AN EXPERIMENTAL STUDY IN THE TEACHING OF GENERAL CHEMISTRY AT THE COLLEGE LEVEL

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Preface

The present problem was first brought to the attention of the writer in 1950 at the First Annual Conference on The Teaching of College Chemistry sponsored by the Division of Chemical Education of the American Chemical Society and held at The Oklahoma Agricultural and Mechanical College. During this conference the need for research, on this and several other problems in the teaching of college chemistry, was discussed.

Since a survey of the literature revealed no report of a study designed to answer this problem, it appeared to the writer that if a satisfactory conclusion were to be reached, an experimental study was necessary. This investigation, therefore, was organized as a cooperative study involving an advisory committee made up of faculty members from the Department of Chemistry and the School of Education at The Oklahoma Agricultural and Mechanical College. This investigation was designed to determine the relative effectiveness of two different organizations for the subject matter of the beginning course in general chemistry. The problem of this study was to determine which, if either, of the two organizations contributed more to the learning outcomes of the general chemistry course.

The 530 students who registered for the beginning general course during the first semester of the school year, 1953-1954, were divided by the enrollment procedure into an experimental and a control group to serve as the experimental samples. This teaching assignment was carried out in the Department of Chemistry at The Oklahoma Agricultural and Mechanical College under actual school conditions.

The writer wishes to express his gratitude to Professor Roy Gladstone for his invaluable help during the course of this study. His keen interest, wise judgment, and generous grants of time are thankfully acknowledged. The writer is indebted to Professor O. C. Dermer for the opportunity given him to teach and to carry out this investigation in the Chemistry Department under most favorable conditions. The author is especially grateful to Professor H. P. Johnston for his constant encouragement and helpful guidance throughout the course of the study. Professors M. R. Chauncey and J. W. Richardson were consulted often and with profit. Thanks are due to the entire general chemistry teaching staff who cooperated wholeheartedly.

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CHAPTER ONE

THE PROBLEM AND ITS SETTING

Introduction

During the past twenty years leaders in the field of chemical education have continually called attention to the importance of emphasizing the theory of atomic structure in the teaching of general chemistry.[1,2] Since this is the basic theory of all chemistry, it is generally agreed that it should be incorporated into the beginning elementary course. A review of the literature reveals, however, that there is little agreement as to the time or place for the most effective introduction of this theory into the beginning course.[3,4] The literature further reveals that not one research study designed to answer this problem has been reported.

General Problem

As the basis for improved practices is provided through educational research, it appeared worth while to undertake an experimental study dealing with two different organizations of the instructional material for the teaching of general chemistry at the college level.

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[1] Sister Mary Martinette, "The Presentation of Atomic Structure to College Freshmen," School Science and Mathematics, XL (December, 1940), 808-814.

[2] J. S. Campbell, "Structural Chemistry," Journal of Chemical Education, XXV (October, 1948), 558-562.

[3] George Wakeham, "A 'Logical' Chemical Course?" Journal of Chemical Education, XX (November, 1945), 536.

[4] Harry H. Sisler, "Why Should Theoretical Principles Be Used in Teaching Elementary Chemistry?" Journal of Chemical Education, XXV (October, 1948), 562-565.

During the first semester of the school year 1953-1954, an experimental investigation was inaugurated which was designed to determine the relative effectiveness of two different organizations of the subject matter of the first course in general chemistry for presentation to students at The Oklahoma Agricultural and Mechanical College.

The equivalent or parallel-group method of research[5,6] was employed, and the two organizations compared were arbitrarily designated as the conventional organization and the principles organization. These organizations were compared with respect to their effectiveness in teaching general knowledge and information, the application of principles, and the ability to use the scientific method of problem-solving. A survey of the literature bearing on this general problem is presented in the following section.

Survey of Related Literature

Numerous studies related to the general problem suggested above have been reported. The abundance of material, however, imposes the necessity of confining this review to the reports bearing most directly on the problem under investigation. For convenience of treatment, the literature bearing on the present problem may be classified roughly into two groups: (1) those reports concerned with objectives, and (2) those reports on experimental studies in the fields of methods and procedures.

Reports Concerned with Objectives. The most important influence upon the identification and the selection of objectives for general

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[5] W. A. McCall, How to Experiment in Education (New York, 1926), pp. 14-36.

[6] C. V. Good, A. S. Barr and D. E. Scates, The Methodology of Educational Research (New York, 1941), pp. 493-495.

chemistry has been exerted by the Committee on Examinations and Tests of the Division of Chemical Education of the American Chemical Society. The report, "Accepted Objectives in the Teaching of General Chemistry,"[7] gives a list of objectives which were essentially derived from the Cardinal Principles of Secondary Education.[8] This list of objectives was divided into three groups and summarized as follows:

> I. A study of chemistry should provide the individual with a broad and genuine appreciation and understanding of the chemical aspects of the universe; of the place of chemistry among the sciences; and of what the developments in chemistry mean in modern social and industrial life.
>
> II. A study of chemistry should provide an opportunity for the acquisition: of experiences in the use and the knowledge of the scientific method of thinking, using chemical problems; of a knowledge of natural laws, important principles and facts; of the ability to draw generalizations from specific experimental data and to apply important principles; and of some skill in laboratory manipulations.
>
> III. A study of chemistry should provide an opportunity for the individual to determine his interests and aptitudes in chemistry as a vocation.[9]

The first form of the American Chemical Society Cooperative Examination in General Chemistry,[10] published in 1935, was based on these objectives. Since 1935, these objectives have been continuously reviewed, and eleven forms of the American Chemical Society Cooperative Examination in General Chemistry have been published. The two latest forms (Form G and Form K) of this examination were used as the principal measuring devices in the present study.

Several other outstanding statements of objectives in science

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[7] Otto M. Smith, Journal of Chemical Education, XII (April, 1935), 180-183.

[8] National Education Association (Washington, D. C., Bulletin, 1918, No. 35).

[9] Smith, op. cit., p. 183.

[10] Prepared jointly by the Cooperative Test Service of the American Council on Education, New York, New York, and the Committee on Examinations and Tests, Division of Chemical Education, American Chemical Society, St. Louis University, St. Louis, Missouri.

teaching that have greatly influenced trends in the teaching of general chemistry have been reported. These reports have served to direct the aims of science education away from formally organized facts and toward concepts, skills, and generalizations which would provide functional information for future life. Specific values of science instruction for these purposes were listed in the report, Reorganization of Science in the Secondary Schools.[11] This report, confined to science in secondary education, was the first comprehensive document to deal exclusively with science education.

The report, "The Place of Science in Education",[12] emphasized the importance of scientific thinking as an objective of science education. One of the most influential reports in science education appeared in 1932 as the Thirty-first Yearbook of the National Society for the Study of Education.[13] This publication advocated that science instruction be organized around broad concepts or generalizations. The objectives of science instruction were stated as being "to give the student a functional understanding of certain science subject matter, and an understanding of the scientific method, with an accompanying scientific attitude." In 1938, the report entitled Science in General Education[14] appeared. This report recommended the centering of science education around broad areas of living and stressed the use of reflective thinking in problem-solving.

In 1947, the National Society for the Study of Education published

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[11] National Education Association (Washington, D. C., Bulletin, 1920, No. 26), 12-14.

[12] American Association for the Advancement of Science, School Science and Mathematics, XXVIII (June, 1928).

[13] A Program for Teaching Science (Bloomington, 1932), p. 57.

[14] Progressive Education Association (New York, 1938), pp. 41 and 46.

its Forty-sixth Yearbook,[15] which dealt with the progress in science education and stressed the importance of a functional understanding of science.

At the present time the trend in science education at the secondary and college levels appears to be toward the more functional point of view.[16] Now among leaders the chief aim of science instruction lies in the modification of the behavior pattern of the individual in such a way as to make his adjustments to the problem of living more effective and more satisfying.

Studies in the Field of Methods and Procedures. Many experimental studies bearing upon various classroom and laboratory problems in the field of science education have appeared. The selections for this review are confined largely to the reports of those studies which made use of a carefully controlled experimental design and which used a refined statistical procedure in evaluating data.

In 1926 Horton used the parallel-group method of research in conducting an experiment to determine the relative effectiveness of the demonstration and laboratory methods of teaching high school chemistry. He used as subjects 464 pupils in 17 classes, divided at random into an experimental group consisting of 85 pupils in four classes, and a control group consisting of 379 pupils in eleven classes. Equivalent initial ability was assumed in both groups.

All the classes were treated in exactly the same manner except that for one period out of five each week the experimental-group

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[15] Science Education in American Schools (Chicago, 1947), pp. 32-33.

[16] Stanley B. Brown, "Trends in Science Education - 1953," The Science Teacher, XXI (February, 1954), 84-85.

students remained in their classrooms and observed teacher-demonstration exercises, whereas the control-group students performed the exercise individually in the laboratory. This procedure constituted the experimental variable.

Two types of tests were used in the evaluation of the experiment (1) the written mid-term and final examinations, and (2) a test of manipulation to be performed individually by the students in the laboratory. Critical ratios between the mean scores for the two groups on the mid-term, final, and laboratory performance tests were 2.50, 2.09, and 4.96 respectively in favor of the experimental group.

This study was important because it was one of the first attempts to measure desirable outcomes of instruction other than factual information and to apply a refined statistical technique in treating the data. It is difficult to understand why the numbers of students in the two groups were not made more nearly equal.

A similar study was conducted by Payne, in which he administered a battery of three standardized tests at the beginning, at the middle, and at the end of the experiment. This procedure, thought not feasible to use in the present study, was noteworthy because it attempted to measure the actual gain in knowledge during the experiment.[18]

More recently, Ward conducted a study in the Fort Worth, Texas, high school system, involving four schools, five teachers, and 420 students. The purpose of this study was to compare the learning

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[17] "Measurable Outcomes of Individual Laboratory Work in High School Chemistry," <u>Contributions to Education</u> No. 303, Columbia University (New York, 1928), pp. 60-90.

[18] "The Lecture Demonstration and Individual Laboratory Methods Compared: (1) The Literature, (2) The Distribution of Time, (3) Experimental," <u>Journal of Chemical</u> Education, IX (May, June and July, 1932), 932-937, 1097-1102 and 1277-1294.

outcomes of two instructional methods which were applied in chemistry classes. The control group was taught by a method based on the philosophy that the mastery of subject matter was the end sought in the teaching of chemistry. The experimental group was taught by a method which made use of learning situations based on interests, needs, and capacities of the students.

In order to reduce the teacher variable, each teacher taught one experimental and one control class. The groups were equated on the basis of I. Q.'s, scholarship averages, and scores made on two tests (1) a functional test and (2) "A test of chemistry information." The same two tests were administered at the end of the semester, and the gain on scores made by the two groups was compared. Thirty-six critical ratios between various pairs of measures were computed. Eleven of these favored the experimental group, but the remaining twenty-five were too small to indicate a true difference.[19]

This study was significant because it was a cooperative study involving several teachers and the entire course in high school chemistry.

Several of the procedures used by Bond in his experiment in the teaching of genetics have significance for the present study.[20] For example (1) as subjects he used 212 college students who were divided by the enrollment procedure into two groups, (2) the total

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[19] "An Experimental Study of Two Methods of Teaching Chemistry in Senior High School," Journal of Experimental Education, XI (September, 1942), 69-80.

[20] "An Experiment in the Teaching of Genetics, with Special References to the Objectives of General Education," Contributions to Education No. 797, Columbia University (New York, 1940).

scores on the American Council on Education Psychological Examination[21] were used to attest the equivalence of the groups, and (3) the testing device was made by selecting certain questions from three forms of the Cooperative Zoology Test.[22]

Several studies dealing with the instructional outcomes of science courses when organized around important scientific principles have been reported. Kilgore used physics classes in four Washington, D. C., high schools to determine how the organization of the physics course around major generalizations would affect the ability of pupils to apply physics principles in new situations.

During the first semester of the school year 1935-1936, four classes in physics, designated as the control group, were given standard instruction in the beginning physics course. The four classes of pupils who enrolled in the same beginning course during the second semester were designated as the experimental group and were taught by a method which stressed the principles of physics. The same four teachers who taught the control groups the first semester taught the experimental groups the second semester. The experimental factor consisted of an attempt to teach the principles of physics.

Ten organizational sheets designed to call attention to the major principles of physics were supplied to each pupil and teacher during the experimental course. The teachers were asked to increase their emphasis on principles while they were teaching these groups.

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[21] L. L. Thurstone and Thelma G. Thurstone, American Council on Education Psychological Examination for College Freshmen, University of Chicago, Cooperative Test Division, Educational Testing Service, Chicago, Illinois.

[22] Cooperative Test Service of the American Council on Education, New York, New York.

The tests used included: (1) a test consisting of certain parts of the Cooperative Physics Test (Form 1934)[23] and (2) a principles test which was constructed by the investigator. These tests were administered to the control groups in January and to the experimental groups in June.

The experimental groups showed superiority over the control group on both tests. The critical ratios between the mean scores on the Cooperative Physics Test and the principles test were .79 and 5.75 respectively in favor of the experimental groups. Kilgore concluded that "learning can be effectively guided in a manner which will aid the pupil in dealing with new problems."[24]

Babitz and Keys used eight classes in two California high schools to form equivalent groups. Four of the classes designated as the control groups received conventional instruction. The other four classes designated as the experimental groups were taught by a method which stressed applications of principles. A test which required the solution of problems in chemistry and an identification of the chemical principles involved was administered to all the pupils at the end of the teaching assignment. The scores on this test revealed that all the experimental groups showed superiority over the control group; however, these differences favoring the experimental groups were too small to be statistically significant.[25]

Mason used two groups of students in the Department of Biological

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[23] Ibid.

[24] "Identification of Ability to Apply Principles of Physics," Contributions to Education No. 840, Columbia University (New York, 1941), pp. 1-19.

[25] Melton Babitz and Noel Keys, "An Experiment in Teaching Pupils to Apply Scientific Principles," Science Education, XXIII (December, 1939), 667-670.

Science at Michigan State College to compare two methods of teaching biological science. The two methods, arbitrarily designated as the scientific thinking method and the descriptive method, were compared with respect to their effectiveness in (1) teaching factual information, (2) meeting the overall objectives in the course, (3) developing the ability to think scientifically, and (4) developing a scientific attitude. The descriptive method stressed the factual information of the course, whereas the scientific thinking method attempted to teach directly for the development of the methods and attitudes of science. Tests used in the study included: (1) the Cooperative College Biology Test,[26] (2) departmental term-end examinations, (3) a departmental Comprehensive Examination in Biological Science - Scientific Thinking Test, (4) Noll's What Do You Think? test,[27] and (5) a test constructed by the investigator designed to measure certain beliefs about biology. The study was continued through three terms, throughout the school year of 1949-1950.

The results of an analysis of the data showed that: (1) the descriptive method was the most effective in teaching factual information as measured by the Cooperative College Biology Test after the first term and also after the second term; but after three terms, the two methods were equally effective; (2) the two methods were equally effective in teaching the overall objectives of the course; (3) the two methods were equally effective in teaching the abilities associated with scientific thinking, as measured by the Comprehensive Examination in Biological Science - Scientific Thinking Test; (4) after

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[26] Cooperative Test Service of the American Council on Education, New York, New York.

[27] Bureau of Publications, Teachers College, Columbia University (New York, 1935).

three terms the scientific thinking method appeared to be slightly more effective, as measured by Noll's What Do You Think? test.[28]

Although this review is far from exhaustive, it is inclusive enough to indicate the nature and extent of the studies in the field of methods and procedures and to exhibit the experimental techniques usually employed by the investigators. There appears to be no overall explanation to account for the fact that some investigators reported significant findings but that others did not. The significant findings reported by some investigators were possibly due to the unequal treatment given the two groups. For example, in the study conducted by Horton the superiority of the experimental group on the laboratory manipulation test could have been expected since the control group had not been in the laboratory prior to the administration of the laboratory performance test; better motivation could have accounted for the superiority of the experimental group in the study reported by Ward, since this group was taught by a method which made use of learning situations based on interests, needs, and capacities of the pupils; the highly significant difference (critical ratio 5.75) in favor of the experimental group on the principles test in physics, as found in the study reported by Kilgore, could have resulted from the fact that the principles test, constructed by the investigator, dealt with the same principles as those given in the organizational sheets and stressed by the teachers in the experimental-group classes. In the present study an effort was made by the investigator to insure equality in the treatment of the two groups whenever possible. Many other studies

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[28] "An Experimental Study in the Teaching of Scientific Thinking in Biological Science at the College Level," Science Education, XXXVI (December, 1952), 270-284.

concerned with methods and procedures have been reported; references to those of significance to this study appear in the bibliography.

Summary

Since the publication of the Cardinal Principles of Secondary Education, several reports have appeared which have served to direct the aims of science education away from formally organized factual information and toward an understanding of the methods of science. The Committee on Examinations and Tests of the Division of Chemical Education of the American Chemical Society has exerted the most important influence upon the general chemistry course. Because of the work of this committee the aims of the general chemistry courses have been more clearly defined, and the American Chemical Society Cooperative Examination in General Chemistry has been designed to measure the attainment of these aims.

The writer has found no report of a study designed to compare the two subject-matter organizations dealt with in the present study. Numerous related studies, however, have appeared. These studies have generally employed the parallel-group method of research. The experimental and control groups were formed by taking a random sampling of a larger population. The scholastic aptitude of groups was most frequently determined by using the scores on the American Council on Education Psychological Examination. The significance of the difference between various statistics was determined by calculating critical ratios between the statistics. In some of the more recent studies, the techniques for analysis of variance and covariance were used in equating groups and in testing hypotheses.[29]

The results of this review of the literature showed that the most effective procedure of science instruction is not clearly defined and that there is need of experimentation to determine the most useful procedures in this field. The present study was devoted to the task of investigating some aspects of this problem.

Restatement of the Problem

This investigation was designed to determine the relative effectiveness of two different organizations of the subject matter of the first course in general chemistry at The Oklahoma Agricultural and Mechanical College. The difference between the two organizations lay principally in the order of presentation of the course content. The first organization, which will be called the conventional organization, was so designed that the student could develop a suitable background in descriptive chemistry before the concepts of atomic structure theory were introduced. The second organization, which will be called the principles organization, introduced the student early in the course to the concepts of atomic structure theory so that these might be used as a basis for developing the other materials of the course. These two organizations were compared with respect to their effectiveness in helping students to develop: (1) the ability to recall factual information, (2) the ability to apply principles, and (3) the ability to use the scientific method of problem-solving, as the achievement of these abilities was determined by the measuring

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[29] Palmer O. Johnson, <u>Statistical Methods of Research</u> (New York, 1949), pp. 246-260.

devices used.

The problem of this study was to determine which, if either, of two organizations of the course content contributed more to the learning outcomes of the general chemistry course. The problem, therefore, became one of testing the null hypothesis:[30] that differences between mean achievement scores of the two groups are no greater than differences which could be expected to arise as a result of chance fluctuations in random sampling.

Need for the Study

The need for the study stemmed from the fact that many chemistry teachers felt that recent theoretical developments, such as the atomic structure theory, should be incorporated into the beginning elementary course.[31,32,33] Other teachers felt, however, that the concepts of atomic structure theory were too difficult for many beginning students to comprehend. They recommended, therefore, that the introduction of these concepts be delayed until the student had developed a background of chemical knowledge sufficient to enable him to understand this complex theory.[34,35] A survey of the literature revealed no report of an attempt that had been made to determine the readiness of beginning

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[30] Henry E. Garrett, Statistics in Psychology and Education (New York, 1953), p. 213.

[31] Harry H. Sisler and Calvin A. Vanderwerf, "Modern Theory: A Tool in Teaching Elementary College Chemistry," Journal of Chemical Education, XX (October, 1943), 479-483.

[32] James S. Coles, Leallyn B. Clapp, and Robert P. Epple, "A New Program for Teaching the Fundamentals of Chemistry in College," Journal of Chemical Education, XXVI (January, 1949), 10-14.

[33] Ernest A. Wildman, "The Need for Modernizing the General Course," Journal of Chemical Education, XII (January, 1935), 11-16.

[34] Ernst A. Hauser, "This Nation Will Need More Chemists After the War," Journal of Chemical Education, XXI (October, 1943), 513-515.

[35] Anthony Standen, "Can Elementary Chemistry Teaching Be Logical?" Journal of Chemical Education, XXII (November, 1945), 554-557.

college chemistry students for the development of these theoretical concepts.

Each of these subject-matter organizations was favored with respect to certain aspects of the psychological laws of learning. The proponents of the conventional organization claimed that this plan was favored in that it started the beginning course with facts and simple concepts which the student could readily learn and understand. As the course progressed, therefore, this organization provided an opportunity for the student to develop a background of chemical knowledge sufficient to enable him to understand the complex atomic structure theory. The advocates of this plan pointed out that psychologists generally agree that students progress through levels of mental maturity and that at each level of mental maturity students are able to participate effectively in certain learning experiences. Other learning experiences may be either too simple or too complex to be effective in the learning process at a given level of maturity.[36] They further point out that facts build together to form concepts; similarly, concepts build together to formulate principles, thus establishing a sort of hierarchy.[37] From this standpoint the conventional organization is sound.

Many important factors also favored the principles organization. In recent years many people have acquired, as a result of the much publicized military and peace-time uses of atomic energy, sufficient knowledge of the component parts of the atom: electrons, protons, neutrons, etc., to enable them to form some elementary concepts of atomic structure. Consequently, many students entering the beginning

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[36] National Society for the Study of Education, Thirty-first Yearbook, p. 5.

[37] *Ibid.*, pp. 3-6.

chemistry course are highly interested in the atom and are eager to expand their knowledge of the atomic structure theory.

It was the hope of the proponents of this organization that, if the complex ideas of atomic structure were clearly and painstakingly presented during this initial period of interest, they might be mastered to the extent of becoming meaningful to the student. This had importance for the present study because an understanding of this theory would give the student a broad picture of chemistry, which might enable him to fit other facts into the picture as they were learned.

Learning organized around such concepts should be important to the learner in at least two respects. First, this kind of organization should help the student to generalize his learning from one content field to another and to apply his knowledge to broader areas of everyday life. Second, research shows that learning acquired under such conditions proves much more permanent than learning concerned with isolated skills and items of unrelated information.[38]

Since each of these organizations had many advantages, their relative effectiveness could be determined only by a carefully executed experiment.

Purpose of the Study

The purpose of this study was to compare two subject-matter organizations with respect to their effectiveness in teaching general knowledge and information, the application of principles, and the ability to use the scientific method of problem-solving when applied in general chemistry

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[38] Ralph W. Tyler, "Permanence of Learning," Journal of Higher Education, IV (April, 1933), 203-204.

classes at The Oklahoma Agricultural and Mechanical College.

Scope of the Study

The teaching function of this study was confined to the first-semester course in general chemistry. During this course, most of the principles of general chemistry are presented. The evaluation instruments were applied at the close of the first-semester course and during the progress of the second-semester course.

Limitations of the Study

This study was limited to the organization of the course content in one particular sequence of courses in the Department of Chemistry at The Oklahoma Agricultural and Mechanical College. The study was further limited to the student population of these courses when taught by the particular instructors involved. The findings, therefore, would apply only to the general population from which these students and instructors came. In order that the findings have more general applicability the experiment would have to be repeated in other schools with different courses, instructors, and students.

CHAPTER TWO

PROCEDURE

The teaching function of this study was confined to the general chemistry course designated as Chemistry 114, the first course of a two-semester sequence of which the remainder is designated as Chemistry 124. This experiment was designed so that two identifiable groups of students were taught for one semester by two different subject-matter organizations, and the learning outcomes resulting from this instruction were measured by suitable measuring devices. This teaching assignment was executed in the Department of Chemistry at The Oklahoma Agricultural and Mechanical College under actual school conditions; therefore, it was impossible to control rigidly all the factors that might have affected the results. In organizing and executing the investigation, several basic assumptions were made.

Basic Assumptions

Some basic assumptions for this investigation were as follows:

1. The course content of general chemistry, viz., the facts, concepts, and principles, was meaningful information and as such could be learned, and the learning outcomes resulting therefrom could be measured by suitable measuring devices.
2. Since the tests used were reliable and valid measures of the objectives of general chemistry, as they appear in the literature, it was assumed that they were suitable measures for the purposes of this study.

3. The design of the experiment and the statistical treatment of the data collected were of such a nature that any significant difference between mean achievement of students in the two groups, as revealed by the scores on the measuring instruments used, may be assumed to be due to differences in the two subject-matter organizations used in this study.

Definition of Terms

Chemical Terms. The following chemical terms were defined in Hackh's Chemical Dictionary as follows:

Chemistry. The fundamental science of the structure of matter and the composition of substances, their transformations, analysis, synthesis, and manufacture.

Science. Systematized and verifiable knowledge reached by observation, measurements and/or experiments.

Fact. Anything that has real existence. Known to be true. A result established by repeated experiment.

Law. A generalized statement of facts or principles.

Natural Law. The formulation of systematized experience as to the workings of nature. For example, the periodic law.

Theory. The reduction of data or facts or a principle, and the demonstration of their interrelations.

Hypothesis. A theory which has not been fully proved by experiment.

Principle. A theory or assumption; a fundamental concept; as the principle of valency.[39]

Statistical Terms. The following statistical terms were defined in the Statistical Dictionary of Terms and Symbols as follows:

Critical Ratio. The difference between two comparable statistics divided by the standard error of the difference. When the difference is between two means, the critical ratio is also called the standard difference.

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[39] Julius Grant, editor (Philadelphia, 1944).

Statistic. A value such as a mean, standard deviation, or correlation coefficient, calculated from an observed sample with a view to characterizing the universe from which the sample is drawn; a function of observations designed to estimate the corresponding value in the universe from which the sample is drawn.

True Difference. The mean of an infinite number of measures of the difference; the difference between the true measures of the two variables or items.

Sample. A finite number of observations or cases, selected from all the cases in a particular universe; part of a total group chosen for investigation.

Arithmetic Mean. The sum of the measures, observations, magnitudes, items, or scores in a statistical series divided by their number of frequency.

Standard Deviation. The square root of the arithmetic mean of the squares of the deviations of the value of the variable from their arithmetic mean; the root mean square of the deviations from the mean.[40]

Other Terms.

General Chemistry, College Chemistry, Beginning College Chemistry, Elementary College Chemistry, and Freshmen Chemistry as used in the literature refer to the introductory college course in general chemistry.

The Parallel or Equivalent Group Method of Research. The method of research in which two or more groups of students, as nearly equivalent as possible in all respects, are used at the same time. The procedure consists of treating the groups in exactly the same manner except for the experimental variable.

Subject-Matter Organizations

This investigation, an experimental study, proposed to determine the relative effectiveness of two ways of organizing instructional materials for teaching general chemistry at the college level. Employing the equivalent or parallel-group method of research, the investigator compared the relative effectiveness of the conventional and principles organizations in achieving the various goals of general chemistry instruction.

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[40] Albert K. Kurtz and Harold A. Edgerton (New York, 1939).

The conventional organization referred to the organization used in presenting general chemistry subject matter to classes at The Oklahoma Agricultural and Mechanical College and at most other schools throughout the country.[41] This organization made use of approximately one semester of general course content in preparing the student for the introduction of the atomic structure theory. In this study the concepts of atomic structure theory were presented to the student in the conventional classes during the last six lectures of the term. The principles organization referred to a newer organization of the subject matter in general chemistry courses. This organization was introduced at the Eighty-eighth National Meeting of the American Chemical Society at Cleveland, Ohio, September 10-14, 1934, during a symposium on "Modernizing the Course in General Chemistry," conducted by the Division of Chemical Education.[42] In essence, this organization introduced the students to the concepts of atomic structure theory early in the beginning course and to the use of these concepts as a basis for the other material of the course. In this study the presentation of the concepts of atomic structure theory to the classes taught by the principles organization came during six lectures starting the second week of the semester. These two organizations were presented in the theory portion of a general course designated as Chemistry 114.

Chemistry 114 and Chemistry 124 were offered to students in the School of Engineering, students in agriculture who plan to study veterinary medicine, science majors from the School of Arts and

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[41] Herbert A. Thelan, "Methodological Study of the Learning of Chemical Concepts of Certain Abilities to Think Critically in Freshmen Chemistry," Journal of Experimental Education, XIII (September, 1944), 53-75.

[42] Ernest A. Wildman, "The Need of Modernizing the General Course," Journal of Chemical Education, XIII (January, 1935), 11-16.

Sciences, pre-medical and pre-nursing students, and education students who plan to teach science courses in the high schools. The weekly program of these courses consisted of two fifty-minute lecture periods, one fifty-minute discussion or quiz period, and one three-hour laboratory period. All course content was introduced in the formal lectures given during the lecture periods. These lectures, basically factual in nature and teacher-centered, provided no classroom participation by the students, whereas the quiz-laboratory classes were informal and student-centered.

Both groups of students used the same textbook, Textbook of Chemistry by Edward Mack et al., Ginn and Company, New York, 1949, and the same laboratory manual, Semimicro Laboratory Exercises in General Chemistry by J. A. Burrows et al., Macmillan Company, New York, 1952.

In order to facilitate the execution of the experiment and to coordinate all phases of the work (lecture, quiz, and laboratory), the investigator prepared a detailed syllabus for each of the subject-matter organizations used (Appendices I and II). The syllabi contained, in brief form, all the subject matter presented during the courses and the order of presentation. Neither of the syllabi followed the exact order of presentation which was given in the textbook. The syllabus used by the control group, however, had almost the same arrangement. The content material of the two syllabi was the same, but the characteristic differences appeared in the organization of the content. The information provided by the syllabi, extensive enough so that the textbook could be used as a reference, tended to reduce the textbook variable. The appropriate mimeographed syllabi were distributed to the students in both groups.

A detailed discussion, outlining the specific characteristics of the two subject-matter organizations, follows:

The Conventional Organization. The order of presentation in the conventional syllabus closely followed that of the textbook, which was representative of texts used in other conventional courses. This 876-page edition was in the line of direct descent of five earlier editions extending back over a period of more than thirty years. The principal ideas of the conventional organization were expressed in the preface of the textbook as follows:

> It has seemed wise to continue the past practice of presenting first a considerable amount of fact about the characteristic behavior and properties of some typical elements and their compounds before considering in detail the theory of atomic structure--in terms of which chemical behaviors and properties are interpreted and explained. But an early, brief preview of the atom is given, and is intended to serve the student as a guiding thread of simple but basic theory to tide him over until the important topic of atomic structure can be developed systematically in later chapters.[43]

The investigator attempted to design the conventional syllabus in order that these ideals might be achieved.

This syllabus contained, in brief form, all the essential information found in the first eighteen chapters of the textbook and had essentially the same arrangement except for minor changes. For example, (1) normal and molar solutions were discussed in the same unit, (2) the nomenclature of acids, bases, and salts was discussed in connection with the chapter dealing with these compounds, and (3) the discussion of the electronic configuration of the atom (Chapter 18) was extended to include the arrangement of the electrons in the sub-shells of the atom. These changes were made in order to secure better organization and to introduce additional information needed in the course. The additional material on atomic structure made possible

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[43] Mack et al., p. vi.

an explanation of the transition elements, the rare earths, and the relationships between their configuration and the modern long form of the periodic table. This information was pertinent to the second-semester course. The transition between the two courses was made at this point; that is, the first-semester course ended by introducing and discussing the information contained in Chapter 18, and the second-semester course started by reviewing this information.

The Principles Organization. The experimental course of this study was organized by the investigator, after analyzing textbooks, studying the literature, and reading several course outlines obtained from other schools. The syllabus, designed to be as representative of this organization as possible, was prepared by selecting the appropriate sequence of topics from the textbook. Extensive references were used in the syllabus to aid the student in his study.

The order of presentation was markedly different in this organization. Chapters One and Two, the introductory chapters, and the first parts of Chapter Three were handled in the usual manner. The brief preview of the atom, given in the latter part of Chapter Three, was used as a point of departure to introduce the complete atomic structure theory as presented in Chapter 18; thus, the last part of Chapter Three, "A Preview of the Atom," the first two parts of Chapter Eight, "The Laws of Chemical Combination," and "Atomic Theory," all of Chapter 18, "The Structure of the Atom," and the additional materials dealing with the sub-shells of electrons, included in the conventionally organized syllabus, were used to develop a complete unit on atomic structure theory. A detailed discussion of the classification of the elements and of the periodic system followed this unit.

An effort was made, both in the syllabus and in the lectures, to present these intangible concepts in such a way as to enable the student to develop a working knowledge of the periodic system and to develop an understanding of the relationship of this system to the electron configuration of the various atoms. The answers were given to the questions: (1) Why do chemical reactions take place? (2) How do chemical reactions take place? The configurations of the transition elements and the rare earths were discussed, and the relationship between their electron arrangements and the long form of the periodic table were explained.

The remaining chapters, which for the most part dealt with descriptive chemistry, were handled in the usual manner, except that an effort was made by the instructors to relate the material, whenever possible, to the atomic structure theory. For example, the halogen and alkali families of elements were studied in the light of this information as were ionization, oxidation-reduction, and electrochemistry, whereas in the conventional procedure these were studied before the introduction of this fundamental theory. The atomic structure theory was thus used as a basis for other materials of the principles course and was repeatedly applied in the explanation of subsequent materials.

Experimental Design

Subjects. Five hundred twenty-five general chemistry students at The Oklahoma Agricultural and Mechanical College were selected as the subjects for this study. The subjects were divided by the enrollment procedure into four lecture sections, and students in two of the lecture sections (Section One and Section Two) were assigned by the same procedure to nine quiz-laboratory classes. This group, using the principles organization, was designated as the experimental group. The students in the other two theory sections (Section Three and Section Four) were likewise assigned to nine quiz-laboratory classes; and this group, taught with the conventionally organized subject matter, was designated as the control group. The lecture sections were taught by two instructors who, along with an additional instructor and five graduate fellows, taught the quiz-laboratory classes.

Control of Variables. Each of the two theory instructors presented lectures to one experimental class and to one control class. The quiz-laboratory instructors were divided into two groups as nearly equal as possible with respect to experience and training. One of these groups taught the experimental quiz-laboratory classes; the other group taught the control classes. No quiz-laboratory instructors were involved with both groups except those who were members of the permanent teaching staff. The quiz-laboratory instructors attended the appropriate lectures for the purpose of obtaining a better understanding of the exact materials introduced and of the teaching method used. These two groups of instructors were given equal treatment with respect to assistance, encouragement, and other motivation factors. The two lecturers familiarized themselves as completely as possible with the two subject-matter organizations and cooperated in selecting the materials and writing the syllabi. Careful agreement was reached between the lecturers as to the amount of time to devote to each phase of the instruction, and a time-table was carefully followed throughout the experiment. All attempts by the lecturers at motivating the students were equally administered to the two groups. A weekly staff meeting was used for the purpose of discussing any questions and for further coordinating all phases of the work. This arrangement of the teaching assignment tended to reduce the teacher variables.

Every precaution was taken to prevent the experiment from affecting the performance of the students and teachers involved. Precaution was taken to prevent the students from knowing that the performances of the two groups would be compared. This was satisfactorily achieved by telling the students that they were competing for grades only with the students in their group, just as though the two organizations constituted two separate courses. This served to prevent students from attempting to change from one group to the other, and no student requested such a change during the study. Further precaution was taken to prevent the quiz-laboratory instructors from recognizing that a comparison was being made between the performances of the two groups of students. The quiz-laboratory instructors and, no doubt, most of the students knew an investigation was being conducted, but there was no evidence of competition between the groups during the experiment. The first-semester examinations, which were not common to the two groups, were made as nearly equivalent as possible with respect to the number and difficulty of the test items used; and when letter grades were assigned, each group of students received approximately the same number of each letter grade.

In order to reduce the variables due to the time of day and time of week for class meetings, the following procedure was used:

(1) Theory Section One, which met each week on Monday and Wednesday at one o'clock, was paired with Theory Section Four, which met each week on Tuesday and Thursday at nine o'clock. Theory Section

Two, which met each week on Tuesday and Thursday at eleven o'clock, was paired with Theory Section Three, which met each week on Monday and Wednesday at eleven o'clock.

(2) The quiz-laboratory classes were divided equally between experimental and control classes for each laboratory session with the following exceptions: two experimental classes, which met each week on Monday morning, were paired with two control classes, which met each week on Friday morning; one experimental class, which met each week on Tuesday morning, and an experimental class, which met each week on Saturday morning, were paired with two control classes, which met each week on Friday afternoon.

Second-Semester Procedure

The second-semester course (Chemistry 124), a continuation of Chemistry 114, was taught by the conventional subject-matter organization. The content of Chemistry 124 stressed descriptive chemistry and the use of the principles of chemistry in solving problems of a chemical nature in metallurgy, agriculture, and other industries. The procedure in this course was that each instructor referred continuously to the periodic table and atomic structure as he discussed the various families and groups of elements as they appear in the periodic system. The students who registered for this course during the experimental study were assigned by the enrollment procedure to three lecture sections and to twelve quiz-laboratory classes. The experimental and control groups, therefore, lost their identities and were taught by the same subject-matter organization. The course was begun by reviewing the material discussed in Chapters 17 and 18, "The Periodic System" and "The Structure of the Atom," and followed the textbook presentation thereafter. All except one of the evaluating instruments were applied during this course.

Examinations

The examinations used for the purpose of evaluating the outcomes of this study were: (1) the American Council on Education Psychological Examination, Form 1948; (2) the American Chemical Society Cooperative Examination in General Chemistry, the first four parts of Form G and Form K; (3) A Test of Scientific Reasoning and Understanding (Physical Science, Form A); (4) a test constructed by selecting certain questions from the American Chemical Society Cooperative Examination in General Chemistry, Form 1944, Form 1946, Form 1948, and Form Z, hereafter referred to as A Test on the Scientific Method; and (5) three departmental intra-semester examinations.

These tests were assumed to be valid instruments. The reasons for this assumption are evident from the discussion of the various tests that follows:

The American Council on Education Psychological Examination. This examination, administered during registration by the college authorities, was used to determine the relative scholastic aptitude of the two groups of students. Recognized as being an effective and widely used device for predicting scholastic aptitude,[44] this test gives a high (approximately .50) coefficient of correlation with the natural sciences.[45]

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[44] Lawrence M. Deridder, "Relationship between Gross Scores on the A.C.E. and Academic Success," Journal of Educational Research, XLVI (January, 1953), 353-358.

[45] Jack W. Dunlap, The 1940 Mental Measurements Yearbook, edited by Oscar E. Buros (Virginia, 1945).

The American Chemical Society Cooperative Examination in General Chemistry, Form G and Form K. As an achievement test designed for use at the end of a two-semester sequence of courses, this examination was used as a measure of the degree of achievement of the overall objectives of the two courses. The first four parts of Form G, covering General Knowledge and Information (25 questions), Applications of Principles (25 questions), Quantitative Applications of Principles (20 questions), and Scientific Method (20 questions), were used as a term-end examination for the first-semester course. Form K, covering General Knowledge and Information (30 questions), Application of Principles (50 questions), and Quantitative Applications of Principles (30 questions), was used as a comprehensive examination at the end of the experiment.

The reliability of the early forms of this examination was determined by using the split-test method.[46] More recently the correlation between the scores on equivalent forms has been used in determining the reliability. The validity of the items of the examination was determined by a method described by Phelan as follows:

> It is assumed that the students who get high scores have a better understanding of the subject, as measured by the test, than those who get low scores. For each item, the percentage of right answers of the high-scoring group is plotted against the percentage of right answers by the low-scoring group. The 45° diagonal of the graph, therefore, represents no discrimination, an item answered correctly by exactly the same percentage of good and poor students. This diagonal is given a value of zero, and the amount of deviation of each item from the diagonal determined its validity index.[47]

Anderson concluded a review of this examination as follows: "In summary, one may reasonably say that this test seems capable of measuring

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[46] Palmer O. Johnson, Statistical Methods in Research (New York, 1949), p. 126.

