FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment: 1

Density of liquids

INTRODUCTION

The density, mass per unit volume, of liquids can be determined like that of solids by measuring both the mass and the volume of a given sample. The density is a characteristic property of a substance; it remains fixed unless the temperature or pressure is changed. For liquids, a relatively small change in temperature can affect the density appreciably, but a pressure change must be quite great to have a measurable effect. As an intensive property the density is independent of the quantity of material measured.

In this experiment, you will practice using a balance to measure mass. In addition, you will learn how to measure volume using a graduated cylinder and then you will determine the density of a) liquid water and b) a salt solution with different concentrations. Once the density of liquid water and its change with temperature are known, you can use the information to find out what volume should be occupied by a known mass of water at a given temperature. Once a data such as concentration vs. density of the salt solution is obtained, it is possible to plot a curve what we called a **calibration curve**. When such a plot is obtained we can determine the density of a substance with a known concentration, or the concentration of a substance with a known density.

PROCEDURE

A)

- Weigh an empty graduated cylinder
- Add about 20 mL of distilled water in this graduated cylinder.
- Record the volume to the nearest 0.1 mL.
- Weigh the graduated cylinder plus water on an analytical balance and record.
- Measure the temperature of the water
- Calculate the density.
- Compare the calculated density of the water with the values given in the Table 1. and calculate the percent error.

B)

- Rinse your graduated cylinder with about 5 mL of the 4 % NaCl solution, and then pour about 20 mL of the 4 % NaCl solution into the graduated cylinder.

- Determine the density of this solution as in Part A.

- Determine the density of 8, 12, and 16 % NaCl solutions similarly.

- Obtain a salt solution of unknown concentration of NaCl from your instructor and determine its density.

- Plot a graph showing the density (on the vertical axis) vs. concentration of NaCl (horizontal axis).

- Using the graph, obtain a value for the concentration of NaCl for your unknown and report it to your instructor.

Table 1.

Temp°C	Density(g/mL)	Temp°C	Density (g/mL)
15	0.9979	26	0.9959
17	0.9977	28	0.9955
18	0.9975	29	0.9952
20	0.9972	31	0.9946
22	0.9968	33	0.9941
24	0.9964	35	0.9935

PRELAB QUESTIONS

1. Write the definition of the density and

a) explain the effect of the temperature and pressure on the density of the liquids.How does the density of liquids changes with the temperature?

b) why do we need a great pressure change to observe a change in the density of the liquids?

c) if we have a gas sample instead of liquids, do we need great pressure change to have a measurable change in the density? Explain briefly.

d) if we have a solid sample instead of liquid or gas, does the density of this solid change if we increase or decrease pressure too much?

2.

a) Suppose that, you measure the temperature of 20 mL of the water as 30°C when it is actually 20°C. Find the minimum percent error in the calculation of the density by using the table given above.

b) Suppose that you determine the mass of 41.2 mL of water as 41.052 g. at 28°C. Calculate the density from the experimental values and match it with the value given in the table. Calculate the percent error.

3. Plot the temperature vs. density values given in the table. Determine the density of the liquid water at 19°C, 27°C, 30°C and 32°C.

4. What is the calibration? Explain briefly.

FE 111 GENERAL CHEMISTRY DATA SHEET

Density of liquids

Experiment : 1 ***Date:***

a)

Mass of empty graduated cylinder : ………………..

Volume of water in the cylinder : ………………..

Mass of graduated cylinder plus water : ………………..

Temperature of water : ………………..

b)

NaCl solution	**4 %**	**8 %**	**12 %**	**16 %**
Mass of grd.cyl.+ known NaCl soln.	………..	………..	………..	………..
Volume of NaCl in cylinder	………..	………..	………..	…………

Mass of grd. cyl+ unknown soln : ………..

Volume of unknown soln : ………..

Name of the student: **Submitted to :**

FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment: 2

Heat and Temperature

INTRODUCTION

Heat differs from temperature in that **heat** is a quantity of energy whereas **temperature** is a measure of the hotness or coldness of an object. Indeed, temperature is a physical property that determines the direction of heat flow. Heat always transfers from hot objects to cold ones. For example, if you touch an ice cube, heat flows out of your hand to the ice, cooling your hand. The **specific heat capacity** is the amount of heat required to raise the temperature of 1g or 1mol of substance by 1^{o}C.

