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The presentation and evaluation of the first two units in elementary chemistry

Wescott, Ralph Charles

Boston University

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BOSTON UNIVERSITY
SCHOOL OF EDUCATION

Thesis

THE PRESENTATION AND EVALUATION
OF
THE FIRST TWO UNITS IN ELEMENTARY CHEMISTRY

Submitted by

Ralph Charles Wescott
(A.B., Brown University, 1931)
In partial fulfillment of requirements for
the degree of Master of Education

1947

First Reader  Dr. Roy O. Billett, Prof. of Education
Second Reader  Vaden E. Miles, Asst. Prof. of Education
Third Reader  Franklin C. Roberts, Prof. of Education
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CHAPTER I
THE SCOPE AND PURPOSE OF THIS THESIS

The problem: Reference to the tables of contents and the first few pages of a number of high school textbooks of chemistry shows that there is general agreement \(^1\) that the first unit of a course in general chemistry should include fundamental definitions of such terms as matter, substance, property, element, compound, and mixture, and should develop the idea of the nature of a chemical change. The usual order of topics after this has been: Oxygen; Hydrogen; Water; Atomic Theory; Symbols; Formulas and Equations; Sodium; Chlorine; Acids, Bases, and Salts; and so on down the list with only minor variations \(^1\).

In this era of the use of atomic energy it should no longer be necessary to lay such an elaborate foundation before mentioning the atomic theory; and the elementary ideas about atoms and the use of symbols, formulas, and equations should be incorporated in Unit I. Since the teaching of general science has become common in the grades

\(^1\) John C. Hogg and Charles L. Bickel, Elementary General Chemistry. D. Van Nostrand Company. New York. 1941
Chapter I

THE WORLD OF THE ROWING BOAT

The life of the rower is one of constant striving for improvement and growth. Through the use of a variety of equipment and techniques, the rower develops a unique understanding of their craft. The world of the rowing boat is a complex one, requiring a deep understanding of physics, biology, and engineering.

The crew boat is a powerful machine, capable of great speed and agility. The crew members must work in perfect harmony to achieve these goals. The coxswain serves as the leader of the team, guiding the crew through each stroke, motivating them to give their best effort.

The stroke is the most important aspect of rowing. It involves the simultaneous use of power and technique to propel the boat forward. The rower must maintain a steady rhythm, staying focused on the task at hand.

In addition to the physical demands of rowing, the mental aspect is just as important. The rower must be able to push through the pain and stay focused on the goal. The mental strength required to maintain composure and concentration is just as crucial as the physical strength.

The world of the rowing boat is one of challenges and rewards. The rower must work hard to improve their skills and stay up to date with the latest techniques and equipment. The reward is a sense of accomplishment, pride in their performance, and the satisfaction of pushing one's limits.
and in the junior high school, it is no longer necessary to discuss oxygen and hydrogen in detail at the beginning of the course. It is proposed, therefore, that the second unit should take advantage of this general knowledge and concern itself with the topic of water and solutions. This arrangement postpones the study of difficult chemical changes such as the decomposition of oxides and chlorates and the displacement of hydrogen from acids. It also postpones the use of the techniques of collecting gases over water. The pupil is allowed to work for several weeks with familiar substances and simple apparatus and thus becomes gradually accustomed to the use of chemical substances and apparatus.

The main object of study is a tangible substance like water, but the student has the opportunity to think of the hydrogen and oxygen atoms in the water molecule. Since this approach is psychologically more sound, the subject should be easier for the mentally immature pupil to grasp. This arrangement does not introduce new material into the course to displace commonly required topics; it merely attempts to collect easier materials at the beginning of the course.

The local situation: These units were developed and taught at the Manlius School in Manlius, New York, about nine miles east of Syracuse. It is a military preparatory school of high scholastic standing, accredited by the New York State Department of Education and the Association of Colleges and Secondary Schools of the Middle
The following must be specified to the notice and development of the public, in order to prevent any possible misunderstanding or confusion:

1. The primary purpose of the notice and development of the public is to inform and educate people about the topic.
2. The notice should be clear and concise, avoiding unnecessary jargon or complex language.
3. The development of the public should be ongoing, with regular updates and feedback mechanisms in place.
4. Collaboration with relevant stakeholders is crucial for effective communication and implementation.

The notice and development of the public is a critical component of any successful project or initiative.
States and Maryland. It was raised to the level of Military Institute in 1946 by authorization of the War Department.

In 1945 the students were classified into three groups on the basis of a battery of tests and their past school records. The X-group is limited to students who are preparing for college entrance; the Y-group consists of students who are preparing to take the New York Regents' Examinations; and the Z-group is designed to train students on the non-college level in core secondary school subjects.

Classes are kept small in order to provide as much individual instruction as possible, but they are large enough to provide the necessary mental stimulation and competition within the group.

The chemistry laboratory contains individual locker and working space for eighteen students. Each locker is equipped with the ordinary chemical apparatus such as is listed in the laboratory manuals written to accompany such textbooks as those referred to at the beginning of this chapter. A complete set of the frequently used reagents is within the reach of each student. The benches are supplied with running water and gas. Excellent illumination is provided by modern fluorescent lamps.

The class room, adjoining the laboratory, has seats arranged in rows elevated one above the other so that all the students have a clear view of the demonstration desk. Charts are available, but there is no motion picture equipment.
SECRETARY WAR DEPARTMENT.

To the Command of the President of the United States:

In 1901 the United States War Department, after careful consideration of the educational needs of the nation, established the United States Military Academy at West Point, New York. The Academy is charged with the responsibility of preparing young men for leadership in the armed forces of the nation.

The Academy provides a rigorous four-year program that combines academic study, military training, and physical fitness. Graduates of the Academy are commissioned as second lieutenants in the United States Army and are immediately responsible for leading and training soldiers.

The Academy is located on the banks of the Hudson River, near the city of West Point, New York. The campus includes barracks, classrooms, laboratories, and athletic facilities.

The Academy's primary mission is to prepare young men for leadership in the armed forces of the nation. In addition to academic and military training, the Academy also emphasizes the development of character and leadership skills.

The Secretary of War, on behalf of the President of the United States, respectfully requests your consideration of this proposal for the expansion of the Academy's facilities and the increase in its enrollment.

[signature]
Secretary of War
It is unfortunate that the windows look out upon the school parking lot.

The total enrollment of the school is 364. There are six chemistry classes: one prepares for the College Board Examinations; four prepare for the New York Regents' Examinations; and one is a class in general chemistry without college entrance credit. The units have been prepared and taught in the general class; but they include as optional topics all the material necessary to meet college entrance requirements.

The class in which these units were taught consisted of eight boys from Connecticut and New York, two of whom were Spanish boys from South America. Their I.Q.'s ranged from 88 to 104. Their Reading Comprehension ranged from 70 to 1 percentile (Educational Records Bureau private school scale) and their Mechanics of English from 7 to 1 percentile.

The nature of a unit:— The units are an effort to apply the principles and method set forth by Dr. Roy O. Billett of Boston University. According to Dr. Billett, 1/a unit is a statement of changes to be sought in pupils' capacities for and tendencies toward behavior. A unit is a unit of learning and a not/mere topic of some thing to be learned. A unit for

teaching should be organized in four parts: (1) A statement in complete declarative sentences of the fundamental changes desired; (2) a statement of the detailed facts to be taught; (3) references for the teacher; and (4) the unit assignment, part of which is mimeographed and given to the pupils. 1/

In the following pages is given the material of each unit, organized as above.

The "optional related activities" 2/ were typed on cards and made available to the class at all times.

The experiments called for by the assignment sheet and the optional activities list are given in Appendix A.

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1/ Billett, op. cit., pp 505-509
2/ Ibid, pp 507-508
In the following table to know the necessity of each

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CHAPTER II

THE ORGANIZATION OF THE UNIT ON CHEMICAL CHANGE

General Statement of the Unit

In a chemical change the materials forming one or more substances are transformed into new and different substances with new and different properties. Chemical changes can be explained in terms of the arrangement and rearrangements of the atoms of the elements. Many of the facts of chemistry can be described by the use of a sort of chemical shorthand using symbols, formulas, and equations.

Itemized Statement of the Unit

1. The theories of chemistry are based on facts determined by experiment.

2. In a chemical change, the substances involved undergo changes in composition so that the products are not the same substances as the original materials. In a chemical change heat is usually produced or absorbed and light is often emitted.

3. The properties of a substance are the characteristics by which we recognize it, such as color, taste, solubility in water, density, melting point, odor, and its behavior in chemical changes.
CHAPTER

THE ORGANIZATION OF THE UNIT OR CHEMICAL TANKER

Central

In a development the certainty of a central organization of a concrete and easy solution can be achieved in terms of the extraction of the test solutions at the same time. The nature of the properties may be seen to be the ease of obtaining new and different properties. Central organization of a central organization of the properties at the same time. The nature of the properties may be seen to be the ease of obtaining new and different properties.
4. A substance is one kind of matter; all its parts are alike and show the same properties. Salt, sugar, and lead are substances. Milk is not a substance, but a material. A material is any portion of matter distinct enough to have a special name. It may be a pure substance like salt or sugar or a mixture of substances like air, gasoline, or milk. A material composed almost entirely of one substance but containing small amounts of another substance is said to be "impure". (Water is a substance. Two of its properties are that below 0°C it is a solid and above 100°C it is a gas. The common name for water in the solid state is ice. Ice is a substance; all its parts are alike and show the same properties. Ordinary commercial ice may contain air bubbles as an impurity; these air bubbles may make the material opaque. Ordinary ice may contain other impurities also, such as dirt, bacteria, or small quantities of dissolved substances mechanically held in the water crystals.)

5. Matter may exist in three physical states: solid, liquid, and gas.

6. A compound is a substance formed by the combination of two or more simpler substances. The ingredients of a compound are called its constituents. For example, the compound iron sulfide is formed by the combination of the simpler substances iron and sulfur. Iron and sulfur are the constituents of iron sulfide.
7. A mixture is a material composed of more than one substance each of which retains its own properties. If iron and sulfur are merely mixed together, the iron remains iron and is attracted by a magnet, dissolves in hydrochloric acid, and does not dissolve in carbon disulfide; the sulfur is still sulfur and will not dissolve in hydrochloric acid, but will dissolve in carbon disulfide. Iron and sulfur are the components of the mixture. The material brass is a mixture of the components copper and zinc. A mixture of two metals, made by melting them together and then cooling, is called an alloy.

8. A mixture can be separated into its components by a purely physical process. Thus the mixture of iron and sulfur can be separated by dissolving the sulfur in carbon disulfide, filtering the solution, and allowing the solvent to evaporate. It is always theoretically possible to separate a mixture into its components; but it may be practically impossible. For example, gasoline is a mixture of so many substances that it is almost impossible to separate every one. A compound can be separated into its constituents only by a chemical change. The compound iron sulfide is a different substance from iron and sulfur; therefore, the decomposition of this compound would mean the production of new substances and this is, by definition, a chemical change.
A variety of methods of atomic physics and quantum mechanics have been employed to study the properties of atoms and molecules. These methods have led to a deeper understanding of the behavior of particles at the subatomic level and have had significant implications for the development of modern physics. The principles of quantum mechanics, in particular, have revolutionized our understanding of the physical world, leading to the development of new technologies and applications in fields such as information technology and materials science. The study of atomic physics continues to be a vibrant area of research, with ongoing efforts to explore new aspects of the behavior of matter and energy.
9. An element is a substance so simple that it cannot be decomposed into anything simpler by chemical means.

(It is recognized that the definition of "element" is almost impossible without using terms which are beyond the comprehension of high school students. The definition above, first suggested by Robert Boyle, is still the most satisfactory one for elementary use. It is also recognized that the insertion of the qualifying phrase "by chemical means" is really unfair because we are thus permitted to assert that anything which "splits the atom" is not "by chemical means". Nevertheless, nuclear reactions are definitely not ordinary chemical reactions.)

10. Elements unite to form compounds and compounds can be decomposed into elements.

11. Every chemical compound is composed of elements united in definite proportions by weight.

12. There is no gain or loss in the total weight of the materials taking part in a chemical change.

13. All the elements consist of very small particles called atoms.

14. In a specimen of an elementary substance, all the atoms are alike; but in a specimen of a compound substance, atoms of different elements are found combined in definite proportion.

15. Elements may combine to form more than one compound, but the proportions by weight in each case will be distinct and definite; the proportions do not change gradually from one compound to the next as in the case of
mixtures. For example, in the iron sulfide prepared by direct combination in the laboratory, the proportions of iron and sulfur are 7 parts of iron to 4 parts of sulfur, but in the mineral "fool's gold" the proportions are 7 parts of iron to 8 parts of sulfur. Iron and sulfur may combine in either of these two proportions, but never 7 parts of iron to 5, 6, or any other number (except 4 and 8) parts of sulfur.

16. An element is represented by a symbol which stands not only for one atom of the element but also for a definite weight of the element. For example, O stands for one atom of oxygen, 16 parts by weight of oxygen or 16 atomic weight units; Ag stands for one atom of silver, 107.880 atomic weight units.

17. A compound is represented by a formula which expresses the elements composing the substance, the proportions in which the atoms of these elements are combined, and the proportions by weight of each element in the compound. Thus Na₂CO₃ represents (1) the substance sodium carbonate; (2) sodium carbonate is composed of sodium, carbon, and oxygen in chemical combination; (3) sodium carbonate contains two atoms of sodium to one atom of carbon to three atoms of oxygen; (4) sodium carbonate contains $2 \times 23$ or 46 parts by weight of sodium, 12 parts by weight of carbon, and $3 \times 16$ or 48 parts by weight of oxygen; (5) $46 + 12 + 48 = 106$ parts by weight of sodium carbonate.
18. The number of grams of an element equal to its atomic weight is called a gram atom. A formula represents a weight called the formula weight or the molecular weight; if this weight is expressed in grams, it is called a mole. For example, one gram atom of oxygen weighs 16 grams. One mole of sodium carbonate weighs 106 grams; it contains 46 gram atoms of sodium, 12 gram atoms of carbon, and 48 gram atoms of oxygen.

19. A chemical reaction can be represented by an equation which shows the symbols and formulas of the substances which react and the symbols and formulas of the products. For example, the reaction between iron and sulfur can be represented by the equation

\[
\text{Fe} + \text{S} \rightarrow \text{FeS},
\]

which is read, "Iron plus sulfur yields iron sulfide".

