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Summary: In recent years there has been a great deal of controversy concerning the education of our better students. This report is designed to identify the academically talented student in chemistry. It also presents laboratory experiments which will motivate and inspire the bright student.

A talented child is one with high learning ability. The generally accepted method of recognizing the academically talented student in chemistry is the use of intelligence quotient tests, class records, teacher evaluation, interest in science, special abilities, and individual characteristics.

The experimental write-up is an important part of the experiment. The student can collect all the information he has obtained in the experiment and determine its accuracy and limitations. The errors and reasons for errors are recorded along with the practical application of the experiment.

The form of the experiments requires the student to do more individual reasoning and outside reading than does the ordinary laboratory experiment. The experiments are selected from all areas of chemistry and equal emphasis is placed upon quantitative, qualitative, and theory.

The equipment necessary to carry out these experiments is available in any well-equipped laboratory. In some cases it is more desirable to have the students devise and make their own equipment. In this way a greater understanding and stimulation will be gained.

Possibilities for enrichment in chemistry is restricted only to the resourcefulness and energy of the teacher and his colleagues.

ADVISER'S APPROVAL ______________________

CHEMISTRY EXPERIMENTS FOR THE ACADEMICALLY TALENTED HIGH SCHOOL STUDENT

By

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CHEMISTRY EXPERIMENTS FOR THE ACADEMICALLY TALENTED HIGH SCHOOL STUDENT

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CHAPTER I

INTRODUCTION

The teacher of the gifted child has a wide choice of methods and materials for enrichment in the field of chemistry. It is the purpose of the writer to emphasize a particular method of experimentation and to develop laboratory experiments which will give the gifted student an opportunity to do more than the so-called "cook-book chemistry."

In this report special attention will be placed on the selection of students for an accelerated course in chemistry. The main emphasis of the report will be to consider a method of laboratory write-up and to select a group of laboratory chemistry experiments with the academically talented student in mind. A list of all materials and apparatus needed for each experiment will also be included.

CHAPTER II

REVIEW OF THE LITERATURE

Definition of the Academically Talented Student

In the past it has generally been concluded that bright students, or gifted students, were eccentric, different, or emotionally unstable. Contrary to this belief, recent studies and reports indicate that the gifted child is physically stronger, socially more secure, and emotionally more stable than the average child of his own chronological age. In addition, he is more alert and eager to learn.[1]

There is frequently a great deal of confusion when one tries to define the academically talented student. In general it has been known that the academically gifted child is one with a high measure of competence in learning ability. The best definition that has been offered is: "the academically talented child is one with great natural intellectual endowment capable of assimilating regular school material at a rapid rate. He never reaches the saturation point, intellectually speaking."[2]

The superior student could possibly be defined as possessing three major characteristics. First, the superior student has a

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[1] Paul Witty, The Gifted Child. (Boston: D. C. Heath & Co. 1951) p. 272.

[2] Roberta Kellogg, "Skills Instruction for the Gifted Child in the Regular Classroom," National Elementary Principal, Vol. XXIX, No. 3, (December, 1949), p. 37.

greater time-span. This means that he can spread his efforts over a longer period, depending less than others upon frequent examination of his results. Secondly, his range of learning is greater. He can maintain an attention which has a wider spread and yet keep all details significantly related to a central purpose. Thirdly, there is continuity in the efforts of the superior mind, a relatedness in the successive acts of attention in which his effort is expended.[3]

Recognizing the Academically Talented Student in Chemistry

The academically talented student may be difficult to recognize and to evaluate. Many different methods have been used to identify or group these persons. The general accepted method for grouping has been based on the results of intelligence quotient tests. In a study during a conference sponsored jointly by the National Education Association and the National Science Teachers Association, it was accepted that the term "academically talented" refers to a student with an intelligence quotient of 120 or more. The group thus included would comprise the top 15 to 20 percent of the high school students in the United States.[4]

Many other standardized tests have been used to identify the rapid learner. They are especially helpful in identifying the shy student or the one who has a language handicap. Some of the available

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[3] Harold E. B. Speight, "Who is the Superior Student," School and Society, Vol. XLVIII, (October, 1938), p. 549.

[4] National Education Association Project on the Academically Talented Student and the National Science Teachers Association. Science for the Academically Talented Student in the Secondary School. (Washington, D. C. National Education Association, 1959), p. 10.

tests measure verbal comprehension, mechanical reasoning, abstract reasoning, spatial visualization and interest. In the above mentioned tests, power tests are superior to speed tests. Ability to solve difficult problems seems to be a better predictor of scientific success than speed in solving simple ones.[5]

The talented student can also be recognized by other characteristics which should enable him to do superior work in the field of chemistry. Extraordinary memory, intellectual curiosity, persistency, insight into abstractions, the ability to do abstract thinking on a high level, the ability to translate data into generalizations, and the ability to apply knowledge to new situations have been used successfully by many teachers in identifying the student with outstanding ability in science.[6]

It has been emphasized that no single test is adequate in recognizing academically talented students. Multiple factors should always be considered in the identification of able chemistry students. These should include general intelligence and other standardized tests scores, past class records, teacher evaluation, expressions of students' interest in science, individual characteristics, and special abilities.[7]

Experimental Write-ups

The academically talented student must be able to communicate.

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[5] Kenneth E. Brown and Phillip G. Johnson, Education for the Talented in Mathematics and Science. U. S. Dept. of Health, Education, and Welfare, Office of Education, Bulletin 1952, No. 15. (Washington, D. C.: Superintendent of Documents, Government Printing Office, 1953.) p. 3-4.

[6] Ibid., p. 6.

[7] National Education Association Project on the Academically Talented Student and the National Science Teachers Association, p. 11.

Written reporting should receive special emphasis, and the student should be learning these skills as he learns to become a scientist. The high school science major should be writing the kind of reports that will later be required of him on the job.

The student should have many laboratory experiences in which he is called upon to act and think as a scientist and to attempt scientific approaches to the solution of problems. The student should understand the role of planned experimentation, observation and recording of data.

Many so-called "open-ended" experiments are now available in the area of chemistry. For instance, when mercuric oxide is used in connection with a discussion of Priestly's discovery of oxygen, students commonly generalize incorrectly that oxides as a group are unstable. One typical open-ended experiment raises the question, "Do Oxides as a Class Release Oxygen When Heated?" Experiments such as this give the student an opportunity to do personal research from which he will gain valuable information.

Writing experimental reports is very essential for the academically talented student. Clarity and honesty must be stressed in reporting the individual project. Early in its pursuit, the problem should be stated concisely, with its limitations and delimitations. The methods should be outlined before the project is begun. The final report should include as its basis a clear statement of the problem and a detailed account of the methods followed, including statistical treatment of the data, results, and conclusions.[8]

A good experimental write-up should include these major points:

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[8] National Education Association Project on the Academically Talented Student and the National Science Teachers Association, p. 33.

I. Title of the Experiment
II. The Problem
III. Apparatus and Materials
IV. Diagram of Apparatus
V. Procedure
VI. Results
VII. Conclusions
    A. Error
    B. Percentage Error
    C. Reason for Error
    D. Practical Application of Experiment
    E. Questions

The laboratory reports should be neat and well written. Experiments which are typewritten or in ink are more easily graded. Written laboratory reports are an important part of the chemistry program and warrant a good deal of the teacher's time in grading and correction.

Experiments in Chemistry

1. Mixtures and Compounds

DISCUSSION

The purpose of this experiment is to illustrate some differences between mixtures and compounds. Mixtures are made up of substances present in varying proportions. Therefore a property such as the color of a mixture may vary with the relative amounts of the substances present. Compounds contain two or more elements that are chemically combined in fixed proportion by weight. The properties of a compound, therefore, are fixed. Mixtures can be separated by mechanical or physical methods; compounds can not.

EXPERIMENTAL

APPARATUS AND MATERIALS

Test tubes, watch glass, Bunsen burner, balance, 1 g. of mercuric oxide, 2 g. of sodium chloride, 0.5 g. of granular zinc, 4 g. of granular lead, 1 g. of sulfur, 20 ml. of carbon disulfide.

PROCEDURE

1. Mixtures

Put a small amount of mercuric oxide in one test tube and about the same volume of sodium chloride in another test tube (approximately the volume of a grain of rice for each substance). Add 10 ml. of water to each; shake, and record which is soluble. Into a third test tube put similarly small amounts of each of the two solids, to make a mixture of the two. Add 10 ml. of water, and shake well. Allow the solid to settle; then separate it from the liquid by decantation. Pour about 1 ml. of the liquid onto a watch glass, and evaporate over a steam bath. Record the results.