In this experiment you will investigate

a) the calibration of a thermometer as a temperature-measuring device
b) the determination of the specific heat capacity of a given metal

Table 2-1 Boiling point of water at various pressures

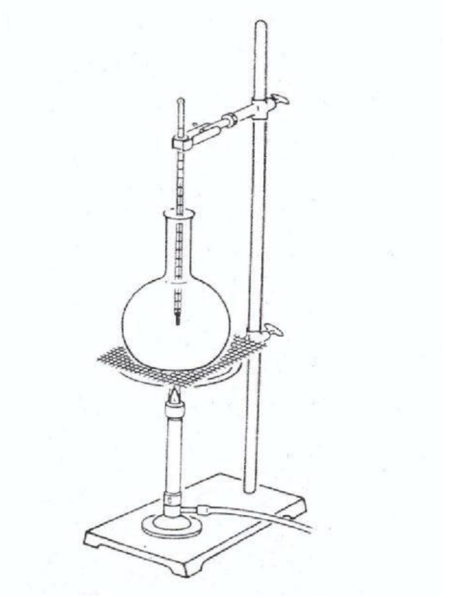
Pressure		
mm-Hg	atm	Bp, ^{o}C
700	0.921	97.7
705	0.928	97.9
710	0.934	98.1
715	0.941	98.3
720	0.947	98.5
725	0.954	98.7
730	0.961	98.9
735	0.967	99.1
740	0.974	99.3
745	0.980	99.4
750	0.987	99.6
755	0.993	99.8
760	1.000	100.0
765	1.007	100.2
770	1.013	100.4

The thermometer can be calibrated by immersing it first in melting ice (0^{o}C) and than in boiling water (100^{o}C at 1 atm pressure). Since the boiling point of water varies with changing barometric pressure. Table 2-1 will enable you to find the true boiling point of water under the conditions of the experiment. To determine the specific heat capacity of a metal, you will heat a weighed amount of metal to a known temperature (that of boiling water), and then place the hot metal in a known amount of water. From the temperature rise of the water, you can calculate the amount of heat transferred from the metal to the water. This can be done directly because it takes 4.2 joules (or 1 calorie) to raise 1 g of water by 1^{o}C.
Since the density of water is almost exactly 1 g per mL. 1 g of water is very nearly 1 mL. The specific heat capacity of the metal is equal to the amount of heat liberated by the the metal divided by its temperature drop times its mass.

PROCEDURE

a) Wash off some crushed ice, and place it in a small beaker. Add distilled water until the ice is nearly covered. Immerse your thermometer in the ice-water mixture as deep as possible; stir gently; record the thermometer reading when it has become constant.

Place about 100 mL of distilled water in your 500 mL flask. Drop in a boiling chip to prevent bumping during boiling. Insert your thermometer into the water. Heat the water to boiling and record the thermometer reading when it has become constant. Record the barometric pressure.



b) Obtain a piece of metal (about 10 g) for which the specific heat is to be found. Weigh it to the nearest 0.1 g. Tie it with a string and suspend it above boiling water in a flask as the thermometer is set in Figure. Do not let the metal get wet. (If water condenses on the metal, pull the metal out, dry, and replace.) Allow sufficient time for the metal to reach the temperature of the vapor from boiling water. Record the barometric pressure. With a graduated cylinder, measure 10.0 mL of distilled water into a test tube. Prop the test tube in the mouth of your 300 mL flask. Measure the temperature of the water to the nearest 0.1^{o}. Withdraw the thermometer, and hold it directly over the test tube so that any drops of water will drain back. Quickly remove the heated metal, and carefully lower it into the water in the test tube. Stir by raising and lowering the metal. Return the thermometer, and note the maximum temperature registered.

PRELAB QUESTIONS

1- Define

Boiling point, Normal Boiling Point, Normal Freezing point 2-
What is the difference between heat capacity and specific heat
capacity ?

FE 111 GENERAL CHEMISTRY DATA SHEET

Heat and Temperature

Experiment : 2 *Date:*

A)

Thermometer reading in ice-water mixture..................

Thermometer reading at water boiling point..................

Barometric pressure..................

B)

Weight of metal …………….

Temperature of the heated metal..................

Temperature of the water in the test tube..................

Temperature of the water and metal mixture..................

