 69-12,929

RICHARDSON, Verlin Homer, 1930-
A STUDY OF THE INQUIRY-DISCOVERY METHOD OF INSTRUCTION IN COLLEGE CHEMISTRY LABORATORY.

The University of Oklahoma, Ph.D., 1969
Education, theory and practice

University Microfilms, Inc., Ann Arbor, Michigan

THE UNIVERSITY OF OKLAHOMA

GRADUATE COLLEGE

A STUDY OF THE INQUIRY-DISCOVERY METHOD OF INSTRUCTION IN COLLEGE CHEMISTRY LABORATORY

A DISSERTATION

SUBMITTED TO THE GRADUATE FACULTY

in partial fulfillment of the requirements for the

degree of

DOCTOR OF PHILOSOPHY

BY

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Norman, Oklahoma

1969

A STUDY OF THE INQUIRY-DISCOVERY METHOD

OF INSTRUCTION IN COLLEGE

CHEMISTRY LABORATORY

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ACKNOWLEDGMENTS

I wish to express my appreciation to Dr. John W. Renner for his untiring guidance in the development of the materials used in this study; to Dr. Bernard Heston, Dr. Howard H. Rowley and Dr. Herbert Hengst for their suggestions and comments during the course of the study; to the many students who participated in the study as experimental subjects; to Mr. Marval Evans for his cooperation and encouragement in working out the mechanical details of the study; and finally to my wife, Mona, for helping with the details of putting the dissertation in the proper order.

TABLE OF CONTENTS

LIST OF FIGURES

LIST OF TABLES

A STUDY OF THE INQUIRY-DISCOVERY METHOD OF INSTRUCTION IN COLLEGE CHEMISTRY LABORATORY

CHAPTER I

THE IMPORTANCE OF INQUIRY, REDISCOVERY AND IMAGINATION IN SCIENCE

Purpose

The purpose of this research is to determine if there is a significant difference in the learning experienced by freshmen college students using an inquiry-discovery centered approach to chemistry laboratory experiments as compared to the learning experienced by a similar group of students using commercially available laboratory experiments.

How the Problem Was Found

The investigator has taught college chemistry for ten years and students have very many times remarked that they had experienced disappointment with the laboratory experiments that were available to them. They were most often disappointed because their instructor often assumed too much prior knowledge of the field of chemistry on their part. Students and investigator alike often felt that the questions and exercises were so rigorous as to be frustrating to the student.

Based on the investigator's observations and student reaction over this ten year period, three principal reasons accounted for student dissatisfaction with commercially prepared laboratory experiments. 1. The experimental procedure often assumes too much prior knowledge on the part of the student and thus the student cannot follow all parts of the experiment with understanding. This was obvious from the questions students always ask about certain parts of the experiment. 2. The experimental procedure often requires more time to complete than can be scheduled. Frequently a large fraction of the class does not complete the experiment even though they work diligently from the beginning of the laboratory period. 3. Frequently the student left the laboratory session not actually aware of the purpose of the experiment. This is shown to be true by: (a) Student performance on laboratory examinations. The student, for example, may have had an experiment on equivalent weight and because he doesn't understand equivalent weight he could not correctly respond to questions dealing with that concept. (b) Evidence of a lack of understanding is offered by student comments that are to be found in the appendix of this study. (c) An example of assuming too much prior knowledge on the part of the student is the procedure found in most commercial manuals that instruct the student to correct for the vapor pressure when dealing with gas pressure. If the student hasn't learned how to correct the pressure of a gas for vapor pressure he doesn't know what to do.

Assuming that the previously stated reasons are valid, a series of experiments was written with very few assumptions about the prior knowledge of the student. The topics dealt with in the experiments were those most often encountered in laboratory manuals and care was taken to see that each experiment did not require more than the allotted two hours. The experiments emphasized the theme of the inquiry-approach to learning. All questions and problems which were a part of an experiment were designed so that the student could respond to them without having had a prior course in chemistry. Since the purpose of this research was to determine the effectiveness of applying the inquiry method to chemistry laboratory experiments, considerable care was taken to be sure that the inquiry approach prevailed throughout the experiments.

The advancement of science has reached such a point that it is no longer possible for a student to learn all the separate detailed facts of a particular subject. With this in mind educators should be concerned with the methods of instruction that will most likely result in a better understanding of fundamental principles of the subject being taught, an increase in the student's ability to think, and developing the student's ability to create knowledge of his own.

Philip Phenix explains the need for rethinking what is taught this way.

> If students in school become familiar with basic modes of thought and investigation, they will be better prepared to cope with a changing world than if they possess only a store of facts and information. Furthermore, by having attention directed to methods, they learn attitudes which will prove useful in adapting not only to changing content but even to changing methods. Perhaps the most important reason for selecting the materials of instruction in order to exemplify methods of inquiry is that these methods are also ways of learning. They are the methods that long experience has shown are most productive of new understanding by workers in the disciplines. Methods are adopted as working procedures as a result of their demonstrated instructiveness to investigators. They are the modes of thought that experts have found most efficient in promoting understanding in their disciplines.[1]

The Educational Policies Commission[2] points out quite clearly that the major purposes of education is to develop the rational powers of the mind of the individual because the ability to utilize those powers is the essence of the ability to think. If a student can realize a gain in his ability to think as a result of this education, then his education has been of maximum benefit to him.

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[1]Philip H. Phenix, Realms of Meaning, McGraw-Hill Book Company, New York, New York, 1964, p. 336.

[2]The Central Purpose of American Education, The Educational Policies Commission of the National Education Association, Washington, D. C., 1962, p. 5.

Researchers such as Dewey[3], Bruner[4], and Renner[5] have made a rather strong case favoring the method of discovery (or inquiry) in science education. The method of inquiry does not suggest that the student should be required to discover for himself every principle of a discipline, but rather some of the laws, theories and generalities, and perhaps more importantly, the ability to create knowledge. The method of discovery is particularly applicable to the sciences because science begins with observation and measurements, is based upon definite laws and often advances by inductive reasoning. The inquiry method is also useful for teaching problem-solving techniques since the student can learn to recognize a problem, collect data related to it and form and test certain hypotheses about it.

The original experiments used in this study were written to assist the student in discovering some of the important laws of chemistry for himself. The questions which he answers will assist him in developing an understanding of the laws being demonstrated.

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[3]John Dewey, Experience and Education, The MacMillian Company, New York, New York, 1938, pp. 51-2.

[4]Jerome S. Bruner, The Process of Education, Harvard University Press, Cambridge, Massachusetts, 1963, p. 21.

[5]John W. Renner, The Thrill of Discovery in Science Education, Educators Progress Service, Randolph, Wisconsin, 1962.

John W. Renner and William B. Ragan, Teaching Science in the Elementary School, Harper and Row, New York, New York, 1968, Chapters 1 and 2.

Conventional laboratory experiments too often do not stimulate the student's imagination nor challenge or develop his capacity to think. In understanding, the principle holds that material for instruction should be selected for its power of stimulating imagination and its importance to the discipline. Science is an area which should readily stimulate the student's imagination. Pasteur, Kepler, Faraday and others were examples of those whose success in scientific achievement depended to a great extent on their imaginative powers and their willingness to observe, measure and experiment, that is to investigate.

In sum, the essence of physical science is the discovery and formulation of general patterns among quantities derived from the processes of observation and physical measurement. A successful science teacher is aware of this and tries to bring about this realization on the part of his students.

CHAPTER II

THE NEED FOR THE STUDY

History demonstrates that man is the principal architect of his environment. For the most part as new discoveries have been made, there have been corresponding improvements for the welfare of mankind. Many of the findings which have resulted in a better way of life for man have come about as a direct result of scholarly research. The recent "knowledge explosion" has contributed greatly to ideas about learning, educational purpose, and science.[1]

For a person to know all the facts in any given field is impossible. In view of this it becomes imperative that a person have learning experiences that increase his rational powers so that the information he possesses can be of more use to him. The Educational Policies Commission expresses the foregoing point of view in this manner:

> Development of rational powers is unfortunately an area of relative neglect in research. The emphasis of recent research has been on the conditions under which learning occurs and on the

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[1]Jerome S. Bruner, The Process of Education, Harvard University Press, Cambridge, Massachusetts, 1963, p. 47.

> pathological aspects of learning in specific situations. Considerable research would need to be done before one could, with reasonable assurance, design a program of study that would develop the ability to make valid inferences.[2]

In order for educational experiences to be of value to the student, there is a need to find a method or methods which will result in the student having a better grasp of the methods of science, a greater ability to think for himself, and an increased understanding of the content involved.

According to the Educational Policies Commission[3] the spread of science and technology has lead educators to realize that educational systems and structures must change. Perhaps this evidence of change makes obvious the fact that experimentation is needed in methods employed in education. Possibly any progress made in the future will be made by men that have the spirit of science within them and are thus willing to investigate. According to the Educational Policies Commission,

> The spirit of science causes a person to long to know and understand, to question all things, to search for data and their meaning, to demand verification, to have respect for logic, and to consider the premises and consequences.[4]

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[2]The Central Purpose of American Education, Educational Policies Commission, National Education Association, Washington, D. C., 1961, p. 13.

[3]Education and the Spirit of Science, Educational Policies Commission, National Education Association, Washington, D. C., 1966, p. 15.

[4]*Ibid.*, p. 15.

The following offers evidence that college graduates are not as knowledgeable in the field of science as might be expected. Concerning the widespread lack of knowledge of science among college graduates, Dr. Glen Seaborg recently said,

> Tens of thousands of young men and women are leaving the halls of higher education each year with allegedly liberal educations; but in fact, they have little or no knowledge of science. If a liberally educated person is one who can make critical judgments of his society and his time, who today is liberally educated if he knows nothing about science? It would be foolhardy and undesirable to try to make every bright student a scientist. It would be impossible to stock a general student's head with scientific facts sufficient for him to be knowledgeable, even for a brief time after graduation, about the broad expanse of science. Yet it is most unfortunate to send him into a world evolving so swiftly under the impact of scientific knowledge without a grasp of scientific method, an elementary understanding of the larger principles of science, an appreciation of the influence of science in philosophy, economics, and history, and a knowledge of the power and dynamics of science in creative evolution.[5]

Since there is a direct relationship between what a student learns in his undergraduate study and his success in professional school,[6] the investigator is interested in seeking methods which will perhaps result in the type of learning on the part of students to which Seaborg referred.

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[5] Dr. Glen Seaborg, Current Issues in Higher Education, "Education for the Third Revolution", Association for Higher Education, Washington, D. C., p. 9.

[6] Dickson, Jordan, Schloerb and Stuit, Predicting Success in Professional Schools, American Council on Education, Washington, D. C., 1949, p. 89.

The review of the literature revealed that there were only four research studies which dealt with the teaching of college chemistry and only two of those dealt with chemistry laboratory instruction.[7] The purpose of Montague's research was to show that the college chemistry laboratory could be used to develop an understanding of problem solving in science. The purpose of Vlassis's study was to determine whether there were any significant differences between the results of a conventional general chemistry laboratory course and a laboratory-centered general chemistry course as far as student interest was concerned.

The data from Montague's research allow him to reach the following conclusions:

1. The ability of the experimental group to solve problems in an actual laboratory situation, as measured by the performance test, was significantly greater than that of the control group.

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[7]George Charen, "A Study of the Effect of Open-ended Experiments in Chemistry on the Achievement of Certain Recognized Objectives of Science Training" (unpublished Ed.D. dissertation, University of Colorado, 1962).

Molly S. Geller, "The Measurement of the Effectiveness of a Teaching Machine Program in the Area of College Chemistry" (unpublished Ph.D. dissertation, New York University, 1963).

Earl J. Montague, "Using the College Chemistry Laboratory to Develop an Understanding of Problem Solving in Science" (unpublished Ph.D. dissertation, The Ohio State University, 1963).

C. G. Vlassis, "A New Approach to General Chemistry Laboratory in College with a Laboratory Manual for General Chemistry" (unpublished Ed.D. dissertation, University of the Pacific, 1963).

2. There was no significant difference between the experimental and control groups in their ability to think scientifically about other areas of science as measured by A Test of Aspects of Scientific Thinking.
3. There was a significant increase in the ability to think scientifically as measured by the differences between the pre-test and post-test scores of A Test of Aspects of Scientific Thinking, for both the experimental and the control groups.
4. The difference between the experimental group and the control group in the ability to think critically about everyday problems, as measured by the Watson-Glazer Critical Thinking Appraisal, favored the experimental group, with the difference being significant at the five percent level.
5. There was no significant difference between the groups in subject matter achievement as measured by the final course examination.[8]

Montague's research measured the achievement of both the control and experimental groups in terms of ability to think scientifically and solve problems in a laboratory situation. The results of the final examination did not show any significant difference between the experimental and control groups in relation to subject matter mastery.

The study conducted by Geller involved the comparison of using a teaching-machine method with the conventional method of instruction for organic chemistry. She evaluated her hypotheses as follows:

1. Immediate learning and retention would be enhanced by machine teaching was not sustained.
2. Students who used teaching machines would show greater interest in organic chemistry was not sustained.
3. Better readers would learn more through the use of teaching machines was not sustained.

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[8]Montague, op. cit., p. 84-85.

4. In each case the experimental and control students achieved equally well.[9]

The study conducted by Charen concerned itself with a comparative investigation of the effectiveness of using conventional laboratory experiments for one group of students and experiments developed by the Manufacturing Chemists' Association for another. The study was conducted using high school students. The conclusions he reached were:

1. Both laboratory techniques resulted in learning.
2. The Manufacturing Chemists' Association experiments were significantly better for promoting understanding of facts.
3. Neither was superior in improving critical thinking in chemistry.
4. Pupils preferred the Manufacturing Chemists' Association experiments to the conventional ones.[10]

The study conducted by Vlassis was of a similar nature to the study carried out by the investigator but the objectives were not the same and the findings were quite different.

> Several methods of evaluating the laboratory-centered general chemistry course in comparison to the conventional general chemistry laboratory course were used. First, the groups were compared on quantitative reasoning ability based on the ACE Psychological Examination. The second evaluation made was on the basis of examinations administered to both groups. Third, the evaluation of student reactions toward the lecture and the laboratory was made.
>
> Since the students' reactions to the laboratory work were not significantly different either it might be assumed that the results of the examinations would not be different. However, from the data collected, the original hypotheses did not hold. The students in the laboratory-centered group did not understand chemical theory as well as the conventional group and they did not show a favorable attitude toward the laboratory work.

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[9]Geller, op. cit., p. 78.

[10]Charen, op. cit., p. 222.

The fact that the laboratory-centered group did not have results as favorable as the conventional group might have been due to the use of the ACE Psychological Examination which has a reliable coefficient of 0.6.

The fact that the attitudes of the two groups were not different may be due to the fact that the teacher might create the same degree of enthusiasm and motivation independently of the method of teaching.[11]

The present investigator based the need for this study on the following:

1. Vlassis has stated that further research in laboratory learning would be desirable.[12]
2. A desire of the investigator to experimentally determine if with the laboratory method as the only variable there was indeed a significant difference between the achievement of experimental and control groups as measured by scores on questions on subject matter dealt with in both laboratory groups.
3. No evidence was found by the investigator that research had previously been done using the inquiry method in general chemistry laboratory instruction.
4. The research works conducted by Vlassis and Montague both failed to show any significantly better performance by the experimental group with respect to mastery of subject matter.[13]

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[11] Dissertation Abstracts, University Microfilms, Inc., 1962, XXIII, p. 412.

[12] Vlassis, op. cit., p. 37.

[13] Ibid., p. 35.

5. None of the aforementioned studies were concerned with the inquiry approach to learning.
6. Advancement of knowledge in science education as in other fields is likely to come as a result of original experimental research conducted in the actual learning situation.
7. The investigator has observed that a teaching method which is more effective than the conventional method of instruction needs to be used in the general chemistry laboratory.

CHAPTER III

THE DESIGN OF THE STUDY

During the Spring semester of 1966 the investigator was a full-time student at Oklahoma University. One of the courses studied during that period was a reading course which dealt with the method of inquiry in instruction. The readings were works written by Renner[1], Elam[2], Ford and Pugno[3], and Karplus[4], all workers in the field of inquiry-method of instruction. As a result of taking this course the writer became interested in the inquiry method and decided to carry on an original study using that method with students in his laboratory sections of beginning college chemistry.

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[1]John W. Renner, Science, Elementary School Children and Learning, Educators Progress Service, Randolph, Wisconsin, 1965.

John W. Renner, The Thrill of Discovery in Science Education, Educators Progress Service, Randolph, Wisconsin, 1961.