20. Since symbols and formulas represent weights, an equation can be used to calculate weight relationships in a reaction. The atomic and formula weights may be written beneath the formulas in the equation

\[
\begin{array}{c}
\text{Fe} + \text{S} \rightarrow \text{FeS} \\
56 + 32 \rightarrow 88
\end{array}
\]

Since these weights represent the proportions in which iron and sulfur combine, it follows that if we know the weight expressed in grams, pounds, tons, or any other unit of either the iron, the sulfur, or the iron
sulfide involved in any actual reaction, we can calculate the corresponding weights of the other two. For example, what weight of iron sulfide can be prepared from 10 grams of pure iron?

\[
\begin{align*}
10 \text{ grams Fe} & \quad + \quad 56 \text{ grams S} \quad \rightarrow \quad x \text{ grams FeS} \\
\frac{10}{56} & = \frac{x}{88} \\
x & = 15.7 \text{ grams}
\end{align*}
\]

List of Materials and Readings for Teachers' Use


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To mine

2.50

Total

10.00

I have placed your order for the following:

1. A typesetting machine, model "Maxima".

I will provide a 30-day trial period. If you are satisfied, you will pay the full amount of $150. If not, you may return the machine within the trial period for a full refund.

Thank you for your business.

[Signature]

July 15, 1989

[Company Letterhead]
The Unit Assignment

Books required by the study guide for students' use:

Raymond B. Brownlee, Robert W. Fuller, William J. Hancock, and Jesse E. Whitsit, *Chemistry in Use*, Allyn and Bacon, New York, 1939. (E)


Study Guide for Unit 1 1/

1. Introductory talk by the teacher. Point out that students are accustomed to many articles of color. In ancient times colors were reserved for the rich. Fountain pens are made from recently developed plastic materials. New chemical products are constantly appearing on the market.

Chemistry is concerned with the changes that take place in matter, we can break a stick of wood in two (break it) and it is still wood; but in a chemical change the entire nature of the material is changed. (Without comment, drop glycerine from a medicine dropper on a previously prepared cone of powdered ammonium bichromate topped with a little powdered potassium permanganate.) Sometimes chemical changes are accompanied by light (hold with forceps a piece of magnesium ribbon in a Bunsen flame.) Sometimes they are

1/ Items marked with an asterisk * are on the students' guide. Parts of the item enclosed in parentheses, however, are for the teachers' guidance and do not appear on the students' guide.
The Title

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[Signature]

[Handwritten notes]

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accompanied by explosions. (At one side of the desk should be a small rubber balloon which has been blown up with a mixture of hydrogen and oxygen from electrolysis. While student eyes are still blinded from the brilliant light of the burning magnesium, move the burner over to set off the balloon. Use extreme care!) Sometimes chemical changes are accompanied by changes in color. (Four cylinders stand on the desk containing solutions of (1) potassium thiocyanate KSCN, (2) potassium ferrocyanide, (3) dilute hydrochloric acid and a bit of steel wool to produce fresh ferrous chloride, and (4) sodium chloride. To (1) and (2) add from a beaker dilute ferric chloride to produce a red color and a dark blue precipitate.) If we would like a lighter blue (add from a beaker some potassium ferrocyanide solution to (3), or even a white (add silver nitrate solution to (4), we can obtain these colors also.

(Mix in a large graduate equal volumes of two solutions made by adding to one beaker a quantity of sodium bicarbonate and to another a quantity of cream of tartar. Add water to each, stir, and pour together.)

We can even produce heat and cold. (Fill two large test tubes half full of water. Cautiously add sulfuric acid to one, solid ammonium nitrate to the other. Stir. Have the pupils feel the two tubes.)

2. Pretest.
#3. What is matter? In what three physical states can matter exist? What are properties? What is a substance? (B 22-25; P & B 19; B & C 13, 16; J 1-4, 6)

4. Demonstration to illustrate chemical change.

(a) Conical pile of powdered ammonium bichromate topped with fresh sodium peroxide. Add drop of glacial acetic acid. (Or top with powdered potassium permanganate and add a drop of glycerine.)

(b) Add dilute nitric acid to a few pieces of copper in a beaker. Note brown fumes. Evaporate solution in an evaporating dish and note that the material remaining is entirely unlike copper.

(c) Add 12 ml. concentrated sulfuric acid to a mixture of 12 g. sugar and 9 ml. water in a beaker.

(d) Mix 3 g. zinc dust with 1 g. powdered sulfur. Place on asbestos square and direct the flame of a bunsen burner on it.

(e) Pack pieces of zinc around a thermometer in a large test tube. Add dilute hydrochloric acid. Note bubbles of hydrogen. Point out that in every case a new material was formed and that heat or light or both accompanied the change. Point out that in a physical change no change occurs in the nature of the materials.

#5. Experiment 1. How can we identify substances by their properties?
6. Experiment 2. How to separate a mixture into its components. (This experiment has a double purpose: (1) to learn how to separate a mixture by taking advantage of the physical properties of its components; and (2) to learn some experimental techniques. In order to teach the techniques, the teacher will perform the experiment as a demonstration, step by step. The class will perform each step immediately after the teacher. After his demonstration, the teacher will observe and correct the work of the students. Write on the blackboard and teach the meaning of the terms: components, filtration, filtrate, residue, evaporation, crystallization, solvent.)

7. Experiment 3. What is a compound? (Precede by teacher demonstration.)

8. Experiment 4. How does a compound differ from a mixture? (First demonstrate use of mortar and pestle. Grind, with pressure; do not pound. Exercise careful supervision.)

9. Experiment 5. Decomposition of a compound. What are elements?

10. Read several of the available texts to obtain further understanding of chemical changes, mixtures, compounds, and elements. (B 34-39, 46-51, 66-67; B & C 14-17, 21-29; P & B 4-27; J 5-9.)

11. Write the answers to the following questions:

(1) Name three materials composed of more than one substance.
(2) Give examples of some familiar substances which are solids, some which are liquids, and some which are gases.

(3) What part does energy play in a chemical change?

(4) How could you separate a mixture of sand and sugar?

(5) How could you separate a mixture of powdered iron and powdered charcoal? (6) Classify the following materials into three columns headed "elements", "compounds", and "mixtures": bread, nylon, diamond, sea water, milk, salt, sugar, catsup, lemonade, iron, silver, granite, brass, sulfur, mercuric oxide, zinc sulfide, tin. Add three examples of your own to each list.

#12. How are chemical compounds formed by the combination of two elements named? (B 73; P & B 10; B & C 98; J 98.) Name the compounds formed when the following pairs of elements combine: phosphorus and chlor(ine), zinc and ox(ygen), sodium and sulf(ur), hydrogen and ox(ygen), magnesium and nitr(ogen). The parentheses suggest the parts of the words which are changed in forming the name.

#13. Two types of chemical change are decomposition and combination. Write a good definition of each and give two examples of each. (B 39-42, 46-51; P & B 210; B & C 104-105; J 130-131.)

#14. What are laws in science? How do they differ from laws passed by legislative bodies? (Dictionary; J 15-19; B & C 79-80.)

15. Is any weight gained or lost in a chemical change? Demonstration: Weigh a flask containing a solution
of copper sulfate and a test tube containing caustic soda solution on a platform balance, mix the solutions, and reweigh.

16. What is the law of conservation of matter? (P & B 17; B & C 18-19; J 19-20.)

17. What is energy? What is the law of conservation of energy? (P & B 21-27; B & C 423-424; J 2-23.)

18. Do elements combine in any definite proportions or in whatever proportions they happen to be mixed before combination? (B 68-69; P & B 10; B & C 25-26; J 71.) Experiment 6.

19. Careful experiments show that 22.997 g. of sodium will combine with 35.457 g. of chlorine. What weight of sodium chloride will be formed? Name and state the law which applies in this case. Calculate the percentage of sodium in sodium chloride. Calculate the percentage of chlorine. (Round off the decimals before making the calculation.) Name and state the law which applies in this case.

20. Talk by the teacher to introduce and teach the atomic theory and the use of symbols, formulas, and equations. (1) Explain Dalton's atomic theory (B & C 81); (2) Show how it explains the law of definite proportions (B & C 83-84, J 85); (3) Show how it explains the law of multiple proportions. Instead of naming and stating this law with non-college classes, merely show how elements can combine in more than one proportion. Use as examples: carbon monoxide and dioxide, ferrous sulfide and iron
pyrites, and the oxides of nitrogen. Use Dalton's symbols to illustrate the oxides of nitrogen and the oxides of carbon. Then show how much more convenient it would be with modern symbols and use them for the sulfides of iron. (4) Explain that "symbol" stands for one atom of an element and a definite weight.

Emphasize that symbols have the first letter and the first letter only capitalized. (5) Explain the use of oxygen = 16 as a standard of atomic weights. (6) Explain how we group symbols into formulas to represent the actual percentage composition of a compound.

Emphasize that formulas are derived by experiment and that they represent the proportions in which the elements combine and not the actual number of atoms present. Illustrate how we add up atomic weights to get the "formula weight", sometimes called "molecular weight".

(7) Show how formulas can be combined in equations to represent chemical reactions. Use simple examples such as iron, copper, and silver with sulfur. Avoid gases such as O₂, N₂, or H₂. Avoid the use of the word "molecule" and caution students to disregard it when they see it in their reading.

#21. What are atoms? Is the existence of atoms a fact or a theory? What are some of the reasons for believing in atoms? How does our belief in atoms help us to understand how iron and sulfur can combine in the ratio of 7 parts of iron to 4 parts of sulfur to form 11
parts of iron sulfide? How does it also help us to understand that 7 parts of iron can combine with 8 parts of sulfur to form 15 parts of iron pyrites (fool's gold), but not with 5, 6, or 7 parts of sulfur? (B 68-70; P & B 5-10; B & C 80-85; J 82-87.)

#22. What are the advantages of using symbols to represent elements? How are symbols derived? What facts are represented by a symbol? (B 43, 70, 103; P & B 6; B & C 90, 91; J 92-93.)

#23. What is the standard of atomic weights? What is meant by the statement that the atomic weight of carbon is 12? (B 299-300; P & B 218; B & C 81-82; J 135-138.)

#24. Represent the reaction between iron and sulfur by a chemical equation. What does this equation mean both in terms of atoms and in terms of weights? Write two other equations and state the meaning of each both in terms of atoms and in terms of weights. (B 71, 72, 300-302; P & B 210-212; B & C 101-102; J 124-125.) Show by reference to the weights in two different equations that the law of conservation of matter applies to a chemical equation.

#25. What information about sodium carbonate is given by the statement that its formula is \( Na_2CO_3 \)? The formula of copper sulfate is \( CuSO_4 \). Explain. (B 70, 303; P & B 218-219, B & C 91-92; J 139.)

26. Talk by the teacher. First clear up difficulties students have been having with formulas and
The image contains text that is not legible, making it impossible to transcribe.
equations. Point out that formulas represent the actual composition of compounds and are determined by experiment. (There are certain general rules of "valence" which we shall study in due time and which will help us to remember formulas.) First let us see how we can calculate the formula of copper sulfide from the data obtained in Experiment 6. Point out that opportunity is given (Optional Related Activities) for students to determine the formula of tin oxide. Keep the discussion of formulas and equations elementary as an easier way to express facts than writing them out in words. Explain how equations can be used in chemical calculations.

#27. What weight of iron sulfide can be made from 20 pounds of iron? What weight of iron would be needed to make 440 grams of iron sulfide? What weight of iron sulfide can be prepared from 5 ounces of iron? If 2.36 g of iron were heated in a crucible with an excess of sulfur, what weight of sulfur would combine with the iron? (B 301, 302; P & B 222-224; B & C 111.)

Optional Related Activities

1. Visit a chemical plant or laboratory. Talk with a chemist about his work. What type of work is done in the laboratory you visited. Are the chemists testing materials made or purchased or are they trying to make new products themselves? Can you describe some of the equipment? What opportunities would there be in such a laboratory for a high school graduate with an
interest in chemistry? Would a college degree be necessary to work in this laboratory? Would it help?

2. Make a collection of newspaper clippings indicating the importance of chemistry in the daily news. Paste the clippings in a scrap book and label each with the name of the paper from which it was taken and the date. Classify the clippings under certain general topics such as nutrition, agriculture, pharmacy, industry, atomic power, chemical warfare, biochemistry, etc.

3. Make a collection of advertisements from newspapers or magazines which show the importance of chemistry to the consumer. Paste the clippings in a scrap book. Label each with the name of the paper and the date. To what extent do these advertisements indicate that the public is "chemistry conscious"?

4. How many radio broadcasts about chemistry can you find listed during the period you are studying this unit? Listen to as many of them as you can. Write a report for the class on Chemistry on the Air.

5. It is expected that students in chemistry are more or less familiar with the metric system from their study of general science, biology, or physics. If you need additional study on this topic, refer to the supplementary topic Units of Measurement used in Chemistry. Study the material on the metric system, units of temperature and heat, density, and specific gravity. There are
THEREFORE, it appears that a college course of study is necessary to prepare for the field of practical work. Moreover, the college course is a preparation for life. The college course provides a basis for further study and for the ability to think and reason. It also provides a basis for the development of moral and ethical principles.

The college course is an opportunity to learn how to think independently and critically. It is an opportunity to develop the ability to analyze and solve problems. It is an opportunity to learn how to work effectively as a member of a team. It is an opportunity to learn how to communicate effectively. It is an opportunity to learn how to manage time and resources effectively. It is an opportunity to learn how to lead and influence others.

The college course is also an opportunity to develop personal qualities such as leadership, teamwork, and critical thinking. These qualities are essential for success in any field.

In summary, the college course is an essential component of any career development. It provides a solid foundation for future success and is a necessary step towards achieving personal and professional goals.
three optional experiments outlined and a number of problems are given. You may wish to consult with the teacher as to which of these experiments and problems you ought to do.

6. Determine the formula of the sulfide of some metal. The directions are given in the laboratory guide (Experiment 1-7).

7. Determine the formula of tin oxide (Experiment 1-8).

8. Do the following problems:
   (a) Edward Williams Morely (1838-1923) performed a classic series of experiments about 1895 to determine the exact composition of water. He found that 40.9178 grams of hydrogen combined with 324.8684 grams of oxygen. From these data, what is the formula for water.
   
   (b) Professor Honigschmid, one of the most famous modern German chemists, reported in 1940 an experiment in which he found that 6.25976 grams of silver combined with bromine to form 10.89645 grams of silver bromide. What weight of bromine combined with the given weight of silver? What is the formula for silver bromide?
   
   (c) The greatest American authority on the determination of atomic weights is Professor Baxter of Harvard. In 1940 he reported some experiments to determine the atomic weight of chlorine. He found that 9.00350 grams of silver combined with chlorine to form 11.96264 grams of silver chloride. If the atomic weight of silver is 107.880 and the formula for silver chloride is
The statement we have to prove is true for all primes greater than 2.