Put a small particle of granular zinc into one test tube, and a similar amount of sulfur into another (approximately the volume of a grain of rice for each substance). Add 5 ml. of 6 N hydrochloric acid (dilute HCl) to each test tube, and watch to see whether there is any reaction, as shown by the evolution of gas bubbles and the disappearance of the solid. Now add 5 ml. of 6 N hydrochloric acid to a test tube containing a mixture of rice size amounts of granular zinc and sulfur. Observe what is happening long enough to decide whether this method can separate one solid from the other.

2. Compounds

Put a particle of granular lead, about half the size of a grain of rice, and approximately the same amount of sulfur into another test tube. Using a hood, and keeping the liquid away from flames, add 10 ml. of carbon disulfide. Be very careful in handling carbon disulfide; it is poisonous and flammable. Shake well until the sulfur is completely dissolved; then separate the solid from the solution by decantation. Pour 1 ml. of the solution onto a watch glass, and set it in the back of the hood to evaporate. Keep it away from flames. What is the residue on the watch glass? Pour the waste carbon disulfide into a bottle designated for that purpose; do not pour it into the sink.

Next, weigh out 2 g. of granular lead and 0.3 g. of sulfur, using a watch glass on the balance pan. Mix them in a test tube and, in a hood, heat the test tube gently, maintaining a glow in the solid if it develops. After all appearance of reaction ceases, heat strongly for five minutes. If any sulfur collects in the cooler part of the tube, heat the tube to drive off or burn the sulfur. Allow to cool. Note the difference in appearance of this solid compared with lead and with sulfur. Using a hood, add 10 ml. of carbon disulfide, to see whether it is still possible to separate the two elements. Pour 1 ml. of the liquid onto a watch glass to evaporate, at the back of the hood as before. Note whether any solid is left after the evaporation.[9]

QUESTIONS

1. Which of the two solids (HgO, $NaCl$) react with HCl?
2. Which of the two solids could be recovered from the mixture unchanged?

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[9] Saverio Zuffanti, Arthur A. Vernon and W. F. Luder, A Laboratory Manual of General Chemistry (Philadelphia and London, 1955), pp. 41-42.

2. Heating Metallic Oxides

DISCUSSION

The purpose of this experiment is to determine the stability of metallic oxides. Find out by testing the gas above the heated solid, to determine whether or not oxygen is given off. It may be possible to tell if a change occurs by the appearance of the oxide before and after heating, and if the splinter glows more brightly or not.

EXPERIMENTAL

APPARATUS AND MATERIALS

Ignition tube, test tube holder, Bunsen burner, matches, splinters, meruric oxide, zinc oxide, copper oxide, magnesium oxide, iron oxide, aluminum oxide, lead dioxide, manganese dioxide, red lead, magnetite, titanium dioxide, lead oxide, (some of these or other oxides).

PROCEDURE

Select an oxide. Take a sample of it about the size of a grain of rice. Place the oxide in an ignition tube. Using a test tube holder, place the tube in the upper part of a Bunsen flame and heat the tube about two minutes.

Test the region above the solid and within the test tube for the presence of oxygen, using a lighted splinter.

Do not heat the oxide long enough to cause the glass to soften and the oxide to stick to the glass.

Repeat the experiment with another oxide. Make a careful record in the writing of results. Heat as many oxides as time permits, recording the effects noticed in all experiments.[10]

QUESTIONS

1. Is there any relationship between the number of atoms of oxygen in the formula and the readiness with which the oxides decompose?

2. From the formulas, calculate the percentage by weight of oxygen in each oxide that was used. Is there any relationship between the percent of oxygen and the ease of decomposition?

3. Look at an Electrochemical Replacement list and find the location of elements whose oxides decompose to give the metal and free oxygen, and those that did not. State any relationship that is found.

4. From results of this experiment, predict how metallic oxides not used should behave when heated. If possible, test an oxide not used previously and verify the conclusion.

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[10] "Scientific Experiments in Chemistry," Student guide HS-S71, Manufacturing Chemists Association, Inc. (Washington, D. C., 1957).

3. Reactions Between Oxides and Water

DISCUSSION

The problem in this experiment is to find which oxides react with water and if they do react, what sort of a compound is formed. In this experiment it would be beneficial to make three classifications: acid-forming, base-forming and no-reaction. When this experiment is completed, make a simplified diagram of a periodic table, using only the elements available in the experiment.

EXPERIMENTAL

APPARATUS AND MATERIALS

Watch glass, burner, matches, deflagrating spoon, red and blue litmus, medicine dropper, glass rod, jar, asbestos, aluminum oxide, barium oxide, calcium oxide, lead oxide, magnesium oxide, manganese dioxide, titanium oxide, mercuric oxide, phosphorus, sulfur, and carbon.

PROCEDURE

Take a sample of the oxide, about the size of a grain of rice, and place it in the center of a watch glass. Add pure water from a medicine dropper until you make a small puddle around the sample. Stir the mixture with a glass rod. Using red and blue litmus paper, test the liquid around the oxide and note any changes and record observations. Before going on to another oxide, make sure watch glass is clean.

To obtain the oxides of phosphorus, carbon and sulfur, place a sample of the basis element about the size of a grain of rice in the bowl of a deflagrating spoon lined with asbestos. Ignite and place in a gas bottle containing a half-inch layer of water. When the burning stops, remove spoon and shake contents of the bottle. With the aid of a glass stirring rod, test the solution on red and blue litmus paper.[11]

QUESTIONS

1. Predict from the conclusions how an oxide not used might react with water. If possible, obtain a sample and test prediction.
2. From the formulas and the results obtained, find the position of the elements of oxides used in the electro-chemical series of elements. Is there a relationship between your results and the position of the elements in this list?
3. When the oxides of phosphorus, carbon and sulfur were combined with water, what compounds were formed?

Not all oxides react with water. The deposits of petrified wood were probably caused this way.

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[11] "Scientific Experiments in Chemistry," Student guide HS-S72, Manufacturing Chemists' Association, Inc., (Washington, D. C., 1957).

4. Determining Activities of Metals

DISCUSSION

The relative activities of many metals can be inferred from observations of the rapidity with which they reduce the hydrogen ion and release hydrogen from acids. Since the rate of hydrogen evolution depends on both the strengths of the acid and the activity of the metal used, one can compare extremely active metals, such as sodium and potassium, by their action on an extremely weak proton donor, water. The use of sodium and potassium in the first part of this experiment should be demonstrated by the instructor only. Somewhat less active metals require water at high temperatures or acids stronger than water for appreciable hydrogen evolution. For those metals which are incapable of releasing hydrogen from acids, comparisons may be based on their reducing actions on cations. For instance, silver is known to be a stronger reducing agent than gold because it liberates metallic gold from solutions of gold salts.

When you have completed experiment arrange the metals tested in order of their activity with the most active element first.

Divide the experiment into two parts allowing an hour on each part.

EXPERIMENTAL

APPARATUS AND MATERIALS

Filter paper, test tubes, Bunsen burner, red and blue litmus, 250 ml. flask, sodium, calcium, magnesium ribbon, granulated magnesium, iron filings, zinc strips, copper strips, aluminum, tin, 0.1 M tin chloride, 0.1 M copper sulfate, 0.1 M silver nitrate, 0.1 M mercuric nitrate, silver coin, mercury.

PROCEDURE

(To be done by instructor): Wrap a piece of sodium in a small bit of dry filter paper. Fill a test tube with water and invert it in a beaker of water. Quickly insert wrapped sodium into mouth of the test tube, keeping the test tube inverted in the beaker of water. After reaction has stopped test the combustibility of the gas. Test the solution in beaker with red and blue litmus. (The remainder to be done by the student):

Drop a small piece of freshly cut calcium into test tube containing 5 ml. of water. Close the end of the tube with the thumb until considerable gas pressure develops, then test gas for combustibility. Test the solution with litmus.

Place about 0.5 g. of granulated magnesium in 100 ml. of distilled water contained in a 250 ml. flask. Close the flask with a stopper equipped with a delivery tube which leads to a small test tube inverted in a beaker of water. Boil the water in the flask and collect the escaping gas in the test tube. When tube is half full of gas test it for combustibility and test solution in flask with litmus.

Hydrochloric acid is a far stronger acid than water. This means that certain metals which react slowly or not at all with water, will replace hydrogen rapidly from hydrochloric acid. It is the hydrogen ion that oxidizes the metals and its concentration is far greater in acid than in water.