[2]Stanley Elam, Education and the Structure of Knowledge, Rand McNally and Co., 1964.

[3]G. W. Ford and L. Pugno, The Structure of Knowledge and the Curriculum, Rand McNally and Co., 1964.

[4]Robert Karplus, One Physicist Looks at Science Education, Science Curriculum Improvement Study, University of California, Berkeley, 1963.

The study consisted of three separate, unique experiments which were done in succession. The three experiments were carried out during two academic years; the 1966-67 school year and the 1967-68 school year. A minimum of one control and one experimental group was involved each time an experiment was conducted and in each experiment more variables were controlled than had been controlled in the previous experiment. For that reason, the design of the study could be entitled the "Progressive Control of Variables." In the final experiment the only variable was the actual materials the students were using. The control groups used the conventional laboratory manuals provided by the college bookstore and the experimental groups used the laboratory procedures developed by the investigator.

Because of the random process of registration the subjects became members of the experimental and control groups by chance. There was, therefore, no reason to believe that one group was academically superior to the other. To test that belief, the null hypothesis was tested, using a t-test, between the ACT natural science scores and composite scores for all groups involved in each experiment. The 0.05 level of significance was used in all cases. The Null Hypothesis was also used to test whether or not there would be differences between the control and experimental groups on the basis of laboratory achievement at the .05 level of significance, as indicated by the t-test. That achievement was measured through performance on a final examination and, in one experiment, pre- and post-laboratory tests.

The first experiment carried out during the Fall of 1966 had the most uncontrolled variables. The control group students could have had any available laboratory instructor, any available lecture instructor, any of the available laboratory assistants, and were not all in the same laboratory section. The experimental group students could have had any available lecture instructor and any of the available laboratory assistants, but they were all in the same laboratory and used the inquiry materials developed by the investigator. Thirty-three experimental group and thirty-one control group students were involved in the Fall of 1966 experiment of the study. The mean final laboratory examination scores for both groups were subjected to the t-test at the 0.05 level of significance. Near the close of the Fall of 1966 semester a science educator not known on the Central State campus recorded interviews with 24 of the experimental group students; their reactions and comments are found in Appendix J.

The second experiment of the study was conducted during the Fall semester of 1967. In that experiment the students of either the control or experimental group could have had any available lecture instructor or laboratory assistant. In the experiment, however, the investigator was the laboratory instructor for both the control and the experimental groups. In other words, this experiment had one less variable than the previous experiment done in the Fall of 1966 semester; i.e., the laboratory instructor was a constant factor.

The second experiment was different from either the first or third experiments of this study in that during the second experiment there were two independent experimental groups and one control group. Again the Null Hypothesis was used to test whether or not differences among the three groups existed (at the 0.05 level of significance) on their performance on a final laboratory examination: the examination was over the same experiments carried out in the laboratory during the semester by the three groups. Since the second experiment involved three independent groups the one-way analysis of variance was used to test for any significant differences among the three groups using F-ratios at the 0.05 level of significance. The findings were then subjected to the t-test to determine among which of the three scores the significant differences existed. The second experiment involved thirty-two control students and twenty-five students in each of the experimental groups.

The third experiment involving one experimental and one control group, was conducted during the Spring of 1968 semester. This final experiment differed from the previous experiments in that eight laboratory exercises performed by the experimental group were matched for content with eight laboratory exercises done by the control group. A pre- and post-test was developed for each set of matched laboratory exercises and administered at the time each exercise was done. The mean gains for both the control and experimental groups of post-test over pre-test were obtained and those data were subjected to a t-test at the 0.05 level of significance. The two groups also took the same final laboratory examination taken by the students involved in the two previous experiments. The mean final laboratory examination scores were subjected to the t-test at the 0.05 level of significance.

The third experiment differed from the previous two experiments in that during the third experiment all students who served as experimental subjects for both the experimental and control groups had the same lecture instructor, the same laboratory instructor, and the same laboratory assistant. The only variable for this experiment was the conventional laboratory approach versus the inquiry method. The Null Hypothesis, that is, there would be no difference between the experimental and control groups on the basis of the mean gains of post-test over pre-test nor any difference between the two groups on the basis of the mean final laboratory examination scores was tested, using the t-test, at the 0.05 level of significance.

In the summer of 1966 the investigator began to develop some experimental procedures for use in the beginning college chemistry laboratory. These were periodically submitted to and constructively criticized by the investigator's committee chairman and then revised according to the criticisms. Once the procedures for these experiments were developed they were not changed during the course of the study. The intention of the investigator was to ask questions of the student that would be thought provoking and at the same time be the kind of question to which the student could respond without having had a previous course in chemistry.

Undoubtedly, there were other experiments which could have been used in place of those that were selected, but the basis for experiment selection was:

1. The investigator's familiarity and previous experience with the procedures and reactions called for in the experiments.

   It was the investigator's judgment that the more knowledge he personally had with the procedures, the more effective he could be in helping the students.

2. The time element involved.

   The investigator felt that the experiments could be effective only so long as the student could work in a pleasant and relaxed atmosphere since in the investigator's opinion the length and complexity are common definite disadvantages of the commercial manuals.

3. Appropriateness to illustrating the desired principle or law.

The investigator selected those experiments which would, in his judgment, clearly reveal the principle involved.

4. The facility with which the experiments could be carried out with available materials and equipment.
5. The variety of types of chemical operations and methods that the experiments provided.

The experiments dealt with such classifications of experiments as analysis, synthesis, type reactions, characteristics of substances, and chemical processes. The experiments selected to be used were chosen with regard to their effectiveness in illustrating concepts such as density, hydration, formula weight, equivalent weight, reactivity of elements and the titration process.

The directions given in the suggested procedure of the experiments were actually tested by the investigator himself in the laboratory prior to issuing the experiment to the students. The purpose of this procedure was to insure that the student was not subjected to unnecessary personal risk, that the experiment did serve to make the desired principle clear, and could easily be carried out in the time allotted.

Every experiment that was developed included at least two textbook references by either page or chapter. One of the references was always the textbook used in the course regardless of which semester the experiments were used. The references gave adequate explanations and quite often illustrative examples. There was a number of other textbooks available in the college library that the student could and probably did use.

Eleven experiments were developed and finally repeatedly used in subsequent semesters by the students in the experimental groups. The titles of the experiments were: 1. Density of Substances. 2. Chemical Operations. 3. Properties of Substances. 4. Preparation and Properties of Oxygen. 5. Reaction of Acids and Bases. 6. Formula Weight of a Liquid. 7. Determination of Water of Hydration. 8. Relative Activity of Some Metals. 9. Equivalent Weight of a Metal. 10. Titration. 11. Solubility of a Salt. Numbers 1, 3, 4, 5, 6, 7, 8 and 9 of the previous list were selected as essentially equivalent experiments to eight of those selected to be done by those students in the control group the Spring semester of 1968.

The investigator taught both the experimental and control groups in the lecture and attempted to give both groups of students equal amounts of information. At the same time the investigator was careful to avoid answering questions in both the experimental and control laboratory sections which could be answered by laboratory observations.

Permission was obtained from the department chairman and the administration to conduct the study. Both the chairman and the administration encouraged the investigator to pursue the study, since the major effort at Central State College is instructional.

The first time that the experiments were used was during the fall semester of 1966. During this semester the investigator taught two laboratory sections for beginning college chemistry as part of the normal teaching load. One laboratory met each Tuesday from 1:30 to 3:30 while the other met each Wednesday from 1:30 to 3:30; both laboratory sections were used. Those students who were in the investigator's lecture sections but that were not in his laboratory sections constituted the control group for the fall semester 1966. That group had other instructors for their laboratory sessions. They used the commercially available laboratory manual that was adopted for use at Central State College. The same laboratory assistant (an upper classman majoring in chemistry) helped in both laboratory sections.

Toward the end of the semester a science educator not known on the campus of Central State College interviewed twenty-four students that had been members of the experimental group. Those interviews were recorded (see Appendix J). The students attended this interview on a voluntary basis.

During the last laboratory session the experimental groups took a final laboratory examination prepared and administered by the investigator. During the last class meeting of the three lecture sections those students in other laboratory sections took the same laboratory final examination.

The scores for this examination are included in the Appendix with the ACT Natural Science Scores and the ACT Composite Scores of the groups.

The next group of students to be participants in this study were those enrolled in the investigator's three sections of beginning college chemistry laboratory meeting at respectively, 7:30 to 9:30 Tuesday, 11:30 to 1:30 Tuesday and 11:30 to 1:30 Thursday during the fall semester 1967. These students had no prior knowledge that they were to be involved in an experiment and as a result could not have enrolled in any of those sections for that reason.

The fall 1967 semester differed from the fall 1966 semester in that during the fall 1967 semester the two experimental and one control laboratory groups were taught by the investigator. During the fall 1966 semester the two experimental laboratory groups were taught by the investigator while the control laboratory group was taught by other staff members of the Chemistry Department. Some of the fall 1967 laboratory students were in other instructor's lecture sections, but the majority of them were in the investigator's lecture sections.

During the fall 1967 semester different laboratory assistants helped in the laboratory sections. The investigator's observations were that these people were of about equal competence in the area of chemistry. The two experimental groups again used the same laboratory procedures developed by the investigator, while the control group used experiments from the commercial manual. The eleven experiments used were those that dealt with essentially the same topics as those used by the experimental group. The following is a list of those experiments used by the control group:

1. The Metric System. The Density of Solids and Liquids.
2. Physical Properties of Substances.
3. Some Elementary Chemical Properties of Substances.
4. Preparation of Pure Substances by Chemical Changes.
5. The Chemistry of Oxygen.
6. The Formation of Salts.
7. The Molal Volume of a Gas.
8. The Formula of a Hydrate.
9. The Activity of Certain Metals and the Stability of their Oxides.
10. The Equivalent Weight of a Metal.
11. Titration of Acids and Bases.

For the same semester that the above list was used by the control group, the following list of experiments was used by experimental groups:

1. Density of Substances.
2. Properties of Substances.
3. Chemical Operations.
4. Solubility of a Salt.
5. Preparation and Properties of Oxygen.
6. Reaction of Acids and Bases.

7. Formula Weight of a Liquid.

8. Determination of Water of Hydration.

9. Relative Activity of Some Metals.

10. Equivalent Weight of a Metal.

11. Titration.

In every case the same or similar chemicals were used to provide the experimental data for the principle under consideration for both the control and experimental groups. In the fifth experiment the experimental group obtained a gas by vaporization of a volatile liquid while the control group obtained oxygen as a gas from the thermal decomposition of a solid chemical compound. The experimental group did an experiment on the solubility of a salt which the control did not do. The main difference between the written materials used by the experimental and control groups was that the experiments used by the experimental groups employed the inquiry approach while the control group used the traditional experiments. Again during the last laboratory meeting of the semester each group took the same final examination as the 1966 students took. These were not returned nor seen again by students.

During the Spring 1968 semester a more carefully regulated enrollment of students was used. Directions were printed in the enrollment schedule for the Spring 1968 semester which regulated the enrollment of students in the investigator's sections of lecture and laboratory. Only those in the investigator's lecture sections could enroll in one of his laboratory sections, and those in the laboratory sections taught by the investigator were required to enroll in one of his lecture sections. This arrangement provided for a better control over the hypothesis being tested and also eliminated the variable of different lecture instructors that was present the previous semester.

The investigator wanted to make a more concrete comparison of experimental and control groups than the comparative final laboratory examination scores for the Spring semester of 1968. To make this possible the investigator developed a series of pre- and post-tests (see Appendix B) with the guidance of his chairman and several committee members. Each of these tests consisted of four question and/or problems for each pre- and post-test except for the eighth one; it had two questions and one problem. The pre- and post-tests were planned so that they would require from 8 to 10 minutes. The pre-test was administered at the end of the class session on Monday and the experiment was conducted by both groups on the following Tuesday. The post-test was administered during the last ten minutes of the laboratory period.

The experiments used by the experimental group during Spring semester 1968 were carefully selected to correspond to experiments used by the control group. The list used by the experimental group included the following:

1. Density of Substances.
2. Properties of Substances.
3. Determination of Water of Hydration.
4. Preparation and Properties of Oxygen.
5. Reactions of Acids and Bases.
6. Equivalent Weight of a Metal.
7. Relative Activity of Some Metals.
8. Formula Weight of a Liquid.
9. Titration.
10. Chemical Operations.
11. Student Designed Experiment.
12. Student Designed Experiment.

The list of experiments used by the control group from the commercial manual for the Spring semester 1968 was:

1. The Metric System. The Density of Solids and Liquids.
2. Some Elementary Chemical Properties of Substances.
3. The Formula of a Hydrate.
4. The Chemistry of Oxygen.
5. The Formation of Salts.
6. Equivalent Weight of a Metal.
7. The Activity of Certain Metals and the Stability of Their Oxides.
8. The Molal Volume of a Gas.
9. Titration of Acids and Bases.
10. Preparation of Pure Substances by Chemical Changes.
11. Student Designed Experiment.

12. Student Designed Experiment.

The first eight experiments from each of the previous lists were carefully matched by topics which the experiments treated. The eight pre- and post-tests found in Appendix B are in the same order as these first eight experiments, correspond with them and were given before (pre-test) the experiment and after the experiment (post-test).

Near the end of the semester two laboratory periods were used by the students (both control and experimental) to design and conduct an experiment on their own (experiments 11 and 12 of the previous lists). The instructor presented them with a problem and instructed them to use whatever materials and equipment they needed. They wrote up a report and turned it in to the instructor at the beginning of the next laboratory period. This reporting procedure was used throughout the semester by both groups.

One of the problems presented to both groups was that of experimentally determining the solubility of potassium chloride in water. No directions whatever were given to either group. They wrote out their own procedure, conducted the experiment, collected their data and wrote up the report. The other problem presented to them was that of being given a solid and a liquid substance and directed to obtain as many true facts concerning their properties as they could during the limits of the laboratory period. Later both groups of reports were graded by the investigator. The first experiment was evaluated on the validity of the procedure used, the accuracy of the results obtained and the student's recognition and statement of the important factors which did or could have had an effect on the results of the experiment. In the second experiment the reports were scored on the basis of the actual number of correct facts reported.

The investigator taught two lecture and laboratory sections of beginning freshman college chemistry during the Spring semester 1968. One laboratory section met from 9:30 to 11:30 Tuesday and the other met from 11:30 to 1:30 Tuesday. Again the decision was made to use the 9:30 section to be the experimental group before the students enrolled but this choice was not known by the students. The same laboratory assistant helped both laboratory sections. This eliminated the variable of different laboratory assistants that existed the previous semester.

In the fall of 1966 thirty-three students completed the semester as members of the experimental group and thirty-one for the control group. In the fall of 1967 thirty-two students completed the semester as members of the control group and twenty-five each for the two experimental groups. In the spring of 1968 thirty-one students completed the semester as members of the experimental group and twenty-five for the control group. A total of 202 students in both experimental and control group for three semesters participated in the study.

DENSITY OF SUBSTANCES

References: Quagliano, James V. Chemistry, Second Edition, Prentice-Hall, Inc., Englewood Cliffs, New Jersey, p. 15.

Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, p. 14.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, p. 4.

Purpose

The purpose of this experiment is to provide the student with a concept of what density means and to show him that information in science can often be gained by indirect methods.

Introduction

Density is a term used in physical measurement in which weight and volume are dependently related. Density is defined as being units of mass divided by units of volume. Density can be expressed in any units of mass measurement related to any units of weight measurement but usually in chemistry mass is expressed in grams and volume in milliliters.

Necessary Materials: One 50 ml. beaker, one 50 ml. graduated cylinder, pieces of aluminum, copper and lead metals, laboratory scales, liquids suitable to be used as liquids with unknown density.

Procedure: Weigh a 50 ml. beaker on a laboratory balance and record the weight. Next pour out 40 mls. of water from your graduated cylinder into the beaker and weigh again and record the weight. Now determine the density of the water.

Next weigh several pieces of copper metal and record the weight. Then place the metal pieces in a known volume of water (use your cylinder) and determine the volume of the metal pieces by noting the volume of water displaced. Now determine and record the density of the copper. Repeat the procedure for the other two metals.

Next obtain a sample of unknown liquid from the storeroom and determine and record its density. Name two liquids that have a density less than that of water. What are two liquids that have a density greater than water? If your experimentally determined density for some of the metals does not agree with the values given in the handbook of chemistry what are some possible explanations? Why should the temperature of the water remain constant during experiments of the kinds used to determine the density of metals? What volume of mercury (density 13.6 gms/ml) would be required to weigh the same as 20 mls. of magnesium (density 1.7 gms/ml.)?

