Name: James Douglas Barlow   Date of Degree: May 28, 1961

Institution: Oklahoma State University

Location: Stillwater, Oklahoma

Title of Report: QUANTITATIVE EXPERIMENTS FOR HIGH SCHOOL
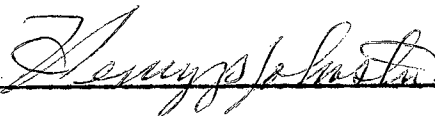CHEMISTRY

Pages in Report: 60   Major Field: Natural Science

Candidate for Degree of Master of Science

Scope and Nature of Report: The purpose of this report was to
prepare a collection of quantitative experiments for high
school chemistry that are interesting, capable of being
done with simple readily available apparatus, and involve
important concepts of chemistry. A check of laboratory
manuals for high school chemistry shows an almost complete
lack of quantitative experiments. A major difficulty has
been the finding of experiments that are both reliable
and capable of being done with simple equipment in a class
period. The experiments were developed with a goal of
causing the students to look for cause-effect relation-
ships in situations in which there are several variables
to be controlled.

Findings and Results: Twelve experiments were developed.
Several of them are open ended in that they raise other
questions that the class or individuals may wish to pur-
sue further. The experiments deal with the following
topics: size of a molecule, common ion effect on solu-
bility, heat of reaction, determination of molecular
weight by freezing point depression, determination of
simplest formula, ionization of a weak acid, effect of
concentration on rate of reaction, vitamin C content
of fruit juices, determination of molecular weight of
a volatile liquid, combining volumes of gases, and
effect of amount of catalyst on reaction rate.

ADVISER'S APPROVAL _____

QUANTITATIVE EXPERIMENTS FOR

HIGH SCHOOL CHEMISTRY

By

JAMES D. BARLOW

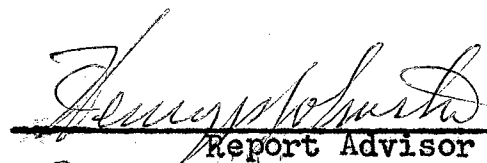Bachelor of Arts

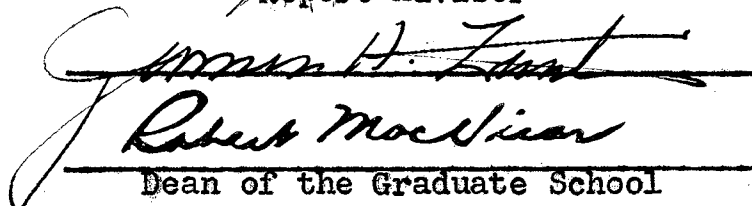Oklahoma State University

Stillwater, Oklahoma

1957

Submitted to the faculty of the Graduate School
of the Oklahoma State University
in partial fulfillment of the requirements
for the degree of
MASTER OF SCIENCE
May, 1961

QUANTITATIVE EXPERIMENTS FOR

HIGH SCHOOL CHEMISTRY

Report Approved:

_____
Report Advisor

_____

_____
Dean of the Graduate School

ACKNOWLEDGEMENTS

TABLE OF CONTENTS

LIST OF FIGURES

CHAPTER I

INTRODUCTION

Laboratory work has long been accepted as an essential learning activity in high school science, but within the past few years there has been growing concern about the value of the laboratory as presently conducted. The concensus of opinion seems to be that the often used fill in the blank, cookbook type of laboratory exercise has little educative value. Many teachers believe that too much time is spent on manipulation and confirming the textbook and not enough attention is given to problem solving. Seldom are students required to formulate hypotheses or determine experimental procedures.[1]

What then should the laboratory experience accomplish? It should require the student to think like a scientist. It should cause the student to look for cause and effect relationships. The laboratory should be a place where the student can take a question or an hypothesis and test them. He should learn the meaning of a real experiment in which there are several variables to be controlled. In short, laboratory work should foster a scientific attitude in the students and

---

[1]Paul Hurd, Science Facilities for the Modern High School (Stanford, 1954), p. 13.

1

it should lead to meaningful generalizations concerning the properties of matter.

The above goals are very good but they are somewhat idealistic in that they can not be fully achieved in every experiment. Rather these are goals to be worked at and approached over a long period of time.

The question now becomes what can the chemistry teacher do to at least partially achieve these goals. We must also keep in mind the lack of equipment and time possessed by the average chemistry teacher.

One suggestion that has been repeatedly made is that of using exercises for which the students do not know the answers.[2] This would usually involve quantitative measurements. It would keep the experiment from being a dull cookbook affair. The student often gets more satisfaction out of his numerical data than from filling in the blanks of a workbook. This is not necessarily the only improvement needed, but it is one of them.

A check of laboratory manuals for high school chemistry discloses an almost complete lack of quantitative experiments. One such manual has only four quantitative exercises out of a total of eighty seven. Most of the experiments could be

[2] Aaron Goff, "Emphasizing the Unknown," Science Teacher, XXII (1955), 291-292.

Henry Gabriel, "Quantitative Work in Freshman Chemistry," Journal of Chemical Education, XXXIII (1956), 42-44.

completed just by using the textbook.[3]

A major difficulty has been the finding of quantitative experiments that are both reliable and capable of being done with simple equipment in a class period.

The purpose of this paper is to prepare a collection of quantitative experiments for high school chemistry that are interesting, can be done with readily available apparatus, and involve important concepts of chemistry.

---

[3]Maurice Evans, Norris Bush, and Ray Easley, Laboratory Problems for Living Chemistry (Boston, 1957).

CHAPTER II

QUANTITATIVE EXPERIMENTS
FOR HIGH SCHOOL CHEMISTRY

Foreword

These experiments are designed for high school chemistry. Some of the experiments can be done individually. The students will need to work in small groups on others. Most of the experiments are suitable for all students but the teacher may wish to reserve some of them for the more able student.

The experiments differ from the ordinary exercises found in laboratory manuals in a number of ways. These experiments come close to being real experiments. The students do not know the outcomes before they enter the laboratory. They can not fill in the answers from the textbook. The pupil must depend on his experimental results and his ability to interpret them. Since the experiments require measurements, they involve more precise manipulations than ordinary exercises. The student is made aware of sources of error and their effect on his results.

Experiment Number One

Size of Molecules

Problem

What is the order of magnitude of the size of a molecule?

General Approach

Molecules and atoms are so small that we can not see them to measure their size. In this experiment you will use a simple indirect method to get an approximate value for the length of a molecule and for Avagadro's number.

A small quantity of oil placed on a water surface will spread out to form a very thin film. This film must be at least one molecule thick. If we could find the thickness of the layer of oil, we would know that the oil molecule is equal to or less than the thickness of the film.

Oleic acid is another material that will form a thin, essentially monomolecular film on water if the water surface is large enough. You will use a very dilute solution of the acid in alcohol to the thinnest possible film (one molecule thick) on a small area.

Apparatus and Materials

    5 ml. pipette or a small graduated cylinder

    100 ml. graduated cylinder

    Large <u>clean</u> shallow tray

    Lycopodium powder or fine chalk dust

    Oleic acid

    Ethyl alcohol

Procedure

1. Measure 5.0 ml. of oleic acid and 95 ml. of alcohol and mix well. Take 5.0 ml. of this solution and mix well with 45 ml. of alcohol. Calculate the amount of oleic acid in each ml. of this final solution.

2. Fill a large clean shallow tray with water to a depth of about one centimeter. Dust the surface of the water lightly with chalk or lycopolium powder. This will serve to make visible any films formed on the water by the alcohol or the oleic acid.

3. Drop one or two drop of alcohol (not the oleic acid solution) on the surface with a medicine dropper. Does the alcohol form a film? Why?

4. Use a clean eyedropper or pipette to apply a drop of the oleic acid solution. What happens? Measure the average diameter of the film and calculate its area. Add another of the solution to the center of the clear area. Is the area twice as large? What does this indicate?

5. Using the same pipette as in step four, find out how many drops of solution it takes to form 5.0 ml. Calculate the volume of one drop and the volume of oleic acid in one drop of the solution.

6. Using the volume of oleic acid in one drop of solution and the area on the water surface covered by that acid, calculate the thickness of the oleic acid film. What does this say about the size of oleic acid molecules?

7. If the layer is one molecule thick and the molecules

are assumed to be cubes, how many molecules would there be
in one ml. of oleic acid?

8. If one ml. of oleic acid weighs 0.89 grams, what
is the mass of one molecule?

9. If the molecular weight of oleic acid is 282, what
does this indicate that Avogadro's number is?

Data Sheet

Calculate the volume of oleic acid in each milliliter
of the solution. Show clearly how you do this.

How do you know that the alcohol doesn't form a film
on the water surface? Why doesn't it?

What is the area covered by the oleic acid in one drop
of the solution? Show your work.

Number of drops in one milliliter is _____.

Calculate the volume of oleic acid in one drop of the
solution. Show your work.

Calculate the length of an oleic acid molecule. Tell
and show how you do this.

Volume of the oleic acid molecule is _____.

Weight of an oleic acid molecule is _____.

Show clearly how you arrive at this.