[47] Earl W. Phelan, "The 1936-1937 College Chemistry Testing Program," Journal of Chemical Education, XIV (December, 1937), 586-590.

quite satisfactorily the objectives of general chemistry as reported in the literature."[48]

A Test on Science Reasoning and Understanding (Physical Sciences, Form A). As one of the measuring devices prepared by the Cooperative Study of Evaluation in General Education (conducted under the auspices of the American Council on Education and directed by Dr. Paul Dressel), this examination was designed to measure the student's ability to use the scientific method of problem-solving in the area of physical science. Specifically, this test was designed to measure the ability of the student to read and to interpret scientific materials in terms of his understandings of the broad principles of science and to apply scientific knowledge to new situations. This examination, consisting of fifty-five multiple choice items, was administered during the fifth week of the second semester. No data are available as to its realiability or validity, although it is known to correlate highly with the American Council on Education Psychological Examination.[49]

A Test on the Scientific Method. This fifty-item test was prepared by selecting certain items from the sections on Scientific Method of four forms of the American Chemical Society Cooperative Examination in General Chemistry (Appendix III), and it was administered and scored according to the instructions with the original forms of the examination. The purpose of this test was to measure the ability of students to apply the scientific method of problem-solving to situations in the fields of general chemistry.

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[48] Kenneth E. Anderson, The Fourth Mental Measurements Yearbook, edited by Oscar E. Buros (New Jersey, 1953).

[49] Paul L. Dressel, private communication, Michigan State College, March 15, 1954.

The two examinations, namely, A Test on Science Reasoning and Understanding and A Test on the Scientific Method, were used in the present study to compare the mean ability of the students in the experimental and control groups in their use of the scientific method in solving problems both in the field of general chemistry and in the area of physical science outside this specific field.

Departmental Intra-semester Tests. Three intra-semester, fifty-item, objective-type tests were prepared by the course instructors and administered during the second-semester course. These tests (Appendices IV, V, and VI) designed to determine the degree to which the students had mastered the course content, were assumed to have curricular validity. Since the students' scores on these tests were used to determine the achievement of the students in the second-semester course, these scores also indicated how well the first-semester course had prepared the students for the learning materials of the second-semester course.

Statistical Methods

The following statistical methods were employed in this study for the purpose of comparing the results of the evaluating devices and thereby comparing the results of the two subject-matter organizations: (1) the arithmetic means were calculated from frequency distributions of the test scores as a measure of central tendency, (2) the standard deviation of the means were calculated as a measure of variability, (3) the standard errors of the means were calculated as a measure of the reliability of the means, (4) the product-moment method was used to calculate the coefficient of correlation between scores on the various tests,[50] (5) the significance of obtained differences between statistics was tested against a null hypothesis, namely, that the true difference was zero, and (6) the test of significance used in testing the null hypothesis was the critical ratio.[51] The critical ratio was defined as the quotient obtained by dividing the obtained difference between two statistics by the standard error of the difference between the two statistics. The standard error of the difference between statistics was calculated by the method described by Lindquist.[52] This method was applicable because the samples were matched by means and standard deviations upon the American Council on Education Psychological Examination scores. The hypothesis held for this study was that achievement is independent of subject-matter organization or that the mean achievement of the students in the two groups was equal. The five per cent level of confidence was used in attempting to refute the null hypothesis, that the outcomes of the two instructional organizations were the same.

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[50] Henry E. Garrett, Statistics in Psychology and Education (New York, 1953), pp. 134-139.

[51] Ibid., p. 215.

[52] E. F. Lindquist, "The Significance of a Difference between 'Matched' Groups," The Journal of Educational Psychology, XXII (March, 1931), 197-204.

CHAPTER THREE

ANALYSIS OF RESULTS

The results of the analysis of the data collected through the application of the foregoing procedure to the problem pertaining to the relative effectiveness of the principles and conventional subject-matter organizations for the teaching of general chemistry are herewith presented.

Equivalency of the Groups

Since in the present study the experimental population was employed for two semesters and since many students were lost to the study for the second semester the results obtained by applying the procedure used in formulating the experimental and control groups are presented according to: (1) the first-semester samples and (2) the second-semester samples.

First-Semester Samples. The 530 student experimental population was distributed in the four theory sections so that two nearly numerically equal groups of students were formed. Table 1 shows these data.

Table 1

Distribution of the First-Semester Experimental Population by Theory Sections

Control Group		Experimental Group	
Theory section three -	139	Theory section one -	126
Theory section four -	126	Theory section two -	134
Total	265	Total	260

Only nineteen women were represented in the experimental population, representing a total of seven in the experimental group and twelve in the other sample.

The data presented in Tables 2 and 3 showed that the two samples of students were statistically equally distributed among the various schools and among the various student classifications.

Table 2

Distribution of the First-Semester Experimental Population
Among Various Schools

School	Control Group	Experimental Group
Engineering	183	199
Arts and Sciences	69	54
Agriculture	7	0
Commerce	4	5
Education	2	1
Home Economics	0	1
Totals	265	260

Table 3

Distribution of the First-Semester Experimental Population
by Student Classification

Classification	Control Group	Experimental Group
Freshmen	215	202
Sophomores	36	46
Juniors	9	9
Seniors	3	3
Totals	265	260

It can be seen from the data summarized in Table 4 that the two population samples, as measured by the <u>American Council on Education Psychological Examination</u>, were equivalent with respect to scholastic aptitude. Since a critical ratio of 1.96 would be required before one could be confident that the obtained difference between the means was a true difference, the critical ratio of 0.82 was not statistically significant at the five per cent level of confidence.

Table 4

A Comparison of the Scores Made by the Two First-Semester Groups on the *A.C.E. Psychological Examination*

Statistic	Control Group	Experimental Group
Number of subjects	239	239
Mean	100.11	98.61
Standard Deviation	19.56	20.30
Standard error of the mean	1.26	1.31
Difference between means	1.50	
Standard error of the difference between means	1.82	
Critical ratio	0.82	

Second-Semester Samples. Approximately forty per cent of the students were lost to the study by the close of the second-semester enrollment period. One hundred and forty-nine of the 237 eligible control-group students and 151 of the 233 eligible experimental-group students registered for the second-semester course. The data given in Table 5, which is arranged according to the first-semester marks, showed the total population loss of each sample.

The data presented in Tables 6 and 7 showed that the second-semester population samples, like the first-semester samples, were statistically equally distributed among the various schools and among the various student classifications.

Table 5

Number of Students Lost to the Second-Semester Samples by First-Semester Marks

Course Grade	Control Group	Experimental Group
A	3	5
B	16	8
C	44	42
D	25	27
F	16	15
W	6	7
WF	0	4
I	6	1
Total loss	116	109

Table 6

Distribution of the Second-Semester Experimental Population Among Various Schools

Schools	Control Group	Experimental Group
Engineering	100	117
Arts and Sciences	42	31
Agriculture	5	0
Commerce	2	2
Education	0	1
Totals	149	151

Table 7

Distribution of the Second-Semester Experimental Population by Student Classification

Classification	Control Group	Experimental Group
Freshmen	123	123
Sophomores	20	22
Juniors	5	5
Seniors	1	1
Totals	149	151

The results of the analysis of the data summarized in Table 8 showed that the difference between mean scores on the American Council on Education Psychological Examination in favor of the control group during the first semester was slightly larger for the second-semester samples. The critical ratio of the difference between the means of 1.22, although not significant at the five per cent level of confidence, indicated there were 78 chances in 100 that the true difference was larger than zero. The critical ratio of the difference between the standard deviations of 1.55 was not significant at the five per cent level of confidence. An inspection of the data presented in Table 4 and Table 8 further shows that the second-semester experimental sample was more variable than the corresponding first-semester sample but that the second-semester control sample was more homogeneous than the corresponding first-semester sample.

Table 8

A Comparison of the Scores Made by the Two Second-semester Groups on the A.C.E. Psychological Examination

Statistic	Control Group		Experimental Group
Number of subjects	140		140
Mean	102.21		99.36
Standard Deviation	17.60		21.21
Standard error of the mean	1.49		1.79
Difference between means		2.85	
Standard error of the difference between means		2.33	
Critical ratio		1.22	

Although the final examination was the only completely common test used during the first semester, the distribution of the first-semester course marks attested the equivalency of the second-semester samples. Table 9 shows these data.

Table 9

Distribution of Course Grades for the Second-Semester Groups in Chemistry 114

Course Grade	Control Group	Experimental Group
A	20	22
B	43	41
C	71	69
D	15	19
Totals	149	151

From the results of the analysis of data just presented, it became evident that the experimental and the control group of students, as employed in this study, were statistically equivalent with respect to the criteria used. The results presented appeared to indicate that these groups of students were samples selected at random from a larger population of students.

Results of the Teaching Experiment

For purposes of evaluation in this study, the values to be derived from a study of general chemistry were divided roughly into the following areas: (1) general knowledge and information, (2) application of principles, and (3) the methodology of science. The area of general knowledge and information was principally concerned with the meaning of important terms, knowledge of important laws, concepts, and principles, and general knowledge of the physical and chemical properties of important elements and their compounds. The second area was concerned with the application of important theories and principles of general chemistry in solving simple problems. In the area of methodology, the chief concern was to give the student an opportunity to develop the abilities identified with the use of the scientific method of problem-solving. These abilities include, among others, the ability to recognize and analyze problems in terms of the student's understanding of the broad generalizations of the science; the ability to distinguish between experimental evidence and theoretical interpretations; and the ability to interpret data and to draw valid conclusions therefrom. An attempt was made to evaluate the achievement of these abilities, both in the subject matter field of general chemistry and in the broader area of physical science outside this specific field.

A summary of the results of this analysis of the data, collected through the administration of the achievement tests described earlier, appears in the following sections.

A.C.S. Cooperative Examination, Form G. The data collected through the administration of the American Chemical Society Cooperative Examination in General Chemistry, Form G, as a term-end examination for the first-semester course, were analyzed; and a summary of the results follow.

1. The product-moment coefficient of correlation between student scores on the American Council on Education Psychological Examination and student scores on this examination were: for students in the experimental group, .49; and for students in the control group, .62.

2. The two subject-matter organizations were equally effective in teaching the overall objectives of the first-semester course as their attainment was measured by this examination. The difference between mean scores of 0.40 in favor of the control group was not statistically significant at the five per cent level of confidence.

The critical ratio of this difference was 0.48, which indicated that there were 67 chances in 100 that the difference was a true difference. The results of the analysis of these data are summarized in Table 10.

For the purpose of comparison, Table 11 shows these data as they applied to the second-semester samples only. The difference between the mean scores of 0.46 in favor of the control group gives a critical ratio of 0.44, which indicated there were 66 chances in 100 that the true difference was greater than zero.

Table 10

A Comparison of the Scores Made by the Two First-Semester Groups on the _A.C.S. Cooperative Examination in General Chemistry, Form G_

Statistic	Control Group		Experimental Group
Number of subjects	246		241
Mean	25.27		24.87
Standard Deviation	10.79		11.30
Standard error of the mean	0.69		0.73
Difference between means		0.40	
Standard error of the difference between means		0.83	
Critical ratio		0.48	
Coefficient of correlation between ACE scores and the _ACS Examination, Form G_		.56	

Table 11

A Comparison of the Scores Made by the Two Second-Semester Groups on the _A.C.S. Cooperative Examination in General Chemistry, Form G_

Statistic	Control Group		Experimental Group
Number of subjects	149		151
Mean	27.56		27.10
Standard Deviation	10.50		11.40
Standard error of the mean	0.86		0.93
Difference between means		0.46	
Standard error of the difference between means		1.05	
Critical ratio		0.44	

3. The two organizations appeared to be equally effective in teaching general knowledge and information, the application of principles, and the scientific method as the attainment of these objectives were measured by each of the different parts of this examination. All the differences between means favored the control group, but none of them were statistically significant at the five per cent level of confidence. The results of the analysis of the data for each of the parts are summarized in Table 12.

Table 12

A Comparison of the Scores Made by the Two First-semester Groups on the Different Parts of the A.C.S. Cooperative Examination in General Chemistry, Form G

Statistic	Control Group		Experimental Group
Part I General Knowledge and Information (25 questions)			
Mean	7.83		7.61
Range	0-16		0-20
Difference between means		0.22	
Part II Application of Principles (25 questions)			
Mean	6.09		5.81
Range	0-19		0-16
Difference between means		0.28	
Part IV Scientific Method (20 questions)			
Mean	5.06		4.95
Range	0-15		0-14
Difference between means		0.11	

4. The comparative value of the two organizations as they relate to scholastic aptitude was determined by comparing the scores made by the upper one-third of the experimental-group students ranked on the basis of the American Council on Education Psychological Examination

scores with the scores made by the upper one-third of the control-group students ranked in the same manner. This was also done for the lowest one-third of the students from the two samples. The results of the analysis of these data showed the two organizations to be equally effective when applied with students of high or low scholastic aptitude. The results of the analysis of these data appear in Table 13.

Table 13

A Comparison of the Mean Scores Made by the Upper One-third and the Lower One-third of the Two Second-semester Groups on the A.C.S. Cooperative Examination in General Chemistry, Form G

Division	Mean Scores Control Group	Mean Scores Experimental Group
Upper one-third	32.71	31.17
Lower one-third	22.00	21.82
Average	27.35	26.49

A Test on Science Reasoning and Understanding. The correlation coefficients between student scores on this examination and student scores on the American Council on Education Psychological Examination were: for the experimental group, .44, and for the control group, .42. The analysis of the data collected through the administration of this test gave results which indicated that the two organizations were equally effective in teaching the scientific method as measured by this test. The difference of 1.15 between the mean scores in favor of the experimental group gave a critical ratio of 1.74, which indicated that there were 90 chances in 100 that the difference between the mean scores was a true difference. Table 14 summarizes these data.

Table 14

A Comparison of the Scores Made by the Two Second-semester Groups on the *Test on Science Reasoning and Understanding, Physical Sciences, Form A*

Statistic	Control Group	Experimental Group
Number of subjects	142	149
Mean	16.10	17.22
Standard Deviation	6.42	5.94
Standard error of the mean	0.55	0.94
Difference between means	1.15	
Standard error of the difference between means	0.66	
Critical ratio	1.74	
Coefficient of correlation between *ACE* scores and *Science Reasoning and Understanding* test scores	.43	

A Test on the Scientific Method. The results of the analysis of the data collected through the use of the examination, A Test on the Scientific Method, showed that the two organizations appeared to be equally effective in teaching the scientific method as the outcomes of this teaching were measured by this test. Here the difference of 0.77 between the mean scores gave a critical ratio of 0.94 in favor of the experimental group, which indicated that there were 83 chances in 100 that the true difference was greater than zero.

The product-moment coefficient of correlation between the scores of the experimental-group students and the control-group students on the *American Council on Education Psychological Examinatinn* and this test were .53 and .34 respectively. The results of the analysis of these data are summarized in Table 15.

Table 15

A Comparison of the Scores Made by the Two Second-Semester Groups on A Test on the Scientific Method

Statistic	Control Group	Experimental Group
Number of subjects	142	151
Mean	19.03	19.80
Standard Deviation	7.53	8.16
Standard error of the mean	0.63	0.66
Difference between means	0.77	
Standard error of the difference between means	0.82	
Critical ratio	0.94	
Coefficient of correlation between <u>ACE</u> scores and the Test on the Scientific Method test scores	.44	

<u>Departmental</u> <u>Intra-semester</u> <u>Examinations</u>. The results of the analysis of the data collected through the administration of the three departmental intra-semester examinations showed that the two subject-matter organizations appeared to be equally effective in preparing the students for the second-semester course. The differences between mean scores for the two groups on the first, second, and third intra-semester examinations gave critical ratios of 0.87, 0.20, and 0.66, respectively, in favor of the experimental group; however, none of these differences were statistically significant at the five per cent level of confidence. Tables 16, 17, and 18 present a summary of these data.

Table 16

A Comparison of the Scores Made by the Two Second-Semester Groups on the First Departmental Intra-semester Examination

Statistic	Control Group		Experimental Group
Number of subjects	148		151
Mean	24.51		25.28
Standard Deviation	7.30		7.75
Standard error of the mean	0.61		0.63
Difference between means		0.77	
Standard error of the difference between means		0.88	
Critical ratio		0.87	

Table 17

A Comparison of the Scores Made by the Two Second-Semester Groups on the Second Departmental Intra-semester Examination

Statistic	Control Group		Experimental Group
Number of subjects	149		151
Mean	27.59		27.79
Standard Deviation	9.72		9.66
Standard error of the mean	0.82		0.79
Difference between means		0.20	
Standard error of the difference between means		1.14	
Critical ratio		0.20	

Table 18

A Comparison of the Scores Made by the Two Second-Semester Groups on the Third Departmental Intra-semester Examination

Statistic	Control Group		Experimental Group
Number of subjects	141		148
Mean	17.47		18.20
Standard Deviation	7.83		8.40
Standard error of the mean	0.66		0.69
Difference between means		0.63	
Standard error of the difference between means		0.96	
Critical ratio		0.66	

American Chemical Society Cooperative Examination, Form K. The analysis of the data collected by the administration of the *American Chemical Society Cooperative Examination in General Chemistry, Form K,* indicated that the two organizations were equally effective in achieving the overall objectives of the courses. The difference of 0.46 between mean scores for the two groups in favor of the experimental group was not significant at the five per cent level of confidence. The critical ratio of the difference of 0.30 indicated there were 62 chances in 100 that the difference represented a true difference. The product-moment coefficient of correlation between student scores on this examination and the *American Council on Education Psychological Examination* were: for the experimental group, .45 and for the control group, .33. The two organizations appeared to be equally effective in teaching either general knowledge and information or the application of principles since there was no significant difference between the mean scores for the two groups on the two parts of this examination. These data are presented in summary form in Tables 19 and 20.

Table 19

A Comparison of the Scores Made by the Two Second-Semester Groups on the *A.C.S. Cooperative Examination in General Chemistry, Form K*

Statistic	Control Group	Experimental Group
Number of subjects	141	150
Mean	29.84	30.20
Standard Deviation	13.80	14.90
Standard error of the mean	1.16	1.22
Difference between means		0.46
Standard error of the difference between means		1.54
Critical ratio		0.30
Coefficient of correlation between *ACE* scores and the *A.C.S. Cooperative Examination in General Chemistry, Form K*		.39

Table 20

A Comparison of the Scores Made by the Two Second-semester Groups on Different Parts of the *A.C.S. Cooperative Examination in General Chemistry, Form K*

Statistic	Control Group		Experimental Group
Part I General Knowledge and Information (30 questions)			
Mean	10.78		10.85
Standard Deviation	5.24		5.84
Difference between means		0.07	
Part II Application of Principles (50 questions)			
Mean	18.95		19.20
Standard Deviation	5.88		6.09
Difference between means		0.25	

Other Results. The distribution of course grades for the second-semester course (Chemistry 124) showed that twelve of the 149 second-semester, control-group students failed to complete the course satisfactorily, and four of the 151 second-semester, experimental-group students failed to complete the course satisfactorily. This difference was largely accounted for by the withdrawal of seven students from the course. The experimental group had a higher percentage of A grades, B grades, and D grades, which showed this group to have been more variable than the control group. Table 21 presents these data.

An inspection of the data pertaining to the number of students lost to each of the groups (Tables 5 and 20) showed that the control group lost a total of 128 students and the experimental group lost a total of 113 students. These losses included all students who failed to complete the year's work satisfactorily. The excessive loss to the control group, which included seven students during the first semester

Table 21

Distribution of Course Grades in Chemistry 124

Course Grade	Control Group	Experimental Group
A	10	16
B	33	42
C	69	57
D	25	32
F	5	3
W	6	0
WF	1	0
I	0	1
Totals	149	151

eight students during the second semester, was nearly equally distributed between the two semesters. These data showed that the control group lost 4.83 per cent more students than were lost to the experimental group. This difference, however, was not statistically significant at the five per cent level of confidence.

Summary of Results

The principal result of this study with respect to the relative effectiveness of the two subject-matter organizations for presenting general chemistry to students at The Oklahoma Agricultural and Mechanical College was that the two organizations were equally effective when employed as they were in this investigation. The first-semester results, as well as the American Council on Education Psychological Examination scores, favored the control group, whereas all the second-semester results favored the experimental group. None of the differences, however, were found to be statistically significant at the five per cent level of confidence. Other findings are summarized as follows:

1. The two organizations were equally effective in teaching knowledge and information as measured by the American Chemical Society Cooperative Examination in General Chemistry and the intra-semester departmental examinations.
2. The two organizations were likewise equally effective in teaching the application of principles as measured by the American Chemical Society Cooperative Examination in General Chemistry.
3. The two organizations appeared to be equally effective in teaching the abilities associated with the scientific method of problem-solving as measured by the tests used in this study.
4. From the results of the analysis of the data summarized in Table 13, it appeared that neither of the two organizations favored the more intelligent or less intelligent students when ranked on the basis of the American Council on Education Psychological Examination scores.
5. From the results of the data presented in Tables 5 and 21 pertaining to the number of students lost to the experiment from each of the samples, there appeared to be a tendency for more students to be lost to the control group. These data showed that the control group contained 128 students who failed to complete the year's work satisfactorily, while the experimental group had a total of 113 students in this category. Thus the control group lost 4.83 per cent more students than were lost to the experimental group. This difference, however, was not statistically significant at the five per cent level of confidence.

6. Apparently each of these organizations was well adapted to its purpose, since neither of them presented any marked difficulty in preparation and administration.

CHAPTER FOUR

SUMMARY AND CONCLUSIONS

Summary. This investigation, an experimental study, was designed to determine the relative effectiveness of two different organizations for the subject matter of the first-semester general chemistry course at the college level. The problem of this study was to ascertain which, if either, of the two organizations of the course content contributed more to the learning outcomes of the general chemistry course.

The equivalent or parallel-group method of research was employed, and the two subject-matter organizations were compared with respect to their effectiveness in teaching general knowledge and information, the application of principles, and the ability to use the scientific method of problem-solving. The 530 students who registered for the beginning general chemistry course during the first semester of the school year 1953-1954 were divided by the enrollment procedure into an experimental and a control group to serve as the experimental samples. Each of the experimental samples was composed of two classes of students, who were taught by two instructors. Each of the instructors taught one control and one experimental class. Two groups of instructors taught the quiz-laboratory classes.

In the instruction of the control group the conventional organization of the course content was used. In this organization almost one semester of general chemistry was studied before the atomic structure theory was introduced into the course. In the instruction of the experimental group, a newer organization was employed, which presented atomic structure theory early in the course and used it as a basis for other materials of the course.

The examinations used for the purpose of evaluating the outcomes of this study are listed as follows: two forms of the American Chemical Society Cooperative Examination in General Chemistry, namely, Form C and Form K; a Test on Science Reasoning and Understanding, Physical Sciences, Form A; a Test on the Scientific Method; and three departmental intra-semester examinations. The means of the two groups on the various tests were found and the significance of the difference calculated using the critical ratio.

Results. The first-semester results, as well as the American Council on Education Psychological Examination scores, favored the control group, but all the second-semester results favored the experimental group. None of the differences, however, were found to be statistically significant at the five per cent level of confidence.

Conclusions. Since none of the differences were significant at the five per cent level of confidence, the null hypothesis was not rejected and it is concluded that neither subject-matter organization has been shown to be superior for instructional purposes.

Implications. From this conclusion it was implied that either organization can be used as effectively as the other in presenting the subject matter of the beginning general chemistry course. Somewhat more to the point, the implications are that general knowledge and information, the application of principles, and the ability to use the scientific method of problem-solving can be taught as effectively by using the principles organization as by using the conventional organization. The usual practice, therefore, of presenting approximately one semester of general course content before introducing the atomic structure theory is no more warranted for a similar population than is the early introduction of atomic theory.

Suggestions for Further Study. Several aspects of the results of this investigation suggest precautions for, and areas of, additional research.

1. In view of the inconclusive results, the advantage enjoyed by the control group during both semesters on the American Council on Education Psychological Examination could have been a critical factor. A more rigid control of equivalence or a different statistical technique would be desirable in future studies.

2. From the results of this study the conclusion was reached that either of these organizations could be used effectively by the instructors while teaching the students in the sequence of courses dealt with in this investigation. These findings, however, would apply only to the general population from which these students and instructors came. Additional research would be necessary to ascertain whether these findings have more general applicability. This, of course, would involve other schools with different courses, instructors, and students. It may be that, if the students were accustomed to the experimental organization, the experimental group would have done better.

3. The fact that all the (non-significant) differences obtained between the various statistics favored the control group during the first semester, whereas they all favored the experimental group during the second semester suggests that a study covering a longer period of time might result in significant findings in favor of the

experimental group.

4. The instructors in this study were, of course, more familiar with the conventional organization. This suggests that a study employing teachers more experienced in handling the principles organization might also result in findings more favorable to the experimental group.

5. In this study syllabi were used in an attempt to minimize the variable favoring the control group introduced through the use of the conventional textbook. This suggests that a study employing two textbooks, one following the conventional and the other the principles organization, might, again, result in findings more favorable to the experimental group.

6. The matter of the transfer of training suggests another area worthy of investigation. A study designed to determine whether the use of either of the two organizations would be superior in helping the student to generalize his learning from one content field to another and to apply this knowledge to broader areas would possibly make a valuable contribution. Similarly, the matter of retention, or permanence, of learning suggests another problem of some importance.

7. The factor of student preference in the matter suggests another area for investigation. A study designed to determine which, if either, of the two organizations would be preferred by students might be desirable. The findings of such a study would have importance because the use of the more desirable organization would possibly result in attracting more students to chemistry as a profession.

BIBLIOGRAPHY

American Association for the Advancement of Science. "Committee Report on the Place of Science in Education." School Science and Mathematics, XXVIII (1928).

Anderson, Kenneth E. The Fourth Mental Measurements Yearbook, edited by Oscar E. Buros. New Jersey: The Gryphon Press, 1953.

Babitz, Melton and Noel Keys. "An Experiment in Teaching Pupils to Apply Scientific Principles." Science Education, XXIII (December, 1939), 667-670.

Boeck, Clarence H. "The Inductive-Deductive Compared to the Deductive-Descriptive Approach to Laboratory Instruction in High School Chemistry." Journal of Experimental Education, XIX (September, 1950-June, 1951), 247-253.

Bond, Austin D. "An Experiment in the Teaching of Genetics, with Special Reference to the Objectives of General Education." Contribution to Education No. 797. New York: Bureau of Publications, Teachers College, Columbia University, 1940.

Brown, Stanley B. "Trends in Science Education - 1953." The Science Teacher, XXI (March, 1954), 84-85.

Campbell, J. A. "Structural Chemistry." Journal of Chemical Education, XXV (October, 1948), 558-562.

Coles, James S., Leallyn B. Clapp, and Robert P. Epple. "A New Program for Teaching the Fundamentals of Chemistry in College." Journal of Chemical Education XXVI (January, 1949), 10-14.

Colyer, Luther M. and Kenneth E. Anderson. "A Comparison of Two Methods of Teaching Formula Writing in High School Chemistry." School Science and Mathematics, LII (January, 1952), 50-59.

Cunningham, Harry A. "Lecture Demonstration Versus Individual Laboratory Method in Teaching - A Summary." Science Education, XXX (March, 1946), 70-82.

Curtis, Francis D. A Digest of Investigations in the Teaching of Science in the Elementary and Secondary Schools. Philadelphia: P. Blakiston's Son and Co., 1926. 341 p.

______. Second Digest of Investigations in the Teaching of Science. Philadelphia: P. Blakiston's Son and Co., 1931. 424 p.

______. *Third Digest of Investigations in the Teaching of Science.* Philadelphia: P. Blakiston's Son and Co., 1939. 419 p.

de Milt, Clara. "The Value of the Historical Approach in the Teaching of General Chemistry." *Journal of Chemical Education*, XXVI (January, 1949), 39-42.

Deridder, Lawrence M. "Relationship Between Gross Scores on the A.C.E. and Academic Success." *Journal of Educational Research*, XLVI (January, 1953), 353-358.

Dunlap, Jack W. *The 1940 Mental Measurements Yearbook*, edited by Oscar K. Buros. Arlington, Virginia: The Gryphon Press, 1945.

Frutchey, F. P. "Retention in High School Chemistry." *Educational Research Bulletin*, Ohio State University, XVI (February, 1937), 34-37.

Garrett, Henry E. *Statistics in Psychology and Education.* New York: Longmans, Green and Co., 1953.

Good, C. V., A. S. Barr and D. E. Scates. *The Methodology of Educational Research*, New York: Appleton Century Co., 1941.

Grant, Julius, editor. *Hackh's Chemical Dictionary*, Third Edition. Philadelphia: The Blakiston Company, 1944.

Hauser, Ernst A. "This Nation Will Need More Chemists After the War." *Journal of Chemical Education*, XXI (October, 1943), 513-515.

Horton, Ralph E. "Measurable Outcomes of Individual Laboratory Work in Highschool Chemistry." *Contributions to Education* No. 303. New York: Bureau of Publications, Teachers College, Columbia University, 1928.

Johnson, Palmer O. *Statistical Methods in Research.* New York: Prentice-Hall, Inc., 1949.

Kilgore, W. A. "Identification of Ability to Apply Principles of Physics." *Contributions to Education* No. 840. New York: Bureau of Publications, Teachers College, Columbia University, 1941.

Kruglak, Haym. "A Comparison of the Conventional and Demonstration Methods in the Elementary College Physics Laboratory." *Journal of Experimental Education*, XX (March, 1952), 293-300.

Kurtz, Albert K. and Harold A. Edgerton. *Statistical Dictionary of Terms and Symbols.* New York: John Wiley and Sons, Inc., 1939.

Lindquist, E. F. "The Significance of a Difference between 'Matched' Groups." *The Journal of Educational Psychology*, XXII (March, 1931), 197-204.

Mack, Edward Jr., *et al*. *Textbook of Chemistry*. Dallas: Ginn and Company, 1949.

Martinette, Sister Mary. "The Presentation of Atomic Structure to College Freshmen." *School Science and Mathematics*, XL (December, 1940), 808-814.

Mason, John M. "An Experimental Study in the Teaching of Scientific Thinking in Biological Science at the College Level." *Science Education*, XXXVI (December, 1952), 270-284.

McCall, William A. *How to Experiment in Education*. New York: The Macmillan Company, 1926.

Nash, Leonard K. "An Historical Approach to the Teaching of Science." *Journal of Chemical Education*, XXVIII (March, 1951), 146-151.

National Education Association. *Cardinal Principles of Secondary Education*. Washington, D. C.: Government Printing Office, 1918.

National Education Association. *Reorganization of Science in Secondary Schools*. Washington, D. C.: Government Printing Office, 1920.

National Society for the Study of Education. *Science Education in*
A *American Schools*. Forty-sixth Yearbook, Part I. Chicago: The University of Chicago Press, 1947.

______. *A Program for Teaching Science*. Thirty-first Yearbook, Part I. Bloomington: Public School Publishing Co., 1932.

Noll, Victor H. *What Do You Think*? New York: Bureau of Publications, Teachers College, Columbia University, 1935.

Osborne, T. R. *et al*. "The Differential Prediction of College Marks by ACE Scores." *Journal of Educational Research*, XLIV (October, 1950), 107-115.

Payne, Virgil Frances. "The Lecture Demonstration and Individual Laboratory Methods Compared: (1) The Literature, (2) The Distribution of Time, (3) Experimental." *Journal of Chemical Education*, IX (May, 1943), 932-937.

Perry, Raymond J. "A Re-evaluation of Teaching First-Year College Chemistry." *Journal of Chemical Education*, XXII (October, 1945), 497-500.

Phelan, Earl W. "The 1936-1937 College Chemistry Testing Program." *Journal of Chemical Education*, XIV (December, 1937), 586-590.

Progressive Education Association. *Science in General Education*. New York: C. Appleton-Century Co., 1938.

Sisler, Harry H. "Why Should Theoretical Principles be Used in Teaching Elementary Chemistry?" *Journal of Chemical Education*, XXV (October, 1948), 562-565.

______, and Calvin A. Vanderwerf. "Modern Theory: A Tool in Teaching Elementary College Chemistry." *Journal of Chemical Education*, XX (October, 1943), 479-483.

Smith, Donald E. and Marvin D. Glock. "Measuring Knowledge and Applications: An Experimental Investigation." *Journal of Experimental Education*, XXI (September, 1952–June, 1953), 327-331.

Smith, Otto M. "Accepted Objectives in the Teaching of General College Chemistry." *Journal of Chemical Education*, XII (April, 1935), 180-183.

Standen, Anthony. "Can Elementary Chemistry Teaching Be Logical?" *Journal of Chemical Education*, XXII (November, 1945), 554-557.

Thelen, Herbert A. "A Methodological Study of the Learning of Chemical Concepts of Certain Abilities to Think Critically in Freshmen Chemistry." *Journal of Experimental Education*, XIII (September, 1944), 53-75.

Tyler, Ralph W. "Permanence of Learning." *Journal of Higher Education*, IV (April, 1933), 203-204.

Wakeham, George. "A Logical Chemistry Course?" *Journal of Chemical Education*, XXII (November, 1945), 536.

Ward, William E. "An Experimental Study of Two Methods of Teaching Chemistry in Senior High School." *Journal of Experimental Education*, XI (September, 1942), 69-80.

Wildman, Ernest A. "The Need of Modernizing the General Course." *Journal of Chemical Education*, XII (January, 1935), 11-16.

Wispe, Lauren G. "Evaluating Section Teaching Methods in the Introductory Course." *Journal of Educational Research*, XLV (November, 1951), 161-186.

APPENDIX

APPENDIX I
Conventional Syllabus

Chemistry 114, Sections 3 and 4

Group A

Syllabus

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Chemistry 114
First Semester 1953-54

A&B

Lecture Text: Textbook of Chemistry by Mack, Garrett, Haskins and Verhoek
Laboratory Manual: Semimicro Laboratory Exercises in General Chemistry
Second Edition: Burrows, Arthur and Smith.

General Information for the Student

The purpose of this course is to give the student an understanding of the principles, theories and laws of chemistry. The course will consist of lectures with demonstrations, recitations, and laboratory work.

The students should:

(1) learn to use the text book as a reference guide in both the theory and laboratory portions of the course.
(2) read over the topics assigned for lecture before going to class, and the topics pertaining to the laboratory work before going to the laboratory.
(3) make notes during lectures to supplement the syllabus, underscoring important points.
(4) as soon as possible after each lecture, read over the syllabus and lecture notes and if there are any points not clearly understood look them up in the text using the index if necessary.
(5) prepare a list of questions, over any material which is not well understood before attending the discussion period.
(6) ask the instructor for individual help on any points that are not clearly understood after the discussion period. (It is not recommended that the student skip over such points, for they may have an important bearing on later subject matter.)
(7) at frequent intervals (possibly once each week) review the syllabus and lecture notes for the material covered to date.
(8) follow the instructions given on the first pages of the laboratory manual in performing the laboratory work unless other instructions are given by the laboratory instructor.

Examinations and Grading

(1) Three one-hour objective-type examinations will be given at previously arranged times during the course. (Possibly during the 5th, 10th and 15th weeks.) These shall account for 50% of the course grade.
(2) Short tests (or quizzes) may be given from time to time during either the discussion periods or the lecture period. Laboratory reports will be collected at not infrequent times to be graded and the quiz-laboratory instructor may use what other methods desired to secure marks on student attitudes, laboratory techniques, etc.
   The short tests, laboratory report grades and teacher estimate grade, etc. will be combined to make up 30% of the course grade.
(3) A two-hour objective-type final examination will be used to account for the other 20% of the course grade.

Chemistry 114 is the first course of a two semester sequence and will be followed by Chemistry 124. Chemistry 124 is a continuation of Chemistry 114 and the same text book and laboratory manual are used for both courses. A comprehensive examination over both courses will be given as the final examination for Chemistry 124.

Chapters I and II

A & B

Introduction - The Science of Chemistry
Matter and Energy

(Students read and study Chapters 1 and 2.)

. Definitions

1. Chemistry - is the branch of natural science which deals with the properties of matter, the changes which matter undergoes, and the laws and theories that govern the changes.
2. Science - is knowledge that has been classified and arranged into a systematic arrangement by subjects.
3. Matter - is defined as any substance that has mass and occupies space (has size) and is capable of undergoing change.
    a. Mass - is the quantity of matter in a body.
    b. Weight - is the attractive force (gravity) which the earth exerts on a body.

Note - For all practical purposes on the earth's surface mass and weight can be used interchangeably.

    c. Law of Conservation of Mass - In ordinary chemical reactions mass is neither created nor destroyed.

4. Substance - The term substance is applied to homogeneous bodies: those which are alike in all their parts.
5. Substances are identified by their properties - distinguishing characteristics such as color, odor, taste, melting point, boiling point, density and the like are physical properties.
6. Chemical Properties - are the essential properties of a substance; if they are changed a new substance is formed.
    a. Chemical Reactions - are the processes that convert substances into other substances (changing chemical properties).
7. Energy - the ability to do work. Work is done when a body is moved through space.

Note - In all physical or chemical changes energy is involved, either energy is applied to cause the change or is released by the change. (familiar manifestations of energy are: heat, light, electricity, etc.)

    a. Kinetic energy - the energy of moving bodies.
    b. Potential energy - energy due to position. (water behind the dam, etc.)
    c. Chemical energy - energy that is stored in a substance due to its chemical composition.
    d. Law of conservation of energy - energy like mass, in all ordinary transformation is neither created nor destroyed.

8. Scientific Laws - The statement of a generalization or a general principle based on a common mode of action of natural things.

. History of Chemistry

(see pages 3-6)

These are only a beginning of terms and words to be defined during the course. he student must strive to understand each work or term as it is introduced, and o review them often enough to retain a working knowledge of them. The student ust be very certain at this time that a good foundation is secured upon which to ase the work that follows.

C. The Scientific Method A & B

1. The accumulation of facts by:
    - a. observation - observing the material in its present form.
    - b. experimentation - this is done by changing the conditions under which the material exists and observing the result. Thus the scientist experiments to establish facts which when established can be classified into general principles or laws.
2. The Use of Laws
    - a. Hypothesis - a tentative but reasonable explanation for observed facts. An attempt to answer the question why? A hypothesis can be tested by experimentation. If it meets the tests the hypothesis is advanced to a theory if it fails to meet the test it is either modified or discarded.
3. Theory - a tested and approved hypothesis. A theory attempts to coordinate into a concise, unified picture all the facts related to a given subject, also certain assumptions may be made in formulating a theory.

D. Units of Measure Used in Chemistry

(see text pages 11 and 842; also inside front cover of laboratory manual and Experiment 1.) Centimeter - gram - second (CGS) system is used in all countries of the world for scientific work.

1. Weight and Volume
    - a. unit of weight, one gram (g.)
    - b. unit of volume, one liter (l.)
2. Length
    - a. units of length are the meter (m), centimeter (cm) and millimeter (mm)
3. Time
    - a. seconds (sec), minutes, etc.
4. Temperature
    - a. Centigrade - usually used in science - degrees centigrade (°C.)
    - b. Fahrenheit - household use - degrees Fahrenheit (°F.)
    - c. Absolute scale of temperature to be considered later

Note - A few important values for converting from one system to another should be memorized by the student. For example - 1 inch = 2.54 cm.; 1 meter = 39.37 inches; 1 ounce = 28.34 grams; 1 liter = 1.0567 quarts.

. Other terms defined

1. Calorie - (page 14) a unit of measuring heat. The calorie is that quantity of heat which will raise the temperature of one gram of water 1°C. at 15°.
2. Atom - minute particles of matter that do not, during natural processes, undergo division. The smallest part of an element to take part in a chemical reaction.
3. Molecule - formed when two or more atoms combine. Molecules are the smallest particles of a chemical substance that can exist alone and which have the characteristic composition and properties of the substance.
4. Density - The quantity of matter per unit of volume. The ratio of the mass of a homogeneous portion of matter to its volume.
5. Specific Gravity - the ratio of the weight of any volume of a substance to the weight of an equal volume of some substance taken as a standard or unit as, usually, water for solids and liquids, and air or hydrogen for gases.

Note - When water is used as the standard its maximum density (at 4°C.) is assigned a value of one. The density of gases usually refer to the weight in grams per liter and the density of solids and liquids refers to the weight in grams per cubic centimeter (g/cc).