Experimentally determined density of water ________________

Metal	Copper	Aluminum	Lead
Weight of metal			
Volume of metal			
Density of metal (experimentally calculated)			
Density of metal (Handbook)			

Density of unknown liquid ________________

CHEMICAL OPERATIONS

Resources: Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, p. 217.

Quagliano, James V. Chemistry, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, p. 206.

Sienko and Plane, Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 365, 369 and 392.

Introduction

When a finely divided solid is suspended in a liquid it can often be removed by passing the liquid through a filtering medium such as asbestos fibers or a filter paper and trapping the solid on the filtering medium. Some substances dissolve in water and could not be removed by this method. Such substances however, can be removed by allowing the solvent to evaporate from the solute (that which is dissolved). Sometimes it is necessary to recover the solvent after it is separated from the solute. One method of separation involves vaporizing the solvent and then cooling the vapor so that it condenses back to a liquid in another separate container.

Some substances can be purified by heating the solid, changing it to a vapor and then condensing the vapor back to a solid that is more pure than the original solid if the original contained impurities. Necessary materials: Filter paper, glass tubing, a test tube, a one-hole stopper to fit the test-tube, matches, burner, watch glass, beaker, funnel, wire triangle, iron ring, ring stand, copper sulfate crystals, sodium chloride, .2M solutions of potassium iodide and lead nitrate, and iodine.

Procedure: Pour 10 mls. of potassium iodide into a beaker and observe its physical appearance. To this add 10 mls. of lead nitrate solution. What do you observe to happen?
Using the equipment provided for you, make a mechanical separation of the solid and liquid. What is the appearance of the liquid that comes through the paper?
What name is given to this liquid?
What is the color of the solid precipitate?
What is the formula of the solid precipitate?

Next put 1 gram of common salt and 1 gram of copper sulfate in 15 mls. of water in a large test tube and fit the test tube with a one-hole stopper and a U-shaped glass tube whose delivery end can be placed in another test tube. Clamp the test tube to the ring stand and heat the tube carefully with the burner in order to boil the liquid. Catch a few mls. of the condensed vapor in another test tube that is placed in a beaker of cold water. Taste a drop of the liquid collected in the test tube in the cold water. Does it taste salty? Should it? Why does the distillate differ in color from the original liquid?

Now clamp your iron ring to the ring stand, place the wire gauze or wire triangle on the iron ring and place an evaporating dish on the wire. Place 2 or 3 crystals of iodine in the evaporating dish and cover the dish with a watch glass containing a few mls. of cold water. Using the burner, heat the evaporating dish for a few minutes and observe any formations on the bottom of the watch glass.

What are the crystals on the watch glass?

What would be a use for this type of process?

What is the name given to the process of going from solid directly to vapor?

How could one separate sand from sugar?

Alcohol from water?

What were the operations carried out in this experiment?

Write an equation for the reaction of the lead nitrate with potassium iodide.

PROPERTIES OF SUBSTANCES

References: Nebergall, Holtzclaw and Schmidt. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, pp. 6-7.

Quagliano, James V. Chemistry, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, pp. 10-11.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, p. 5.

Purpose

The purpose of this experiment is to give the student some experience making observations related to the properties of substances.

Necessary materials: Powdered sulfur, aluminum, and iron, liquid mercury, solid iodine, ethyl alcohol, matches, carbon disulfide, and ordinary bar magnet, paraffin wax, beakers, evaporating dishes, mortar and pestle, test tubes, burner, test tube holder, and pieces of magnesium ribbon.

Procedure: Test the magnetic properties of both iron and sulfur powder. What did you observe?

Next test the solubility of both the iron and sulfur powder in carbon disulfide (only a pinch of each of the powders is necessary). What are their solubilities in carbon disulfide? (BE VERY CAREFUL NOT TO BRING CARBON DISULFIDE NEAR AN OPEN FLAME AS IT IS DANGEROUSLY FLAMMABLE). Now take a quantity of sulfur about equal in volume to a pea and twice this amount of powdered iron and mix the two in a test tube. Heat to redness and continue to heat until no further activity within the test tube can be observed, place the hot tube on the base of the iron ring stand and allow it to cool and prepare another mixture of aluminum and sulfur but this time use a ratio of two parts of sulfur to one part aluminum. Heat until apparent reaction is complete and place the test tube again on the ring stand base and allow to cool.

After the two tubes have cooled, test the iron sulfur combination for magnetic properties. What did you observe? Did you observe that the iron still possessed its original properties? Next add a few drops of water to each of the test tubes containing the iron-sulfur and aluminum-sulfur combinations. Did you observe any change in either case as evidenced by visible reaction or evolution of any gases?

What would you conclude as to the stability of the two metal-sulfur combinations in the presence of water?

Next place a globule of mercury the size of a BB in an evaporating dish and add an equal volume of solid iodine and mix the two with a stirring rod. Add a drop of alcohol to facilitate reaction. What is the most obvious change that occurs?

Are the visible characteristics of the two elements still apparent?

Now take a small piece of paraffin and add about an equal quantity of sulfur and heat until the paraffin melts. Quickly pour the liquid out on a piece of paper towel. Are any of the characteristics of the sulfur still present? Treat some paraffin and iron powder in the same way. After allowing a little time for cooling, test the magnetic properties of a small piece of the mixture containing the iron. What did you observe?

Take a small piece of the iron-sulfur combination prepared earlier and see if it has solubility in carbon disulfide. Now list what you think were chemical changes.

Take a piece of magnesium ribbon about one inch in length and hold with the tongs in the hottest part of the bunsen flame. Is there evidence of a chemical change? What is the evidence?

What were the sources of energy that caused the chemical changes you observed?

What are the basic differences between chemical changes and physical changes?

Write chemical equations for what you believe are chemical reactions observed in today's experiment.

THE PREPARATION AND PROPERTIES OF OXYGEN

References: Quagliano, James V. Chemistry, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, Chapter 6.

Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, Chapter 5.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 333-345.

Purpose

The purpose of this experiment is to give the student some knowledge of the most commonly used method of preparing oxygen in the laboratory and the properties of some of the representative compounds of oxygen.

Necessary materials: One large pyrex test tube, a one-hole stopper to fit the large test tube, glass and rubber tubing, ring stand, test tube clamp, burner, five gas bottles, a pneumatic trough, matches, potassium chlorate, manganese dioxide, sulfur, charcoal, phosphorus, steel wool, calcium metal turnings, red and blue litmus paper strips, five glass plates and a deflagrating spoon.

Procedure: Set up an oxygen generator consisting of the one-hole stopper, a large test tube, glass tubing and rubber tubing. Mix thoroughly 5 grams of potassium chlorate with

1 gram of manganese dioxide and place it in the test tube in such a way that it is somewhat spread out in the test tube rather than being in a compact lump in the very bettom of the tube. Next fill the five gas bottles with water and place a glass plate over the mouth of each bottle, invert them and place in the partially filled pneumatic trough, then remove the glass plates. What now keeps the water in the bottles?

Next make sure that the gas delivery tube is in the mouth of the bottle and start heating the chlorate-manganese dioxide mixture to produce the oxygen which will replace the water from the bottles. It may be necessary to momentarily remove the flame from the tube of the generator if the gas production becomes too rapid.

As each bottle of gas is obtained slide the glass plate over the mouth of the bottle while it is still inverted in the water and then invert the bottle and place it upright on the desk leaving the glass plate on it.

When the five bottles of gas are obtained turn off the burner and after allowing several minutes for cooling, dismantle the generator. Next take a small quantity of sulfur (about equal to the volume of a match head) and place it in a deflagrating spoon and heat over the flame until it ignites and thrust it into a bottle of oxygen. What do you observe to happen?

What reason can you offer for this observation?

Place a little distilled water in the bottle (about ¼ inch in the bottle) and replace the glass plate and shake vigorously for a few seconds and then test the solution with both blue and red litmus paper strips. What do you observe? How do you account for it being as you observe? Suggest a formula for this compound formed and verify your choice by reference to a text.

Next heat about an equal quantity of phosphorus in the deflagrating spoon and plunge it into another bottle of oxygen. What compound do you think is produced here? Test it in the same way with water and litmus paper strips as in the case with the sulfur. What do you observe about its basicity or acidity? Write an acceptable formula for an oxide of phosphorus and then balance an equation showing its reaction with water.

Repeat the above procedure using a small quantity of charcoal in the spoon and the third bottle of oxygen and again test the pH of its solution and record the results. Write an equation showing the reaction of water and this compound of carbon.

Next heat a small ball of steel wool about the size of a marble until it glows and quickly plunge it into the fourth bottle of oxygen. What do you observe to happen? What property of oxygen does this illustrate? Why does steel wool better serve to demonstrate this than an iron nail for instance?

Heat a few calcium metal turnings in the deflagrating spoon until very hot and then plunge them into the fifth bottle of oxygen. Observe the reaction. After allowing a few minutes for cooling add some water and test again with litmus paper. What do you observe?

How can you account for the pH of this solution being different from that of the others you tested?

What rather general rule could you formulate concerning the acidity and basicity of different oxides?

What general properties of oxygen have you found? Be sure to describe them thoroughly, and give equations for all reactions observed in this experiment.

THE REACTION OF ACIDS AND BASES

References: Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, pp. 330-40.

Quagliano, James V. Chemistry, Second Edition, Prentice-Hall, Inc., Englewood Cliffs, New Jersey, pp. 460-482.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 226-30.

Purpose

The purpose of this experiment is to provide chemical evidence of the differences between acids and bases. The reactions carried out in this experiment will result in the formation of salts that can be isolated.

Introduction

Acids and bases can be thought of as being opposites in chemistry. Acids and bases are sometimes distinguished from each other by using small quantities of organic compounds called "indicators" that have a different molecular arrangement in acid medium than in basic medium, and also different colors.

Acids can be considered to be substances that yield hydronium ions upon reaction with water and bases are those that yield hydroxide ions upon reaction with water. In this experiment the acidity or basicity of several chemical compounds will be verified and salts will be produced by neutralizations.

Necessary materials: Magnesium metal strips, copper metal turnings, iron powder, calcium metal turnings, watch glass, evaporating dish, beakers, test tubes, solutions of KOH, $Ca(OH)_2$, NaOH, HCl, H_2SO_4, $HC_2H_3O_2$, NH_4OH, methyl orange and phenophthalein, solid Na_2CO_3. Procedure: Pour about 5 ml of each of the solutions of KOH, $Ca(OH)_2$, NaOH, NH_4OH, HCl, H_2SO_4, and $HC_2H_3O_2$ into separate test tubes and add about 3 drops of phenolphthalein solution to each of them and observe and record the resulting colors. Empty the test tubes and repeat the above procedure using instead this time 3 drops of methyl orange solution and observe and record the colors. What accounts for the colors of all these to fall into two groups?

What two classes of substances are being dealt with here?

Add about 3 ml of NaOH to about 3 ml of HCl and place in an evaporating dish and heat gently until the water is evaporated. What is the white substance in the dish? Write an equation that represents its formation.

Now pour about 5 ml of HCl into each of four test tubes and to one add some calcium metal and observe the reaction. Write an equation for the reaction. To the second tube add some magnesium metal. Observe the reaction and write an equation to show the reaction. To the third test tube add some iron powder and observe any reaction. What does its speed of reaction compared to the other two metals already tested imply?

Observe the same tube 10 minutes later and see if any or all of the iron has disappeared. In the fourth tube put in some copper metal and observe. Is there any apparent reaction? Would you expect a reaction? Why or why not?

Drop a piece of magnesium into 5 ml of the $HC_2H_3O_2$ solution and compare the rate of reaction to that of the reaction with the HCl. How do you account for the comparatively different rates?

Next take about a gram of Na_2CO_3 and add slowly enough HCl to get a good reaction and to cause all the solid Na_2CO_3 to disappear. Then evaporate the solution to dryness. What is the white residue that is left?

Write an equation to express the reaction that occurred upon adding the HCl to the Na_2CO_3. Write a statement about the relationships of acids, bases and salts.

Complete and balance:

(a) $Fe_2O_3 + HCl \longrightarrow$

(b) $Ca(OH)_2 + H_3PO_4 \longrightarrow$

(c) $Ca(OH)_2 + H_2SO_3 \longrightarrow$

(d) $Mg(OH)_2 + HCl \longrightarrow$

(e) $Ba(OH)_2 + H_3PO_4 \longrightarrow$

(f) $Fe_2O_3 + HBr \longrightarrow$

(g) $HC_2H_3O_2 + MgO \longrightarrow$

(h) $ZnO + HNO_3 \longrightarrow$

FORMULA WEIGHT OF A LIQUID

References: Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, pp. 133-4.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 148-9.

Purpose

The purpose of this experiment is to give the student an actual experience in finding the formula weight of an unknown liquid by relating the weight of a gas to a known volume at a measured temperature and pressure.

Introduction

If the volume that a known weight of vapor occupies at a known pressure and temperature can be determined, then it is possible to determine experimentally the formula for molecular weight of the compound. This is done by using the factors of the ideal gas law which follow:

P = pressure measured in atmospheres (1 atmosphere = 760 mm mercury pressure)

V = volume measured in liters

n = number of moles (grams/ grams per formula weight)

R = .082 liter atmospheres/ degree mole (the gas constant)

T = the temperature measured in degrees Kelvin (degrees C. + 273)

(The value of .082 liter atmospheres/degree mole is derived from the comparison of the gram molar volume of any gas at standard temperature, 22.4 liters, to the correction factor for correcting degrees centigrade to degrees Kelvin, 273.)

$$\frac{22.4 \text{ liters/mole}}{273 \text{ degrees/atmosphere}} = .082 \text{ liter atmosphere/ degree mole}$$

<u>Necessary materials</u>: One 250 ml. erlenmeyer flask, one one-hole stopper, one glass jet tip, burrette clamp, ring stand, iron ring, wire gauze, matches, bunsen burner, laboratory scales, wall barometer, an unknown organic liquid, and a thermometer.

<u>Procedure</u>: Place the glass jet tip in a one-hole stopper and place the stopper in the erlenmeyer flask (250 ml) and weigh this assembly as accurately as possible on the laboratory balance. Record this information. Now take the flask to the stockroom and obtain about 4 ml of an organic liquid. Place it in the weighed flask. Replace the stopper

in the weighed flask and set up your apparatus such that the flask containing the unknown organic liquid can be held submerged as far as possible in water in a beaker which is to be heated. Now bring the water to a boil and continue to boil for about five minutes after there are no liquid drops remaining in the flask. Record the temperature of the boiling water. How could the temperature of the boiling water have been determined if a thermometer was not available? Now allow the flask to cool somewhat and remove it from the water and dry it off carefully and weigh it again as accurately as you can. Record this weight. Now determine the volume of the flask assembly by filling the assembly completely full of water and weighing. (At room temperature one gram of water is very close to one ml of volume). Read and record the barometric pressure.

Why did it not matter that we know the exact amount of liquid that was placed in the flask at the start? What parts of the ideal gas law are known to us from the data we gathered? Now algebraically rearrange the ideal gas law equation to equate the known values to the unknown molecular weight. Other than errors in weighing, what might cause your experimentally calculated molecular weight to be greater than its actual value? Show your calculations and your experimentally determined molecular weight. What is the maximum boiling point that a compound may have and still be used in this experiment?

Problems:

(1) If 22 grams of a gas occupies 24 liters of volume at 700 mm of mercury pressure and at $265^{\circ}C.$, what is its molecular weight?

(2) What volume will 13 grams of benzene (C_6H_6) occupy at $100^{\circ}C.$ and at 380 mm of mercury pressure?

DETERMINATION OF WATER OF HYDRATION

References: Nebergall, Holtzclaw and Schmidt, General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, p. 168.

Quagliano, James V. Chemistry, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, p. 234.

Sienko and Plane, Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 344-5.

Purpose

The purpose of this experiment is to study the amount of water in a specific hydrate. The ratio of combined water to salt is constant and this can be shown by experiment.

Introduction

Certain metallic salts upon crystallizing from water solution-form hydrated salts. Such a compound has a definite percentage of combined water in its crystal, although there has been no chemical reaction between the water and the salt. The percentage of water in a hydrate can be determined by first weighing a sample, then heating it strongly to drive off water and weighing again. The percentage of water in the hydrate can then be determined.

Necessary materials: Hydrated crystals of copper sulfate and barium chloride, matches, wire triangle, iron ring, ring stand, laboratory balance, crucible, burner, crucible tongs, and spatula.