Place in separate test tubes a piece of zinc, copper, aluminum, (which has been cleaned with sand paper), tin foil, iron wire, and a piece of magnesium. The metals should all be about the same size. Add to each test tube 5 ml. of dilute 1 M hydrochloric acid. Observe rates of gas evolution and arrange in order of their activity.

For those metals which do not displace hydrogen from hydrochloric acid, a comparison of their chemical activity may be observed in their ability to displace other metals from aqueous solutions of their salts.

Hold a bright piece of copper in 5 ml. of a 0.1 M solution of tin chloride. Record your observations. Now hold a bright piece of tin foil in a 0.1 M solution of copper sulfate. Record your observations. Determine which is more active, copper or tin.

Using silver, mercury and 0.1 M solution of silver nitrate and mercuric nitrate, determine activity of these two metals.

Devise a test to determine the order of activity of the metals silver, mercury, tin and copper. Carry out these tests and note the results.[12]

QUESTIONS

1. Write equations for each one of the reactions in this experiment.
2. Arrange all the elements tested in order of their activity with the most active first.
3. If there are discrepancies between your results and the theoretical results explain the possible deviation.

When copper is placed in a concentrated solution of hydrochloric acid, nothing happens. However, when placed in a dilute solution of nitric acid, a reaction takes place.

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[12] William Hered and William H. Nebergall, Basic Laboratory Studies in College Chemistry, (Boston, 1957), p. 40.

5. The Relationship Between Volume and Pressure of a Gas

DISCUSSION

When the temperature is maintained at a constant value, the volume of a gas varies inversely as the pressure; that is, doubling the pressure will halve the volume; if the pressure is increased threefold, the volume becomes one third of the original volume, and so on. The relationship may be expressed mathematically as $PV = K$ where K is a constant. In other words, for a given mass of a certain gas at constant temperature, the product of pressure and volume should be a constant. As one variable is changed, the other must change in such a way that the product is a constant value. All gases do not conform exactly to this relationship, but approximate it quite closely.

EXPERIMENTAL

APPARATUS AND MATERIALS

Glass tubing 10 mm. in diameter, one meter long, barometer, meter stick, mercury.

PROCEDURE

Shape a piece of glass tubing in the form of a (J). Close off the short end of the tube and add just enough mercury to the tube to fill the bent portion. Adjust mercury levels so they are the same height in either arm. Measure with a meter stick the length in mm. of the confined gas. The pressure of the confined gas is the same as the atmospheric pressure since the levels are the same. Read the barometric pressure and record results.

Add more mercury to the long arm of the tube. By measuring the column of air in the short arm, determine its new volume. The new pressure is found by taking the difference of the mercury levels and adding this to the barometer reading.

Repeat this step several times and tabulate the results. Calculate the product of the pressure and volume in each instance and record.[13]

QUESTIONS

1. How would the change in pressure affect the density of a gas?
2. How is a barometer used in the laboratory?
3. How can Boyles Law be applied in the laboratory?

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[13] G. Brooks King and William E. Caldwell, Laboratory Experiments in General Chemistry (New York, 1956), p. 67.

6. Determination of the Formula of a Hydrate

DISCUSSION

Many salts that have been crystallized from water solutions, although perfectly dry, yield large quantities of water when heated. When the water is driven off, the crystal structure of the solid changes. These salt hydrates conform to the law of definite composition: the number of moles of water driven off per mole of anhydrous salt is some simple number. Apparently the water was present as an integral part of the crystal structure. Analyze a hydrate and calculate its formula.

EXPERIMENTAL

APPARATUS AND MATERIALS

Crucible and cover, balance, clay triangle, Bunsen burner, desiccator, barium chloride hydrate.

PROCEDURE

Read over experiment. Decide what data will be collected, and prepare a table.

Weigh a dry crucible and cover to the limits of the balance available. Place about 2 g. of barium chloride hydrate in crucible and weigh. place the covered crucible on the clay triangle, and warm it slowly with a low temperature Bunsen flame until most of the water has been driven from the crystal. Increase the temperature gradually until the bottom of the crucible is red hot, and heat for ten minutes more. Allow the crucible to cool in a desiccator and weigh at once. Heat again to redness for ten minutes. Cool in the same manner as before, and weigh. Repeat until your weighings agree (to within 0.5 mg. if weighed to 0.1 mg.).

From the weights calculate the number of moles of water that were combined with one mole of barium chloride.

(Optional)

Repeat above experiment using a hydrate of unknown composition. Report the percentage of water in the hydrate. The instructor will give the formula of the anhydrous salt. Calculate the formula of the hydrate.[14]

QUESTIONS

1. Why do you keep the crucible covered?

2. If after ignition the crucible is allowed to stand exposed to the air, it recovers its former weight. Why?

3. Do results agree exactly with those for a definite hydrate? If not, explain any deviation. Compare results with others to see if the deviations are always in the same direction.

There are some compounds that will crumble and turn to powder when exposed to the air while other compounds will collect water and appear to dissolve themselves.

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[14] J. A. Campbell and L. E. Steiner, Laboratory Experiments in General Chemistry (New York, 1955), p. 27.

7. Making a Solubility Curve

DISCUSSION

This experiment is designed to show the relationship between temperature and the solubility of a salt in water. A number of solutions containing the same solute saturated at a given temperature will always contain the same number of grams of solute per 100 grams of water. By weighing a solution saturated at a given temperature and then evaporating to dryness and weighing again the grams of solute per 100 grams of water can be calculated. Around the class the same experiment is being carried out but at different temperatures. A graph can be made of the class results. Such a graph is a solubility curve.

EXPERIMENTAL

APPARATUS AND MATERIALS

Ringstand and ring, Bunsen burner, beaker, evaporating dish, wire screen, tongs, balance, solution of potassium dichromate, thermometer.

PROCEDURE

Prepare a hot water bath for use in evaporating water from the solution. Start heating water so no time is lost in evaporating the water from solution.

From a number of solutions, containing potassium dichromate at different temperatures, select one and add about 20 ml. to a previously weighed evaporating dish. Record temperature of solution. Then evaporate the solution to dryness over a beaker of boiling water. Finally remove the last traces of water by placing the dish on a wire screen over a small Bunsen flame. Be careful to avoid excess heating which may cause splattering or even decomposition of the salt. Cool and weigh dish which now contains the dry salt. Decide what data will be needed to determine the weight of solute per 100 grams of water. Devise a table which will show a record of the tabulations. From the class results using values for temperature on the horizontal axis and solubility in gm. per 100 gm. of water on the vertical axis plot the curve.[15]

QUESTIONS

1. Can it be assumed that the same sort of graph is typical of the solubility of all salts in water?

2. If the graph shows a straight line portion, what is the meaning of this portion?

3. If the salt used in this experiment should be a hydrate, what difference in method if any, would be used?

Various salts can be obtained in relatively pure form from sea water. The Dead Sea is an example of this.

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[15] "Scientific Experiments in Chemistry," Student guide HS-S78, Manufacturing Chemists Association, Inc. (Washington, D. C., 1957).

8. How Many Atoms in a Molecule of Hydrogen?

DISCUSSION

This experiment is designed to familiarize the student with the relationship between electrons supplied in a circuit and hydrogen ions discharged. The hydrogen is generated by the reaction $H^+ + e = H$, the reduction of hydrogen ions. A sample of hydrogen is collected and the number of hydrogen atoms in the sample is calculated.

The student should be familiar with a few physical relationships before tackling this experiment.

Experimentally, Faraday found that for all elements, the charge necessary to liberate a mass (in grams), equal numerically to the chemical equivalent, is the same for all substances. "This charge is known as a Faraday. It is the quantity of electricity that will liberate a gram-atom of any substance that has a valence of 1."[16] Hence, $Q = 96{,}500$ Coulombs. From the equation, $Q = Ne$, where (N) is the number of electrons and (e) is the charge on the electron, we can find the total charge (Q). From the equation $Q = It$, the charge added to the system can be determined. A milliammeter and a timing device which records in seconds is all that is required.

In the conclusion, determine the number of hydrogen atoms contained in the gas sample. Assume that hydrogen is monatomic. Determine the volume that should be occupied by the number of hydrogen atoms of your sample. Compare actual volume with the calculated volume. On

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[16] Frank M. Durbin, Introduction to Physics, (Englewood Cliffs, New Jersey, 1955) p. 365.

this basis, draw a conclusion as to the number of hydrogen atoms per molecule.

