

FE 111 GENERAL CHEMISTRY LAB SHEET

Experiment: 1

Introduction to Chemistry Laboratory

INTRODUCTION

The following instructions are to acquaint and guide you in the procedure to be followed in performing and reporting laboratory experiments.

A. PROMPTNESS: At the beginning of each laboratory period some time will be devoted to a discussion of the work for the day. If tardy, you will, of course, miss some of the discussion. In addition, the assignment for the following week will be given during this discussion period.

B. PREPARATION: Your assignments will be given one week in advance, it is expected that you will acquaint yourself with the objective and procedure for each experiment to be performed in any given week. Careful planning of your work will enable you to understand each experiment and to complete it with a minimum of wasted effort.

C. CLEANLINESS AND ORDERLINESS: The condition of your laboratory table and equipment is some indication as to whether you are a good student. Therefore, your instructor will consider these factors in his (her) appraisal of your work.

D. CHEMICALS: Use only the quantity of each chemical that is necessary for each experiment. Do not pipette from reagent bottles directly. Do not pour chemicals back into reagent bottles.

E. QUESTIONS: Do not hesitate ask questions to your instructor. However, you are expected to make certain you can not answer questions with some additional thought before you monopolize his (her) time.

F. RECORDING DATA: There are at least good for writing a good laboratory notebook.

a) This is a laboratory course. Therefore, your work in the laboratory, which is represented partially by this record of your experience in the laboratory, is the most important factor in your course grade.

b) More important to you is the fact that you learn a great deal about experiment when you "write it up".

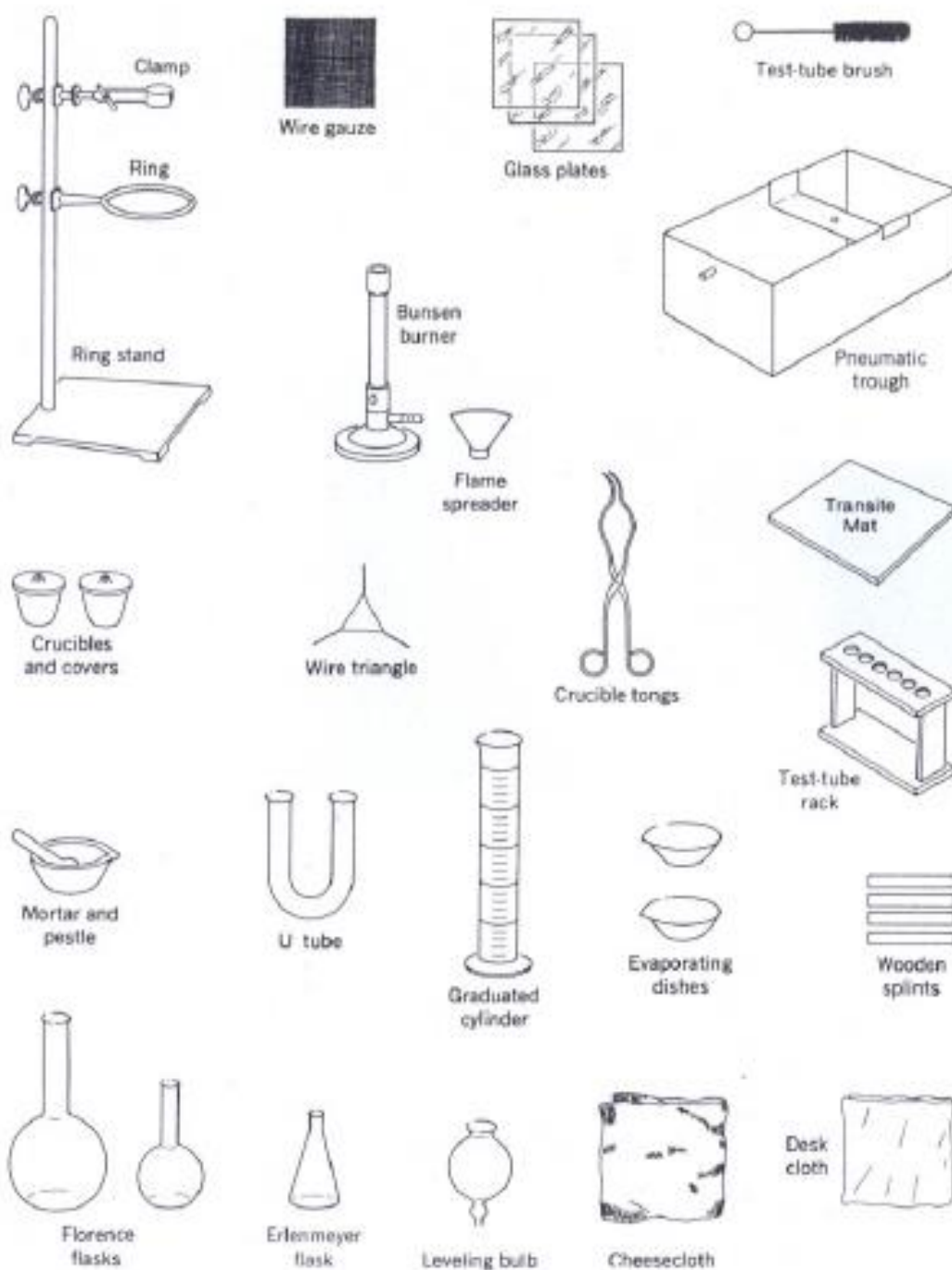
c) Most important of all; you learn how to write a good research report; a good research report is an essential requirement for a good scientist. A research report is valuable to only when you can tell at a glance exactly what you did on a particular date why you did it. Your book should be intelligible to another scientist as well, particularly your instructor in this course.