Procedure: Obtain about 5 grams of copper sulfate crystals. Using the triple beam balance, weigh your sample as accurately as you can. Place the crystals in the crucible and weigh again. Record the two weights. Place the crucible on the wire triangle on the iron ring and adjust the height of the ring on the ring stand so that the inner blue cone of the burner flame will just touch the bottom of the crucible. Heat the crystals for about 25 minutes. Why is occasional stirring necessary? (Be sure to use care so none of the crystals will be lost from the crucible.) After heating for the specified time remove the crucible from the flame and place on the iron base of the ring stand to cool. Record the weight after cooling.

Repeat the above procedure only this time use crystals of barium chloride hydrate and while these crystals are heating calculate the formula for the hydrated crystals of copper sulfate. (Hint: You now have available the weight of combined water since this is equal to the weight loss. The molecular weight of water is 18 and the formula weight of $CuSO_4$ is 160.) If you are not able to make the calculations ask the instructor or laboratory assistant to help you. Make the same determination for the barium chloride hydrate and calculate its formula from your experimental data.

What might you assume a substance to be if, upon heating it lost weight? If your calculations indicate a lower degree of hydration than is actually the case ($CuSO_4 \cdot 5H_2O$, $BaCl_2 \cdot 2H_2O$) what might be the source of error? If your calculations indicate a higher degree of hydration than is the actual case how might you account for this?

RELATIVE ACTIVITY OF SOME METALS

References: Nebergall, Holtzclaw and Schmidt, General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, p. 110.

Quagliano, James V. Chemistry, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, p. 135.

Sienko and Plane, Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, p. 310.

Purpose

The purpose of this experiment is to study the relative chemical reactivity of some of the common metals.

Necessary materials: Solid potassium, sodium and zinc metals, powdered or granular magnesium, sinc, and iron, tin sheet, and copper turnings, hydrochloric acid (0.1M), litmus paper strips, test tubes, matches, burner and tongs, iron nails, 0.1M $CuSO_4$ and 0.1M $AgNO_3$.

Procedure: Place a piece of potassium metal the size of a large BB shot, which has been wrapped in dry filter paper (use your tongs), into the mouth of a test tube which is filled with water and inverted in a beaker of water. Be careful that the water is not allowed to run out of the test tube. Observe the reaction and test the liquid in the beaker with litmus paper after the reaction has ceased. Keeping the test tube upside down, bring it to the flame of your burner. What do you observe?

Considering the reactants here and your observations, write an equation that represents the reaction between potassium and water.

Using a piece of sodium metal about the same size as that of the potassium and a fresh beaker of water, do the same experiment you did with the potassium and record your observations.

Place 0.6 gram of granulated magnesium in 100 mls of water in a 250 ml erlenmeyer flask. Close the flask with a stopper equipped with a delivery tube which leads to a small test tube which has been filled with water and inverted in a water-filled beaker. Boil the water magnesium mixture in the flask and collect the escaping gas in the test tube. When the test tube is half full of gas bring it to the flame in an inverted position. What did that tell you?

Now place about 5 mls of 0.1M HCl in each of four test tubes and place in the separate tubes about a gram of the metals zinc, iron, tin and copper. Observe and record any reactions that occur in a period of 5 minutes.

Place an iron nail in a small beaker containing copper sulfate solution and carefully observe for five minutes. Next place a small strip of copper metal in a small test tube containing 5 mls of 0.1 M $AgNO_3$ solution and observe for 10 minutes.

On the basis of your observations work out a list showing the relative chemical activity of the metals tested in this experiment. Work out correct equations for all the reactions that you observed to occur. List three things or conditions that will speed up chemical reactions.

THE EQUIVALENT WEIGHT OF A METAL

References: Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, pp. 136-7.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, p. 121.

Purpose

The purpose of this experiment is to provide the student with experimental procedure for determining equivalent weight of an element. The student will also learn the relationship between equivalent weight and combining valence of an element.

Introduction

There are two standards for comparison in expressing the equivalent weight of an element; hydrogen and oxygen. That amount of an element that will combine with or be displaced by one gram of hydrogen or eight grams of oxygen is said to be its equivalent weight.

The valence of an element is always a small whole number and is related to the equivalent weight of an element according to the following formula:

$$\text{Valence} = \frac{\text{Atomic weight}}{\text{Equivalent weight}}$$

In this experiment a weighed quantity of a metal will be dissolved in acid and the resulting hydrogen gas trapped in a gas measuring tube and its volume carefully measured. From the relationship of weight of metal to the volume of hydrogen gas produced at standard temperature and pressure, the equivalent weight of the metal will be determined. The metal used will be either aluminum or zinc. The student will be able to determine which it is by the volume of gas produced from a known quantity of the metal.

Necessary materials: Some pieces of metallic element, a 100 ml gas measuring tube, a one-hole stopper to fit the gas tube, a ring stand, burrette clamp, concentrated HCl, either a 400 or 600 ml beaker, a laboratory balance, and thermometer.

Procedure: Put approximately 200 ml of water in a 400 or 600 ml beaker and place it on the ring stand base. Weigh out from about .22 to .24 grams of the metal as accurately as you can and put it in the gas measuring tube. Next fill the gas measuring tube to about 65 ml with distilled water and then finish filling the tube with concentrated HCl. Carefully insert the one-hole stopper into the tube and quickly invert the tube (so that the closed end is up) and place the open end in the beaker and clamp the tube in a vertical position to the ring stand.

After all the metal has been dissolved by the acid, bring the level of the water in the tube to that of the water in the beaker by either adding water to the beaker or raising the gas measuring tube. When the two levels of the water are the same, read and record the volume of the confined gas in the tube. Take the temperature of the water in the beaker and record the corresponding vapor pressure found on page 621 of your text. Now set up your method of correcting the gas volume to standard temperature and pressure (be sure to make the correction for the vapor pressure of the water) and have it approved by the instructor or laboratory assistant before making calculations.

Now let us suppose that you used .24 grams of the metal and obtained 80 mls of hydrogen corrected to standard temperature and pressure. The method used to calculate the experimentally determined equivalent weight of the metal would be as follows:

$$\frac{11.2 \text{ liters/equivalent (hydrogen)}}{.08 \text{ liters produced}} \times .24 \text{ grams metal} =$$

33.6 gms/equivalent

How can you be sure that this relationship of three quantities gives the correct equivalent weight? Balance the equation for the reaction between aluminum and HCl. Balance the equation for the reaction between zinc and HCl. For which of these metals would you get the quantity of gas given in the above example if .24 grams of metal were used?

Determine the equivalent weight of the metal from your experimental data. From your experimentally determined equivalent weight and your decision as to whether the metal is zinc or aluminum, can you now see the relationship of the valence, equivalent weight and atomic weight of the metal and the equation for its reaction with HCl?

Divide the atomic weight of the metal used in your experiment by its valence and what do you have? Compare this with your experimentally determined value. How do they compare? What might account for the differences?

TITRATION

References: Nebergall, Schmidt and Holtzclaw. General Chemistry, Second Edition, D. C. Heath and Company, Boston, Massachusetts, p. 192.

Sienko and Plane. Chemistry, Third Edition, McGraw-Hill Book Company, New York, New York, chapter 10.

Purpose

The purpose of this experiment is to investigate the process of titration and carry out an analysis using titration.

Necessary materials: One 50 ml burrette, commercial vinegar, sodium hydroxide pellets, laboratory scales, a standardized solution of HCl approximately .15M strength, phenolphthalein, and a stirring rod.

Procedure: Using a piece of slick paper on the balance pan, weigh out about 3 grams of sodium hydroxide pellets and dissolve them in 200 mls of distilled water. Now using your burrette, accurately measure 40 mls of the standardized acid into each of the two 250 ml beakers and add four drops of phenolphthalein solution to the acid. Now discard the remaining standardized acid solution from the burrette and rinse it out with tap water and then with a little distilled water. Why is this necessary?

Fill your burrette with the sodium hydroxide solution and slowly add the hydroxide solution to each of the 40 ml quantities of standardized acid to the phenolphthalein endpoint. Record the quantity of sodium hydroxide solution used for each acid sample, calculate the molarity of the sodium hydroxide solution from the two determinations and record the molarity.

You are now ready to do an analysis. Weigh a 100 ml beaker and record its weight. Measure carefully 5 ml of vinegar into the beaker, weigh again and record the weight of the beaker and vinegar. Add about four drops of phenolphthalein and using the sodium hydroxide solution titrate to the phenolphthalein endpoint. What do you find the molarity of the vinegar to be? What is the percentage by weight of acetic acid in the vinegar? Why could you use an acetic acid solution for food preservation but not a hydrochloric acid solution of the same molarity? What is the reason for the basic difference between acetic and hydrochloric acid?

Problems:

(1) What is the percentage of a 10 gram sample of propanoic acid (CH_3H_2COOH) that requires 30.6 mls of 4M NaOH for neutralization?

(2) How many mls of a 1.5M NaOH solution would be required to neutralize 75 grams of 4% acetic acid solution?

SOLUBILITY OF A SALT

References: Sienko and Plane. _Chemistry_, Third Edition, McGraw-Hill Book Company, New York, New York, pp. 221-6.

Quagliano, James V. _Chemistry_, Second Edition, Prentice-Hall Inc., Englewood Cliffs, New Jersey, p. 31.

Purpose

To learn how the solubility of a salt can be determined by experiment.

Introduction

The solubility of a salt is usually determined by evaporating a known weight of a saturated solution and weighing the salt that remains. A numerical ratio of the weight of salt is usually expressed per one hundred grams of water.

Necessary materials: Thermometer, potassium chloride and the usual student desk equipment.

Procedure: Place about 25 mls of distilled water and about 8 grams of potassium chloride in a 250 ml erlenmeyer flask. Place a clean stopper in the flask and shake it for 8 to 10 minutes. At this time check to see if any undissolved salt remains in the flask. If there is no salt, more must be added and the contents shaken for a few more minutes. Now record the temperature of the solution. Weight carefully an empty evaporating dish and filter the solution into the evaporating dish. Weigh the dish and contents carefully. Record this weight. Place the dish on a beaker of boiling water (the latter serving as a steam bath) and evaporate to dryness. Now carefully heat the evaporating dish placed on wire gauze supported by the iron ring. Don't heat too strongly. Let the evaporating dish cool and weigh. Record the weight. Now rinse out the evaporating dish and dry it. Now place about 250 mls of tap water in a 600 ml beaker and place about 25 ml of distilled water in the 250 ml erlenmeyer flask and add about 12 grams of potassium chloride to the flask. Heat the water in the beaker to about 60 degrees C. and try to maintain that temperature for about 8 minutes while the flask containing the salt solution is kept submerged in the warm water in the flask. During this time stir the solution with a stirring rod. Take the final temperature of the solution and again filter the solution into the previously dried and weighed evaporating dish. Weigh the dish and contents and record the weight. Again evaporate the solution to dryness and weigh the salt and dish again. Record this weight. Now determine the ratio of the weight of salt to weight of water at the two different temperatures. Should there be a different weight of salt dissolved in the water at the different temperatures? What factors influence the solubility of a salt in water? Determine the solubility of potassium chloride in terms of grams of salt per 100 grams of water at the two different temperatures. Why must there be undissolved salt in the solution at the time it is filtered?

Data and Calculations	Room Temp.	Higher Temp.
a. Temperature of saturated solution	$^{o}C.$	$^{o}C.$
b. Weight of dish and solution	gm	gm
c. Weight of dish	gm	gm
d. Weight of solution (b-c)	gm	gm
e. Weight of dish and dry salt	gm	gm
f. Weight of KCl in the solution (e-c)	gm	gm
g. Weight of water in the solution	gm	gm
h. Grams KCl/gm H_2O		
i. Solubility of KCl in gm/100 gm H_2O		

CHAPTER IV

DATA GATHERED AND INTERPRETATIONS

At the end of each of the three semesters during which this study was conducted a final laboratory examination was given to those students comprising the control and experimental groups. These three semesters were the Fall of 1966, the Fall of 1967 and the Spring of 1968. The final laboratory examination was taken by four experimental groups and three control groups during those three semesters. During the Fall of 1966 and the Spring of 1968 there was one control and one experimental group each time. During the Fall of 1967 there were two experimental groups and one control group.

Tables 1, 2 and 3 of this chapter present the t-values resulting from statistical comparisons of final laboratory examination scores, ACT Natural Science scores and ACT Composite scores. The mean final examination score for the experimental group for the Fall of 1966 was 25.90 and that for the control group was 20.90 (see Figure 1). The mean ACT Natural Science score for the experimental group was 74.73 and for the control 67.29. The mean ACT Composite

Table 1.--T-test values of findings for fall semester 1966

Year	1966
t-test of means of final examination scores	3.29*
t-test of means of ACT Natural Science scores	1.57
t-test of means of ACT Composite scores	1.05
Mean final examination score of experimental group	25.90
Mean final examination score of control group	20.90

* .05 t 50 to 60 = 2.01 to 2.00

Table 2.--One-way analysis of variance test for achievement on final laboratory examination

Fall Semester 1967

	Control	Experimental Group 1	Experimental Group 2
Mean Scores	17.06	20.48	21.00

	Examination Scores	ACT Natural Science Scores	ACT Composite Scores
F-ratios	7.65*	.173	.100

TEST FOR SIGNIFICANT GAPS

	Experimental Group 1 and Control Group	Experimental Group 2 and Control Group	Experimental Group 1 and 2
t-values	3.15**	3.60**	.42

* $_{.05}F_{2,79} = 3.11$

** $_{.05}t_{79} = 1.99$

Table 3.--T-test values of findings for spring semester 1968

Year	1968
t-test of means of final examination scores	2.03*
t-test of means of ACT Natural Science scores	.725
t-test of means of ACT Composite scores	.555
Mean final examination score of experimental group	21.58
Mean final examination score of control group	18.44

* $.05\ t_{50\ \text{to}\ 60} = 2.01$ to 2.00

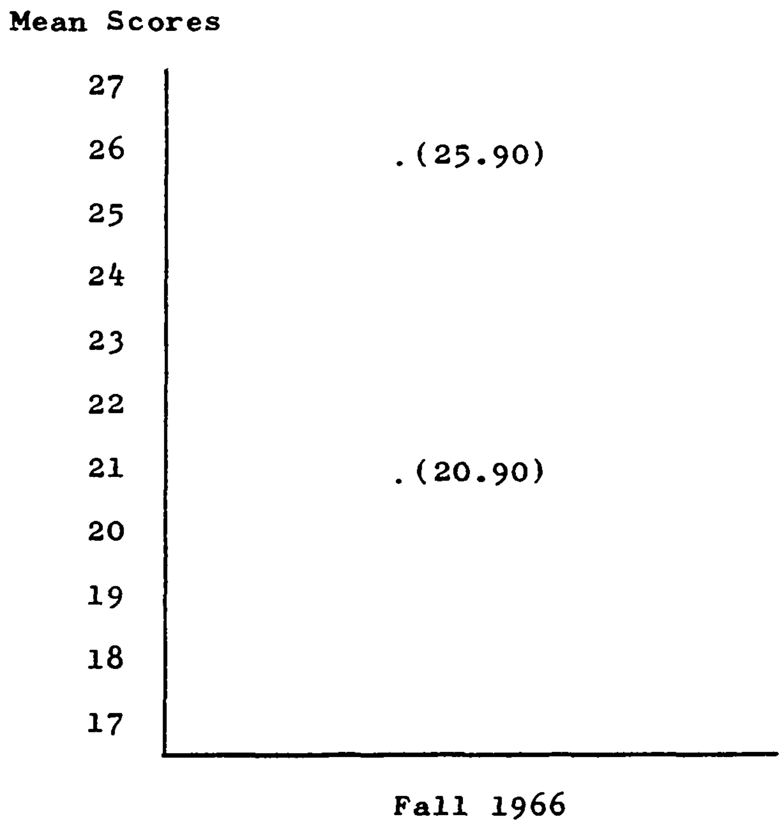

Fig. 1.--Achievement on laboratory final examination

score for the experimental group was 75.21 and for the control group 70.22 (see Appendix C and D). A statistical t-test was used to compare these mean scores. It was found that the t-value for the final laboratory examination mean scores was equal to 3.29; the .05 level of significance was used to minimize the probability of committing a Type II statistical error.[1] The t-value for the ACT Natural Science mean scores was equal to 1.57 and the t-value for the ACT Composite mean scores was equal to 1.05.

The findings just presented can be interpreted to mean that there are no significant differences between the ACT Composite scores and the ACT Natural Science scores for control and experimental groups for the Fall of 1966. On the basis of the ACT score comparisons the groups are similar. There was a significant difference between the two groups on the basis of their mean final laboratory examination scores. This could be interpreted to mean that the experimental group demonstrated a better mastery of subject matter than did the control group as measured by the questions used on the final laboratory examination.