APPARATUS AND MATERIALS

Two small platinum electrodes connected to copper wires encased in waterproof, acidproof insulation; two 50 ml. burettes, battery jar, source of direct current, milliammeter, copper connecting wire, rubber tubing, hose clamps, timing device for recording time in seconds, thermometer, barometer, water, and sulfuric acid.

PROCEDURE

Fill the battery jar about half full of water. Fill the burettes with dilute acid above the graduated portion. Close the valves or clamps on the burettes. Connect the milliammeter, the two electrodes, and the source of direct current in series with a switch. Slowly add sulfuric acid with considerable stirring. Close the switch and watch carefully for evidence of current flow. If source of current can be varied, adjust the flow so that a convenient number of milliamperes registers on the ammeter. Let the current flow for about two minutes. Notice the reading of the milliammeter from time to time. A constant flow of current is desired. The solution is now saturated with gas. Let the gas rise out of both tubes and then record the level of the liquid in each tube. Start the flow of current again. Note carefully the time that the current starts to flow. Be prepared to measure time. Let it continue to flow until between 40 and 50 ml. of hydrogen has been collected, or there is sufficient clearance to level the liquids inside the burette with the water level in the battery jar. Turn off the current and note the time. After all bubbles have gone out of the liquid, record the level in the tube containing the hydrogen. Also record the time interval. Read the temperature and pressure and change the volume of hydrogen to S. T. P. Take account of pressure, temperature, water-vapor pressure, and difference in level of liquids (if any).[17]

QUESTIONS

1. Show by equations the reactions at the anode and the cathode.
2. What is meant by oxidation and reduction?
3. What changes would have to be made to determine the atoms in a molecule of oxygen?

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[17] "Scientific Experiments in Chemistry," Student guide HS-S 834, Manufacturing Chemists Association, Inc. (Washington, D. C., 1958).

9. Composition and Simplest Formula of Tin Oxide

DISCUSSION

Equations for chemical reactions and formulas for compounds can be used in making important calculations. For example, an equation can be used to calculate the proportions by weight of reactants and products in a chemical reaction, and the formula can be used to calculate the percentage composition by weight of a pure compound.

In the following experiment calculate the percentage composition of the tin oxide from your results and from the percentage composition determine the simplest formula of the tin oxide.

The following is a list of four rules for solving problems in which the simplest formula is to be found.

1. Divide the percentage by weight of each element by its atomic weight.
2. Find the ratio of these quotients by dividing each one by the smallest quotient.
3. If the ratios are fractions, reduce them to the smallest whole numbers possible.
4. Write the symbol of each element with that subscript number following it which corresponds to the ratio found.[18]

EXPERIMENTAL

APPARATUS AND MATERIALS

Ring stand with ring, crucible, balance, tin foil, concentrated nitric acid, Bunsen burner.

PROCEDURE

Weigh the crucible on the balance. Add about one gram of tin foil and weigh the crucible and foil. Record all weights in your

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[18] Charles E. Dull, William O. Brooks and H. Clark Metcalf, Modern Chemistry, (New York, 1953) p. 133.

results.

Place the crucible on the wire triangle and place near an open vent or hood. Add 5 ml. of concentrated nitric acid very slowly. Pour the acid over every part of the metal.

Hold the Bunsen burner in one hand and warm the crucible with a low flame. Stop heating if evidences of splattering take place. When the contents are dry, use a larger flame and heat for ten minutes. Do not put crucible in desiccator at any time. When the contents are cold weigh the crucible and determine the weight of the tin oxide. Record all results and complete the requirements of the experiment.

QUESTIONS

1. Write the equation for the reaction that took place.
2. What was the brown gas liberated in the experiment?
3. Do the results of this experiment show the true formula of tin oxide? Explain.

10. Gram Molecular Volume

DISCUSSION

A gram molecular weight (mole) of any gas occupies a volume of 22.4 l. at standard conditions of temperature and pressure. Thus 2 g. of hydrogen, 32 g. of oxygen, or 28 g. of nitrogen occupy 22.4 l. at standard conditions. By determining the weight of 22.4 l. of a substance in the form of a gas at standard conditions, we find directly the molecular weight. Of course, we might determine the weight of a much smaller volume of the gas and calculate the weight of 22.4 l. This volume, 22.4 l., is termed the gram molecular volume or molar volume of a gas.

Usually it is not convenient to measure a gas at standard conditions, 760 mm. pressure and 0° C. temperature. However, the volume of gas under any other set of conditions may be measured, and the volume then calculated for standard conditions with the aid of the gas laws.

In this experiment, we shall measure the volume of a known weight of oxygen, and calculate the volume at standard conditions of 32 g. of oxygen. This volume, barring experimental error, should be the gram molecular volume, namely 22.4 l.

EXPERIMENTAL

APPARATUS AND MATERIALS

Balance, evaporating dish, ring stand, test tube holder, test tube, 500 ml. flask, pneumatic trough, thermometer, potassium chlorate, manganese dioxide.

PROCEDURE

Set up a test tube oxygen generator. Clean and dry the test tube thoroughly, add to it approximately 0.5 g. of manganese dioxide, which has been heated moderately in an evaporating dish to remove moisture. Weigh accurately and record the weight. Place not more than 1-2 g. of potassium chlorate in the tube and accurately weigh again. Mix intimately the potassium chlorate and manganese dioxide. Clamp the tube in a nearly horizontal position and attach to the delivery tube. Leave a channel above the solid in the tube.

Fill an approximately 500 ml. flask with water and invert it in a pan or pneumatic trough. Adjust the flask so that the evolved oxygen may be collected by water displacement. Heat the test tube generator slowly. Begin heating above the solid and gradually heat downward toward the end of the tube. Collect all the gas which is driven over (including the air in the test tube and delivery tube). Remove the delivery tube from the water immediately upon cessation of heating.

Cool the tube and weigh. Record the weight. Describe to your instructor how you may determine the volume of gas in the partially filled flask. If he approved your method, proceed with the measurement. Record the temperature of the water in the trough, as well as the barometric pressure.

Record all data in a systematic form taking into consideration vapor pressure of water, barometric pressure, volume of gas collected, temperature of water, weight of oxygen liberated. Determine from the results the volume of oxygen at S. T. P., the weight of 1 liter of oxygen and the weight of 22.4 l. of oxygen.[19]

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[19] King and Caldwell, p. 41.

QUESTIONS

1. Calculate the weight of 100 liters of carbon dioxide gas at S. T. P.

2. Calculate the volume of oxygen obtained at S. T. P. by heating 245 g. of $KClO_3$.

3. A steel cylinder contains 20 lbs. of liquid carbon dioxide. When the cylinder is opened, the carbon dioxide gas assumes S. T. P. conditions. How many cubic feet volume would the gas then occupy?

11. The Equivalent and Atomic Weight of a Metallic Element

DISCUSSION

In a quantitative analysis the chemist measures the relative weights of elements which have combined with one another or replaced one another in compounds. If the chemist knows the equivalent weight of one of the elements, he can then calculate accurate values for the equivalent weights of other elements. If he also knows the number of equivalents in an atomic weight, that is, the oxidation number, he can then get an accurate value for the atomic weight of the element.

The amount of hydrogen gas produced when an acid reacts with a known weight of an oxidizable metal will provide us with a method of replacement. From the volume of gas liberated and the weight of the metal the hydrogen can be found. From the oxidation number and the value of the equivalent weight an accurate value of the atomic weight can be found.

EXPERIMENTAL

APPARATUS AND MATERIALS

Gas measuring tube, barometer, large battery jar, thermometer, magnesium ribbon (ribbon 5 cm. long gives a sufficient amount of hydrogen), copper wire, balance.

PROCEDURE

Fill a battery jar 3/4 full of water. Cut off 5 cm. of magnesium ribbon; weigh, and roll into a coil. Tie a piece of copper wire around the center of coil. Pour about 5 ml. of concentrated hydrochloric acid into the gas measuring tube. Then slowly fill the tube with water, being careful to avoid mixing the water and acid. Suspend the coil of magnesium in the tube about 5 cm. beneath the surface. Quickly place your thumb over the end of the tube, and invert it into the battery jar. Support the tube in place and fasten the copper wire so that the magnesium ribbon will be held 5-6 cm. inside the tube. The heavy acid will diffuse down the tube and react with the metal, releasing hydrogen.