Calculate Avogadro's number. Show your work clearly.

Why does the oleic acid spread to form a monomolecular layer? Why doesn't it dissolve in the water?

Experiment Number Two

Effect of Other Substances on Solubility

Problem

Is the solubility of a substance influenced by the presence of other substances in solution? If yes, what kinds of substances effect each other and how?

General Approach

In this experiment you will determine the solubility of calcium iodide in water and in solutions of calcium bromide, potassium bromide, and potassium iodide. The solubility will not be the same. The scientist tries to find some constant relationships to unite the apparent variations of nature. You should try to find some mathematical expression for the solubility of calcium iodide that is constant in each case. You may need to refer to reference texts for help.

Apparatus and Materials

100 ml. graduated cylinder

10 ml. pipette

Evaporating dish

Balance

Calcium bromide

Calcium iodide

Potassium bromide

Potassium iodide

Rubber stoppers for the cylinders

Procedure

1. Prepare solutions in each of four graduated cylinders as follows:

a. 20.0 grams of potassium bromide in enough distilled water to make 50.0 ml. of solution.

b. 20.0 grams of calcium bromide in enough water to make 50.0 ml. of solution.

c. 20.0 grams of potassium iodide in enough water to make 50.0 ml. of solution.

d. 50.0 ml. of water only.

Label the cylinders and stopper them to prevent evaporation.

2. Add 20.0 grams of calcium iodide to each of cylinders and agitate to aid dissolving. If all the salt dissolves then add about five grams more. Continue adding calcium iodide until some of it remains undissolved in the bottom of the cylinder after setting for a day with periodic mixing.

3. After you are satisfied that the solutions are saturated with calcium iodide, you have the problem of determining the amount of the substances present in each solution. Determine the volume of each solution (don't forget to subtract the volume of undissolved salt in the bottom of the cylinder). You can now calculate the concentration of potassium bromide, potassium iodide, and calcium bromide.

4. Determine the total concentration of dissolved salt in each solution by removing 10.0 ml. of clear solution with a pipette and evaporating the solution to recover the salt.

You should rinse the pipette out into the evaporating dish so that none of the salt is left in the pipette. The dish should be large enough that it is no more than half filled by the solution. Evaporate cautiously so that you do not lose salt by splattering. You should now be able to calculate the concentration of each kind of salt dissolved in each solution.

5. Your job now is to find something about the quantity of calcium chloride that is the same in all the solutions. Maybe the total weight or number of calcium and chloride ions in each solution is constant.

6. Describe your attempts to solve the problem on an additional data sheet. Give all your calculations and their results.

This experiment contains a number of sources of error. Indicate what some of these are and what might be done to eliminate them.

Data Sheet

    Solution (a)

    Final volume of solution   _____

    Concentration of potassium bromide (grams/ml.) _____

    Weight of dissolved salt in 10.0 ml.   _____

    Weight of calcium iodide in 10.0 ml.   _____

    Concentration of calcium iodide (grams/ml.)   _____

    Solution (b)

    Final volume of solution   _____

    Concentration of calcium bromide (grams/ml.)   _____

    Weight of dissolved salt in 10.0 ml.   _____

Weight of calcium iodide in 10.0 ml. _____

Concentration of calcium iodide (grams/ml.) _____

Solution (c)

Final volume of solution _____

Concentration of potassium iodide (grams/ml.) _____

Weight of dissolved salt in 10.0 ml. _____

Weight of calcium iodide in 10.0 ml. _____

Concentration of calcium iodide (grams/ml.) _____

Solution (d)

Weight of calcium iodide in 10.0 ml. _____

Concentration of calcium iodide (grams/ml.) _____

Notes to Teacher

This experiment is not intended to be assigned to all students. It is intended to be an original investigation for the interested superior student.

The following data are for a temperature of $0^{\circ}C$. The solubility product of calcium iodide is given by

$$(Ca) \times (I)^2 = 45$$

where ( ) indicate concentration of the ion in moles per liter. The solubility of calcium iodide in 50 ml. of

1. water - 33 grams

2. water and 20 grams of potassium bromide - 33 grams

3. water and 20 grams of calcium bromide - 25 grams

4. water and 20 grams of potassium iodide - 22 grams

Experiment Number Three

Measurement of a Change in Chemical Energy[1]

Problem

How much heat is produced by the neutralization of hydro-
chloric acid with sodium hydroxide?

General Approach

The heat change of a chemical reaction that occurs in
water solution can be found by measuring the temperature change
of the solution as the reaction proceeds. In this experiment
you will neutralize an acid with a base and determine the heat
of reaction in water solution. The heat liberated by this
exothermic reaction will warm the solution.

The unit of heat energy is the calorie. A calorie is
the amount of heat it takes to raise the temperature of one
gram of water one degree centigrade. Thus the heat absorbed
by water as its temperature increases is equal to the temp-
erature increase times the weight of the water. A small part
of the heat given off by the reaction is used to heat the
beaker that contains the solution. It takes about 0.17 cal-
orie to heat one gram of glass through one centigrade degree.
Thus the heat absorbed by the beaker is 0.17 of the product
of the weight of the beaker and the temperature increase.

_____

[1]Alfred Garret, Joseph Haaskins, Thor Rubin, and Frank
Verhoek, Chemistry for the Laboratory (Boston, 1957), p. 19.

Apparatus and Materials

    1.0 N hydrochloric acid and sodium hydroxide

    400 ml. beaker

    100 ml. graduated cylinder

    2-200 ml. beakers

    Thermometer

    Balance

Procedure

    1. Measure into separate beakers 100 ml. of 1.0 N HCl and 100 ml. of 1.0 N NaOH. Determine the temperature of each of these solutions and record. They should both be at room temperature.

    2. Place the thermometer in a clean 400 ml. beaker that has been previously weighed. Place the beaker on a folded towel to insulate it from the table top.

    3. Pour the acidic and basic solutions into the beaker. Stir with the thermometer and record the highest temperature reached.

    4. Calculate the total heat produced by the reaction by adding together the heat absorbed by the solution and the beaker as they warm up.

    5. Calculate the amount of heat produced when one mole of sodium hydroxide reacts with one mole of hydrochloric acid.

    6. Find out the accepted heat value of this reaction and calculate your per cent error.

Data and Results

    Initial temperature of solution and beaker   _____

Final temperature of solution and beaker   _____

Weight of beaker   _____

Calculate the amount of heat absorbed by the water.

Calculate the amount of heat absorbed by the beaker. Show your work clearly.

Calculate the amount of heat that is released when one mole of sodium hydroxide neutralized one mole of hydrochloric acid. Show how you do this.
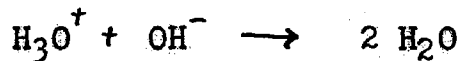
Write the equation for this reaction and include the heat of reaction on the proper side.

Accepted value for the heat of reaction   _____

Per cent difference of experimental value from the accepted value   _____

What are some sources of error in this experiment?

What do you think the heat of this reaction is? Explain.

$$H_3O^+ + OH^- \longrightarrow 2 H_2O$$

Notes to the Teacher

The basic reaction is that of hydrogen (or hydronium) ion with hydroxyl ion to produce water. The accepted value for the heat of reaction is 13,600 calories per mole of water formed from a mole of each of the ions. The temperature of the solution will increase about six centigrade degrees.

Experiment Number Four

Determination of the Molecular Weight of Sulfur

Problem

How can we use the molal freezing point constant to find the molecular weight of a substance?

General Approach

When a substance is dissolved in a liquid the freezing point of the solution is lower than that of the pure liquid. The amount of the lowering depends on the concentration of solute particles. An Avagadro's number of particles in one kilogram of water lowers the freezing point by 1.86 C^o. This suggests a method of determining molecular weights.

In this experiment you will determine the molecular weight and formula for sulfur by observing the freezing point lowering produced by dissolving sulfur in napthalene. Water can not be used as the solvent because sulfur is not soluble enough in it. Napthalene is used because of its convenient melting point (80°C.) and its large molal freezing point constant (6.8 C^o.). You will prepare a solution of known concentration and measure its freezing point. Using this you can calculate the weight of sulfur needed to lower the freezing point of 1000 grams of napthalene 6.8 C^o. The calculated weight is the gram molecular weight of sulfur as it occurs in this solution. You can now use the atomic weight to determine the number of atoms per molecule.

This is a general method that can be used with other

substances and other solutions.

Materials and Apparatus

Thermometer

Large beaker

Ring stand and ring

Test tube clamp

Asbestos pad

Balance

Bunsen Burner

Napthalene

Sulfur

Large test tube

Procedure

1. Set up a beaker of water on a ring stand as if you were going to heat the water.

2. Weigh a clean dry large test tube as accurately as you can. Add about twenty grams of napthalene and again weigh the tube carefully. Record the weights.

3. Clamp the test tube in the beaker of water. The water level in the beaker should be higher than the napthalene in the test tube.

4. Heat the water until all the napthalene has melted and then remove the burner. Fit the test tube with a two hole stopper. There should be a thermometer through one hole that extends down into the napthalene. The other hole should contain a nichrome or steel wire with a horizontal loop on it. This will serve as a stirrer.