For the Fall semester of 1967 there were two experimental groups and one control group that took the same final laboratory examination as that taken by the two groups during

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[1]Henry E. Garrett, Statistics in Psychology and Education, Longmans, Green and Company, New York, New York, 1961, p. 219.

the Fall of 1966. Since there were three independent groups this time it was necessary to use a one way analysis of variance rather than a t-test to search for any existing statistically significant differences. The findings are presented in Table 2 and Figure 2.

The mean final examination scores for the three groups for the Fall of 1967 semester were 17.06 for the control group, 20.48 for the experimental group 1 and 21.00 for the experimental group 2. The ACT Natural Science mean score for the control group was 68.93, for the experimental group 1 it was 72.12 and for the experimental group 2 it was 69.00. The ACT Composite mean score for the experimental group 1 was 71.16, for the experimental group 2 it was 69.08 and for the control group 68.81 (see Appendix E, F and G).

The F-ratios obtained from the one way analysis of variance were 7.65 for the final laboratory examination scores, .173 for the ACT Natural Science scores and .100 for the ACT Composite scores (see Table 2). At the .05 level of significance, there is no significant difference among the scores of the three independent groups on the basis of ACT Composite scores nor was there any significant difference among the three independent groups on the basis of the ACT Natural Science scores. The ACT-scores comparisons show that the groups were similar.

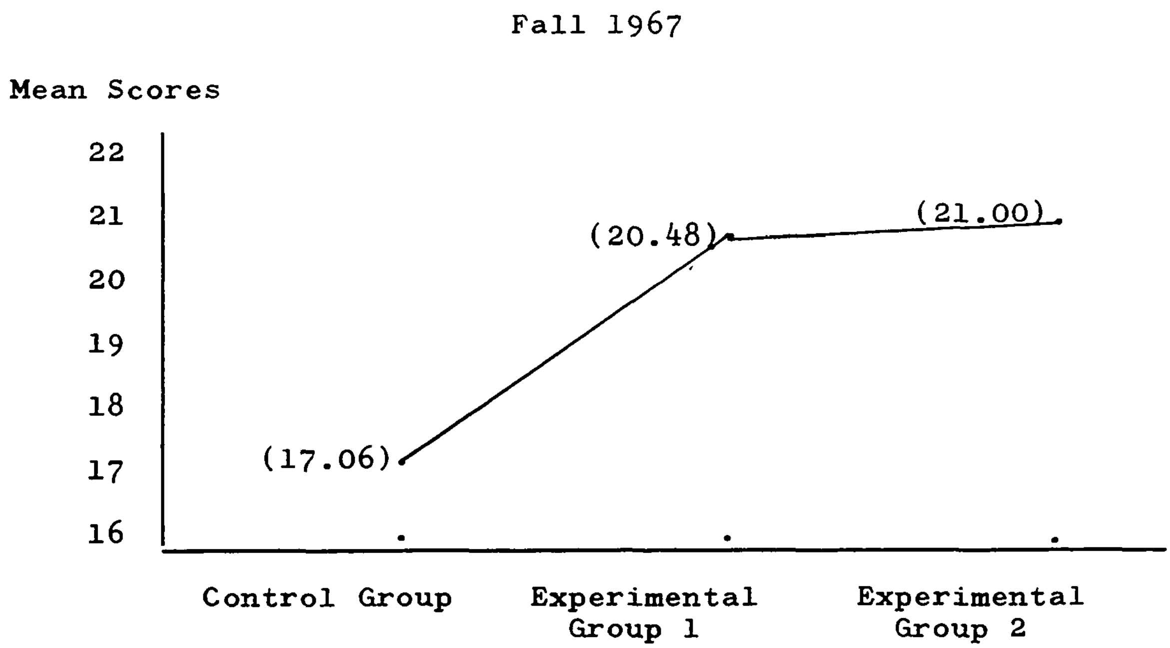

Fig. 2.--Mean scores for final laboratory examination

The F-ratio value of 7.65 shows that a significant difference exists among the three groups on the basis of mean final laboratory examination scores. Since the F-value is significant it was necessary then to isolate and determine where the difference lies. This was done by the use of the t-test. The t-test for the control and experimental group 1 yielded a result of 3.15, for the control group and experimental group 2 the t-value was 3.60 and for the experimental group 1 and the experimental group 2 the t-value was 0.42. These findings show that the achievement of the two experimental groups on the final laboratory examination was not significantly different but that both of the experimental groups made a significantly better achievement on the final laboratory examination than did the control group.

The same final laboratory examination was used during the Spring of 1968 with one control group and one experimental group. The mean final laboratory examination scores for the two groups were 18.44 for the control group and 21.58 for the experimental group (see Figure 3). The mean ACT Natural Science score for the control group was 65.04 and for the experimental group it was 68.77. The mean ACT Composite score for the control group was 66.12 and for the experimental group it was 68.97 (see Appendix H and I).

The t-values obtained for the two groups in the 1968 Spring semester were 2.03 for the means of final laboratory examination scores, and .725 for the means of ACT

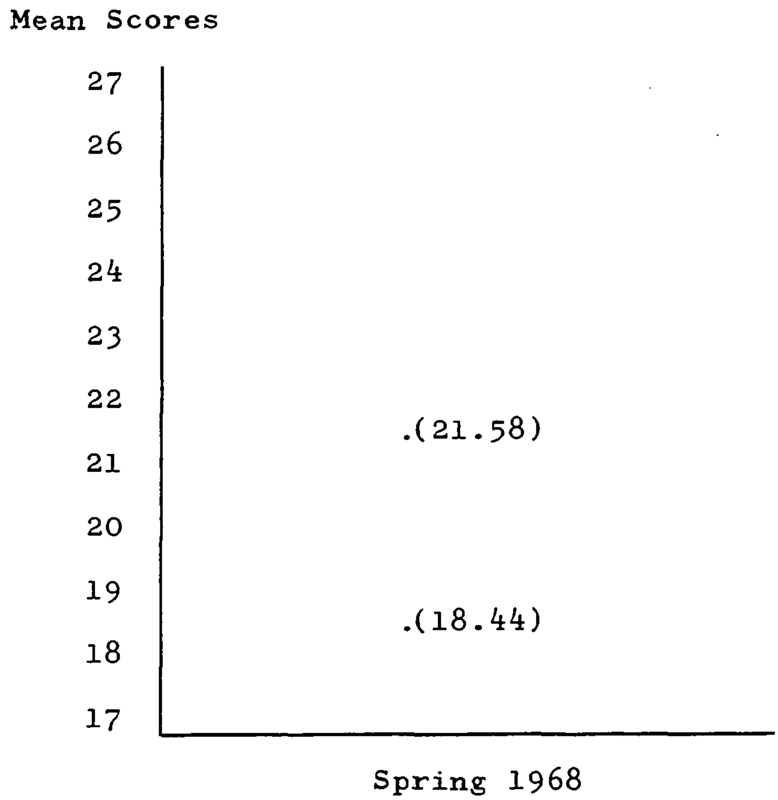

Fig. 3.--Achievement on laboratory final examination

Natural Science scores and .555 for means of ACT Composite scores (see Table 3). These data show that there was not a statistically significant difference between the two groups on the basis of ACT Composite scores nor on the basis of ACT Natural Science scores. Here again, the comparison of the ACT scores shows the groups to be similar. There was, however, a significant difference between the two groups on the basis of their mean final laboratory examination scores in favor of the experimental group.

In each of the three semesters the experimental group made a statistically significantly better score than the control group for that semester on the basis of the mean final laboratory examination scores. At the same time within each semester there was not a statistically significant difference between experimental and control groups on the basis of ACT Natural Science scores nor on the basis of ACT Composite scores.

At the end of the Fall semester of 1967 the experimental and control students took a thirty-one question test. These questions were then used during the Spring Semester of 1968 in a pre- and post-test situation. The results of the test for the students enrolled during the Fall semester of 1967 students were broken down into eight smaller tests to correspond to individual experiments and used as pre- and post-tests during the Spring semester of 1968.

The means of the scores for the control and experimental groups and the t-values obtained from t-tests for each of the eight tests are recorded in tabular form in Table 4 of this chapter. In all eight tests the mean scores of the experimental group exceeded that of the control group. Also in six of the eight tests the experimental group scored statistically significantly better than the control group at the .05 level of significance. Table 4 also shows that the experimental and control groups were not statistically significantly different on the basis of ACT Natural Science scores nor on the basis of ACT Composite scores.

In the Spring semester of 1968 there was one control group and one experimental group. The results of their achievement on the laboratory final examination have been presented (see Appendix H and I). The performance of the control and experimental groups was compared in a different way than the control and experimental groups of any other semester. Since the control and experimental groups took both pre- and post-tests during the Spring semester, the mean gain for each of the groups on the eight pre- and post-tests during the Spring semester, the mean gain for each of the groups on the eight pre- and post-tests was computed; also the t-value was obtained for each of the eight pre- and post-test mean gains.

The data in Table 5 show that at the .05 level of significance, the experimental group experienced a

Table 4.--Mean scores for examination given at the end of the 1967 fall semester

Test Number	1	2	3	4	5	6	7	8
Mean for Control Group	1.12	2.33	1.00	1.52	1.75	0.94	1.52	1.48
Mean for Experimental Group	1.69	2.61	2.42	2.23	2.12	1.12	2.23	1.73
t-test for examination scores	4.45*	1.47	10.60*	5.88*	2.58*	1.81	4.93*	2.02*
t-test for ACT Natural Science scores	0.20							
t-test for ACT Composite scores	1.21							

* .05 t 50 to 60 degrees of freedom = 2.01 to 2.00

Table 5.--Data obtained from same questions as for Table 3 but used as post- and pre-test questions during spring 1968

Pre- and Post-test number	1	2	3	4	5	6	7	8
Mean gain by Control group	0.74	0.48	1.60	0.39	0.54	1.63	1.52	1.48
Mean gain by Experimental group	1.68	1.34	2.46	1.08	0.97	2.53	2.03	1.77
t-test value for gain	7.76*	9.47*	5.19*	8.15*	5.48*	5.19*	2.56*	3.50*

* .05 t 50 to 60 degrees of freedom = 2.01 to 2.00

statistically significantly better gain on the post-test over the pre-test than did the control group in all eight cases.

During the Fall semester of 1967 the students took all of the pre- and post-test questions as one test and when the results were broken down to correspond to individual experiments the experimental group had a mean score that was statistically significantly better than that of the control group in all except numbers 2 and 6 of the eight pre- and post-tests. In the Spring semester of 1968 the experimental group experienced a statistically significantly better gain based on comparative mean gains than did the control group in all eight of the pre- and post-tests. These findings are shown in graphic form in Figures 4 and 5.

Two weeks before the end of the laboratory schedule for the Spring semester of 1968 the experimental and control groups were both presented with the problem of experimentally finding the solubility of potassium chloride in water. The students of both groups completely developed their own procedures and carried out the determination.

The next laboratory period the students of both groups turned in their reports for grading. They were graded on a 100 point basis as explained in chapter three. The mean score for the control group was 81.8 and the mean score for the experimental group was 84.3. Although the mean score for the experimental group was higher it was not

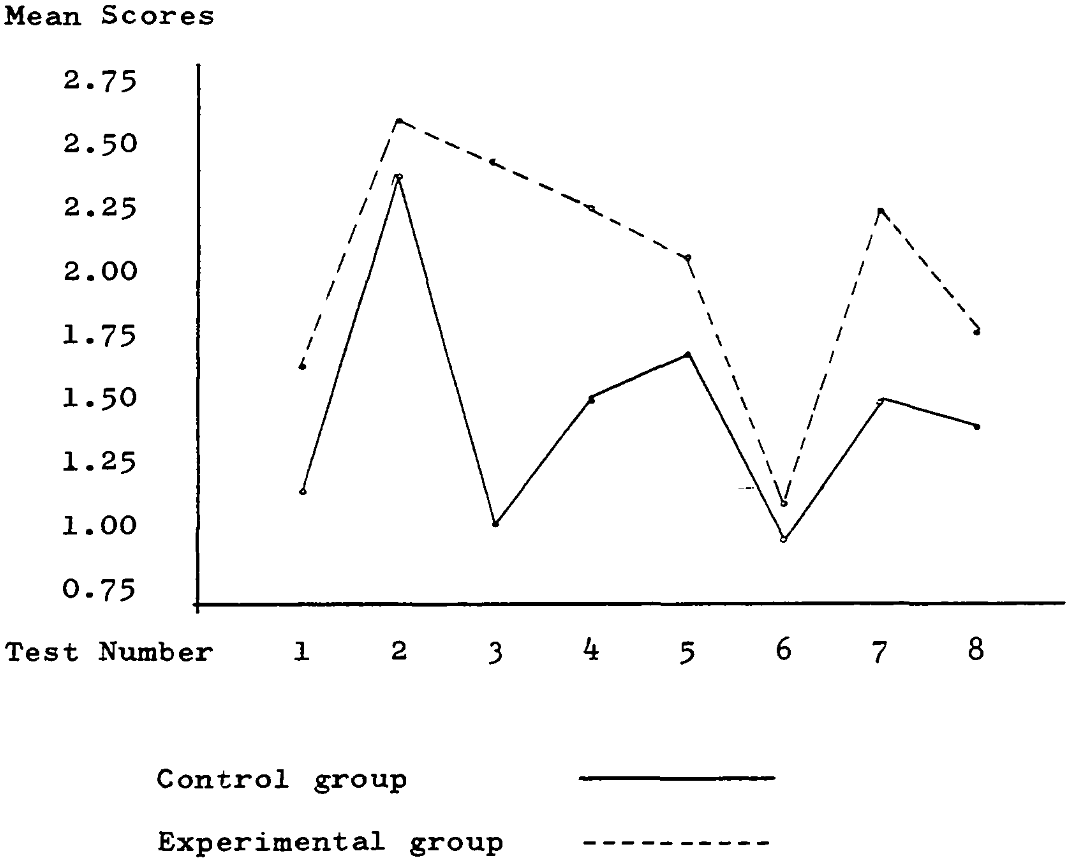

Fig. 4.--Mean scores for examination given at the end of the 1967 fall semester

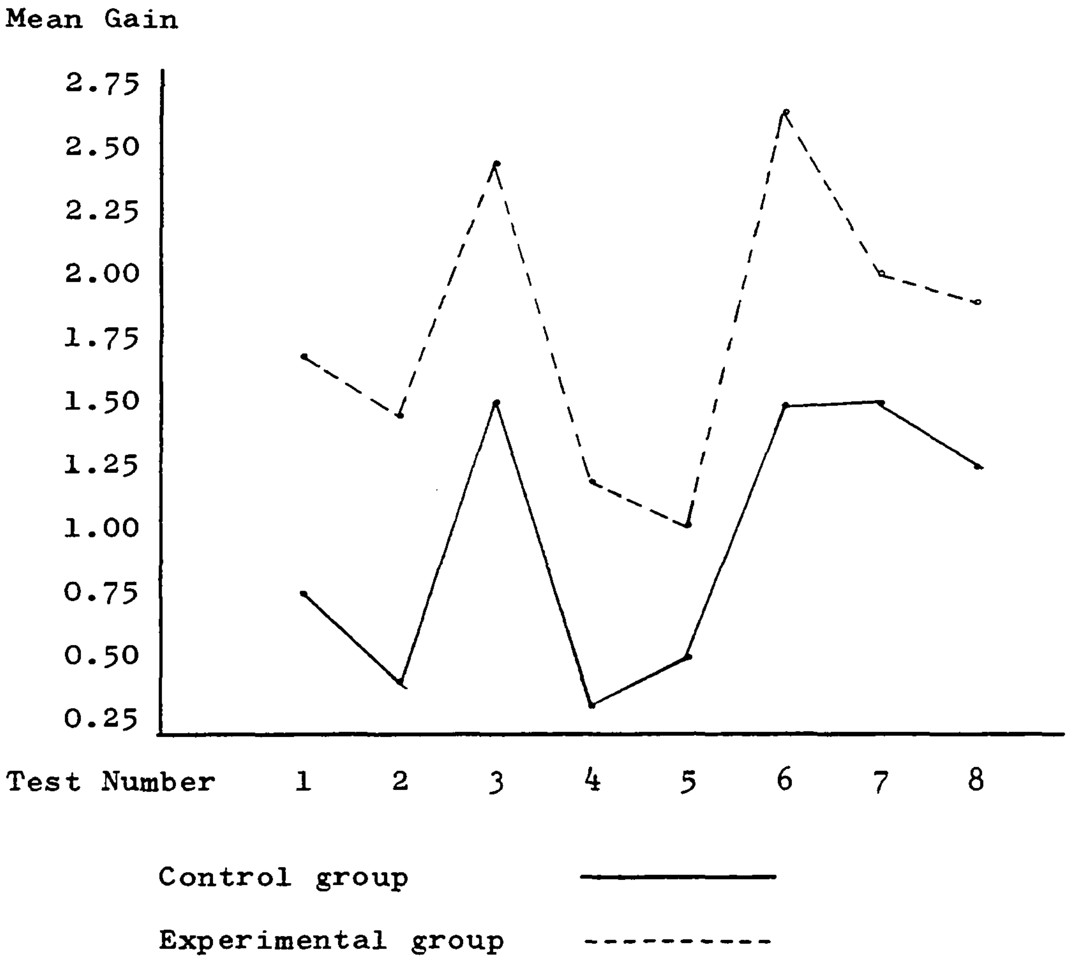

Fig. 5.--Mean gains on pre- and post-tests spring semester 1968

statistically significant at the .05 level (the t-value was equal to .434). These data might be interpreted to mean that there may be some laboratory procedures which do not lend themselves to the inquiry approach anymore than the conventional method.