When the action has ceased, all the metal having reacted, adjust the tube so that the water levels are the same inside and outside the tube. Read the volume of the hydrogen obtained from the reaction. Take the temperature of the water and consider it to be the temperature of the gas collected. Read the barometer and consult a water vapor pressure table for the pressure. Find the pressure due to the dry gas alone by subtracting the pressure due to aqueous tension from the barometer reading. Making further corrections for pressure and temperature, find the volume of the hydrogen under standard conditions. From this standardized volume determine the weight of hydrogen generated. Using the known weight of magnesium and the weight of hydrogen, calculate the weight of magnesium which would replace 1.008 grams of hydrogen. From this information determine the valence of magnesium.[20]

QUESTIONS

1. How many gram atoms of hydrogen will be liberated by treatment of one gram atom of each of the following metals with excess acid:

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[20] Theodore E. Eckert, Harley K. Lyons, Wallace H. Strevell, Discovery Problems in Chemistry (New York, 1957) p. 293.

zinc, aluminum, magnesium, sodium, calcium?

2. 0.5 g. of a metal, when treated with excess hydrochloric acid, liberates 784 ml. of hydrogen gas measured at 20° C. and 700 mm. pressure. Determine the equivalent weight of the metal.

3. If in the calculations one uses 0.09 g. instead of 0.0899 grams as the weight of a liter of hydrogen, what percentage error will be introduced?

The experimental equivalent weight is not equal to the theoretical value.

12. Conductivity of Solutions

DISCUSSION

Certain classes of substances, when dissolved in water, cause changes in the boiling and freezing points of water which yield anomalous values for molecular weight in that the calculated molecular weight is smaller than that corresponding to the simplest formula. These so-called "abnormal" solutes also show high electrical conductivity when dissolved in water. In this experiment conductivity will be the only factor considered.

In determining conductivity several factors must be kept in mind. The electrodes must come in complete contact with the substance tested. The tumbler containing the substance should always contain the same volume and electrodes immersed the same distance. Rinse electrodes with clear water after each test.

EXPERIMENTAL

APPARATUS AND MATERIALS

Conductivity apparatus, dry sodium chloride, distilled water, tap water, sodium chloride 0.1 N, copper sulfate 0.1 N, ammonium hydroxide 0.1 N, calcium hydroxide saturated, potassium hydroxide 0.1 N, sodium hydroxide 0.1 N, acetic acid (glacial), acetic acid 1 N, hydrochloric acid 0.1 N, nitric acid 0.1 N, sulfuric acid 0.1 N, sucrose, ethyl alcohol, carbon tetrachloride.

PROCEDURE

The conductivity apparatus is wired so that the electrodes are in series with the lamp filament and the voltage source. Test the conductivity of each substance listed in the materials. If the substance is a good conductor the lamp will glow brightly. With a fair conductor the lamp will glow dimly, and with poor or non-conductors, the lamp will not glow at all. Tabulate all results and separate compounds into acids, bases and salts.[21]

QUESTIONS

1. For those solutions showing ionization write the ionic equation.
2. What general classes of compounds are good electrolytes?

The conductivity of ortho phosphoric acid is not three times that of hydrochloric acid even though it contains three times as many hydrogen atoms.

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[21] Ibid., p. 99.

13. Neutralization by Titration

DISCUSSION

The process of determining the concentration of unknown solutions by reaction with a solution of known composition is termed titration. In this experiment an unknown sample of acid will be determined from a known sample of base. To determine when the base has exactly neutralized the acid, an indicator solution which changes color when equivalent amounts of acid and base are present may be used. The color change is termed the end point of the titration. Different indicators change color at varying pH values; for example, phenolphthalein changes from colorless to pink at a pH of about 9, while methyl red or methyl orange changes from red to yellow at a pH of about 5. The correct choice of an indicator is determined by the pH value of the solution when equivalent quantities of acid and base are present. In this experiment phenolphthalein will be used as the indicator.

The concentration of a titrating solution is usually expressed in terms of normality, which is defined as the number of equivalents of the acid or base per liter of solution. It has been pointed out in a previous experiment that an equivalent of an acid is that weight of the acid in grams which contains 1.008 g. of replaceable hydrogen, or an equivalent of a base; that weight of the base in grams which reacts with 1.008 g. of hydrogen ion.

The following relationships are valuable in understanding the theory of titrations and in making the necessary calculations involved in this quantitative analysis procedure.

By definition, the following two relationships exist:

1. $\text{Normality} = \frac{\text{Number of equivalent weights}}{\text{Number of liters of solution}}$

2. $\text{Number of equivalent weights} = \frac{\text{Grams}}{\text{Equivalent weight}}$

In titration, when the neutralization point is reached, the same number of equivalent weights of acid and base have been added to the solution.

3. Number of equivalent weights of acid = number of equivalent weights of base. By rearranging equation (1) we obtain equation (4).

4. Normality X number of liters of wolution = number of equivalent weights. Substituting this equation into (3), equation (5) becomes:

5. Normality of acid X liters of acid = normality of base X liters of base.

In this experiment a basic solution will be made up, the concentration of which will be known. The student will then determine the composition of an unknown acid solution.[22]

EXPERIMENTAL

APPARATUS AND MATERIALS

2 burettes, ring stand, 2 burette clamps, 2 Erlenmeyer flasks, 2 stirring rods, dilute hydrochloric acid, 1 N solution of sodium hydroxide, phenolphthalein solution.

PROCEDURE

Mount two clean burettes in a vertical position on a ring stand. In one burette add a dilute solution of hydrochloric acid (1 part concentrated hydrochloric to 10 parts water). Fill the other burette with

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[22] King and Caldwell, p. 141.

a 1 N solution of sodium hydroxide. This normal solution must contain exactly one molecular weight in grams of sodium hydroxide in one liter of solution.

It is not necessary for the burettes to be filled to the zero point if an accurate record is made of the initial reading. The readings should be taken at the bottom of the meniscus for uniformity in all results.

Allow about 10 ml. of the acid solution to drop into a small, clean Erlenmeyer flask. Place a sheet of white paper under the flask containing the acid solution. Add two drops of phenolphthalein solution. Slowly drop sodium hydroxide solution from the other burette into this flask, rotating or stirring the flask constantly. When a pink color tends to persist as the base comes in contact with the indicator, add the basic solution slowly, drop by drop. When a single drop of the basic solution causes the entire liquid in the flask to change and remain a faint pink, the neutralization is complete. This condition is known as the end point. If by accident too much base is added, return to the burette filled with acid and add a few drops of acid to bring the solution back to its colorless condition and then return to the base. When you are satisfied with the results take the readings of both burettes and record.

The titration should be repeated several times in order to obtain valid results.

Set up a data sheet with the essential results tabulated and determine the weight of acid in 1 ml. of solution, the weight of acid in 1 liter of solution.[23]

QUESTIONS

1. Why is it not necessary to fill the burettes full to begin the experiment?

2. Explain how one might run subsequent trials and alleviate the chance of error in color difference between solutions.

3. If an error is made in titrating too much of one solution, how can this situation be corrected?

When a salt such as Na_2CO_3 is placed in a beaker of water the solution becomes alkaline.

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[23] Ames and Jaffe, p. 93.

14. Acids, Bases, Salts

DISCUSSION

Acids, bases and salts constitute by far the largest group of inorganic compounds. Acids are substances which yield protons in solution and are defined as proton donors. A base is a substance which accepts protons. In water solution, the principal acceptor of protons is the hydroxyl ion. Since hydroxides yield this ion in aqueous solution, metallic hydroxides such as $NaOH$ and $Ca(OH)_2$ are the more common basic substances. Salts are compounds in which a metal ion is combined with a nonmetal ion or radical.

EXPERIMENTAL

APPARATUS AND MATERIALS

Test tubes, evaporating dish, Bunsen burner, medicine dropper, splinters, dilute hydrochloric acid, phenolphthalein, methyl orange, zinc, calcium carbonate, sodium sulfite, calcium oxide, sodium hydroxide, lead oxide, litmus paper, ammonium hydroxide, sodium sulfate, potassium hydroxide.

PROCEDURE

1. Acids: Test a very dilute solution of hydrochloric acid with two drops of the indicator, methyl orange. Test a second portion with two drops of phenolphthalein. Record the results.

Recall the generalization developed in Experiment 3 in which the products of combustion of sulfur, phosphorus, and carbon with water formed acids.

Review Experiment 4 and tabulate the results in this experiment as the action of acids on metals.

To a pinch of sodium carbonate in a test tube, add a few drops of dilute hydrochloric acid. Repeat the above, using sodium sulfite and hydrochloric acid. Observe the odor of the gases evolved and note chemical reaction.

From the above list of results summarize acids in conclusion of this experiment.