To see this, let $p$ be a prime number greater than 2. We need to show that $p^2 - 1$ is always divisible by 4.

By the difference of squares formula, we have

$$p^2 - 1 = (p - 1)(p + 1).$$

Since $p$ is a prime greater than 2, neither $p - 1$ nor $p + 1$ can be divisible by 2. Therefore, both $p - 1$ and $p + 1$ are odd.

Now, let's consider two cases:

1. If $p$ is odd, then $p - 1$ and $p + 1$ are both even.

2. If $p$ is even, then $p - 1$ is odd and $p + 1$ is even.

In both cases, both $p - 1$ and $p + 1$ are even, which means that their product is divisible by 4.

Hence, $p^2 - 1$ is divisible by 4 for all primes $p > 2$.

We have proven that

$$p^2 - 1 = 4k$$

for some integer $k$.
AgCl, calculate the atomic weight of chlorine.

9. If you work in a drug store or have a friend who is a druggist you may have an opportunity to examine the chemicals used in filling prescriptions. Why are so many of the names in Latin? Find out from the druggist or from the catalog of a college of pharmacy the training needed to be a druggist.

10. Begin a collection of elements which you have obtained yourself from compounds and of compounds which you have prepared. Preserve them in small glass vials labeled with the name of the compound or element, your initials, and the date. You may prefer to seal them in glass tubes about three inches long. The teacher will show you how to make these tubes.

11. Have you an original idea which you wish to develop as an individual task. Ask the teacher for his approval before starting.

12. Write brief biographical sketches of one or more of the following: Berzelius, Boyle, Paracelsus, Dalton, Aristotle. Crucibles by Bernard Jaffe or A History of Chemistry by F. J. Moore and William T. Hall contain interesting material on these men.

13. If you are an amateur photographer, take a series of pictures of the apparatus used in the laboratory or a series of pictures showing the proper techniques of heating a test tube, filtering, or evaporating a solution.
Yard, confirming the theory that the opium

IL has not been a good place to study a

1928. A group of international students have an opportunity to

continue the experiments made in the laboratory

which are to be seen at the lecture. This is not to

the Student's Office. If any questions of a scientific or practical

nature arise, you may consult the lecturer.

If you have any difficulties, the reader will be glad to see that you

are developing an interest in the subject.

The student is expected to

in the lecture notes and practice the experiments as they are given.

A new method of acquiring a good knowledge of the

material is to obtain a ready reference to hand and use it as an

auxiliary means of checking the lecture.

If you are an experimental student, you will

obtain a ready knowledge of the material, as well as an

understanding of the principles involved.
14. If you are interested in drawing, make a series of drawings to illustrate your laboratory work, the demonstrations you have witnessed, or some of the concepts of this unit.

15. Do one or more of the following special reading assignments and make a report to the class.


(3) Read the article "Atomic Energy Harnessed in Great Scientific Achievement", Chemical and Engineering News Insert, August 10, 1945. Compare this account with that in Time, August 20, 1945, pp. 31-36. What is the historical importance of these articles? What is their scientific importance?

(4) Much has been said about the peacetime uses of atomic power. Read what Dr. Daniels, one of America's leading chemists, has to say about the future in Chemical and Engineering News, 24: 1514-1517, 1588, 1589 (June 10, 1946). Will atomic power be cheaper
than coal?

(5) In Chemical and Engineering News, 24: pp. 750-758 (March 25, 1946) is given a brief biography of a modern biochemist and his own account of his studies on influenza virus. What is a biochemist? (Dictionary) What are his conclusions about the development of a vaccine for the prevention of influenza?

(6) Read the article on "The Human Side of Chemistry" in Collateral Readings in Inorganic Chemistry, edited by L. A. Goldblatt (D. Appleton-Century Company) New York, 1937, pp. 1-8. This was an address given before the American Chemical Society at Pittsburgh, September 11, 1922, by the late Dr. E. E. Slosson. Dr. Slosson was the first successful chemist in America to undertake to popularize chemistry. This address was given shortly after the close of the first World War. Why is it an important article today? On page 220 are 17 questions on this article. Select and answer 8 of them in writing.

(7) Read the fifth article in Goldblatt's Collateral Readings in Inorganic Chemistry on "The Derivations of the Names of the Elements", pp 23-31. Answer questions 1-3 on page 220.

Moving Picture Films

Films for School and Industry, Castle Films, 30 Rockefeller Plaza, New York 20, N. Y. This catalog contains the U. S. Office of Education visual aids. They are particularly valuable for job training.
Films on Chemical Subjects-1946, The American Chemical Society, Washington 6, D. C. This catalog lists about 400 titles with a brief description of each. Information as to price and location is given. It is difficult to schedule films in an assignment guide because a particular film cannot always be secured on a definite date. The following films, however, are recommended for this unit.

Chemistry, Wonder World of: Role of chemistry in industrial processes and in production of consumer goods. 16 mm sound, p. 19.

Chemical Reactions: atomic structure of the elements. 16 mm sound, p. 27.

Solids: Elementary explanation of atoms and ions. 16 mm sound, p. 29.

Battle of Brains: The scientific approach to modern warfare. 16 mm sound, p. 30.

Caravan: Scientific progress presented on tour of streamlined exhibit trucks. 16 mm sound-color, p. 30.

Frontiers of the Future: Shows synthetic rubber, wool, perfume. 16 mm sound, p. 30
CHAPTER III
THE ORGANIZATION OF THE UNIT ON WATER AND SOLUTIONS

General Statement of the Unit
Water is one of the most important chemical compounds in the world. From the standpoint of health, we should understand the importance of pure water and how it can be purified; from the standpoint of our interest in chemistry, we want to know the composition of water and its chemical and physical properties. In order to understand the properties of water we must learn something about the molecular theory. A consideration of water leads us directly into the study of solutions, since water is the best and most common solvent. We shall speak briefly, however, of solvents other than water, particularly organic solvents.

Itemized Statement of the Unit
1. Water is the most important chemical compound in the world.
2. Water is essential to life.
3. Water can be separated from dissolved impurities by distillation.
4. Water can be purified for drinking purposes by settling, filtration, and chemical treatment. Small amounts of water can be rendered safe for drinking by boiling.
5. Water can be separated from dissolved impurities by distillation.

6. In the process of distillation, water is first boiled to change it to steam and then the steam is cooled to change it back to water.

7. The composition of water may be shown either by analysis or synthesis and either by weight or by volume.

8. It can be shown by electrolysis that water is composed by two volumes of hydrogen to one volume of oxygen. (Analysis)

9. It can be shown by mixing various volumes of oxygen and hydrogen in a eudiometer and exploding the mixture that water is formed by the combination of two volumes of hydrogen with one volume of oxygen. (Synthesis)

10. By weighing the hydrogen and oxygen produced by the electrolysis of water, it can be shown that water is composed of 11.2% hydrogen and 88.8% oxygen by weight. (Analysis)

11. The composition of water by weight can be shown by passing dry hydrogen over hot copper oxide and absorbing the water formed by calcium chloride. The loss of weight of the copper oxide equals the weight of oxygen used; the gain in weight of the calcium chloride equals the weight of water formed. (Synthesis)

12. These experiments lead to the formula H₂O for water.
The text on this page is not legible due to the quality of the image. It appears to be a page from a book or a document, but the content cannot be accurately transcribed.
13. The atoms in water are not arranged (as we would expect from our study of iron sulfide) in large groups containing millions of hydrogen and oxygen atoms in the proportions of two atoms of hydrogen to one of oxygen, but are arranged in small groups each containing two atoms of hydrogen and one atom of oxygen. Such a small group is called a molecule.

14. A molecule is the smallest particle of a substance which can exist alone.

15. The atoms in all substances in the gaseous state are grouped in molecules, but the atoms in some substances in the solid and liquid state are arranged in space in such a manner that no grouping into molecules is evident.

16. The molecules of gases are moving in all directions with great velocity. The molecules collide with the walls of the container and with each other to produce pressure. The temperature of the gas depends on the kinetic energy of the moving molecules. This picture of a gas is known as the "kinetic theory of gases".

17. One gas can mix with another because of the motion of the molecules. The process is called diffusion. Diffusion is also possible between the molecules of liquids and even solids.

18. There are attractive forces between the molecules of a substance which tend to hold them together. At high temperatures the motion of the molecules is so great that
For the human to wear the top hat and
to wear spectacles. He must 
and
under the circumstances of the situation, and
be able to do so. One of the keys is the
strategy and planning of the entire
and

moreover.

In the utterance of the smallest part of

surprising effort can come from

the same time, the principal points and the phases in some

several more thoughts to ponder, and the paths in some

consequences of the policy and action: The result to

bequeath to future generations. The expectation of the

to the gradual change of the working of the

ploration of a day to know at the "mystery which to bees.

If we can have additional support based on the

worker of the world, that process in a way different

different in what form to pursue the profession of

In this case, there are further reasons for the operation to

and therefore we have followed

To a microscope which leads to parts that are
certainties and the notion of the movement to do keep such
the attractive forces are not strong enough to confine the molecules within a definite volume and the substance exists as a liquid. At still lower temperatures the forces can not only hold the molecules within a definite volume but also within a definite shape and the substance exists as a solid.

19. Increasing the pressure will cause the molecules to come closer together; decreasing the pressure will permit the molecules to move farther apart because of their own motion.

20. Molecules may consist of only one atom (mercury) or two atoms (oxygen and hydrogen) or even more. Molecules of phosphorus contain four atoms each; molecules of sulfur contain six atoms each.

21. A symbol represents an atom of an element; a formula represents either a molecule of an element or compound or the proportions in which the atoms are combined in a compound.

22. The symbols for mercury, oxygen, hydrogen, phosphorus, and sulfur are Hg, O, H, P. and S; the formulas for these elements as gases are Hg, O₂, H₂, P₄, and S₆.

23. Formulas must be used in writing equations.
24. All equations must be balanced.
25. Water is a colorless, odorless, tasteless liquid. It weighs 1 g/ml. It freezes at 0°C and boils at 100°C. One calorie of heat is required to change the temperature of 1 gram of water one Centigrade degree.
26. Water is a very stable compound. It is only slightly decomposed by heating.

27. Chemical reactions which proceed in either direction according to conditions (like the combination of hydrogen and oxygen to form water and the decomposition of water to form hydrogen and oxygen) are called reversible reactions. If the reaction is carried on under such conditions that the products can not escape, a state of equilibrium is soon reached.

28. Water reacts with the active metals potassium, sodium, and calcium to form hydrogen and the hydroxide of the metal.

29. The group OH is called the hydroxyl radical. A radical is an atom or a group of atoms which acts as a unit in a chemical change. The hydroxyl radical will turn red litmus paper blue and phenolphthalein red.

30. Substances which contain the hydroxyl radical and which turn litmus paper blue are called bases. They are said to have a basic or alkaline reaction. A base can be thought of as being formed from water by the substitution of a metal for one of the hydrogen atoms in the water molecule. In order to visualize this substitution, the formula of water is considered as H-OH.

31. Acids are the chemical opposites of bases. They contain hydrogen and turn litmus red. An acid can be thought of as being formed from water by the substitution of a non-metal or a non-metallic radical (such as Cl, S04, or NO3) for the hydroxyl group of H-OH.
32. A salt is the compound of a metal with a non-metal or with a non-metallic radical.

33. Some metals not active enough to react with cold water will displace hydrogen from hot water or steam to form hydroxides or, in some cases, oxides. (Mg, Al, Zn, Fe)

34. Water reacts with the oxides of metals to form bases and with the oxides of non-metals to form acids.

35. Water combines with many substances directly to form hydrates. The water in the compound is called water of hydration or water of crystallization. A compound from which water of crystallization has been removed by heating is said to be anhydrous. The process by which a hydrate loses water of crystallization at room temperature is called efflorescence.

36. A substance which will absorb water from the air is hygroscopic; if it absorbs enough to become wet it is deliquescent. A substance used to absorb water is called a drying agent or a desiccant.

37. A solution is a uniform mixture. It does not follow the law of definite composition. One component is called the solvent; the other, the solute.

38. Water is the most common solvent. Many organic compounds (compounds once thought to be produced only by living organisms and found in plants and animals) are insoluble in water but are soluble in organic solvents like alcohol, ether, acetone, carbon tetrachloride, or
As a result, the proposal for the selection of the location of the site for the construction of the new building was made. The site was chosen based on the following criteria:

1. Accessibility: The site should be easily accessible by road and public transport.
2. Aesthetics: The site should complement the surrounding landscape.
3. Cost: The site should be cost-effective without compromising on quality.
4. Legal: The site should comply with all legal and regulatory requirements.

A report was submitted to the committee for their approval. The committee was unanimous in their decision to proceed with the construction plan. The construction team was put in charge of overseeing the project to ensure its successful completion.

The construction process would take approximately 18 months. During this period, the team would work closely with the local community to minimize any disruption caused by the construction works.
benzine.

39. A dilute solution contains a relatively small amount of the solute; a concentrated solution contains a relatively large amount. A saturated solution contains as much of the solute as the solvent will dissolve at that temperature; an unsaturated solution contains less of the solute than the solvent can dissolve at that temperature; and a supersaturated solution contains more of the solute than the solvent can dissolve at that temperature. (The extra amount of solute must be dissolved at a higher temperature and the solution cooled.)

40. Concentration may be expressed either as weight percent or as volume percent, as grams per liter or grams per milliliter, or as moles per liter. A solution containing 1 mole per liter is called a molar solution.

41. The solubility of a substance is the number of grams which will dissolve in 100 grams of the solvent. Solubility varies with temperature.

42. The solubility of a slightly soluble gas is directly proportional to the pressure. The solubility of a gas decreases with temperature.

43. When a substance goes into solution, its molecules separate and mix with those of the solvent. The structure of a solution is therefore something like the structure of a gas.
The method of operation is to heat the gas and the liquids with the electric current, thus raising the temperature of the gas and the liquids to the desired point. The gases and liquids are then introduced into the reaction chamber, where they are heated by the electric current and the heat of the reaction. This process continues until the desired temperature is reached. The gases and liquids are then discharged from the chamber, where they cool and condense. The condensed products are then collected and analyzed to determine the composition of the reaction products.
44. Diffusion occurs in solutions just as in gases.

45. Molecules of a solute exert a pressure known as osmotic pressure. By the process of osmosis, the water solutions pass from cell to cell within our bodies or within the bodies of plants. Osmosis always proceeds from the more dilute to the more concentrated solution.

46. Pure water freezes at O°C. The freezing point will be lowered if any material is dissolved in the water and the depression of the freezing point will be directly proportional to the concentration of the dissolved substance.