LABORATORY SAFETY

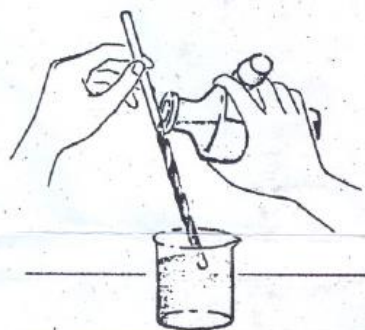
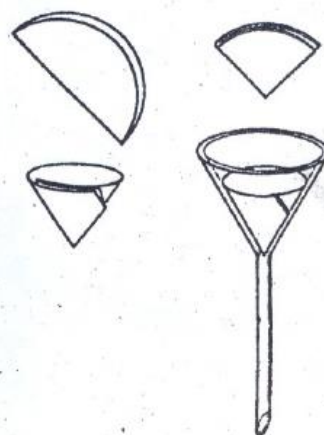
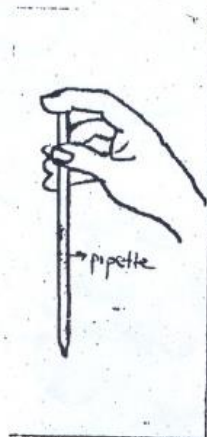
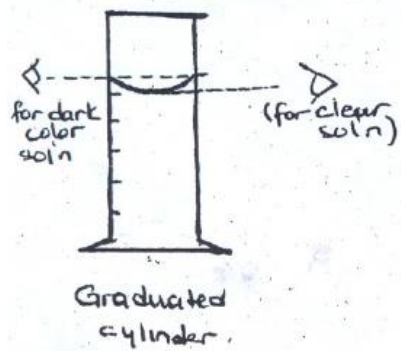
The chemical laboratory has some hazards and a laboratory accident can not be predicted . To avoid those following safety rules should be considered.

1. **Responsible behavior is essential.** The dangers of spilled acids and chemicals and broken glassware created by thoughtless actions are too great to be tolerated.
2. **Wear suitable clothing and shoes.** Shorts or bare feet leave large areas of skin exposed to potential chemical burns. Long, full sleeves may drag in your chemicals or catch on your equipment ; both the sleeves and the experiment could be ruined. Long hair should not be worn loose, especially when you are using a burner or centrifuge.
3. **Take precautions to avoid cutting yourself.** Discard cracked or chipped glassware.
4. **Dispose of waste chemicals as directed.** Do not return any substances to reagent bottles. Usually, your instructor will ask you place waste chemicals in waste jars or to pour them down a sink with plenty of running water.
5. **Perform no unauthorized experiments.**
6. **Do not smoke in the laboratory at any time.**
7. **Report all injuries to your instructor at once.**
8. **Always pour concentrated acid into water (not water into acid).**

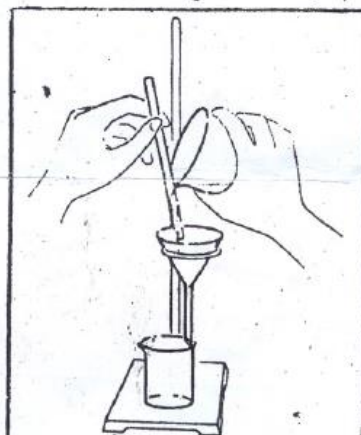
typical chemistry laboratory equipment







preparation of filter paper



Laboratory Rules

You must;

- come to laboratory in time,
- wear your laboratory coat before entering the laboratory,
- collect your hair tightly,
- read and learn your lab. sheets before coming to the laboratory

You will perform totally 9 experiments for FE 111 Lab. And if you do not attend 3 or more experiment you will fail directly.