5. Stir the napthalene with an up and down motion as the water and napthalene cool. Keep track of the temperature as it falls. Record it each thirty seconds. As the napthalene solidifies the temperature will remain constant. This is the freezing point of napthalene. Record.

6. Weigh accurately a piece of glazed paper. Place about two grams of sulfur on the paper and reweigh. Record.

7. Remelt the napthalene and add all the sulfur to the test tube. Continue to heat and stir until all the sulfur dissolves.

8. Remove the burner and allow the solution to cool. Stir continuously and read the thermometer as before to determine the freezing point of the solution. Use the temperature at which you first see solid forming as the freezing point. Remelt the napthalene before removing the thermometer and stirrer.

Data Sheet

    Freezing point of napthalene   _____

    Freezing point of solution   _____

    Weight of napthalene in solution   _____

    Weight of sulfur in solution   _____

Calculate the number of moles of sulfur that you had in your solution for each 1000 grams of napthalene. Show how you do this clearly.

Calculate the number of grams of sulfur per 1000 grams of napthalene in the solution you made.

You have now calculated the number of moles and the number of grams of napthalene per 1000 grams of solvent. Deduce from this the molecular weight of sulfur.

If the atomic weight of sulfur is 32, what is the formula of a molecule? Do you think that sulfur would always have the same formula?

Why do we take the freezing point of sulfur in napthalene solution when it first begins to freeze rather than toward the end?

Notes to the Teacher

The thermometers should be read or estimated to the nearest 0.1°. The temperature lowering will be about 2.5 C$^{\circ}$ if the specified amounts are used.

The molecular formula of sulfur should be S_8.

Experiment Number Five

Determination of the Simplest Formula[2]

Problem

Can we actually find the formula of a compound by the methods described in the textbook?

General Approach

Nitric acid is a strong oxidizing agent. When it is added to tin, the metal is converted to the oxide and the brown fumes of nitrogen dioxide are liberated. This gas is poisonous. The reaction should be performed under a hood or in a well ventilated room.

In this experiment you are to find the percentage composition of tin oxide and from this calculate the simplest formula of the compound.

Materials and Apparatus

Tin

Nitric acid

Evaporating dish

Ring stand, ring, and asbestos pad

Balance

Procedure

1. Carefully weigh a clean dry evaporating dish to the nearest 0.01 gram. Record.

---

[2]Fred Weisbruch, *Semimicro Laboratory Exercises for High School Chemistry* (Boston, 1956), p. 67.

2. Add about two grams of powdered or mossy tin to the dish and reweigh. Record.

3. Add about five ml. of 7-8 N nitric acid to the tin. Allow the reaction to take place under a hood.

4. When the reaction has nearly ceased, begin heating the dish with a <u>low</u> flame. Gradually increase the heating until most of the liquid has evaporated. Finally heat with the full strength of the flame for five minutes. The oxide must be completely dry before you stop heating. Allow to cool.

5. Examine the residue. Can you see any tin left unreacted. If so, add more nitric acid and heat again. If you are satisfied that the reaction is complete, weigh the dish and its contents and record.

Data and Results

Weight of evaporating dish _____

Weight of dish and tin _____

Weight of tin used _____

Weight of tin oxide and dish _____

Weight of tin oxide formed _____

Weight of oxygen in tin oxide _____

Calculate the per cent of tin in the tin oxide.

Calculate the per cent of oxygen in the tin oxide.

Calculate the simplest formula of tin oxide. Show and label your work clearly.
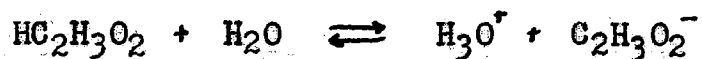
Experiment Number Six

Ionization of a Weak Acid

Problem

How does the ionization of an acid change as the concen-
of the acid and its salts changes?

General Approach

Acetic acid ionizes in water solution as indicated in
the following reaction.

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

The two arrows indicate that this is a reversible reaction
that reaches equilibrium. At equilibrium the rate of the re-
verse reaction is equal to the rate of the forward reaction.
This causes the total amount of each reactant or product to
remain constant. This does not mean that the concentration
of substances on the left side of the equation is equal to
the concentration of those on the right side.

The position of the equilibrium can be expressed in a
number of ways. It can be expressed as the per cent of ioni-
zation. This is equal to the amount of acetic acid ionized
divided by the total amount of acetic acid. These amounts
are usually expressed in moles per liter.

The second method of expressing the position of the
position of the equilibrium involves the use of the ionizat-
ion constant (K) which is derived from the law of mass action.

$$K = \frac{(H_3O^+)(C_2H_3O_2^-)}{(HC_2H_3O_2)}$$

The parenthesis indicate that the concentration in moles per liter is used in the equation. You must use the total hydronium and acetate ion concentration in the solution regardless of the source of the ion. The concentration of unionized acetic acid is used as the denominator.

In this experiment you will prepare solutions of acetic acid of known concentration and measure the hydrogen ion concentration. From this you will be able to calculate the ionization constant and the per cent ionization. Study the following examples closely for the techniques you will use in doing your own calculations.

Example 1. Suppose that you measure the pH of a 1.0 molar solution of acetic acid and it is 4.0. What is the per cent of ionization and the ionization constant?

First we must find the hydrogen ion concentration of the solution. The pH is defined by the relationship

$$pH = 1/\log (H_3O^+) \quad \text{or} \quad (H_3O^+) = 1/\text{antilog } pH.$$

The antilogarithm of 4 is 10^4 so that (H_3O) equals $1/10^4$ or 0.0001 moles/ liter. Since the ionization reaction produces equal numbers of hydronium and acetate ions, the acetate ion concentration must be the same.

The concentration of unionized acetic acid is 1.0 minus 0.0001 moles/liter. This is 0.9999 moles/liter. But our data does not contain this many significant figures so it is just as accurate and more convenient to use 1.0 mole/liter as the acetic acid concentration.

We can now use these values to calculate the per cent

of ionization and the ionization constant.

$$K = \frac{(H_3O^+)(C_2H_3O_2^-)}{(HC_2H_3O_2)} = \frac{10^{-4} \times 10^{-4}}{1.0} = 10^{-8}$$

Per cent ionization $= \dfrac{10^{-4}}{1.0} = 10^{-4} = 0.01\%$

Example 2. You have one liter of solution that contains one mole of acetic acid and one mole of sodium acetate. Its pH is 5.4. What is the per cent of ionization and the ionization constant?

$(H_3O^+) = 1/$ antilog $5.4 = 1/2.5 \times 10^5 = 4.0 \times 10^{-6}$

The acetate ion concentration comes from the ionized acetic acid and from the sodium acetate. The total acetate ion concentration must equal

$(C_2H_3O_2^-) = 4.0 \times 10^{-6}$ plus $1.0 = 1.000006$ moles/liter. We will use 1.0 mole/liter for the acetate ion concentration since our date is not accurate to seven significant figures. The unionized acetic acid concentration is still approximately 1.0 mole/liter since the part that ionized to produce hydrogen ions is not a significant part of this. Substituting these values in the equations for the ionization constant and per cent of ionization gives the following.

$$K = \frac{(H_3O^+)(C_2H_3O_2^-)}{(HC_2H_3O_2)} = \frac{4.0 \times 10^{-6} \times 1.0}{1.0} = 4.0 \times 10^{-6}$$

Per cent of ionization $= \dfrac{4.0 \times 10^{-6}}{1.0} = 0.0004\%$

The values in these examples are not to be taken as necessarily realistic.

Materials and Apparatus

Balance

100 ml. graduated cylinder

Beakers

pHydrion paper or other accurate universal indicators

Sodium acetate

Acetic acid

Procedure

1. Take 10.0 ml. of 1.0 molar solution of acetic acid and measure its pH with pHydrion paper. Your instructor will will show you how to use this indicator.

2. Add 90.0 ml. of distilled water to the 10.0 ml of acetic acid solution. Stir well. This dilutes it to a 0.10 molar solution. Measure and record its pH.

3. Take 10.0 ml. of this 0.10 molar solution (save the other 90 ml.) and add 90.0 ml. of distilled water to it. This produces a 0.010 molar solution. Measure and record its pH.

4. Take 10.0 ml. of 1.0 molar acetic acid and add 90.0 ml. of distilled water and one tenth of a gram molecular weight of sodium acetate. This gives you a solution that is 0.10 molar acetic acid and 0.1 molar sodium acetate. Measure and record its pH.

5. Take 10.0 ml. of the solution from step four and add the 90 ml. of solution that you saved from step three. This gives you a solution that is 0.10 molar acetic acid and 0.01 molar sodium acetate.

6. Calculate the hydrogen ion concentration for each solution.

7. Calculate the acetate ion concentration for each solution. The acetate ion concentration will equal the hydrogen ion concentration if it is a solution of the acid only. The acetate ion concentration is increased by the amount of sodium acetate dissolved in the solution.

8. The acetic acid concentration will equal the molarity of the solution minus the hydronium ion concentration. if the hydronium ion concentration is very much smaller than the molarity, then it can be ignored.