2. Bases: In separate test tubes, place a portion (about the size of a pea) of calcium oxide, sodium hydroxide and lead oxide. Add 5 ml. distilled water, shake well, then test portions of the three solutions with litmus paper, methyl orange indicator, and phenolphthalein. Record colors of the indicators in results.

Similarly test 2 ml. of dilute ammonium hydroxide. Try the feeling of the solution of a base. Cautiously taste a dilute solution of a base. Record the different properties of bases as a systematic table.

3. Salts: Place a small piece of zinc in a dilute solution of hydrochloric acid. When reaction stops evaporate the solution and note the residue. Test the escaping gas with a lighted splinter.

Place 5 ml. of sodium hydroxide solution (5 per cent) in an evaporating dish. Dilute 5 ml. of dilute hydrochloric acid solution with 20 ml. of water, and stir thoroughly. To the sodium hydroxide add 2 drops of methyl orange solution. Add the hydrochloric acid solution until an end point is reached. Cautiously evaporate the solution to dryness over a low flame and taste the residue.

From step 1 involving acids, note the reaction of an acid with sodium sulfite, a salt.

To 5 ml. of sodium sulfate solution, add 2 ml. of potassium hydroxide solution. Note any chemical reaction.

From the experiments conducted on acids, bases and salts make a data sheet with all the results observed. Write out a conclusion for each of the three types of compounds.[24]

QUESTIONS

1. How might one distinguish between acids, bases and salts.
2. Will all metals react with hydrochloric acid and produce a salt? Explain.
3. Devise a method for identifying the salts of carbonate and sulfites.

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[24] King and Caldwell, pp. 79-80.

15. Oxidation and Reduction

DISCUSSION

Combination, decomposition, and various other chemical changes involve at least two elements changing in valence or oxidation number.

Reactions in which a transfer of electrons take place are termed oxidation-reduction. Oxidation is defined as a loss of electrons or a gain in valence and reduction as a gain of electrons or a loss in positive valence. These are mutually dependent processes and both occur simultaneously. The substance in the reaction which gives up electrons is referred to as a reducing agent, and the substance which takes up these electrons is referred to as the oxidizing agent.

The ease with which electrons are given up or accepted by a given substance is a measure of its reducing or oxidizing ability.

In the following experiment a metal is oxidized to form an ion, and then the ion is reduced to form the original metal.

EXPERIMENTAL

APPARATUS AND MATERIALS

Medicine dropper, evaporating dish, Bunsen burner, test tube, beaker, one-hole stopper, delivery tube, watch glass, copper, dilute nitric acid, concentrated nitric acid, powdered charcoal, ammonium hydroxide, lime water (calcium hydroxide solution).

PROCEDURE

Place a small strip of copper which weighs about 3 g. in an evaporating dish. Add dilute nitric acid until the metal is about one-half covered. Add a few drops of concentrated nitric acid to hasten the start of the chemical reaction. Note the color of the gas bubbles in the solution and also after they break above the solution. Avoid inhaling the gas.

When the reaction is complete remove any excess metal from dish and wash with water. Heat the solution until about one-half of the liquid is evaporated. Control the heat so no splattering will take place.

Remove a small sample of the warm solution and place it on a watch glass. When solution has cooled and crystals form note their shape and color. In the conclusion of this experiment tell how this might be a way of identifying substance.

Withdraw a small sample of solution from the evaporating dish by means of a medicine dropper. Place this in about 5 ml. of water in a test tube. Add a little ammonium hydroxide and shake the mixture. Then continue adding ammonium hydroxide and shaking mixture until a second change occurs and the solution is clear and blue. This result is typical for one ion only. Consult reference material for complete reaction.

Heat the crystals and liquid which remain in evaporating dish. Notice the change in color of the residue after it apparently was dry. Notice the color of the gas escaping and try to identify it.

The dark solid residue in dish is cupric oxide. Mix this solid with about an equal volume of powdered charcoal. Transfer this mixture to a hard glass test tube containing a one-hole stopper and delivery tube. Allow the delivery tube to extend below the surface of a solution of calcium hydroxide in another test tube. Slant the test tube with the mixture slightly down and begin heating gently at first and then more strongly. When a change occurs in the calcium hydroxide remove the stopper from the mixture and heat until a glow appears throughout the mixture. Allow the test tube to cool with its open end slightly downward. After it has cooled, dump the contents into a beaker half full of water. Place the beaker under a slow stream of running water, and stir until all unused carbon is washed away. Identify the substance in the bottom of the beaker.[25]

QUESTIONS

1. Show how the metal was oxidized, and the resulting compound of copper was reduced.
2. Tell how oxidation and reduction were accomplished in this experiment in terms of electron transfer.
3. Account for the fact that the weight of the final metal is less than the original metal.
4. Write equations for the entire series of reactions.

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[25] "Scientific Experiments in Chemistry," Student guide HS-S 7-17, Manufacturing Chemists' Association, Inc., (Washington, D. C., 1958).

16. Preparation of Iodoform

DISCUSSION

The purpose of this experiment is to give the student an understanding of organic reactions. One of the marked differences between organic and inorganic reactions is the percentage yield. If no side reactions occur and experimental technique is flawless the entire process would be 100%. In inorganic reactions the results come nearer to this reality than organic reactions.

In this experiment the student will prepare iodoform from which he will determine the theoretical and actual yield.

EXPERIMENTAL

APPARATUS AND MATERIALS

500 ml. flask, filter paper, funnel, desiccator, microscope, test tube, watch glass, small bottle with lid, acetone, 8 N sodium hydroxide, ethyl alcohol.

PROCEDURE

Place 10 g. of iodine in a 500 ml. flask, then all 10 g. of acetone. Make a solution of 20 ml. of 8 N sodium hydroxide solution and 80 ml. of water. Add this in small portions to the iodine, meanwhile shaking the flask continually. Keep the flask cool by periodically holding it in a stream of cold water.

When all brown color has been removed from the liquid and no scraps of unattacked iodine are present on the bottom of the flask, reject the balance of the sodium hydroxide solution and set the preparation aside.

Prepare the funnel for filtration. After the preparation has stood for about 5 minutes, decant most of the liquid from the yellow precipitate and remove the rest by the use of the funnel. Place the liquid at once in the bottle labeled "Iodoform filtrate."

Wash the precipitate on the funnel with a little water, then withdraw all water possible. Place the precipitate on a dry filter paper on a watch glass in the desiccator. The desiccator should be charged with calcium chloride.

The filter paper used in the funnel will retain some iodoform which it was impossible to remove. Crush the paper into a ball, drop it into a test tube, and add about 5 ml. of ethyl alcohol. Warm this to the boiling point, then shake it well. Agitate the filter paper and press it with a stirring rod. Repeat the heating and stirring procedure 10-15 times, then filter the solution into a clean test tube; cork this and let it stand.

If no crystals appear upon cooling, heat the solution and add water drop by drop until a very slight cloudiness is seen in the hot solution, then again insert the cork and let the tube stand. Finally chill it in cold water. Insert a pipet, draw up a few drops of solution (with suspended crystals), and place this on a microscope slide. Observe the crystals with a low-power objective. Place a sketch of the crystals in your report.

The iodoform should be dry at the next laboratory period. Weigh the sample and hand it in in one of the small sample bottles. Note on the label: desk number, experiment number, your name, weight of the product, and percentage yield.[26]

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[26] E. Wertheim, Experiments in Organic Chemistry, (Philadelphia, 1950), pp. 63-65.

QUESTIONS

1. Write a complete set of equations for the production of iodoform from iodine, acetone, and alkali.

2. What reason can you assign for the use of acetone in this experiment in excess of the requirements?

3. What becomes of the sodium acetate produced in the reaction? How might it be recovered? Would this recovery be economical? Why?

4. Why do we save the filtrate from the reaction mixture? In what form are the iodine atoms in this filtrate?

5. Describe a simple test which would prove that iodoform contains halogen.

6. What chemicals would be used to prepare bromoform by the method of this experiment?

7. Describe a method by which the crude iodoform could have been purified.

17. Preparation of Colloids

DISCUSSION

A colloid is a suspension of infinitely small clumps of molecules of an insoluble material in a suspending medium. The suspended clumps are intermediate in size between molecules and precipitates. Since they are clumps of molecules they are larger than molecules. They are smaller than precipitates, and, unlike precipitates, they remain in suspension.

Colloidal suspensions can be prepared by two general methods. In the first, called the condensation method, molecules are allowed to agglomerate until they become clumps of colloidal size. In the second, called the dispersal method, precipitates or masses of material are reduced to colloidal size.