Chapter 3
Elements, Compounds, Mixtures, Solutions
The Atomic Theory

. Four Groups of Substances

1. *Elements* - are substances incapable of being broken down into simpler substances by ordinary chemical means. The molecules of an element consist of only one kind of atoms. The elements are the fundamental substances from which all other substances are made.
2. *Compounds* - are composed of molecules that are all alike but each molecule contains at least two different kinds of atoms. A compound is a substance composed of molecules which are all alike.
3. *Mixtures* - a mixture is matter composed of two or more different kinds of molecules.
4. *Solution* - a homogeneous molecular mixture. (solutions will be discussed in Chapter 11)

Note - Mixtures may be homogeneous (alike in all their parts) or heterogeneous (not alike in all their parts) but solutions are homogeneous mixtures.

Note - Atoms of different elements combine to form compounds. There are 92 naturally occurring elements but millions of compounds.

. The Atomic Theory (pages 28 and 128)

1. "All matter is made up of unit particles called atoms."
2. "The atoms of a particular element have the same weight, or at least an average weight characteristic of that element, while atoms of a different element have a different average weight."
3. "In chemical reactions whole atoms, never fractions of atoms, combine in definite ratios, or separate, or change place."

. Application of the Atomic Theory (page 28)

1. *The law of definite composition*. A compound always contains the same elements, and the ratio (by weight) in which these elements are combined in the compound is always the same. (This led Dalton to advance his atomic theory.)
2. *The law of definite proportions*. The proportion by weight of the substances involved in a specific chemical change are definite.

Molecules of Elements

1. *Monoatomic* - one atom per molecule.
2. *Diatomic* - two atoms per molecule (H_2, O_2, N_2 for example).

Note - In monoatomic elements the atom and molecule are identical. Some elements have four or even eight atoms per molecule.

Symbols

1. For brevity and convenience each element is represented by a symbol.
2. The symbols consist of one of the following:
    - a. the first letter (capitalized) in the name of the element. Ex. C for carbon, N for nitrogen, etc.
    - b. the first letter (capitalized) in the name followed by an appropriate second letter. Ex. Ca for calcium; Cl for chlorine, etc.

Note - In some cases the symbol is derived from the Latin name of the element. Ex. Fe for ferrum (iron), Cu for cuprum (copper), etc.

3. The symbol also represents a definite quantity of an element. Ex. S represents sulfur and also 32 grams (a gram-atomic weight) of sulfur.

. Formula

1. The formula shows the number and kinds of atoms in a molecule of a substance. The formula also represents a definite weight of the substance. (see Chapter 9)

quation

1. Chemical equations are used to indicate chemical reactions just as symbols desighate elements and formula indicate compounds.

A Preview of the Atom

. General Information About the Atom

1. Actual dimensions and masses of atoms are too small to be seen with the most powerful optical microscope or their weight detected with the most sensitive balance.
2. A system of relative weights based upon an arbitrairly fixed standard is used.
3. Atomic weights - the relative weights of the various atoms based upon the oxygen atom = 16.0000 as the standard. An atomic weight unit is one-sixteenth of the atomic weight of oxygen.

Note - For convenience exponential notations are used for very large and very small numbers (page 31).

. The Inner Structure of the Atom

1. The atom is a miniature solar system.
2. Major components
   a. The central compact nucleus, positively charged and containing most of the mass of the atom, is composed principally of protons and neutrons.
   b. A cloud of planetary electrons revolving about the nucleus as the earth and other planets revolve about the sun.

. Some Properties of Three Component Particles of the Atom(page 322)

Name	Charge	Mass (O = 16)
The proton (p)	positive by one unit	Large 1.007582
The neutron (n)	no charge	Large 1.00893
The electron (e)	negative by one unit	Very small 0.000549

. Electron Levels

1. The electrons are in spherical shells or levels about the nucleus (see diagrams page 34).
2. The number of electrons in the outer most level largely determines the chemical properties of the atoms.
3. The atoms of some elements loose electrons from their outer shell leaving a positively charged particle (ion) the atoms of other elements add electrons into their outer shell forming negative particles (ions). Then due to the attraction between the oppositely charged particles a compound is formed.

. Valence - the combining capacity of elements (pages 130 and 154)

1. Electrovalent compounds - caused by atoms transferring electrons to secure the stable shell of eight electrons.
2. Covalent compounds - formed due to atoms sharing electrons to secure stable shells of eight electrons.

Oxygen and Ozone

. History of Oxygen
  1. Priestley and Schelle discovered oxygen, Priestley is usually given credit for the discovery. Why?
  2. Lavoisier investigated "burning" and named oxygen. Students get details from text (page 39).

. Occurrence of Oxygen
  1. Most abundant element. About one-half of the earths mass is oxygen.
  2. Occurs both in the free state and the combined state (compounds).

. Preparation of Oxygen
  1. <u>Heating oxygen compounds</u> (HgO, $KClO_3$, $NaNO_3$, etc.) to liberate part or all of the oxygen.

$$2HgO \longrightarrow 2Hg + O_2$$
$$2KClO_3 \longrightarrow 2KCl + 3\,O_2$$

Note - The rate of the decomposition of $KClO_3$ is greatly increased by the addition of manganese dioxide (MnO_2). This is one of many examples in which the rate of chemical reaction is influenced by a foreign substance. Such a substance is called a <u>catalyst</u> and may be defined as "a substance which alters the speed of a chemical reaction without itself being permanently changed." (see page 42).

  2. <u>Preparation from sodium peroxide.</u>
     a. A convenient but expensive method

$$2Na_2O_2 + 2H_2O \longrightarrow 4NaOH + O_2$$

  3. <u>Preparation by the electrolysis of water</u>
     a. This preparation will be demonstrated. (see diagram of apparatus page 43 of text.)
  4. Preparation from the fractional distillation of air
     a. Define distillation and fractional distillation (page 44).
     b. Nearly all the commerical oxygen used in the United States is prepared by this method.

Note - Fractional distillation of air and the electrolysis of water are commercial methods of preparing oxygen. The other methods listed are laboratory methods. Commerical methods are performed cheaply and on a large scale. Laboratory methods are performed conveniently and on a small scale.

. Properties of Oxygen
  1. Physical properties (How oxygen behaves by itself)
     a. Oxygen is a colorless, tasteless, gas which weighs 1.429 gram/liter and is slightly soluble in water.
     b. Oxygen atoms, like all active gases, form diatomic molecules (O_2). Molecules containing more than two atoms are not uncommon (O_3, P_4, etc. (page 44).
  2. Chemical Properties (How oxygen reacts with other substances)
     a. Relative inactive at ordinary temperatures but very active at high temperatures.
     b. Oxygen reacts with most of the other elements both metals and non-metals to form oxides.

c. Oxygen reacts with fuels containing carbon and hydrogen to give water, carbon dioxide (CO_2) and heat. (combustion is discussed in Chapter 30.)

Note - Every chemical reaction either absorbs (takes on) heat or evolves (gives off) heat. All reactions that absorb heat are known as endothermic reactions; all reactions that evolve heat are known as exothermic reactions.

. Other Related Information

1. A simple test for oxygen employs a glowing splint.
2. Oxidation - the process or change that takes place when oxygen combines with another substance.
3. Oxidizing agent - oxygen or any compound that supplies oxygen for oxidation.
4. Oxide - a compound composed of two elements, one of which is oxygen.

Note - There are two classes of oxides, metallic and non-metallic oxides. The properties which distinguish the two classes of oxides also serve to distinguish metallic and non-metallic elements. The most striking of these properties is their reaction with water. When dissolved in water metallic oxides form bases while non-metallic oxides form acids (Chapter 13).

5. Combustion - an oxidation reaction accompanied by the evolution of heat and light.
6. Kindling temperature - the temperature at which a substance will undergo self-sustaining combustion.
7. Spontaneous combustion - when a substance takes fire on its own accord - essential conditions for spontaneous combustion are: (a) an existing slow oxidation reaction and (b) good heat insulation (page 40).

. Rate of Chemical Reaction

1. Factors that affect the rate of chemical reaction are: (page 49)
   a. the nature of the reacting substance - size of particles, solutions,etc.
   b. temperature - the higher the temperature the faster the reaction
   c. concentration - the greater the amount in a given volume the faster the reaction.
   d. amount of surface exposed - the greater the surface area exposed the faster the reaction.
   e. Catalyst - will increase or decrease the rate.

. Importance of Oxygen (see page 50-51)

1. Essential to life
2. Used in production of energy from fuels, foods, etc.

Ozone (O_3)

. Preparations

1. Ozone is formed by sunlight in the upper layers of the atmosphere.
2. Laboratory preparation - ozone is formed when an electric spark is passed through oxygen (air).

. Properties

1. Physical - pale-blue colored gas with a powerful odor.
2. Chemically - ozone is like oxygen except more active and is a good bleaching agent.

. Uses

1. Oxidizing agent
2. Bleaching agent
3. Disinfectant
4. Deodorizer
5. Sterilizing drinking water

Chapter 7
Gases and the Gas Laws: The Kinetic Molecular Theory

ote - Substances, in general, are capable of existing in three different forms or states - solids, liquids, and gases.

. The Characteristics of the Gaseous State

1. Depending upon the temperature all substances theoretically could exist as gases.

   Note - Some substances decompose on being heated before they reach the liquid or gaseous states.
2. The volume of a given mass of gas is not fixed but it expands to fill whatever container it is in. It takes the shape of the container.
3. Gases diffuse or spread through space. The molecules are in rapid motion. The speed of the molecules varies directly with the temperature.
4. The distance between gaseous molecules is large thus the molecules can be pressed closer together by pressure; gases are compressible.
5. Due to the molecules striking the walls of the container gases exert a steady pressure.
   a. The pressure of gases is measured by a barometer.

   Note - To provide a uniform basis for the comparison of the volumes of gases standard conditions of temperature and pressure (S.T.P.) have been established. Standard temperature is the freezing point of water 0°C. (273°A.) and the pressure 760 mm. of mercury (the average atmospheric pressure at sea level and termed one atmosphere) is standard pressure.

. The Gas Laws

All gaseous substances or mixtures behave alike with respect to changes in temperature or pressure, their solubility, rate of diffusion, etc. These common modes of behavior for all gases are known as the gas laws. These laws are generally known by the name of the man who first noted the regularities.

1. Boyle's Law - the volume of a gas varies inversely as the pressure, if the temperature remains constant (page 58). (for example if the pressure on a volume of gas is doubled the volume will be reduced by one-half, etc.) (solve typical problems).
2. Charles's Law - the volume of gas at constant pressure varies directly as the absolute temperature.
3. The Absolute Temperature Scale (page 62)
   a. Absolute zero = -273°C.
   b. Absolute zero is the temperature at which a gas would occupy no space.
   c. Absolute degrees are equal to centigrade degrees.
   d. To change absolute° to centigrade° add algebracially 273° to °C. (solve typical problems and also solve problems involving both pressure and temperature changes.)
4. Dalton's Law - partial pressures of gases - in a mixture of gases every exerts the same pressure that it would if it alone were confined in the same volume, and the total pressure is the sum of these partial pressures.
   a. This law must always be applied when gases are collected over water (see Fig. 47, page 66).

   Note - As a minimum, the student should be able to calculate the volume of a gas at standard conditions from measurements made at laboratory conditions and over water, and vice-versa. The student should learn to reason out each problem on the basis of the law being applied, memorization of formulas is discouraged.

5. Graham's Law - The rate of diffusion of gases
    a. Gaseous molecules will spread or separate to fill whatever space they are in; this is called diffusion.
    b. "The rate of diffusion of a gas is inversely proportional to the square root of its density or molecular weight." (work typical problems.)

Note - The lighter a gas, the faster it diffuses.

6. Avogadro's Law
    a. "Equal volumes of gases, under the same conditions of temperature and pressure, contain the same number of molecules."
7. Applications of Avogadro's Law
    a. By measurements it can be shown that 32 grams of oxygen (one gram molecular weight) occupys 22.4 liters at S.T.P. Therefore by applying Avogadro's law 22.4 liters of any other gas at S.T.P. will contain the same number of molecules as are in 22.4 liters of oxygen. Then a gram molecular weight of any gas will contain the same number of molecules. This information can be used in determining molecular weights of gases. (Chapter 8)
    b. The density (weight/liter) of a gas can be determined by dividing the molecular weight by 22.4 or the molecular weight can be determined by using the weight/liter etc.

. The Kinetic-Molecular Theory of Gases

1. Principle concepts of the Kinetic-molecular theory
    a. All gases are made up of extremely minute particles (molecules), separated at distances that are large compared with the size of the molecules themselves.
    b. These molecules are in motion, and move in straight lines until they suffer a change of direction as a result of a collision with another molecule or with the walls of the containing vessel.
    c. The collisions are perfectly elastic; that is, there is no average loss of energy as a result of them.
    d. The kinetic energy of the moving molecules is proportional to the absolute temperature.

Note - From these concepts the student should be able to account for such facts as:

(1) gases are highly compressible
(2) gases diffuse and fill any container
(3) gases do not settle
(4) mixtures of gases are homogeneous
(5) heavier gaseous molecules diffuse more slowly than lighter molecules.

Chapter 6
Hydrogen and Hydrogen Peroxide

. History of Hydrogen
  1. Cavendish a famous English scientist prepared pure hydrogen in 1766 and called it "inflammable" air.
  2. Lavoisier recognized hydrogen as a constituent of water and named it hydrogen from the Greek words "water former".

. Occurrence
  1. Hydrogen makes up about 1% of the earth's crust, mostly as compounds and is found free in volcanic gases and in the upper atmosphere.
  2. Common compounds of hydrogen are water, petroleum, natural gas and many others.

. Preparation
  1. Laboratory methods
     a. moderately active metals (Zn, Mg, Fe, Al, etc.) react with acids to give hydrogen.

        Ex. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

     b. very active metals (Na, K, etc.) react with water to give hydrogen.

        Ex. $2Na + 2H_2O \longrightarrow 2NaOH + H_2 \uparrow$

     c. certain elements (Zn, Al, Li) will react with strong bases to give hydrogen.

        Ex. $Al + NaOH \longrightarrow Na_3AlO_3 + H_2 \uparrow$

  2. Commercial methods
     a. Electrolysis of water (see preparation of oxygen)

        Ex. $H_2O + \text{elect.} \longrightarrow H_2 \uparrow + O_2 \uparrow$

     b. When steam (H_2O) is passed over red hot coke (C) hydrogen is formed.

        Ex. $H_2O + C \longrightarrow CO + H_2 \uparrow$

     c. When steam is passed over hot iron hydrogen is formed.

        Ex. $Fe + H_2O \longrightarrow Fe_3O_4 + H_2 \uparrow$ (at high temp.)

. Electrochemical series (activity series of the metals) page 82.
  1. An arrangement of the metals in order of chemical activity usually hydrogen is used as the reference element.

  Note - The most active elements are at the top of the table, the least active at the bottom. Hydrogen has an intermediate position and any element above it will replace hydrogen from dilute acids (H_2SO_4, HCl, $HC_2H_3O_2$) metals below hydrogen in the series will not replace hydrogen from these acids.

  Note - Hydrogen, which is a reducing agent, cannot be prepared from oxidizing acids such as nitric acid.

  Note - Any free metal will replace any of the metals below it in the activity series from their compounds.

  Ex. $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu \downarrow$ $Zn + AgNO_3 \longrightarrow Zn(NO_3)_2 + Ag \downarrow$

  Note - Arrows following a formula in equation for reactions as in the above equations indicate a precipitate when pointing downward and a gas when pointing upward.

E. Properties of Hydrogen

1. Physical properties - hydrogen like oxygen, is odorless, colorless, and tasteless. It is the lightest element weighing less than 0.09 grams/liter at S.T.P. It is slightly soluble in water.
2. Chemical properties
   - a. molecular hydrogen (H_2) is not very active and requires either heat or a catalyst.
   - b. Hydrogen combines with many of the other elements to form hydrides. Ex. NaH, CaH_2, NH_3, H_2S, HCl, HI, H_2O, etc.
   - c. Hydrogen acts to reduce metallic oxides to give free metals and water.

$$\text{Ex.}\quad CuO + H_2 \longrightarrow Cu + H_2O$$

Note - A compound from which oxygen is removed is said to be reduced, and the substance that removes the oxygen is called the reducing agent. Thus in the above reaction hydrogen is the reducing agent and the copper oxide is reduced. Therefore hydrogen is a reducing agent.

   - d. Hydrogen reacts with "unsaturated" organic compounds, such as liquid vegetable oils, to change them to solid fats by a process called hydrogenation. Products produced by this process are shortening, oleomargarine etc.

. Uses of Hydrogen (page 90)

1. Hydrogen is used
   - a. in producing ammonia Ex. $N_2 + 3H_2 \longrightarrow 2NH_3$
   - b. for "hardening" fats for foods.
   - c. for hydrogenation of coal to form lubricating oils, gasoline, etc.
   - d. as a reducing agent.
   - e. for many laboratory purposes.

Hydrogen Peroxide

. General

Oxygen and hydrogen combine to form another compound other than water, which is called hydrogen peroxide (H_2O_2). In this compound just twice as much oxygen is combined with the hydrogen as in water.

. Preparation

1. Hydrogen peroxide can be prepared by:
   - a. treating sodium peroxide with hydrochloric acid.

$$\text{Ex.}\quad HCl + Na_2O_2 \longrightarrow H_2O_2 + NaCl$$

   - b. treating barium peroxide (BaO_2) with sulfuric acid.

$$\text{Ex.}\quad H_2SO_4 + BaO_2 \longrightarrow BaSO_4 + H_2O_2$$

. Properties of Hydrogen Peroxide

1. Hydrogen peroxide is a colorless, sirupy liquid which decomposes readily into oxygen and water.
2. Since oxygen is readily liberated it is a strong oxidizing agent.

. Uses of Hydrogen Peroxide are:

1. oxidizing agent
2. bleaching agent
3. rocket fuel
4. germicide

Chapter 7

The Atmosphere, Nitrogen, Carbon Dioxide, Water and the Rare Gases

. The Atmosphere

1. The term atmosphere is applied to the mixture of gases, called air, that surrounds the earth.
2. The substances present in the air are:
   nitrogen, oxygen, argon, carbon dioxide, hydrogen, neon, helium, krypton, xenon and water vapor (see table - page 97 for percentages)

. The function of some of the components in the atmosphere are:

1. Nitrogen - (nitrogen will be studied later - chapter 19)
   a. serves to dilute the oxygen and as an insulating material
2. Oxygen - (see section on oxygen, Chapter 4)
3. Water Vapor
   a. Water vapor in the air accounts for all precipitation (rains, snow, etc.)
   b. Water vapor serves to regulate the temperature
4. Carbon dioxide (will be studied later - Chapter 24)
   a. Carbon dioxide is used by plants in the process of photosyntheses. The process by which carbohydrate foods are made.

Water

. Importance of Water

1. Water is essential to life (the human body is almost 70% water)
2. The properties of water are used as standards for:
   a. defining calories
   b. reference points on centigrade thermometers
   c. unit of volume, the liter
   d. relative densities - the specific gravity of liquids and solids are usually referred to water at 4°C. as having a density of one.
   e. specific heat - the specific heat of water is one calorie

Note - The specific heat of a substance is the number of calories of heat required to raise the temperature of one gram of the substance 1°C.

. Physical Properties of Water

1. Water is a colorless, odorless, tasteless compound.
2. The heat of vaporization of water, that is, the heat required to evaporate one gram, is 539 calories, which by comparison is very high.
3. Water has its maximum density at 4°C. thus ice freezes on top of a body of water. Why is this important?
4. The density of ice is less than the desnity of water. Why is this important?
5. <u>The heat of fusion</u> - that is the amount of heat, measured in calories, involved when a gram of water freezes or a gram of ice melts. The heat of fusion of water is 80 cal/g.
6. Water is an excellent solvent

Chemical properties of water

1. Water is very stable toward heat
2. Water reacts with oxides (chapter 4) to form acids and bases.
3. Water reacts with some salts by <u>hydrolysis</u> to form acids or bases. (Chapter 22, page 434)
4. Water reacts with many salts to form hydrates (page 110)
5. Water has a catalytic effect upon many chemical reactions.

The Rare Gases (see page 115)

Chapter 8

Law of Chemical Combination; Atomic Theory; Atomic Weights and Molecular Weights; Isotopes

A. Laws of Chemical Combination

1. The law of conservation of mass

   "When chemical reactions take place, the total mass of matter remains unchanged."

   Note - There are two general methods of procedure, analysis and synthesis, used to determine how much of each element is present in a compound. In the method of analysis a known weight of a pure compound is broken down into its elements, and the weights of the elements are determined. In the method of synthesis the pure compound is built up from its elements, and the weight relationships between the elements and the compound are obtained.

2. The law of definite composition

   "In any pure compound the elements are united in definite proportions by weight."

3. The law of combining weights, or equivalent weights (page 126)

   "To every element may be assigned an experimental number which, in itself, or when multiplied by some small whole number, 2, 3, 4, 5 ... expresses the weight by which the element enters into combination with other elements."

4. The law of multiple proportions

   "If the same elements, A and B, form more than one compound, the weights of A combined with a fixed weight of B in the compounds are in ratios of small whole numbers.

. Atomic Theory - (see page 128 and page 3 of the syllabus)

The simplest conclusion to be derived from these laws is that matter is made of minute particles called atoms. Atoms do not divide in ordinary chemical changes but take part in chemical reactions as unit particles.

. Atomic Weights, Combining Weights, and Valence

1. The atomic weight unit is equal to $1/16$ the weight of oxygen atom; thus, the weight of an oxygen atom is arbitrarily assigned the weight 16 atomic weight units. By using this procedure the atomic weights of all the atoms are one or greater. Oxygen was chosen here because it combines with most of the other elements.

2. The combining weight of an element is the same as the equivalent weight which may be defined as the weight in grams of the element which combines with 8 grams of oxygen or its equivalent (1.008 grams of hydrogen, 35.45 grams of chlorine, etc.)

   Note - The atomic weight of the atom of an element is its combining weight multiplied by its valence, (combining weight x valence = atomic weight) the valence may be found by dividing the atomic weight by the combining weight (atomic weight/combining weight = valence.)

3. The <u>valence</u> of an atom is the number of hydrogen atoms with which it will combine or will replace. <u>Valence</u> is the chemical name for the <u>combining capacity</u> of the <u>elements</u>.

   Note - For an accurate method of determining equivalent weights see Experiment 14 (page 61) in the laboratory manual.

D. Gram-atomic weights

1. The gram-atomic weight of an element is the atomic weight expressed in grams.

Note - For convenience the chemist usually expresses atomic weights in grams although pounds, tons, ounces, grains, or any other units of weight could just as well be used.

E. Gay-Lussac's Law of Combining Volumes

1. "When two gases combine, there is a whole-number ratio between their volumes, and between the volume of either one of them and that of the product if it too is a gas."

F. Avogadro's Law

1. Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of molecules (see Chapter 5)

G. Molecular Weight and Gram Molecular Weight

1. The molecular weight is the sum of the atomic weights in the molecule
2. The gram molecular weight is the same number of grams of the substance as the molecular weight. Ex. carbon dioxide has a molecular weight of 44; this means that a molecule of CO_2 will weigh 44 when an atom of oxygen weighs 16, or to find the molecular weight of H_2SO_4 take the sum of the atomic weights

   $H = 1 \quad S = 32 \quad O = 16$ or $2H = 2 \quad S = 32$

   $4 \times O = 64$, then $2 + 32 + 64 = 98$ molecular weight of H_2SO_4. Thus 98 grams is a gram molecular weight of the substance.

   The gram molecular weight (G.M.W. or mole) of any substance is the molecular weight expressed in grams.

Note - A gram molecular weight of all substances have the same number of molecules. This number is 6.023×10^{23} and is known as Avogadro's number.

3. The gram-molecular volume is the space occupied by a gram-molecular weight of a compound. For solids and liquids this volume varies from substance to substance, but the gram-molecular volume of all gases are the same when under conditions of equal temperature and pressure. The gram-molecular volume of gases can be measured and is found to be 22.4 liters - therefore a gram molecular weight of any gas will occupy 22.4 liters of space at S.T.P.

Note - The student should be able to calculate the molecular weight of a gas if the weight per liter is given, or to calculate the density (wt./l.) if the molecular weight is known.

Note - Very accurate atomic weights can be calculated by using the mass spectrometer (page 143-44).

4. Atoms of the same elements which have different masses are called <u>isotopes</u>. The difference in masses is due to some atoms having more neutraons in the nucleus than others, while the number of protons and electrons are the same, therefore their chemical properties are the same, the only variation is in the weight of the nucleus.

Chapter 9

Formula, Valence, Equation, Chemical Calculations

A. Formulas of Compounds (see Chapter 3)

1. The formula indicates the composition of the compound by showing the following
    a. The elements present in the compound.
    b. The number of atoms of each element present in the compound.
    c. The combining proportions (by weight) of the elements present in the compound.
    d. The molecular weight of the compound.
    e. The weight of a volume of gas.
2. A formula represents
    a. a single molecule of the compound
    b. A definite weight of the compound (the gram-molecular weight or one mole of the compound). If the compound is in the gaseous state the formula will also represent the volume occupied by the compound.

Kinds of Formulas are:

1. **Emperical** - which give the ratio of the different atoms present in the compound.
2. **Molecular** - which give a true representation of the molecules.
    Ex. the molecular formula for ethylene gas is C_2H_4. Its emperical formula is CH_2.
3. **Structural** - structural or graphic formulas show the relationship of the atoms to each other in the compound. (see page 154 - these will be studied later.)

Valence - the combining capacity of the elements.

1. Since atoms in stable molecules do not fall apart there must be some force holding them together. The two kinds of valences (covalence and electrovalence) were mentioned in Chapter 3.
    Note - The driving force in the combination of atoms is a tendency for them to attain a stable configuration of electrons such as the inert gases possess. There are two ways this stable configuration can be attained, (1) by electron transfer (electrovalence) and (2) by sharing electrons (covalence).
2. **Covalence** is defined as "the number of hydrogen atoms or other univalent atoms that the atom can hold in combination, and as the number of pairs of electrons shared with other atoms.
    Note - Univalent, bivalent, and trivalent atoms are atoms with a valence of one, two or three respectively.
3. **Electrovalence** - The number of electrons which an element must gain or lose in order to reach the stable, rare gas configuration. Electrovalence forms either positive Na^+, Ca^{++}, Mg^{++}, Al^{+++}, etc. or negative Cl^-, $S^=$, $N^\equiv$, etc. particles called ions.

Radical ions or Complex Ions

1. A group of two or more atoms, which are chemically combined and stay together through many chemical reactions, acting as a single atom is a radical. Examples of radicals are: SO_4 - sulfate; NH_4 - ammonium; NO_3 - nitrate; OH - hydroxide; CO_3 - carbonate.
    Note - Radicals have valences like other atoms and form positive or negative ions. Ex. $SO_4^=$ sulfate ion; NH_4^+ ammonium ion; NO_3^- nitrate ion, etc.

The student should memorize the names and valences of the ions given in the able on page 156.

. Equations

1. Chemical reactions are represented by equations.
2. The equation shows by symbols and formulas the rearrangement of atoms and molecules during a chemical reaction.
3. The steps in writing equations are:
    a. "recall what substances react (the reactants) and what products are formed."
    b. Write down the formulas of the reactants to the left and the formulas of the products to the right of the reaction symbol, usually shown as an arrow (---->).
    c. Then balance the equation by writing coefficients before the formulas to show the correct ratios. When the equation is balanced the same total number of atoms of each of the elements must be shown in the initial substances and in the products. (Law of conservation of mass)

. Generalized statements of chemical facts are useful in predicting what products will be formed when a reaction takes place.

1. Some generalized statements of facts are:
    a. Oxides of metals react with water to form bases:

    Ex. $CaO + H_2O \longrightarrow Ca(OH)_2$

    b. Oxides of non-metals react with water to form acids:

    Ex. $SO_3 + H_2O \longrightarrow H_2SO_4$

    c. The metals above hydrogen in the electromotive series will replace hydrogen from acids:

    Ex. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

    d. Very active metals will replace one half of the hydrogen from water:

    Ex. $2Na + 2HOH \longrightarrow 2NaOH + H_2 \uparrow$

    e. The oxides of the noble metals can be decomposed by heat to yield oxygen and the free metal:

    Ex. $2HgO \xrightarrow[\Delta]{} 2Hg + O_2 \uparrow$

Note - More information about reactions are conveyed by certain other conventions. For example in the reaction

$$2KClO_3 \xrightarrow[\Delta]{MnO_2} 2KCl + O_2 \uparrow$$

the MnO_2 is used as a catalyst and the Δ indicates heat was applied; the upward pointing arrow $\uparrow$ indicates a gas (O_2) passed off. A downward pointing arrow $\downarrow$ in an equation indicates a precipitate is formed. In addition to these, sometimes the heat that is absorbed or evolved during a reaction is indicated by showing the number of calories involved.

Ex. $Fe + S \longrightarrow FeS + 24{,}000$ calories

or $C + CO_2 \longrightarrow 2CO - 41{,}250$ calories

In the first reaction 24,000 cal. of heat were produced. In the second equation 41,250 cal. of heat was absorbed.

. Weight Calculations From Equations

Note - At this stage of the course it is very desirable that each student become quite adept at working so called "yield" problems. Being able to solve problems of this type enables the student to calculate how much material he can prepare from a given weight of starting material or how much of the reactants are required to produce a given amount of products.

Here again it is suggested that the blind use of proportions by the student is not good. Most problems in freshman chemistry are solved by reasoning based on the laws or principles being applied and by elementary arithmetic. A good procedure to follow in working problems is to write down the information given and what information is needed and what you are trying to find. An example of a typical problem and the steps used to solve it follows.

Problem - What weight of oxygen can be prepared by heating 40 grams of potassium chlorate ($KClO_3$)?

Solution -

1. Write and balance the equation $2KClO_3 \longrightarrow 2KCl + 3\ O_2$

Note - The balanced equation tells us that two G.M.W. (or moles) of $KClO_3$ decomposes to give 3 G.M.W. of oxygen, or 2 x 122.6 = 245.2 grams of $KClO_3$ will yield 3 x 32 = 96 grams of oxygen.

2. Find the amount of oxygen one gram of $KClO_3$ will produce. one gram of $KClO_3$ gives 96/245.2 = 0.3915 g. of O_2 (by simple arithmetic).
3. Find the number of grams produced by 40 grams of $KClO_3$.
   0.3915 x 40 = 15.66 grams of O_2 answer

Note - This problem could be solved as readily to find the number of liters of oxygen at S.T.P. by substituting the gram molecular volume for its equal, the gram molecular weight.

Note - The only way to become efficient at problem solving is by practice. You will find problems of this type frequently in the remainder of the course and in the course for next semester.

Chapter 10

The States of Matter and Their Transitions: Crystal Structure

<u>Kinetic Energy of Particles</u> - the tendency of the particles in any body of matter to fly apart.

1. The Kinetic energy of molecules, ($\frac{1}{2}mV^2$) is proportional to the absolute temperature (Chapter 5).

<u>Van der Waal's Forces</u> - the tendency of the particles in any body of matter to be drawn together by attractive forces.

Note - The Van der Waal's forces are very great between particles that are close together and fall off rapidly with increasing distances between particles. Thus they are effective only when the particles are very close together.

<u>Liquid-Vapor Equilibrium</u>

1. Energy relations in states of matter
    - a. In <u>gases</u> the kinetic energy predominates, or is greater than the Van der Waal's forces. Thus gases diffuse to fill any container.
    - b. In <u>liquids</u> the two forces (kinetic and attractive) are approximately equal, allowing the molecules to move readily over each other.
    - c. In <u>solids</u> the attractive energy predominates and holds the molecules in rigid position confining their motion to vibrations.

Note - Since the kinetic energy is proportional to the absolute temperature the state of any body of matter is determined by the temperature.

Transition Between the Liquid and Gaseous State

<u>Evaporation</u>

1. The molecules in a liquid move about and molecules with the greatest kinetic energy escape into the space above the liquid.
2. Since the molecules with the greatest kinetic energy escape leaving the slower molecules, the average kinetic energy is decreased - thus the remaining liquid is cooled.
3. The rate of evaporation (the tendency for the liquid molecules to escape) increases with an increase in temperature. Why?

Vapor Pressure

1. An evaporation in a closed container will establish equilibrium between the liquid and gaseous phase. (The rate of evaporation equals the rate of condensation.)
2. The pressure exerted by a vapor in equilibrium with its liquid in a closed vessel is defined as the <u>vapor pressure</u> which is usually expressed in millimeters of mercury.

Note - The boiling point of a liquid is defined as the temperature at which the vapor pressure equals the atmospheric pressure.

Note - Some liquids change readily to the vapor state when left in an open container because their vapor pressure is high, thus their boiling points are low, such liquids are said to be volatile. Liquids with opposite properties are said to be non-volatile.

Heat of Vaporization and Condensation

1. The <u>heat of evaporation</u> (vaporization) is defined as the number of calories of heat required to change one gram of substance from the liquid state to the vapor state without changing the temperature.

Note - The heat of vaporization of water is 539 calories/gram. Conversely, the condensation of a gram of a gas or vapor to a liquid liberates a quantity of heat exactly equal to the heat of vaporization. This is called the heat of condensation.

D. Critical Points

1. Critical Temperature - "The critical temperature of a gas is the temperature above which that gas cannot be liquified no matter how great a pressure is applied."
2. Critical Pressure - "The pressure required to liquify a gas at its critical temperature."

Note - The kinetic energy of gases can be decreased by applying pressure and cooling; thus gases can be liquified, giving off heat. The same amount of heat is absorbed when the process if reversed, the mechanical refrigeration operates on this principle. (Page 178)

Transition Between Solids and Liquids or Gases

. Transition Between Solids and Liquids or Gases

1. Two forms of solid substances are:
    a. Crystaline solids - with definite crystaline form
    b. Amorphous solids - without definite crystaline form
2. Freezing point
    a. The freezing point is defined as "the temperature at which both solid and liquid will remain in contact with each other without change of temperature".
    b. The melting point is the temperature at which a solid will change to a liquid.
3. Heat of fusion - The heat of fusion is defined as the amount of heat that is absorbed when one gram of a solid at the freezing point changes to a liquid at the same temperature.

Note - The heat of fusion of water is 80 cal. The same quantity of heat is given up when a gram of liquid changes to a solid at the same temperature. This is the heat of solidification.

4. Solids have vapor pressures - In solids some of the molecules have enough kinetic energy to escape. The vapor pressures of solids are usually very small; however some solids (iodine, "dry ice", etc.) have relatively high vapor pressures.
    a. Sublimation - the process of changing a solid directly to a gas, and condensing the gas back to the solid state.
    b. Heat of sublimation - The heat of sublimation is the number of calories required to change one gram of substance from the solid to the gaseous state without changing the temperature.

Note - The process of converting a liquid to a gas and condensing it back to the liquid state is called distillation. Sublimation and distillation are purification processes.

Note - The student should practice solving problems involving specific heat, heat of vaporization, etc. For ex. How many calories of heat would be required to change 10 grams of ice at 0° to steam at 100°C? (for other problems see page 190.)

. Crystalographic Systems

1. Each individual compound has a definite crystaline structure, therefore compounds may be classified by the crystal lattice they form.
2. Every crystal has a definite geometric form. There are six crystalographic systems: (see pages 183-88)

a. Isometric or regular system b. tetragonal system c. orthorombia system d. monoclinic system e. triclinic system f. hexagonal system.

Chapter 11
Solutions

A. Solution - a homogeneous molecular mixture, the composition of which may be varied continuously within limits.

1. The two constituent parts of a solution are:

a. Solute - the component that is dissolved.

b. Solvent - the component that does the dissolving

Note - When it is difficult to distinguish the solvent and solute, the component that is used chemically may be defined as the solute and the other component the solvent.

2. Solubility - The amount of solute that can be dissolved in a given amount of solvent. The solubility is usually expressed as grams of solute/100 grams of solvent.

B. Types of solutions in terms of physical state of components

1. Nine possible types of solutions and examples are:

a. gases in gases - the atmosphere
b. liquid in liquids - alcohol-water
c. gases in liquids - foams
d. solids in liquids - sugar-water
e. solids in solids - alloys of metals
f. solids in gases - smoke in air
g. liquids in gases - fog or clouds in air
h. gases in solids - pumice stone
i. liquids in solids - jellies

2. Relative importance of the types of solutions (the first five are the most important)

a. Mixtures of gases in gases were discussed in Chapter 5 (Dalton's Law, page 65). Gases mix with each other in all proportions.

b. Solutions of gases in liquids - a liquid will dissolve a limited amount of gas depending upon the following factors.

(1) Effect of specific properties - (see page 193).

(2) Effect of pressure - gases are most soluble at higher pressures.

Note - A statement of this: "The weight of gas dissolved by a definite quantity of a given solvent is directly proportional to the pressure, provided the temperature remains constant", is known as Henry's Law.

(3) Effect of Temperature - with a raise in temperature gases become less soluble.

c. Solutions in liquids in liquids have the following important properties:

(1) The vapor pressure of a solution made by mixing two liquids is never as great as the sum of the two vapor pressures of the liquids before they are mixed.

(2) Since the vapor pressure is lowered the boiling point is increased. (see Chapter 10)

(3) Fractional distillation - two liquids whose boiling points are different can be separated by distillation called fractional distillation.

(4) A constant-boiling solution - the vapors given off have the same percentage composition as the solution, therefore the concentration of the solution is unchanged or the component parts cannot be separated by distillation.

d. Solutions of <u>solids</u> <u>in</u> <u>liquids</u>

- (1) In general solubility increases with an increase in temperature. there are some exceptions to this rule, for ex. Calcium acetate is more soluble in cold water. (Explain the solubility curves, Fig. 115, page 197).
- (2) The <u>heat</u> <u>of</u> <u>solution</u> is the amount of heat which is absorbed or evolved when one gram of a solid dissolves; any solid that absorbs heat when it dissolves will be more soluble in hot water and the reverse is also true.
- (3) The <u>heat</u> <u>of</u> <u>dilution</u> is the heat involved when a concentrated solution is diluted.

C. Concentration of Solutions

The term concentration denotes "the quantity of solute dissolved in a given quantity of solvent or of solution" (page 199).

1. <u>Concentrated</u> <u>and</u> <u>dilute</u> - are loose qualitative terms that are used for comparisons.
2. <u>Molar</u> <u>solution</u> - a gram molecular weight of solute per liter of solution.
3. <u>Normal</u> <u>solution</u> - a gram equivalent weight of solute per liter of solution. (see Chapter 13, page 236).
4. <u>Molal</u> <u>solution</u> - a gram molecular weight per 1000 grams of solvent.
5. <u>Standard</u> <u>solution</u> - any solution whose exact concentration (molarity or normality) is known (see page 236).
6. <u>Saturated</u> <u>solution</u> - a solution that has all the solute the solvent will hold, or a solution in which the dissolved solute is in equilibrium with undissolved solute.
7. <u>Super</u> <u>saturated</u> <u>solution</u> - a solution that has more solute than the solvent will ordinarily hold - Explain how to make.
8. <u>Unsaturated</u> <u>solution</u> - a solution that has less solute in the solvent than it would hold under these conditions.

Note - In expressing concentrations the word "per" is always used. "Per" can be translated into mathematical terms by means of the division sign ÷, so that the above may be expressed as:

$$\text{concentration} = \frac{\text{quantity}}{\text{volume}}$$

$$\text{molarity} = \frac{\text{number of moles}}{\text{number of liters}}$$

$$\text{normality} = \frac{\text{number of equivalents}}{\text{number of liters}}$$

$$\text{molality} = \frac{\text{grams of solute in 1000 g. of water}}{\text{gram-mol. wt. of solute}}$$

when weight, molarity and volume are required:
grams = mol. wt. x molarity x volume in liters
grams = eq. wt. x normality x volume in liters

$$M = \frac{\text{grams}}{\text{mol. wt. x volume}}$$

$$N = \frac{\text{grams}}{\text{eq. wt. x volume}}$$

Colligative Properties of Solutions

1. Those properties of a solution that depend on the number of solute particles in a given amount of solvent.

a. Lowering of vapor pressure

(1) Raoult's law - the lowering of the vapor pressure of a solvent is directly proportional to the weight of the solute dissolved is a definite weight of solvent.