9. Calculate the ionization constant and the per cent of ionization for the acetic acid in each of the five solutions.

Data and Results

Prepare on a separate sheet a table that gives the following information about each solution: molarity of the solution with respect to acetic acid and sodium acetate, pH, concentration of hydronium and acetate ions, ionization constant, and per cent of ionization of the acetic acid. Show your calculations clearly and labeled.

Your data should indicate that the hydronium ion concentration decreases as sodium acetate is added to the solution. Explain why.

Why are the ionization constants of weak acids used more than the per cent ionization?

What happens to the per cent ionization as the acid solution becomes more dilute? Why?

If the rates of the opposing reactions are equal at equilibrium, why doesn't the concentration of the reactants

and the products have to be equal?

Note to the Teacher

pHydrion papers are available that cover the pH range from one to thirteen and give distinct color changes for a changes of 0.4 pH units.

Experiment Number Seven

Effect of Concentration on Rate of Reaction[3]

Problem

What is the relationship between the concentration of a substance and the rate of a chemical reaction in which it takes part?

General Approach

The speed of a chemical reaction is influenced by many things. Some of these are concentration of reactants, temperature, nature of the reaction, presence of catalyst, and the presence of competing reactions. In this exercise you are going to vary the concentration of one of the reactants and study the resulting rates of reaction. You will need to keep the other factors that influence the speed of the reaction constant.

It is ordinarily difficult to determine the speed of a reaction. In this experiment you will use a series of reactions. The first reaction is relatively slow. The second takes place rapidly and only after the first is completed. A product of the second reaction causes a color change in the solution. We can measure the time of the slow reaction because it takes up practically all of the time from the mixing of the chemicals to the appearance of the color change.

---

[3]Richard Siegel, Scientific Experiments in Chemistry (Washington, D. C., 1958), HS-T 712.

You will run the reaction a number of times with varying concentrations to determine the relationship between the concentration of one of the reactants and the speed of the reaction.

The reactions involved in this experiment can be expressed by the following equations.

$$3 \ H_2SO_3 + HIO_3 \longrightarrow 3 \ H_2SO_4 + HI \qquad (1)$$

$$5 \ HI + HIO_3 \longrightarrow 3 \ I_2 + 3 \ H_2O \qquad (2)$$
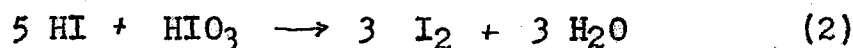
Reaction (1) is slow and reaction (2) is instantaneous. Since sulfurous acid is a stronger reducing agent than hydriodic acid, reaction (2) can not occur until all the sulfurous acid is used up in reacting with the iodic acid. There will be an excess of iodic acid in each of the solutions so that reaction (2) can occur. As soon as reaction (1) is complete, reaction (2) takes over, producing enough iodine in a small fraction of a second to react with starch to produce the dark blue color that characterizes the starch-iodine complex. This color marks the end of the reaction. You will vary the concentration of the iodic acid and note its effect on the reaction time.

Materials and Apparatus

 100 ml. graduated cylinder

 250 ml. beaker

 Clock or watch with second hand

 Potassium iodate solution

 Sulfurous acid-starch solution

 Stirring rod

Procedure

1. Measure 30.0 ml. of sulfurous acid-starch solution in a dry graduated cylinder (or one previously rinsed with sulfurous acid). Pour into a 250 ml. beaker.

2. Rinse the cylinder and add 10.0 ml. of water to the beaker. Rinse the cylinder with about 5 ml. of iodic acid solution. Discard the rinse. Measure 60.0 ml. of the iodic acid solution.

3. Be prepared to note the time. Add the iodic acid solution quickly to the beaker, stirring well, and noting the time when the solution was all added to the contents of the beaker. Determine the time that passes from the addition of the acid to the appearance of the dark blue color.

4. Repeat this procedure using the other volumes indicated in the data. Keep your procedure as nearly constant as possible.

Data and Results

	H_2SO_3 (ml)	H_2O (ml)	HIO_3 (ml)	Total (ml)	Time (sec.)	Speed (1/time)
(a)	30	10	60	100		
(b)	30	20	50	100		
(c)	30	30	40	100		
(d)	30	40	30	100		
(e)	30	50	20	100		

Calculate the speed of the reaction to the nearest 0.001 by calculating the fractional part of the reaction that occurs per second. Record above.

Can you see any relationship between the speed of the reactions and the concentration hydriodic acid? If so, what? Why is this relationship so?

Make a graph by plotting concentrations of HIO_3 against the speed of the reaction. What does this graph indicate about the relationship between these two variables?

Consult your text or references to find out what the Law of Mass Action is. Explain how it applies in the experiment you have just done.

Suppose you had a reaction

$$A + B \longrightarrow$$

and you did not know how many molecules of each are involved in the equation. Suppose that you doubled the concentration of A and the reaction went four times as fast. How many molecules of A are in the equation?

Notes to the Teacher

The "iodic acid" solution is prepared by dissolving two grams of potassium iodate in a liter of distilled water. The acid is formed when this solution is added to the sulfurous acid which contains an excess of hydrogen ion.

The sulfurous acid-starch solution is made as follows. Add about ten ml. of cold water to five grams of starch (not precooked type) in a 600 ml. beaker and stir to a paste. Add with stirring 200 ml. of boiling water. Add 200 ml. of cold water and allow to cool. Add two grams of anhydrous sodium sulfite and six ml. of 6N sulfuric acid. Add enough water to make about two liters of solution. This solution

should be fresh when used. It should be shaken to distribute the starch.

This experiment uses measuring techniques not often used in high school. The measurement of concentration is only a relative one. This will need explanation to the class.

This experiment can be adapted to class periods of varying lengths. The class can be broken into groups with each group determining the time for only a few concentrations if the period is short.

Special mention should be made of the technique of pre-rinsing graduated cylinders and the reasons for it.

The following is a typical set of time intervals measured under the conditions given in the procedure and data sheet: (a) 70 seconds, (b) 84 seconds, (c) 104 seconds, (d) 140 seconds, and (e) 208 seconds.

Experiment Number Eight

Vitamin C Content of Fruit Juices[4]

Problem

How is the vitamin C content of fruit juices influenced by the source and handling?

General Approach

There are several substances in fruit juices that will react with iodine to reduce it to the iodide ion. The most important of these is ascorbic acid or vitamin C. We can get a good approximation of the concentration of vitamin C in fruit juices by titrating with iodine solution. The end point is shown by starch which turns dark blue in the presence of iodine left in solution after all the vitamin C has been oxidized. Instead of using an iodine solution of known strength and calculating the amount of vitamin C present using the chemical reaction, you will titrate a solution of known vitamin C content with your iodine solution in order to calibrate it for titration of the fruit juices.

Apparatus and Materials

Vitamin C tablets of known content

6N hydrochloric acid

Burette

Starch solution

---

[4]Ernestine Long, Scientific Experiments in Chemistry (Washington, D. C., 1958), HS-T 719.

Iodine solution

Beakers

10 ml. pipette or graduated cylinder

Portions of fresh, canned, frozen, and stale juices from several fruits.

Procedure

1. Dissolve a 50 mg. tablet of vitamin C in about ten milliliters of distilled water in a small beaker. Acidify with two drops of hydrochloric acid. Add two droppers-full of starch solution.

2. Fill your burette nearly full of iodine solution. Allow a little of the solution to run out (not into the vitamin solution) of the burette to fill the tip. Read and record the level of iodine solution in the burette. Remember to read at the bottom of the meniscus.

3. Add iodine solution drop by drop to the vitamin C solution until a faint violet color persists. A white paper under the beaker will help you to see the end point. The flask should be constantly swirled to keep the contents well mixed.

4. Read the burette and subtract the initial reading to find the smount of iodine solution used. Rinse your flask and repeat the titration of another vitamin C tablet until you get two values very close together. Use the average of the "good" values to calculate the weight of vitamin C that reacts with 1.0 ml. of the standard iodine solution.

5. You are now to determine the vitamin C content of

fruit juices. The class has been provided with portions of juices from several fruits that have been handled by various methods. Since you will not have time to titrate samples of all the juices, your instructor will assign you do some of these while others titrate portions of the other juices. The data from each of the titrations will be made available to all members of the class.

6. If you are assigned a fresh fruit, you will need to cut the fruit and squeeze out the juice. Filter the juice to remove pulp before taking a sample to titrate.

7. Measure 10.0 ml. of juice you have been assigned. The pipette or cylinder should be clean and dry or else freshly rinsed with the juice you are going to titrate. Put the sample in a clean but not necessarily dry beaker or flask. Add two droppers-full of starch solution as indicator. Add the iodine solution with burette slow and with swirling until a faint violet color persists. Repeat until you get two volumes of iodine solution needed for this juice that are close together.

8. Repeat with other juices.

9. Obtain data of the other students titrations.

10. Calculate the amount of vitamin C in ten ml. of the various juices. Make a neat table of the findings. Include these headings: Kind of juice, Treatment of juice, Average volume of iodine used, and Amount of vitamin C per 10 ml.

Conclusions

Analyze the table. Draw conclusions about the relative

merit of various juices and methods of storing them. Compare with values given in nutrition charts.