Name of the student: **Submitted to :**

FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment 3

Charles's Law

INTRODUCTION

Charles's law gives the relation between the volume and temperature of a gas at constant pressure. It states that the volume of a given mass of gas is directly proportional to the absolute temperature when pressure is constant. Mathematically we may state this in either of two ways:

$V = kT$ (P constant) or $V_1/T_1 = V_2/T_2$ (P constant)

In this experiment the student will heat the air inside a flask to the boiling temperature of water. Then he will cool the flask by placing it in a cold water bath. As the air inside the flask cools, its volume will decrease, and water will be drawn into the flask. The volume of this water is the difference between the initial and final air volumes.

One correction must be made. The water drawn into the flask has a vapor pressure (about 20 mm). Therefore, the air in the flask at the end of the experiment will be mixed with water vapor. The pressure of air:

$$P_{air} = P_{atm} -- P_{H_2O}$$

Now we can use Boyle's law to calculate the volume the air would occupy if $P_{air} = P_{atm}$. Let's call this volume V_2. Then

$$P_{atm} V_2 = (P_{atm} - P_{H_2O})(V_1 - V_{H_2O})$$

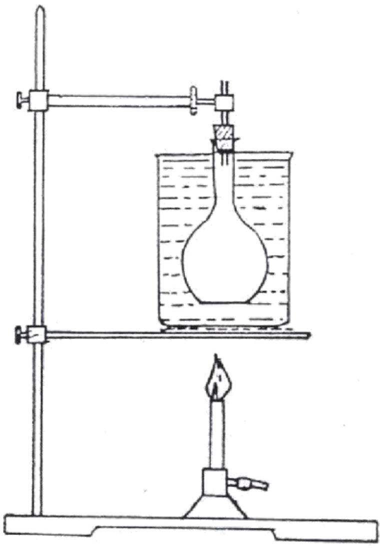
where V_{H_2O} is the volume of water drawn into the flask and V_1 is the volume of the flask.

Now it should be possible to compare V_1/T_1 (the volume and temperature of the air at the boiling temperature of water) with V_2/T_2 (where T_2 is room temperature). Since in both cases $P = P_{atm}$, if Charles's law is obeyed, we will find that

$$V_1/T_1 = V_2/T_2$$

PROCEDURE

Take a clean 250 mL flask and make sure it contains no moisture by passing it back and forth through a (**luminous)** bunsen burner flame, until it is completely dry. (You have to hold the flask in your **hand** and heat the area that contains moisture). Take a one-holed rubber stopper that fits the flask and insert a 7-8 cm length of fire-polished glass tubing so that one end is even with the bottom of the stopper. Fit the stopper **firmly** into the flask and mark with a pencil where it comes in contact with the flask. Place the flask in a 1000 mL beaker and set up apparatus as shown below. The flask should be almost completely immersed in water.



Gently boil the water for about 10 minutes. Record the temperature of the boiling water (T_1) and the atmospheric pressure (P_{atm}). Turn the burner off and cover the open end of the glass tube with your finger. Loosen the clamp and, keeping your finger over the glass tube, turn the flask upside down and immerse it in a room temperature water bath. Remove your finger and keep the flask completely submerged for 10 minutes. At the end of this time, raise the inverted flask until the water levels inside and outside are the same. Cover the end of the glass tube with your finger, remove the flask from the bath, and restore it to an upright position. Record the temperature of the water bath (T_2). Pour the water from the flask into a graduated cylinder and record its volume (V_{H_2O}).

Fill the flask with water and place the rubber stopper in its original position (using the pencil mark). The glass tube should be completely filled with water, and no air bubbles should be trapped around the bottom of the stopper. Dry the outside of the stopper, then remove the stopper and carefully drain the contents of the glass tube into the flask. Pour the water from the flask into a graduated cylinder and record its volume (V_1).

From the appendix of your textbook obtain the vapor pressure of water (P_{H_2O}) at room temperature T_2.

CALCULATIONS

Remember to convert T_1 and T_2 to the absolute temperature scale. Now calculate V_2, as shown in the introductory discussion. Calculate and compare V_1/T_1 and V_2/T_2.

QUESTIONS

1- Assuming your values of V_1, T_1, and T_2 were correct, what is the theoretical value of V_2 ? (Calculate from Charles's law). Calculate your percent error.

$$\% \text{ error} = \frac{\text{difference between } V_2 \text{ (theoretically) and } V_2 \text{ (experimentally)}}{V_2 \text{ (theoretically)}} \times 100$$

2- How would the experimental results be affected if the flask contained moisture at the beginning of the experiment ?