Note - The kind or size of particles make no difference as long as the substance is not volatile or does not separate into ions.

b. Elevation of the boiling point

(1) If the vapor pressure is lowered the boiling point will be proportionally raised. In a one molal solution the boiling point would be raised 0.518°. In a 2 molal solution the boiling point elevation would be twice as great (2 x 0.518= 1.036°) etc.

c. Lowering of freezing point

(1) The freezing point depression is proportional to the concentration of the solution. For ex. a 1 molal solution will freeze at a temperature 1.86° lower than the pure solvent.

Note - Practical applications of these properties are anti-freeze mixtures for automobile radiators, cooling mixtures such as ice cream salt, determination of molecular weights of compounds, etc.

. Osmosis

1. If a semipermeable membrane separates two water solutions of different concentrations, the rate of diffusion of water molecules out of the concentrated solution is less than in the opposite direction. The net result is water passes from the dilute to the concentrated solution. This phenomenon is known as osmosis.

   (A semipermeable membrane is one that allows one component of a solution, but not the other, to pass through it.)

2. Osmotic pressure - That pressure which, when applied to a solution, will just prevent the entrance of solvent into it through a semipermeable membrane is the osmotic pressure of the solution.

3. The laws of osmotic pressure show that osmotic pressure is:
   - a. Approximately proportional to concentration.
   - b. Independent of the nature of the solute particles.
   - c. Proportional to the absolute temperature.

   or

   "The osmotic pressure of a solution is approximately equal to the gas pressure which the solute would exert if it were a gas occupying the same volume as the solvent, at the same temperature."

Chapter 12
Ionization

A. Ionization, Electrolytes and Non-electrolytes
  1. Ionization - the term that is applied to the process that takes place when some molecules dissociate into charged particles, called ions, upon being dissolved in certain solvents (usually water).
  2. Electrolyte - a substance which, when dissolved in certain solvents, or when fused (melted), dissociates into ions.
  3. Non-electrolyte - a substance that does not separate into ions when dissolved in certain solvents such as water.

B. Abnormal Properties of Solutions of Electrolytes
  1. In solutions of electrolytes the boiling point and freezing point etc. are abnormally affected because, it is recalled, the colligative properties of solutions are a result of the number of particles present; therefore if a substance ionizes to form two particles per molecule the change in boiling point etc. would be twice as great.

C. Theory of Ionization
  1. Arrhenius Theory (1887) makes the following assumptions:
    a. Molecules of electrolytes, when dissolved in water dissociate into charged particles called ions.
    b. There is the same number of positive as negative charges on the ions, that is, the solution as a whole is neutral.
    c. The process of ionization is reversible so that such solutions may contain both ions and molecules in a state of equilibrium.
    d. Chemical reactions in ionic solutions take place between the ions, the molecules will supply the ions however until they are used up.
    e. Ions are just as effective as molecules in changing the colligative properties of solutions.

D. Evidence to Support the Theory of Ionization
  1. Solutions of electrolytes will conduct an electric current.
  2. The abnormal lowering of the freezing point and elevation of the boiling point.
  3. Abnormal osmotic pressure.
  4. High melting points of certain compounds.

E. The Degree or Percentage of Ionization
  1. Strong electrolytes - a substance that is almost completely ionized or almost all of the molecules are separated into ions.
  2. Weak electrolytes - a substance that is relatively only slightly ionized, or most of the substance remains as molecules.
  3. The degree of ionization can be determined by the abnormality of the colligative properties.
     For ex. - If KCl ionizes to form two particles per molecule

$$KCl \longrightarrow K^+ - Cl^-$$

the boiling point would be elevated exactly twice as much and the freezing point depressed twice as much as they would be if no ionization took place. Thus the colligative properties are proportional to the degree of ionization.

F. Interionic Attraction

1. Unlike electrical charges attract each other and the attraction is a function of the distance between the charged particles. Thus in concentrated solutions the attraction between oppositely charged ions is very great because the ions are close together. In dilute solutions the attraction is small because the particles are far apart.
2. The electrostatic forces are greater, the higher the charge on the ions. For ex. the attraction between Mg^{++} and $SO_4^{=}$ is greater than the attraction between Na^+ and Cl^-

Note - In ionic solutions chemical reactions take place between the ions, the properties of ionic solutions are actually the properties of the ions (see page 216).

G. Electrolytic Conduction - the movement of charged particles through a solution resulting in an electric current.

1. The conducting of a solution will depend on:
   a. The number of ions present.
   b. The speed of each ion present
   c. The charge carried by each ion

Electrochemistry

A. Electrolysis - the decomposition of a compound by an electric current.

1. Definitions
   a. Cathode - the electrode by which electrons enter the solution, the negative electrode.
   b. Anode - the electrode by which electrons leave the solution, the positive electrode.
   c. Anion - the ions attracted by the anode, negatively charged ions.
   d. Cation - the ions attracted by the cathode, positively charged ions.

Note - Electrolysis takes place in water solutions or in melted compounds. The anions and cations move toward the anode and cathode respectively At the electrode a transfer of electrons take place. For ex. in the electrolysis of melted NaCl the reactions are:

$$\text{cathode} - Na^+ + e^- \longrightarrow Na$$
$$\text{anode} - 2Cl^- \longrightarrow 2e + Cl_2$$

   e. Voltaic cell - is a device used to convert chemical energy into electrical energy. The reaction takes place spontaneously. The two reactants must be separated into two partial reactions, in which electrons are lost by one substance and gained by the other. The two electrodes must be connected by an external wire to complete the circuit.
   f. Electrolytic cell - Electrical energy is applied from an outside source (a battery) causing a non-spontaneous reaction to take place. Example - Any typical electrolysis set-up (see Fig. 124, page 217).

2. Application of Electrolysis
   a. Electroplating
   b. Electrotyping
   c. Electrorefining of metals
   d. Electroreductions of metallic compounds
3. Faraday's Laws of Electrolysis are:
   a. "The weight of a given element liberated at an electrode in electrolysis is directly proportional to the quantity of electricity passed through the solution."

  - b. "The weight of different elements liberated by the same quantity of electricity are proportional to the equivalent weights of the elements." Thus the quantity of electricity required to deposit one gram-equivalent weight of an element is the same for all elements.

4. Other electrical terms are:
   - a. <u>Coulomb</u> - the quantity of electricity conducted by an electric current of one ampere flowing for one second (amp. x sec. = coulombs).
   - b. <u>Faraday</u> - 96,500 coulombs - the amount of electricity required to liberate one gram-equivalent weight of substance at an electrode.

Note - This is obvious since a gram molecular weight of any substance contains the same number of molecules (Avagadro's number).

   - c. <u>Ohm</u> - unit of resistance.
   - d. <u>Ampere</u> - rate of flow.
   - e. <u>Volt</u> - electromotive force (e.m.f.) or potential - The electrical pressure under which the current flows.

B. Batteries

1. Definition - a battery is a group of cells - a cell consists of an electrolyte and a single pair of electrodes, one of which is capable of being oxidized and the other reduced.

Note - Oxidation is defined as a loss of electrons from an atom or ions, and reduction is the process of gaining electrons. Thus an oxidation cannot be separated completely from a reduction. In order for one substance to loose electrons another must gain them (see page 211).

2. <u>The acid cell</u> - this set-up will be demonstrated and explained in theory (see page 223).
3. <u>The dry cell</u> -
   - a. Cathode is a zinc vessel.
   - b. Anode is a carbon rod.
   - c. Electrolyte is ammonium chloride (NH_4Cl).
   - d. Manganese dioxide (MnO_2) is used as a depolarizer (see polarizer page 224).
   - e. The reactions are: Neg. pole $Zn \longrightarrow Zn^{++} + 2e^-$

$$\text{Pos. pole } NH_4^+ + 2MnO_2 + 2e^- \longrightarrow 2NH_3 + H_2O + Mn_2O_3$$

Note - The dry cell gives a small current for a short period of time otherwise it becomes polarized, that is, the carbon electrode becomes covered by hydrogen.

4. <u>The Daniell Cell</u> (crowfoot cell)
   - a. Cathode - metallic zinc in $ZnSO_4$ solution.
   - b. Anode - sheet copper in $CuSO_4$ solution

Note - The less dense $ZnSO_4$ solution floats on top of the more dense $CuSO_4$ solution.

   - c. The reactions are: Neg. pole $Zn \longrightarrow Zn^{++} + 2e^-$

$$\text{Pos. pole } Cu^{++} + 2e^- \longrightarrow Cu \quad \text{or} \quad Zn + Cu^{++} \longrightarrow Zn^{++} + Cu$$

Note - The Zn goes into solution and the copper plates out thus the Zn becomes lighter in weight while the copper gains.

5. <u>Lead Storage Battery</u>
   - a. Cathode - a lead plate covered with spongy lead.
   - b. Anode - a lead plate covered with lead dioxide (PbO_2).
   - c. Sulfuric acid is used as the electrolyte.

d. The reactions are:

$$\text{Neg. pole } Pb + SO_4^{=} \dashrightarrow PbSO_4 + 2e^-$$

$$\text{Pos. pole } PbO_2 + SO_4^{=} \dashrightarrow PbSO_4 + 2H_2O$$

Note - The chief difference in the storage battery and the others is that the storage battery can be recharged, the chemical reactions that produce the electrons are reversed by applying an electric current and the chemicals are returned to their original condition. The overall reactions in the lead storage battery are:

$$Pb + PbO_2 + 2H_2SO_4 \underset{\text{charge}}{\overset{\text{discharge}}{\rightleftharpoons}} 2PbSO_4 + 2H_2O$$

The density of the H_2SO_4 is used to indicate the charge of the battery.

6. <u>The Edison Cell</u>
   - a. Neg. electrode - iron
   - b. Pos. electrode - nickel covered with NiO_2
   - c. Electrolyte is sodium hydroxide
   - d. The reactions are: $NiO_2 \cdot 2H_2O + Fe \dashrightarrow Fe(OH)_2 + Ni(OH)_2$

Note - The lead storage bettery gives a high potential over a short period of time, each cell produces two volts. The Edison cell is more durable and does not suffer from disuse, it produced about 1.3 volts/cell.

The Electromotive Series

(Review electrochemical series Chapter 6, page 82.)

Note - As indicated in Chapter 6, there are differences in the tendencies of elements to give up valence electrons to become ions. The relative activities of the elements have been determined and the elements arranged in an activity series with the most active metal at the head of the series. It is obvious that elements that give up their valence electrons with the most force will accept them back only when a greater force is applied. Therefore the most active metal will require the greatest electromotive force to cause its ions to accept back the valence electrons to become metals again, and the activity series is also an electromotive-force series.

1. Standard electrodes potentials
   - a. The hydrogen electrode is arbitrarily assigned the value of zero and all others are referred to it as a basis.
2. Applications of the electromotive series
   - a. The electromotive force of any cell may be determined by subtracting algebraically the potentials of its two electrodes.

$$\text{Ex. } Fe \dashrightarrow Fe^{++} + 2e^-; \quad E = 0.44 \text{ volts}$$

$$2Ag^+ + 2e^- \dashrightarrow 2Ag; \quad E = -0.80 \text{ volts}$$

$$0.44 + (-0.80) = 1.24 \text{ volts} \quad \text{ans.}$$

   - b. The fact that iron will replace silver from its salts is also indicated.
   - c. The electromotive series also applied to non-metals the same as it does to metals.

Chapter 13
Acids, Bases and Salts

A. Introduction

1. Three important classes of electrolytes

a. Acids - "any compound that gives hydrogen ions (H^+) in solution." An acid is formed when a non-metal oxide is dissolved in water.

$$\text{Ex. } SO_3 + H_2O \longrightarrow H_2SO_4$$

Note - In water solution, the hydrogen ion is combined with a molecule of water, i.e., it is hydrated ($H^+ + H_2O \longrightarrow H_3O^+$) producing a hydronium ion (H_3O^+). For convenience, the hydrogen ion, in solution will be represented by H^+, but it is to be understood that such an ion is actually hydrated.

b. Base - "Any compound that gives hydroxyl ions (OH^-) in water solution." A base is formed when a metallic oxide is dissolved in water.

$$\text{Ex. } CaO + H_2O \longrightarrow Ca(OH)_2$$

$$\text{metal oxide + water} \longrightarrow \text{a base}$$

Note - In some cases bases may be referred to as alkalis and as hydroxides.

c. Salts - An electrically neutral compound composed of the metal (cation) of any base and the non-metal (anion) of any acid.

Note - Acids combine with bases to produce salts and water.

$$\text{Acid + base} \longrightarrow \text{salt + water}$$

B. Acids

1. Properties of acids - i.e. the hydrogen ion in solution

a. Taste - dilute acids have a sour taste.

b. Action on indicators - acids change the color of certain indicators, (blue litmus to red for example).

Note - Indicators are organic acids, called dyes, which, give one color in acid solution and another color in basic solutions due to rearrangement of the valence bonds in the molecule.

c. Action on metals

(1) metals above hydrogen in the electromotive series replace hydrogen in acids to form salt and hydrogen.

$$\text{Ex. } Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

(2) Some metals react with oxidizing acids (HNO_3, H_2SO_4, etc.) by oxidation-reduction.

$$\text{Ex. } Cu + HNO_3 \longrightarrow Cu(NO_3 + NO_2 + H_2O \text{ (no } H_2 \text{ produced)}$$

d. Action on hydroxides of metals (bases)

(1) Acids react with hydroxides of metals to form a salt and water.

$$\text{Ex. } Mg(OH)_2 + H_2SO_4 \longrightarrow MgSO_4 + 2H_2O$$

$$\text{base + acid} \longrightarrow \text{salt + water}$$

Note - The student must observe that the properties of acids and bases as given here refers to a water solution of the ions of these electrolytes and not to the molecular compounds.

2. Naming (nomenclature) of acids (see page 275)

a. Binary acids - only two elements present. Ex. HCl, H_2S, HCN, etc. Name by using the prefix "hydro" and the name of the other element, changed to end in the suffix "ic". Ex. HCl - hydrochloric acid, H_2S - hydrosulfuric acid.

Note - Hydrocyanic acid actually contains three elements, but by definition a radical (in this case CN) can be considered as an element. (see Chapter 9)

b. Ternary acids - acids composed of three elements, hydrogen, oxygen, and an acid-forming element (S, N, P, Cl, etc.). Ex. H_2SO_4, HNO_3, H_2CO_3, etc. These acids are named by using the name of the acid-forming element with its ending changed to denote the valence of the non-metal (acid-forming) element.

(1) suffixes

"ic" is used to denote the higher valence or the more common acid of the series.

"ous" is used to denote the acid in which the non-metal has the next lower valence than the "ic" acid.

(2) prefixes

"hypo" is used with the suffix "ous" to denote an acid in which the valence of the non-metal is less than it is in the "ous" acid.

"per" is used with the suffix "ic" to denote an acid in which the valence of the non-metal is greater than it is in the "ic" acid. Examples: HNO_2 - Nitrous acid; H_2SO_3 - sulfurous acid; HNO_3 - nitric acid; H_2SO_4 - sulfuric acid. There are four oxygen acids of the non-metallic (acid-forming) element chlorine: $HClO$ - hypochlorous acid; $HClO_3$ - chloric acid; $HClO_2$ - chlorous acid; $HClO_4$ - per chloric acid.

Note - The student should be able to tell which of these acids are the more or most common and what the valence of the non-metal is in all cases.

. Bases

1. Properties of bases - i.e. the OH^- ion in solution
   - a. Taste - bases have a bitter taste.
   - b. Feel - bases have a slick, slippery feeling on the skin.
   - c. Bases turn the indicator litmus blue.
2. Naming (nomenclature) of bases

The metallic element is named followed by the work hydroxide - the OH^- ion is named as an atom, or as a unit. Ex. NaOH - sodium hydroxide; $Ca(OH)_2$ - calcium hydroxide. If the metal has more than one valence, its name is changed to end in "ic" for the higher valence and "ous" for the lower valence. Ex. CuOH - cuprous hydroxide (lower valence); $Cu(OH)_2$ - cupric hydroxide (higher valence).

Note - Acids may be classed as:
monohydrogen (monobasic) acid - HCl
dihydrogen (dibasic) acid - H_2SO_4
trihydrogen (tribasic) acid - H_3PO_4
depending upon the number of replaceable hydrogens.
Bases may be classed as:
monohydroxyl (monoacid) base - NaOH
dihydroxyl (diacid) base - $Ca(OH)_2$
trihydroxyl (triacid) base - $Al(OH)_3$

. Neutralization

1. Heat of neutralization
   - a. The heat of neutralization of all acids are the same showing that the only reaction is between the H^+ and OH^- to produce water.
2. Titration of acid and bases
   - a. The porcess of determining just how much of a standard solution is required to react with a sample being analyzed is called a "titration".

- b. This important branch of analytical chemistry is called "volumetric analysis".
- c. The point at which the desired reaction is complete is indicated by an indicator and is called the "end point".
- d. If the concentrations are expressed in normalities it is true that (ml. of acid) x (N of acid) = (ml. of base) x (N of base)
  For examples see page 237 and Exp. 22. The student should become adept at working problems of this type.

E. Salts

- 1. The following types of reactions produce salts
  - a. Metals replacing hydrogen from acids - $Mg + 2HCl \longrightarrow MgCl_2 + H_2$
  - b. Certain metals with strong bases - $Zn + 2NaOH \longrightarrow Na_2ZnO_2 + H_2$
  - c. All neutralization reactions
  - d. Certain salts with strong bases - $CuCl_2 + 2NaOH \longrightarrow Cu(OH)_2 + 2NaCl$
  - e. Certain salts with certain acids - $2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$
  - f. Some salts with other salts - $BaCl_2 + K_2SO_4 \longrightarrow 2KCl + BaSO_4$
- 2. Naming (nomenclature) of salts
  - a. The name of the metal is written first (If the metal shows two valences the "-ous" and "-ic" ending are used to indicate the lower and higher valences, respectively.) followed by the name of the acid, from which the salt was formed, with the following suffixes:
    - (1) if a binary acid the ending "-ide" is used.
    - (2) other acids whose names end with "-ous" and "-ic" give salts whose names end in "-ite" and "-ate" respectively.
- 3. Varieties of salts are:
  - a. Normal salts - "A normal salt is a compound built up wholly of the cations of any base and the anions of any acid." (see page 238)
  - b. Acid salts - "An acid salt is a salt composed of both metallic and hydrogen cations, together with the anions of an acid."
    Ex. $NaHCO_3$, $KHSO_4$, etc.
  - c. Basic salts - "A basic salt is composed of metallic cations, together with both hydroxyl anions of a base and anions of an acid."
    Ex. $PbOHNO_3$, $BiOHCl_2$, etc.
  - d. Double salts - A combination of two salts (Ex. $Cu(AsO_2)_2 \cdot Cu(C_2H_3O_2)_2$, $KCl \cdot MgCl_2$ etc. (see page 630).
  - e. Complex salts - Examples - $Ag(NH_3)_2Cl$, $Cu(NH_4)_4Cl_2$, K_2SiF_6, etc. (see page 631)

Chapters 14 and 15

The Halogen Family, The Hydrogen Halides, Oxidation-Reduction Reactions; The Oxygen Compounds of the Halogens

The Halogen Family

. General Information

1. The elements fluorine, chlorine, bromine, iodine and astatine are members of the halogen family.
2. Their most common valence is -1 (Cl^-, Br^-, I^-, etc.)
3. As the atomic weights increase the physical properties change--Fluorine and chlorine are gases, bromine is a liquid, and iodine a solid.
4. The chemical activity decreases as atomic weights increase--Fluorine is most active, then chlorine, bromine, etc.

. Preparation

1. Laboratory
    - a. Electrolysis of fused (melted) salts
        - (1) Fluorine - this is the only method for preparing fluorine because it is the strongest chemical oxidizing agent. Potassium hydrogen fluoride is fused, and the electrolysis carried out in a copper cell. The cell reaction is:

$$2KHF_2 \xrightarrow{e} KF + H_2 + F_2$$

        - (2) Chlorine, bromine, and iodine can be prepared by the same method (similar equations for bromine and iodine)

$$NaCl \xrightarrow{e} Na + Cl_2$$

        - (3) The halogens other than fluorine can be prepared by the electrolysis of aqueous solutions of their salts.

$$2NaCl + 2H_2O \xrightarrow{e} 2NaOH + H_2 + Cl_2$$

(similar equations for bromine and iodine)

    - b. Reaction of the hydro-acids with oxidizing agents
        - (1) The halogens other than fluorine can be prepared by this method.

$$2HCl + MnO_2 \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

$$2KMnO_4 + 16HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$$

(similar equations for bromine and iodine)

Note - HI is a strong reducing agent, and HBr is strong enough as a reducing agent to reduce conc. H_2SO_4.

$$2HBr + H_2SO_4 \longrightarrow Br_2 + SO_2 + 2H_2O$$

    - c. Replacement reactions:
        - (1) Bromine and iodine are prepared by replacing them from their compounds with chlorine.

$$2KBr + Cl_2 \longrightarrow 2KCl + Br_2$$

$$2KI + Cl_2 \longrightarrow 2KCl + I_2$$

2. Commercial Preparation
    - a. The commercial preparations of <u>fluorine</u> and <u>chlorine</u> are the same as the laboratory methods.
    - b. <u>Bromine</u> is recovered from sea water by oxidizing it to free bromine by treating it with chlorine.

$$2Br^- + Cl_2 \longrightarrow 2Cl^- + Br_2$$

The bromine is blown out with air and treated with sulfur dioxide to reduce the bromine to hydrogen bromide.

$$Br_2 + SO_2 + 2H_2O \longrightarrow H_2SO_4 + 2HBr$$

The HBr is dissolved in water to separate it from the air then treated with chlorine to form free bromine again which is distilled out and collected (see page 265).

c. **Iodine** is prepared by the reduction of sodium iodate ($NaIO_3$) with sodium hydrogen sulfite ($NaHSO_3$):

$$NaIO_3 + NaHSO_3 \longrightarrow 2Na_2SO_4 + 2NaHSO_4 + H_2O + I_2$$

Iodine is also obtained from the iodides in certain sea weeds (kelps)

d. **Asatine** is not found in nature, but is produced by nuclear reactions.

. Properties of the Free Elements

1. **Fluorine** (page 262) is a pale-yellow gas, 1.3 times as heavy as air, and forms diatomic molecules (F_2). It is the most active element.
2. **Chlorine** (page 247) is a greenish-yellow gas (Cl_2) with a characteristic suffocating odor. It is 2.49 times as heavy as air, and is easy to liquify.
3. Chemically Chlorine:
   - a. combines with most of the metals directly.
   - b. combines with most of the non-metals. $2Sb + 3Cl_2 \longrightarrow 2SbCl_3$
   - c. combines with hydrogen. $Cl_2 + H_2 \longrightarrow 2HCl$
   - d. combines with the hydrogen in hydrogen compounds.
     $H_2S + Cl_2 \longrightarrow 2HCl + S$ $2NH_3 + 3Cl_2 \longrightarrow 6HCl + N_2$
     (its reaction with turpentine ($C_{10}H_{16}$) will be demonstrated)
   - e. action upon water $Cl_2 + H_2O \longrightarrow HCl + HClO$
   - f. Chlorine water is a powerful bleaching agent.
4. **Bromine** (page 268) is a dark-red liquid whose density is 3.119 which vaporizes readily to a reddish-brown gas. Chemically bromine is like chlorine except it is less active. Bromine water is a good oxidizing agent.
5. **Iodine** is a purplish-black solid. Its density is 4.93, melts at 113.5 and is purified by sublimation. Chemically it is like chlorine and bromine except less active.

. Comparative Chemical Activity of the Halogens

The comparative activity of the halogens is well indicated by their reaction with water.

1. Fluroine reacts vigorously with water. $2F_2 + 2H_2O \longrightarrow 4HF + O_2$
2. Chlorine reacts readily with water in well defined steps.
   - a. $Cl_2 + H_2O \longrightarrow HCl + HClO$
   - b. $2HClO \longrightarrow 2HCl + O_2$

   (The decomposition of HClO is hastened by a high temperature and light.)
3. Bromine reacts with water to a very limited extent.

   $$Br_2 + H_2O \longrightarrow HBr + HBrO$$
4. Iodine does not react with water.

Note - As a general rule the more chemically active an element is, the more stable its compounds will be. Thus fluorides are the most stable, then chlorides, bromides, and iodides.

E. Important Uses of the Halogens

1. <u>Fluorine</u>
    a. Preparation of fluorine - containing hydrocarbons (freon CCl_2F_2, see page 179)
    b. Fluoridation of drinking water
    c. Separation of uranium isotopes
2. <u>Chlorine</u> - one of the most important industrial chemicals
    a. Bleaching agent - wood pulp, paper, and cotton cloth
    b. manufacturing of organic compounds, dyes, drugs, explosives
    c. war gases - phosgene ($COCl_2$) chloropicrin CCl_3NO_2 etc.
    d. Purification of drinking water
    e. Preparation of bleaching powder ($CaOCl_2$)
3. <u>Bromine</u>
    a. Production of ethylene bromide ($C_2H_4Br_2$) used in gasoline.
    b. Manufacturing of organic compounds, dyes, drugs, etc.
    c. Preparation of light sensitive AgBr for photography.
4. <u>Iodine</u>
    a. Most important use is as an antiseptic "Tincture of Iodine" (alcohol)
    b. Iodine is used in testing for starch. Starch-Potassium iodide paper used to test for fluorine, chlorine, bromine, ozone, and hydrogen peroxide.

The Hydrogen Halides

A. General Information

1. The hydrogen halides are the compounds formed when hydrogen combines with a member of the halogen family.

B. Methods of Preparation

1. <u>Hydrogen fluoride</u> - by the action of a strong non-volatile acid on a metallic fluoride.

$$CaF_2 + H_2SO_4 \text{ (conc.)} \longrightarrow CaSO_4 + 2HF$$

Note - The formula for hydrogen fluoride is sometimes written as H_2F_2 because the molecules tend to associate. In the liquid state the formula may be H_6F_6. However for our purposes the formula HF will be used.

2. <u>Hydrogen chloride</u> - Two general methods of preparation are available.
    a. The action of a strong non-volatile acid on a metallic chloride.

$$NaCl + H_2SO_4 \text{ (conc.)} \longrightarrow NaHSO_4 + HCl$$

    b. By the direct union of the elements $H_2 + Cl_2 \longrightarrow 2HCl$

Note - This is the only hydrogen halide that can be prepared satisfactorily by this method. Fluorine is difficult to prepare and reacts violently with hydrogen. While bromine and iodine react only with difficulty with hydrogen.

3. Hydrogen bromide - Two general methods of preparation are available.
    a. Reaction of a bromide with a strong non-volatile, non-oxidizing acid.

$$2NaBr + H_3PO_4 \longrightarrow Na_2HPO_4 + 2HBr$$

    b. The hydrolysis of a non-metallic bromide such as PBr_3

$$PBr_3 + 3H_2O \longrightarrow H_3PO_3 + HBr$$

4. Hydrogen iodide is prepared by the same two methods as HBr.
    a. $2NaI + H_3PO_4 \longrightarrow Na_2HPO_4 + 2HI$
    b. $PI_3 + 3H_2O \longrightarrow H_3PO_3 + 3HI$

. Properties of the hydrogen halides
1. Stability decreases rapidly in the order HF, HCl, HBr, and HI.
2. As reducing agents they decrease in the order HI, HBr, HCl, and HF.
3. Dry hydrogen halides are covalent compounds.
4. Reaction with water

$$HF + H_2O \longrightarrow H_3O^+ + F^-$$
$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
$$HBr + H_2O \longrightarrow H_3O^+ + Br^-$$
$$HI + H_2O \longrightarrow H_3O^+ + I^-$$

Note - The aqueous solutions of these compounds are called hydrofluoric, hydrochloric, hydrobromic and hydroiodic acids. All are strong acids except HF which is weak due to association of its molecules. Hydrofluoric acid reacts with glass but can be stored in wax bottles. The reaction with glass is represented by the equation:

$$CaSiO_3 + 6HF \longrightarrow CaF_2 + SiF_4 + 3H_2O$$

. Uses of the Hydrogen Halides
1. Hydrochloric acid can be used in most cases requiring a strong, cheap acid.
2. HCl is used to catalyze the hydrolysis of starch to glucose.
3. HCl, HBr and HI are all reducing agents (see methods of preparing chlorine etc.)
4. Hydrobromic and hydroiodic acids have little use as acid.
5. Hydrofluoric acid is used in etching glass.

Oxidation-Reduction Reactions

. The Nature of Oxidation-Reduction Reactions (see pages 211, 246 and 254)
1. The oxidation states (or valences) of some of the elements involved in the reaction are changed.
    For example the reaction between a solution of bromine in water and metallic copper as represented by the equation: $Cu + Br_2 \longrightarrow CuBr_2$, or ionically the equation is:

$$Cu^0 + Br_2^0 \longrightarrow Cu^{++} + 2Br^-$$

    Now suppose that we examine each of the reactants to see what had happened to it.

Note - The student should recall that:
- a. The valence of an unreacted atom is zero.
- b. A typical metal gets its positive valence by losing electrons, whereas,
- c. A non-metal gets its negative valence by taking on electrons.
- d. The algebraic sum of positive and negative valences of all the atoms in a compound must equal zero.
- e. Hydrogen usually has a valence of +1 while oxygen usually has a valence of -2.

The copper reacts by losing electrons thus: $Cu^0 \longrightarrow Cu^{++} + 2e$ and the bromine reacts by gaining two electrons thus:

$$Br_2 + 2e^- \longrightarrow 2Br^-$$

Note - Loss of electrons is oxidation; gain of electrons is reduction, therefore the copper is oxidized by the bromine and the bromine is reduced by the copper. The copper is the reducing agent and the bromine is the oxidizing agent.

2. Oxidation-Reduction Reactions Include:
   a. Combination reactions $Cu + Br_2 \longrightarrow CuBr_2$
   b. Decomposition reactions $2HI \longrightarrow H_2 + I_2$
   c. Displacement reactions $Zn + CuSO_4 \longrightarrow Cu + ZnSO_4$
   d. Complex reactions such as:

$$2KMnO_4 + 16\ HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$$

. Balancing Oxidation-Reduction Reactions (see pages 246, 257 and 258)

Note - Simple oxidation-reduction reactions (the first three examples above) can be balanced easily by the <u>inspection method</u>. The more complex types (the fourth example above) are often difficult to balance by this method. However there are several methods of balancing these complex equations all of which are based upon the fact that "the number of electrons yielded by the reducing agent must, in the balanced equation, be equal to the number of electrons accepted by the oxidizing agent. The "valence-electron method" has many advantages and will be outlined here.

1. The valence-electron method

Suppose that it is desired to balance the equation:

$$CuO + NH_3 \longrightarrow Cu + N_2 + H_2O$$

a. Notice the change in valence that certain ones of the atoms undergo. This might be indicated thus:

$$\overset{+2}{Cu}O + N\overset{-3}{H_3} \longrightarrow \overset{0}{Cu} + \overset{0}{N_2} + H_2O$$

decrease of 2

increase of 3

or $\overset{+2}{Cu} + 2e \longrightarrow Cu^{o}$ $\quad \overset{-3}{N} \longrightarrow 3e + N$

b. Then multiply these equations by the smallest numbers that will give equal numbers of electrons in the two equations.

$$3\overset{+2}{Cu} + 6e \longrightarrow 3\overset{0}{Cu} \qquad 2\overset{-3}{N} - 6e \longrightarrow \overset{0}{N_2}$$

c. Add the equations together (note that the electrons cancel out)

$$3\overset{+2}{Cu} + \cancel{6e} \longrightarrow 3\overset{0}{Cu}$$
$$2\overset{-3}{N} - \cancel{6e} \longrightarrow \overset{0}{N_2}$$
$$3\overset{+2}{Cu} + 2\overset{-3}{N} \longrightarrow 3Cu + N_2$$

d. Now that the number of atoms of copper and nitrogen are fixed substitute into the original equation and complete the balance by inspection.

$$2CuO + 2NH_3 \longrightarrow 3Cu + N_2 + 3H_2O$$

Notice that there are 3 oxygen atoms and 6 hydrogen that combine to form water ($3H_2O$).

Note - Notice that when the oxidizing agent is multiplied by the valence change of the reducing agent and the reducing agent is multiplied by the valence change of the oxidizing agent the equation is balanced. This will be true in all cases except where not all of one of the reactants react by oxidation-reduction.

For example in the reaction illustrated by the equation:

$$\overset{0}{Cu} + H\overset{+5}{N}O_3 \longrightarrow \overset{+2}{Cu}(NO_3)_2 + \overset{+2}{N}O + H_2O$$

increase of 2

decrease of 3

Since not all of the nitric acid reacted by oxidation-reduction the valence change of the copper (2) indicates the amount of NO formed and not the amount of HNO_3 used, all of the Cu was oxidized so the valence change (3) of the nitrogen shows the amount of Cu used. Now the equation:

$$3Cu + HNO_3 \longrightarrow CuNO_3 + 2NO + H_2O$$

can be written and the balance completed by inspection.

Oxygen Compounds of the Halogens

A. General Information

1. In addition to a valence of -1 the halogens, except fluorine, show valences of +1, +3, +5 and +7.
2. The corresponding compounds of chlorine are illustrated by the acids $HClO$, $HClO_2$, $HClO_3$, and $HClO_4$.
3. The <u>oxygen acids</u> of the halogens have these common properties
   - a. Except for HClO, they are strong acids
   - b. They are unstable toward heat
   - c. They are strong oxidizing agents

Note - Since the oxy-acids of the halogens are unstable toward heat they have little use as acids but both the acids and their salts which are more stable, are used as oxidizing agents.

B. Important Oxy-acids of the Halogens

1. Hypochlorous acid (HClO)
   - a. Preparation

$$Cl_2 + H_2O \rightleftharpoons HCl + HClO$$

   - b. Properties: oxidizing and bleaching agent

$$HClO \longrightarrow HCl + (O)$$

   - c. Uses
     - (1) preparation of bleaching powder ($CaOCl_2$)

$$Ca(OH)_2 + Cl_2 \longrightarrow CaOCl_2 + H_2O$$

     - (2) preparation of hypochlorite salts which are good oxidizing agents and bleaching agents.

2. Chloric Acid: ($HClO_3$)
   - a. Preparation

$$Ba(ClO_3)_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HClO_3$$

   - b. Uses - The chlorates (salts of chloric acid) are strong oxidizing agents, (recall preparation of oxygen) and are used in "weed killers".

3. Perchloric acid ($HClO_4$)
   a. Preparation:

$$4KClO_3 \longrightarrow KCl + 3KClO_4$$

$$KClO_4 + H_2SO_4 \longrightarrow KHSO_4 + HClO_4$$

   b. Properties - the perchlorates are the most stable salts of the oxy-acids of chlorine. They are good oxidizing agents and $KClO_4$ is nearly insoluble in water.
   c. Uses - The perchlorates are used as oxidizing agents and in quantative analysis of potassium.
4. Iodic acid (HIO_3)
   a. Preparation

$$NaIO_3 + H_2SO_4 \longrightarrow NaHSO_4 + HIO_3$$

sodium iodate ($NaIO_3$) is found in Chile saltpeter and is the principle source of iodine.

Chapter 16
The Alkali Family of Metals

. Introduction and general properties

1. Metals are those elements which react with oxygen to form basic anhydrides. non-metals react with oxygen to form acid anhydrides.
2. Metals are mallable, and have a metallic luster while non-metals are brittle and possess no luster.
3. The alkali metals show the following properties:
    a. The density increases as the atomic number increases (Li 0.53, Cs 1.90)
    b. The melting point decreases as the atomic weight increases (Cs 26, Li 186)
    c. Atomic and ionic radii increase as the atomic number increases.
    d. The electrode potential decreases as the at. wt. increases. (Cs gives up its electron the most readily)
    e. Li, Na, and K are lighter than water.
    f. The alkali metals are soft and easy to cut; the atoms are large.
    g. The alkali metals are very active and are not found free in nature.
    h. The order of chemical activity is Cs, Rb, Li, K, Na.
    i. The compounds of the alkali metals are very stable and are soluble in water.
        (1) oxides cannot be decomposed by heat.
        (2) carbonates decompose only at high temperatures.
        (3) hydroxides melt without decomposing
4. Chemical properties are:
    a. Valence - electrovalence of +1 formed by loss of the one valence electron.
    b. The replacement of hydrogen in water to form bases.

    Ex. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

    c. The alkali metals form the strongest bases.
    d. All alkali metals corrode readily in air to form oxides.
    e. The free metals are strong reducing agents (they give up electrons readily).

. Occurrence

1. Sea water - 2.8% Na; 0.08% K.
2. Igneous rocks - 2.85% Na; 2.60% K.
3. Potassium is present in plants
4. Salt beds and salt seas
5. Oilwell brines
6. Order of abundance - Na, K, Li, Rb, Cs.

. Preparation of Free Metals

1. Recall methods of preparation of halogens from melted salts.

    $$2NaCl \xrightarrow{elect.} 2Na + Cl_2$$

2. Electrolysis of fused hydroxides

    $$2NaOH \xrightarrow{elect.} Na + O_2 + H_2$$

Note - In these electrolytic processes (Costner Process for fused hydroxides and the Downs Cell for fused chlorides) care must be taken to prevent the products from mixing and reacting.

3. Cs can be prepared by reacting its salts with calcium and separating the Cs by distillation.

4. Potassium can be prepared by a similar procedure.

D. Uses of the Alkali Metals (not very useful as metals)
    1. Sodium is the most useful, some of its uses are:
        a. Sodium vapor lamps.
        b. Core for exhaust valves in aircraft engines.
        c. Used in the manufacturing of many chemicals, soap, dyes, organic chemicals, etc.
    2. Cesium and rubidium are used in photoelectric cells.

E. Compounds of the Alkali Metals
    1. Lithium compounds
        a. Lithium chloride - used in preparing other compounds and in air conditioning.
        b. LiBr and Li_2CO_3 - used in medicine and in ceramics for glasses and glazes.
    2. Sodium compounds
        a. Sodium chloride (table salt)(NaCl) (See Chapter 13, page 231) essential to the diet. Widely distributed in nature, used to prepare metallic sodium and chlorine and most compounds of the two elements, -- HCl, Na_2SO_4, etc.
        b. Sodium hydroxide (NaOH)
            (1) Prepared by two processes:
                (a) Treating calcium hydroxide with sodium carbonate

$$Ca(OH)_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaOH$$

                recover NaOH of filtration and evaporization.
                (b) Electrolysis of sodium chloride solution:
                (see prep. of chlorine, Chapter 14, page 244)
            (2) Uses of sodium hydroxide: (see table page 292)
                (a) manufacturing of hard soaps
                (b) refining of petroleum
                (c) preparation of wood pulp for paper making
                (d) rayon, chemicals, textiles, etc.
        c. Sodium sulfate (Na_2SO_4) - a byproduct of the manufacture of acids other than H_2SO_4.
            (1) Uses are
                (a) manufacturing of rayon, glass, and paper
                (b) medicine - "Glauber's salt". ($Na_2SO_4 \cdot 10\ H_2O$)
        d. Sodium Peroxide (Na_2O_2) used in preparation of oxygen (see page 43).
        e. Sodium Thiosulfate ($Na_2S_2O_3$)
            (1) Preparation: By adding sulfur to a solution of sodium sulfite

$$Na_2SO_3 + S \longrightarrow Na_2S_2O_3$$

            (2) Used in photography - "Hypo" and to remove excess chlorine "antichlor" from any substance bleached with chlorine.
        f. Sodium Carbonate (soda ash)(Na_2CO_3)
            (1) Preparation - Solvay Process
                (a) Raw materials: NaCl, NH_3, CO_2, $CaCO_3$, H_2O
                (b) Products are: Na_2CO_3, NaHCO, $CaCl_2$ and NaOH.

$$NH_3 + H_2O + CO_2 \longrightarrow NH_4HCO_3$$

$$NH_3HCO_3 + NaCl \longrightarrow NaHCO_3 + NH_4Cl$$

$$2NaHCO_3 + heat \longrightarrow Na_2CO_3 + H_2O + CO_2$$

$$2NH_4Cl + CaO \longrightarrow CaCl_2 + 2NH_3 + H_2O$$

$$CaCO_3 \longrightarrow CaO + CO_2$$

$$Na_2CO_3 + Ca(OH) \longrightarrow CaCO_3 + 2NaOH$$

Overall the reaction is:

$CaCO_3 + 2NaCl \longrightarrow CaCl_2 + Na_2CO_3$, which is an impossible reaction.