What assumptions do you have to make in order to use this method of finding the amount of vitamin C content?

Why were you not asked to add hydrochloric acid to the fruit juices as you did with the vitamin C tablet?

Teacher's Notes

You can get vitamin C tablets at a drug store. They should be of known quantity and approximately 50 mg.

You should prepare enough of the iodine solution to last throughout the experiment for the whole class. It will take a minimum of 150 ml. for each student. Use 0.6 grams of iodine for each liter of solution. Dissolve the iodine in a little ethyl alcohol (5 ml for each 0.6 grams of iodine). Then add water to get a final solution with 0.6 grams/ liter. The use of the alcohol makes the end point sharper.

The starch indicator solution is made by starting with a paste of Faultless starch with a small amount of water and boiling for five minutes.

It will take about 80 ml. of the iodine solution to react with 50 mg. of vitamin C. It would be well to advise the students of this so that they won't take so long to reach the end point.

You will need to advise the students carefully about the techniques of titration.

You should prepare or have the students bring in the juice of a fruit (orange for example) that has been handled several

ways, i. e., fresh, frozen (follow directions on can), canned, exposed to air in refrigerator for a week, exposed to air at room temperature for a week, and boiled as when sterilized for babies. You can use several different canned fruit juices to compare the various kinds of fruit juices.

This experiment assumes that (1) no iodine is consumed by anything in the solution except vitamin C, (2) the uptake of vitamin C is proportional to the vitamin C content, and (3) no iodine reacts with the starch until all the vitamin C is oxidized.

The students should be discouraged from adding the iodine too fast. The starch will turn blue before the end point and it will take a few minutes for the color to disappear.

Experiment Number Nine

Determination of Percentage Composition and Simplest Formula[5]

Problem

How can we determine the percentage composition and simplest formula of two substances using simple apparatus?

General Approach

A familiar reaction in chemistry is the reduction of copper oxide with hydrogen. When hydrogen is passed over heated copper oxide, the metal is set free and hydrogen oxide (water) formed.

You will prepare hydrogen and dry it by passing it over calcium chloride. The hydrogen is used to reduce copper oxide. The percentage composition of copper can be calculated from the weight of copper oxide used and the weight of copper after reduction.

The water formed is absorbed by the calcium chloride of the second drying tube. The weight of water formed is the increase in weight of this second drying tube. The weight of oxygen in this water is the oxygen lost by the copper. You should now be able to calculate the percentage composition of water.

You can now calculate the simplest formula of copper oxide and water.

_____

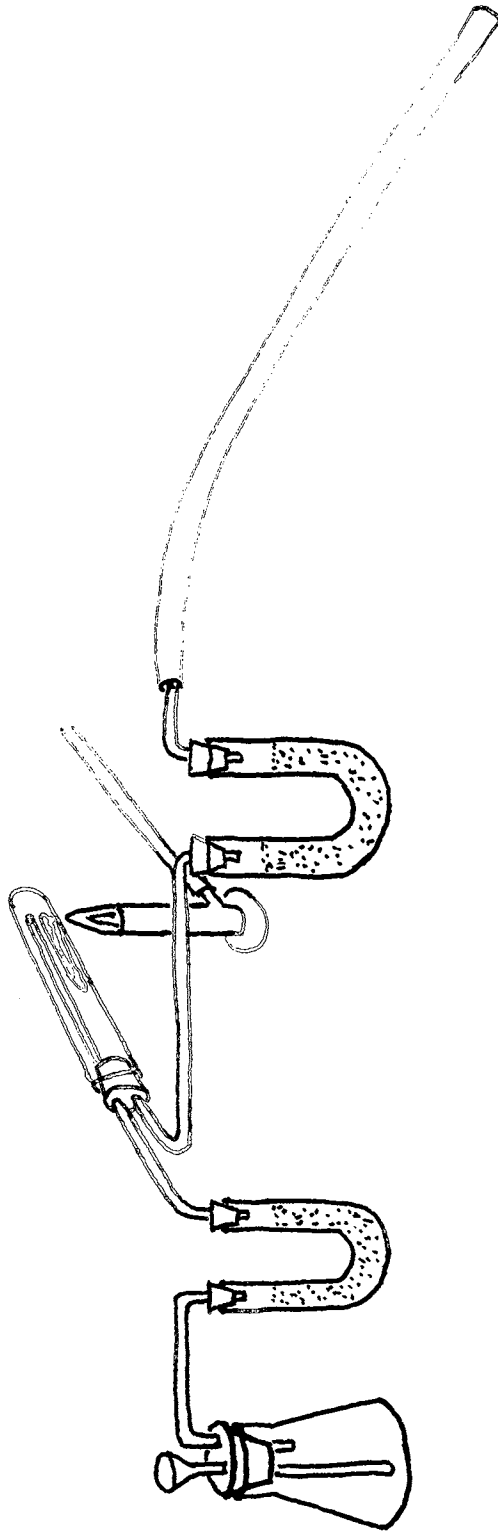[5]Fred Weisbruch, Semimicro Laboratory Exercises in High School Chemistry (Boston, 1956), pp. 69-70.

Figure 1. Apparatus for Determining Percentage Composition

Apparatus and Materials

250 ml. Bottle or flask with two hole stopper

Thistle tube

Glass tubing

U-tubes

Test tube with stopper (two hole)

Rubber tubing

Balance

Zinc

Sulfuric acid

Calcium chloride, anhydrous granular

Procedure

1. Set up the apparatus as in the diagram on the previous page.

2. Place about 20 grams of mossy zinc in the flask. The thistle tube must extend to near the bottom of the flask. The glass tube should end just below the stopper. Wrap a towel around the generator. The rubber tube from the second U-tube should extend well away from any flame.

3. Fill the U-tubes about 2/3 full of anhydrous calcium chloride. Weigh the second one as accurately as possible. Bottles may be used in the place of U-tubes. In this case one of the glass tubes should extend to the bottom of the bottle and the other should end just below the stopper.

4. Weigh the test tube. It should be dry and clean.

5. Add about five grams of finely powdered cupric oxide to the test tube and weigh again. Place the test tube into

position with the bottom end somewhat lower than the top end.

6. Be careful to see that the glass tube that delivers hydrogen into the test tube does not touch the copper oxide.

7. Add enough water to cover the zinc and the bottom of the thistle tube. Add 15-20 drops of concentrated sulfuric acid. Add more acid from time to time to maintain a slow steady production of hydrogen.

8. Wait about three minutes for the air to be driven out of the apparatus. Collect a test tube of hydrogen from the final rubber tube by water displacement. Test to see if it explodes when ignited. If it does then air is still in the apparatus. When you are sure that all the air is out of the apparatus, begin heating the copper oxide. Continue until all the copper oxide has disappeared and only copper remains.

9. At the end drive all of the water out of the test tube and glass tubing by heating them cautiously. Do not hold the Bunsen burner flame directly on the stopper except in passing. Allow to cool and stop the generator by adding water.

10. Weigh the test tube and the second U-tube.

11. Perform the calculations necessary to determine the percentage composition and simplest formula of the compounds involved.

Data and Results

  Weight of test tube and copper oxide   _____

  Weight of test tube and copper   _____

  Weight of test tube   _____

Weight of second U-tube before reaction   _____

Weight of second U-tube after reaction   _____

Weight of copper oxide used   _____

Weight of copper in copper oxide   _____

Weight of oxygen in copper oxide   _____

Per cent of oxygen in copper oxide   _____

Per cent of copper in copper oxide   _____

Simplest formula of copper oxide   _____

Weight of water formed   _____

Weight of oxygen in the water   _____

Weight of hydrogen in the water   _____

Per cent of hydrogen in the water   _____

Per cent of oxygen in the water   _____

Simplest formula for water   _____

Compare your results with the actual percentage composition of water and cupric oxide.

What are the sources of error in this experiment?

Write the equation for the reaction. Label the oxidizing agent and the reducing agent.

Why is hydrogen allowed to flow over the reduced copper while it cools?

What is the purpose of the first drying tube?

Experiment Number Ten

Determination of the Molecular Weight of a Volatile Liquid

Problem

How can we find the molecular weight of a volatile liquid.

General Approach

One of the generalizations you have learned is that one mole of a gas has a volume of 22.4 liters at standard temperature and pressure.

In this experiment you will vaporize a liquid and find the weight of known volume of the vapor at a known temperature and pressure. You will use the gas laws to calculate the volume of the vapor at standard conditions if it did not condense. Then use this volume and its weight to calculate the weight of 22.4 liter of vapor at standard conditions. This will be the gram molecular weight.