Your total lab. mark will be calculated with; reports (30%), quiz (30%) and final exam (40%) marks.

Experiment: 2

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.

Temperature (°C)	Density (g/ml)
15	0.9979
17	0.9977
18	0.9975
20	0.9972
22	0.9968
24	0.9964

Temperature (°C)	Density (g/ml)
26	0.9959
28	0.9955
29	0.9952
31	0.9946
33	0.9941
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 change 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 : 2

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 :

Experiment: 3

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 1oC.

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 points of water at various pressures

Pressure		Boiling Point
mm-Hg	atm	°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
765	1.007	100.2
770	1.013	100.4

The thermometer can be calibrated by immersing it first in melting ice (0°C) and than in boiling water (100°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

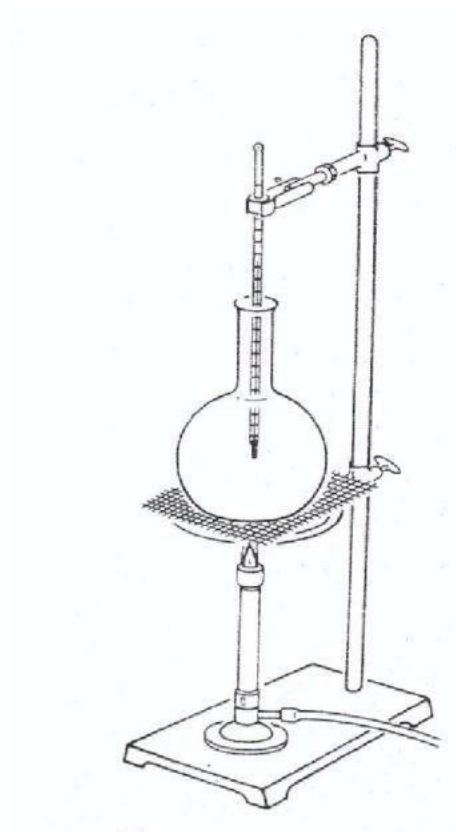
water. This can be done directly because it takes 4.2 joules (or 1 calorie) to raise 1 g of water by 1°C.

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 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°. 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 : 3

Date:

A)

Thermometer reading in ice-water mixture:

Thermometer reading at water boiling point :

Barometric pressure:

B)

Weight of metal:

Temperature of heated metal:

Temperature of the water in the test tube:

Temperature of the water and metal mixture:

Name of the student:

Submitted to :

Experiment: 4

Heats of Reaction (Exothermic or Endothermic Reactions)

INTRODUCTION

When a process occurs in which the system absorbs heat, the process is called endothermic. During an endothermic process, such as a melting of ice, heat flows into the system from its surroundings. If we, as a part of surroundings, touch a container in which ice is melting, it feels cold to us because heat has passed from our hands to the container. In an endothermic process, the temperature of the mixture in the container decreases. Assume that the system consist of the chemical reactants and products. The solvent in which they are dissolved is a part of the surroundings. Heat flows from the solvent, as a part of surroundings, into the system as reactants are converted to products. Thus, the temperature of the solution drops.

A process in which the system loses heat is called exothermic. During an exothermic process heat flows out of the system and into the surroundings. Heat flows to the solvent, again as a part of the surroundings, into the system as reactants are converted to products. Thus, the temperature of the mixture in the container increases.

The formation of a solution can be either exothermic or endothermic. For example, when magnesium sulphate, MgSO_4 , is added to water, the resultant solution gets quite warm. In contrast, the dissolution of ammonium nitrate, NH_4NO_3 , is endothermic; the temperature drops.

In this experiment, you will determine exothermic and endothermic reactions according to temperature change after addition of a solute into the water.

PROCEDURE

A)

- Put 5 ml of water in a test tube
- Record the temperature of the water
- Add 1 g of sodium chloride (NaCl)
- Carefully stir, using the thermometer and record the temperature again

B)

- Put 5 ml of water in a test tube
- Record the temperature of the water
- Add 1 g of urea
- Carefully stir, using the thermometer and record the temperature again

C)

- Put about 5 ml of copper (II) sulphate (CuSO_4) in a test tube
- Record the temperature of the CuSO_4
- Add a spatula measure of powdered zinc (Zn)
- Carefully stir, using the thermometer and record the temperature again

QUESTIONS

1. Which reaction is endothermic and which is endothermic?
2. Write the reaction for part C.
3. Write the type of reaction in the part C.