It is the object of this experiment to show how representative colloids can be prepared.

EXPERIMENTAL

APPARATUS AND MATERIALS

Bunsen burner, 250 ml. beaker, medicine dropper, test tube, mortar and pestle, slide projector, 0.2 M ferric chloride, 0.02 M arsenic oxide, saturated hydrogen sulfide solution, kerosene, soap solution, carbon black, gelatin, 1% gelatin solution, potassium chromate, ferric thiocyanate.

PROCEDURE

1. Condensation Methods

a. In preparing colloidal ferric hydroxide heat 100 ml. of water and, while maintaining the boiling add 0.2 M ferric chloride dropwise at the rate of about 1 drop per second until a deep wine red suspension of ferric hydroxide is obtained. The ferric hydroxide molecules agglomerate to form clumps of molecules. These clumps absorb ferric ions from the solution and as a result become stabilized. Save the colloid for part 3.

b. In the preparation of colloidal arsenic sulfide mix 25 ml. of a solution of arsenic oxide in water with 10 ml. of hydrogen sulfide solution. The yellow color is due to the presence of colloidal arsenic sulfide. The arsenic sulfide molecules agglomerate to form clumps which are stabilized by adsorption of sulfide ions. Save the colloid for part 3.

2. Dispersion Methods

a. Place about 10 ml. of water in a test tube, add 10 drops of soap, and one ml. of kerosene and shake. This produces an emulsion which is a form of colloid.

The mass of kerosene has been broken down into fine particles by the shaking process. The soap lowers the surface tension of the water and also is absorbed by the tiny oil particles.

b. In preparing a suspension of carbon black in a mortar, add 2 ml. of water and 10 drops of soap solution, and grind to an intimate mixture. Add 50 ml. of water, mix thoroughly by grinding and transfer to a small beaker. A colloidal suspension of carbon in water has been formed.

3. The Tyndall Effect

Place 5 ml. of the colloidal suspensions of ferric hydroxide and arsenic sulfide in two test tubes. Place in two other test tubes a solution of 0.2 M potassium chromate and a solution of ferric thiocyanate. Hold all four tubes in a beam of light from a projecting lantern or spotlight. Record observations. Account for the difference in the appearance of the two colloids as compared with the two true solutions.

Allow the beam from the lantern to fall upon the smoke produced by bringing near each other pieces of filter paper soaked with hydrochloric acid and ammonium hydroxide. Record and account for what was observed.[27]

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[27] C. H. Sorum, A Laboratory Manual of General Chemistry, (Englewood Cliffs, N. J., 1955), pp. 206-207.

18. Separation of Cations in Qualitative Analysis

DISCUSSION

The purpose of this experiment is to acquaint the student with the method of separating the cations.

Cation analysis is more systematic than is anion analysis. Consequently, the first step will be to see into what analytical groups the cations may be divided on the basis of similarities in reaction with various reagents. The ions tested are antimonous, cobaltous, ferric, lead, cupric, magnesium, mercurous, mercuric, strontium, sodium, and potassium.

Be sure to record formulas, colors, and appearances of all precipitates. Describe precipitates as "crystalline," "powdery," "curdy," "flocculent," or "gelatinous." Also, when colored substances appear, note the hue and intensity of the color.

Fill in the data as experiment proceeds. Write equations for reactions that occur. Where ions are involved write ionic equations.

EXPERIMENTAL

APPARATUS AND MATERIALS

Test tube rack, ten test tubes, hydrogen sulfide generating tube, Bunsen burner, 3 N nitric acid, 3 N hydrochloric acid, ammonium nitrate (solid), ammonium hydroxide, ammonium sulfide, 0.7 N ammonium hydroxide, 1 N ammonium carbonate, ions of antimonous, cobaltous, ferric, lead, magnesium, mercurous, mercuric, strontium, sodium, potassium, cupric.

PROCEDURE

To eight test tubes filled with 1 ml. of water, add one drop of a test solution to each test tube.

Group I. Note carefully any solutions in which a precipitate appeared when the drop of test solution was diluted with water. Such salts hydrolyze in water to form basic salts. (Reminder: Describe the appearance and color of the precipitates.) Now add 3 N nitric acid, drop by drop with shaking and warming, until any basic salts redissolve.

To each of the tubes add one drop of 3 N hydrochloric acid and shake. Allow the precipitates to settle. If no precipitate has appeared in the lead ion solution, add one more drop of the lead ion solution and shake. To each of the tubes add one more drop of 3 N hydrochloric acid. Does more precipitate appear? Add a third drop of acid to each tube. The ions that form insoluble chlorides are called "Group I" in the qualitative analysis. Empty and wash the tubes containing precipitates. (Mercury and its compounds should always by dumped into crocks, never the sink, as mercury amalgamates with the lead joints.) Replace the water and test solution, and set them as "Group I" at the head of the set of tubes. Save all the tubes and their contents for Group II.

What evidence do you have of the relative solubilities of the chlorides of the Group I cations?

Group II. Add three drops of 3 N nitric acid to the tubes containing the Group I cations. (Add three drops of hydrochloric acid to the other tubes in Group I.) Draw down some glass tubing to make a long tip, attach with rubber tubing to the hydrogen sulfide gas line in the hood, and pass hydrogen sulfide gas into each of the tubes, washing off the bubbling tip during each transferral from one tube to the next. The white precipitate in the test tube which contained ferric ion is sulfur. Write the equation for the formation of this precipitate. Which of the cations form insoluble sulfides in acid solution? What is the acidity of the solutions? The tubes with precipitates, except for those containing Fe^{+++} or the Group I cations, contain Group II cations; set them as "Group II" next to the Group I tubes. Empty and rinse all the tubes.

Which of the Group II salts hydrolyzed when diluted with water? Did any salts not in this group hydrolyze? (See also Table for Group I). The hydrolysis reaction in an unknown is an indication of the presence of one or more of these ions.

Group III. Replace the water and test solutions in all the tubes (adding HNO_3 as in Group I if necessary to clarify any ppt.) and to each add 0.3 g. of solid ammonium nitrate. (Compare with 0.3 g. sample of NH_4NO_3 on exhibition.) Now add 3 N ammonia until the solutions are slightly alkaline. Which of the cations form insoluble hydroxides in the presence of an excess of ammonium ions? Set the tubes with precipitates, except those in Groups I and II, together as "Group III". Refill the tubes containing precipitates with the original solutions and 0.3 g. of ammonium nitrate.

What would have happened if you had used ammonium chloride instead of ammonium nitrate? In the qualitative analysis ammonium chloride is used, but Groups I and II have been removed before Group III is precipitated.

Group IV. To each of the tubes add two or three drops of ammonium sulfide reagent. (Ammonium sulfide should be prepared fresh each time it is needed by bubbling hydrogen sulfide through 3 N ammonia solution for a minute or two.) Test the reagent with litmus. Examine the precipitates in the Group IV tubes. How do they compare with the precipitates in Group III? Which of the tubes contain precipitated sulfides? Set these tubes except those belonging to Groups I, II and III apart as "Group IV." Empty and wash all tubes.

Group V. Refill the tubes with original solution and add 0.3 g. of solid ammonium nitrate to each. Then add 7 N ammonia until the solutions are just alkaline. To the tubes not containing insoluble hydroxides at this point add 1 N ammonium carbonate as long as a precipitation continues. Which of the cations form insoluble carbonates under these conditions? Set the tubes not already in Groups I-IV as "Group V."

Exercise. Which of the cations formed no precipitates in any of the preceding experiments? The cations in these tubes are classed in "Group VI." Since hydrogen ions and ammonium ions are added in the procedures, an unknown solution must be tested for these before the group separations are made. Therefore, these two ions are classed as "Group O."

Summarize your observations in the results, entering each cation in the proper qualitative group and indicating the formula and the color of each precipitate. Write the equations for the reactions.

Locate, on the periodic table, each of the elements whose ions are listed in the results. Mark all members of each Analytical Group with the same characteristic marking. Note the correlation between analytical groups and position in the periodic table.[28]

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[28] J. A. Campbell and L. E. Steiner, <u>Laboratory Experiments in General Chemistry</u>, (New York, 1955), pp. 154-55.

19. Test for Anions in Qualitative Analysis

DISCUSSION

The analysis of anions is similar to the analysis of cations in that the various anions studied are divided into groups, each group containing ions possessing similar chemical properties. However, since all the anions that constitute a particular group will not form insoluble compounds on the addition of a given precipitating agent, each group must be taken separately using portions of the original solution.