Apparatus and Materials

    Large beaker

    300 ml. Erlenmyer flask with one hole rubber stopper

    Glycerine

    Thermometer

    Barometer

Procedure

1. Fit a clean dry 300 ml. Erlenmyer flask with a one hole stopper that has been fitted with a short piece of bent glass tubing. The end of the glass tube should not extend beyond the lower surface of the stopper. Rub a thin film of
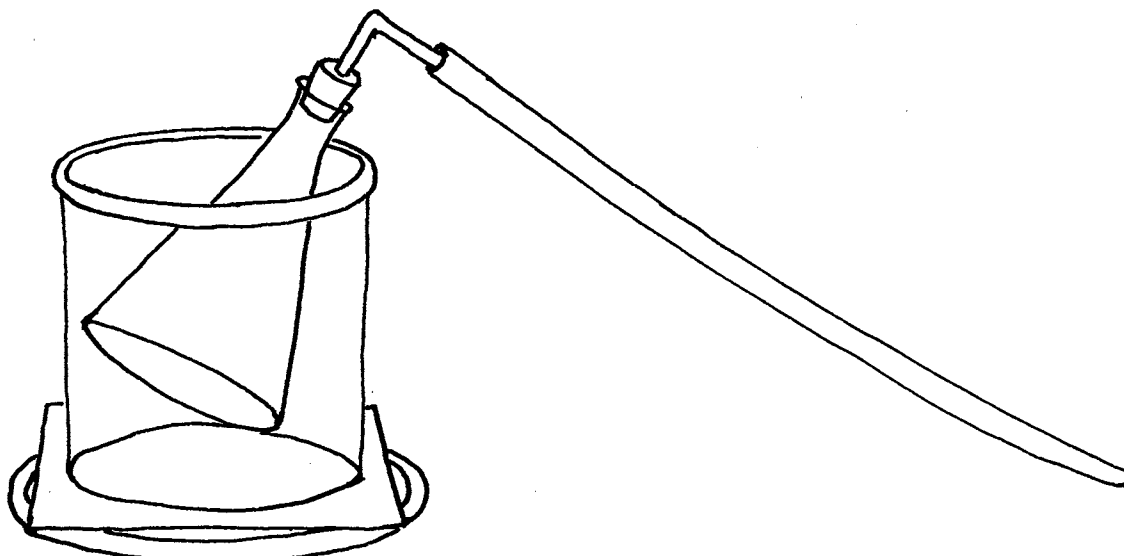
Figure 2. Apparatus for Molecular Weight Determination

glycerine over the stopper so that the stopper won't absorb some of the unknown. Attach a long rubber tube to the glass tube to carry the vapor to the floor of the room away from any flame.

2. Add about ten milliliters of a pure unknown liquid to the flask and stopper tightly. Position the flask in the beaker on a ring stand so that it can be heated. Clamp the flask in an inclined position as low as possible in the beaker.

3. Add water to the beaker so that water surrounds the flask as much as possible. Tear a hole in a sheet of paper and place it around the neck of the flask so that water can not splatter onto the stopper.

4. Heat the water to a gentle boil. The liquid in the flask will be vaporized and escape from the flask taking the air with it. There will be left only enough vapor to produce a pressure equal to atmospheric pressure. The vapor will be

at the temperature of the boiling water. Measure its temperature and record as the temperature of the vapor. Continue to heat for three minutes past the time when all the liquid in the flask and glass has vaporized. Remove the rubber tube from the glass tube.

5. Remove the flask from the boiling water and allow it to cool to room temperature. The vapor will condense. Dry the water off of the outside of the flask and weigh. This is the weight of the flask, stopper with glass tube, and condensed vapor. The unknown liquid should be poured out of the flask and the liquid wetting the flask allowed to evaporate. This evaporation can be hastened holding the flask in boiling water. Weigh the flask and stopper with glass tube when you have them cool and dry inside and out. The difference between these two weights is the weight of the vapor in the flask at atmospheric pressure and at the boiling point of water.

6. Read and record the atmospheric pressure from the barometer found in the laboratory.

7. Determine the volume of the flask by filling it with water and inserting the stopper and tube tighly into the flask so that the tube is filled with water. If you will measure the volume of this water, you will have the volume of vapor held by the flask.

Data and Results

Weight of flask and condensed vapor   _____

Weight of flask   _____

Weight of vapor in the flask   _____

Temperature of the vapor when it filled the flask _____

Pressure of the vapor when it filled the flask _____

Volume of the vapor when it filled the flask _____

Calculate the volume this vapor would occupy at $0^\circ C$. and 760 mm. Hg. pressure. Show your work.

Calculate the weight of this vapor needed to have 22.4 liters of it at standard temperature and pressure.

What is the molecular weight of your liquid? Explain.

Find out the formula of the unknown from your instructor and calculate its molecular weight from the atomic weights.

Calculate the per cent by which your answer differs from the molecular weight according to the formula.

Is it proper to neglect the weight of the air in the flask? Why?

Would the following increase, decrease, or have no effect on the molecular weight as determined in this experiment? Explain your answers.

1. The unknown was impure and part of it did not vaporize.

2. The temperature of the boiling water was actually higher than the thermometer indicated.

3. The stopper absorbed some of the unknown liquid.

4. The glass tube was not filled with water when the volume of the apparatus was determined.

Notes to the Teacher

It is desirable to weigh as accurately as possible. Weighing to the nearest 0.01 gram is good but to the nearest

0.05 gram is acceptable.

There must be good ventilation in the laboratory so that there will not be a dangerous build up of combustible vapor. Be very cautious.

There are a number of organic liquids with boiling points lower than water that are suitable as unknowns. Some of these are carbon tetrachloride, benzene, ethyl acetate, methyl acetate, acetone, chloroform, cyclohexane, cyclopentane, and ethyl iodide.

Experiment Number Eleven

Combining Volumes of Gases[6]

Problem

How can we use Gay-Lussac's law of combining volumes of gases and Avogadro's hypothesis to determine the product of a chemical reaction between gases?

General Approach

Gay-Lussac showed that when some or all of the reactants or products of a reaction are gases, the volumes of the gases are in the ratio of small whole numbers. The volumes must be measured at the same temperature and pressure. Thus one liter of nitrogen will react with three liters of hydrogen to produce two liters of ammonia under the proper conditions.

Avogadro explained this by saying that equal volumes of gases contain equal numbers of molecules at the same temperature and pressure. It takes three liters of hydrogen to react with one liter of nitrogen because it takes three molecules of hydrogen to react with one molecule of nitrogen to produce two molecules of ammonia.

In this experiment you will react two gases (oxygen and nitric oxide). By measuring the volumes of each that are used, you will know how many molecules of one it takes to react with the other. By measuring the volume of the product,

_____

[6]J. R. Schwench and Raymond Martin, Basic Principles of Experimental Chemistry (Englewood Cliffs, 1958), pp. 136-139.

you will know the relative number of molecules of the product produced. This is enough information to deduce the formula of the product and write the balanced equation for the reaction.

Apparatus and Materials

One liter flasks or bottles with stoppers

Glass and rubber tubing

Pinch clamps

Test tube and stopper (one hole)

Pneumatic trough

Potassium chlorate

Manganese dioxide

Sodium Nitrite

Hydrochloric acid

Procedure

1. Assemble the apparatus (see figure 3) for generation of oxygen and collecting it by displacing water. Use a one liter flask and fill it full of water. Invert it over the end of the rubber tube so that oxygen will displace the water.

2. Place about six grams of potassium chlorate and two grams of manganese dioxide in the test tube. Mix by shaking.

3. Heat the mixture gently allowing the first gas produced to escape in order to flush the air out of the flask. Collect a flask of oxygen, stopper it, and save for later use.