One week from the end of the laboratory schedule for the Spring of 1968, the experimental and control groups were both presented with a solid and a liquid and were instructed to record as many facts about the two substances as they could based on laboratory observations. The mean number of correctly observed and recorded facts for the experimental group for the last mentioned experiment was 9.54 and that for the control group was 7.00. The t-value obtained from testing the two mean scores had a value of 4.31. This is a significantly better score in favor of the experimental group. The interpretation here might be that those students accustomed to using inquiry methods in the laboratory are more skillful in making observations than are those students that are accustomed to using the conventional laboratory method.

Figure 4 illustrates that the mean score for the experimental group exceeds the mean score for the control group on all of the pre- and post-test questions for the students making up the Fall semester of 1967. The scores on pre- and post-tests number two and six however, are not significantly different. Examination of Figure 5 reveals that the curves that result from plotting mean gains on post-tests for the experimental and control groups are very similar in shape and that the experimental group for the Spring semester of 1968 has mean scores that exceed the mean scores of the control group. In all eight cases the difference is statistically significant.

Considering the eight pre- and post-tests for two semesters (1967 and 1968), the two student designed experiments and the three semesters of comparative final laboratory examination scores, the experimental group made a statistically better achievement than the control group in eighteen of twenty-one instances. The interpretation to be made here is that there is strong evidence to show that on the basis of the manner in which both control and experimental groups were evaluated, the experimental group achieved significantly better than the control group.

CHAPTER V

CONCLUSIONS

The experimental groups made a significantly better mean score on the final laboratory examination than did the control groups by semesters for each of the three semesters involved in the study. The investigator concludes that the students in the experimental group had experiences that better equipped them with the ability to find meaningful relationships with data supplied them in hypothetical situations. The investigator also concludes that the experimental students had experiences that enhanced their reasoning ability more than did the experiences of the control students since many of the questions to which both groups responded required the student to interpret the data available to him in the light of the basic principles of chemistry.

When the pre- and post-test questions were given as one larger test at the end of the semester for the Fall semester of 1967, the experimental group made a significantly better score than did the control group in six out of eight of the pre- and post-tests. In view of the fact that the experimental group for the Fall semester of 1968 scored significantly better than the control group in terms of mean gains of the post-test over pre-test, the investigator concludes that the use of the questions as a final examination was not as valid a measure as were the pre- and post-test situations. Also when the comparison was made for all eight of the pre- and post-test mean gains treated as one unit, the mean gain for the experimental group was 1.73 and for the control group, 1.05. The value for the experimental group by this comparison was still significantly better statistically than the control group.

The findings of this study also show that on the basis of the gain as measured by the use of pre- and post-tests for eight experiments in general college chemistry an experimental group using laboratory experiments based upon the principle of inquiry made in all eight cases significantly better gains than did control groups who were using conventional laboratory experiments.

The value of the t-test (7.76, see Table 5) obtained by a comparison of the mean gains of the experimental and control groups for the experiment on density is significant. Since the same materials were used by both the control and experimental groups for the density experiment the possible conclusion might be that those students using the form written by the investigator gained a better understanding of the relationship between units of mass and units of volume than did the control group which used the conventional laboratory manual. There is also a possibility that the experimental group also gained a better understanding of the Archimedes principle and its application to experimental determination of density for solid objects.

The value of the t-test (9.47, see Table 5) obtained by a comparison of the mean gains of the experimental and control groups for the experiment on the properties of substances is a significant value. Both experiments (the experimental and control) dealt primarily with chemical and physical changes. The findings indicate that the experimental group possibly learned to make a better distinction between chemical and physical changes than did the control group.

The third matched experiment dealt with the preparation and properties of oxygen. The value of the t-test obtained from a comparison of mean gains of the control and experimental groups had a value of 5.19 (see Table 5). The same method was used in both cases for the preparation of oxygen gas but there seems to be a possibility that the experimental group may have gained a better understanding of the nature of the oxides produced by burning various elements in oxygen gas than did the control group.

The fourth experiment of the eight matched experiments dealt with reactions of acids and bases. The value of the t-test obtained from a comparison of mean gains of the control and experimental groups was 8.15 (see Table 5).

This was the second highest t-value of the eight matched experiments. The findings would seem to offer evidence that the experimental group may have gained a better understanding than the control group in relation to the main distinction between acids and bases, quantitative calculations of acids and bases and identifying characteristics of weak acids and strong acids.

The fifth matched experiment dealt with the experimental determination of the gram-molar volume of a gas in both the experimental and control groups. As stated in chapter three the control group obtained oxygen as a gas by thermal decomposition of a chemical solid while the experimental group obtained a gas by the vaporization of a volatile liquid. The findings indicate (t-value of 5.48, see Table 5) that the experimental group may have gained a better understanding of how to use experimentally-obtained data to determine the molecular weight of a gaseous substance.

The sixth experiment dealt with the experimental determination of water of hydration of crystallized substances. The same chemicals were used by both the experimental and control groups but the experimental group used the inquiry method while the control used the conventional procedure provided by the department adopted manual. The t-value (5.19, see Table 5) offers evidence that the experimental group may have gained a better understanding of how to calculate the number of moles of water of hydration in a salt using experimentally-obtained data. At the conclusion of the Fall semester of 1967 when the same pre- and post-test questions were used as one larger examination the experimental group scored better than the control group on the questions pertaining to this particular experiment but not statistically significantly better.

The seventh experiment had as an objective that of helping the student to develop a relative activity list of some elements from experimental observations. The t-value of 2.56 (see Table 5) was statistically significant. There is a possibility that the experiment developed by the investigator is more effective in helping the student to see the relative activity of some of the common elements by means of the reactions they carried out in the experiment than is the experiment provided by the conventional manual. The procedure written by the investigator makes use of a minimum but sufficient number of chemical reactions to demonstrate the relative chemical activity of several metals.

The eighth matched experiment dealt with the topic of equivalent weight. The t-value obtained from the comparison of mean gains for the experimental and control groups was 3.50 (see Table 5). This evidence indicates that there is a possibility that the experimental group learned better than the control group how to determine the equivalent weight of an element as shown by their reaction to equivalent weights determined from formulas, equations of reactions and experimentally obtained data. All the t-values obtained by comparisons of experimental and control groups on the basis of mean gains of post-tests over pre-tests are listed in Tables 4 and 5.

In the first of the two student designed experiments, that of experimentally determining the solubility of potassium chloride, the experimental group made a better mean score than the control (84.3 and 81.8) but there was not a statistically significant difference in these scores. The conclusions might be drawn here that the experimental group had no significantly better ability to design the experimental procedure and obtain the experimental data than did the control group. Perhaps more investigation needs to be done to determine if some types of experiments are more conducive to inquiry methods than others.

In the second of the two student-designed experiments carried out by both the control and the experimental groups, the mean score for the experimental group was 9.54 and that for the control group was 7.00. These values were obtained from the actual correctly observed and recorded observations by the students during the laboratory period. The value of the t-test for these scores was 4.32. That t-value is statistically significant. This offers evidence that possibly the experimental group developed keener powers of observation as a result of their experiences of using the inquiry-centered approach. There is a possibility also that the experimental group was more careful to record their data since they may have become more aware than the control group of the importance of recorded data in the area of scientific investigation.

In all three semesters involved in this study the experimental group scored statistically significantly better on the laboratory final examination than did the control group. This evidence allows the conclusion to be drawn that the experimental group had experienced a greater increase in knowledge of the subject matter. The conclusion can be drawn that the experimental group had experiences that resulted in a better development of the rational powers than did the control group. The argument to support this contention is that both groups took the same pre- and post-tests and yet the experimental group experienced a greater mean gain of post-test scores over pre-test scores than did the control group in all eight cases.

The ability to answer the questions and work the problems on the post-tests depended both on the student's observations during the experimentation and his power of reasoning stimulated by the questions asked in the experiment itself. In cases where data was supplied for a hypothetical situation, the correct answer required the use of the student's rational powers to relate, that is, analyze, compare, classify, synthesize and generalize the data correctly. In some other cases the correct solution depended upon the student's ability to determine the properties of a substance based upon his laboratory experiences. If the student had obtained information from previous experiences that was useful to him in subsequent situations then this could have been helpful in developing his rational powers. If he could recognize and apply pertinent observations toward the development of an important principle, this also could be attributed to reasoning. The investigator concludes that the students in the experimental groups had laboratory experiences that were more meaningful to them than the conventional laboratory experiences were to the students in the control groups. As a result of these experiences the students in the experimental groups apparently were better able to think critically about a problem, recall previous experiences and relate experimentally obtained data in a productive manner. The investigator concludes that those students in the experimental group may have been more observant than the control group and may have better realized how to interpret the data they collected than did the control group as a result of having used the inquiry-centered experiment.

The statistical treatment of the data allows the rejection of the Hypothesis of no difference at the 0.05 level of significance in achievement between the control and experimental groups on the laboratory final examination for all three semesters. Also the hypothesis of no difference between the control and the experimental group on the basis of mean gains of post-tests over pre-tests can be rejected for experiment three.

On the basis of the interpretation of the data for the first experiment in which there was no control over lecture instructor, laboratory assistant, nor laboratory instructor for the control group and no control over the lecture instructor or laboratory assistant for the experimental group, the experimental group performed better on the final laboratory examination than the control group.

For the second experiment the statistical treatment of the data allows the conclusion to be drawn that even in a situation in which there were three independent and homogeneous groups, the two experimental groups performed significantly better than the control group and yet there was no statistical difference between the achievement of the two experimental groups. This offers further evidence of the effectiveness of the inquiry method as an aid to learning.

Concrete evidence of the superiority of the inquiry method of laboratory instruction over that of the conventional method is offered by the interpretation of the data collected during the third experiment. In a situation in which both the control and experimental groups responded to eight identical pre- and post-tests, the experimental subjects achieved a statistically significantly better mean gain than the control subjects on all eight pre- and post-tests. In addition to this the achievement of the experimental subjects on the same final laboratory examination as used in experiment one and two was significantly better than that of the control subjects.

The findings and their statistical interpretation justifies the conclusion that in a variety of actual learning situations regardless of the control of the inherent variables, the inquiry method of learning as opposed to the conventional laboratory procedure resulted in significantly better performance, on the measures used, by the students using the inquiry method.

The evidence also seems to indicate that the experimental group developed a better understanding of the principles of chemistry as brought out in the experiments used than did the control group. The investigator concludes that there is sufficient evidence to indicate that possibly the inquiry method of instruction is more effective in helping the student to learn the subject matter in chemistry laboratory than is the conventional laboratory method.

Twenty-four students of the experimental group of the Fall semester of 1966 were interviewed by a science educator not known on the Central State campus and all stated that they felt that the laboratory experiments had provided them with worthwhile learning experiences. A majority of them indicated that they would like to have another course in which the laboratory experiments used the inquiry approach. A few had been in a chemistry course previously that used the conventional laboratory experiments and all of these stated that they preferred the inquiry-centered experiments to the conventional ones. The actual interviews are recorded in the Appendix.

The investigator concludes that the inquiry-method of instruction in beginning chemistry laboratory is a successful method and one that appeals to the student's imagination and stimulates his thinking. If this does indeed stimulate the student's thinking it may be that this will be of use to him beyond the academic limits of chemistry.

CHAPTER VI

SUMMARY AND RECOMMENDATIONS

This study had its origin during the year 1966. The investigator became interested in the use of inquiry in instruction and with the supervision of his chairman developed a series of experiments to be used in beginning college chemistry laboratory.

These experiments were used by four different and independent groups of students over a period of two academic years involving three semesters. Each time that a group of students used these experiments there was another group, the control group, that used the experiments provided in a commercially available laboratory manual. Both groups of students conducted experiments that dealt with the same principles of chemistry.

Each time that the experiments were used the details relating to the administration of the mechanics of laboratory operation were refined in the interest of making the best possible use of available time. The experiments developed by the investigator, however, were not altered during the course of the study. Finally, during the third semester that the experiments were used (Spring semester 1968) the only variable that remained was the inquiry method being used for the experimental group and the conventional approach for the control group. During this semester both groups had the same textbook, lecture instructor, laboratory instructor and laboratory assistant.

Two hundred two students were involved in the study over the three-semester period. A greater number than this started at the beginning but some dropped out during each of the semesters and thus did not take the final examination nor complete the series of experiments.

Eleven experiments were developed by the investigator and used during the three semesters that the study was conducted. Eight of these experiments dealt with topics identical to eight of those in the commercial manual. The pre- and post-test questions were developed for these eight experiments. During the first two semesters the control and experimental groups were compared statistically on the basis of their ACT Natural Science scores, ACT Composite scores and their scores on a laboratory final examination developed by the investigator. The findings indicated that there was no statistically significant difference between the experimental and control groups on the basis of ACT Natural Science scores and ACT Composite scores. These data show that the two groups were academically similar and the results obtained, therefore, cannot be attributed to the intellectual superiority of the experimental group.

During the Spring semester 1968 one control and one experimental group participated in the study. Each group took a series of pre- and post-test questions developed to correspond with eight of the eleven experiments. These were given in addition to the final laboratory examination that the other two previous semester's students had taken. The experimental and control students for this semester were compared statistically on the basis of their mean gains on the performance of post-test compared to pre-test, scores on the final laboratory examination, ACT Natural Science scores and the ACT Composite scores.

The investigator found that for each of the three semesters the experimental groups scores statistically significantly better than did the control groups on the final laboratory examination. During the Fall semester 1967 the experimental group did not score statistically significantly better than the control group for the pre- and post-tests number 2 and number 6. During the Spring semester 1968 the experimental group failed to score statistically significantly better than the control group on the first of two student designed experiments in which the students experimentally determined the solubility of potassium chloride in water. The experimental group for the Spring semester 1968 experienced a significantly better gain as determined by the pre- and post-tests. The final analysis of this is that the experimental group scored statistically better than the control group in eighteen of twenty-one measures.

Each semester the students were free to enroll in any laboratory section of their choice. The investigator then randomly selected the sections that would be the control and experimental groups. For each semester the intellectual levels of the experimental and control groups were not statistically significantly different as judged by their ACT Composite scores and their ACT Natural Science scores.

Concerning the taped interviews conducted by a science educator not known on the campus of Central State College, the investigator found that:

1. All twenty-four students from the experimental group of the Fall semester of 1966 that were interviewed stated that the laboratory experiments provided them with worthwhile learning experiences.

2. Sixty-six per cent of them would prefer to use the inquiry approach in the laboratory for any subsequent chemistry course that they might take.

3. Three stated that they felt that they had gained knowledge of learning how to learn, as a result of this laboratory experience.

4. Seven students stated that they felt that there was a prevailing atmosphere in the laboratory that was conducive to learning.

The following general evaluations were made by the investigator during the course of this study.

1. For each of the three semesters the control and experimental groups took the same final laboratory examination. These scores were subjected to statistical analysis and it was found that the experimental group scored statistically significantly better than the control group for each of the three semesters.

2. The students for the Fall semester of 1967 took a test composed of the pre- and post-test questions used for eight different experiments during the Spring semester of 1968. On the basis of these eight pre- and post-tests, the experimental group scored statistically significantly better than did the control group for six of the eight.

3. The students for the Spring semester of 1968 took the eight different pre- and post-tests and the mean gains for the experimental and control groups were statistically compared. For this semester the experimental group scored statistically better than the control group in all eight cases. On the basis of two student-designed experiments the experimental group made a higher mean score than the control group on both experiments. On only one of the two, however, did the experimental group score statistically better than the control group.

4. A tape recording was made of the reactions of twenty-four students from the experimental group for the

Fall semester of 1966. The purpose of this was to attempt to determine the effectiveness, in the opinion of the students, of the experiments developed by the investigator.

