the equilibrium reaction whose concentration is most directly changed by adding HCl to the mixture: __________________________________

Direction the reaction shifted on adding HCl (circle one): left right Explain the direction of the shift

using Le Châtelier's principle:………………………………………………………………………

………

………

………

………

Adding Na_4EDTA

Change on adding Na_4EDTA solution: ____

Component of the equilibrium reaction whose concentration is directly changed by adding Na_4EDTA to the mixture: __________________________________

Direction the reaction shifted on adding Na_4EDTA (circle one): left right Explain the direction of the

shift using Le Châtelier's principle: ………………………………………………………………

………

………..

Changing the Temperature

Color of indicator in the room temperature solution: ___________________________________

Color change on heating: ____

Color change on cooling: ____

On heating, the reaction shifted to the (circle one): left   right

Net ionic equation for this equilibrium, with change in indicator colors, and with heat as a reactant or product: ……………………………………………………………………………………………………

………

………

………

According to the reaction above, this process is (circle one): exothermic   endothermic