3- Why was it necessary to match the water levels inside and outside the cooled flask before removing it from the water bath ?

FE 111 GENERAL CHEMISTRY DATA SHEET

Charles's Law

Experiment : 3 *Date:*

Temperature of the Boiling water (T_1) : ....

Atmospheric Pressure (P_{atm}) : ....

Temperature of Water Bath (T_2) : ....

Volume of Water in the Flask (V_{H_2O}) : ....

Vapor Pressure of Water at T_2 (P_{H_2O}) : ....

Volume of the Flask (V_1) : ....

Name of the student: **Submitted to :**

FE 111 GENERAL CHEMISTRY LAB SHEET

Experimet: 4

Chemical Equilibrium

INTRODUCTION

Chemical reactions occur so as to approach a state of chemical equilibrium. The equilibrium state can be characterized by specifying its equilibrium constant, i.e., by indicating the numerical value of the **equilibrium constant expression**. In this experiment you will determine the value of the equilibrium constant for the reaction between ferric ion, Fe^{3+} and isothiocyanate ion, SCN^-.

$$Fe^{3+} + SCN^- =\!= FeSCN^{2+}$$

for which the equilibrium condition is

$$\frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]} = K$$

To find the value of K, it is necessary to determine the concentration of each of the species Fe^{3+}, SCN^-, and $FeSCN^{2+}$ in the system at equilibrium. This will be done colorimetrically, taking advantage of the fact that $FeSCN^{2+}$ is the only highly colored species in the solution.

The color intensity of a solution depends on the concentration of the colored species and on the depth of solution viewed. Thus 2 cm of a 0.1 M solution of a colored species appears to have the same color intensity as 1 cm of a 0.2 M solution. Consequently, if the depths of two solutions of unequal concentrations are chosen so that the solutions appear equally colored, then the ratio of concentrations is simply the inverse of the ratio of the two depths ($h_1 M_1 = h_2 M_2$). It should be noted that this procedure permits only a comparison between concentrations. It does not give an absolute value of either one of the concentrations. To know absolute values it is necessary to compare with a standard of known concentration.

For color determination of $FeSCN^{2+}$ concentration, you must have a standard solution in which the concentration of $FeSCN^{2+}$ is known. Such a solution can be prepared by starting with a small known concentration of SCN^- and adding such a large excess of Fe^{3+} that essentially all the SCN^- is converted to $FeSCN^{2+}$. Then, the concentration of $FeSCN^{2+}$ may be calculated from the equation:

$$[FeSCN]^{2+} = \frac{C_{NaSCN} x V_{NaSCN} x \frac{1\,mol\,FeSCN^{2+}}{1\,mol\,NaSCN}}{V_{total}}$$

PROCEDURE

- Obtain five clean test tubes. Rinse with distilled water and let drain.
- Give numbers to each test tube.
- Add 5 mL of 0.0020 M NaSCN solution to each.
- To the first test tube, which will serve as your standard, add 5 mL of 0.20 M $Fe(NO_3)_3$.
- For the remaining four tubes proceed as described below.
- Add 10 mL of 0.20 M $Fe(NO_3)_3$ to your graduated cylinder.
- Add 15 mL distilled water to make the volume exactly 25 mL. Stir this solution thoroughly.
- Pour 5 mL of the this solution into the second test tube.
- Pour out the half of the remaining 20 mL of the solution in graduated cylinder so that the volume will be exactly 10 mL.
- Make the volume again 25 mL with distilled water and mix thoroughly.
- Pour 5 mL of this solution into the third test tube.
- Pour out half of the solution in the graduated cylinder until exactly 10 mL remain.
- Add 15 mL of distilled water and mix thoroughly.
- Pour 5 mL of this solution into the fourth test tube.
- Again pour out half of the solution in the graduated cylinder until exactly 10 mL solution remain.
- Add 15 mL of distilled water and mix thoroughly.
- Pour 5 mL of this solution into the fifth test tube.

The next step is to determine the relative $FeSCN^{2+}$ concentration in each test tube. To do this you will compare the color intensity in test tube 1 with that in each of other four test tubes.

- Take test tube 1 and test tube 2.
- Hold them side by side.
- Wrap a piece of white paper around them.

- Look down through the solution towards the white background that your table makes.

- If color intensities appear identical record this fact.

- If not, pour out some of the standard in the test tube 1 until the colour intensities are the same.

- Measure the heights of solutions in the two tubes being compared.