FE 111 GENERAL CHEMISTRY DATA SHEET

Heats of Reaction

Experiment : 4

Date:

DATA

Initial solution	Temperature of solution (°C)	Solid added	Final temperature of mixture (°C)	Temperature change	Type of reaction

Name of the student:

Submitted to :

Experiment: 5

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:

$$V = kT \text{ (P constant) or } V_1/T_1 = V_2/T_2 \text{ (P constant)}$$

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:

$$P_{\text{air}} = P_{\text{atm}} - P_{\text{H}_2\text{O}}$$

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

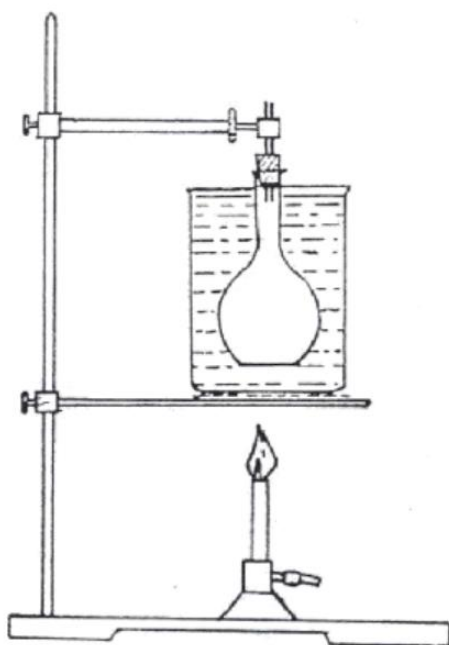
$$P_{\text{atm}} V_2 = (P_{\text{atm}} - P_{\text{H}_2\text{O}}) (V_1 - V_{\text{H}_2\text{O}})$$

where $V_{\text{H}_2\text{O}}$ 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 $P = P_{\text{atm}}$, 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 : 5

Date:

Temperature of the Boiling water (T1):

Atmospheric Pressure (P_{atm}):

Temperature of Water Bath (T2) :

Volume of Water in the Flask (V_{H₂O}):

Vapor Pressure of Water at T₂ (P_{H₂O}):

Volume of the Flask (V1):

Name of the student:

Submitted to :

Experiment: 6

Solutions

INTRODUCTION

A **solution** is a homogeneous mixture of two or more substances. The substance present in the greatest is usually called **solvent**. The other substances in the solution are called the **solutes**; they are said to be dissolved in the solvent. The solubility is the maximum amount of solute that will dissolve in a given amount of solvent at a specified temperature. Some substances can be dissolve in each other without limit. Liquids that mix in all proportions are said to be **miscible**. For example, we can prepare an ethanol water solution with a mass per cent ethanol ranging essentially from 0 to 100%. Ethanol and water are completely miscible. Most substances, though, have limited solubility. Chemists refer to substances as soluble, slightly soluble or insoluble in qualitative sense. A substance is said to be soluble (s) if a fair amount of it visibly dissolves when added to solvent. If not, the substance is described as slightly soluble (ss) or insoluble (i).

The extent to which one substance dissolves in another depends on the **nature of both the solute and solvent. It also depends on temperature.**

In this experiment you will study various factors which affect the solubility and the rate of the solution.

PROCEDURE

A. Nature of Solute and Solvent

- Test the solubility of sugar, sodium sulphate (Na_2SO_4), sodium chloride, paraffin and iodine in each of the following solvents: water and alcohol. Use 2-3 ml of solvent and a lump of solute not more than half of size of a pea.
- Record the results of your test in a table designating the substance as: soluble (s), slightly soluble (ss) and insoluble (l) in each of the solvents.

B. State of Subdivision

- Take copper sulphate and put one test tube as crystal form the other test tube as powder form.
- Add 2 ml of water to each of the test tubes and shake.
- Observe what happens.

PRELAB QUESTIONS

1. What are the factors that affect the solubility? Explain each factor.
2. What is the meaning of the term non-volatile?
3. How can a non-volatile solute affect the boiling temperature?
4. How does the solubility of a liquid and a gas changes with temperature? Does it increase or decrease?