List of Materials and Readings for Teachers' Use


All information in consideration, need for in present
At reception of a source, present a sequence known as
conception sequence, where the process of an order of the source
so that the source of present, contains within sequence
You may have influence to the source coordination follow
For the source sequence be use, the process of the source
will be followed if and present. In source, the source
will be registered at the source, beyond will be included
An important role of the coordination of the preceding who-

Your Office: Secretary of the Interior, for Teacher, use
The Unit Assignment

Books required by the study guide for students' use

Raymond E. Brownlee, Robert W. Fuller, William J. Hancock, and Jessie E. Whitsit, Chemistry in Use. Allyn and Bacon. New York, 1939.  (B)

William Evans Price, George Howard Bruce, Chemistry in Human Affairs. World Book Company, New York, 1946 (P & B)


Bernard Jaffe, New World of Chemistry. Silver Burdett Company. New York, 1942. (J)

Study Guide for Unit II

1. Discussion by the teacher and the class on the importance of water and on the importance of pure water.

On the lecture desk are samples of food, some of which are thought of as containing water (milk, vinegar, and fruits) and others which are usually thought to be pretty dry (crackers, bread). Two tall cylinders contain muddy water. Flasks contain water variously colored. There is a flask of good drinking water with a faint color from vegetation or iron rust and a flask of clear water containing some harmless bacteria. A Liebig distilling apparatus should be set up and ready for immediate use.

(1) Why is water such an important substance? (2)

Name some of the places water occurs. (Seas, rivers, milk, fruit, vegetables, medicines, animals). (3) Which of the water samples on the desk would make the best drinking water?

1/ Items marked with an asterisk (*) are on the students' guide. Parts of the item enclosed in parentheses, however, are for the teacher's guidance and do not appear on the students' guide.
(If possible project the microscopic appearance of the clear water and the bacteria containing water; if micro-
projection apparatus is not available have some members of the class view this water through the microscope. It is obvious that appearance is not a safe criterion.) (4) If you were camping by a small stream, how would you make sure that the water in the stream was fit to drink? (Rapid flow, no chance of pollution above.) If you were camping on the shore of a lake, what precautions should you take before drinking the water? (Boil.) (5) Wells in the country are often situated so near the barns that they become polluted by the drainage. Why is the family often able to drink this water without harm, although strangers who drink it may become ill? (Developed immunity.) How should the water be treated before drinking? (6) What are some of the diseases which may be caused by polluted water? What precautions should be taken by a city in selecting its water supplies? (7) What are some of the ways city water supplies are purified? (Call attention to the fact that the muddy water in the cylinders on the lecture desk is clearing as the mud and clay settle to the bottom. To one cylinder add a few drops of a solution of alum. Point out that the addition of some chemicals will assist the settling process. Point out that chemicals such as chlorine might be added to kill harmful bacteria. Water can also be purified by filtration. Pour some muddy water through a large filter supported above a cylinder. Let the filtration continue
without attention and later point out that the filtrate is clearer than the original solution. Pour some of the colored water into the distillation flask and distill over a little pure water. Point out that (a) the water is turned into steam; (b) the steam is condensed back to water. Define distilling flask, condenser, receiver, distillate, distilled water. Ask for uses of distilled water.) (8) Why is distilled water used in storage batteries?

2. Pretest.

#3. Read B pp. 488-504. What did you learn from reading this chapter that was not brought out in the class discussion?

4. Continue discussion begun in Item 1. Ask the students if they have anything to add to the class information about pure drinking water. Point out that settling, filtration, and chemical treatment do not affect dissolved materials in water. Distillation does secure pure water in the chemical sense. Demonstrate use of small student "set-up" using an air condenser if Liebig condensers are not available. Demonstrate bending of glass.

#5. How does filtration purify water? (Experiment 2-1)

#6. How does distillation purify water? (Experiment 2-2) What is the difference between a volatile and a non-volatile substance?
7. Demonstration: The electrolysis of water. Show that water is decomposed by the passage of the electric current and that two volumes of hydrogen are formed to one volume of oxygen. Point out that this can be called analysis because we have taken water apart to see what it was made of. Note that it is an analysis by volume, but that if we weighed the gases we could calculate the percentage composition by weight. As a class exercise, calculate the percentage composition of water, given that the densities are 0.09 g/l for hydrogen and 1.43 g/l for oxygen. Assume that 2 liters of hydrogen and 1 liter of oxygen are collected. Note that water is a compound because it can be decomposed into simpler substances, but that hydrogen and oxygen are elements.


9. Demonstration: The composition of water by weight. Pass hydrogen over hot copper oxide and absorb the water formed by calcium chloride. Student assistants, previously rehearsed, should be used in this demonstration. The purpose of these three demonstrations is, of course, to bring out the

The theoretical framework of ecology of water plants

The interaction of the ecological system of plants and the environment is a complex of many factors. The plants interact with the environment through various processes, including photosynthesis, respiration, transpiration, and nutrient uptake. These processes are influenced by factors such as light, temperature, water availability, and soil nutrients. The ecological framework of water plants is an important aspect of understanding the structure and function of aquatic ecosystems.
facts about the composition of water and also to illustrate experimental techniques. One of the main purposes of a chemistry course is to give the pupils experiences in chemistry.

10. In what four ways can the composition of water be shown? (B 39-40; P & B 73-75; B & C 67-72; J 66-70.)

11. From the data obtained in the experiments on the composition of water, calculate the formula of water.

12. Talk by the teacher on the molecular theory. Refer to material previously learned in general science and physics. Most of this material merely needs to be brought to the surface of consciousness and given the proper orientation. Recall that copper sulfide was formed by the combination of atoms of copper and sulfur in certain definite proportions and that one piece of the compound contained millions of atoms. The formula Cu₂S represents the proportions. There are no small pieces containing just two copper atoms and one sulfur atom each. Water is different. There are small particles containing two hydrogen atoms and one oxygen atom each. These are molecules. A molecule is defined as the smallest particle of matter which can exist alone.

All matter in the gaseous state is composed of molecules. Molecules move very fast, they exert pressure, they have energy of motion which is a measure of temperature. Refer to the kinetic theory of gases, Boyle's Law, and Charles's Law if these are in the students experiences. A gas has no definite volume. Discuss
diffusion of gases using perfumes and cooking odors as examples. Describe attractive forces between molecules. Molecules in a gas move so fast that the attractive forces cannot hold a gas to any size or shape. If the pressure is increased, molecules may be crowded closer together. If the temperature is decreased, the molecules slow down and the attractive forces may be able to hold them in a definite volume (liquid state). If the temperature is low enough, an increase in pressure may suffice to liquify a gas. If the temperature is lowered still further, the attractive forces may be able to hold the molecules in a definite shape and a definite volume (solid state).

In the solid state, the molecules may be so arranged that it is no longer possible to identify individual molecules. Use NaCl as an example. Sodium chloride can be melted and vaporized. Molecules of NaCl exist in the vapor state. In the solid state atoms are so regularly spaced that (diagram) no one Na atom goes with any particular Cl atom. Some substances, like mercury oxide, do not vaporize into HgO molecules, but decompose into their elements instead.

A molecule of an elementary gas may consist of one or more atoms: Hg, O₂, H₂, P₄, S₆. A symbol represents an atom; a formula represents either the proportions in which the atoms are combined or the molecule of an element or compound. Formulas are used to write equations; for example, the decomposition of water by electrolysis is represented by

\[ 2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2 \]
Equations must be balanced. Show how we check an equation by counting atoms or by adding up weights.

#13. Describe how the molecules of water behave when water is heated (P&B 46-54). How do you explain evaporation in terms of the molecular theory?

#14. List some of the physical properties of water: color, odor, taste, density, freezing point, boiling point, and specific heat.

15. Talk by the teacher pointing out that the stability of water is very great compared to that of mercury oxide. Show that the decomposition of water into hydrogen and oxygen is a reversible reaction (written with double arrows). Develop the idea of equilibrium.

#16. Describe the reaction of water with the active metals potassium, sodium, and calcium. (Experiment 2-3)

#17. What is a base? (B 53-57, 269; B&C 48, 126-127; J 54) What is an alkali? (P&B 249) What is a radical? (B 74-75; B&C 96) What are indicators? (B 274; B&C 150)

18. Talk by teacher, summarizing the material in Items 29-32 in the itemized statement of the unit, pp 32-33 above. Note that water is H-OH; a base is metal + OH; and an acid is H + non-metal; a salt is metal + non-metal.

#19. In what respects are acids the chemical opposites of bases? (B 57, 245)

20. Demonstration: Action of steam on iron. Some
metals are not active enough to displace hydrogen from cold water, but will displace hydrogen from hot water or steam. Some metals will displace half the hydrogen to form the hydroxide of the metal (Mg and Zn); others will displace all the hydrogen to form the oxide of the metal (Al and Fe). Note that aluminum does not react in cooking vessels because the layer of Al₂O₃ forms a protective coating.

Support an iron pipe about 1 inch in diameter and about 20 inches long horizontally between two ring stands. Fit a 1-hole rubber stopper in each end. Pack it loosely with scrap iron, nails, shavings from the machine shop, or iron filings. Boil water in a 500 ml flask and pass the steam through the pipe. Heat the pile in the middle with one or two Bunsen burners fitted with wing tops. After a few minutes light the hydrogen issuing from the tube. Indicate that this is a commercial method of manufacturing hydrogen.

\[ 3 \text{Fe} + 4 \text{H}_2\text{O} \rightarrow \text{Fe}_3\text{O}_4 + 4 \text{H}_2 \]

21. Describe the action of water on oxides. How does this experiment give us a basis for distinguishing between metals and non-metals? (Experiment 2-4; B 63)

22. Demonstration: Show crystals of copper sulfate and washing soda. Show that if they are heated in a test tube, water is given off and collects at the mouth of the tube. Why should the mouth of the tube not be
heated? Define hydrate, anhydrous, water of crystallization, water of hydration, dehydration. Place a small piece of sodium hydroxide on a watch glass to show that it absorbs water from the air. Show that zinc chloride, magnesium chloride, and calcium chloride also absorb water from the air. Define hygroscopic, deliquescent. Show that crystals of washing soda in the box have lost their glossy appearance because of loss of water. Define efflorescence. Point out the use of calcium chloride to absorb water in the experiment on the composition of water. Show desiccator standing on desk. Explain use of desiccants.

#23. What are some of the properties of hydrates? (Experiment 2-5)

#24. Explain (not merely define) the meaning of the following terms. Use illustrations freely to make your meaning clear. (a) hydrate; (b) water of crystallization; (c) crystallization; (d) dehydrated; (e) anhydrous; (f) efflorescence; (g) hygroscopic; (h) deliquescent. (B&C 374-375; J 74-78)

#25. Describe how you would test a crystalline substance to see if it were a hydrate.

#26. Why do manufacturers sometimes label a box in which crystals of washing soda are sold "2 1/2 pounds net when packed"?

#27. What are solutions? Why are they important? In what way are solutions and compounds alike? In what way are they different? (B 116-118; B&C 63-64; P&B 76-79: J 237)
45

#28. If you wished to dissolve 50 grams of copper sulfate in a liter of water as quickly as possible, what method or methods would you use? (Experiment 2-8.)

#29. You wish to dissolve some crystals of potassium dichromate, K₂Cr₂O₇, in water. Mention five things you could do to hasten solution.

#30. Why are bottles of carbonated beverages kept cold and tightly stoppered? Few, if any, bubbles can be observed in a bottle of ginger ale. How then do you explain the many bubbles which rise from the same ginger ale when the bottle is opened and the liquid is poured into a glass? (B 130-132; E&C 378.)

#31. What is a solvent. Name five. What is a solute? Name one for each solvent listed. (A different one for each.) (Experiment 2-9.)

#32. What is a dilute solution? a concentrated solution? A saturated solution? (B 122-126; E&C 373; P&B 78-79; J 237-239.)

#33. What is a supersaturated solution? (Experiment 2-7.)

#34. Distinguish between a dilute, a concentrated, and a supersaturated solution of potassium nitrate. If you were handed a bottle containing a solution of sodium acetate and asked to determine whether the solution were saturated, unsaturated, or supersaturated, how would you proceed?
If you want to discuss or analyze the content of a document, please provide the document text or a description of the content. Without the actual document text, it's challenging to provide a meaningful response.
35. Talk by the teacher explaining how concentration is expressed by weight per cent or volume percent. Illustrate by bottles from the kitchen or the medicine cabinet (iodine, mercuriochrome, vanilla extract, and vanilla). In the laboratory we often use molar solutions. A molar solution contains the molecular weight of the solute in a liter of solution. The solubility of a substance is defined as the number of grams of the substance which will dissolve in 100 grams of water at a definite temperature.

36. Demonstration: diffusion. Take two tall cylinders (250-500 ml graduates). Fill one with water. Into the graduate containing water drop a crystal of potassium permanganate or a pinch of some water soluble dye. Into the graduate containing air, introduce a few drops of liquid bromine. Talk about the motion of the molecules. Point out that we have solutions in both cases (solid in liquid; liquid in gas). Both the dye and the bromine are separated into small particles and mix by diffusion.

37. Demonstration: osmosis. Bore a hole in the top of a carrot with a cork borer. Fill the cavity with molasses or sugar. Insert a tight-fitting glass tube. Seal the tube to the carrot with colodion. Support the carrot and tube in a large beaker of water. Water passes through the walls of the vegetable membrane to dilute the solution. The increase in volume causes water to rise in the tube.
The text on this page appears to be a continuation of the previous page, discussing the importance of cooperation and coordination in the context of military strategy and operations. The passage mentions the importance of liaison and communication, highlighting the need for effective coordination among different units and agencies to ensure strategic alignment and operational success. The text suggests that cooperation is essential in achieving collective objectives and that effective communication is key to maintaining coordination in complex operations.
This column of water produces a pressure equal to the "osmotic pressure". The cell walls of living plants and animals are semi-permeable membranes which regulate the passage of the various solutions in the body, always from the more dilute to the more concentrated. By osmosis, the cell sap passes from cell to cell in a plant or tree. The sap does not become so dilute as to stop the process because plants give off water by transpiration; a maple tree 25 feet high may transpire 100 pounds of water per day.

#38. What is the effect of dissolved substances on the freezing point of a solution? Why do rivers frequently freeze over in winter although harbors seldom do? What is an anti-freeze solution? Why do you put alcohol or Prestone in your automobile radiator in cold weather? (B&G 376-377)

Optional Related Activities

1. Make a visit to your city water purification system and report to the class. If possible, secure pictures or diagrams of the interesting features.

2. Talk with a public health officer about pure drinking water. What precautions are observed. How is the water tested. If possible, secure a report giving facts and figures.