- Repeat this comparison for the remaining three test tubes and record the heihts of both the standard solution and solution being compared.

RESULTS AND CALCULATIONS

1- Calculate the concentration of Fe^{3+} and SCN^{-} in each of the five tubes, assuming that no $FeSCN^{2}+$ had been formed as $[Fe^{3+}]_0$ and $[SCN^{-}]_0$. (Use $M_1V_1 = M_2V_2$).

2- Assume that all the SCN^{-} in the test tube 1 was converted to $FeSCN^{2+}$. Then calculate the concentration of $FeSCN^{2+}$, $[FeSCN^{2+}]$, in each of the other test tubes from

$$(h_1) M_1 = (h_2) M_2$$

3- Calculate $[Fe^{3+}]$ and $[SCN^{-}]$ from

$$[Fe^{3+}] = [Fe^{3+}]_0 - [FeSCN^{2+}] \quad \text{and} \quad [SCN^{-}] = [SCN^{-}]_0 - [FeSCN^{2+}]$$

4- Determine the K value separately for each tube by using the equilibrium constant expression.

PRELAB QUESTIONS

1- Define the following terms.

Equilibrium Constant, Le Chatelier’s Principle, Chemical Equilibrium, Homogeneous Equilibria and Heterogeneous Equilibria

2- Explain the effect of concentration, temperature, pressure and catalyst according to Le Chatelier’s principle on chemical equilibrium.

3- Write a brief procedure for the experiment in your own words.

QUESTIONS

1- What are possible sources of error in this experiment ?

2- Why are the values of K determined for test tubes 3, 4, and 5 probably more reliable than that determined for tube 2 ?

3- Chose one of the K values which you think has the least error.

4- How valid is the assumption that all the SCN^- in the test tube 1 was converted to $FeSCN^{2+}$? Using the K value you have choosen in the above question calculate the error by this assumption.

FE 111 GENERAL CHEMISTRY DATA SHEET

Chemical Equilibrium

Experiment : 4 ***Date:***

Test tubes	2	3	4	5
Height of Solution (h_2)				
Height of Standard (h_1)				

Name of the student: **Submitted to :**

FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment: 5

pH and Indicators

INTRODUCTION

An aqueous solution can be acidic, neutral or basic depends on the hydronium ion concentration. We can quantitatively describe the acidity by giving the hydronium ion concentration. But because these concentration values may be very small, it is often more convenient to give the acidity in terms of the pH, which is defined as the negative of the logarithm of the molar hydronium ion concentration:

$$pH = -\log [H_3O^+]$$

For a solution in which the hydronium ion concentration is 1.0×10^{-3} M, the pH is

$$pH = -\log (1.0 \times 10^{-3}) = 3.00$$

Note that the number of places after the decimal point in the pH equals the number of significant figures reported in the hydronium ion concentration.

A neutral solution, whose hydronium ion concentration at 25 ^oC is 1.00×10^{-7} M, has a of pH 7.000. For an acidic solution, the hydronium ion concentration is greater than 1.00×10^{-7} M, so the pH is less than 7.000. Similarly, a basic solution has a pH greater than 7.000.

The pH of a solution can be accurately measured by a **pH meter**. This instrument consists of specially designed electrodes that are dipped into the solution. Although less precise, **acid-base indicators** are often used to measure pH, because they usually change color within a small pH range. Acid-base indicators are organic compounds of complex structure that change color in solution as the pH changes. For example, if a drop of compound methly red (indicator) is added to an aqueous solution, it will turn the solution red if the pH is below 4.8 but yellow if the pH is above 6.0 (the solution will take on an intermediate color if the pH is between 4.8 and 6.0). Not that methly red cannot tell the difference between a pH of 6.0 and one of 8.0 or 10.0. It will be yellow in all three cases. But it can tell if the pH is greater than 6.0 or less than 4.8.

Indicators are weak acids or weak bases. Since they are intensely colored only a few drops of a dilute solution of an indicator must be used in any determination. Hence, the acidity of the solution is not significantly altered by the addition of the indicator.

The indicator exists in two forms: a weak acid, represented symbollically as HIn and having one color, and its conjugate base, represented as In^- and having a different color. Therefore, the color change of an indicator involves establishment of an equilibrium between an acid form and a base form. The equilibrium is

$$\underset{\text{acid color}}{HIn(aq)} + H_2O(l) ==== H_3O^+(aq) + \underset{\text{base color}}{In^-(aq)}$$

An increasing in the concentration of H_3O^+ shifts the equilibrium to the left, the color of HIn is observed. On the other hand, addition of OH^- decreases the concentration of H_3O^+. The equilibrium shifts to the right and the color of In^- is observed.