DISCUSSION QUESTIONS

1. Discuss the solubility of sugar, sodium sulphate, sodium chloride, paraffin and iodine in water and alcohol in terms of nature of solute and solvent by using the table obtained from part A.
2. Compare and explain the rate of solution of the powder and the single crystal in water in part B of the experiment.

FE 111 GENERAL CHEMISTRY DATA SHEET

Solutions

Experiment : 6

Date:

A)

	Na ₂ SO ₄	NaCl	paraffin	iodine	sugar
water					
alcohol					

B)

Solubility of powder:

Solubility of a single crystal:

Name of the student:

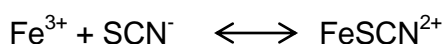
Submitted to :

Experiment: 7

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^- .



for which the equilibrium condition is;

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

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 ($h_1 \cdot M_1 = h_2 \cdot M_2$). 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:

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

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 $\text{Fe}(\text{NO}_3)_3$.
- For the remaining four tubes proceed as described below.
- Add 10 mL of 0.20 M $\text{Fe}(\text{NO}_3)_3$ to your graduated cylinder.
- Add 15 mL distilled water to make the volume exactly 25 mL, stir this solution thoroughly.
- Pour 5 mL of 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 heights 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 test tubes, assuming that no FeSCN^{2+} had been formed as $[\text{Fe}^{3+}]_0$ and $[\text{SCN}^-]_0$. 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+} , in each of the other test tubes from ($h_1M_1=h_2M_2$)
3. Calculate $[\text{Fe}^{3+}]$ and $[\text{SCN}^-]$ from;

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

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- Choose 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 chosen in the above question calculate the error by this assumption.

FE 111 GENERAL CHEMISTRY DATA SHEET

Chemical Equilibrium

Experiment : 7

Date:

Test tubes	2	3	4	5
Height of solutions (cm)				
Height of standard (cm)				

Name of the student:

Submitted to :

Experiment: 8

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:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

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

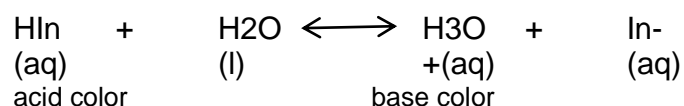
Note that the number of places after the decimal point in the pH equals the number of significant figures reported in the hydronium ion concentration.

A neutral solution, whose hydronium ion concentration at 25°C is $1.00 \times 10^{-7} \text{ M}$, has a pH 7.00. For an acidic solution, the hydronium ion concentration is greater than $1.00 \times 10^{-7} \text{ M}$, so the pH is less than 7.00. Similarly, a basic solution has a pH greater than 7.00.

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 methyl 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). Note that methyl 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 symbolically 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



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 into the third test tube and dilute it exactly 10 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, $\text{HC}_2\text{H}_3\text{O}_2$, solution. Test it with each of the indicators 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 into the third test tube and dilute it exactly 10 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 solution.
- 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 reliable? 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 CaCl₂ 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 in your own words

FE 111 GENERAL CHEMISTRY DATA SHEET

pH and Indicators

Experiment : 8

Date:

A)

pH	2	3	4	5
Tymol blue				
Methyl orange				

pH of Unknown acid:

1.0 M Acetic Acid:

0.01 M Acetic Acid:

B)

pH	2	3	4	5
phenolphthalein				
alizarin yellow				
tropolin O				

pH of Unknown base:

Name of the student:

Submitted to :

Experiment: 9

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 Na₂CO₃ 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 CO₂ 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, Na₂CO₃.
- 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.

PRELAB QUESTIONS

1. Explain the meaning of the following terms; Titration, End point, Standard solution, Meniscus
2. Describe the preparation of 500 mL of 0.30 M standard NaOH solution from solid NaOH (mw:40)
3. Write a brief procedure for the experiment in your own words.

QUESTIONS

1. How does dissolved CO₂ affect the end point in the titration of Na₂CO₃ with HCl? Explain by using chemical formulas. Does the pH decrease or increase in the presence of dissolved CO₂?
2. Why did you carry out the same procedure with two times for the titration of Na₂CO₃ 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 Titrations

Experiment : 9

Date:

For the standardization of HCl solution

Volume of Na ₂ CO ₃ solution	
Initial buret reading	
Final buret reading	
Volume of HCl used for Na ₂ CO ₃	

For NaOH solution

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

For unknown solution

Volume of unknown solution	
Initial buret reading	
Final buret reading	
Volume of HCl used for unknown	

Name of the student:

Submitted to :