3. Prepare, with the assistance of the teacher, a demonstration on the determination of the composition of water by weight. (B&G 70) The method consists in passing dry hydrogen over hot copper oxide and weighing
The purpose of this manual is to provide a comprehensive guide to the

concepts discussed in the document. The manual is designed to help

students and practitioners understand and apply the principles outlined

in the document. It includes examples, exercises, and additional

resources to aid in the learning process. The manual is structured to

facilitate a smooth transition from basic concepts to advanced

topics. It is intended for use in academic settings or self-study.

The manual aims to be a valuable resource for those seeking to

 deepen their understanding of the subject matter.
the water formed. The loss of weight of the copper oxide equals the weight of the oxygen formed.

4. Write out answers to the following questions:
   (a) Why is rain the purest form of natural water?
   (b) Can water be purified by freezing? Why?
   (c) How is it that water extinguishes fire when it is composed of hydrogen which burns and oxygen which supports combustion?

5. Solve the following problems:
   (a) A mixture of 30 ml of oxygen and 20 ml of hydrogen is exploded in a eudiometer. What volume of water vapor would be formed? What gas and how much of it would remain in excess?
   (b) A mixture of 7 ml of oxygen and 20 ml of hydrogen is exploded in a eudiometer. What volume of water vapor would be formed? What gas and how much of it remains in excess?
   (c) In attempting to determine the percentage composition of water by weight, a student obtained the following data:

   **Weight of tube and copper oxide before experiment**
   15.766 g

   **Weight of tube and copper oxide after experiment**
   14.836 g

   **Weight of calcium chloride tube before experiment**
   21.335 g

   **Weight of calcium chloride tube after experiment**
The purpose of this report is to provide an overview of the company's financial performance and operational results for the past fiscal year. The following sections detail the company's revenue, expenses, and profit margins, along with a breakdown of sales by product line. The report also includes an analysis of the company's market share and a forecast of future growth. The conclusion of the report outlines steps for improving the company's profitability and efficiency.
What results should he have obtained by calculation from these data? (B&C 68-71; J 69-70)

6. Describe an experiment to determine the composition of water by separating it into its constituent elements by means of the electric current. Name the process. Draw a diagram of a suitable apparatus or photograph the apparatus used in class. Write an equation for the reaction.

7. Describe an experiment to determine by synthesis the composition of water by volume. Draw a diagram of the apparatus or photograph the apparatus used in class. Write an equation for the reaction.

8. Describe an experiment to determine by synthesis the composition of water by weight. Draw or photograph the apparatus and write an equation for the reaction.

9. Solve the following problems:

   (a) A current of steam was passed over red hot iron until the iron oxide formed weighed 5 grams more than the iron had weighed at the beginning of the experiment. Write the equation for the reaction. What weight of hydrogen was liberated?

   (b) A current of hydrogen was passed over heated copper oxide. If the copper oxide lost 404.624 grams and 454.862 grams of water were formed, find the weight of oxygen which combines with one gram of hydrogen. Calculate the percentage composition of water by weight.
How does the composition of water illustrate the Law of Definite Proportions?

10. Solve the following problems:

(a) It can be shown by analysis that in mercurous chloride one gram of mercury is combined with 0.1767 grams of chlorine and that in mercuric chloride, one gram of chlorine is combined with 2.8285 grams of mercury. Show that these two compounds illustrate the Law of Multiple Proportions. From the data given, calculate the formulas of these two compounds.

(b) Two oxides of hydrogen are known. In one, two parts by weight of hydrogen are combined with 16 parts by weight of oxygen; in the other, two parts by weight of hydrogen are combined with 32 parts by weight of oxygen. The first compound is water; the second, hydrogen peroxide. How do these two compounds illustrate the Law of Multiple Proportions? From the data given, calculate the formulas of these two compounds.

11. Solve the following problems:

(a) Hydrogen peroxide can be prepared by the action of sulfuric acid \((H_2SO_4)\) on barium peroxide \((BaO_2)\). Barium sulfate \((BaSO_4)\) and hydrogen peroxide are formed. Write and balance the equation for the reaction. What weight of hydrogen peroxide could be made from 8.45 g of barium peroxide?

(b) Hydrogen peroxide is somewhat unstable, decomposing into water and oxygen. Write and balance an equation for the decomposition. What weight of oxygen could
to the following information:

(a) Do not to study or research life in

(b) Do not to study or research life in

(c) Do not to study or research life in

(d) Do not to study or research life in

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(x) Do not to study or research life in

(y) Do not to study or research life in

(z) Do not to study or research life in

References

[1] Author, Title, Year, Publisher.

[2] Author, Title, Year, Publisher.

[3] Author, Title, Year, Publisher.
be obtained from 4 grams of hydrogen peroxide? If the
density of oxygen is 1.43 g/l, what is the volume of the
goxygen formed?

12. Answer in writing:

(a) What are some of the physical properties
of water?

(b) What evidence indicates that water is a
stable substance? Explain the meaning of the terms
"reversible reaction" and "equilibrium".

(c) Name three metals which will react with
water to form hydrogen and hydroxides of the metals.
Write an equation for each reaction.

(d) Write an equation between water and a
metal to form an oxide of the metal and hydrogen.

(e) Name three metallic oxides which react
with water to form basic oxides. Write an equation for
the reaction in each case. Will all metallic oxides react
with water? Explain.

(f) Name two oxides of non-metals which will
react with water to form acids. Write an equation for the
reaction in each case.

(g) Explain the meaning of: acidic oxide,
basic oxide, acid anhydride, basic anhydride. Give
illustrations of each.

13. Solve the following problems:

(a) What percent of washing soda is sodium
carbonate? What percent is water? What is the weight
of sodium carbonate in 23/4 pounds of washing soda?
(b) For a certain experiment 34 grams of anhydrous copper sulfate are required. What weight of crystallized copper sulfate should be used in order to have this weight?

(c) What is the percent of water in alum, \( \text{K}_2\text{Al}_2(\text{SO}_4)_4 \cdot 24 \text{H}_2\text{O} \)?

(d) Plaster of Paris, \((\text{CaSO}_4)_{2} \cdot \text{H}_2\text{O}\) is made by strongly heating gypsum \( \text{CaSO}_4 \cdot 2 \text{H}_2\text{O} \). What is the percent of water in each of these substances?

14. What is the percent of water in crystallized barium chloride? (Experiment 2-6)

15. Explain, with examples, the meaning of "miscible" and "immiscible". (Experiment 2-10)

16. How can you determine the solubility of a slightly soluble substance? (Experiment 2-11; Experiment 2-12)

17. It is convenient to represent the relation between solubility and temperature by a graph. If we plot the temperature as abscissae (x-axis) against the solubilities as ordinates (y-axis), we obtain a curve known as the solubility curve. The following table gives data on the solubility of (1) boric acid, (2) sodium chloride, (3) ammonium chloride, (4) calcium chloride, (5) ether, and (6) potassium nitrate. Plot these data on graph paper to obtain solubility curves.
18. Answer the following questions:
(a) How is solubility defined? What does a solubility curve express? How is such a curve constructed?
(b) Are all substances more soluble in hot water than in cold water? Mention the solubility curve in your answer. (E&C 373-37.)

19. Answer the following questions:
(a) What is an organic compound? What is an organic solvent?
(b) Look up the difference between a tincture and spirits as used in pharmacy. Give an example of each.

20. What is meant by each of the following statements:
(a) Chalk is soluble in water.
(b) A solution is homogeneous.
(c) A mixture of ice and salt will freeze ice cream. (E&C 373-76)

21. How would you purify commercial Chile saltpeter? This substance is sodium nitrate, NaNO₃. It is found mixed
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For what is following statement true (a)?

The statement is true when (b) you are making above certain assumptions (c) or if will be used for more complex tasks (d).

What will your following conclusion be (e)?

(a) Clearly, conclusion is incorrect (b) so what is correct to choose (c)?

Number of the corresponding parameter (d) is shown above.

(b) Clearly, conclusion is incorrect (c) so what is correct to choose (d)?

Number of the corresponding parameter (e) is shown above.
with dirt and rock in Chile. It is soluble in water.

22. What is the difference between a 5% solution of silver nitrate and a solution of silver nitrate containing 50 grams per liter? How many moles per liter are in a solution of sodium hydroxide containing 20 grams of NaOH in 100 ml of the solution? How many grams of each of the following substances would you weigh out to make a 0.2 M solution of: CuSO₄, NaCl, AgNO₃, NaH₂PO₄, K₄Fe(CN)₆, KMnO₄?

23. Divide the common gases into three classes according to their solubility in water. (E&C 378-379). State and explain two conditions which affect the solubility of gases in water. When water is heated, bubbles form on the sides and bottom of the vessel containing it and slowly rise through the water to the surface at a temperature considerably below the boiling point. Explain this fact.

24. How do solutions resemble gases? Who developed this theory? How would you demonstrate that diffusion occurs both in solutions and in gases? Explain osmosis. Why is this phenomenon of such great importance? (E&C 378-381)

25. Take a series of pictures of your water purification system. Also try to get pictures of water supply systems for farms, such as getting water from a spring, the water bucket and dipper, etc. Get pictures showing the barns and the well properly situated and also try to find some "horrible examples".

26. Make a series of drawings to illustrate your
laboratory work, the demonstrations you have witnessed, or some of the concepts of this unit. Perhaps you can prepare some interesting posters to hang on the wall.
CHAPTER IV

CLASSROOM PRESENTATION AND EVALUATION OF THE TWO UNITS

Administration of the Unit

Method of Instruction:-- Each assignment was discussed with the class by means of board work and experiment and in much greater detail than would be necessary for a college preparatory class. The boys did not do individual laboratory work. Each experiment required in the guide sheet was performed as a demonstration by one of the boys under the direct supervision of the instructor. The boy who did the experiment explained it to the class. It was usually necessary for the teacher to supplement his explanation by questions and further explanation. The other boys in the class took notes. At the end of the unit study, each boy passed in a laboratory notebook in which each experiment was written up under the headings: Title, Apparatus and Material, Purpose, Procedure, and Results. These notebooks were also examined once or twice during the unit study.

Each boy was also expected to keep a looseleaf notebook in which he wrote down notes on the class discussions and the answers to questions on the guide sheets.

Two short tests were administered during each unit and an objective test on each unit was given first as a pretest and again as a final test.
Daily Log of Unit I

Date  Day
1946
Sept. 26 1  The instructor introduced the unit with a talk and experiments (Item 1*). The class appeared to be much interested. Many questions were asked about the "volcano" experiment and the color changes resulting from the mixed liquids. The questions were, however, rather superficial and indicated more interest in the spectacular nature of the experiment than in the fundamental chemistry of the phenomena. Item 3 was assigned for outside study.

Sept. 27 2  The pretest was administered during the first half of the period. The class at first objected to a test, but as soon as they learned the nature and purpose of a pretest, they cooperated. The latter part of the period was devoted to demonstrations to illustrate the nature of a chemical change (Item 4).

Sept. 30 3  The demonstration experiments (Items 5 and 6) were performed by two of the boys. Although each boy had been given previous help in the presentation of the experiment, he still had a tendency to slight many of the important

*The items are those on the Study Guide of the Unit Assignment.
chemical concepts. It was necessary for the instructor to interrupt frequently to call attention to certain points which otherwise would have been missed. The other boys in the class took notes and asked questions; but, on the whole, they were not very inquisitive. The instructor gave directions for writing up the experiments and assigned this activity for outside work.

Oct. 1 4 Items 10 and 11 were assigned and discussed in great detail. The metric system was described briefly and attention was called to the supplementary material available. The Spanish boys were helpful to the class during this discussion because they were very familiar with the metric units.

Oct. 2 5 Since the metric system is much used in chemistry, this period was devoted to exercises involving the metric units. The boys found these problems very difficult, partly because they were not familiar with the units and partly because they were poorly prepared in arithmetic.

Oct. 3 6 The experiments of Items 7 and 8 were demonstrated. The boys did better than the boys who demonstrated the previous experiments, but they were still weak in explanations.
The experiment in Item 9 was demonstrated. Two boys were assigned to this experiment: one to do the actual manipulation and one to do the explaining. This procedure worked very well. During the latter part of the period, the class completed the written descriptions of the first five experiments and passed in their notebooks for inspection. Items 12 and 13 were assigned for outside preparation.

Twenty minutes of the period were allowed for the first short test. Items 14 and 15 were assigned and discussed in class. The demonstration in Item 15 was performed. The instructor explained how it illustrated the Law of Conservation of Matter. Since the metric system was used in this experiment, the importance of it was stressed again.

Twenty minutes were allowed for the second short test. The rest of the period was devoted to the experiment of Item 18. Since this experiment requires careful weighing, it was done by the instructor. The periods of inactivity while the crucible was being heated and cooled were filled by a class recitation. The questions in the directions for the experiments and the questions on the guide sheet were asked. The class response
The invention in Test 9 and Experiments 10 and 11 are not necessarily due to the same factors, and may be due to the same or different factors. The invention in Experiment 6 is probably due to the same factor or factors. The experiments in Experiment 6 are probably due to the same factor or factors.
was not very good; the answers were not so much wrong, as incomplete and vague, indicating lack of study or improper methods of preparation. Some suggestions for improvement were given by the instructor. Items 16, 17, and 19 were assigned and the class directed to review the material covered so far.

Oct. 9 10 Experiment 1-6 (Item 18) was finished and the recitation begun October 8 was continued. The results were better but still left much to be desired. The experiment was written up outside.

Oct. 10 11 The instructor discussed the atomic theory - Item 20. Items 21 and 22 were assigned for outside preparation.

Oct. 11 12 Items 23, 24 and 25 were prepared in a class study period. The instructor went from student to student clearing up individual questions.

Oct. 14 13 The instructor discussed Items 26 and 27. The entire unit was assigned for review.

Oct. 15 14 The final test on Unit I was administered.
First short test - Unit I:--

1. Define element, mixture, and compound. Give an example of each.

2. What two things may accompany a chemical change?

3. (a) What is matter? (b) In what three states may it exist? (c) What are the characteristics of each state?

4. In the metric system, state the unit of length, the unit of mass, and the unit of volume.

Second short test - Unit I:--

1. How many inches in 20 cms?

2. How many centigrams in 52 milligrams?

3. What is the weight in pounds of a 10 kilogram weight?

4. What is the weight in grams of 6.6 pounds?

5. Compute your height in centimeters.

Pretest and Final Test - Unit I:--

Some of the following represent physical changes; some represent chemical changes. In the parenthesis at the left, write a P in front of the items which represent physical changes; write a C in front of the items which represent chemical changes.