In view of the statistical data gathered in the course of this study the investigator concluded that the inquiry method of instruction is more effective in helping the student to learn the essential principles of chemistry in the beginning chemistry laboratory than the conventional method.

More research needs to be done in the area of beginning college chemistry laboratory since the investigator found that in three of twenty-one instances, the achievement of the experimental group was not statistically significantly better than the control group. Research should be done to try to determine if some experiments lend themselves more to the inquiry method than others since the investigator found that for one student-designed experiment the experimental group did not score statistically significantly better than the control group. Also, the investigator developed a minimum of experiments for only one semester. Another set of inquiry-centered experiments should be developed for the second semester of beginning college chemistry laboratory to determine their effectiveness in helping the student to learn the principles of chemistry. The investigator has also observed that many times students are not very enthusiastic about the laboratory in organic chemistry;

research should be done to develop and test inquiry-centered experiments for that learning experience.

The investigator also recommends that more extensive research be done using the inquiry method of instruction in other academic fields of study to determine if this method is effective in those areas in developing the student's ability to think.

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APPENDICES

APPENDIX A

FINAL LABORATORY EXAMINATION

I. Select the best answer.

____ 1. The density of a liquid is (a) volume divided by mass (b) always less than 1 gram/milliliter (c) mass divided by volume (d) volume divided by weight.

____ 2. If five milliliters of mercury (specific gravity 13.6) has the same mass as 14 milliliters of a second liquid, the density of the second liquid is approximately (a) 38.6 grams/milliliter (b) 4.72 grams/milliliter (c) 7.42 grams/milliliter (d) 2.74 grams/milliliter.

____ 3. The specific gravity of (a) benzene (b) alcohol (c) water (d) carbon tetrachloride is equal to one.

____ 4. In which of the following does a chemical reaction not occur? (a) melting paraffin and adding powdered sulfur (b) heating to redness a mixture of iron and sulfur powder (c) grinding in a mortar some iodine and mercury (d) adding water to aluminum sulfide.

____ 5. The following reaction: $Al_2S_3 + 3H_2O \rightarrow Al_2O_3 + H_2S$ is an example of (a) hydrolysis (b) a single replacement reaction (c) neutralization (d) a reaction that does not actually occur.

____ 6. The most volatile substances of the following is (a) sulfur (b) iodine (c) mercury (d) carbon disulfide (all at room temperature).

____ 7. An essentially insoluble substance is (a) $BaCl_2$ (b) PbI_2 (c) NaCl (d) KNO_3.

____ 8. A substance that sublimes easily is (a) iodine (b) hydrogen sulfate (c) silicon (d) sodium chloride.

____ 9. The best method for separating pure water from salt water is (a) filtration (b) precipitation (c) sublimation (d) distillation.

____ 10. In preparing magnesium oxide by heating the metal in air, an impurity that is most likely to form is (a) $MgCl_2$ (b) Mg_2N_3 (c) Mg_3N_2 (d) $Mg(NO_3)_2$.

____ 11. Heating an active metal in air would result in a chemical reaction known as (a) synthesis (b) analysis (c) neutralization (d) hydrolysis.

____ 12. The equivalent weight of magnesium as determined from the equation shown in number 20 is (a) 9 grams/equivalent (b) 24 grams/equivalent (c) 8 grams/equivalent (d) 12 grams/equivalent.

____ 13. A substance whose crystals contain water of hydration is (a) barium chloride (b) sodium chloride (c) potassium nitrate (d) sulfur.

____ 14. A compound that would be likely to effloresce would be (a) NaCl (b) $CaCl_2$ (c) $Na_2CO_3 \cdot 10H_2O$ (d) H_2SO_4.

____ 15. One could experimentally determine the water of hydration in (a) crystals of copper (II) sulfate (b) crystals of sodium chloride (c) crystals of $C_6H_{12}O_6$ (d) crystallized iodine, by heating enough to drive off all the water of hydration.

____ 16. The oxide of (a) magnesium (b) sodium (c) selenium (d) aluminum would form an acid in water solution.

____ 17. The method most usually used to produce oxygen in the laboratory is that of heating (a) $KClO_3$ and MnO_2 (b) $KClO_4$ (c) KNO_3 with MnO_2 (d) K_3PO_3 and MgO.

____ 18. An aqueous solution of (a) K_2O (b) BaO (c) P_2O_5 (d) CaO would be acidic.

____ 19. The solubility of most salts (a) decreases with an increase in temperature (b) increases with an increase in temperature (c) does not change with temperature changes.

____ 20. Based on the reactions represented by the following equations, the metal with an equivalent weight of 9 grams/equivalent is (a) Mg (b) Na **(c)** Ca (d) Al.

(a) $Mg + H_2SO_4 \longrightarrow MgSo_4 + H_2$

(b) $2Na + 2HCl \longrightarrow 2NaCl + H_2$

(c) $Ca + HCl \longrightarrow CaCl_2 + H_2$

(d) $2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$

II. Give an appropriate answer to the following questions.

1. What is the proper unit of pressure measurement for using the ideal gas law?
2. What volume would one equivalent of hydrogen gas occupy at standard temperature and pressure?
3. Suppose you increase the volume of a gas confined in a cylinder with frictionless piston at constant pressure, what else must also change?
4. What is the gram molecular volume for any gas at standard temperature and pressure?
5. What results when acids react with bases?
6. What would be the nature of a clear solution that turned red when phenolphthalein was added?
7. What would be common to all acid solutions in water?
8. What determines whether an acid is weak or strong?
9. What is the most active of the usually available laboratory metals?
10. Name a metal that would not dissolve in hydrochloric acid solution.
11. What would be a metal less active than hydrogen but more active than silver?
12. What is obtained by dividing the atomic weight of a metal by its equivalent weight?

13. What might have been the metal that in reaction with hydrochloric acid solution required .22 grams of the metal to produce 75 milliliters of hydrogen gas at standard temperature and pressure?

14. Household vinegar is approximately what per cent acetic acid?

15. What is the molarity of an acetic acid solution (CH_3COOH) if it contains 45 grams of acetic acid per liter of solution?

16. What method would be used to standardize an acid solution if a basic solution of predetermined strength was used?

17. What would be the percentage purity of an acetic acid solution (10 grams) if it required 25 milliliters of 4 molar NaOH solution for neutralization?

18. If an acid solution of H_2SO_4 is 2 molar, what would be its normality in a complete neutralization reaction?

19. What is the proper dimension for volume in the ideal gas law formula?

20. What is the normality of a base for which 20 milliliters require 55 milliliters of .1N acid for neutralization?

APPENDIX B

PRE- AND POST-TEST QUESTIONS USED DURING SPRING SEMESTER 1968

Test 1

1. Tell how you would experimentally determine the density of an odd shaped solid object.

2. The density of mercury is 13.6 grams/ml and that of water is 1 gram per milliliter. What volume of mercury would have the same weight as one kilogram of water?

3. What would be the best dimensions for expressing the density of a gas in scientific notation?

4. Twenty-five milliliters of carbon tetrachloride weigh forty grams. A solid object dropped in the carbon tetrachloride in a graduated cylinder caused a volume increase of 12 mls. How much less would the object weigh in this liquid if suspended by a thread than it would weigh in air?

Test 2

1. Magnesium when heated hot enough in air gives a bright display of light and heat. What kind of a change is this?

2. How would you defend your answer to number 1?

3. It is known that iron will corrode in the air at normal temperature. What kind of change is this?

4. How might you experimentally prove that your answer to number 3 is correct?

Test 3

1. A crystallized substance upon heating was observed to suffer a weight loss. What does this information tell you about its original composition?

2. Crystallized copper sulfate lost 1.8 grams of weight when heated from a 5 gram sample. What is the mole ratio of salt to water in the original substance?

3. What are some explanations for an experimentally determined formula of a hydrate being in error?

4. Crystallized copper sulfate hydrate is blue. Upon being heated it turns white. What causes this change?

Test 4

1. How do aqueous solutions resulting from dissolving an oxide of a non-metal differ from that of the solution of a metal oxide?

2. Oxygen can be prepared by heating $KClO_3$ in a test tube. What should be added to the $KClO_3$ to make the reaction more effective?

3. A substance contains 48% oxygen and gives up one-third of its oxygen when heated. What weight of oxygen could be obtained by heating 10 grams of the substance?

4. Sulfur is observed to burn brilliantly in pure oxygen. The resultant product is dissolved in distilled water. What is the nature of the solution as shown by an indicator?

Test 5

1. What is the basis for distinction between acids and bases?

2. What results from the neutralization of an acid with a base?

3. What accounts for the difference in strength of two acids?

4. Forty milliliters of .1M NaOH are mixed with 50 milliliters of 1M HCl. What is the molarity of the resultant solution as an acid?

Test 6

1. How could you experimentally determine the equivalent weight of a metal?

2. If a gram-atom of a metal in a chemical reaction caused the production of 22.4 liters of hydrogen at standard temperature and standard pressure, what is the ratio of equivalent weight to atomic weight of the metal?

3. The formula ratio of a metal oxide is M_2O_3. If half the atomic weight of oxygen is an equivalent weight, what fraction of the atomic weight of the metal represents its equivalent weight?

4. Based on the reaction represented by the following equation, what is the equivalent weight of calcium?

$$Ca + 2HCl \longrightarrow H_2 + CaCl_2$$

Test 7

1. Compare the reaction of sodium metal and cold water with that of magnesium and cold water.

2. Which of the following reactions will actually occur?

   (a) $Cu^{++} + Fe \longrightarrow Fe^{++} + Cu$

   (b) $Cu + Fe^{++} \longrightarrow Fe + Cu^{++}$

3. Which metal of the following shows no reaction with HCl solution? Magnesium, aluminum, zinc, or copper.

4. A piece of copper metal is placed in a solution of silver nitrate and silver metal is obtained. Copper metal is not obtained when a piece of silver is placed in a copper sulfate solution. What is the explanation for these observations?

Test 8

1. A certain liquid is more volatile than water. How does its vapor pressure compare with the vapor pressure of water at the same temperature?

2. You are going to vaporize a liquid by submerging a flask containing the liquid in boiling water. What would be the upper limit of boiling point allowable for the liquid being tested?

3. Data from an experiment shows that .8 gram of a vapor at 740 mm of mercury pressure and at temperature of 372°K (or absolute) occupied .25 liter volume. Using the equation of state, show how you could experimentally determine the molecular weight of the substance.

APPENDIX C

FINAL LABORATORY EXAMINATION

Fall 1966

EXPERIMENTAL GROUP

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	20	83	63
2	22	49	35
3	27	93	78
4	27	90	75
5	28	80	82
6	31	93	92
7	29	90	86
8	35	55	86
9	34	93	95
10	23	71	68
11	29	59	58
12	20	68	73
13	27	87	68
14	19	75	78
15	21	94	85
16	19	68	58
17	20	75	68
18	25	99	97
19	24	71	82
20	20	79	68
21	23	75	86
22	18	90	86
23	30	75	86
24	27	75	82
25	22	95	92
26	27	87	68
27	26	79	58
28	30	75	82
29	26	49	63
30	29	79	73
31	32	75	86
32	37	44	52
33	28	71	73
	Mean = 25.90	Mean = 74.73	Mean = 75.21

APPENDIX D

FINAL LABORATORY EXAMINATION

Fall 1966

CONTROL GROUP

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	24	39	58
2	27	93	97
3	26	99	98
4	14	93	89
5	17	60	65
6	15	93	89
7	24	99	98
8	17	87	89
9	25	93	86
10	22	1	4
11	30	39	78
12	23	55	58
13	23	90	92
14	24	90	92
15	28	22	35
16	16	71	78
17	15	71	73
18	27	88	70
19	15	79	70
20	23	93	95
21	27	95	86
22	23	28	28
23	16	68	66
24	22	55	52
25	12	59	68
26	20	59	78
27	20	26	24
28	20	93	78
29	20	26	58
30	17	59	63
31	16	55	68
	Mean = 20.90	Mean = 67.29	Mean = 70.22

APPENDIX E

FINAL LABORATORY EXAMINATION

Fall 1967

CONTROL GROUP

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	17	87	86
2	16	90	86
3	16	98	86
4	23	75	63
5	11	68	63
6	14	87	78
7	12	44	68
8	16	83	78
9	18	87	86
10	27	90	92
11	17	32	35
12	18	79	68
13	15	79	73
14	16	49	68
15	16	49	63
16	12	55	52
17	19	83	82
18	18	79	52
19	10	17	52
20	16	96	92
21	12	68	68
22	21	55	78
23	13	39	63
24	15	90	82
25	23	79	82
26	18	75	68
27	13	44	35
28	8	83	82
29	8	28	26
30	20	44	30
31	18	87	92
32	15	87	73
	Mean = 17.06	Mean = 68.93	Mean = 68.81

APPENDIX F

FINAL LABORATORY EXAMINATION

Fall 1967

EXPERIMENTAL GROUP 1

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	21	79	78
2	17	90	82
3	18	87	86
4	24	83	86
5	16	17	35
6	23	93	92
7	25	93	95
8	21	87	82
9	21	87	92
10	20	93	89
11	19	93	82
12	19	71	73
13	33	90	86
14	17	30	41
15	19	55	46
16	18	44	52
17	27	93	89
18	21	35	41
19	21	93	95
20	15	44	24
21	25	69	28
22	19	35	52
23	15	87	89
24	17	68	86
25	21	87	78
	Mean = 20.48	Mean = 72.12	Mean = 71.16

APPENDIX G

FINAL LABORATORY EXAMINATION

Fall 1967

EXPERIMENTAL GROUP 2

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	27	79	78
2	31	98	95
3	20	68	73
4	18	79	41
5	15	95	86
6	23	93	86
7	26	35	63
8	15	44	35
9	14	64	41
10	23	90	89
11	16	64	68
12	18	68	68
13	16	59	52
14	10	44	18
15	35	60	82
16	23	71	92
17	30	83	89
18	18	55	58
19	14	87	78
20	15	49	58
21	26	87	82
22	22	90	97
23	14	17	38
24	26	87	78
25	30	59	82
	Mean = 21.00	Mean = 69.99	Mean = 69.08

APPENDIX H

FINAL LABORATORY EXAMINATION

Spring 1968

EXPERIMENTAL GROUP

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	26	44	46
2	24	68	68
3	17	79	82
4	27	75	67
5	19	75	52
6	20	95	90
7	23	87	82
8	16	44	41
9	13	55	58
10	16	83	73
11	12	59	63
12	15	55	63
13	32	87	78
14	27	83	73
15	10	39	58
16	19	87	73
17	25	71	78
18	23	87	73
19	20	44	58
20	31	85	89
21	21	79	68
22	34	79	86
23	16	39	63
24	28	75	73
25	15	29	61
26	24	26	30
27	18	68	73
28	26	93	95
29	22	90	86
30	33	93	97
31	17	59	41
	Mean = 21.58	Mean = 68.77	Mean = 68.97

APPENDIX I

FINAL LABORATORY EXAMINATION

Spring 1968

CONTROL GROUP

Student	Raw Score	ACT Natural Science Score	ACT Composite Score
1	20	71	82
2	11	17	30
3	12	49	41
4	8	71	46
5	18	70	69
6	12	75	58
7	18	83	82
8	10	39	41
9	10	49	46
10	23	79	68
11	23	90	86
12	20	83	80
13	26	98	92
14	12	38	40
15	21	64	82
16	19	60	62
17	26	79	86
18	19	59	52
19	26	90	89
20	19	49	63
21	22	68	86
22	24	93	89
23	22	49	63
24	17	39	52
25	23	64	68
	Mean = 18.44	Mean = 65.04	Mean = 66.12

APPENDIX J

TAPED INTERVIEWS

Interview Number 1

Interviewer: Have any of you had high school chemistry?

First student: Yes.

Second student: No.

Third student: No.

Interviewer: Were there differences between the high school laboratory and the laboratory part of this course?

First student: Yes. We used a manual in high school in which we filled in blanks. I liked this procedure better as there was a better correlation with the lecture. I often didn't understand the purpose of some of the experiments in our high school course.

Interviewer: Were the experiments too elementary?

First student: No.

Second student: I feel that they were just right for me.

Third student: The experiment on chemical and physical changes seemed too elementary to me, but I think the rest were about right.

Interviewer: What did you do with the information you obtained during the laboratory period?

First student: I always wrote up the report of the experiment right away.

Second student: I usually waited a day or so and then wrote up the report.

Third student: I referred to the textbook for those things I didn't completely understand and then wrote up the report, usually the next day.

Interviewer: Was the textbook of value to you in your learning?

Third student: Yes.