(2) Uses:
  (a) $Na_2CO_3 \cdot 10\ H_2O$ - washing soda
  (b) glass, soap, chemicals, etc. (see table, page 295)

g. Sodium Hydrogen Carbonate (bicarbonate of soda)(baking soda)($NaHCO_3$)
  (1) Preparation: Solvay Process above, and by treating Na_2CO_3 solution with carbon dioxide:

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

  (2) Uses: baking and medicine.

h. Sodium Nitrate ($NaNO_3$) (Chile saltpeter)
  (1) Uses: fertilizer

i. Sodium Cyanide (NaCN)
  (1) Preparation

$$NaNH_2 + C \longrightarrow NaCN + H_2$$

  (2) Uses: dissolves gold, good poison.

j. Sodium Phosphates (will be studied with phosphorus (Chapter 19)

3. Potassium Compounds:

a. Potassium Hydroxide (Caustic Potash) (KOH) Same properties as NaOH. NaOH is cheaper and is usually used.

b. The potassium halides (KCl, KBr, KI)
  (1) very similar in properties to the corresponding sodium compounds.
  (2) the bromide and iodide are used in medicine and in the preparation of salts used in photography.

c. Potassium Sulfate (K_2SO_4)
  (1) Preparation:

$$KCl + H_2SO_4 \longrightarrow K_2SO_4 + 2HCl$$

  (2) Used in fertilizer and in preparing baking powder.

d. Potassium Carbonate (potash)(K_2CO_3)
  (1) Preparation:

$K_2S + CaCO_3 \longrightarrow K_2CO_3 + CaS$ and from wood ashes.

  (2) Uses: manufacturing of hard glass, and soft soap.

e. Potassium Nitrate (saltpeter)(KNO_3)
  (1) Preparation:

$$NaNO_3 + KCl \longrightarrow KNO_3 + NaCl$$

  (2) Uses:
    (a) Preparation of O_2: $2KNO_3 \longrightarrow 2KNO_2 + O_2$
    (b) Good oxidizing agent
    (c) Black gun powder - KNO_3, S, C.
    (d) Fertilizer, fireworks, medicine.

). Compounds of Rubidium and Cesium are used in radio tubes and photoelectric cells.

The Spectroscope

The Spectrum

1. Every element produces light of a characteristic wave length; no two elements give the same wave length.
2. A <u>spectrum</u> is any series of radiant (light) energies arranged in order of wave lengths.
3. Elements can be identified by their spectrum.
   Ex. Li - red, Na - yellow, K - violet, Cs - blue, Rb - blue-red, etc.

Kinds of Spectrum

1. The Emission Spectra:
   The spectrum formed by the light given off when any substance is heated.
2. The Absorption Spectra:
   When a white light (which gives a continuous spectrum) is passed through a gas, the gas absorbs the wave lengths which it would emit if luminous. The interruptions in the continuous spectrum will indicate the gas present. Helium was discovered in the sun's atmosphere before it was on earth by this procedure.

Chapter 17
The Classification of the Elements

The Periodic Law

A. Introduction

1. The purpose of the classification of the elements is to facilitate their study.
2. Atomic number - (see Chapter 3, page 32) the number of protons in the nucleus of the atoms of an element.

B. Origin of the Periodic System

1. Dobereiner Triads (1829)
    - a. The elements could be arranged into triads with similar properties - The atomic weight of one was the mean of the other two. Ex. Cl - 35.5; Br - 79.9; I - 126.9.
2. Newlands Law of Octaves (1836)
    - a. The English Chemist Newland stated that "the eighth element, starting from a given one, is a kind of repetition of the first, like the eight notes in music. (Note - the rare gases were not considered by Newland.)
3. Mendeleev and the Periodic Law
    - a. If the elements are arranged in order of increasing atomic weights, there is a recurrence of chemical and physical properties. This generalization may be stated as follows: "The physical and chemical properties of the elements are periodic functions of their atomic weights." This is known as the periodic law.
    - b. Mendeleev, a Russian chemist, announced this law in 1869 at the same time the German chemist Lothar Meyer announced it in Germany.
    - c. Mendeleev arranged the elements according to their atomic weights which gave some discrepancies (see A and K, Co and Ni, Te and I.)
    - d. Mosley (1913) determined atomic numbers by bombarding metallic elements with electrons and measuring the length of X-rays given off.

Note - When the two electrons in the K shell are bombarded with electrons they give off X-rays. The wave length of the X-rays is a function of the positive charge on the nucleus, the greater the charge the more firmly and closely the electrons are held thus giving off shorter wave lengths. Hydrogen with one proton in its nucleus gives the longest X-rays, then He, Li, Be, B, etc. follow. Thus each element differs from the one preceeding it by having one more proton in its nucleus. The number of protons in the nucleus is the atomic number. The atomic number was found to be the most fundamental property of the elements. When the elements were arranged in the order of increasing atomic number the defects in the Mendeleev arrangement were corrected. (A = 18, K = 19; Co = 27, Ni = 28; Te = 52, I = 53).

- e. Modern Periodic Law: The properties of the elements are a periodic function of the atomic number.
- f. Isotopes - atoms of the same element which have different atomic weights are known as isotopes.

Note - Isotopes of potassium have weights of 39, 40, 41, the average weight being 39.1 because the isotope 39 is the most abundant. Isotopes of argon have weights of 36, 38 and 40, the average weight being 39.94 because isotope 40 is the most abundant. Thus in the original periodic table, potassium preceeded argon but the at. no. of all argon atoms is 18 and the at. no. of all potassium atoms is 19 which corrects the defect in the original periodic table. The same explanation accounts for the other discrepancies.

C. General Aspects of the Periodic Table

Note - The extended or long form of the periodic table is to be used in this course and is shown on page 316 of the text book and on the lecture room wall.

1. Groups - The vertical columns are called groups. Groups 1 to 7 consist of a main group and a sub-group which are called families and are designated as Family A and Family B.
   Ex. Group I A (group one Family A) is the alkali family Li, Na, K, Rb, Cs and Fr. (Main group)
   Group IB (group one Family B) is the silver family Cu, Ag, and Au. (sub-group).

Note - In the short form of the periodic table the two families of each group are shown in the same column in the extended form they are shown in separate columns. This is the essential difference in the two forms.

Note - The only relationship between the elements in the two families of any group is that they will have the same number of valence electrons. For similarities between members of a family recall the study of the halogen family and the alkali family.

2. Periods - Each horizontal row is a period. There is a total of six periods:
   a. One short period of two elements H and He.
   b. Two short periods of eight elements each starting with Li (at. no. 3) and Na (at. no. 11).
   c. Two long periods of 18 elements each starting with K (at. no. 19) and Rb (at. no. 37).
   d. One long period of 32 elements starting with Cs (at. no. 55) and including 14 rare-earth elements.
3. The zigzag line that descends the table toward the right separates the metals (to its left) from the non-metals (to its right).
4. The rare earth elements (57-71) and the acetinide series (89-98) are shown separately at the bottom of the chart.

D. Value of the Periodic Table

1. Aid to memory
2. Makes possible the study of families of elements instead of single elements.
3. In the past it has been useful in "checking of experimental data".
4. Used to "predict new elements".

Note - This chapter serves merely to introduce some of the <u>basic</u> <u>generalizations</u> of the periodic system. Many references will be made to it next semester. The student should learn these as a start and add to the knowledge as the year progresses.

Chapter 18
The Structure of the Atom

Note - This is an expansion of some of the material presented earlier. The student should review "a preview of the atom", Chapter 3, page 30 before studying this chapter.

A. The Inner Structure of the Atom (the atom is a minature solar system)
  1. An atom consists of:
    a. A minute compact nucleus made up of protons and neutrons, which carries a positive charge equal to the number of protons and gives the atoms most of its mass or weight.
    b. Planetary electrons arranged in a definite manner into shells of energy levels about the nucleus. The electrons give size or volume to the atom and are equal in number to the protons in the nucleus making the atom electrically neutral.

B. The Components of the Atom are:
  1. <u>The electron</u> - a particle of negative electricity, with a charge of one and a mass of 1/1837 that of the hydrogen atom. Electrons are present in all matter and are responsible for negative charges.
  2. <u>The proton</u> - a positive unit of electricity with a mass of 1836/1837. The proton is a hydrogen ion (H^+).
  3. <u>The neutron</u> - a neutral particles with a mass of one.
  4. <u>The positron</u> - a particle with a unit positive charge, and a mass equal to the electron, or a positively charged electron.

. The Arrangement of Electrons in Atoms (The Bohr Theory)
  1. It is assumed that the planetary electrons are arranged in concentric orbits or shells, called energy levels, about the nucleus (page 32).
  2. These shells or energy levels, in order of increasing energy, are given numbers 1, 2, 3, 4, etc. or the letters K, L, M, N, etc.
  3. Each shell is divided into subshells, which are designated in order of increasing energies, as s, p, d, f, subshells.

  Note - Electrons always occupy the position with the lowest energy. For convenience it is assumed that each succeeding element is formed by adding one more planetary electron to the atoms of the previous element.

  4. The electrons are definitely arranged in the shells as follows:
    a. The K shell (with only an s subshell) is filled with two electrons.
    b. The L shell (with s and p subshells) is filled with eight electrons.
    c. The M shell (with s, p, and d subshells) is filled with 18 electrons.
    d. The N shell (with s, p, d, and f subshells) is filled with 32 electrons.

  Note - The maximum number of electrons in a given shell is given by the formula:

$$2N^2 = \text{max. no. of electrons}$$

N is the shell number when K = 1, L = 2, M = 3, etc. Thus the
1st shell can hold $2(1^2) = 2$ electrons
2nd shell can hold $2(2^2) = 8$ electrons
3rd shell can hold $2(3^2) = 18$ electrons, etc.

  Note - Since the energies of the various subshells increase within the shells in the order s, p, d, f above the second shell, the s subshell of the next higher shell will have a lower energy state than the d or f subshell in the preceeding energy level.

Ex. 4s has a lower energy state than 3d, or 6s has a lower energy state than 4d which is lower than 5p. Therefore since electrons always occupy the lowest energy position a new shell of electrons is started before the third and fourth shells are completed. However a given shell above the first will fill the s and p subshells (8 electrons) before a new shell is started. This gives rise to the transition elements including the rare earths etc.

5. Valence shell - since the electrons in the outer shell of an atom determine the chemical properties of the atoms they are called the valence electrons. The valence shells of atoms tend to form stable configurations of eight electrons by giving up or accepting electrons (oxidation-reduction) (electrovalence) or by sharing electron (covalence). Thus this is of primary importance to the chemist.
   a. Electrovalent compounds are formed by electrovalence or by transferring electrons. No atom ever completely gives up more than three electrons to form a positive ion also no atom gains completely more than three electrons to form a negative ion.
   b. Covalent compounds are formed by atoms sharing electrons to complete the valence shells. In sharing the electrons are not completely lost or gained by either atoms.

The Electronic Structure of Some Elements in Tabular Form

At. No.	Element	K shell	L shell		M shell			N shell				O shell				P shell			
		s	s	p	s	p	d	s	p	d	f	s	p	d	f	s	p	d	f
1	H	1																	
2	He	2																	
3	Li	2	1																
4	Be	2	2																
5	B	2	2	1															
..	..	..	..	..															
10	Ne	2	2	6															
11	Na	2	2	6	1														
..	..	..	..	..	..														
18	A	2	2	6	2	6													
19	K	2	2	6	2	6		1											
20	Ca	2	2	6	2	6		2											
21	Sc	2	2	6	2	6	1	2											
..	..	..	..	..	..	..	..	..											
28	Ni	2	2	6	2	6	8	2											
29	Cu	2	2	6	2	6	10	1											
30	Zn	2	2	6	2	6	10	2											
31	Ga	2	2	6	2	6	10	2	1										
..	..	..	..	..	..	..	..	..	..										
36	Kr	2	2	6	2	6	10	2	6										
37	Rb	2	2	6	2	6	10	2	6			1							
..	..	..	..	..	..	..	..	..	..			..							
54	Xe	2	2	6	2	6	10	2	6	10		2	6						
55	Cs	2	2	6	2	6	10	2	6	10		2	6			1			
..	..	..	..	..	..	..	..	..	..	..		..	..			..			
81	Tl	2	2	6	2	6	10	2	6	10	14	2	6	10		2	1		
82	Pb	2	2	6	2	6	10	2	6	10	14	2	6	10		2	2		
83	Bi	2	2	6	2	6	10	2	6	10	14	2	6	10		2	3		
Total/shell		2	8		18			32				18				5			

APPENDIX II
Principles Syllabus

Chemistry 114, Sections 1 and 2

Group B

Syllabus

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Chemistry 114
First Semester 1953-54 A & B

ecture Text: Textbook of Chemistry by Mack, Garrett, Haskins and Verhoek
aboratory Manual: Semimicro Laboratory Exercises in General Chemistry
Second Edition; Burrows, Arthur and Smith.

General Information for the Student

The purpose of this course is to give the student an understanding of the rinciples, theories and laws of chemistry. The course will consist of lectures ith demonstrations, recitations, and laboratory work.

The students should:

(1) learn to use the text book as a reference guide in both the theory and laboratory portions of the course.
(2) read over the topics assigned for lecture before going to class, and the topics pertaining to the laboratory work before going to the laboratory.
(3) make notes during lectures to supplement the syllabus, underscoring important points.
(4) as soon as possible after each lecture, read over the syllabus and lecture notes and if there are any points not clearly understood look them up in the text using the index if necessary.
(5) prepare a list of questions, over any material which is not well understood before attending the discussion period.
(6) ask the instructor for individual help on any points that are not clearly understood after the discussion period. (It is not recommended that the student skip over such points, for they may have an important bearing on later subject matter.)
(7) at frequent intervals (possibly once each week) review the syllabus and lecture notes for the material covered to date.
(8) follow the instructions given on the first pages of the laboratory manual in performing the laboratory work unless other instructions are given by the laboratory instructor.

Examinations and Grading

1) Three one-hour objective-type examinations will be given at previously arranged times during the course. (Possibly during the 5th, 10th and 15th weeks.) These shall account for 50% of the course grade.

2) Short tests (or quizzes) may be given from time to time during either the discussion periods or the lecture period. Laboratory reports will be collected at not infrequent times to be graded and the quiz-laboratory instructor may use what other methods desired to secure marks on student attitudes, laboratory techniques, etc.

The short tests, laboratory report grades and teacher estimate grade, etc. will be combined to make up 30% of the course grade.

3) A two-hour objective-type final examination will be used to account for the other 20% of the course grade.

Chemistry 114 is the first course of a two semester sequence and will be ollowed by Chemistry 124. Chemistry 124 is a continuation of chemistry 114 and the ame text book and laboratory manual are used for both courses. A comprehensive xamination over both courses will be given as the final examination for Chemistry .24.

Chapters I and II

A & B

Introduction - The Science of Chemistry
Matter and Energy

(Students read and study Chapters 1 and 2.)

A. Definitions

1. Chemistry - is the branch of natural science which deals with the properties of matter, the changes which matter undergoes, and the laws and theories that govern the changes.
2. Science - is knowledge that has been classified and arranged into a systematic arrangement by subjects.
3. Matter - is defined as any substance that has mass and occupies space (has size) and is capable of undergoing change.
    a. Mass - is the quantity of matter in a body.
    b. Weight - is the attractive force (gravity) which the earth exerts on a body.

Note - For all practical purposes on the earth's surface mass and weight can be used interchangeably.

    c. Law of Conservation of Mass - In ordinary chemical reactions mass is neither created nor destroyed.

4. Substance - is the term applied to homogeneous bodies: those which are alike in all their parts.
5. Substances are identified by their properties - distinguishing characteristics such as color, odor, taste, melting point, boiling point, density, and the like are physical properties.
6. Chemical Properties - are the essential properties of a substance; if they are changed a new substance is formed.
    a. Chemical Reactions - are the processes that convert substances into other substances (changing chemical properties).
7. Energy - the ability to do work. Work is done when a body is moved through space.

Note - In all physical or chemical changes energy is involved, either energy is applied to cause the change or is released by the change. (familiar manifestations of energy are: heat, light, electricity, etc.)

    a. Kinetic energy - the energy of moving bodies.
    b. Potential energy - energy due to position. (water behind the dam, etc.)
    c. Chemical energy - energy that is stored in a substance due to its chemical composition.
    d. Law of conservation of energy - energy like mass, in all ordinary transformation is neither created nor destroyed.

8. Scientific Laws - The statement of a generalization or a general principle based on a common mode of action of natural things.

. History of Chemistry

(see pages 3-6)

These are only a beginning of terms and words to be defined during the course. he student must strive to understand each work or term as it is introduced, and o review them often enough to retain a working knowledge of them. The student ust be very certain at this time that a good foundation is secured upon which to ase the work that follows.

C. The Scientific Method A & B

1. The accumulation of facts by:
    - a. observation - observing the material in its present form.
    - b. experimentation - this is done by changing the conditions under which the material exists and observing the result. Thus the scientist experiments to establish facts which when established can be classified into general principles or laws.
2. The Use of Laws
    - a. Hypothesis - a tentative but reasonable explanation for observed facts. An attempt to answer the question why? A hypothesis can be tested by experimentation. If it meets the tests the hypothesis is advanced to a theory if it fails to meet the test it is either modified or discarded.
3. Theory - a tested and approved hypothesis. A theory attempts to coordinate into a concise, unified picture all the facts related to a given subject, also certain assumptions may be made in formulating a theory.

D. Units of Measure Used in Chemistry

(see text pages 11 and 842; also inside front cover of laboratory manual and Experiment 1.) Centimeter - gram - second (CGS) system is used in all countries of the world for scientific work.

1. Weight and Volume
    - a. unit of weight, one gram (g.)
    - b. unit of volume, one liter (l.)
2. Length
    - a. units of length are the meter (m), centimeter (cm) and millimeter (mm)
3. Time
    - a. seconds (sec), minutes, etc.
4. Temperature
    - a. Centigrade - usually used in science - degrees centigrade (°C.)
    - b. Fahrenheit - household use - degrees Fahrenheit (°F.)
    - c. Absolute scale of temperature to be considered later

Note - A few important values for converting from one system to another should be memorized by the student. For example - 1 inch = 2.54 cm.; 1 meter = 39.37 inches; 1 ounce = 28.34 grams; 1 liter = 1.0567 quarts.

E. Other terms defined

1. Calorie - (page 14) a unit for measuring heat. The calorie is that quantity of heat which will raise the temperature of one gram of water 1°C. at 15°.
2. Atom - minute particles of matter that do not, during natural processes, undergo division. The smallest part of an element to take part in a chemical reaction.
3. Molecule - formed when two or more atoms combine. Molecules are the smallest particles of a chemical substance that can exist alone and which have the characteristic composition and properties of the substance.
4. Density - the quantity of matter per unit of volume. The ratio of the mass of a homogeneous protion of matter to its volume.
5. Specific Gravity - the ratio of the weight of any volume of a substance to the weight of an equal volume of some substance taken as a standard or unit as, usually, water for solids and liquids, and air or hydrogen for gases.

Note - When water is used as the standard its maximum density (at 4°C.) is assigned a value of one. The density of gases usually refer to the weight in grams per liter and the density of solids and liquids refers to the weight in grams per cubic centimeter (g/cc).

Chapters 3, 8, and 18

Groups of Substances, Laws of Chemical Combination, and the Atomic Theory

A. Four Groups of Substances

1. <u>Elements</u> - are substances incapable of being broken down into simpler substances by ordinary chemical means. The molecules of an element consist of only one kind of atoms. The elements are the fundamental substances from which all other substances are made.
2. <u>Compounds</u> - are composed of molecules that are all alike but each molecule contains at least two different kinds of atoms. A compound is a substance composed of molecules which are all alike.
3. <u>Mixtures</u> - a mixture is matter composed of two or more different kinds of molecules.
4. <u>Solution</u> - a homogeneous molecular mixture. (solutions will be discussed in Chapter 11)

Note - Mixtures may be homogeneous (alike in all their parts) or heterogeneous (not alike in all their parts) but solutions are homogeneous mixtures.

Note - There are 92 naturally occurring elements. The elements are of exceptional interest to us in our study of chemistry because all other substances are made from them. Elements combine to form compounds; most natural substances are mixtures of these compounds.

By chemical operations it is possible to determine how much of each element is present in a compound. There are two general methods of procedure, analysis and synthesis, for carrying out such operations. In the method of analysis a known weight of a pure compound is broken down into its elements, and the weights of the elements determined. In the method of synthesis the compound is built up from its elements and the weight relationships between the elements and the compound are obtained. By using these methods chemists have determined the exact composition of hundreds of compounds. From the facts thus obtained four general laws of chemical combination have been formulated.

B. Laws of Chemical Combination (see pages 119-126)

1. <u>The law of conservation of mass</u> (see pages 10 and 119-120)
   If the total weight of the reactants which enter into a chemical change is compared with the total weight of the products of that change, the two are found to be identical. Thus when carbon burns in oxygen, carbon and oxygen are the reactants; carbon dioxide the product, and 12 grams of carbon combines with 32 grams of oxygen to give 44 grams of carbon dioxide.

$$\text{12 g.} + \text{32 g.} \longrightarrow \text{44 g.}$$

$$C + O_2 \longrightarrow CO_2$$

REACTANTS YIELD PRODUCT

2. <u>The law of definite composition</u> or <u>the law of definite proportions</u>. (see pages 23, 29 and 122.)
   A compound always contains the same elements, and the ratio (by weight) in which these elements are combined in the compound is always the same.
   For Ex. All samples of water contain 11.19% by weight of hydrogen and 88.81% of oxygen.
3. <u>The law of combining weights</u> (see pages 124-127.)
   "To every element may be assigned an experimental number which, in itself, or when multiplied by some small whole number, 2,3,4,5, ... expresses

the weight by which the element enters into combination with other elements." The combining weight or equivalent weight of any element is the weight of the element which will combine with 8.0000 parts by weight of oxygen or with 1.0080 parts of hydrogen.

4. The Law of multiple proportions (page 127)
"If the same elements A and B, form more than one compound, the weight of A combined with a fixed weight of B in the compound are in ratio of small whole numbers.

. The Atomic Theory (see pages 27-28 and 128)

The simplest conclusion to be derived from these laws of chemical combination is that matter is made up of minute particles called atoms. (This was first stated by John Dalton in 1805 and is now known as Dalton's Atomic Theory of Matter.) In a given element the atoms have a fixed or constant weight and chemical combination is the union of these individual units, or atoms, each carrying into the union its fixed weight. When, as in chemical changes, elements are transferred from one compound to another, the transfer is made in terms of these individual atoms, each with its constant weight.

This conclusion is so well supported that the existence of atoms is now accepted as a fact, and Dalton's atomic theory is the basic theory upon which all chemistry rests.

1. The main points of the atomic theory are:
    a. "All matter is made up of unit particles called atoms."
    b. "The atoms of a particular element have the same weight, or at least an average weight characteristic of that element, while atoms of a different element have a different average weight."
    c. "In chemical reactions whole atoms, never fractions of atoms, combine in definite ratios, or separate, or change places."
    d. "Atoms of the same elements can unite in more than one ratio to form more than one compound."

The Atomic Theory

. Introduction

The original atomic theory pictured the atoms as minute fragments of matter, solid, homogeneous, and unchanging. The only property accounted for by it was the mass of the atoms, which was constant for all atoms of the same element and different for all atoms of different elements. It made no attempt to explain the chemical nature of elements or to answer the questions: why and how atoms combine to form compounds? To answer these and other questions we first must become familiar with the internal structure of the atom.

. General Information and Definitions

1. Actual dimension and mass of any single atom is too small to be seen with the most powerful optical microscope or its weight detected with the most sensitive balance.
2. A system of relative weights based upon an arbitrairly fixed standard is used.
3. Atomic Weights - the atomic weight unit is equal to 1/16 the weight of the oxygen atom; thus, the weight of an oxygen atom is arbitrairly assigned the weight 16 atomic weight units. By using this procedure the atomic weights of all the atoms are one or greater. Oxygen was chosen here because it combines with most of the other elements.

Note - A table of the elements, including symbols, atomic numbers, and atomic weights is found on the inside of the back cover of the text book.

4. Gram Atomic Weight - The atomic weight expressed in grams. Ex. at. wt. of oxygen = 16, the gram atomic weight = 16 grams.
5. Symbols (page 29)
    a. For brevity and convenience each element is represented by a symbol.
    b. The symbols consist of one of the following:
        (1) the first letter (capitalized) in the name of the element. Ex. C for carbon, N for nitrogen, etc.
        (2) the first letter (capitalized in the name followed by an appropriate second letter. Ex. Ca for calcium; Cl for chlorine etc.

Note - In some cases the symbol is derived from the Latin name of the element. Ex. Fe for ferrum (iron), Cu for suprum (copper) etc.

    c. The symbol also represents a definite quantity of an element. Ex. S represents sulfur and also 32 grams (a gram-atomic weight) of sulfur.

6. Formula (pages 29 and 30)
    a. The formula shows the number and kinds of atoms in a molecule of a substance. The formula also represents a definite weight of the substance. (see Chapter 9)
7. Equation (page 30)
    a. Chemical equations are used to indicate chemical reactions just as symbols designate elements and formula indicate compounds.
8. Valence - the combining capacity of elements (see pages 130 and 154).

The Structure of the Atom
(Pages 30-37 and 321-324)

The Inner Structure of the Atom (the atom is a minature solar system).

1. An atom consists of: (page 324)
    a. A minute, compact nucleus made up of protons and neutrons, which carries a positive charge equal to the number of protons and gives the atom most of its mass or weight.
    b. Planetary electrons arranged in a definite manner into shells or energy levels about the nucleus. The electrons give size or volume to the atom and are equal in number to the protons in the nucleus making the atom electrically neutral.

The Components of the Atom are: (page 322)

1. The electron - a particle with a negative charge of one and a mass of 1/1837 that of the hydrogen atom. Electrons are present in all matter and are responsible for negative charges such as electric current.
2. The proton - a positive unit of electricity with a mass of 1836/1837. The proton is a hydrogen ion (H^+).
3. The neutron - a neutral particle with a mass of one.
4. The positron - a particle with a unit positive charge, and a mass equal to the electron, or a positively charged electron.

Note - The nuclear charge, the most fundamental of all the properties of an element, is called its atomic number.

Some Properties of Three Component Particles of the Atom (page 322)

Name	Charge	Mass (O = 16)
The proton (p)	positive by one unit	Large 1.007582
The neutron (n)	no charge	Large 1.00893
The electron (e)	negative by one unit	Very small 0.000549

The Arrangement of Electrons in Atoms (The Bohr Theory) (page 329)

1. It is assumed that the planetary electrons are arranged in concentric orbits or shells, called energy levels, about the nucleus (page 32).
2. These shells or energy levels, in order of increasing energy, are given numbers 1, 2, 3, 4, etc. or the letters K, L, M, N, etc.
3. Each shell is divided into subshells, which are designated in order of increasing energies, as s, p, d, f subshells.

Note - Electrons always occupy the position with the lowest energy. For convenience it is assumed that each succeeding element is formed by adding one more planetary electron to the atoms of the previous element.

4. The electrons are definitely arranged in the shells as follows:
    a. The K shell (with only an s subshell) is filled with two electrons.
    b. The L shell (with s and p subshells) is filled with eight electrons.
    c. The M shell (with s, p, and d subshells) is filled with 18 electrons.
    d. The N shell (with s, p, d, and f subshells) is filled with 32 electrons.

Note - The maximum number of electrons in a given shell is given by the formula: (page 333)

$$2N^2 = \text{max. no. of electrons}$$

N is the shell number when K = 1, L = 2, M = 3, etc. Thus the:
1st shell can hold $2(1^2) = 2$ electrons
2nd shell can hold $2(2^2) = 8$ electrons
3rd shell can hold $2(3^2) = 18$ electrons

Note - Since the energies of the various subshells increase within the shells in the order s, p, d, f above the second shell the s subshell of the next higher shell will have a lower energy state than the d or f subshell in the preceeding energy level.

Ex. Subshell s of N level has a lower energy state than subshell d of M shell. Since electrons always occupy the lowest energy position a new shell of electrons is started before the d or f subshell in the lower shell is filled. However all shells above the first will fill the s and p subshells (8 electrons) before a new shell is started. Referring to the table (next page) observe that the nineteenth electron in the potassium atom goes into the s subshell of N shell, as does the 20th electron in clacium thus filling this subshell. In the scandium atom the 21st electron goes into the d subshell of M shell. Scandium is the first of nine elements to be formed by adding electrons to an inner level. Elements included in such a series are known as transition elements. (see table on page 316 and page 333). Thus elements 21 to 29 are transition elements as are elements 39 to 47 and 57 to 79.

This is very important to the chemist because just as the mass of an atom is determined by the number of protons and neutrons in the nucleus the atomic volume and the chemical behavior of an atom is determined by the number and arrangement of the electrons about the nucleus. In fact, only the electrons in the outer electron level (or shell) are concerned in chemical reactions. The electrons in the outer shell determine the combining capacity (or valence) of the atom, and are called valence electrons. Thus the transition elements which have either one or two valence electrons are very similar in chemical properties.

The Electronic Structure of Some Elements in Tabular Form

At. No.	Element	K shell	L shell		M shell			N shell				O shell							
		s	s	p	s	p	d	s	p	d	f	s	p	d	f	s	p	d	f
1	H	1																	
2	He	2																	
3	Li	2	1																
4	Be	2	2																
5	B	2	2	1															
..	..	..	..	..															
10	Ne	2	2	6															
11	Na	2	2	6	1														
..	..	..	..	..	..														
18	A	2	2	6	2	6													
19	K	2	2	6	2	6		1											
20	Ca	2	2	6	2	6		2											
21	Sc	2	2	6	2	6	1	2											
..	..	..	..	..	..	..	..	..											
28	Ni	2	2	6	2	6	8	2											
29	Cu	2	2	6	2	6	10	1											
30	Zn	2	2	6	2	6	10	2											
31	Ga	2	2	6	2	6	10	2	1										
..	..	..	..	..	..	..	..	..	..										
36	Kr	2	2	6	2	6	10	2	6										
37	Rb	2	2	6	2	6	10	2	6			1							
..	..	..	..	..	..	..	..	..	..	..	..	..							
54	Xe	2	2	6	2	6	10	2	6	10		2	6						
55	Cs	2	2	6	2	6	10	2	6	10		2	6			1			
..	..	..	..	..	..	..	..	..	..	..	..	..	..	..	..	..			
81	Tl	2	2	6	2	6	10	2	6	10	14	2	6	10		2	1		
82	Pb	2	2	6	2	6	10	2	6	10	14	2	6	10		2	2		
83	Bi	2	2	6	2	6	10	2	6	10	14	2	6	10		2	3		
..	..	..	..	..	..	..	..	..	..	..	..	..	..	..	..	..	..		
86	Rn	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6		
Total/shell		2		8		18			32				18			8			

Atomic Structure and Chemical Behavior
(pages 33-36 and 334-344)

A. Introduction

Now that we have some knowledge of the arrangement of the electrons in atoms we are now ready to consider the question, "Why do atoms combine with one another?" (page 334).

B. The inert gases are unique in being incapable of combining with other atoms, they have no valence.

Their electron structure is unique in that the outer (valence shell) level is complete. Each of these elements have eight electrons in the outermost electron shell (except He which is complete with two electrons in its only shell). The outermost shell never contains more than eight electrons and each time this maximum is reached the element is an inert gas. It may be assumed that in the inert gases the electron arrangement is a stable configuration. (Notice the atomic structure of elements number 2, 10, 18, 36, 54, and 86 in the table on page 7 of this syllabus.) Since only the inert gases have this stable configuration, all other atoms can be considered as being unstable in this respect. As a result of this instability all other atoms tend to combine with one another so as to attain the stable inert gas, configuration.

. Now we are ready to consider the question, "How do atoms combine with one another?"

Two general ways in which atoms attain stable configurations through chemical combinations are:

(1) by losing electrons to, or gaining them from, other atoms and
(2) by sharing electrons with other atoms.

1. Electrovalence - (page 35 and 335) the number of electrons which an element must gain or lose in order to reach the inert gas configuration.
    a. Electrovalence is positive when electrons are given off and negative when electrons are accepted.
2. Covalence - (page 36 and 335) when elements reach a stable configuration by sharing electrons. The number of pairs of electrons shared is the covalence. Usually atoms share a pair of electrons, one of which is provided by each atom. There is another type called the co-ordinate covalent bond which is formed by one of the elements providing both electrons.

. Atomic Structure and Properties of the Elements.

1. Elements whose atoms have one, two and three electrons in the valence shell (outermost shell) will tend to react by giving up the valence electrons forming positive ions.

Note - Ions are electrically charged particles and are represented by affixing the appropriate sign to the symbol of the atom, thus Na^+; Ca^{++}, Cl^-, $O^=$, $S^=$, etc. (page 33).

2. Elements whose atoms have seven or six electrons (just one or two short of the stable configuration of eight electrons) will tend to react by accepting electrons to form negative ions.

Note - Ions that bear unlike charges are attracted together to form compounds

$$Na^+ + Cl^- \dashrightarrow NaCl$$

sodium ion + chloride ion ----> sodium chloride

3. Factors affecting electron-attractive powers are: (page 339)
    a. Effect of the charge on the nucleus
        (1) the greater the charge on the nucleus the greater the attraction
    b. Distance of the electron from the nucleus
        (1) the closer the electrons are to the nucleus the greater the attraction.
    c. Intervening shells of electrons
        (1) as the number of electrons between the valence electrons and the nucleus become greater the force holding the valence electrons becomes less.
4. Metals and Nonmetals (page 342)
    a. Those elements which easily lose electrons to form positive ions are called metals.
    b. Elements that tend to gain electrons to form negative ions are called nonmetals.

Chapter 17

Atomic Structure and the Periodic Table
(pages 307-319)

. Introduction - Nearly 100 different elements and millions of compounds are now known. The task of learning the characteristics of the elements and their compounds would be very difficult if a method of classification were not available. However, when the elements are arranged in the order of their increasing <u>atomic number</u>, it is found that similar properties recur at regular intervals, and the elements fall naturally into groups. The arrangement thus obtained is called a <u>Periodic Table</u> (page 307).

Note - The <u>atomic number</u> of an element is the number of positive charges on the nucleus of the atoms of that element. The number of positive charges on the nucleus is equal to the number of protons in the nucleus which also equals the number of planetary electrons.

. The Periodic Law (page 311)

"The physical and chemical properties of the elements are periodic functions of their atomic numbers."

. General Aspects of the Periodic Table

Note - The extended or long form of the periodic table is to be used in this course and is shown on page 316 of the text book and on the lecture room wall.

1. <u>Groups</u> - the vertical columns are called groups. Groups 1 to 7 consist of a main group and a sub-group which are called families and are designated as Family A and Family B.
   Ex. Group I A (group one Family A) is the alkali family Li, Na, K, Rb, Cs and Fr (main group)
   Group I B (group one Family B) is the silver family Cu, Ag, and Au. (sub-group).

Note - In the short form of the periodic table the two families of each group are shown in the same columns; in the extended form they are shown in separate columns. This is the essential difference in the two forms.

Note - The only relationship between the elements in the two families of any group is that they have the same number of valence electrons.

2. <u>Periods</u> - Each horizontal row is a period. There is a total of six periods:
   a. One short period of two elements H and He.
   b. Two short periods of eight electrons each, starting with Li (at. no. 3) and Na (at. no. 11).
   c. Two long periods of 18 elements each, starting with K (at. no. 19) and Rb (at. no. 37).
   d. One long period of 32 elements starting with Cs (at. no. 55) and including 14 rare-earth elements.
3. The zigzag line that descends the table toward the right separates the metals (to its left) from the nonmetals (to its right).
4. The rare-earth elements (57-71) and the actinide series (89-98) are shown separately at the bottom of the chart.

Note - The long periods are due to the transition elements.

. Origin of the Periodic System

1. Dobereiner Triads (1829)
   a. The elements could be arranged in triads with similar properties - the atomic weight of one was the mean of the other two,
      Ex. Cl - 35.5; Br - 79.9; I - 129.9.

2. Newlands Law of Octaves (1863)
    a. The English chemist Newland stated that "the eighth element, starting from a given one, is a kind of repetition of the first, like the eight notes in music. (Note - The rare gases were not considered by Newland.)
3. Mendeleev and the Periodic Law
    a. Mendeleev, a Russian chemist, announced this law in 1869 at the same time the German chemist Lothar Meyer announced it is Germany.
    b. Mendeleev arranged the elements according to their atomic weights which gave some discrepancies (see A and K, Co and Ni, Te and I.)
    c. Moseley (1913) determined atomic numbers by bombarding metallic elements with electrons and measuring the length of X-rays given off.

Note - When the two electrons in the K shell are bombarded with electrons they give off X-rays. The wave length of the X-rays is a function of the positive charge on the nucleus, the greater the charge the more firmly and closely the electrons are held thus giving off shorter wave lengths. Hydrogen with one proton in its nucleus gives the longest X-rays, then He, Li, Be, B etc. follow. Thus each element differs from the one preceeding it by having one more proton in its nucleus.

    d. Isotopes - atoms of the same element which have different atomic weights are known as isotopes.

. Value of the Periodic Table

1. Provides a convenient and comprehensive classification of the elements, making possible the study of families of elements instead of single elements.
2. Another extremely useful way of using the Periodic Table is for the grouping together of the members of a family in such a way that the chemistry of each element is remembered by association with that of others.
3. In the past it has been useful in "checking" of experimental data."
4. Used to "predict new elements".

Note - This chapter serve merely to introduce some of the basic generalizations of the periodic system. Many references will be made to it during this and next semester.

Chapter 4

Oxygen and Ozone

n Introduction to Descriptive Chemistry

We have already mentioned that oxygen is used as a basis for determining atomic eights and combining weights (Chapter 3) because it combines with most of the other lements. The compounds formed, when oxygen combines with the other elements, is ften among the most important compounds of these elements. Oxygen is the first lement of Group 6B of the Periodic Table. It has six electrons in the L shell and s a nonmetal with a valence of negative two. It forms compounds by electrovalence, ovalence and co-ordinate covalence.

Oxygen combines with the food in the body to provide the energy to support all nimal life. It combines with fuels to provide energy for heat and power (see page 0-51). Oxygen is not a typical element therefore it is not studied as a member of family of elements. Oxygen, hydrogen, and their compound water (H_2O) are all typical and will be studied separately.

. History of Oxygen

1. Priestley and Schelle discovered oxygen, Priestley is usually given credit for the discovery. Why?
2. Lavoisier investigated "burning" and named oxygen. (page 39)

. Occurrence of Oxygen

1. Most abundant element. About one-half of the earth's mass is oxygen.
2. Occurs both in the free state and the combined state (compounds).

. Preparation of Oxygen

1. Heating oxygen compounds (HgO, $KClO_3$, $NaNO_3$, etc.) to liberate part or all of the oxygen.

$$2HgO \longrightarrow 2Hg + O_2 \qquad 2KClO_3 \longrightarrow 2KCl + 3O_2$$

Note - The rate of the decomposition of $KClO_3$ is greatly increased by the addition of manganese dioxide (MnO_2). This is one of many examples in which the rate of chemical reaction is influenced by a foreign substance. Such a substance is called a <u>catalyst</u> and may be defined as "a substance which alters the speed of a chemical reaction without itself being permanently changed." (see page 42)

2. <u>Preparation from sodium peroxide</u>.
   a. A convenient but expensive method

$$2Na_2O_2 + 2H_2O \longrightarrow 4NaOH + O_2$$

3. <u>Preparation by the electrolysis of water</u>
   a. This preparation will be demonstrated (see page 43)
4. <u>Preparation from the fractional distillation of air</u>
   a. Define distillation and fractional distillation (page 44).
   b. Nearly all the commerical oxygen used in the United States is prepared by this method.