4. Next prepare nitric oxide (NO). Place 25 grams of sodium nitrite and 30 ml. of water in a 250 ml. flask. Fit a two hole stopper on the flask. Through one hole there should
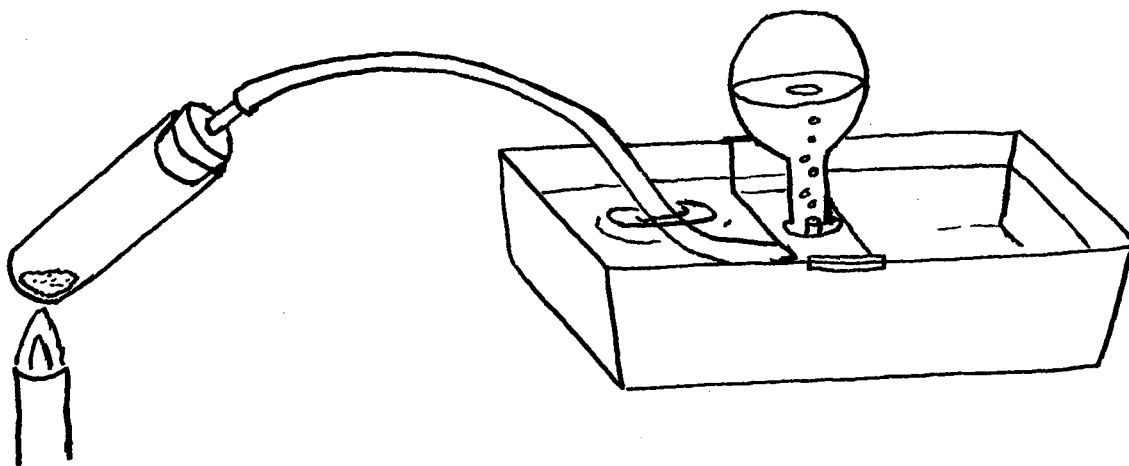
Figure 3. Apparatus for Generation of Oxygen
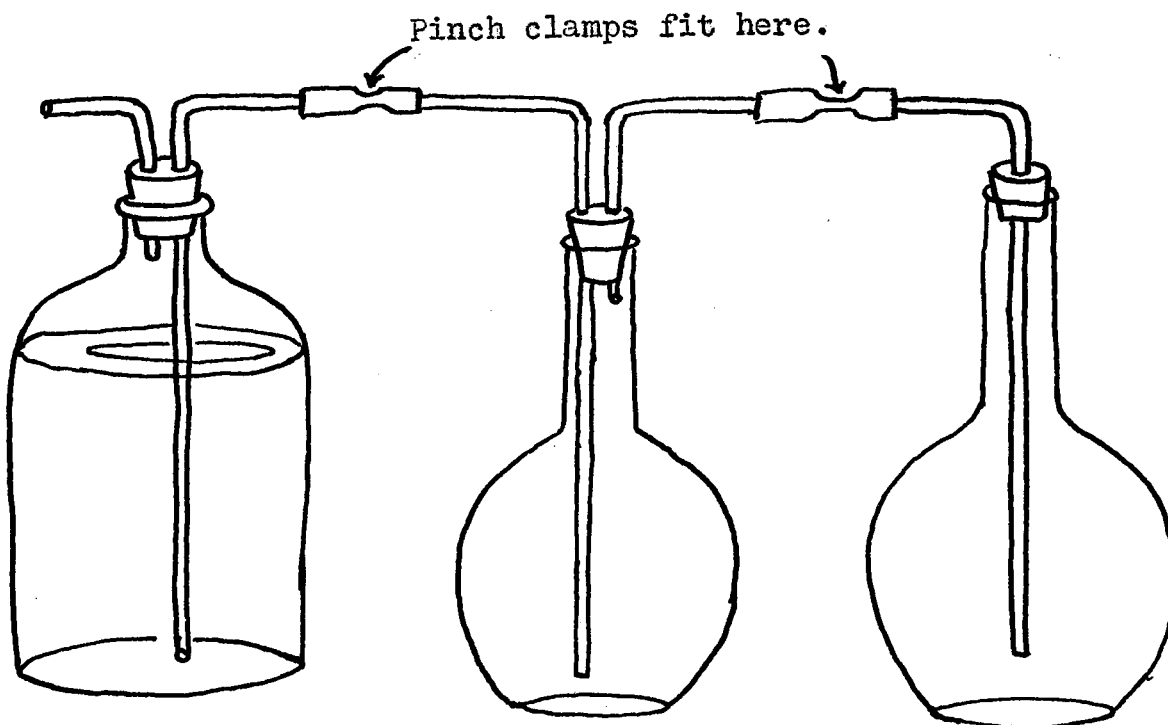
Pinch clamps fit here.

Figure 4. Apparatus for Reacting Gases

be a thistle tube that extends down below the level of water in the flask. Through the other hole in the stopper there should be a short glass tube that connects to a rubber tube that acts as a delivery tube. Place under a hood if one if available.

5. Add about 25 ml. of 6N hydrochloic acid through the thistle tube. A mixture of colorless NO and brown NO_2 is produced.

6. When the air has been swept out of the generator, bubble the gas into an inverted one liter flask full of water in your pneumatic trough. The NO_2 will dissolve in the water and the insoluble NO will fill the flask. Do not breathe these gases.

7. Remove the delivery tube immediately when the flask is filled with the colorless NO and stopper it quickly. Leave the generator under the hood or add water and a dilute base to kill the reaction.

8. Prepare the apparatus as in figure 4. The pinch clamps should be in place keeping the connecting tubes closed. Notice that the last stopper is one holed. Use a bottle that will hold about 1000 ml. of water.

9. The oxygen and nitric oxide flasks should be put in place quickly to aviod escape of gases. The oxygen is placed in the center next to the water bottle. The stoppers must be tight.

10. Remove the pinch clamps and blow gently on the short glass tube of the water bottle. This will force water over

into the oxygen bottle which in turn forces oxygen over into
the nitric oxide flask where it reacts with this gas. How do
you know that a reaction is taking place?

11. Continue gently forcing water into the oxygen flask
as long as gentle pressure suffices.

12. Allow the apparatus to set for a minute to equalize
the pressure and temperature throughout. Then place the
pinch clamps in place.

13. You must now determine the volume of oxygen, nitric
oxide, and product involved in the reaction. You may assume
that all the nitric oxide has taken part in the reaction if
you has any oxygen left over at the end of the reaction.

14. The volume of oxygen taking part in the reaction is
equal to the volume of water that has been forced over into
the oxygen flask. Why? Measure this with a graduated cylin-
der.

15. Since the product is a colored gas, you can see the
volume occupied by it. Do not breathe this gas. Fill the
flask with water to the level of the product and measure the
volume of the water in order to get a numerical value.

16. The volume of nitric oxide you began with is not
exactly one liter if you used a one liter flask. Fill the
flask full of water and measure the volume of water it holds.

Data and Results

    Volume of nitric oxide used in the reaction   _____

    Volume of oxygen used in the reaction   _____

    Volume of product formed   _____

Determine the ratio of the volumes by dividing each of the above volumes by the smallest of the three. These should be very close to whole numbers. If so, round off to the whole numbers. If not, consult your instructor. Show your work here.

The ratio of the volumes of oxygen, nitric oxide, and product is ___:___:___. Therefore, what is the number of molecules of oxygen, nitric oxide, and product in the balanced equation.

___ O_2 + ___NO ⟶ ___product
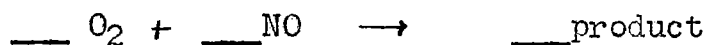
You should now be able to figure out the formula of the product. Write the complete balanced equation for the reaction below.

What other evidence do you have that you formula for the product is correct?

How do you know a reaction took place when the oxygen and nitric oxide were mixed?

The ratio of the volumes probably was not exactly small whole numbers. This indicates errors. What are some sources of error in this experiment? What remedies can you suggest?

What assumptions must be made in order for this experiment to be valid?

Notes to the Teacher

The reaction involved in this experiment is

O_2 + 2NO ⟶ 2NO_2
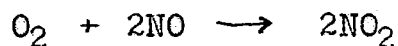
This equation indicates that the product should just fill the flask that the nitric oxide was in. The volume of oxygen

used should be about 500 ml.

If the students have not had previous experience at making oxygen or collecting a gas by water displacement, then they will need more instruction than found in the experiment.

This experiment assumes the validity of Avagadro's hypothesis and Gay-Lussac's law and that there is only one product whose volume is not less than that of the nitric oxide.

Experiment Number Twelve

Effect of Amount of Catalyst on Reaction Rate[7]

Problem

What is the relationship between the concentration of catalyst and the rate of reaction? If we double the amount of catalyst, does the reaction go twice as fast?

General Approach

The decomposition of hydrogen peroxide to water and oxygen is catalyzed by manganese dioxide. The rate of the reaction depends upon several variables such as temperature, concentration of hydrogen peroxide, and concentration of catalyst. In this experiment you will keep the first two constant and determine the effect of changing the concentration of the catalyst on the rate of reaction. You will measure the rate of the reaction by measuring the rate at which oxygen is produced.

Apparatus and Materials

Large test tube with a two-hole stopper

Thistle tube

Glass and rubber tubing

Pneumatic trough

Test tubes

Ring stand with test tube clamps

---

[7]Elbert C. Weaver, Scientific Experiments in Chemistry (Washington, D. C., 1957), HS-T73.
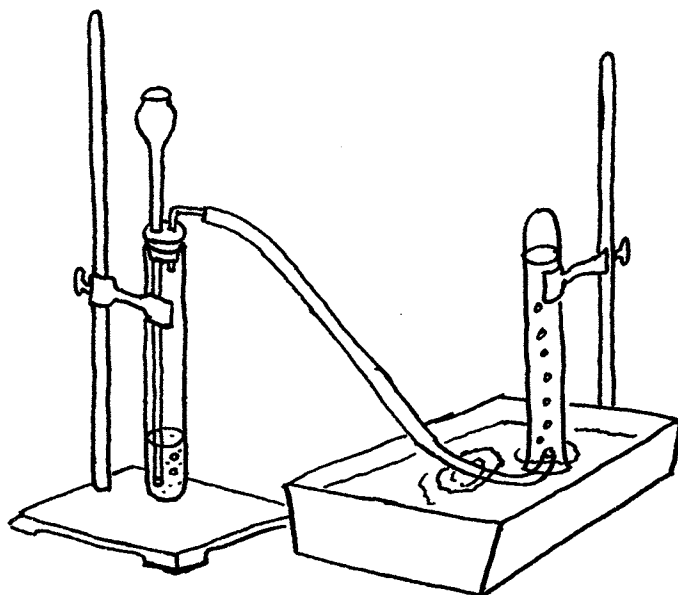
Figure 5. Apparatus for Decomposition of Hydrogen Peroxide

Stopwatch or clock with second hand

Small graduated cylinder or 10 ml. pipette

Hydrogen Peroxide solution

Manganese Dioxide

Procedure

1. Assemble the apparatus as shown in the diagram. Protect your hand with a towel when inserting the tube into the stopper. Lubricate the glass with glycerine before inserting. Both test tubes can be supported from the same ring stand. The thistle tube should extend down to the bottom of the test tube.