Interviewer: What did this laboratory require you to do most? (No response to this question.)

Interviewer: Would you say that this procedure was "cook book"?

First student: It seemed so to me.

Second student: No.

Third student: It made one think and tie all the parts together. No, I would not say that it was "cook book".

Interviewer: Would you say that the laboratory was a worthwhile learning experience?

First student: Yes.

Second student: Yes, definitely.

Third student: Yes, I felt that it was.

Interviewer: Did the laboratory precedure require specific answers?

First student: No.

Second student: No.

Third student: The generalities came from the specifics and I felt there was quite a bit of flexibility in the writing of the report.

Interviewer: Suppose that you could go in and do the laboratory experiment first and then have the lecture. Would you like this method?

(All three agreed that they would prefer the lecture first.)

Interviewer: Do you have any criticisms of this laboratory experience?

First student: I would like to have the experimental procedure sheet a week ahead.

Second student: I have no criticisms.

Third student: I liked the laboratory quite well.

Interviewer: Would you like to have this type of laboratory in future chemistry courses or would you like to have the traditional laboratory?

First student: I would like to try the traditional laboratory to see how they compare.

Second student: I liked this type and would like the same kind again.

Third student: I liked this method quite well.

Interview Number 2

Interviewer: Have any of you had high school chemistry?

First student: Yes.

Second student: Yes.

Third student: No.

Interviewer: What difference if any did you notice between this laboratory experience and your high school course?

First student: I felt that this was more to the point and provided more of the important details.

Second student: I felt that the topics dealt with in the lecture became alive and meaningful in the laboratory.

Interviewer: Did you feel that the experiments were too elementary?

First student: No. I felt that they were the ones we should have had since they went right along with the topics being discussed in the class.

Second student: A couple of them were the same as we had done in high school but most of them dealt with topics that were new to me and as a result I found them to be quite interesting.

Third student: Since I had not had high school chemistry I felt that they were all worthwhile and necessary for me.

Interviewer: Think if you will about what you did from one laboratory experiment to the other.

First student: I used my findings in the laboratory to write up my laboratory report and also found that the laboratory made the problems in the lecture easier to understand.

Second student: I collected facts during the laboratory and then wrote the report as soon as I could.

Third student: The questions asked in the experiment caused me to think more deeply about the experiment and I learned by this.

Interviewer: Did the directions in the procedure tell you what conclusions should be drawn from the experiment?

(All agreed that they did not.)

Interviewer: Did you feel that this laboratory experience was worthwhile?

First student: Yes.

Second student: Yes.

Third student: I feel that I got more out of the laboratory than I did from the lecture.

Interviewer: Do you think that this was better than your laboratory experience in high school chemistry?

First student: Yes.

Second student: It was more general and dealt with important topics.

Interviewer: Would you want to continue this type laboratory in the next course?

First student: I would like to take the traditional laboratory so that I could compare the two.

Second student: I would agree with him (student number one) in wishing to try the usual laboratory experiments.

Third student: I would like to continue with this method.

Interviewer: What do you think would make the laboratory better?

First student: I would like to see less detail in the introduction.

(No comment from second and third students.)

Interviewer: Was there a good correlation of lecture and laboratory?

First student: I was not in Mr. Richardson's lecture section but there was still a good correlation between the laboratory experiments and the topics dealt with in my lecture section.

(Second and Third students both stated that there was a good correlation.

Interview Number 3

Interviewer: Did you know that you were an experimental group?

(All stated that they did.)

Interviewer: Had you previously had chemistry in high school?

First student: Yes.

Second student: Yes.

Third student: No.

Interviewer: Was this procedure different than that you had in your high school laboratory?

First student: Yes.

Interviewer: How was it different?

First student: It seemed that the high school laboratory was based more on getting an exact answer to each question.

Second student: In my high school course it appeared that we were just supposed to get the blanks filled in and get finished and get out. Here more variation was allowed and there seemed to be a purpose.

Interviewer: Of what good was this?

Second student: This caused me to be more interested in finding out for myself.

Interviewer: Did this stimulate your imagination?

Second student: Yes. I had this course at another college last year and there was no flexibility there. I feel that this laboratory that I am in now caused me to think more.

Interviewer: Do you think that the experiments were too elementary?

First student: I felt that they were just right.

Second student: I liked the laboratory as it was.

Third student: No. I had not had chemistry before and the topics we dealt with in the laboratory seemed easy for me to see. I felt that the pace was good.

Interviewer: Was there good correlation between the laboratory and the lecture?

First student: Yes.

Second student: Yes, there was good correlation.

Third student: I had a different lecture instructor but there was still a good correlation.

Interviewer: Has this been a worthwhile learning experience?

First student: I don't believe it was any more so than the lecture, but it was good. It aroused my curiosity.

Second student: I could understand the concepts better that were taken up in the laboratory.

Third student: I agree with him (second student).

Interviewer: Now let's suppose that we were to turn the procedure around and have the laboratory first and then a lecture over the subject dealt with in the laboratory. What would you think of this?

First student: It might be pretty good. You possibly would be able to better understand the topics dealt with in lecture.

Second student: Maybe you could better see the reason for the mistakes you made in the laboratory.

Third student: I believe I would prefer it the way it was done.

Interviewer: If you take another course of chemistry, would you like to have this same kind of a laboratory procedure?

First student: I would like this same kind of procedure.

Interviewer: Why?

First student: I am not impressed with the commercially prepared manual used here at Central.

Second student: I would prefer the procedure that we used.

Interviewer: Does it stimulate you to think?

Second student: Yes. It does.

Third student: I would prefer to use this type of procedure. It amounted to more than just looking for an answer. I felt that these experiments gave me more freedom to think.

Interviewer: Can you make suggestions for the improvement of this program?

First student: I would like to have the reports we turn in returned to us.

(No comments from second and third students.)

Interview Number 4

Interviewer: Have any of you had chemistry in high school?

First student: Yes.

Second student: Yes.

Third student: No.

Interviewer: Did you notice any differences between your high school laboratory and this present laboratory?

First student: These were more exacting. They required more effort on my part.

Second student: This was better in that we had more experiments and they dealt with topics in greater depth.

Interviewer: Did you feel that the experiments in Mr. Richardson's laboratory were too elementary?

First student: No. I found that I often had to use references to answer the questions in the procedure.

Second student: No. I had to use the data collected in the laboratory experiment to arrive at the logical conclusion.

Third student: I don't think they were too elementary. After I did the laboratory experiment the lecture seemed easier for me to understand.

Interviewer: Was the laboratory pertinent to the lecture in the course?

First student: Yes.

Second student: It agrees well with the lecture.

Third student: The experiments went right along with the lecture and I found that this format stimulated me to think.

Interviewer: Did this procedure provide the answers to the questions?

First student: This method led one to the answers rather than actually give the answer.

Interviewer: What criticisms (constructive as well) do you have of this laboratory experience?

First student: I disliked having to wait in line quite often to obtain supplies.

Second student: I would liked to have had more help in the laboratory.

Third student: I particularly liked the short discussion that was given to us before we did each experiment.

Interviewer: If you could turn this experience around and have the laboratory first and then the discussion, would you like this?

First student: I would like to discuss the experiment afterward.

Second student: I think I would prefer it the way it was.

Third student: I like it as it is.

Interview Number 5

Interviewer: Have you had high school chemistry before?

(All three students replied that they had.)

Interviewer: Did you notice any differences between the high school laboratory and your present laboratory?

First student: I attended a small school and we had very little laboratory at all.

Second student: We did very simple experiments in high school and seldom had laboratory.

Third student: I was in an extremely large class and the laboratory was too crowded. The laboratory experiments were too complex for me to understand. It was the CBA (Chemical Bond Approach).

Interviewer: Were the experiments too elementary?

First student: Most were right for me.

Second student: I felt that they were just right.

Third student: They were not too elementary for me.

Interviewer: Were the experiments pertinent to the lecture?

First student: They seemed to help crystallize what we were talking about in the lecture.

Second student: Most of them were well correlated.

Third student: They seemed to fit quite well even though I had different lecture and laboratory instructors.

Interviewer: Do you feel that this has been a worthwhile experience?

First student: Yes. The experiments helped me understand certain chemical processes.

Interviewer: Did this format help you to learn chemistry?

First student: Yes.

Second student: I felt like I understood the principle after doing the experiment.

Third student: I think this was a worthwhile experience. Since my laboratory background was so limited, I gained a basis for understanding chemistry.

Interviewer: Did this require you to take information from the experiment and formulate an answer or conclusion that you were not aware of before?

First student: Yes, in most cases.

Second student: It usually was a case where you arrived at an answer as a result of the data gathered.

Interviewer: Did it help your ability to think analytically?

First student: Yes.

Second student: Yes.

Third student: I honestly didn't spend enough time outside the laboratory period for it to be of real benefit to me.

Interviewer: Do you feel that this was a worthwhile college experience?

First student: Yes, I do. This is much better than a situation where you fill in the blanks in a notebook.

Second student: You feel that there is actually an atmosphere of learning in this type laboratory experience.

Interviewer: Student number three, you mentioned that you had CBA chemistry in high school and there was no manual there or here, and yet you say this was better. Why?

Third student: This was more on my level of understanding. The first course was above my head.

Interviewer: What would you think of a situation where you had the laboratory first and then the lecture?

First student: I wouldn't like it.

Second student: I think you must have some knowledge of the chemistry involved in an experiment before you go into the laboratory. Otherwise, it might be dangerous.

Interviewer: Would you like to see the next course conducted like this was?

First student: Yes, I would prefer this.

Second student: From what I have heard of the other procedure, I would prefer this.

Third student: Yes.

Interview Number 6

Interviewer: Did you have high school chemistry?

First student: No.

Second student: Yes.

Third student: Yes.

Interviewer: Did you have laboratory as part of your high school course?

Second student: Yes, but I didn't especially appreciate my high school course and I made a D in it.

Third student: Yes. I had a good teacher in high school chemistry but I felt that I learned a lot in this course we are in now because you really had to think a lot. Filling in the blanks in our high school course was easier to do but I didn't learn as much.

Interviewer: Do you feel that the experiments were too elementary?

First student: I don't think I can say as I didn't know that much about chemistry to begin with.

Interviewer: Did you have a difficult time with this laboratory?

First student: No, not especially.

Second student: I felt some of the answers to the questions in the body of the experiment were too obvious.

Interviewer: Were the questions graduated in order of increasing difficulty?

First student: Yes. Even the experiments dealt with progressively more difficult concepts.

Interviewer: Did the laboratory correlate with the lecture?

First student: Our laboratory experiments followed the lecture quite well. The laboratory experiments sort of showed you how it was.

Interviewer: What would you think if you had laboratory work first and then the lecture?

First student: I believe I would like that.

Second student: I suppose it would be useful but I feel that one needs to know something about the specific situation before going into the laboratory itself.

Third student: I would prefer it the way it is.

Interviewer: Did you like the procedure? What did you do? Did it stimulate you?

First student: It wasn't easy but since I was on my own I had to think more about it and I feel that I learned more.

Interviewer: What did you do with the information you got?

First student: I put it together to try to reach a conclusion.

Second student: I didn't actually do much work outside the laboratory.

Third student: I used references when I needed to and looked up information I needed to answer some of the questions.

Interviewer: Do you feel that the format given you gave you a procedure for writing up the experiment?

(All agreed that it was some help.)

Interviewer: Are there any comments you wish to make?

First student: I would like to see the lecture follow the laboratory better.

Interviewer: Would you like this kind of experimental procedure in the next course?

First student: Yes I would. Working answers to questions is better than filling in blanks.

Second student: Yes I would.

Third student: Yes.

Interviewer: Has this been a worthwhile experience?

First student: It sure has been for me. I thought that I would not like chemistry at all and now I see that it is a wonderful system of matter.

Second student: I think it is real interesting and I do feel that it was definitely worthwhile.

Third student: Yes. I enjoyed it.

Interview Number 7

Interviewer: Have all of you had high school chemistry?

(All three said that they had.)

Interviewer: How did this laboratory compare to your high school laboratory?

First student: This is different than what we had in high school in that we used a laboratory book in high school. The answers are not as cut and dried in this procedure as in the high school course. Here you are more on your own. The high school manual was the type in which you filled in the blanks.

Second student: My reaction is the same as his. We filled in the blanks. I liked this method better.

Interviewer: Were you told what to find in this procedure?

First student: No. This procedure was much less confining.

Third student: For me this was better. I like for the teacher that teaches the course to write the laboratory procedure. The teacher then knows the solution to situations that come during laboratory experiments.

Interviewer: Were you more dependent upon yourselves in this laboratory?

First student: Yes.

Interviewer: What did you do with the information you obtained in the laboratory?

First student: I used it to write up the account of the experiment and then reach a conclusion from the data.

Interviewer: Did you find that the laboratory procedure told you how to reach the correct conclusion?

Second student: No. It sort of led you gently to a logical conclusion.

Third student: I would agree with him.

Interviewer: Would you say that this procedure gave you freedom?

(All three agreed that it did.)

Interviewer: Do you believe this has been a worthwhile learning experience?

First student: Yes. It has contributed a lot to my knowledge of chemistry. I know better how to make certain preparations and calculations as a result of this experience.

Second student: This is better than lecture class because we find out about things that we ourselves want to know.

Third student: I would say it was a good experience because of the greater freedom to draw our own conclusions. We dealt with more important basic concepts of chemistry. I also feel that I can use this method for other problems in other courses and come to a successful solution on my own.

Interviewer: Would you say that this laboratory experience was helpful to you in learning how to learn?

First student: Yes. I can't remember all the facts but I can remember the procedure. My high school course didn't do this for me.

Interviewer: If you were going to have another course in chemistry, would you want to have this same type of laboratory?

First student: Yes. I feel that it is a good experience in learning how to learn.

Second student: Yes. I would like to have the same method again.

Third student: I definitely would.

Interviewer: Is there anything you would like to see as improvements?

First student: A little more imagination could be used. The experiments could be changed gradually sothat the student becomes progressively more on his own.

Second student: I can't think of any changes right now.

Third student: I have no suggestions.

Interviewer: Would you like to have laboratory then the lecture instead of the way it is presently?

First student: I believe I would like that.

Second student: No. I don't think I would like it that way.

Third student: No. I wouldn't.

Interview Number 8

Interviewer: Did you have high school chemistry?

(First and Second students had, the Third did not.)

Interviewer: Did you notice any differences between the high school laboratory experiments and those of this laboratory?

First student: Yes. Here we had more time for help and the chance to ask questions. In our high school course there was not enough time. It was a sink or swim situation.

Second student: Here you are more on your own and the success of the experiment depends on your findings during the laboratory period.

Interviewer: What did you do with the information obtained from the experiment from the time you left the laboratory until you returned next time?

First student: I would sometimes seek help from the instructor if I needed explanation and then I would write up the experiment.

Second student: I didn't follow up on this after the laboratory period as I should have.

Third student: I did a rough draft of the report right away and then I came back later and rewrote it in better form.

Interviewer: Did you find that the experiment write-up told you what the conclusions were supposed to be?

(All replied that it did not.)

Third student: The procedure and questions asked helped one to arrive at the logical conclusion. The questions asked were of a probing nature and if you could answer the questions you could then quite easily reach the conclusion.

Interviewer: Do you think that the laboratory was pertinent to or sufficiently related to the lecture?

First student: Yes, I do.

Second student: I had another lecture instructor, but there was still good correlation.

Third student: Mr. Richardson was the instructor for both laboratory and lecture and the laboratory went right along with the lecture.

Interviewer: If you were to take another semester of chemistry would you want to have this kind of a laboratory again?

First student: Yes sir. I think that too many times things are taken for granted that we know when we actually don't know them. This procedure helped me to understand many things well.

Second student: I would like this kind of procedure as I got more out of it than another laboratory I took.

Interviewer: Would you say that this experience helped you in your thinking?

First student: Yes, very much.

Interviewer: Suppose you had a course where the laboratory came first and then the lecture. What would you think of that?

First student: I don't think I would like that as well. I believe that a person needs to know as much about a chemical reaction as possible before going into the laboratory and doing the experiment.

Second student: I would not like that arrangement. I like the short discussion that is given before the experiment.

Third student: I like it better the way it is.

Interviewer: Do you have any contributions to make that would make the laboratory better the next time it is done?

First student: I would like to have more laboratory periods per week.

Second student: I sometimes need more than two hours time.

Interviewer: Were the experiments too elementary?

First student: I thought that the first two were.

Second student: I suppose those were necessary so that we could develop the correct technique.

Thirst student: The first two were somewhat elementary but as a result I feel that I learned what I needed to know so that I could successfully deal with the rest of the experiments.