Note - Fractional distillation of air and the electrolysis of water are commercial methods of preparing oxygen. The other methods are laboratory methods. Commerical methods are performed cheaply and on a large scale. laboratory methods are performed conveniently and on a small scale.

. Properties of Oxygen

1. Physical properties (How oxygen behaves by itself)
   a. Oxygen is a colorless, odorless, tasteless, gas which weighs 1.429 gram/liter and is slightly soluble in water.

- b. Oxygen atoms, like all active gases, form diatomic molecules (O_2). Molecules containing more than two atoms are not uncommon (O_3, P_4, etc. page 44).

2. Chemical Properties (How oxygen reacts with other substances)
   - a. Relatively inactive at ordinary temperatures but very active at high temperatures.
   - b. Oxygen reacts with most of the other elements both metals and non-metals to form oxides.
   - c. Oxygen reacts with fuels containing carbon and hydrogen to give water, carbon dioxide (CO_2) and heat.

Note - Every chemical reaction either absorbs (takes on) heat or evolves (gives off) heat. All reactions that absorb heat are endothermic reactions; all reactions that evolve heat are exothermic reactions.

Other Related Information

1. A simple test for oxygen employs a glowing splint
2. Oxidation - the process or change that takes place when oxygen combines with another substance.
3. Oxidizing agent - oxygen or any compound that supplies oxygen for oxidation.
4. Oxide - a compound composed of two elements, one of which is oxygen.

Note - There are two classes of oxides, metallic and non-metallic oxides. The properties which distinguish the two classes of oxides also serve to distinguish metallic and non-metallic elements. The most striking of these properties is their reaction with water. When dissolved in water metallic oxides form bases while non-metallic oxides form acids. (Chap. 13).

5. Combustion - an oxidation reaction accompanied by the evolution of heat and light. (Combustion is discussed in Chapter 30.)
6. Kindling temperature - the temperature at which a substance will undergo self-sustaining combustion.
7. Spontaneous combustion - when a substance takes fire on its own accord - essential conditions for spontaneous combustion are: (a) an existing slow oxidation reaction and (b) good heat insulation (page 40).

Rate of Chemical Reaction

1. Factors that affect the rate of chemical reaction are: (page 49)
   - a. The nature of the reacting substance - size of particles, solutions, etc.
   - b. Temperature - the higher the temperature the faster the reaction.
   - c. Concentration - the greater the amount in a given volume the faster the reaction.
   - d. Amount of surface exposed - the greater the surface area exposed the faster the reaction.
   - e. Catalyst - will increase or decrease the rate.

Ozone (O_3)

Preparations

1. Ozone is formed by sunlight in the upper layers of the atmosphere.
2. Laboratory preparation - ozone is formed when an electric spark is passed through oxygen (air).

Properties

1. Physical - pale-blue colored gas with a powerful odor.
2. Chemically - ozone is like oxygen except more active and is a good bleaching agent.

Uses

1. oxidizing agent
2. Bleaching agent
3. Disinfectant
4. Deodorizer
5. Sterilizing drinking water

Chapter 5

Gases and the Gas Laws: The Kinetic Molecular Theory

Note - Substances, in general, are capable of existing in three different forms or <u>states</u> - solids, liquids, and gases.

A. The Characteristics of the Gaseous State

1. Depending upon the temperature all substances theoretically could exist as gases.

Note - Some substances decompose on being heated before they reach the liquid or gaseous states.

2. The volume of a given mass of gas is not fixed but it expands to fill whatever container it is in. It takes the shape of the container.
3. Gases diffuse or spread through space. The molecules are in rapid motion. The speed of the molecules varies directly with the temperature.
4. The distance between gaseous molecules is large thus the molecules can be pressed closer together by pressure; gases are compressible.
5. Due to the molecules striking the walls of the container gases exert a steady pressure.
   a. The pressure of gases is measured by a barometer.

Note - To provide a uniform basis for the comparison of the volumes of gases <u>standard conditions of temperature and pressure</u> (S.T.P.) have been established. Standard temperature is the freezing point of water 0°C. (273°A.) and the pressure 760 mm. of mercury (the average atmospheric pressure at sea level and termed <u>one atmosphere</u>) is <u>standard pressure</u>.

B. The Gas Laws

All gaseous substances or mixtures behave alike with respect to changes in temperature or pressure, their solubility, rate of diffusion, etc. These common modes of behavior for all gases are known as the gas laws. These laws are generally known by the name of the man who first noted the regularities.

1. <u>Boyle's Law</u> - the volume of a gas varies inversely as the pressure, if the temperature remains constant (page 58). (for example if the pressure on a volume of gas is doubled the volume will be reduced by one-half, etc.) (solve typical problems).
2. <u>Charles's Law</u> - the volume of gas at constant pressure varies directly as the absolute temperature.
3. <u>The Absolute Temperature Scale</u> (page 62)
   a. Absolute zero = -273°C.
   b. Absolute zero is the temperature at which a gas would occupy no space.
   c. Absolute degrees are equal to centigrade degrees.
   d. To change absolute° to centigrade° add algebracially 273° to $^\circ$C. (solve typical problems and also solve problems involving both pressure and temperature changes.)
4. <u>Dalton's Law</u> - partial pressures of gases - in a mixture of gases every gas exerts the same pressure that it would if it alone were confined in the same volume, and the total pressure is the sum of these partial pressures.
   a. This law must always be applied when gases are collected over water (see Fig. 47, page 66).

Note - As a minimum, the student should be able to calculate the volume of a gas at standard conditions from measurements made at laboratory conditions and over water, and vice-versa. The student should learn to reason out each problem on the basis of the law being applied, memorization of formulas is discouraged.

5. Graham's Law - The rate of diffusion of gases
    a. Gaseous molecules will spread or separate to fill whatever space they are in; this is called diffusion.
    b. "The rate of diffusion of a gas is inversely proportional to the square root of its density or molecular weight." (work typical problems.)

Note - The lighter a gas, the faster it diffuses.

6. Avogadro's Law
    a. "Equal volumes of gases, under the same conditions of temperature and pressure, contain the same number of molecules."
7. Applications of Avogadro's Law
    a. By measurements it can be shown that 32 grams of oxygen (one gram molecular weight) occupys 22.4 liters of S.T.P. Therefore by applying Avogadro's law 22.4 liters of any other gas at S.T.P. will contain the same number of molecules as are in 22.4 liters of oxygen. Then a gram molecular weight of any gas will contain the same number of molecules. This information can be used in determining molecular weights of gases. (Chapter 8)
    b. The density (weight/liter) of a gas can be determined by dividing the molecular weight by 22.4 or the molecular weight can be determined by using the weight/liter etc.

. The Kinetic-Molecular Theory of Gases

1. Principle concepts of the Kinetic-molecular theory
    a. All gases are made up of extremely minute particles (molecules), separated at distances that are large compared with the size of the molecules themselves.
    b. These molecules are in motion, and move in straight lines until they suffer a change of direction as a result of a collision with another molecule or with the walls of the containing vessel.
    c. The collisions are perfectly elastic; that is, there is no average loss of energy as a result of them.
    d. The kinetic energy of the moving molecules is proportional to the absolute temperature.

Note - From these concepts the student should be able to account for such facts as:

(1) gases are highly compressible
(2) gases diffuse and fill any container
(3) gases do not settle
(4) mixtures of gases are homogeneous
(5) heavier gaseous molecules diffuse more slowly than lighter molecules

Chapter 6

Hydrogen and Hydrogen Peroxide

. History of Hydrogen
1. Cavendish, a famous English scientist, prepared pure hydrogen in 1766 and called it "inflammable" air.
2. Lavoisier recognized hydrogen as a constituent of water and named it hydrogen from the Greek words "water former".

. Occurrence
1. Hydrogen makes up about 1% of the earth's crust, mostly as compounds and is found free in volcanic gases and in the upper atmosphere.
2. Common compounds of hydrogen are water, petroleum, natural gas and many others.

. Preparation
1. Laboratory methods
    a. moderately active metals (Zn, Mg, Fe, Al, etc.) react with acids to give hydrogen.

       Ex. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

    b. very active metals (Na, K etc.) react with water to give hydrogen.

       Ex. $2Na + 2H_2O \longrightarrow 2NaOH + H_2 \uparrow$

    c. certain elements (Zn, Al, Li) will react with strong bases to give hydrogen.

       Ex. $Al + NaOH \longrightarrow Na_3AlO_3 + H_2 \uparrow$

2. Commercial methods
    a. Electrolysis of water (see preparation of oxygen)

       Ex. $H_2O + \text{elect.} \longrightarrow H_2 \uparrow + O_2 \uparrow$

    b. When steam (H_2O) is passed over red hot coke (C) hydrogen is formed.

       Ex. $H_2O + C \longrightarrow CO + H_2 \uparrow$

    c. When steam is passed over hot iron hydrogen is formed.

       Ex. $Fe + H_2O \longrightarrow Fe_3O_4 + H_2 \uparrow$ (at high temp.)

. Electrochemical series (activity series of the metals) page 82.
1. An arrangement of the metals in order of chemical activity usually hydrogen is used as the reference element.

Note - The most active elements are at the top of the table the least active at the bottom. Hydrogen has an intermediate position and any element above it will replace hydrogen from dilute acids (H_2SO_4, HCl, $HC_2H_3O_2$) metals below hydrogen in the series will not replace hydrogen from these acids.

Note - Hydrogen, which is a reducing agent cannot be prepared from oxidizing acids such as nitric acid.

Note - Any free metal will replace any of the metals below it in the activity series from their compounds.

Ex. $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu \downarrow$ $Zn + AgNO_3 \longrightarrow Zn(NO_3)_2 + Ag \downarrow$

Note - Arrows following a formula in equation for reactions as in the above equations indicate a precipitate when pointing downward and a gas when pointing upward.

. Properties of Hydrogen

1. Physical properties - hydrogen like oxygen, is odorless, colorless, and tasteless. It is the lightest element weighing less than 0.09 grams/liter at S.T.P. It is slightly soluble in water.
2. Chemical properties -
   a. molecular hydrogen (H_2) is not very active and requires either heat or a catalyst.
   b. Hydrogen combines with many of the other elements to form hydrides.
   Ex. NaH, CaH_2, NH_3, H_2S, HCl, HI, H_2O, etc.
   c. Hydrogen acts to reduce metallic oxides to give free metals and water.

$$\text{Ex.} \quad CuO + H_2 \longrightarrow Cu + H_2O$$

Note - A compound from which oxygen is removed is said to be reduced, and the substance that removes the oxygen is called the reducing agent. Thus in the above reaction hydrogen is the reducing agent and the copper oxide is reduced. Therefore hydrogen is a reducing agent.

   d. Hydrogen reacts with "unsaturated" organic compounds, such as liquid vegetable oils, to change them to solid fats by a process called hydrogenation. Products produced by this process are shortening, oleomargarine, etc.

. Uses of Hydrogen (page 90)

1. Hydrogen is used
   a. in producing ammonia Ex. $N_2 + 3H_2 \longrightarrow 2NH_3$
   b. for "hardening" fats for foods.
   c. for hydrogenation of coal to form lubricating oils, gasoline, etc.
   d. as a reducing agent.
   e. for many laboratory purposes.

Hydrogen Peroxide

. General

Oxygen and hydrogen combine to form another compound other than water, which is called hydrogen peroxide (H_2O_2). In this compound just twice as much oxygen is combined with the hydrogen as in water.

. Preparation

1. Hydrogen peroxide can be prepared by:
   a. treating sodium peroxide with hydrochloric acid.
   Ex. $HCl + Na_2O_2 \longrightarrow H_2O_2 + NaCl$
   b. treating barium peroxide (BaO_2) with sulfuric acid.
   Ex. $H_2SO_4 + BaO_2 \longrightarrow BaSO_4 + H_2O_2$

. Properties of Hydrogen Peroxide

1. Hydrogen peroxide is a colorless, sirupy liquid which decomposes readily into oxygen and water.
2. Since oxygen is readily liberated it is a strong oxidizing agent.

. Uses of Hydrogen Peroxide are:

1. oxidizing agent 2. bleaching agent 3. rocket fuel 4. germicide

Chapter 7

The Atmosphere, Nitrogen, Carbon Dioxide, Water and the Rare Gases

A. The Atmosphere
1. The term atmosphere is applied to the mixture of gases, called air, that surrounds the earth.
2. The substances present in the air are:
   nitrogen, oxygen, argon, carbon dioxide, hydrogen, neon, helium, krypton, xenon and water vapor (see table, page 97 for percentages)

B. The function of some of the components in the atmosphere are:
1. Nitrogen - (nitrogen will be studied later - chapter 19)
   a. serves to dilute the oxygen and as an insulating material
2. Oxygen - (see section on oxygen (Chapter 4)
3. Water Vapor
   a. Water vapor in the air accounts for all precipitation (rains, snow, etc.)
   b. Water vapor serves to regulate the temperature
4. Carbon dioxide (will be studied later - Chapter 24)
   a. Carbon dioxide is used by plants in the process of photosyntheses. The process by which carbohydrate foods are made.

Water

A. Importance of Water
1. Water is essential to life (the human body is almost 70% water)
2. The properties of water are used as standards for:
   a. defining calories
   b. reference points on centigrade thermometers
   c. unit of volume, the liter
   d. relative densities - the specific gravity of liquids and solids are usually referred to water at 4°C. as having a density of one
   e. specific heat - the specific heat of water is one calorie

Note - The specific heat of a substance is the number of calories of heat required to raise the temperature of one gram of the substance 1°C.

B. Physical Properties of Water
1. Water is a colorless, odorless, tasteless compound.
2. The heat of vaporization of water, that is, the heat required to evaporate one gram, is 539 calories, which by comparison is very high.
3. Water has its maximum density at 4°C. thus ice freezes on top of a body of water. Why is this important?
4. The density of ice is less than the density of water. Why is this important?
5. The heat of fusion - that is the amount of heat, measured in calories, involved when a gram of water freezes or a gram of ice melts. The heat of fusion of water is 80 cal/g.
6. Water is an excellent solvent

C. Chemical properties of water
1. Water is very stable toward heat
2. Water reacts with oxides (chapter 4) to form acids and bases
3. Water reacts with some salts by <u>hydrolysis</u> to form acids or bases (Chapter 22, page 434)
4. Water reacts with many salts to form hydrates (page 110)
5. Water has a catalytic effect upon many chemical reactions.

The Rare Gases (see page 115)

Chapter 8

Law of Chemical Combination; Atomic Theory; Atomic Weights and Molecular Weights; Isotopes

A. Laws of Chemical Combination (see page 3 of syllabus)

B. Atomic Theory (see page 4 of syllabus)

C. Atomic Weights, Combining Weights, and Valence (page 130)

1. The atomic weight unit is equal to 1/16 the weight of the oxygen atom; thus, the weight of an oxygen atom is arbitrarily assigned the weight 16 atomic weight units. By using this procedure the atomic weights of all the atoms are one or greater. Oxygen was chosen here because it combines with most of the other elements.
2. The combining weight of an element is the same as the equivalent weight which may be defined as the weight in grams of the element which combines with 8 grams of oxygen or its equivalent (1.008 grams of hydrogen, 35.45 grams of chlorine, etc.)

Note - The atomic weight of the atom of an element is its combining weight multiplied by its valence, (combining weight x valence = atomic weight) the valence may be found by dividing the atomic weight by the combining weight (atomic weight/combining weight = valence.)

3. The valence of an atom is the number of hydrogen atoms with which it will combine or will replace. Valence is the chemical name for the combining capacity of the elements.

Note - For an accurate method of determining equivalent weights see Experiment 14 (page 61) in the laboratory manual.

D. Gram-atomic Weights

1. The gram-atomic weight of an element is the atomic weight expressed in grams.

Note - For convenience the chemist usually expresses atomic weights in grams, although pounds, tons, ounces, grains, or any other units of weight could just as well be used.

E. Gay-Lussac's Law of Combining Volumes (page 131)

1. "When two gases combine, there is a whole-number ratio between their volumes, and between the volume of either one of them and that of the product if it too is a gas."

F. Avogadro's Law (page 131)

1. Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of molecules (see Chapter 5).

G. Molecular Weight and Gram Molecular Weight (page 136)

1. The molecular weight is the sum of the atomic weights in the molecule.
2. The gram molecular weight is the same number of grams of the substance as the molecular weight. Ex. carbon dioxide has a molecular weight of 44 this means that a molecule of CO_2 will weigh 44 when an atom of oxygen weighs 16, or to find the molecular weight of H_2SO_4 take the sum of the atomic weights

$H = 1 \quad S = 32 \quad O = 16$ or $2H = 2 \quad S = 32$

$4O = 64$ then $2 + 32 + 64 = 98$ molecular weight of H_2SO_4. Thus 98 grams is a gram molecular weight of the substance.

The gram molecular weight (G.M.W. or mole) of any substance is the molecular weight expressed in grams.

Note - A gram molecular weight of all substances have the same number of molecules. This number is 6.023×10^{23} and is known as <u>Avogadro's number</u>.

3. The <u>gram-molecular</u> <u>volume</u> is the space occupied by a gram-molecular weight of a compound. For solids and liquids this volume varies from substance to substance but the gram-molecular volume of all gases is the same when under conditions of equal temperature and pressure. The gram-molecular volume of gases can be measured and is found to be 22.4 liters - therefore a gram molecular weight of any gas will occupy 22.4 liters of space at S.T.P.

Note - The student should be able to calculate the molecular weight of a gas if the weight per liter is given, or to calculate the density (wt./1.0 if the molecular weight is known.

Note - Very acurate atomic weights can be calculated by using the mass spectrometer (page 143-144).

4. Atoms of the same elements which have different masses are called <u>isotopes</u>. The difference in masses is due to some atoms having more neutrons in the nucleus than others, while the number of protons and electrons are the same, therefore their chemical properties are the same, the only variation is in the weight of the nucleus.

Chapter 9
Formula, Valence, Equation, Chemical Calculations

Formulas of Compounds (see Chapter 3 and page 150)

1. The formula indicates the composition of the compound by showing the following
    - a. The elements present in the compound.
    - b. The number of atoms of each element present in the compound.
    - c. The combining proportions (by weight) of the elements present in the compound.
    - d. The molecular weight of the compound,
    - e. The weight of a volume of gas.
2. A formula represents
    - a. A single molecule of the compound
    - b. A definite weight of the compound (the gram-molecular weight or one mole of the compound). If the compound is in the gaseous state the formula will also represent the volume occupied by the compound.

Kinds of Formulas are:

1. <u>Emperical</u> - which give the ratio of the different atoms present in the compound.
2. <u>Molecular</u> - which give a true representation of the molecules.
   Ex. The molecular formula for ethylene gas is C_2H_4. Its emperical formula is CH_2.
3. <u>Structural</u> - structural or graphic formulas show the relationship of the atoms to each other in the compound. (see page 154 - these will be studied later.)

Radical Ions or Complex Ions

1. A group of two or more atoms, which are chemically combined and stay together through many chemical reactions, acting as a single atom is a radical.
   Examples of radicals are: SO_4 - sulfate; NH_4 - ammonium; NO_3 - nitrate; OH - hydroxide; CO_3 - carbonate.

Note - Radicals have valences like other atoms and from positive or negative ions. Ex $SO_4^=$ sulfate ion; NH_4^+ ammonium ion; NO_3^- nitrate ion, etc.

The student should memorize the names and valences of the ions given in the table on page 156.

E. Equations (page 157)
  1. Chemical reactions are represented by equations.
  2. The equation shows by symbols and formulas the rearrangement of atoms and molecules during a chemical reaction.
  3. The steps in writing equations are:
    a. Recall what substances react (the reactants) and what products are formed.
    b. Write down the formulas of the reactants to the left and the formulas of the products to the right of the reaction symbol, usually shown as an arrow ($\longrightarrow$).
    c. Then balance the equation by writing coefficients before the formulas to show the correct ratios. When the equation is balanced the same total number of atoms of each of the elements must be shown in the initial substances and in the products. (Law of conservation of mass)

F. Generalized statements of chemical facts are useful in predicting what products will be formed when a reaction takes place.
  1. Some generalized statements of facts are:
    a. Oxides of metals react with water to form bases:

$$\text{Ex. } CaO + H_2O \longrightarrow Ca(OH)_2$$

    b. Oxides of non-metals react with water to form acids:

$$\text{Ex. } SO_3 + H_2O \longrightarrow H_2SO_4$$

    c. The metals above hydrogen in the electromotive series will replace hydrogen from acids:

$$\text{Ex. } Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$$

    d. Very active metals will replace one half of the hydrogen from water:

$$\text{Ex. } 2Na + 2HOH \longrightarrow 2NaOH + H_2 \uparrow$$

    e. The oxides of the noble metals can be decomposed by heat to yield oxygen and the free metal:

$$\text{Ex. } 2HgO \longrightarrow 2Hg + O_2 \uparrow$$

Note - More information about reactions are conveyed by certain other conventions. For example in the reaction

$$2KClO_3 \xrightarrow[\Delta]{MnO_2} 2KCl + O_2 \uparrow$$

the MnO_2 is used as a catalyst and the Δ indicates heat was applied; the upward pointing arrow $\uparrow$ indicates a gas (O_2) passed off. A downward pointing arrow $\downarrow$ in an equation indicates a precipitate is formed. In addition to these, sometimes the heat that is absorbed or evolved during a reaction is indicated by showing the number of calories involved.

$$\text{Ex. } Fe + S \longrightarrow FeS + 24{,}000 \text{ calories}$$

$$\text{or } C + CO_2 \longrightarrow 2CO - 41{,}250$$

In the first reaction 24,000 cal. of heat was produced. In the second equation 41,250 cal. of heat were absorbed.

G. Weight calculations from equations (page 161)

Note - At this stage of the course it is very desirable that each student become quite adept at working so called "yield" problems. Being able to solve problems of this type enables the student to calculate how much material he can prepare from a given weight of starting material or how much of the reactants are required to product a given amount of products.

Here again it is suggested that the blind use of proportions by the student is not good. Most problems in freshmen chemistry are solved by reasoning based on the laws or principles being applied and by elementary arithmetic. A good procedure to follow in working problems is to write down the information given and what information is needed and what you are trying to find. An example of a typical problem and the steps used to solve it follows.

Problem - What weight of oxygen can be prepared by heating 40 grams of potassium chlorate ($KClO_3$)?

Solution -

1. Write and balance the equation $2KClO_3 \longrightarrow 2KCl + 3O_2$

Note - The balanced equation tells us that two G.M.W. (or moles) of $KClO_3$ decomposes to give 3 G.M.W. of oxygen, or 2 x 122.6 = 245.2 grams of $KClO_3$ will yield 3 x 32 = 96 grams of oxygen.

2. Find the amount of oxygen one gram of $KClO_3$ will produce. one gram of $KClO_3$ gives 96/245.2 = 0.3915 g. of O_2 (by simple arithmetic).
3. Find the number of grams produced by 40 grams of $KClO_3$.
   o.3915 x 40 = 15.66 grams of O_2 answer

Note - This problem could be solved as readily to find the number of liters of oxygen at S.T.P. by substituting the gram molecular volume for its equal the gram molecular weight.

Note - The only way to become efficient at problem solving is by practice. You will find problems of this type frequently in the remainder of the course and in the course for next semester.

Chapter 10

The States of Matter and Their Transitions: Crystal Structure

. Kinetic Energy of Particles - the tendency of the particles in any body of matter to fly apart.

1. The Kinetic energy of molecules ($\frac{1}{2}mV^2$) is proportional to the absolute temperature (chapter 5).

. Van der Waal's Forces - the tendency of the particles in any body of matter to be drawn together by attractive forces.

Note - The Van der Waal's forces are very great between particles that are close together and fall off rapidly with increasing distances between particles. Thus they are effective only when the particles are very close together.

. Liquid-Vapor Equilibrium

1. Energy relations in states of matter
   - a. In gases the kinetic energy predominates, or it's greater than the Van der Waal's forces. Thus gases diffuse to fill any container.
   - b. In liquids the two forces (kinetic and attractive) are approximately equal, allowing the molecules to move readily over each other.
   - c. In solids the attractive energy predominates and holds the molecules in rigid position confining their motion to vibrations.

Note - Since the kinetic energy is proportional to the absolute temperature the state of any body of matter is determined by the temperature.

Transition Between the Liquid and Gaseous State

. Evaporation

1. The molecules in a liquid move about and molecules with the greatest kinetic energy escape into the space above the liquid.
2. Since the molecules with the greatest kinetic energy escape leaving the slower molecules the average kinetic energy is decreased - thus the remaining liquid is cooled.
3. The rate of evaporation (the tendency for the liquid molecules to escape) increases with an increase in temperature. Why?

. Vapor Pressure

1. An evaporation in a closed container will establish equilibrium between the liquid and gaseous phase. (The rate of evaporation equals the rate of condensation.)
2. The pressure exerted by a vapor in equilibrium with its liquid in a closed vessel is defined as the vapor pressure which is usually expressed in millimeters of mercury.

Note - The boiling point of a liquid is defined as the temperature at which the vapor pressure equals the atmospheric pressure.

Note - Some liquids change readily to the vapor state when left in an open container because their vapor pressure is high, thus their boiling points are low, such liquids are said to be volatile. Liquids with opposite properties are said to be non-volatile.

. Heat of Vaporization and Condensation

1. The heat of evaporation (vaporization) is defined as the number of calories of heat required to change one gram of substance from the liquid state to the vapor state without changing the temperature.

Note - The heat of vaporization of water is 539 calories/gram. Conversely, the condensation of a gram of a gas or vapor to a liquid liberates a quantity of heat exactly equal to the heat of vaporization. This is called the <u>heat of condensation</u>.

. Critical Points
  1. <u>Critical Temperature</u> - "The critical temperature of a gas is the temperature above which that gas cannot be liquified no matter how great a pressure is applied".
  2. <u>Critical Pressure</u> - "The pressure required to liquify a gas at its critical temperature".

  Note - The kinetic energy of gases can be decreased by applying pressure and cooling; thus gases can be liquified, giving off heat. The same amount of heat is absorbed when the process is reversed, the mechanical refrigeration operates on this principle. (Page 178)

Transition Between Solids and Liquids or Gases

. Transition Between Solids and Liquids or Gases
  1. Two forms of solid substances are:
     a. <u>Crystaline solids</u> - with definite crystaline form
     b. <u>Amorphous solids</u> - without definite crystaline form
  2. Freezing point
     a. The <u>freezing point</u> is defined as "the temperature at which both solid and liquid will remain in contact with each other without change of temperature".
     b. The <u>melting point</u> is the temperature at which a solid will change to a liquid.
  3. <u>Heat of Fusion</u> - The heat of fusion is defined as the amount of heat that is absorbed when one gram of a solid at the freezing point changes to a liquid at the same temperature.

  Note - The heat of fusion of water is 80 cal. The same quantity of heat is given up when a gram of liquid changes to a solid at the same temperature. This is the heat of <u>solidification</u>.

  4. <u>Solids have vapor pressures</u> - In solids some of the molecules have enough kinetic energy to escape. The vapor pressures of solids are usually very small; however some solids (iodine, "dry ice", etc.) have relatively high vapor pressures.
     a. <u>Sublimation</u> - the process of changing a solid directly to a gas, and condensing the gas back to the solid state.
     b. <u>Heat of Sublimation</u> - The heat of sublimation is the number of calories required to change one gram of substance from the solid to the gaseous state without changing the temperature.

  Note - The process of converting a liquid to a gas and condensing it back to the liquid state is called <u>distillation</u>. Sublimation and distillation are purification processes.

  Note - The student should practice solving problems involving specific heat, heat of vaporization, etc. For ex. How many calories of heat would be required to change 10 grams of ice at 0° to steam at 100°C.? (for other problems see page 190.)

. Crystalographic Systems
  1. Each individual compound has a definite crystaline structure, therefore compounds may be classified by the crystal lattice they form.
  2. Every crystal has a definite geometric form. There are six crystalographic systems: (see pages 183-88)
     a. isometric or regular system b. tetragonal system c. orthorombia system d. monoclinic system e. triclinia system f. hexagonal system

Chapter 11

Solutions

Solution - a homogeneous molecular mixture, the composition of which may be varied continuously within limits.

- 1. The two constituent parts of a solution are:
    - a. Solute - the component that is dissolved.
    - b. Solvent - the component that does the dissolving.

Note - When it is difficult to distinguish the solvent and solute, the component that is used chemically may be defined as the solute and the other component the solvent.

- 2. Solubility - The amount of solute that can be dissolved in a given amount of solvent. The solubility is usually expressed as grams of solute/100 grams of solvent.

Types of solutions in terms of physical state of components

- 1. Nine possible types of solutions and examples are:
    - a. gases in gases -- the atmosphere
    - b. liquid in liquids -- alcohol-water
    - c. gases in liquids -- foams
    - d. solids in liquids -- sugar-water
    - e. solids in solids -- alloys of metals
    - f. solids in gases -- smoke in air
    - g. liquids in gases -- fog or clouds in air
    - h. gases in solids -- pumice stone
    - i. liquids in solids -- jellies
- 2. Relative importance of the types of solutions (the first five are the most important)
    - a. Mixtures of gases in gases were discussed in Chapter 5 (Dalton's Law, page 65). Gases mix with each other in all proportions.
    - b. Solutions of gases in liquids - a liquid will dissolve a limited amount of gas depending upon the following factors:
        - (1) Effect of specific properties - (see page 193).
        - (2) Effect of pressure - gases are more soluble at higher pressures.

Note - A statement of this: "The weight of gas dissolved by a definite quantity of a given solvent is directly proportional to the pressure, provided the temperature remains constant", is known as Henry's Law.

        - (3) Effect of Temperature - with a raise in temperature gases become less soluble.
    - c. Solutions of liquids in liquids have the following important properties:
        - (1) The vapor pressure of a solution made by mixing two liquids is never as great as the sum of the two vapor pressures of the liquids before they are mixed.
        - (2) Since the vapor pressure is lowered the boiling point is increased. (see Chapter 10)
        - (3) Fractional distillation - two liquids whose boiling points are different can be separated by distillation called fractional distillation.
        - (4) A constant-boiling solution - the vapors given off have the same percentage composition as the solution, therefore the concentration of the solution is unchanged or the component parts cannot be separated by distillation.

d. Solutions of solids in liquids

(1) In general solubility increases with an increase in temperature. There are some exceptions to this rule, for ex. Calcium acetate is more soluble in cold water. (Explain the solubility curves (Fig. 115, page 197).

(2) The heat of solution is the amount of heat which is absorbed or evolved when one gram of a solid dissolves, any solid that absorbs heat when it dissolves will be more soluble in hot water and the reverse is also true.

(3) The heat of dilution is the heat involved when a concentrated solution is diluted.

. Concentration of Solutions

The term concentration denotes "the quantity of solute dissolved in a given quantity of solvent or of solution" (page 199).

1. Concentrated and dilute - are loose qualitative terms that are used for comparisons.
2. Molar solution - a gram molecular weight of solute per liter of solution.
3. Normal solution - a gram equivalent weight of solute per liter of solution. (see Chap. 13, page 236)
4. Molal solution - a gram molecular weight per 1000 grams of solvent.
5. Standard solution - any solution whose exact concentration (molarity or normality) is known (see page 236).
6. Saturated solution - a solution that has all the solute the solvent will hold, or a solution in which the dissolved solute is in equilibrium with undissolved solute.
7. Super saturated solution - a solution that has more solute than the solvent will ordinarily hold - explain how to make.
8. Unsaturated solution - a solution that has less solute in the solvent than it would hold under these conditions.

Note - In expressing concentrations the word "per" is always used. "Per" can be translated into mathematical terms by means of the division sign ÷, so that the above may be expressed as:

$$\text{concentration} = \frac{\text{quantity}}{\text{volume}}$$

$$\text{molarity} = \frac{\text{number of moles}}{\text{number of liters}}$$

$$\text{normality} = \frac{\text{number of equivalents}}{\text{number of liters}}$$

$$\text{molality} = \frac{\text{grams of solute in 1000 g. of water}}{\text{gram-mol. wt. of solute}}$$

when weight, molarity and volume are required:
grams = mol. wt. x molarity x volume in liters
grams = eq. wt. x normality x volume in liters

$$M = \frac{\text{grams}}{\text{mol. wt. x volume}}$$

$$N = \frac{\text{grams}}{\text{eq. wt. x volume}}$$

. Colligative Properties of Solutions

1. Those properties of a solution that depend on the number of solute particles in a given amount of solvent.
    - a. Lowering of vapor pressure
        - (1) Raoult's Law - the lowering of the vapor pressure of a solvent is directly proportional to the weight of the solute dissolved in a definite weight of solvent.

   Note - The kind or size of particles make no difference as long as the substance is not volatile or does not separate into ions.
    - b. Elevation of the boiling point
        - (1) If the vapor pressure is lowered the boiling point will be proportionally raised. In a one molal solution the boiling point would be raised 0.518°. In a 2 molal solution the boiling point elevation would be twice as great ($2 \times 0.518 = 1.036^\circ$) etc.
    - c. Lowering of freezing point
        - (1) The freezing point depression is proportional to the concentration of the solution. For ex. a 1 molal solution will freeze at a temperature 1.86° lower than the pure solvent.

   Note - Practical applications of these properties are anti-freeze mixtures for automobile radiators, cooling mixtures such as ice cream salt, determination of molecular weights of compounds, etc.

. Osmosis

1. If a semipermeable membrane separates two water solutions of different concentrations, the rate of diffusion of water molecules out of the concentrated solution is less than in the opposite direction. The net result is water passes from the dilute to the concentrated solution. This phenomenon is known as osmosis.

   (A semipermeable membrane is one that allows one component of a solution, but not the other, to pass through it.)
2. <u>Osmotic</u> <u>pressure</u> - That pressure which, when applied to a solution, will just prevent the entrance of solvent into it through a semipermeable membrane is the osmotic pressure of the solution.
3. The laws of osmotic pressure show that osmotic pressure is:
    - a. Approximately proportional to concentration.
    - b. Independent of the nature of the solute particles.
    - c. Proportional to the absolute temperature.

   or

   "The osmotic pressure of a solution is approximately equal to the gas pressure which the solute would exert if it were a gas occupying the same volume as the solvent, at the same temperature."

Chapter 12
Ionization

Ionization, Electrolytes and Non-electrolytes

1. Ionization - the term that is applied to the process that takes place when some molecules dissociate into charged particles, called ions, upon being dissolved in certain solvents (usually water).
2. Electrolyte - a substance which, when dissolved in certain solvents, or when fused (melted), dissociates into ions.
3. Non-electrolyte - a substance that does not separate into ions when dissolved in certain solvents such as water.

Abnormal Properties of Solutions of Electrolytes

1. In solutions of electrolytes the boiling bpoint and freezing point etc. are abnormally affected because, it is recalled, and colligative properties of solutions are a result of the number of particles present; therefore if a substance ionizes to form two particles per molecule the change in boiling point etc. would be twice as great.

Theory of Ionization

1. Arrhenius Theory (1887) makes the following assumptions:
   - a. Molecules of electrolytes, when dissolved in water dissociate into charged particles called ions.
   - b. There is the same number of positive as negative charges on the ions, that is, the solution as a whole is neutral.
   - c. The process of ionization is reversible so that such solutions may contain both ions and molecules in a state of equilibrium.
   - d. Chemical reactions in ionic solutions take place between the ions, the molecules will supply the ions however until they are used up.
   - e. Ions are just as effective as molecules in changing the colligative properties of solutions.

Evidence to Support the Theory of Ionization

1. Solutions of electrolytes will conduct an electric current.
2. The abnormal lowering of the freezing point and elevation of the boiling point.
3. Abnormal osmotic pressure.
4. High melting points of certain compounds.

The Degree or Percentage of Ionization

1. Strong electrolytes - a substance that is almost completely ionized or almost all of the molecules are separated into ions.
2. Weak electrolytes - a substance that is relatively only slightly ionized, or most of the substance remains as molecules.
3. The degree of ionization can be determined by the abnormality of the colligative properties.
   For ex. - If KCl ionizes to form two particles per molecule

$$KCl \longrightarrow K^+ + Cl^-$$

The boiling point would be elevated exactly twice as much and the freezing point depressed twice as much as they would be if no ionization took place. Thus the colligative properties are proportional to the degree of ionization.

F. Interionic Attraction

1. Unlike electrical charges attract each other and the attraction is a function of the distance between the charged particles. Thus in concentrated solutions the attraction between oppositely charged ions is very great because the ions are close together. In dilute solutions the attraction is small because the particles are far apart.

2. The electrostatic forces are greater, the higher the charge on the ions. For ex. The attraction between Mg^{++} and $SO_4^{=}$ is greater than the attraction between Na^+ and Cl^-.

Note - In ionic solutions chemical reactions take place between the ions, and the properties of ionic solutions are actually the properties of the ions (see page 216).

G. Electrolytic Conduction - the movement of charged particles through a solution resulting in an electric current.

1. The conducting of a solution will depend on:
   - a. The number of ions present.
   - b. The speed of each ion present.
   - c. The charge carried by each ion.

Electrochemistry

A. Electrolysis - the decomposition of a compound by an electric current.

1. Definitions
   - a. Cathode - the electrode by which electrons enter the solution, the negative electrode.
   - b. Anode - the electrode by which electrons leave the solution, the positive electrode.
   - c. Anion - the ions attracted by the anode, negatively charged ions.
   - d. Cation - the ions attracted by the cathode, positively charged ions.

Note - Electrolysis takes place in water solutions or in melted compounds. The anions and cations move toward the anode and cathode respectively. At the electrode a transfer of electrons take place. For ex. in the electrolysis of melted NaCl the reactions are:

cathode - $Na^+ + e^- \longrightarrow Na$

anode - $2Cl^- \longrightarrow 2e + Cl_2$

   - e. Voltaic cell - is a device used to convert chemical energy into electrical energy. The reaction takes place spontaneously. The two reactants must be separated into two partial reactions, in which electrons are lost by one substance and gained by the other. The two electrodes must be connected by an external wire to complete the circuit.
   - f. Electrolytic cell - Electrical energy is applied from an outside source (a battery) causing a non-spontaneous reaction to take place. Example - Any typical electrolysis set-up (see Fig. 124, page 217).

2. Application of Electrolysis
   - a. Electroplating
   - b. Electrotyping
   - c. Electrorefining of metals
   - d. Electroreductions of metallic compounds

3. Faraday's Laws of Electrolysis are:
   - a. "The weight of a given element liberated at an electrode in electrolysis is directly proportional to the quantity of electricity passed through the solution".

  - b. "The weight of different elements liberated by the same quantity of electricity are proportional to the equivalent weights of the elements". Thus the quantity of electricity required to deposit one gram-equivalent weight of an element is the same for all elements.
4. Other electrical terms are:
  - a. __Coulomb__ - the quantity of electricity conducted by an electric current of one ampere folowing for one second (amp. x sec. = coulombs).
  - b. __Faraday__ - 96,500 coulombs - the amount of electricity required to liberate one gram-equivalent weight of substance at an electrode.

Note - This is obvious since a gram molecular weight of any substance contains the same number of molecules (Avegadro's number).

  - c. __Ohm__ - unit of resistance.
  - d. __Ampere__ - rate of flow.
  - e. __Volt__ - electromotive force (e.m.f.) or potential - The electrical pressure under which the current flows.

Batteries

1. Definition - a battery is a group of cells - a cell consists of an electrolyte and a single pair of electrodes, one of which is capable of being oxidized and the other reduced.

Note - Oxidation is defined as a loss of electrons from an atom or ions, and reduction is the process of gaining electrons. Thus an oxidation cannot be separated completely from a reduction. In order for one substance to loose electrons another must gain them (see page 211).