The anions to be considered are $CO_3^=$, $S^=$, $SO_3^=$, $SO_4^=$, Cl^-, Br^-, I^-, and NO_3^-, and these are classified into the following groups:

Group I, the Volatile Group, contains carbonate, sulfide, and sulfite ions.

Group II, the Barium Nitrate Group, includes sulfates.

Group III, the Silver Nitrate Group, includes chlorides, bromides and iodides.

Group IV, the Nitrate Group, contains the nitrate ion from which the group derives its name.

Tests for the individual ions are listed in the procedure. From these individual tests the student should devise a method for separating the ions into groups.

Write equations for all reactions that take place.[29]

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[29] Carl Likes and Aubrey Harvey, First Year Qualitative Analysis, (New York, 1947), p. 98.

EXPERIMENTAL

APPARATUS AND MATERIALS

Test tube rack, test tubes, silver nitrate solution, nitric acid, carbon disulfide, chlorine water, hydrochloric acid, lead acetate, paper, barium chloride, potassium permanganate, lime water, concentrated sulfuric acid, freshly prepared ferrous sulfate solution.

PROCEDURE

1. Test for a chloride: To the unknown in solution add some silver nitrate solution. A white precipitate which turns purple in sunlight and is insoluble in nitric acid shows the presence of a chloride.

2. Test for a bromide: To the unknown solution add some chlorine water, 2 ml. of carbon disulfide and shake the solution. An orange-red color in the bottom layer shows that free bromine has been liberated from a bromide.

3. Test for an iodide: To the unknown solution add some chlorine water, 2 ml. of carbon disulfide and shake the solution. A purple color in the bottom layer shows that free iodine has been liberated from an iodide.

4. Test for a sulfide: To some of the solid unknown add a few drops of hydrochloric acid. The evolution of hydrogen sulfide gas shows the presence of a sulfide. This will turn lead acetate paper black.

5. Test for a sulfate: To the unknown solution add a little barium chloride solution. A white precipitate insoluble in hydrochloric acid shows the presence of a sulfate.

6. Test for a sulfite: To the solid unknown add a little

hydrochloric acid. The evolution of sulfur dioxide gas with its suffocating odor is a test for a sulfite. This gas also decolorizes a dilute solution of potassium permanganate.

7. Test for a carbonate: To some of the solid unknown add some hydrochloric acid. The evolution of an odorless gas which when passed into lime water causes it to turn milky shows the presence of a carbonate.

8. Test for a nitrate: To some of the unknown solution add some freshly prepared ferrous sulfate solution. Then carefully pour down the side of the tube some concentrated sulfuric acid. A brown ring in between the two layers indicate the presence of a nitrate.[30]

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[30] Ames and Jaffe, p. 249.

Apparatus and Materials Required

Apparatus:

- asbestos
- balance
- barometer
- beaker, 250 ml.
- beaker, 400 ml.
- bottles, 2 oz. w-lid
- bottles, wide mouth 8 oz.
- Bunsen burner
- burettes, 50 ml.
- clamp, test tube
- clamps, tubing
- clay triangle
- conductivity apparatus
- crucible and cover
- deflagrating spoon
- desiccator
- evaporating dish
- filter paper
- flask, 250 ml.
- flask, 500 ml.
- flask, Erlenmeyer
- funnel
- gas measuring tube
- glass rod
- jar, battery
- litmus, blue
- litmus, red
- medicine dropper
- meter stick
- microscope
- milliammeter
- mortar and pestle
- platinum electrode
- pneumatic trough
- rack, test tube
- ring stand
- ring, 4 in.
- stoppers, rubber
- test tubes
- thermometer, Centigrade
- tongs, utility
- tubing, glass

- tubing, rubber
- watch glass
- watch, stop
- wire screen

Materials:

- acetic acid (glacial)
- acetic acid 0.1 N
- acetone
- alcohol, ethyl
- aluminum
- aluminum oxide
- ammonium carbonate, 1 N
- ammonium hydroxide, 0.1 N
- ammonium nitrate
- ammonium sulfide
- antimonous chloride
- arsenic oxide 0.02 M
- barium chloride
- barium oxide
- calcium chips
- calcium carbonate
- calcium chloride
- calcium hydroxide
- calcium oxide
- carbon (charcoal)
- carbon (lampblack)
- carbon disulfide
- carbon tetrachloride
- chlorine water
- cobaltous chloride
- copper wire
- copper strips
- copper oxide
- cupric sulfate 0.1 M
- cupric sulfate 0.1 N
- ferric chloride 0.2 M
- ferric oxide
- ferric sulfate
- ferric thiocyanate
- ferrous sulfate (fresh)
- gelatin
- hydrochloric acid
- hydrogen sulfide
- iodine
- iron filings
- kerosene
- lead, (granulated)
- lead acetate
- lead acetate paper
- lead monoxide
- lead oxide (red lead)
- magnesium (granulated)
- magnesium (ribbon)
- magnesium nitrate

magnesium oxide
magnetite
matches
manganese dioxide
mercury
mercuric nitrate 0.1 M
mercuric oxide
mercurous nitrate
methyl orange
nitric acid
phenolphthalein
phosphorous
potassium chlorate
potassium chloride
potassium dichromate
potassium hydroxide, 0.1 N
potassium permanganate
soap solution
sodium
sodium chloride (dry)
sodium chloride 0.1 N
sodium hydroxide 0.1 N
sodium hydroxide 8 N
sodium sulfate
sodium sulfite
silver coin
silver nitrate
stannous chloride 0.1 M
strontium nitrate
sucrose
sulfur (rolled)
sulfuric acid
tin, strips
tin, foil
titanium dioxide
zinc (granulated)
zinc oxide

SUMMARY AND CONCLUSIONS

This report is designed to identify the academically talented student, and to recognize this student in the field of chemistry. It also deals with the problem of providing the talented student with experiments which will motivate and inspire him to greater understanding in the chemistry laboratory.

In summary, a talented child is one with high learning ability. He has a greater time-span, his range of learning is greater, and he can keep details significantly related.

The generally accepted method of recognizing the academically talented student in chemistry is with the use of intelligence quotient tests. An intelligence quotient of 120 or more will comprise this group of students. However, there is no single method of identifying the talented child. Additional factors which should be considered are general intelligence, class records, teacher evaluation, interest in science, special abilities and individual characteristics.

It is the belief of the writer that "enrichment" in chemistry comes from the course which is designed to give the student an opportunity to think for himself and is then given occasion to put his thoughts in writing. Using the laboratory write-up as outlined in this report, the student may devise his own method of tabulating results. It is necessary for the student to first evaluate the experiment and to determine what information is actually required and what can be omitted in the results. The conclusion of the write-up is most important. In this section the student can collect all the information he has obtained in the experiment and determine its accuracy and limitations. Requiring the student to write reasons for his errors also gives him an opportunity to review his technique and to discover more efficient methods of carrying out experiments. In giving the practical application of experiments the student can see value in what he is actually doing. Many students become disinterested if they feel that a chemistry experiment has no practical use.

The experiments in this report do not differ in theory with most chemistry laboratory experiments. However, the form of the experiments requires the student to do more individual reasoning and outside reading than does the experiment found in the ordinary high school laboratory manual. Also, in some cases, the experiments will require more than one laboratory period to complete. Before starting an experiment, it is advisable for the student to determine a suitable stopping point which will enable him to continue during the following laboratory period.

The equipment necessary to carry out these experiments can be found in most well-equiped high school chemistry laboratories. If an analytical balance is not available, a good triple beam balance will serve the purpose. In cases where a source of direct current is not available, a storage battery will suffice. In many cases, it is desirable to have students devise and make their own equipment. By building their equipment, they will receive greater understanding and stimulation.

The teacher's role in writing of laboratory reports is very important. Grading of experimental write-ups is tedious and requires a great deal of patience. The experiments should never be undertaken if the reports are not corrected. The experience of the writer is that initial experiments which are well-graded make subsequent experiments easier to grade. Allowing students to read good reports will also make grading of future reports much easier.

It might be said that the sole purpose of this report is to help bright students receive extra stimulation in the field of chemistry experimentation. Most methods of providing this motivation do not require expenditures beyond those made by a first-class school system. It calls for more finesse than finance to give enrichment to talented students. Possibilities for enrichment in chemistry for the academically talented student is restricted only to the individual resourcefulness and energy of the teacher and his colleagues. Motivating students to think rather than to merely memorize is the key concept.

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