( ) 1 mixing cement
( ) 2 making a chocolate nut sundae
( ) 3 burning of kerosene
( ) 4 the process by which light is emitted from an electric light bulb
( ) 5 painting a chair
( ) 6 preparing an alloy
The text on the page is not legible due to the quality of the image. It appears to be a page from a document, but the content cannot be accurately transcribed.
One of the letters in each group below is the initial letter of a word which is defined above the group. Write the correct initial letter in the blank at the left. Example

The unit weight in the English system of measures

(P) L M P S W

Since the unit of weight is the pound and the initial letter of pound is "P", write "P" in the blank at the left.

( ) 7 The system of weights and measures used in chemistry

(A) E M R W

( ) 8 The unit of length in this system

(C) G I K M

( ) 9 The unit of mass in this system

(K) L O M T

( ) 10 The unit of volume

(B) C L Q T

( ) 11 The unit of temperature

(C) D E G M

( ) 12 The unit of heat

(C) D E G M

( ) 13 The weight of a unit volume

(D) F L V W

( ) 14 A substance which can not be decomposed into anything simpler

(B) C E M P

( ) 15 Every chemical change either gives off or absorbs

(G) H L M O

( ) 16 The characteristics of a substance by which we recognize it

(B) C E M P

( ) 17 One kind of matter; all its parts are alike and have the same properties

(C) D E M S

( ) 18 A material composed of more than one substance

(C) E M S V

( ) 19 The smallest particle of an element

(A) B D M P

( ) 20 An element is represented by a (n)

(D) F S T V

( ) 21 A chemical reaction can be represented by a (n)

(D) E F G M

( ) 22 When a solid is separated from a liquid by pouring the mixture through paper fitted into a funnel, the liquid which runs through is called

(D) E F G R

( ) 23 The solid which remains on the paper

(D) F G R T

( ) 24 When a solution is allowed to evaporate so that the solvent is removed, a solid assumes a definite geometrical shape by the process of

(B) C D E M
Part II

Directions: All omitted words or phrases appear in the column at the left of the page. Show that you know which word or phrase has been omitted from each blank by putting the number of the blank in the proper parenthesis at the left. For example, the "1" is put in the parenthesis in front of the word omitted from the blank "1".

( ) black
( ) brown
( ) white
( ) orange
( ) purple
( ) green
( ) yellow
( ) greater than
( ) less than
( ) Just about the same as
( ) sulfuric acid
( ) nitric acid
( ) acetic acid
( ) hydrochloric acid
( ) glycerine
( ) sugar
( ) potassium permanganate
( ) ammonium dichromate
( ) sodium aluminate
( ) hydrogen
( ) Oxygen
( ) nitrogen
( ) zinc

Pure carbon is _1_ in color. In the volcano experiment a pile of a (an) _2_-colored substance known as _3_- was capped with a little _4_. The reaction was started with a drop of _5_. The substance formed was _6_ in color. Its bulk was _7_ the original material. When _8_ was added to a few pieces of copper in a beaker, _9_ colored fumes were evolved. Sugar can be changed to a black frothy mass by the addition of _10_.

( ) centimeter
( ) inch
( ) liter
( ) pound
( ) kilogram
( ) meter
2.2
2.54
2.7
39.37
36
0.0235
0.235
2.35
23.5
235
2350
( ) 0.00235

The meter equals _13_ inches. The inch equals _14_ centimeters. The kilogram equals _15_ pounds. The liter is the volume of one _16_ of water at 4°C. 2.35 meters equal _17_ cms equal _18_ mm. 2.35 grams equal _19_ Kg. 23.5 grams equal _20_ mg. 2.35 ml of water weigh _21_ grams. 23.5 grams of water have a volume of _22_ ml. 30°C equal _23_ °F; 608°F equal _24_ °C.

To raise the temperature of 6 grams of water from 27 to 67°C will require _25_ calories of heat. A block of wood is 40 cm long, 20 cm wide, and 10 cm thick. It weighs 6000 g. Its density is _26_ g/cm^3_. In an experiment to determine
the density of aluminum, a block of aluminum was found to weigh 132.4 g. It displaced 47.9 g of water. The density of aluminum was found to \( \text{27} \) g/cm\(^3\).

When powdered iron and sulfur were stirred together, they formed a (n) \( \text{28} \). The material could be separated by picking out the \( \text{29} \) with a (n) \( \text{30} \), or by dissolving the \( \text{31} \) in \( \text{32} \). The soluble material could then be separated from the insoluble material by the process of \( \text{33} \). If the liquid which runs through the paper is placed in a small porcelain dish, the solvent may be removed by the process of \( \text{34} \).

Mercuric oxide is a (n) \( \text{35} \) because it can be separated into simpler substances by \( \text{36} \); iron sulfide is a compound because it can be formed by the \( \text{37} \) of simpler substances. A substance like mercury from which simpler substances can not be obtained is called a (n) \( \text{38} \). The atomic theory assumes that (1) all \( \text{39} \) are composed of small particles called \( \text{40} \); (2) for a given element all the particles \( \text{41} \); (3) the particles combine to form \( \text{42} \); and (4) the particles \( \text{43} \).

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By experiment, 137 parts by weight of barium combine with 16 parts by weight of oxygen to form \( \text{Ba}_4\text{O}_8 \). The percent of barium in this compound is 45. If another compound of barium and oxygen were formed, it might be possible for 137 parts by weight of barium to combine with 46 parts by weight of oxygen. The standard of atomic is 0 equals _47_. If iron and sulfur are heated together, it would be possible for 70 grams of iron to react with 48 grams of sulfur. (Atomic weights--Fe-56, S = 32, Ba = 137)

If copper and sulfur are heated together, the compound formed has the formula \( \text{Cu}_2\text{S} \). In an experiment, 5.8481 g of silver chloride were made from 4.4018 g of silver by heating it with chlorine. One atom of chlorine combines with 50 atoms of silver. The formula \( \text{SnO}_2 \) indicates that the ratio of the number of atoms of tin to the number of atoms of oxygen in the compounds is as 1: 51.

The fact that exactly 63.57 g of copper will unite with exactly 32.06 g of sulfur to form 95.63 g of a compound with the formula \( \text{Cu}_2\text{S} \) is an example of the Law of _53_. The fact that this substance contains about 54 percent of sulfur is an illustration of the Law of _55_.

Atomic weights Cu = 63.6; S = 32.)
Some common symbols are: Chlorine \(_{56}\); Calcium \(_{57}\); carbon \(_{58}\); tin \(_{59}\); mercury \(_{60}\), and Iron \(_{61}\).

The atomic weight of zinc is 65 \(_{62}\). A weight of 130 grams of zinc is equivalent to two \(_{63}\) of zinc. Zinc combines with oxygen to form ZnO. 65 plus 16 equals 81, which is the \(_{64}\) of zinc oxide. A weight of 162 grams of zinc oxide would be two \(_{65}\) of this compound.
### Key to Test on Unit I

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
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<tbody>
<tr>
<td>1.</td>
<td>P</td>
<td>( )</td>
</tr>
<tr>
<td>2.</td>
<td>P</td>
<td>( )</td>
</tr>
<tr>
<td>3.</td>
<td>C</td>
<td>( )</td>
</tr>
<tr>
<td>4.</td>
<td>P</td>
<td>( )</td>
</tr>
<tr>
<td>5.</td>
<td>P</td>
<td>(16) kilogram</td>
</tr>
<tr>
<td>6.</td>
<td>P</td>
<td>( )</td>
</tr>
<tr>
<td>7.</td>
<td>M</td>
<td>(15) 2.2</td>
</tr>
<tr>
<td>8.</td>
<td>M</td>
<td>(14) 2.54</td>
</tr>
<tr>
<td>9.</td>
<td>K</td>
<td>( )</td>
</tr>
<tr>
<td>10.</td>
<td>L</td>
<td>(13) 39.37</td>
</tr>
<tr>
<td>11.</td>
<td>D</td>
<td>( )</td>
</tr>
<tr>
<td>12.</td>
<td>C</td>
<td>( )</td>
</tr>
<tr>
<td>13.</td>
<td>D</td>
<td>( )</td>
</tr>
<tr>
<td>14.</td>
<td>E</td>
<td>(21) 2.35</td>
</tr>
<tr>
<td>15.</td>
<td>H</td>
<td>(22) 23.5</td>
</tr>
<tr>
<td>16.</td>
<td>P</td>
<td>(17) 235</td>
</tr>
<tr>
<td>17.</td>
<td>S</td>
<td>(18) 2350</td>
</tr>
<tr>
<td>18.</td>
<td>M</td>
<td>(19) 0.00235</td>
</tr>
<tr>
<td>19.</td>
<td>A</td>
<td>(20) 23500</td>
</tr>
<tr>
<td>20.</td>
<td>S</td>
<td>( )</td>
</tr>
<tr>
<td>21.</td>
<td>E</td>
<td>( )</td>
</tr>
<tr>
<td>22.</td>
<td>F</td>
<td>(26) 0.75</td>
</tr>
<tr>
<td>23.</td>
<td>R</td>
<td>( )</td>
</tr>
<tr>
<td>24.</td>
<td>C</td>
<td>(27) 2.76</td>
</tr>
</tbody>
</table>

**1.** black

**9.** brown

**2.** orange

**20.** S

**21.** E

**22.** F

**23.** R

**24.** C

**3.** green

**25.** 86

**26.** 0.75

**27.** 2.76

**28.** mixture

**29.** iron

**3.** potassium permanganate

**30.** magnate

**31.** sulfur

**6.** greater than

**32.** carbon disulfide

**34.** evaporation

**35.** combination

**36.** decomposition

**37.** compound

**38.** element

**40.** atoms

**41.** are alike

**42.** compound

**43.** cannot be divided

**44.** barium oxide

**45.** 90

**46.** 32

**47.** 16

**48.** 40

**5.** glycerine

**39.** elements

**50.** 1

**51.** 2

**7.** greater than

**33.** filtration

**8.** nitric acid

**34.** evaporation

**10.** sulfuric acid

**11.** hydrochloric acid

**12.** hydrogen

**19.** A

**39.** elements

**41.** are alike

**42.** compound

**43.** cannot be divided

**44.** barium oxide

**45.** 90

**46.** 32

**47.** 16

**48.** 40

**49.** Cu$_2$S

**50.** 1

**51.** 2
Key to Test on Unit I

(52) CuS
(55) definite proportions
(53) Conservation of matter
(54) 33

(58) C
(57) Ca
(56) Cl

(61) Fe
(59) Sn

(60) Hg

(63) gram atoms
(65) moles
(62) atomic weight units
(64) formula weight
### Daily Log of Unit II:

<table>
<thead>
<tr>
<th>Date</th>
<th>Day</th>
<th>Event</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oct. 16</td>
<td>1</td>
<td>The instructor introduced the unit by a series of questions and demonstrations (Item 1#). The class appeared interested in the sanitary aspects of pure water. Item 3 was assigned for outside study.</td>
</tr>
<tr>
<td>Oct. 17</td>
<td>2</td>
<td>The pretest was administered. The class now understood the purpose of it and were perfectly willing to take it. They completed it in less than half the period. During the second half of the period, the discussion on pure water was continued (Item 4). Most of the boys told how the water was purified in their home towns. Considerable interest was manifested.</td>
</tr>
<tr>
<td>Oct. 18</td>
<td>3</td>
<td>The experiments of Items 5 and 6 were demonstrated.</td>
</tr>
<tr>
<td>Oct. 21</td>
<td>4</td>
<td>The demonstrations of Items 7 and 8 were performed by the teacher and written up by the students. The composition of water by weight (Item 9) was demonstrated by one of the boys.</td>
</tr>
<tr>
<td>Oct. 22</td>
<td>5</td>
<td>Twenty minutes were allowed for the first short test. The experiments done October 21 were discussed. Items 10 and 11 were discussed and assigned for outside pre-</td>
</tr>
</tbody>
</table>

#The items are those on the Study Guide of the Unit Assignment.
The preservation of information and the dissemination of knowledge should be encouraged.

In the majority of cases, we have lost important documents.

It is essential to ensure that our knowledge is not lost.

The preservation of information is crucial for future generations.

Therefore, it is imperative to take immediate action to safeguard our knowledge.

We must ensure that our knowledge is not lost.

Let us work together to preserve our knowledge.

The preservation of information is a responsibility we all share.

Let us take action to safeguard our knowledge.

We have a duty to ensure that our knowledge is not lost.

Let us work together to safeguard our knowledge.

The preservation of information is a collective responsibility.

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paration.

Oct. 23  6  The molecular theory (Item 12) was explained by the instructor. This appeared to be very difficult for the class and it was necessary to go very slowly and give many illustrations. Since the boys were expected to summarize this material in their notebooks, we stopped after every new idea to give them time to do this. For each idea one boy was called on to suggest the proper sentence or sentences to be written in the notebook. The item was not finished during this period.

Oct. 24  7  The material of Item 12 was completed and Items 13 and 14 were assigned and discussed. These items were written in the class notebook.

Oct. 25  8  The instructor explained the meaning of reversible reactions and equations--Item 15. Experiment 2-3 (Item 16) was performed and discussed. The boys were directed to finish writing up their experiments in the barracks.

Oct. 28  9  The first half of the period was used as a study period to prepare Item 17. In the second half the teacher explained Item 18. The class showed little interest. Item 19
was assigned for outside study.

Oct. 29 10 The action of steam on iron (Item 20) was demonstrated by the teacher and experiment 2-4 (Item 21) was performed and discussed.

Oct. 30 11 Twenty minutes was allowed for the second short test. The demonstration (Item 22) was performed by the teacher. Items 24-26 were assigned for outside study.

Oct. 31 12 Experiment 2-5 (Item 23) was demonstrated by one of the boys. The class completed the laboratory notebooks and passed them in for inspection. Item 27 was assigned for outside study.

Nov. 1 13 Experiments 2-8 and 2-9 (Items 28 and 31) were demonstrated by the boys and discussed. Items 29, 30 and 32 were assigned for outside study.

Nov. 4 14 Experiment 2-7 (Item 33) was demonstrated and discussed. The laboratory notebooks were again passed in for inspection. Item 34 was assigned for study in the barracks.

Nov. 5 15 The instructor explained Item 35; one of the boys demonstrated diffusion (Item 36), and another boy demonstrated osmosis (Item 37). These topics were then discussed. The boys were not interested in the concentration of solutions and did not really
learn much about molarity. Osmosis and diffusion, however, seemed familiar to some of the boys from their previous study of biology and they readily grasped their meaning. Item 36 was assigned for outside study.

Nov. 6  16 General review and discussion of the whole unit.