2. __The acid cell__ - this set-up will be demonstrated and explained in theory (see page 223).
3. __The dry cell__ -
  - a. Cathode is a zinc vessel.
  - b. Anode is a carbon rod.
  - c. Electrolyte is ammonium chloride (NH_4Cl)
  - d. Manganese dioxide (MnO_2) is used as a depolarizer (see polarizer page 224).
  - e. The reactions are: Neg. pole Zn ----- $Zn^{++} + 2e^-$
    Pos. pole $NH_4^+ + 2MnO_2 + 2e^-$ ----- $2NH_3 + H_2O + Mn_2O_3$

Note - The dry cell gives a small current for a short period of time otherwise it becomes polarized, that is, the carbon electrode becomes covered by hydrogen.

4. __The Daniell Cell__ (crowfoot cell)
  - a. Cathode - metallic zinc in $ZnSO_4$ solution.
  - b. Anode - sheet copper in $CuSO_4$ solution.

Note - The less dense $ZnSO_4$ solution floats on top of the more dense $CuSO_4$ solution.

  - c. The reactions are: Neg. pole $Zn \longrightarrow Zn^{++} + 2e^-$
    Pos. pole $Cu^{++} + 2e^- \longrightarrow Cu$ or $Zn + Cu^{++} \longrightarrow Zn^{++} + Cu$

Note - The Zn goes into solution and the copper plates out thus the Zn becomes lighter in weight while the copper gains.

5. __Lead Storage Battery__
  - a. Cathode - a lead plate covered with spongy lead.
  - b. Anode - a lead plate covered with lead dioxide (PbO_2).
  - c. Sulfuric acid is used as the electrolyte.
  - d. The reactions are: Neg. pole $Pb + SO_4^= \longrightarrow PbSO_4 + 2e^-$
    Pos. pole $PbO_2 + SO_4^= \longrightarrow PbSO_4 + 2H_2O$

Note - The chief difference in the storage battery and the others is that the storage battery can be recharged, the chemical reactions that produce the electrons are reversed by applying an electric current and the chemicals are returned to their original condition. The overall reactions in the lead storage battery are:

$$Pb + PbO_2 + 2H_2SO_4 \underset{charge}{\overset{discharge}{\rightleftarrows}} 2PbSO_4 + 2H_2O$$

The density of the H_2SO_4 is used to indicate the charge of the battery.

6. The Edison Cell
   - a. Neg. electrode - iron
   - b. Pos. electrode - nickel covered with NiO_2
   - c. Electrolyte is sodium hydroxide
   - d. The reactions are: $NiO_2 + Fe \longrightarrow Fe(OH)_2 + Ni(OH)_2$

Note - The lead storage battery gives a high potential over a short period of time, each cell produces two volts. The Edison cell is more durable and does not suffer from disuse, it produced about 1.3 volts/cell.

. The Electromotive Series

(Review electrochemical series Chapter 6, page 82.)

Note - As indicated in Chapter 6, there are differences in the tendencies of elements to give up valence electrons to become ions. The relative activities of the elements have been determined and the elements arranged in an activity series with the most active metal at the head of the series. It is obvious that elements that give up their valence electrons with the most force will accept them back only when a greater force is applied. Therefore the most active metal will require the greatest electromotive force to cause its ions to accept back the valence electrons to become metals again, and the activity series is also an electromotive-force series.

1. Standard electrodes potentials
   - a. The hydrogen electrode is arbitrarily assigned the value of zero and all others are referred to it as a basis.
2. Applications of the electromotive series
   - a. The electromotive force of any cell may be determined by subtracting algebraically the potentials of its two electrodes.

Ex. $Fe \longrightarrow Fe^{++} + 2e^-$; $E = 0.44$ volts
$2Ag^+ + 2e^- \longrightarrow 2Ag$; $E = -0.80$ volts
$0.44 + (-0.80) = 1.24$ volts ans.

   - b. The fact that iron will replace silver from its salt is also indicated.
   - c. The electromotive series also applied to non-metals the same as it does to metals.

Chapter 13
Acids, Bases and Salts

A. Introduction

1. Three important classes of electrolytes
   - a. <u>Acids</u> - "any compound that gives hydrogen ions (H^+) in solution." An acid is formed when a non-metal oxide is dissolved in water.

$$\text{Ex. } SO_3 + H_2O \longrightarrow H_2SO_4$$

Note - In water solution, the hydrogen ion is combined with a molecule of water, i.e., it is hydrated ($H^+ + H_2O \longrightarrow H_3O^+$) producing a hydronium ion (H_3O^+). For convenience, the hydrogen ion, in solution will be represented as H^+, but it is to be understood that such an ion is actually hydrated.

   - b. <u>Base</u> - "Any compound that gives hydroxyl ions (OH^-) in water solution." A base is formed when a metallic oxide is dissolved in water.

$$\text{Ex. } CaO + H_2O \longrightarrow Ca(OH)_2$$
$$\text{metal oxide } + \text{ water} \longrightarrow \text{a base}$$

Note - In some cases bases may be referred to as <u>alkalis</u> and as <u>hydroxides</u>.

   - c. <u>Salts</u> - An electrically neutral compound composed of the metal (cation) of any base and the non-metal (anion) of any acid.

Note - Acids combine with bases to produce salts and water.

$$\text{Acid} + \text{base} \longrightarrow \text{salt} + \text{water}$$

B. Acids

1. Properties of acids - i.e. the hydrogen ion in solution
   - a. <u>Taste</u> - dilute acids have a sour taste.
   - b. <u>Action on indicators</u> - acids change the color of certain indicators, (blue litmus to red for example).

Note - Indicators are organic acids, called dyes, which, give one color in acid solution and another color in basic solutions due to rearrangement of the valence bonds in the moelcule.

   - c. Action on metals
     - (1) metals above hydrogen in the electromotive series replace hydrogen in acids to form a salt and hydrogen.

$$\text{Ex. } Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

     - (2) Some metals react with oxidizing acids (HNO_3, H_2SO_4, etc.) by oxidation-reduction.

$$\text{Ex. } Cu + HNO_3 \longrightarrow Cu(NO_3) + NO_2 + H_2O \text{ (no } H_2 \text{ produced)}$$

   - d. Action on hydroxides of metals (bases)
     - (1) Acids react with hydroxides of metals to form a salt and water.

$$\text{Ex. } Mg(OH)_2 + H_2SO_4 \longrightarrow MgSO_4 + 2H_2O$$
$$\text{base } + \text{ acid} \longrightarrow \text{salt} + \text{water}$$

Note - The student must observe that the properties of acids and bases as given here refers to a water solution of the ions of these electrolytes and not to the molecular compounds.

2. Naming (nomenclature) of acids (see page 275)
    a. <u>Binary acids</u> - only two elements present. Ex. HCl, H_2S, HCN, etc. Name by using the prefix "hydro" and the name of the other element, changed to end in the suffix "ic". Ex. HCl - hydrochloric acid, H_2S - hydrosulfuric acid.

Note - Hydrocyanic acid actually contains three elements, but by definition a radical (in this case CN) can be considered as an element. (see Chapter 9)

    b. <u>Ternary acids</u> - acids composed of three elements, hydrogen, oxygen, and an acid-forming element (S, N, P, Cl, etc.). Ex. H_2SO_4, HNO_3, H_2CO_3, etc. These acids are named by using the name of the acid-forming element with its ending changed to denote the valence of the non-metal (acid-forming) element.
        (1) suffixes
            "ic" is used to denote the higher valence or the more common acid of the series.
            "ous" is used to denote the acid in which the non-metal has the next lower valence than the "ic" acid.
        (2) prefixes
            "hypo" is used with the suffix "ous" to denote an acid in which the valence of the non-metal is less than it is in the "ous" acid.
            "per" is used with the suffix "ic" to denote an acid in which the valence of the non-metal is greater than it is in the "ic" acid. Examples: HNO_2 - Nitrous acid; H_2SO_3 - sulfurous acid; HNO_3 - nitric acid; H_2SO_4 - sulfuric acid. There are four oxygen acids of the non-metallic (acid-forming) element chlorine: $HClO$ - <u>hypo</u>chlor<u>ous</u> acid; $HClO_3$ - chlor<u>ic</u> acid; $HClO_2$ - chlor<u>ous</u> acid; $HClO_4$ - <u>per</u> chlor<u>ic</u> acid.

Note - The student should be able to tell which of these acids are the more or most common and what the valence of the non-metal is in all cases.

. Bases

1. Properties of bases - i.e. the OH^- ion in solution
    a. <u>Taste</u> - bases have a bitter taste.
    b. <u>Feel</u> - bases have a slick, slippery feeling on the skin.
    c. Bases turn the indicator litmus blue.
2. Naming (nomenclature) of bases
    The metallic element is named followed by the work hydroxide - the OH^- ion is named as an atom, or as a unit. Ex. $NaOH$ - sodium hydroxide; $Ca(OH)_2$ - calcium hydroxide. If the metal has more than one valence its name is changed to end in "ic" for the higher valence and "ous" for the lower valence. Ex. $CuOH$ - cuprous hydroxide (lower valence); $Cu(OH)_2$ - cupric hydroxide (higher valence.)

Note - <u>Acids</u> may be classed as:
monohydrogen (monobasic) acid - HCl
dihydrogen (dibasic) acid - H_2SO_4
trihydrogen (tribasic) acid - H_3PO_4
depending upon the number of replaceable hydrogens.
<u>Bases</u> may be classed as
monohydroxyl (monoacid) base - $NaOH$
dihydroxyl (diacid) base - $Ca(OH)_2$
trihydroxyl (triacid) base - $Al(OH)_3$

. Neutralization

1. Heat of neutralization
    a. The heat of neutralization of all acids are the same showing that the only reaction is between the H^+ and OH^- to produce water.

2. Titration of acid and bases
    a. The process of determining just how much of a standard solution is required to react with a sample being analyzed is called a "<u>titration</u>"
    b. This important branch of analytical chemistry is called "<u>volumetric analysis</u>".
    c. The point at which the desired reaction is complete is indicated by an <u>indicator</u> and is called the "end point".
    d. If the <u>concentrations</u> are expressed in <u>normalities</u> it is true that (ml. of acid) x (N of acid) = (ml. of base) x (N of base).
    For examples see page 237 and Exp. 22. The student should become adept at working problems of this type.

. Salts

1. The following types of reactions produce salts
    a. Metals replacing hydrogen from acids - $Mg + 2HCl \longrightarrow MgCl_2 + H_2$
    b. Certain metals with strong bases - $Zn + 2NaOH \longrightarrow Na_2ZnO_2 + H_2$
    c. All neutralization reactions
    d. Certain salts with strong bases - $CuCl_2 + 2NaOH \longrightarrow Cu(OH)_2 + 2NaCl$
    e. Certain salts with certain acids - $2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$
    f. Some salts with other salts - $BaCl_2 + K_2SO_4 \longrightarrow 2KCl + BaSO_4$

2. Naming (nomenclature) of salts
    a. The name of the metal is written first (If the metal shows two valences the "-ous" and"-ic" ending are used to indicate the lower and higher valences, respectively.), followed by the name of the acid, from which the salt was formed, with the following suffixes
        (1) if a binary acid the end "-ide" is used.
        (2) other acids whose names end with "-ous" and "-ic" give salts whose names and in "-ite" and "-ate" respectively.

3. Varieties of salts are:
    a. <u>Normal salts</u> - "a normal salt is a compound built up wholly of the cations of any base and the anions of any acid." (see page 238)
    b. <u>Acid salts</u> - "an acid salt is a salt composed of both metallic and hydrogen cations, together with the anions of an acid."
    Ex. $NaHCO_3$, $KHSO_4$, etc.
    c. <u>Basic salts</u> - "a basic salt is composed of metallic cations, together with both hydroxyl anions of a base and anions of an acid."
    Ex. $PbOHNO_3$, $BiOHCl_2$, etc.
    d. <u>Double salts</u> - a combination of two salts (Ex. $Cu(AsO_2)_2 \cdot Cu(C_2H_3O_2)_2$, $KCl \cdot MgCl_2$ etc. (see page 630).
    e. Complex salts - Examples - $Ag(NH_3)_2Cl$, $Cu(NH_4)_4Cl_2$, K_2SiF_6, etc. (see page 631)

Chapters 14 and 15

The Halogen Family, The Hydrogen Halides, Oxidation-Reuction Reactions; The Oxygen Compounds of the Halogens

The Halogen Family

A. General Information

1. The elements fluorine, chlorine, bromine, iodine and astatine are members of the halogen family.
2. Their most common valence is -1 (Cl^-, Br^-, I^-, etc.)
3. As the atomic weights increase the physical properties change--Fluorine and chlorine are gases, bromine is a liquid, and iodine a solid.
4. The chemical activity decreases as atomic weights increase--Fluorine is most active, then chlorine, bromine, etc.

B. Preparation

1. Laboratory
    - a. Electrolysis of fused (melted) salts
        - (1) Fluorine - this is the only method for preparing fluorine because it is the strongest chemical oxidizing agent. Potassium hydrogen fluoride is fused, and the electrolysis carried out in a copper cell. The cell reaction is:

$$2KHF_2 \xrightarrow{e} KF + H_2 + F_2$$

        - (2) Chlorine, bromine, and iodine can be prepared by the same method

$$NaCl \xrightarrow{e} Na + Cl_2$$

(similar equations for bromine and iodine)

        - (3) The halogens other than fluorine can be prepared by the electrolysis of aqueous solutions of their salts.

$$2NaCl + 2H_2O \xrightarrow{e} 2NaOH + H_2 + Cl_2$$

(similar equations for bromine and iodine)

    - b. Reaction of the hydro-acids with oxidizing agents
        - (1) The halogens other than fluorine can be prepared by this method

$$2HCl + MnO_2 \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

$$2KMnO_4 + 16HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$$

(similar equations for bromine and iodine)

Note - HI is a strong reducing agent, and HBr is strong enough as a reducing agent to reduce conc. H_2SO_4.

$$2HBr + H_2SO_4 \longrightarrow Br_2 + SO_2 + 2H_2O$$

    - c. Replacement reactions:
        - (1) Bromine and iodine are prepared by replacing them from their compounds with chlorine.

$$2KBr + Cl_2 \longrightarrow 2KCl + Br_2$$

$$2KI + Cl_2 \longrightarrow 2KCl + I_2$$

2. Commercial Preparation
    - a. The commercial preparations of <u>fluorine</u> and <u>chlorine</u> are the same as the laboratory methods.

b. Bromine is recovered from sea water by oxidizing it to free bromine by treating it with chlorine.

$$2Br^- + Cl_2 \longrightarrow 2Cl^- + Br_2$$

The bromine is blown out with air and treated with sulfur dioxide to reduce the bromine to hydrogen bromide.

$$Br_2 + SO_2 + 2H_2O \longrightarrow H_2SO_4 + 2HBr$$

The HBr is dissolved in water to separate it from the air then treated with chlorine to form free bromine again which is distilled out and collected (see page 265).

c. Iodine is prepared by the reduction of sodium iodate ($NaIO_3$) with sodium hydrogen sulfite ($NaHSO_3$):

$$NaIO_3 + NaHSO_3 \longrightarrow 2Na_2SO_4 + 2NaHSO_4 + H_2O + I_2$$

Iodine is also obtained from the iodides in certain sea weeds (kelps)

d. Asatine is not found in nature, but is produced by nuclear reactions.

. Properties of the Free Elements

1. Fluorine (page 262) is a pale-yellow gas, 1.3 times as heavy as air, and forms diatomic molecules (F_2). It is the most active element.
2. Chlorine (page 247) is a greenish-yellow gas (Cl_2) with a characteristic suffocating odor. It is 2.49 times as heavy as air, and is easy to liquify.
3. Chemically Chlorine:
    a. combines with most of the metals directly.
    b. combines with most of the non-metals. $2Sb + 3Cl_2 \longrightarrow 2SbCl_3$
    c. combines with hydrogen. $Cl_2 + H_2 \longrightarrow 2HCl$
    d. combines with the hydrogen in hydrogen compounds.
    $H_2S + Cl_2 \longrightarrow 2HCl + S$ $\quad 2NH_3 + 3Cl_2 \longrightarrow 6HCl + N_2$
    (its reaction with turpentime ($C_{10}H_{16}$) will be demonstrated)
    e. action upon water $Cl_2 + H_2O \longrightarrow HCl + HClO$
    f. Chlorine water is a powerful bleaching agent.
4. Bromine (page 268) is a dark-red liquid whose density is 3.119 which vaporizes readily to a reddish-brown gas. Chemically bromine is like chlorine except it is less active. Bromine water is a good oxidizing agent.
5. Iodine is a purplish-black solid. Its density is 4.93, melts at 113.5 and is purified by sublimation. Chemically it is like chlorine and bromine except less active.

. Comparative Chemical Activity of the Halogens

The comparative activity of the halogens is well indicated by their reaction with water.

1. Fluorine reacts vigorously with water. $2F_2 + 2H_2O \longrightarrow 4HF + O_2$
2. Chlorine reacts readily with water in well defined steps.
    a. $Cl_2 + H_2O \longrightarrow HCl + HClO$

    b. $2HClO \longrightarrow 2HCl + O_2$
    (the decompositive of HClO is hastened by a high temperature and light.)

3. Bromine reacts with water to a very limited extent.

$$Br_2 + H_2O \longrightarrow HBr + HBrO$$

4. Iodine does not react with water.

Note - As a general rule the more chemically active an element is, the more stable its compounds will be. Thus fluorides are the most stable, then chlorides, bromides, and iodides.

. Important Uses of the Halogens

1. <u>Fluorine</u>
    - a. Preparation of fluorine - containing hydrocarbons (freon CCl_2F_2, see page 179)
    - b. Fluoridation of drinking water
    - c. Separation of uranium isotopes
2. <u>Chlorine</u> - one of the most important industrial chemicals.
    - a. bleaching agent - wood pulp, paper, and cotton cloth.
    - b. manufacturing of organic compounds, dyes, drugs, explosives
    - c. war gases - phosgene ($COCl_2$) chloropicrin CCl_3NO_2 etc.
    - d. Purification of drinking water
    - e. Preparation of bleaching powder ($CaOCl_2$)
3. <u>Bromine</u>
    - a. Production of ethylene bromide ($C_2H_4Br_2$) used in gasoline.
    - b. Manufacturing of organic compounds, dyes, drugs, etc.
    - c. Preparation of light sensitive AgBr for photography.
4. <u>Iodine</u>
    - a. Most important use is as an antiseptic "Tincture of Iodine" (alcohol)
    - b. Iodine is used in testing for starch. Starch-potassium iodide paper used to test for fluorine, chlorine, bromine, ozone, and hydrogen peroxide.

The Hydrogen Halides

. General Information

1. The hydrogen halides are the compounds formed when hydrogen combines with a member of the halogen family.

. Methods of Preparation

1. <u>Hydrogen</u> <u>fluoride</u> - by the action of a strong non-volatile acid on a metallic fluoride.

$$CaF_2 + H_2SO_4 \text{ (conc.)} \longrightarrow CaSO_4 + 2HF$$

Note - The formula for hydrogen fluoride is sometimes written as H_2F_2 because the molecules tend to associate. In the liquid state the formula may be H_6F_6. However for our purposes the formula HF will be used.

2. <u>Hydrogen</u> <u>chloride</u> - Two general methods of preparation are available.
    - a. The action of a strong non-volatile acid on a metallic chloride.

$$NaCl + H_2SO_4 \text{ (conc.)} \longrightarrow NaHSO_4 + HCl$$

    - b. By the direct union of the elements $H_2 + Cl_2 \longrightarrow 2HCl$

Note - This is the only hydrogen halide that can be prepared satisfactorily by this method. Fluorine is difficult to prepare and reacts violently with hydrogen. While bromine and iodine react only with difficulty with hydrogen.

3. Hydrogen bromide - Two general methods of preparation are available.
   a. Reaction of a bromide with a strong non-volatile, non-oxidizing acid.

$$2NaBr + H_3PO_4 \longrightarrow Na_2HPO_4 + 2HBr$$

   b. The hydrolysis of a non-metallic bromide such as PBr_3

$$PBr_3 + 3H_2O \longrightarrow H_3PO_3 + HBr$$

4. Hydrogen iodide is prepared by the same two methods as HBr.
   a. $2NaI + H_3PO_4 \longrightarrow Na_2HPO_4 + 2HI$
   b. $PI_3 + 3H_2O \longrightarrow H_3PO_3 + 3HI$

C. Properties of the hydrogen halides
1. Stability decreases rapidly in the order HF, HCl, HBr, and HI.
2. As <u>reducing agents</u> they decrease in the order HI, HBr, HCl, and HF.
3. Dry hydrogen halides are covalent compounds.
4. Reaction with water

$$HF + H_2O \longrightarrow H_3O^+ + F^-$$
$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
$$HBr + H_2O \longrightarrow H_3O^+ + Br^-$$
$$HI + H_2O \longrightarrow H_3O^+ + I^-$$

Note - The aqueous solutions of these compounds are called hydrofluoric, hydrochloric, hydrobromic and hydroiodic acids. All are strong acids except HF which is weak due to association of its molecules. Hydrofluoric acid reacts with glass but can be stored in wax bottles. The reaction with glass is represented by the equation:

$$CaSiO_3 + 6HF \longrightarrow CaF_2 + SiF_4 + 3H_2O$$

D. Uses of the Hydrogen Halides
1. Hydrochloric acid can be used in most cases requiring a strong, cheap acid.
2. HCl is used to catalyze the hydrolysis of starch to glucose.
3. HCl, HBr, and HI are all reducing agents (see methods of preparing chlorine etc.)
4. Hydrobromic and hydroiodic acids have little use as acid.
5. Hydrofluoric acid is used in etching glass.

Oxidation-Reduction Reactions

. The Nature of Oxidation-Reduction (see pages 211, 246 and 254)
1. The oxidation states (or valences) of some of the elements involved in the reaction are changed.

   For example the reaction between a solution of bromine in water and metallic copper as represented by the equation: $Cu + Br_2 \longrightarrow CuBr_2$, or ionically the equation is:

$$Cu^0 + Br_2^0 \longrightarrow Cu^{++} + 2Br^-$$

   Now suppose that we examine each of the reactants to see what has happened to it.

Note - The student should recall that:

- a. The valence of an unreacted atom is zero.
- b. A typical metal gets its positive valence by losing electrons,
- c. A non-metal gets its negative valence by taking on electrons.
- d. The algebraic sum of positive and negative valences of all the atoms in a compound must equal zero.
- e. Hydrogen usually has a valence of +1 while oxygen usually has a valence of -2.

The copper reacts by losing electrons thus: $Cu^{o} \longrightarrow Cu^{++} + 2e$ and the bromine reacts by gaining two electrons thus:

$$Br_2 + 2e^- \longrightarrow 2Br^-$$

Note - Loss of electrons is oxidation; gain of electrons is reduction, therefore the copper is oxidized by the bromine and the bromine is reduced by the copper. The copper is the reducing agent and the bromine is the oxidizing agent.

2. Oxidation-Reduction Reactions Include:
   - a. Combination reactions $Cu + Br_2 \longrightarrow CuBr_2$
   - b. Decomposition reactions $2HI \longrightarrow H_2 + I_2$
   - c. Displacement reactions $Zn + CuSO_4 \longrightarrow Cu + ZnSO_4$
   - d. Complex reactions such as:

$$2KMnO_4 + 16\ HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$$

3. Balancing Oxidation-Reduction Reactions (see pages 246, 257 and 258)

Note - Simple oxidation-reduction reactions (the first three examples above) can be balanced easily by the inspection method. The more complex types (the fourth example above) are often difficult to balance by this method. However there are several methods of balancing these complex equations all of which are based upon the fact that "the number of electrons yielded by the reducing agent must, in the balanced equation, be equal to the number of electrons accepted by the oxidizing agent." The "valence-electron method" has many advantages and will be outlined here.

1. The valence-electron method

Suppose that it is desired to balance the equation:

$$CuO + NH_3 \longrightarrow Cu + N_2 + H_2O$$

a. Notice the change in valence that certain ones of the atoms undergo. This might be indicated thus:

$$\overset{+2}{CuO} + \overset{-3}{NH_3} \longrightarrow \overset{0}{Cu} + \overset{0}{N_2} + H_2O$$

decrease of 2

increase of 3

or $\overset{+2}{Cu} + 2e \longrightarrow Cu^{o}$ $\qquad$ $\overset{-3}{N} \longrightarrow 3e + N$

b. Then multiply these equations by the smallest numbers that will give equal numbers of electrons in the two equations.

$$\overset{+2}{3Cu} + 6e \longrightarrow \overset{0}{3Cu} \qquad \overset{-3}{2N} - 6e \longrightarrow \overset{0}{N_2}$$

c. Add the equations together (note that the electrons cancel out)

$$\overset{+2}{3Cu} + \cancel{6e} \longrightarrow \overset{0}{3Cu}$$

$$\overset{-3}{2N} - \cancel{6e} \longrightarrow \overset{0}{N_2}$$

$$\overset{+2}{3Cu} + \overset{-3}{2N} \longrightarrow \overset{0}{3Cu} + \overset{0}{N_2}$$

d. Now that the number of atoms of copper and nitrogen are fixed substitute into the original equation and complete the balance by inspection.

$$3CuO + 2NH_3 \longrightarrow 3Cu + N_2 + 3H_2O$$

Notice that there are 3 oxygen atoms and 6 hydrogen atoms that combine to form water ($3H_2O$)

Note - Notice that when the oxidizing agent is multiplied by the valence change of the reducing agent and the reducing agent is multiplied by the valence change of the oxidizing agent the equation is balanced. This will be true in all cases except where not all of one of the reactants react by oxidation-reduction.

For example in the reaction illustrated by the equation:

$$\overset{0}{Cu} + \overset{+5}{HNO_3} \longrightarrow \overset{+2}{Cu(NO_3)_2} + \overset{+2}{NO} + H_2O$$

increase of 2

decrease of 3

Since not all of the nitric acid reacted by oxidation-reduction the valence change of the copper (2) indicates the amount of NO formed but not the amount of HNO_3 used, all of the Cu was oxidized so the valence change (3) of the nitrogen shows the amount of Cu used. Now the equation:

$$3Cu + HNO_3 \longrightarrow CuNO_3 + 2NO + H_2O$$

can be written and the balance completed by inspection.

Oxygen Compounds of the Halogens

A. General Information

1. In addition to a valence of -1 the halogens, except fluorine, show valences of +1, +3, +5 and +7.
2. The corresponding compounds of chlorine are illustrated by the acids $HClO$, $HClO_2$, $HClO_3$, and $HClO_4$.
3. The oxygen acids of the halogens have these common properties
   - a. Except for $HClO$, they are strong acids
   - b. They are unstable toward heat
   - c. They are strong oxidizing agents

Note - Since the oxy-acids of the halogens are unstable toward heat they have little use as acids but both the acids and their salts which are more stable, are used as oxidizing agents.

B. Important Oxy-acids of the Halogens

1. Hypochlorous acid ($HClO$)
    - a. Preparation

$$Cl_2 + H_2O \rightleftharpoons HCl + HClO$$

    - b. Properties: oxidizing and bleaching agent.

$$HClO \longrightarrow HCl + (O)$$

    - c. Uses
        - (1) preparation of bleaching powder ($CaOCl_2$)

$$Ca(OH)_2 + Cl_2 \longrightarrow CaOCl_2 + H_2O$$

        - (2) preparation of hypochlorite salts which are good oxidizing agents and bleaching agents.
2. Chloric Acid: ($HClO_3$)
    - a. Preparation

$$Ba(ClO_3)_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HClO_3$$

    - b. Uses - The chlorates (salts of chloric acid) are strong oxidizing agents, (recall preparation of oxygen) and are used in "weed killers".
3. Perchloric Acid ($HClO_4$)
    - a. Preparation:

$$4KClO_3 \longrightarrow KCl + 3KClO_4$$

$$KClO_4 + H_2SO_4 \longrightarrow KHSO_4 + HClO_4$$

    - b. Properties - the perchlorates are the most stable salts of the oxy-acids of chlorine. They are good oxidizing agents and $KClO_4$ is nearly insoluble in water.
    - c. Uses - The perchalorates are used as oxidizing agents and in quantative analysis of potassium.
4. Iodic acid (HIO_3)
    - a. Preparation

$$NaIO_3 + H_2SO_4 \longrightarrow NaHSO_4 + HIO_3$$

sodium iodate ($NaIO_3$) is found in Chile saltpeter and is the principle source of iodine.

Chapter 16
The Alkali Family of Metals

A. Introduction and general properties

1. Metals are those elements which react with oxygen to form basic anhydrides. Non-metals react with oxygen to form acid anhydrides.
2. Metals are malleable, and have a metallic luster while non-metals are brittle and possess no luster.
3. The alkali metals show the following properties:
    a. The density increases as the atomic number increases (Li 0.53, Cs 1.90)
    b. The melting point decreases as the atomic weight increases (Cs 26, Li 186).
    c. Atomic and ionic radii increase as the atomic number increases.
    d. The electrode potential decreases as the at. wt. increases. (Cs gives up its electron the most readily)
    e. Li, Na, and K are lighter than water.
    f. The alkali metals are soft and easy to cut; the atoms are large.
    g. The alkali metals are very active and are not found free in nature.
    h. The order of chemical activity is Cs, Rb, Li, K, Na.
    i. The compounds of the alkali metals are very stable and are soluble in water.
        (1) oxides cannot be decomposed by heat.
        (2) carbonates decompose only at high temperatures.
        (3) hydroxides melt without decomposing.
4. Chemical properties are:
    a. <u>Valence</u> - electrovalence of +1 formed by loss of the one valence electron.
    b. The replacement of hydrogen in water to form bases.

$$\text{Ex. } 2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

    c. The alkali metals form the strongest bases.
    d. All alkali metals corrode readily in air to form oxides.
    e. The free metals are strong reducing agents (they give up electrons readily).

B. Occurrence

1. Sea water - 2.8% Na; 0.08% K.
2. Igneous rocks - 2.85% Na; 2.60% K.
3. Potassium is present in plants.
4. Salt beds and salt seas.
5. Oilwell brines.
6. Order of abundance - Na, K, Li, Rb, Cs.

C. Preparation of Free Metals

1. Recall methods of preparation of halogens from melted salts.

$$2NaCl \xrightarrow{elect.} 2Na + Cl_2$$

2. Electrolysis of fused hydroxides

$$2NaOH \xrightarrow{elect.} Na + O_2 + H_2$$

Note - In these electrolytic processes (Castner Process for fused hydroxides and the Downs Cell for fused chlorides) care must be taken to prevent the products from mixing and reacting.

   3. Cs can be prepared by reacting its salts with calcium and separating the Cs by distillation.
   4. Potassium can be prepared by a similar procedure

D. Uses of the Alkali Metals (not very useful as metals)
   1. Sodium is the most useful, some of its uses are:
      a. Sodium vapor lamps.
      b. Core for exhaust valves in aircraft engines.
      c. Used in the manufacturing of many chemicals, soap, dyes, organic chemicals, etc.
   2. Cesium and rubidium are used in photoelectric cells.

E. Compounds of the Alkali Metals
   1. Lithium compounds
      a. <u>Lithium chloride</u> - used in preparing other compounds and in air conditioning.
      b. <u>LiBr</u> and <u>Li_2CO_3</u> - used in medicine and in ceramics for glasses and glazes.
   2. Sodium compounds
      a. <u>Sodium chloride</u> - (table salt)(NaCl) (See chapter 13, page 231) essential to the diet. Widely distributed in nature, used to prepare metallic sodium and chlorine and most compounds of the two elements, -- HCl, Na_2SO_4, etc.
      b. <u>Sodium hydroxide</u> (NaOH)
         (1) Prepared by two processes:
            (a) Treating calcium hydroxide with sodium carbonate

$$Ca(OH)_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaOH$$

recover NaOH by filtration and evaporazation.
            (b) Electrolysis of sodium chloride solution: (see prep. of chlorine, Chapter 14, page 244)
         (2) Uses of sodium hydroxide: (see table, page 292)
            (a) Manufacturing of hard soaps.
            (b) Refining of petroleum.
            (c) Preparation of wood pulp for paper making.
            (d) Rayon, chemicals, textiles, etc.
      c. <u>Sodium sulfate</u> (Na_2SO_4) - a byproduct of the manufacture of acids other than H_2SO_4.
         (1) Uses are
            (a) Manufacturing of rayon, glass, and paper.
            (b) Medicine - "Glauber's salt". ($Na_2SO_4 \cdot 10\ H_2O$)
      d. <u>Sodium peroxide</u> (Na_2O_2) used in preparation of oxygen (see page 43).
      e. <u>Sodium thiosulfate</u> ($Na_2S_2O_3$)
         (1) Preparation: By adding sulfur to a solution of sodium sulfite

$$Na_2SO_3 + S \longrightarrow Na_2S_2O_3$$

         (2) Used in photography - "Hypo" and to remove excess chlorine "antichlor" from any substance bleached with chlorine.
      f. <u>Sodium carbonate</u> (soda ash)(Na_2CO_3)
         (1) Preparation - Solvay Process
            (a) Raw materials: NaCl, NH_3, CO_2, $CaCO_3$, H_2O
            (b) Products are - Na_2CO_3, $NaHCO_3$, $CaCl_2$ and NaOH.

$$NH_3 + H_2O + CO_2 \longrightarrow NH_4HCO_3$$

$$NH_3HCO_3 + NaCl \longrightarrow NaHCO_3 + NH_4Cl$$

$$2NaHCO_3 + \text{heat} \longrightarrow Na_2CO_3 + H_2O + CO_2$$

$$2NH_4Cl + CaO \longrightarrow CaCl_2 + 2NH_3 + H_2O$$

$$CaCO_3 \longrightarrow CaO + CO_2$$

$$Na_2CO_3 + Ca(OH) \longrightarrow CaCO_3 + 2NaOH$$

Overall the reaction is:

$$CaCO_3 + 2NaCl \longrightarrow CaCl_2 + Na_2CO_3,$$

which is an impossible reaction.

(2) Uses:
- (a) $Na_2CO_3 \cdot 10H_2O$ - washing soda
- (b) glass, soap, chemicals, etc. (see table, page 295)

g. Sodium Hydrogen Carbonate (bicarbonate of soda)(baking soda)($NaHCO_3$)

(1) Preparation: Solvay Process above, and by treating Na_2CO_3 solution with carbon dioxide:

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

(2) Uses: baking and medicine.

h. Sodium Nitrate ($NaNO_3$)(Chile saltpeter)

(1) Uses - fertilizer

i. Sodium Cyanide (NaCN)

(1) Preparation

$$NaNH_2 + C \longrightarrow NaCN + H_2$$

(2) Uses - dissolves gold, good poison.

j. Sodium Phosphates (will be studied with phosphorus (Chapter 19)

3. Potassium Compounds:

a. Potassium Hydroxide (Caustic Potash) (KOH) Same properties as NaOH. NaOH is cheaper and is usually used.

b. The potassium halides (KCl, KBr, KI)
- (1) Very similar in properties to the corresponding sodium compounds.
- (2) The bromide and iodide are used in medicine and in the preparation of salts used in photography.

c. Potassium Sulfate (K_2SO_4)

(1) Preparation:

$$KCl + H_2SO_4 \longrightarrow K_2SO_4 + 2HCl$$

(2) Used in fertilizer and in preparing baking powder.

d. Potassium Carbonate (potash)(K_2CO_3)

(1) Preparation:

$$K_2S + CaCO_3 \longrightarrow K_2CO_3 + CaS$$

and from wood ashes.

(2) Uses - manufacturing of hard glass, and soft soap.

e. Potassium Nitrate (saltpeter)(KNO_3)

(1) Preparation:

$$NaNO_3 + KCl \longrightarrow KNO_3 + NaCl$$

(2) Uses:
- (a) Preparation of O_2: $2KNO_3 \longrightarrow 2KNO_2 + O_2$
- (b) Good oxidizing agent
- (c) Black gun powder - KNO_3, S, C.
- (d) Fertilizer, fireworks, medicine.

. Compounds of Rubidium and Cesium are used in radio tubes and photoelectric cells.

The Spectroscope

A. The Spectrum

1. Every element produces light of a characteristic wave length, no two elements give the same wave length.
2. A <u>spectrum</u> is any series of radiant (light) energies arranged in order of wave lengths.
3. Elements can be identified by their spectrum.
   Ex. Li - red, Na - yellow, K - violet, Cs - blue, Rb - blue-red, etc.

B. Kinds of Spectrum

1. The Emission Spectra:
   The spectrum formed by the light given off when any substance is heated.
2. The Absorption Spectra:
   When a white light (which gives a continuous spectrum) is passed through a gas, the gas absorbs the wave lengths which it would emit if luminous. The interruptions in the continuous spectrum will indicate the gas present. Helium was discovered in the sun's atmosphere before it was on earth by this procedure.

APPENDIX III

A Test on the Scientific Method

CHEMISTRY 124

Final Examination Spring 1954

Form B

Name____________________________ Lab. Inst. ___________________________

This examination consists of four parts. Directions are given at the beginning of each part. Read them carefully and proceed at once to answer the question. Score = R - W/4. Do not mark on the test sheets. There is only one correct answer for each question.

Part One

Items 1 through 18: Classify each of these statements according to five numbered catagories below, and indicate the number of your classification by blackening the corresponding space on your answer sheet.

(1) An exact scientific law.
(2) An approximate scientific law.
(3) A theoretical explanation advanced to explain or account for observed experimental results.
(4) A correct definition.
(5) A false statement.

1. A solid crystal of NaCl contains Na^+ and Cl^- ions, which separate when the crystal dissolves in water.

2. For a given metal, the product of specific heat and atomic weight is a constant.

3. At constant temperature, the volume of a given weight of any real gas will vary inversely with the pressure.

4. Density is equal to mass divided by volume.

5. Combination of non-metals with each other is made possible by the sharing of electrons.

6. Molecules of all real gases have some attraction for each other as all known gases have been liquefied.

7. In a given isolated system the total sum of mass plus energy, when expressed in consistent units, will remain unchanged during physical and chemical changes.

8. An atom of uncombined bromine has 7 electrons in its outer energy level or shell and forms NaBr with an atom of sodium by the transfer of 1 electron from the sodium atom.

9. A good reducing agent is one which has a great tendency to acquire electrons.

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10. Zero on the centigrade scale represents the temperature at which ice and water are in equilibrium at 760 mm mercury pressure of air.

11. The properties of the elements are periodic functions of their atomic numbers.

12. When sodium hydroxide solution reacts with nitric acid solution the OH^- ions from the base unite with the H^+ (or H_3O^+) ions from the acid to form H_2O.

13. At a pressure of 760 mm of mercury, 100° C is the boiling point of pure water.

14. A normal solution contains one gram-molecular weight in one liter of water.

15. At constant temperature, the solubility, by weight, of a gas in a liquid is directly proportional to the pressure of the gas.

16. A free atom of an element having a valence of -1 has one electron in the outer level or orbit.

17. For a given electrolyte, the amount of material liberated is directly proportional to the amount of electricity passing through the cell.

18. An aqueous solution of 0.1 normal acetic acid does not contain as many H^+ (or H_3O^+) ions per liter as an aqueous solution of 0.1 normal hydrochloric acid.

Part 2

<u>Items 19 through 27 refer to the following experiment</u>: In studying methods for making pine wood fire-resistant, white pine sawdust was treated with the substances listed below. It was then ignited, and the time of duration of flame and per cent loss in weight during the combustion period were determined.

Substance Added	Amount Added	Flame Duration	Weight Loss
Nothing added	0	5.66 min	78%
NH_4Cl	5%	4.66 min	70%
	10%	3.92 min	43%
	15%	1.33 min	19%
$(NH_4)_2HPO_4$	5%	2.75 min	43%
	10%	2.28 min	25%
	15%	1.37 min	20%
$ZnCl_2$	5%	3.00 min	50%
	10%	2.00 min	25%
	15%	1.50 min	20%

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Using these data, mark the degree of correctness of items 19 through 27 as follows:

(1) The statement is <u>true</u>.
(2) The statement is <u>probably true</u>; additional data would be necessary for a final decision.
(3) The statement is <u>impossible to judge</u>; the experiment provides no evidence upon which to make a prediction of the results to be expected in this case.
(4) The statement is probably false; additional data would be necessary for a final decision.
(5) The statement is <u>false</u>.