In this experiment solutions of known pH will be prepared and the behavior of various indicators at different pH values observed. Then you will estimate the pH of two unknown solutions by using indicators. The degree of dissociation of a weak acid will also be determined with the help of indicators.

PROCEDURE

a)

- Obtain eight clean test tubes.
- Put 9.0 mL of 0.10 M HCl solution into the first tube. (pH of this solution is 1.0)
- Put 1.0 mL of 0.10 M HCl solution into the second test tube and dilute it to exactly 10.0 mL with distilled water. (pH of this solution is 2.0)
- Take exactly 1.0 mL of solution from the second test tube and put it into the third test tube and dilute it to exactly 10.0 mL with distilled water. (pH of this solution is 3.0).
- In a similar way prepare solutions of pH 4 in the fourth test tube.
- Place half of these solutions in another four clean, dry test tubes. Now you have two sets of solution.
- Add one drop of **thymol blue** indicator in each test tube of first set. (pH 1 through 4).
- Similarly add one drop of **methyl orange** indicator in each tube of second set.
- Shake and observe the color produced in each case. Record the colors observed.
- Now obtain a sample of an acid unknown, test it with each of the indicators, and estimate its pH.
- Take 5 mL of 1.0 M acetic acid, $HC_2H_3O_2$, solution. Test it with each of the indica-tors and estimate its pH to the nearest 0.5 pH unit.
- Take 5 mL of 0.01 M acetic acid solution and determine its pH similarly.

b)

- Obtain twelve clean test tube
- Put 9.0 mL of 0.10 M NaOH solution into the first tube. (pH of this solution is 13)
- Put 1.0 mL of 0.10 M NaOH solution into the second test tube and dilute it to exactly 10.0 mL with distilled water. (pH of this solution is 12)
- Take exactly 1.0 mL of solution from the second test tube and put it into the third test tube and dilute it to exactly 10.0 mL with distilled water. (pH of this solution is 11).
- In a similar way prepare solutions of pH 10 in the fourth test tube.
- Divide each of the four of the solutions above into three parts. Now, you have three sets of solutions.
- Add one drop of **phenolphthalein** indicator in each test tube of first set
- Add one drop of **alizarin yellow** indicator in each test tube of second set
- Add one drop of **tropolin O** indicator in each test tube of third set
- Prepare a chart which shows the color of each of these indicator at pH 10 through 13.
- Now obtain sample of unknown base and determine its pH.

RESULTS AND CALCULATIONS

From each of your two pH measurements for acetic acid calculate the dissociation constant of acetic acid. Which is more realiable? When you dilute 1.0 M acetic acid to 0.1 M, did [H_3O^+] change by a factor of 10? Explain.

PRELAB QUESTIONS

1- Define: pH and pK, acidity, indicator, strong acid and strong base, weak acid and weak base

2- What is the difference between the pH and acidity

3- Find the indicators and their working pH range.

4- Suppose that 100 mL of 0.5 M HAc (Acetic acid) and 50 mL of 2 M $CaCl_2$ are mixed at room temperature to prepare a brine solution for pickling. What will be the theoretical pH of this mixture. (K_a for HAc = 1.8×10^{-5})

5- Write a brief procedure for the experiment **<u>in your own words</u>**

FE 111 GENERAL CHEMISTRY DATA SHEET

pH and Indicators

Experiment : 5 ***Date:***

Part a:

pH	1	2	3	4
Thymol Blue				
Methyl orange				

pH of Unknown:
pH of 1.0 M Acetic Acid :
pH of 0.01 M Acetic Acid :

Part b:

pH	1	2	3	4
Phenolphthalein				
Alizerin Yellow				
Tropolin O				

pH of Unknown :

Name of the student: **Submitted to :**

FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment: 6

Acid - Base Titration

INTRODUCTION

Titration is a process which is carried out by the carefully controlled addition of one of solution to another. Therefore, two solutions are used during the titration, which are **standard** and **sample** solutions. The **standard solution** is a reagent of known concentration that is used to carry out a volumetric analysis. The standard solution is gradually added to a second solution known as **sample solution** which is often placed in an erlenmeyer flask. The standard solution is placed in a container known as a buret which is used to measure the volume of standard solution added during the titration. A titration is performed by slowly adding a standard solution from a buret to a solution of the analyte until the reaction between the two is complete. The volume needed to complete the titration is determined from the difference between the initial and final buret readings.