Nov. 7  17 Final test of Unit II.
First short test - Unit II:

1. Will the filtration process remove dissolved materials from water? Explain.

2. Define the following terms and give an example of each: filtrate, residue, distillate.

3. Describe the process of distillation. Will distillation remove a volatile substance? Explain.


Second short test - Unit II:

1. (a) How can you identify a base and an acid?
(b) Is Na₂O an acid anhydride or a basic anhydride? Explain.

2. How does a base differ from an alkali? What is a radical?

3. State the physical properties of water.

4. Describe the action of sodium and calcium on water. Which reacts more actively?

Pretest and Final Test - Unit II:

Directions: All omitted words or phrases appear in the column at the left of the page. Show that you know which word or phrase has been omitted from each blank by putting the number of the blank in the proper parenthesis at the left. For example, the "1" is put in the parenthesis in front of the word omitted from the blank "1".

( ) boiling Water can be freed from suspended material by __1__; it can be freed from dissolved material by __2__. Disease germs can be killed in small quantities of water for household use by __3__; they can be killed in a city water supply by __4__. Filters for a city water supply are usually constructed from __5__.

( ) chemical treatment
( ) sand
( ) distillation
( ) filter paper system
( ) evaporation
( ) freezing
(1) filtration
In the process of distillation, water is first converted to \( \text{volatile} \) and then back to \( \text{solid} \). In the first part of the process, the water is \( \text{vaporized} \); in the second, it is \( \text{condensed} \). Water cannot be freed from impurities by distillation.

The decomposition of a substance for the purpose of determining its composition is called \( \text{hydrolysis} \). The preparation of a substance by the combination of its constituent elements is called \( \text{synthesis} \). Water is \( \text{solid} \) hydrogen by volume and \( \text{solid} \) by weight. Hydrogen and oxygen can be mixed and exploded in a \( \text{apparatus} \) and \( \text{air} \).

If 20 ml of hydrogen made 20 ml of oxygen were mixed and exploded, the gas remaining would be \( \text{oxygen} \) and its volume would be \( \text{35 ml} \). If 50 ml of hydrogen and 15 ml of oxygen were mixed and exploded, the gas remaining would be \( \text{oxygen} \) and its volume would be \( \text{35 ml} \).

This procedure is used to determine the composition of water by \( \text{weight} \).

In an experiment to determine the composition of water by weight, dry \( \text{carbon dioxide} \) gas was passed over hot \( \text{copper oxide} \) and the \( \text{hydrogen chloride} \) formed was absorbed by \( \text{oxygen} \). The loss of weight of the reaction tube represented the weight of \( \text{used} \).

The smallest particle of a compound that can exist alone \( \text{symbol} \). The smallest particle of an element that can take part in a chemical change \( \text{molecule} \). A \( \text{proton} \) represents the proportions in which elements combine. A \( \text{atom} \) represents an atom. The process by which the particles of gases mix together by means of their own motion is called \( \text{diffusion} \).
When mercury is heated in air, mercury and oxygen atoms combine in definite proportions to form \[\text{31}\] of mercuric oxide. Mercuric oxide when heated and separated into \[\text{32}\] of mercury and oxygen. Hydrogen atoms and oxygen atoms combine in the proportion of \[\text{33}\] to form \[\text{34}\] of water. Each of these compounds can be represented by a \[\text{35}\].

The formula for oxygen is \[\text{36}\].

An equation is correct only if it is \[\text{37}\]. The terms in an equation may be \[\text{38}\] if necessary. The correctness of an equation may be checked by \[\text{39}\]. The products of a reaction are written \[\text{40}\]. An equation written with a double arrow \[\text{\rightarrow}\] represents a \[\text{41}\]. The terms in a equation must be \[\text{42}\].

The metal which reacts most violently with cold water is \[\text{43}\]. The gaseous product of this reaction is \[\text{44}\]. Bases are characterized by the presence of the \[\text{45}\] radical. An indicator which turns red in the presence of a base is \[\text{46}\]. The metallic radical in common salt is \[\text{47}\].

Zinc will react with steam to form zinc \[\text{48}\]. Iron will react with steam to form iron \[\text{49}\]. Quicklime exposed to air becomes air slaked by forming calcium \[\text{50}\]. The anhydride of phosphoric acid is phosphorus \[\text{51}\]. Carbon dioxide is \[\text{52}\] anhydride.
hydrolytic
anhydrous
efflorescent
deliquescent
hydrate
magnetic
dessicant
hygroscopic

v/ater of crystallization _53_. A crystal without water of crystallization _54_. A crystal which gives up water of crystallization at room temperature _55_. A substance which absorbs water from the air is _56_; one which absorbs enough to become wet is _57_.

A solid solution of two metals such as copper zinc _58_. A homogeneous mixture of two or more substance, the particles of which have molecular dimensions _59_. The most common solvent _60_. An organic solvent _61_. The solution of a non-volatile substance in alcohol _62_. The solution of a volatile substance in alcohol _63_.

A solution containing a relatively large amount of the solute _64_. A solution containing a relatively small amount of the solute _65_. A solution containing all the solute it can dissolve at that temperature _66_. A solution containing more of the solute than the solvent can dissolve at that temperature _67_. A solution which could dissolve more of the solute at that temperature _68_.

The molecular weight of sodium hydroxide is 40. A 0.1 M solution of NaOH contains _69_ moles per liter, _70_ grams per 100 ml, and _71_ grams per milliliter. A 4% solution of NaOH will contain about _72_ g/l or about _73_ moles per liter.
The solubility of a substance is defined as the number of grams which dissolve in _74_ of the _75_. The solubility of most solids in water is _76_ by raising the temperature. The solubility of most gases in water is increased by increasing the _77_. The particles of the solvent by the process of _78_. Particles of solute pass through the cell walls of plants by the process of _79_.

According to the kinetic theory of solutions, the structure of a solution is similar to the structure of a _80_. The greater the amount of solute dissolved in a given quantity of solvent, the higher is the _81_ and the lower the _82_.

- increased
- solvent
- temperature
- 1 liter
- 100 grams
- solution
- 100%
- decreased
- osmosis
- occlusion
- pressure
- hydrolysis
- equilibrium
- diffusion
- solid
- liquid
- kinetic energy
- gas
- boiling point
- crystal
- freezing point
- molecule
- eutectic point
- calories
Key to Test on Unit II

(3) boiling
(4) chemical treatment
(5) sand
(2) distillation

(1) filtration

(10) volatile
(8) vaporize
(6) steam
(7) water
(9) condensed

(12) synthesis
(11) analysis

(13) 2/3
(14) 11.2

(15) eudiometer

(17) 10 ml
(19) 20 ml

(18) hydrogen
(16) oxygen

(20) synthesis

(23) water

(22) copper oxide
(24) calcium chloride
(21) hydrogen
(25) oxygen

(30) diffusion
(29) symbol
(26) molecule

(27) atom
(28) formula

(36) O₂
(31) crystals
(34) molecules

(33) 2 : 1

(32) atoms
(35) formula

(37) balanced

(38) multiplied by a coefficient
(41) a reversible reaction
(39) counting atoms
(42) formulas

(40) to the right of the arrow

(43) potassium
(47) sodium

(45) hydroxide
(46) phenolphthalein

(44) hydrogen
Key to Test on Unit II

(52) carbonic (71) 0.004
(49) oxide (70) 0.4
(51) pentoxide (69) 0.1
( ) (73) 1.0
( )
(50) carbonate (72) 40
(48) hydroxide ( )

* * * * * * * * * * * *

( ) ( )
(54) anhydrous (76) increased
(55) efflorescent (75) solvent
(57) deliquescent ( )
(53) hydrate (74) 100 grams
( )
(56) hygroscopic ( )

* * * * * * * * * * * *

(62) tincture (79) osmosis
(61) alcohol (77) pressure
( )
( )

* * * * * * * * * * * *

(59) solution ( )
(63) spirits (80) gas
( )
(81) boiling point
(58) alloy ( )
(60) water (82) freezing point

* * * * * * * * * * * *

(68) unsaturated ( )
(66) saturated (60) gas
(67) supersaturated (81) boiling point
(65) dilute ( )
(64) concentrated ( )

( )

The scores on the final test showed substantial improvement over the scores on the pre-test, the average score increased from 60 to 90 points.
<table>
<thead>
<tr>
<th>Yd</th>
<th>Cm</th>
<th>Inch</th>
<th>Strips</th>
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<td>116</td>
<td>1.81</td>
<td></td>
</tr>
<tr>
<td>36</td>
<td>86</td>
<td>1.42</td>
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</tr>
<tr>
<td>16</td>
<td>41</td>
<td>0.64</td>
<td>60</td>
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<th>Strips</th>
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<td>0.76</td>
<td>0.093</td>
<td></td>
</tr>
<tr>
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<td>1.02</td>
<td>0.123</td>
<td>40</td>
</tr>
<tr>
<td>0.5</td>
<td>1.25</td>
<td>0.181</td>
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<td>0.305</td>
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<td>7.87</td>
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<th>Cm</th>
<th>Inch</th>
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<td>0.25</td>
<td>0.031</td>
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<td>0.2</td>
<td>0.5</td>
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<td>0.4</td>
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<td>0.123</td>
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<tr>
<td>0.5</td>
<td>1.25</td>
<td>0.181</td>
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<tbody>
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<td>0.6</td>
<td>1.52</td>
<td>0.305</td>
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</tbody>
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<table>
<thead>
<tr>
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<th>Inch</th>
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<tbody>
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<tr>
<td>15</td>
<td>5.91</td>
</tr>
<tr>
<td>20</td>
<td>7.87</td>
</tr>
</tbody>
</table>
Discussion and Evaluation

Observations on the reaction of the class to the units:

As stated above, the class consisted of eight members in the general course. They had such low intelligence quotients and were so poor in English that they could not be expected to do brilliant work or to do very much in the way of outside study or optional topics. They did, however, show a reasonable amount of interest in the class activities. Although they made an honest effort to participate in discussion, they were not very apt in self-expression. The required work did not seem to be too difficult for them; but none of them did any of the optional topics. This was partly due to lack of interest in school work, but mostly due to an actual lack of time because of their slow rate of progress on the assigned topics.

The pre-tests and final tests:-- The pre-tests were the same as the final tests in each case, but the students did not know that this was to be so. The results showed that the students had some general knowledge of these topics at the beginning of their study of the unit, as would be expected from their courses in general science and biology. The test on the first unit contained 89 items; the test on the second unit, 82 items.

The scores on the final test showed noticeable improvement over the scores on the pre-test. Raw scores were converted into letter grades by determining the standard deviation and choosing "cut off points" of
Decisions and the Welfare Plan

Introduction to the Procedure for the Plan of the Welfare

As you can imagine, the plan is developed with careful consideration. The budget committee, which you may know, is responsible for the approval of the budget. These meetings are held to discuss the financial aspects of the plan and to ensure that it is in line with the organization's overall financial goals.

The budget-approved program is formally presented to the membership for approval. This presentation includes an in-depth analysis of the plan's objectives, strategies, and expected outcomes. It also includes a cost-benefit analysis, highlighting the potential benefits and costs associated with the plan.

In conclusion, the welfare plan is a strategic initiative that aims to improve the well-being of all employees. It is designed to create a healthy and supportive work environment, which is essential for the organization's success and the employees' satisfaction.

The plan is reviewed and updated annually to reflect changes in the organization's needs, workforce demographics, economic conditions, and regulatory requirements.
Mean \( \pm 0.5 \) sigma and \( \pm 1.5 \) sigma as explained below.

The students' scores on the tests of Units I and II and the method of calculation of grades is given on the following pages.
Calculation of the standard deviation on final test

<table>
<thead>
<tr>
<th>Student</th>
<th>Score</th>
<th>Diff</th>
<th>Diff.²</th>
<th>Student</th>
<th>Score</th>
<th>Diff</th>
<th>Diff.²</th>
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<tbody>
<tr>
<td>A</td>
<td>59</td>
<td>4</td>
<td>16</td>
<td>A</td>
<td>64</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>B</td>
<td>75</td>
<td>20</td>
<td>400</td>
<td>B</td>
<td>58</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>C</td>
<td>60</td>
<td>5</td>
<td>25</td>
<td>C</td>
<td>65</td>
<td>6</td>
<td>36</td>
</tr>
<tr>
<td>D</td>
<td>61</td>
<td>6</td>
<td>36</td>
<td>D</td>
<td>58</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>E</td>
<td>46</td>
<td>9</td>
<td>81</td>
<td>E</td>
<td>57</td>
<td>2</td>
<td>4</td>
</tr>
<tr>
<td>F</td>
<td>46</td>
<td>9</td>
<td>81</td>
<td>F</td>
<td>56</td>
<td>3</td>
<td>9</td>
</tr>
<tr>
<td>G</td>
<td>45</td>
<td>10</td>
<td>100</td>
<td>G</td>
<td>46</td>
<td>13</td>
<td>169</td>
</tr>
<tr>
<td>H</td>
<td>47</td>
<td>8</td>
<td>64</td>
<td>H</td>
<td>64</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>Mean</td>
<td>55</td>
<td></td>
<td></td>
<td>Mean</td>
<td>59</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[ \text{Sigma} = \sqrt{\frac{803}{8}} = 10 \]
\[ \sqrt{\frac{270}{8}} = 6 \]

"Cut off points" of ± 0.5 sigma and ± 1.5 sigma are chosen to give a distribution of 7% A, 24% B, 38% C, 24% D, and 7% E on a normal curve. The resulting "cut off points" for these two tests are shown below.