19. NH_4Cl is the cheapest salt to use in making white pine fire-resistant.

20. When the amount added is 15%, NH_4Cl is more effective fireproofing than $ZnCl_2$ as far as weight loss is concerned.

21. Treatment with $NH_4H_2PO_4$, in place of $(NH_4)_2HPO_4$, would make white pine sawdust burn less readily than the untreated sawdust.

22. For $ZnCl_2$, the amount of weight loss is directly proportional to the time of flame duration.

23. By adding enough of any of the three chemicals to white pine sawdust, the time of flame duration could be made less than 1.3 minutes.

24. These three chemicals would prevent the rotting of white pine.

25. These chemicals would accelerate the rate of burning of sawdust from yellow pine.

26. Under the same conditions, 12% $(NH_4)_2HPO_4$ would result in a 23% weight loss during combustion.

27. Na_2CO_3 is the best material for fireproofing wood.

Part 3

<u>Items 28 through 41 are grouped in pairs</u>. Preceding each pair of items is a statement. The items require you to select the best <u>experimental evidence</u> and the best <u>theoretical explanation</u> for each statement.

STATEMENT: Sodium chloride in the pure form contains sodium and chloride ions.

28. What is the best experimental evidence for this statement? (1) Sodium chloride is very soluble in water.

29. What is the best theoretical explanation of this statement?

(2) Sodium has a great tendency to lose electrons and chlorine has a great tendency to gain electrons.
(3) Fused sodium chloride is a good conductor of electricity.
(4) A molar solution of sodium chloride gives twice the freezing point depression of a molar solution of sugar.
(5) In electrolysis of an aqueous solution of sodium chloride, sodium ions travel toward the negative electrode and chloride ions travel toward the positive electrode.

STATEMENT: Under certain conditions H_2O_2 can act as an oxidizing agent; under other conditions, as a reducing agent.

30. What is the best experimental evidence for this statement?

31. What is the best theoretical explanation of this statement?

(1) H_2O_2 will decolorize $KMnO_4$ solutions in the presence of an acid and will turn black lead sulfide to a white compound.
(2) An atom within a compound can sometimes attain a more stable electronic structure either by gaining or by losing electrons.
(3) The oxygen atoms in H_2O_2 share electrons; therefore, H_2O_2 is a covalent compound.
(4) H_2O_2 is a good bleaching agent.
(5) Peroxides are stronger oxidizing agents than are oxides.

STATEMENT: Copper is more reactive chemically than is gold.

32. What is the best experimental evidence for this statement?

33. What is the best theoretical explanation of this statement?

(1) Copper loses electrons more easily than does gold.
(2) Gold oxide requires less heat per gram mole for decomposition than does copper oxide.
(3) The copper ion has less tendency to lose electrons than has the gold ion.
(4) Gold is alloyed with copper.
(5) The copper atom may lose as many as two electrons, whereas the gold atom may lose as many as three electrons.

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STATEMENT: Many of the chemical properties of the elements vary periodically with the atomic numbers.

34. What is the best experimental evidence for this statement?

35. What is the best theoretical explanation of this statement?

(1) Number of electrons in the outermost shell of the atom vary in a regular manner.
(2) Structure of the nucleus.
(3) Similarity in chemical properties of the elements in a given family.
(4) Fact that many elements have isotopes
(5) Number of protons in the nucleus.

STATEMENT: The formula of hydrogen sulfide gas is H_2S.

36. What is the best experimental evidence for this statement?

37. What is the best theoretical explanation of this statement?

(1) When hydrogen sulfide is burned in air or oxygen, H_2O and SO_2 are formed.
(2) Free hydrogen gas consists of H_2 molecules, but sulfur is solid.
(3) At standard conditions the weight of 22.4 liters of hydrogen sulfide gas is 34 grams, of which 32 grams is sulfur.
(4) Sulfur and hydrogen atoms combine in a ratio of two to one.
(5) The sulfur atom has six electrons in its outer level and obtains the two electrons needed for a stable arrangement by combining with two hydrogen atoms.

STATEMENT: Molecules of higher-than-average kinetic energy are involved in chemical reactions.

38. What is the best experimental evidence for this statement?

39. What is the best theoretical explanation of this statement?

(1) A rise in temperature causes an increase in the vapor pressure of a liquid.
(2) A small rise in temperature causes a rather large increase in the velocity of most chemical reactions.
(3) Collision between molecules will not result in reaction unless enough energy is available for "breaking up" the colliding particles.

(4) At the same temperature light molecules have a higher average velocity than heavy molecules.
(5) Some reactions take place only by the absorption of energy.

STATEMENT: In magnesium chloride the valence of magnesium is 2.

40. What is the best experimental evidence for this statement?

41. What is the best theoretical explanation of this statement?

(1) Magnesium combines with chlorine, which has the formula Cl_2.
(2) Two faradays of electricity are required to liberate one gram atomic weight of magnesium from molten magnesium chloride.
(3) Magnesium vapor contains 2 atoms per molecule.
(4) The nucleus of the magnesium atom contains two protons.
(5) The magnesium atom has 2 electrons in the outermost shell.

Part 4

<u>Items 42 to 50</u>: Mark the one best answer for each of the following questions.

42. "The valence of hydrogen is one." Which of the following best characterizes this statement?
    (1) The statement is true by observation.
    (2) The statement is true because of the use of hydrogen as a standard for measuring valence.
    (3) The statement is a consequence of the composition of water as shown by the formula H_2O.
    (4) The statement follows from the electron theory of valence as the H atom has the same valence as the alkali metals.
    (5) The statement is true because the atomic number decreases by one when a hydrogen nucleus (proton) is knocked out of an atom.

43. When sodium hydroxide solution is added to magnesium sulfate solution, a white precipitate of magnesium hydroxide is obtained. When sodium hydroxide solution is added to an "unknown" solution, a white precipitate is obtained. In order to conclude that the unknown solution contains magnesium ion it must be assumed that:
    (1) NaOH is more soluble than $Mg(OH)_2$.
    (2) Na_2SO_4 is soluble in water.
    (3) $Mg(OH)_2$ is insoluble in water.

(4) NaOH forms no white precipitate with any other ion except Mg^{++}.
(5) Zn^{++}, which forms white $Zn(OH)_2$, is not present in the unknown.

44. "The reaction of zinc and an aqueous solution of hydrochloric acid involves a transfer of electrons." Which of the following is the best experimental evidence for this statement?
(1) Zinc is a reducing agent; chlorine is an oxidizing agent.
(2) The reduction of zinc with hydrochloric acid generates considerable heat.
(3) The freezing point of water is lowered by dissolving zinc chloride in it.
(4) The electrons of zinc are free as shown by the electrical conductivity of zinc.
(5) A galvanic cell with electrodes of zinc and carbon, containing the reactants in separate beakers connected by a salt bridge can be made to produce an electric current.

45. How would you determine experimentally whether the compound XCl_2 is electrovalent or covalent?
(1) Test its solubility in water.
(2) Determine its percentage composition.
(3) Test its electrical conductivity in the fused state.
(4) Determine whether the compound has an electrical charge.
(5) See whether it will react with "electron-sharing" or with "electron-transfer" compounds.

46. Which of the following is explained by kinetic molecular theory as applied to gases?
(1) The high compressibility of gases.
(2) The solubility of gases in water.
(3) The color of gases.
(4) The combustibility of gases.
(5) The chemical activity of gases.

47. An aqueous solution of NaCl contains positive sodium ions and negative chloride ions. This explains most directly the fact that:
(1) The solution is neutral to litmus.
(2) The sodium metal will react with HCl solution.
(3) The solution conducts electricity.
(4) Chlorine is used for bleaching
(5) The solution has a lower vapor pressure than pure water.

48. The inert gases in Group 0 of the Periodic Table have the most stable electronic arrangements. This explains most directly why these elements
(1) Are inactive chemically
(2) Are, with one exception, not radioactive.
(3) Are gases.
(4) were discovered so comparatively recently.
(5) may be identified by use of a spectroscope.

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49. Calcium metal has only one more electron than potassium metal in its outer ring or energy level. This explains why
    (1) the atomic weight of Ca is 40, that of K is 39.
    (2) the valence of Ca is one less than that of K.
    (3) the valence of K is one less than that of Ca.
    (4) Ca^{++} causes hardness in water, K^{+} does not.
    (5) $CaCO_3$ is much less soluble than K_2CO_3.

50. Crowding more molecules in a given space will increase the number of chances for collision. This explains most directly
    (1) the law of mass action.
    (2) why mercury is more dense than aluminum.
    (3) why water causes a mixture of solid $BaCl_2$ to react with solid $CuSO_4$.
    (4) why reactions between gases often occur more rapidly than reactions between liquids.
    (5) why liquids tend to evaporate.

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APPENDIX IV

First Departmental Intra-semester Examination

CHEMISTRY 124
First Hour Examination
Spring 1954

Name ______________________

DIRECTIONS: Mark the one best answer on the test sheets and on the answer sheets. Score = R - W/4

1. The elements of group five, (N,P,As etc.) have a lowest oxidation state of:
   (1) +5 (2) +3 (3) zero (4) -3 (5) -5

2. The elements of group five have five valence electrons and hence their maximum oxidation state is:
   (1) -3 (2) zero (3) +1 (4) +3 (5) +5

3. Ammonia can be prepared by heating ammonium chloride (NH_4Cl) with:
   (1) H_2SO_4
   (2) hydrochloric acid
   (3) calcium hydroxide
   (4) gelatin
   (5) nitric acid

4. Which of the following would not be found in a bottle of ammonia water?
   (1) water (H_2O)
   (2) molecular ammonia (NH_3)
   (3) ammonium ion (NH_4^+)
   (4) Hydronium ion (H_3O^+)
   (5) hydroxyl ion (OH^-)

5. The valence of nitrogen in ammonia is:
   (1) -3 (2) -1 (3) zero (4) +3 (5) +5

6. To calculate the density (grams per liter) of gaseous ammonia at S.T.P. you would:
   (1) divide its molecular weight by its valence.
   (2) divide its molecular weight by 22.4.
   (3) multiply its molecular weight by its valence.
   (4) add the atomic weights
   (5) dissolve it in water

7. The elements of the nitrogen family (N, P, As etc.)
   (1) are all metals
   (2) are all nonmetals
   (3) all form volatile hydrides
   (4) all have three valence electrons
   (5) all form oxides that are acid anhydrides

8. Most of the industrial ammonia is prepared in this country by:
   (1) the destructive distillation of soft coal
   (2) the Haber Process
   (3) thunderstorms
   (4) bacteria
   (5) heating NH_4Cl with $Ca(OH)_2$

9. With increase in the atomic weight the members of the nitrogen family:
   (1) become more metallic
   (2) form stronger acids
   (3) have more valence electrons
   (4) become more nonmetallic
   (5) have fewer electrons in the outer levels

(COPY)

10. The "fixation" of nitrogen is accomplished in nature by:
(1) photosynthesis in green plants
(2) nitrogen dissolving in rain-water
(3) bacteria in the roots of legume plants
(4) the burning of nitrogen in gas wells
(5) none of these

11. The products formed in the reaction $PCl_3 - H_2O \longrightarrow$ ________ + _____ are:
(1) $PO + HClO_2$ (4) $HCl + H_3PO_3$
(2) $PH_3 + HCl_3O$ (3) $HClO + PClO$ (5) $HCl + H_3PO_4$

12. Rock phosphate has the formula:
(1) NaH_2PO_4 (4) H_3PO_4
(2) $Ca_3(PO_4)_2$ (3) $Ca(H_2PO_4)$ (5) PCl_5

13. The formula for disodium phosphate is:
(1) Na_3PO_4 (4) NH_4Cl
(2) Na_2HPO_4 (3) NaH_2PO_4 (5) $NaNO_3$

14. The formula for phosphine is:
(1) HPO_3 (4) PH_3
(2) $H_4P_2O_7$ (3) PCl_3 (5) H_3PO_4

15. When a piece of copper is heated in the air:
(1) it is reduced (4) metallic copper is deposited
(2) nitric acid is formed
(3) copper oxide (CuO) is formed (5) no reaction takes place

16. How many grams of sodium nitrate ($NaNO_3$, molecular weight 85) would be required to produce 20 grams of nitric acid (Mol. Wt. 63). The reaction is represented by the equation: $NaNO_3 + H_2SO_4 \longrightarrow NaHSO_4 + HNO_3$.
(1) 6.5 grams (4) 27 grams
(2) 13.5 grams (3) 18 grams (5) 36 grams

17. Mark the same answer as in preceding question.

18. The molecular weight of ammonia (NH_3) is 17. How many grams of ammonia would be present in 1 liter of 0.25 N ammonium hydroxide.
(1) 2.12 grams (4) 17 grams
(2) 4.25 grams (3) 8.5 grams (5) none of these

19. Mark the same answer as in the preceding question.

20. A six normal (6N) sulfuric acid solution (H_2SO_4) is:
(1) 6 molar (4) 2 molar
(2) 12 molar (3) 3 molar (5) none of these

21. A solution that contains one gram equivalent weight of solute per liter of solution is:
(1) one molar (2) one normal (3) saturated (4) concentrated
(5) none of these

(COPY)

22. Heating antimony sulfide (Sb_2S_3) in the air causes the formation of:
(1) $Sb + SO_2$ (4) $SbO + S$
(2) $Sb_2O_3 + SO_2$ (3) $SbNO_3 + H_2O$ (5) no new substances

23. Antimony is used in making type metal because:
(1) it forms a soft alloy with lead (4) it is non-poisonous
(2) it forms alloys that expand on cooling (5) it has a low melting point
(3) it forms a hard alloy with lead

24. One use of bismuth is:
(1) making lead shot (4) for fertilizer
(2) making insect sprays (5) there are no commercial uses for bismuth
(3) making low melting point alloys

25. One of the chief uses of arsenic or its compounds is:
(1) fertilizer (4) as catalysts
(2) insect poisons (5) none of these
(3) making bearing metal

26. If 11.2 liters (STP) of ammonia gas was dissolved in enough water to make one liter of solution, the normality of the solution would be:
(1) 0.3 N (2) 0.4 N (3) 0.5 N (4) 0.6 N (5) 0.7 N

27. The pale yellow color which appears in nitric acid, as prepared in the laboratory, is due to:
(1) nitrogen oxide (4) nitrogen pentioxide
(2) nitric oxide (3) nitrogen dioxide (5) nitrous oxide

28. The valence state of nitrogen in nitric acid (HNO_3) is:
(1) -3 (2) -3 (3) zero (4) -5 (5) none of these

29. A general method for preparing volatile acids is:
(1) to treat their salts with sulfuric acid
(2) to treat chlorides with silver nitrate
(3) to dissolve ammonia in water
(4) to dissolve metallic oxides in water
(5) to heat HNO_3 and their salts together

30. Sulfur burns in the air to form
(1) sulfuric acid (4) sulfur dioxide
(2) sulfurous acid (3) sulfur trioxide (5) hydrogen sulfide

31. Sulfur changes from the rhombic form to the monoclinic form if the temperature is raised to 95.5°. This temperature is known as the:
(1) melting point of sulfur (4) boiling point of sulfur
(2) transition temperature of sulfur (5) critical temperature of sulfur
(3) Frasch process

(COPY)

32. The Frasch process for mining sulfur depends upon the fact that:
(1) sulfur is a nonmetal
(2) sulfur is lighter than water
(3) sulfur is soluble in water
(4) sulfur has a low melting point
(5) sulfur floats on water

33. Hydrogen sulfide will not
(1) precipitate metal ion (i.e. sulfides)
(2) oxidize SO_2
(3) tarnish silver
(4) produce an acid solution
(5) form salts such as NaHS

34. Dried Fruits may be bleached and preserved at the same time by treating them with:
(1) H_2S (2) SO_2 (3) SO_3 (4) H_2SO_3 (5) H_2SO_4

35. The most difficult but vital step in any process for preparing sulfuric acid is:
(1) getting sulfur to combine with oxygen
(2) overcoming the low solubility of SO_3 in water
(3) combining SO_2 with oxygen
(4) concentrating the acid
(5) keeping the temperature constant

36. One advantage of the contact process for making sulfuric acid is:
(1) the catalyst is easily poisoned
(2) the raw materials need not be pure
(3) it gives a pure concentrated acid
(4) the up-keep of the plant is small
(5) the acid is impure and dilute but well suited for the fertilizer industry

37. Concentrated sulfuric acid is used to make other acids because it:
(1) is an oxidizing agent
(2) has a high boiling point
(3) is a dehydrating agent
(4) is a strong acid
(5) is dibasic

38. In the "lead chamber" process for making sulfuric acid:
(1) the acid produced is of high purity (100%)
(2) the catalyst is an oxide of nitrogen
(3) the process requires no catalyst
(4) the product is unfit for conversion into fertilizer-grade ammonium sulfate
(5) sulfur dioxide is reduced

39. In testing for sulfuric acid and sulfates which of the following would you use to <u>prove</u> the presence of the sulfate ion?
(1) lime water
(2) sodium chloride
(3) barium chloride
(4) litmus
(5) something else

40. In the neighborhood of room temperature the rate of a chemical reaction is approximately doubled if the temperature is:
(1) doubled
(2) increased one degree
(3) raised to 100° C.
(4) raised 10°
(5) increased to the boiling point of the solution

(COPY)

41. The expression $[H_2]$ would be read to mean:
    (1) the hydrogen ion concentration
    (2) the hydrogen gas concentration
    (3) weight of hydrogen
    (4) the molar concentration of hydrogen ions
    (5) the normality of the hydrogen

42. A necessary condition for chemical equilibrium is that
    (1) the temperature be above 25°
    (2) the rates of the opposing reactions be unequal
    (3) gases be involved in the reaction
    (4) some of the products of the reaction be removed
    (5) reactions be reversible

43. Iron is known to be a good catalyst for the combination of H_2 and N_2. If an iron catalyst is added to a mixture of H_2, N_2 and NH_3 at equilibrium the result would be:
    (1) no change in the point of equilibrium
    (2) a shift toward formation of more NH_3
    (3) a slowing down of the rate of decomposition of NH_3
    (4) a halt in the reactions
    (5) none of these

44. According to the "mass law" the relative quantities of products and reactants at equilibirum depend upon:
    (1) the total masses of the reactants and products
    (2) the concentrations of reactants and products
    (3) the concentration of the reactants only
    (4) the molecular weights of the products
    (5) all of these

45. Addition of 0.1 N $NaC_2H_3O_2$ to 0.1 N acetic acid will
    (1) increase the ionization
    (2) make the acetate ion 0.2 N
    (3) lower the pH
    (4) make the acetate ion approximately 0.1 N
    (5) make the hydrogen ion approximately 0.1 N

46. Which of the following is an incorrect answer to the statement: "A reaction between ions will go to completion if:
    (1) a very volatile product is formed
    (2) a product is precipitated
    (3) a slightly ionized molecule is formed
    (4) a strongly complexed ion is produced
    (5) only highly ionized products are formed

47. A neutral solution:
    (1) has a pH greater than 7
    (2) contains no OH^- ions
    (3) has equal OH^- and H^+ concentrations
    (4) cannot be a solution of NaCl in H_2O
    (5) is more basic than one whose pH is 10

(COPY)

48. A solution of $AgNO_3$ was added to a NaCl solution and a white precipitate formed. Ammonia solution was added and the precipitate dissolved because?
    (1) NH_4OH is a weak electrolyte (4) NH_3 is a volatile compound
    (2) $NaNO_3$ is a weak electrolyte (5) Ag ions form a complex with NH_3
    (3) $NaNO_3$ forms a complex with NH_3

49. A buffer solution:
    (1) is used to polish brass objects
    (2) is added to explosives to decrease their sensitivity toward violent decomposition
    (3) is a solution of nearly constant pH
    (4) is one containing a strong acid and a salt of that acid
    (5) none of these

50. In the reaction $N_2 + O_2 + \text{heat} \rightleftharpoons 2NO$ an increase in total pressure will:
    (1) not affect the equilibrium (4) favor the left-to-right reaction
    (2) favor the right-to-left reaction (5) none of these
    (3) act as a catalyst

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APPENDIX V

Second Departmental Intra-semester Examination

CHEMISTRY 124
Second Hour Examination
Spring 1954

DIRECTIONS: Mark the <u>one</u> best answer on the test sheets and on the answer sheets. Score = R - W/4

1. Which of the following would be considered proof of a colloidal solution:
(1) no particles visible under a magnifying glass
(2) boiling point of the solution essentially that of the solvent
(3) color in the solution
(4) no precipitation on long standing
(5) a Tyndall cone under a bright light

2. Colloid solutions of a salt differ from true solutions of the same salt with respect to:
(1) presence of charged particles (4) kind of matter in solutes
(2) transparency (5) none of these
(3) molecular weight of solute particles

3. Which of the following is <u>not</u> a colloidal system:
(1) fog (2) jelly (3) corn syrup (4) coffee (5) all are colloidal

4. Charcoal removes the color from sugar solutions because:
(1) sugar is a nonelectrolyte (4) charcoal has a large surface area
(2) charcoal is black (5) charcoal has a low density and will
(3) charcoal is carbon float on the solution

5. The charge observed on many colloidal particles:
(1) is necessarily negative
(2) is of the same magnitude for each particle
(3) is of the same sign for all materials
(4) decreases their stability
(5) none of these

6. Two solutions are mixed forming an insoluble compound. Which of the following is true:
(1) colloids cannot be obtained since the product is insoluble
(2) colloids cannot be obtained unless a protective agent is added
(3) colloids will always be obtained in reactions producing an insoluble compound
(4) colloids will be formed if the crystal growth is not stopped
(5) colloids will be formed if the solutions are stirred

7. Compounds that contain carbon and hydrogen only are called:
(1) organic chemistry (4) hydrides
(2) carbohydrates (3) hydrocarbons (5) petroleum

8. Organic chemistry is a study of:
(1) petroleum (4) the compounds of
(2) medicine carbon
(3) the compounds of carbon and hydrogen (5) biological chemistry

(COPY)

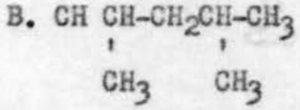
9. Compounds that have the same molecular formula but different structural formulas are:
(1) isotopes (4) isomers
(2) isotonic (3) isooctane (5) dimers

10. Open chain hydrocarbons whose names end with -ane are:
(1) saturated (4) isomers
(2) synthetic (3) unsaturated (5) alcohols

Referring to the following molecular formulas answer questions 11 - 16.

A. $C_nH_{2n} + 2$ C. $H_2C = CH_2$

B.



$CH\ \underset{\overset{|}{CH_3}}{CH}-CH_2\underset{\overset{|}{CH_3}}{CH}-CH_3$

D. HC≡CH E. C_nH_{2n}

11. A saturated hydrocarbon is represented by:
(1) A (2) C (3) D (4) E (5) none of these

12. The formula for ethene is represented by:
(1) A (2) B (3) C (4) D (5) none of these

13. Acetylene is represented by:
(1) A (2) B (3) C (4) D (5) none of these

14. The correct IUC name for B is:
(1) hexane (3) 2,2-dimethyl butane
(2) 2,4-hexane (4) 2,4-dimethyl pentane

15. The IUC name for D is:
(1) acetylene (2) ethyne (3) ethene (4) ethylene (5) none of these

16. The compound that is used to "ripen" citrus fruits is represented by the formula in:
(1) A (2) B (3) C (4) D (5) E

17. The radical formed from saturated hydrocarbons by loss of a hydrogen atom is called:
(1) alkyl radical (4) alkyne radical
(2) alkane radical (3) alkene radical (5) none of these

18. Organic compounds whose names end with -yne are:
(1) saturated hydrocarbons (4) alcohols
(2) hydrocarbons with double bonds (5) acids
(3) unsaturated hydrocarbons with triple bonds

19. Alcohol that is rendered unfit for beverages by adding poisonous chemicals is called:
(1) rubbing alcohol (4) denatured alcohol
(2) absolute alcohol (3) wood alcohol (5) none of these

(COPY)

20. Which of the following is the most active chemically?
(1) ethane (4) octane
(2) propane (3) propene (5) hexane

21. The process of uniting small molecules to make large ones is called:
(1) isomerization (4) pyrolysis
(2) polymerization (3) hydrolysis (5) analysis

22. Protein foods are formed from:
(1) esters (4) alcohols
(2) amino acids (3) fatty acids (5) carbohydrates

23. Soap is prepared by:
(1) reacting an alcohol with hard water
(2) reacting an ester with a strong base
(3) reacting a fat with an organic acid
(4) reacting a base with an acid
(5) hydrolyzing a high molecular weight acid

Listed below are several general formulas for derivatives of hydrocarbons using R as an H or any alkyl radical. Refer to them in answering questions 24-29.

A. $RCHO$ B. ROR C. ROH D. $RCOOH$ E. $RCOOR$

24. Alcohols are represented by:
(1) A (2) B (3) C (4) D (5) none of these

25. Compounds whose names end with -oic are represented by:
(1) A (2) B (3) C (4) D (5) E

26. Fats are represented by the general formula in:
(1) A (2) C (3) D (4) E (5) none of these

27. Ethers are represented by:
(1) A (2) B (3) C (4) D (5) E

28. Esters are represented by:
(1) A (2) B (3) C (4) D (5) E

29. Aldehydes are represented by:
(1) A (2) B (3) C (4) D (5) E

30. A $KMnO_4$ solution is decolorized by a sample of petroleum. This shows that:
(1) the petroleum sample is a member of the methane series
(2) the petroleum sample is unsaturated
(3) the petroleum sample is a hydrocarbon
(4) the petroleum sample is volatile
(5) the petroleum sample is straight-run gasoline

31. The product formed when a primary alcohol is oxidized one-step is:
(1) an ether (4) an aldehyde
(2) an ester (3) an acid (5) a ketone

(COPY)

32. The lead in an ordinary lead pencil is:
(1) diamond (4) graphite
(2) pure metallic lead (3) amorphous carbon (5) none of these

33. Organic acids react with alcohols to give:
(1) ethers (4) salts
(2) esters (3) ketones (5) soap

34. The name of the compound represented by the formula $(C_6H_{10}O_5)_x$ is:
(1) grain alcohol (4) starch
(2) sugar (3) milk-sugar (5) rayon

35. Proteins are made by polymerizing:
(1) starch molecules (4) sugar molecules
(2) organic acid molecules (5) CO_2 molecules
(3) amino acid molecules

36. If bromine is added to $CH_2 = CH_2$
(1) CH_2Br-CH_2Br will form (4) $CBr_2 = CBr_2 = CBr_2$ and H_2 will form
(2) $CHBr = CHBr$ and H_2 will form
(3) $CBr_2 = CH_2$ and HBr will form (5) no reaction will occur

37. Which of the following will form a salt with sodium hydroxide?
(1) CH_3CH_3 (4) $HCOOH$
(2) $CH_2 = CH_2$ (3) CH_3OH (5) none of these

38. Colloid particles are larger than ordinary molecules. This explains why:
(1) a colloidal suspension usually shows turbidity when a beam of light is allowed to pass through it
(2) most colloidal particles move in an electric field towards an electrode
(3) colloid particles are in constant motion
(4) soap emulsifies oil
(5) such colloids as gelatin act as "protective agents"

39. Which of the following would be most effective in removing undesirable odors from the air circulating in a closed refrigerator?
(1) solid calcium chloride (4) a mixture containing chlorophyll
(2) activated charcoal
(3) concentrated sulfuric acid (5) rock salt

40. Which of the following ions may form a curdy precipitate with ordinary soap?
(1) Na^- (2) $CO_3^=$ (3) $SO_4^=$ (4) NO_3^- (5) Ca^{--}

41. The carbon in the wood of a growing tree comes from
(1) the decayed matter in the soil
(2) the carbon dioxide in the air
(3) limestone and other carbonates around the roots of the tree
(4) artificial fertilizers
(5) none of the above

(COPY)

42. Which of the following will react with water to produce a gas that will burn in air?
(1) hydrogen perixide (4) zinc metal
(2) potassium chloride (3) calcium carbonate (5) calcium carbide

43. A gas made by the reaction of steam with incandescent coke:
(1) producer gas (4) coal gas
(2) acetylene (3) natural gas (5) water gas

44. Which of the following would be formed by the complete combustion of a liquid fuel containing C, H, and O?
(1) CO_2 and H_2O (4) C and H_2O
(2) CH_3OH (3) CO and H_2 (5) CO and H_2O

45. Which of these statements indicates a result of the process of photosynthesis?
(1) nitrogen is fixed
(2) carbohydrates and free oxygen are formed from CO_2 and H_2O
(3) the carbon in decaying organic matter is released as CO_2
(4) the N_2 of the air is changed to ammonia
(5) energy from sunlight is given out in the form of heat from growing green plants

46. Which of the following compounds is a paraffin (methane series) hydrocarbon?
(1) C_5H_{12} (4) C_6H_6
(2) $C_5H_{11}OH$ (3) $(C_2H_5)_2O$ (5) C_6H_5Cl

47. Which of the following compounds is an organic acid?
(1) $(CH_3)_2CO$ (4) $CH_3{\cdot}CHO$
(2) $C_{12}H_{23}COOH$ (3) CH_3OH (5) C_5H_{12}

48. Which of the following compounds will react most rapidly with bromine?
(1) CH_4 (4) C_2H_6
(2) C_6H_6 (3) C_2H_4 (5) $C_2H_4Cl_2$

49. Which of the following compounds is an alcohol?
(1) $C_3H_5(OH)_3$ (4) $C_2H_5{\cdot}O{\cdot}C_2H_5$
(2) C_2H_5CHO (3) C_6H_{14} (5) HCOOH

50. Which of the following statements applies in general to all organic compounds?
(1) none are soluble in water (4) they possess high melting points
(2) they are covalent (5) they form no metallic compounds
(3) they ionize readily

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APPENDIX VI

Third Departmental Intra-semester Examination

CHEMISTRY 124
Third Hour Exam
Spring 1954

DIRECTIONS: Mark the one best answer on the test sheets and on the answer sheets. Score = R - W/4

1. Metals are distinguished from nonmetals because they
   (1) form oxides
   (2) are ordinarily solids
   (3) occur in nature as salts
   (4) have electrons in their outer atomic shells
   (5) have low electrical resistance

2. All naturally occurring compounds of metals are
   (1) ores
   (2) salts
   (3) minerals
   (4) bases
   (5) alloys

3. Roasting is a metallurgical process whereby:
   (1) oxides are heated until melted
   (2) sulfides may be converted to oxides
   (3) coke is used to reduce the ore to the metal
   (4) metals are formed into alloys
   (5) none of these

Use is made in questions 4 through 8 of the following series of elements arranged in order of decreasing activity: Ca, Al, Cr, Fe, Cd, Sn, Pb, H, Cu, Bi, Hg, Ag, Pt. In the following questions if the statement is true mark answer No. 1 and if false mark answer No. 5.

4. This is the order of increasing ease of reduction.

5. All the metals are found in the uncombined state in nature, but more often in the cases of Ag and Pt.

6. This is the order of increasing ease of reaction with oxygen.

7. All the metals can be reduced by electrolysis of certain of their compounds.

8. The metals toward the middle of the series are more likely to be found as sulfides than those at either end.

9. Alloys:
   (1) are electrovalent compounds
   (2) have properties identical with those of the component metals
   (3) may be solid solutions
   (4) cannot be made containing more than two metals
   (5) none of these

10. A sample of anhydrous $AlCl_3$ is desired. Which reaction would be best suited
   (1) hydrochloric acid + aluminum
   (2) aluminum hydroxide + hydrochloric acid
   (3) aluminum + chlorine
   (4) barium chloride solution + aluminum sulfate solution
   (5) aluminum nitrate solution + hydrogen chloride gas

(COPY)

11. The following oxides were gently heated. Which decomposes most easily?
(1) BaO (4) CuO
(2) Fe_2O_3 (3) Ag_2O (5) Al_2O_3

12. Which of the following would be soluble in water?
(1) $BaSO_4$ (4) $CaCO_3$
(2) Ag_3PO_4 (3) $Pb(NO_3)_2$ (5) ZnS

13. Which would be the strongest base?
(1) $Al(OH)_3$ (4) $Zn(OH)_2$
(2) $Ba(OH)_2$ (3) $Fe(OH)_3$ (5) $Ni(OH)_2$

14. Which would be insoluble in water?
(1) $(NH_4)_2SO_4$ (4) $NiSO_4$
(2) $Mg(NO_3)_2$ (3) $FeCl_3$ (5) $PbCO_3$

15. In the metallurgy of iron the coke added to the blast furnace:
(1) makes steel (4) serves as fuel to melt the iron
(2) reduces iron oxides (5) none of these
(3) oxidizes the molten iron

16. Steel never:
(1) contains carbon (4) is made by the Bessemer Process
(2) is harder than pure iron
(3) contains any metal other than iron (5) is pure iron

17. The tempering of steel involves:
(1) addition of coke to the molten steel
(2) separating the slag from the metal
(3) addition of limestone
(4) allowing cementite crystals to grow from solution to iron
(5) producing a more brittle steel by heat treatment

18. The rusting of iron:
(1) involves the formation of $Fe(OH)_3$
(2) is a problem of scientific interest only
(3) cannot be controlled by painting the surface
(4) occurs more readily with pure iron than with less pure iron
(5) none of these is correct

19. Iron salts:
(1) do not hydrolyze readily (4) have a single valence
(2) are called cupric and cuprous (5) none of these is correct
(3) may act as oxidizing and reducive agents

20. Iron stains:
(1) are caused by ferrous hydroxide
(2) may be removed by oxidizing agents
(3) are caused by hydrolysis of ferric compounds
(4) easily wash out of fabrics with distilled water
(5) none of these is correct

(COPY)

21. Transition metals:
    (1) have inner electron shells with more than 8 but less than their maximum number
    (2) generally have colorless ions
    (3) have a single valence
    (4) are relatively poor catalysts
    (5) none of these

22. Which of the following metals are transition metals:
    (1) zinc (4) cobalt
    (2) cadmium (3) aluminum (5) barium

23. Which of the following does <u>not</u> apply to nickel:
    (1) imparts corrosion resistance to iron alloys
    (2) Ni^{++} in water is green
    (3) forms complex ions readily
    (4) catalyzes the hydrogenation of oils to give fats
    (5) none of these

24. Which of the following has the lowest electrical resistance:
    (1) iron (3) steel (5) graphite
    (2) copper (4) nickel (6) sand

25. In the electrorefining of copper:
    (1) impure copper is made the positive electrode
    (2) solium sulfate is the electrolyte
    (3) a high voltage is required
    (4) molten cryolite (Na_3AlF_6) is the electrolyte
    (5) pure copper is dissolved at the negative electrode

26. Which of the following pairs are not matched:
    (1) Cu^{++} - oxidizing agent
    (2) Ag^{+} - difficult to reduce
    (3) Al^{+++} - insoluble hydroxide
    (4) Zn^{++} - oxide used in paints
    (5) Hg_2^{+} - found in calomel

27. Silver:
    (1) is soluble in hydrochloric acid
    (2) will be precipitated from solution by metallic zinc
    (3) forms a large number of soluble salts
    (4) lies above hydrogen in the activity series
    (5) carrodes rapidly in moist air

28. Which of the following is <u>not</u> true:
    (1) $AgCl$ - soluble in ammonium hydroxide
    (2) $AgBr$ - affected by light
    (3) $AgNO_3$ - precipitated from solutions of $AgOH$ and HNO_3
    (4) Au - alloys readily with Ag
    (5) Ag - readily amalgamated

(COPY)

29. One could not distinguish between Cu^{++} and Ag^{+} in a solution by:
(1) adding HNO_3
(2) adding HCl
(3) color of solution
(4) adding a piece of zinc
(5) none of these

30. In photography the developer:
(1) dissolves out the unexposed AgBr
(2) is an oxidizing agent
(3) stops the action of light in darkening the film
(4) reduces the exposed AgBr to silver
(5) none of these

31. Which of the following elements are incorrectly matched in valence:
(1) Ag - (+1)
(2) Zn - (+2)
(3) Cu - (+1, +2)
(4) Hg - (+2, +3)
(5) Ca - (+2)
(6) Al - (+3)

32. Which of the following is incorrectly written:
(1) $Al(NO_3)_3$
(2) Mg_2CO_3
(3) AgI
(4) $CoCl_2$
(5) CuO

33. Sodium bicarbonate:
(1) is Na_2CO_3
(2) decomposes on heating
(3) gives a red flame test
(4) is unreactive toward most acids
(5) none of these

34. Alum ($KAl(SO_4)_2 \cdot 12\ H_2O$) is used in baking powder because:
(1) it contains $SO_4^{=}$
(2) it contains K^+
(3) it is an acid
(4) it hydrolyzes in water giving an acid solution
(5) KOH is a strong base

35. Which of the following are similarities between the alkaline earth metals and the members of the zinc family:
(1) valence
(2) uses
(3) ability to form complex ions
(4) flame tests
(5) reactivity

36. Which of the following metals would not react with HCl to give H_2:
(1) Zn
(2) Al
(3) Mg
(4) Fe
(5) Cu

37. On the basis of the periodic table which of the following predictions is incorrect:
(1) $B(OH)_3$ would be a stronger base than $Al(OH)_3$
(2) $Ba(OH)_2$ would be a stronger base than $Mg(OH)_2$
(3) Mg would be a more active metal than Al
(4) Ca should resemble Ra closer than Al
(5) none of these

38. Which of the following reactions indicates amphoteric behavior:
(1) $ZnCl_2 + H_2S \longrightarrow ZnS + 2HCl$
(2) $NaOH + HCl \longrightarrow NaCl + H_2O$

(COPY)

(3) $3NaOH + Al(OH)_3 \longrightarrow Na_3AlO_3 + 2H_2O$
(4) $Al + Cl_3 \longrightarrow AlCl_3$
(5) $Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3$

39. It is desired to manufacture a white paint pigment. Which of the following reactants would be the most suitable:
(1) $Ca(NO_3)_2 + Na_2CO_3$
(2) $CdCl_2 + H_2S$
(3) $Al_2(SO_4)_3 + NaOH$
(4) $BaCl_2 + H_2SO_4$
(5) $BaS + ZnSO_4$

40. The protection of iron from rusting by covering with zinc is called
(1) pickling
(2) cracking
(3) refining
(4) oxidizing
(5) galvanizing

41. Ruby, Topaz, and sapphire are
(1) ZnO
(2) CaO
(3) Al_2O_3
(4) Fe_2O_3
(5) CuO

42. Hg_2Cl_2 is:
(1) Bauxite
(2) quicklime
(3) lithopone
(4) Calomel
(5) borax

43. The "setting" of plaster of Paris involves:
(1) formation of limestone
(2) formation of aluminum silicate
(3) oxidation of lime
(4) drying out of the plaster
(5) formation of gypsum

44. The use of borax in glazing pottery and in ironware enamels depends upon:
(1) its containing oxygen
(2) its high melting point
(3) the formation of a glassy solid on solidifying
(4) its solubility in water
(5) none of these

45. What weight of aluminum metal could be obtained on electrolysis of 2000 pounds of aluminum oxide which was 80% pure?
(1) 847
(2) 1230
(3) 629
(4) 1138
(5) 720

46. Mark the same answer as in question 45.

In questions 47 through 50 complete and balance the indicated equations. Add together the number of molecules of both reactants and products and mark the corresponding answer.

47. $Al + HCl \longrightarrow$ (1) 5 (2) 10 (3) 9 (4) 13 (5) 8

48. $Zn(OH)_2 + NaOH \longrightarrow$ (1) 4 (2) 6 (3) 8 (4) 7 (5) 12

(COPY)

49. $Fe_2(SO_4)_3$ + H_2O ----> (1) 10 (2) 8 (3) 14 (4) 11 (5) 12

50. $Na(HCO_3)_2$ + heat ----> (1) 5 (2) 4 (3) 6 (4) 8 (5) 9

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Vita

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Thesis: AN EXPERIMENTAL STUDY IN THE TEACHING OF GENERAL CHEMISTRY AT THE COLLEGE LEVEL

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THESIS TITLE: AN EXPERIMENTAL STUDY IN THE TEACHING
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