2. You will need portions of manganese dioxide as follows: 4.0, 2.0, 1.0, 0.5, 0.25, 0.125, 0.63, and 0.031 grams. If a suitable balance is not available, then weigh out eight

grams. Divide this into equal parts with a spatula; divide one of these portions in half, and so on until you have piles of 4, 2, 1, 1/2, 1/4, 1/8, 1/16, and 1/32 gram.

3. Place the four gram portion in the test tube and put the stopper and thistle tube in position. Place the delivery under the mouth of a test tube which is full of water and inverted in the pneumatic trough so that the oxygen produced will be collected by displacement of the water.

4. Add 10 ml. of 1% hydrogen peroxide solution to the thistle tube. Note the time to the second. After collecting oxygen for exactly two minutes remove the rubber tube from the mouth of the inverted test tube.

5. Mark the volume of gas collected in the test tube with a wax pencil or a rubber band. Remove the test tube and fill with water to the pencil mark. Measure the volume of this water with a graduated cylinder. This is the volume of the gas you collected. But this includes 10 ml. of air that was displaced from the generator when you added the peroxide. Subtract the 10 ml. to get the volume of oxygen produced. Record these results on the data sheet.

6. Repeat these steps for each of the other portions of manganese dioxide. Clean and dry the oxygen generator between runs. Record data.

7. Make a graph of the data in the table. Let the weight of manganese dioxide be the horizontal axis and the volume of oxygen produced be the vertical axis.

Data and Results

Weight of MnO$_2$	Volume of Oxygen Produced
4 grams	
2 grams	
1 gram	
1/2 gram	
1/4 gram	
1/8 gram	
1/16 gram	
1/32 gram	

Does it appear from a graph of the data above that doubling the concentration of the catalyst doubles the rate of reaction?

What interpretation can you give as to the relationship?

Predict by using your graph the volume of oxygen you would get if you used 1.5 grams or 0.75 gram of manganese dioxide in 10 ml. of the peroxide solution. Carry out the experiment and describe how the result compares with your prediction.

Why do you not have to correct the gas volumes to standard temperature and pressure or for water vapor pressure?

Notes to the Teacher

You should try out this experiment before the students so it in order to make sure that the test tube is large enough to hold the oxygen produced with the four grams of catalyst.

You can purchase the peroxide at a drug store. Dilute

Volume
of
Oxygen
in ml.
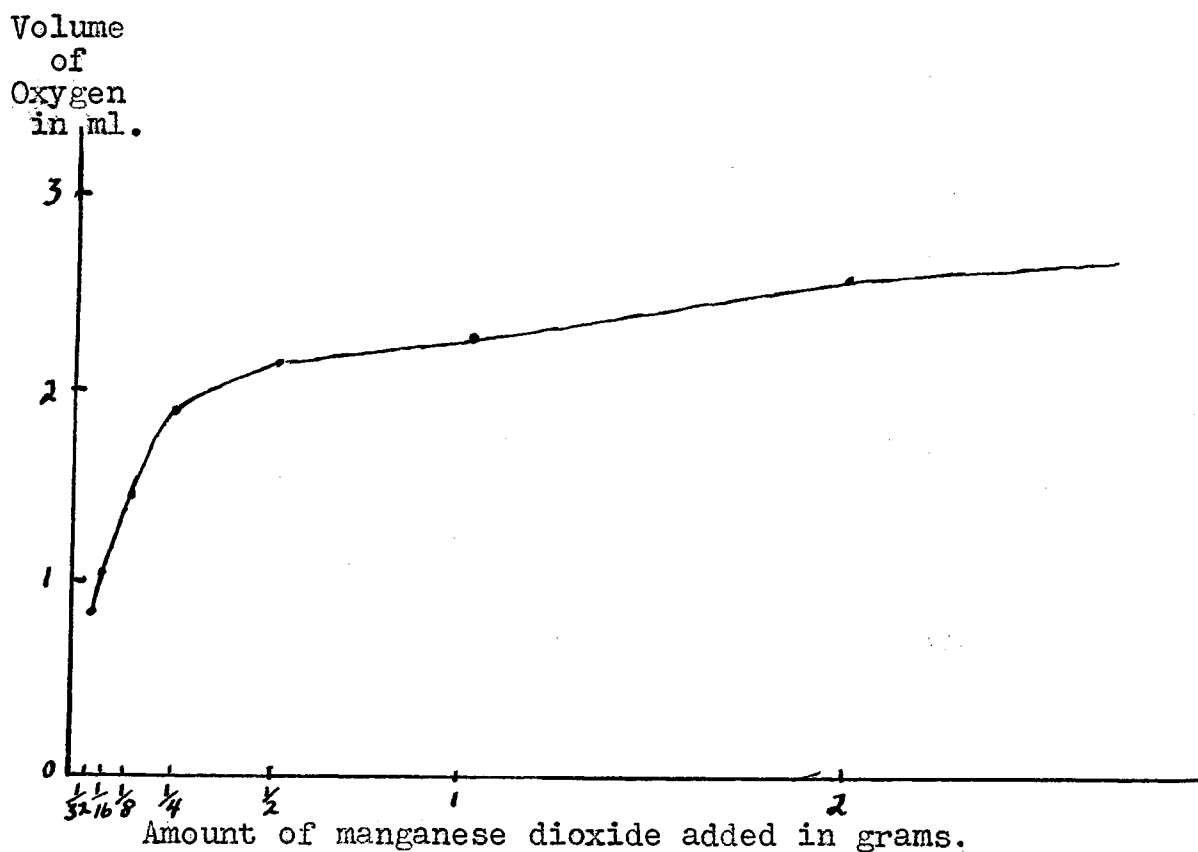

Amount of manganese dioxide added in grams.

Figure 6. Typical Catalyst Activity Curve

it to about 1% concentration.

If you have one arm test tubes these will be preferable to use as oxygen generators. A small inverted graduated cylinder can be used to collect the oxygen and the volume read directly.

The graphs of the students should be similar to the one given in figure 6. Some students will draw the curve through the origin. This is a chance for the teacher to emphasize that the reaction takes place in the absence of the catalyst. The catalyst just speeds the reaction.

CHAPTER III

CONCLUSION

The purpose of this report was to devise a number of quantitative experiments suitable for high school chemistry. These experiments should involve important concepts and be capable of being done with simple apparatus in the short class period of the high school. The experiments were difficult to find and develop. Very few experiments were found in high school chemistry manuals. Experiments used in college chemistry require too much time and involve apparatus not available to the high school chemistry class.

This report has included twelve experiments developed and adapted by the author. They are interesting and challenging but not too difficult. The ideas involved are important and well developed in the experiments. Several of them are open ended in that they raise other questions that the class or individuals may wish to pursue further.

BIBLIOGRAPHY

Campbell, J. A. "Preparation for Graduate Work." Journal of Chemical Education, XXXVII (February, 1960), 71.

Drago, Russell S., and Theodore L. Brown. Experiments in General Chemistry. Boston: Allyn and Bacon, 1960.

Evans, Maurice, Norris Bush, and Ray Easley. Laboratory Problems for Living Chemistry. Boston: Ginn and Company, 1957.

Gabriel, Henry. "Quantitative Work in Freshman Chemistry." Journal of Chemical Education, XXXIII (January, 1956), 42-44.

Garrett, Alfred, Joseph Haskins, Thor Rubin, and Frank Verhoek. Chemistry for the Laboratory. Boston: Ginn and Company, 1957.

Goff, Aaron. "Emphasizing the Unknown." Science Teacher, XXII (November, 1955), 291-292.

Hurd, Paul. Science Facilities for the Modern High School. Stanford: Stanford University Press, 1954.

Schwench, J. Rae, and Raymond M. Martin. Basic Principles of Experimental Chemistry. Englewood Cliffs: Prentice-Hall Inc., 1958.

Scientific Experiments in Chemistry. Washington: Manufacturing Chemists' Association, Inc., 1958.

Weisbruch, Fred T. Semimicro Laboratory Exercises for High School Chemistry. Boston: D. C. Heath Company, 1956.

VITA

James Douglas Barlow

Candidate for the Degree of

Master of Science

Report: QUANTITATIVE EXPERIMENTS FOR HIGH SCHOOL CHEMISTRY

Major Field: Natural Science

Biographical:

Personal Data: Born near Ramona, Oklahoma, October 15, 1934, the son of Clarence Earl and Edna Doherty Barlow

Education: Attended grade school in Ramona and Collinsville, Oklahoma; graduated from Collinsville High School in 1952; received the Bachelor of Arts degree from the Oklahoma State University, with a major in Chemistry, in May, 1957; completed requirements for the Master of Science Degree in May, 1961.

Professional Experience: Taught at Stillwater High School, Stillwater, Oklahoma, 1957-1960; member of National Science Teachers Association, National Education Association, and Phi Kappa Phi.