The **equivalence point** in a titration is reached when the amount of added titrant is chemically equivalent to the amount of analyte in the sample. For example, the equivalence point in the titration of sulfuric acid, H_2SO_4, with sodium hydroxide, NaOH, is reached after introduction of two moles of base for each mole of acid.

The equivalence point of a titration is a theoretical point that cannot be determined experimentally. It can only be estimated by observing some physical change associated with the condition of equivalence. In acid – base titrations, acid base indicators are used for this purpose. Indicators are often added to the analyte solution to produce an observable physical change (color change) at or near the equivalence point. This change is called the **end point** for the titration.

In an **acid – base titration**, a solution containing a known concentration of base is slowly added to an acid (or an acid is added to a base). Acid – base indicators can be used to observe a color change in the sample solution at the end point.

In this experiment, you will first determine the concentration of an acid (part a). In part b the standardized acid will be used to determine the molarity of a sodium hydroxide, NaOH, solution. In part c, you will determine the amount of sodium hydroxide in an unknown solution.

PROCEDURE

a) Determination of Concentration of HCl Solution:

- Using a pipet place 10.0 mL of a known concentration of Na_2CO_3 solution in each of two flasks.

- Add two drops of methyl orange indicator to each of the flasks and mix the solution with a gentle shaking.

- Fill your buret with HCl.

- Place one of the Erlenmeyer flasks under the tip of the buret.

- Handle the stopcock of the buret with your left hand while shaking the flask gently with your right hand.

- Carefully run HCl from the buret into the Na_2CO_3 solution until the color change observed. (End point).

- The end point is the first appearance of a permanent red color.

- To prevent the effect of dissolved CO_2 on the pH, boil the mixture for a few minutes.

- After cooling the solution add a few drops of HCl until the end point is observed again.

- Record the volume of HCl added.

- Titrate the solution in the second flask following exactly the same procedure. Try to detect the end point as closely as possible.

- Calculate the molarity of the HCl solution.

2. Quantitative Determination of NaOH:

- Using a pipet, add 10.0 mL of 0.1 M NaOH to each of two erlenmeyer flasks.

- Add two drops of methyl orange to each flask and titrate with standardized HCl just as you did with sodium carbonate, Na_2CO_3.
- It is not necessary to boil the solution.

- Calculate the molarity of the NaOH solution and compare your result with the actual concentration.

3. Unknown

You will be given an unknown solution of NaOH. Titrate this solution just as you did above and calculate the number of grams of NaOH in the solution.

RESULTS

1. Volume of Na_2CO_3 solution used :

   Number of moles of Na_2CO_3 :

   Volume of HCl used :

Molarity of HCl :

Average :

2. Volume of NaOH solution used :

Molarity of HCl :

Volume of HCl used :

Molarity of base :

Average :

3. Volume of HCl used :

Molarity of acid used :

Mass of base (unknown) :

PRELAB QUESTIONS

1- Explain the meaning of the following terms

a) Titration

b) End Point

c) Standard Solution

d) Meniscus

2- Describe the preparation of 500 mL of 0.30 M standard NaOH solution from solid NaOH, (fw = 40.0).

3- Write a brief procedure for the experiment ***in your own words.***

OUESTIONS

1- How does dissolved CO_2 affect the end point in the titration of Na_2CO_3 with HCl? Explain by using chemical formulas. Does the pH decrease or increase in the presence of dissolved CO_2?

2- Why did you carry out the same procedure with two times for the titration of Na_2CO_3 and NaOH? Were the results same with each other? If not what can be the reasons of obtaining different results?

FE 111 GENERAL CHEMISTRY DATA SHEET

Acid-Base Titration

Experiment: 6 ***Date:***

For the standardization of HCl solution

	1	2
Volume of Na_2CO_3 solution		
Initial buret reading		
Final buret reading		
Volume of HCl used for Na_2CO_3		

For NaOH solution

	1	2
Volume of NaOH solution		
Initial buret reading		
Final buret reading		
Volume of HCl used for NaOH		

For unknown solution

	1	2
Initial buret reading		
Final buret reading		
Volume of HCl used for unknown		

Name of the student: **Submitted to:**