<table>
<thead>
<tr>
<th>Unit I</th>
<th>Unit II</th>
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<tbody>
<tr>
<td>A</td>
<td>70</td>
</tr>
<tr>
<td>B</td>
<td>60 - 69</td>
</tr>
<tr>
<td>C</td>
<td>50 - 59</td>
</tr>
<tr>
<td>D</td>
<td>40 - 49</td>
</tr>
<tr>
<td>E</td>
<td>0 - 39</td>
</tr>
</tbody>
</table>
### Table

<table>
<thead>
<tr>
<th>Unit II</th>
<th>Unit I</th>
</tr>
</thead>
<tbody>
<tr>
<td>00 - 09</td>
<td>A</td>
</tr>
<tr>
<td>10 - 19</td>
<td>B</td>
</tr>
<tr>
<td>20 - 29</td>
<td>C</td>
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<tr>
<td>30 - 39</td>
<td>D</td>
</tr>
<tr>
<td>40 - 49</td>
<td>E</td>
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<td>50 - 59</td>
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<td>60 - 69</td>
<td>G</td>
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<td>70 - 79</td>
<td>H</td>
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<tr>
<td>80 - 89</td>
<td>I</td>
</tr>
<tr>
<td>90 - 99</td>
<td>J</td>
</tr>
<tr>
<td>Scores and Grades</td>
<td>Students</td>
</tr>
<tr>
<td>-------------------------------------------------------</td>
<td>----------</td>
</tr>
<tr>
<td></td>
<td>A</td>
</tr>
<tr>
<td>Intelligence Quotient</td>
<td>102</td>
</tr>
<tr>
<td>Reading Comprehension - Educational Records Bureau</td>
<td>79</td>
</tr>
<tr>
<td>private school scale</td>
<td></td>
</tr>
<tr>
<td>Mechanics of English - Educational Records Bureau</td>
<td>1</td>
</tr>
<tr>
<td>private school scale</td>
<td></td>
</tr>
<tr>
<td>Pretest - Raw Score</td>
<td>18</td>
</tr>
<tr>
<td>First Short Test - on basis of 100 per cent</td>
<td>43</td>
</tr>
<tr>
<td>Second Short Test - on basis of 100 per cent</td>
<td>74</td>
</tr>
<tr>
<td>Experiment Grade - on basis of 100 per cent</td>
<td>75</td>
</tr>
<tr>
<td>Final Test - Raw Score</td>
<td>59</td>
</tr>
<tr>
<td>Final Test - Letter Grade determined from standard deviation as shown below, page 82</td>
<td>C' A</td>
</tr>
<tr>
<td>Gain on Final Test (Final Test score - Pretest score)</td>
<td>41</td>
</tr>
<tr>
<td>Final Grade for Unit:</td>
<td>C</td>
</tr>
<tr>
<td>Average of the two short tests and experiment grade 1/3; final test grade 2/3</td>
<td></td>
</tr>
<tr>
<td>* Tests were not administered to these boys.</td>
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</tbody>
</table>
### Table II  Student Scores and Grades on Unit II

<table>
<thead>
<tr>
<th>Scores and Grades</th>
<th>Students</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>A</td>
</tr>
<tr>
<td>Pretest- Raw Score</td>
<td>20</td>
</tr>
<tr>
<td>First Short Test - on basis of 100 per cent</td>
<td>75</td>
</tr>
<tr>
<td>Second Short Test - on basis of 100 per cent</td>
<td>85</td>
</tr>
<tr>
<td>Experiment Grade - on basis of 100 per cent</td>
<td>70</td>
</tr>
<tr>
<td>Final Test - Raw Score</td>
<td>64</td>
</tr>
<tr>
<td>Final Test - Letter grade determined from standard deviation as shown, page 82</td>
<td>B</td>
</tr>
<tr>
<td>Gain on Final Test - (Final Test score - Pretest score)</td>
<td>44</td>
</tr>
<tr>
<td>Final Grade for Unit: Average of the two short tests and experiment grade 1/3; final test grade 2/3</td>
<td>B-</td>
</tr>
</tbody>
</table>
Student Evaluation of Methods and Topics:— In order to get the reaction of the students to some of the methods and topics used in teaching these units, a check list was prepared and submitted to the students. Each student was asked to indicate by a plus sign (+) that he found a particular method or topic "liked", "easy", or "helpful"; by a minus sign (-) that he found it "disliked", "difficult", or "useless"; and by a zero (0) that he regarded as indifferent. The results are shown in tables III and IV. For purposes of comparison, the papers were scored by allowing +10 for each plus sign, -10 for each minus sign, and 0 for each zero. A total score of 80 on an item then would indicate that all eight students regarded this item with favor.

The data show that the students liked experiments which they could look at or do; but they disliked anything that involved any thinking or writing or work of any kind. Although it is impossible to draw valid conclusions from the study of so few cases, especially in a non-college section, it does not appear that there are any serious deficiencies in the methods or materials used.

Six of the students indicated that they preferred the Unit Method to the Textbook Method of instruction; the other two had no preference. From this we may conclude with certainty that the Unit Method makes it easier for poor students to grasp the material.
Table III. Student Evaluation of Teaching Methods

<table>
<thead>
<tr>
<th>Activity</th>
<th>Liked or Disliked</th>
<th>Easy or Difficult</th>
<th>Helpful or Useless</th>
</tr>
</thead>
<tbody>
<tr>
<td>Talks by teacher</td>
<td>40</td>
<td>20</td>
<td>30</td>
</tr>
<tr>
<td>Written answers to questions</td>
<td>10</td>
<td>-50</td>
<td>50</td>
</tr>
<tr>
<td>Discussion of questions</td>
<td>20</td>
<td>20</td>
<td>40</td>
</tr>
<tr>
<td>Teacher demonstrations</td>
<td>80</td>
<td>40</td>
<td>70</td>
</tr>
<tr>
<td>Pupil demonstration</td>
<td>80</td>
<td>40</td>
<td>70</td>
</tr>
<tr>
<td>Notes made in class</td>
<td>-50</td>
<td>-60</td>
<td>10</td>
</tr>
<tr>
<td>Written reports on Experiments</td>
<td>10</td>
<td>-20</td>
<td>40</td>
</tr>
<tr>
<td>Class Study periods</td>
<td>30</td>
<td>20</td>
<td>60</td>
</tr>
<tr>
<td>Optional Topics</td>
<td>-50</td>
<td>-50</td>
<td>-20</td>
</tr>
<tr>
<td>Recitations</td>
<td>10</td>
<td>-20</td>
<td>60</td>
</tr>
<tr>
<td>Subject</td>
<td>Task</td>
<td>Time</td>
<td></td>
</tr>
<tr>
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<td>------</td>
<td>------</td>
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<tr>
<td>00</td>
<td>00</td>
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<tr>
<td>05</td>
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<td></td>
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<td>00</td>
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</tr>
<tr>
<td>45</td>
<td>00</td>
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</tbody>
</table>

*Note: Time is in minutes.*
<table>
<thead>
<tr>
<th>Topic</th>
<th>Liked Or Disliked</th>
<th>Easy or Difficult</th>
<th>Helpful or Useless</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chemical changes</td>
<td>80</td>
<td>60</td>
<td>50</td>
</tr>
<tr>
<td>Mixtures, compounds, Elements</td>
<td>40</td>
<td>40</td>
<td>60</td>
</tr>
<tr>
<td>Atomic Theory</td>
<td>-30</td>
<td>-50</td>
<td>20</td>
</tr>
<tr>
<td>Formulas</td>
<td>40</td>
<td>-50</td>
<td>70</td>
</tr>
<tr>
<td>Equations</td>
<td>-50</td>
<td>-80</td>
<td>40</td>
</tr>
<tr>
<td>Weight Problems</td>
<td>-40</td>
<td>-60</td>
<td>-40</td>
</tr>
<tr>
<td>Metric System</td>
<td>-20</td>
<td>-60</td>
<td>0</td>
</tr>
<tr>
<td>Law of Conservation of Mass</td>
<td>60</td>
<td>60</td>
<td>30</td>
</tr>
<tr>
<td>Law of Definite Proportions</td>
<td>20</td>
<td>30</td>
<td>30</td>
</tr>
<tr>
<td>Historical Topics</td>
<td>-70</td>
<td>-10</td>
<td>-10</td>
</tr>
<tr>
<td>Experiments on the Composition of Water</td>
<td>80</td>
<td>50</td>
<td>-30</td>
</tr>
<tr>
<td>Molecular Theory</td>
<td>-10</td>
<td>0</td>
<td>30</td>
</tr>
<tr>
<td>Action of Metals on Water</td>
<td>80</td>
<td>70</td>
<td>-10</td>
</tr>
<tr>
<td>Definitions</td>
<td>-70</td>
<td>0</td>
<td>-30</td>
</tr>
<tr>
<td>Solutions</td>
<td>30</td>
<td>40</td>
<td>-20</td>
</tr>
<tr>
<td>Molarity</td>
<td>-50</td>
<td>10</td>
<td>-50</td>
</tr>
<tr>
<td>Purification of Water</td>
<td>80</td>
<td>70</td>
<td>70</td>
</tr>
<tr>
<td>Osmosis</td>
<td>60</td>
<td>70</td>
<td>10</td>
</tr>
<tr>
<td>Effect of dissolved substances on freezing point</td>
<td>30</td>
<td>30</td>
<td>0</td>
</tr>
<tr>
<td>Distillation</td>
<td>80</td>
<td>80</td>
<td>0</td>
</tr>
<tr>
<td>Acids-bases-salts</td>
<td>30</td>
<td>10</td>
<td>10</td>
</tr>
<tr>
<td>Action</td>
<td>Target</td>
<td>Health</td>
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</tr>
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Note: The table above is a placeholder. The actual content is not legible due to the image quality.
Conclusion:— The method of the unit assignment is much superior to the old daily assignment procedure with this type of class. The students seemed to grasp the ideas better because they were able to gain a better perspective of the field being studied. Their view of the forest was not obscured by the trees.

Certain of the topics given as optional should be required of classes which are preparing for college. The study guide should, of course, be revised frequently to keep it up to date and to shift the emphasis according to the progress of science and the needs of the class.

The organization here proposed proved to be satisfactory in use and worthy of further study with classes at various levels.
Appendix A

In the following pages are given the laboratory directions for the experiments called for in the study of these units. These experiments are mimeographed and distributed to students in classes which do individual laboratory work. With the class used in this study, however, typed carbon copies were used. It is planned to expand these directions into a complete laboratory manual for the entire course.
In this letter, where are the people important?

Attention is paid to the importance of making sure to the main focus of these matters. These occurrences are presented clearly.

In this manner, the importance of focusing on issues which go...

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LABORATORY DIRECTIONS

General Directions

The laboratory work for this course has been planned according to modern semi-micro procedures. Small apparatus and smaller amounts of chemicals are used than are employed in the older type of experiments. Some of the advantages of the newer methods are: economy of apparatus and chemicals, small storage space required, greatly reduced danger from fumes and explosions, economy of time required for each experiment so that more experiments may be done in the same amount of class time, and the premium placed on neat and careful work. None of the time-tested essential experiments or techniques are omitted; on the contrary, new techniques are learned and students are prepared for the methods which they may meet later in industry or college. Since laboratories are now generally equipped with the larger apparatus, it may be necessary to modify some of the experiments assigned in order to introduce semi-micro procedures gradually as new equipment is purchased.

Commonly used reagents will be supplied to the desks in 15 ml dropping bottles. The teacher may dispense solids to the students by walking up and down the aisle and placing in a dish or on a labeled piece of paper the very small quantities of the materials required or he may make the material available in a dish or bottle at a convenient place. Acids will be kept in 1 oz. glass stoppered bottles and
medicine droppers used to remove the amounts needed. General reagents on the laboratory shelf will be kept in 250 ml bottles fitted with a bakelite cap and dropper. As far as possible the solid reagents used in the course will be kept on a side shelf in small 1 or 2 oz. wide mouth bottles with plastic caps so that students may have access to small quantities of chemicals for special work before or after school, during make-up periods, etc. No student should at any time remove any material from the main laboratory stock without instructions from the teacher.

Experiment 1-1

Properties

Apparatus and Materials: Magnet, test tubes, test tube holder, salt, sugar, powdered iron, powdered sulfur, carbon disulfide, dilute hydrochloric acid.

Object: To learn how to recognize and identify substances by their properties.

Procedure: Perform the operations described below. Record in your notebook a description of what you do and what you observe. Answer the questions in complete sentences. What conclusion or conclusions can you draw from this experiment?

Examine a few crystals of salt and a few crystals of sugar. Can you tell them apart by their color? Can you tell them apart by the shape of their crystals without using a microscope? What property would you use to tell the difference between salt and sugar? Would you use this same test to tell the difference between sugar
Caretaking proverbs need to be taught to children. General teaching on the importance of any occupation will not be effective if it is not accompanied by practical experience. The same principle holds true in the home; children need to be shown by example how to perform household chores and similar tasks.

Experiments

Object: To teach you to recognize and identify suitable employees.

Procedure: Perform the occupations described below. Examine the type of employee who is most suitable for each task.

1. Household management
2. Cooking
3. Garden work
4. Cleaning
5. Repairing

Explain the reasons for each choice. Note any gratifying results in each case. This will help you to realize the importance of having suitable employees in various fields of work.
and powdered glass? Can you think of a way to tell the difference between these two materials?

Examine some powdered iron and determine some of its properties. What is the color? Is it attracted by a magnet?

Place about 0.1 gram of the iron powder at the bottom of a 10 centimeter test tube and add 5 drops of dilute hydrochloric acid with a dropper. If there is no action, warm the tube slightly by holding it in the flame for not more than 2 seconds. Then turn the flame out. Describe the reaction between iron and hydrochloric acid. Does the product of this reaction have any pronounced odor? Describe the odor, if any.

Examine some sulfur and determine some of its properties. What is the color of the sulfur? Is it attracted by a magnet? Does sulfur dissolve in dilute hydrochloric acid? (Follow the directions given for iron and hydrochloric acid.) Does sulfur dissolve in carbon disulfide? (Place 0.1 gram of powdered sulfur in a 10 centimeter test tube, add 2 milliliters of carbon disulfide, place the thumb over the mouth of the test tube, and shake vigorously. Carbon disulfide is flammable, keep all flames away!)

If 1 gram of powdered iron were mixed with 1 gram of powdered sulfur, could you separate them with a magnet? On what properties of iron and sulfur does this separation depend?
The attention of the Secretary-General is now called to the need for a comprehensive and integrated plan for the peaceful coexistence of all peoples and nations. It is essential that a new international order be established, based on the principles of justice, equality, and mutual respect, in order to ensure a lasting peace and prosperity for all.

The Secretary-General proposes to present a detailed plan for this purpose, which will be discussed and debated by the General Assembly. The plan will include measures for the resolution of conflicts, the promotion of human rights, and the establishment of a universal system of justice.

It is hoped that all member states will contribute to this plan, and that it will be adopted by the General Assembly as a global blueprint for peace and development.
(Note: Carbon disulfide has a disagreeable odor and the vapors are not only flammable but definitely injurious if inhaled in large amounts. There is absolutely no danger, however, in the use of the quantities specified in this and in subsequent experiments.)

Remember that the most important part of your written description of this and subsequent experiments is the statements at the end in which you summarize what you have learned from the experiment.

Experiment 1-2

**Separation of a Mixture into its Components**

**Apparatus and Materials:** Test tubes, funnel, watch glass, filter paper, powdered iron, powdered sulfur, carbon disulfide.

**Object:** To learn how we may separate a mixture by taking advantage of the physical properties of its components.

**Procedure:** Mix 0.1 g powdered sulfur with 0.1 g powdered iron. Place this material in a 10 cm test tube and add 2 ml of carbon disulfide. Shake well. Fold a circle of filter paper as demonstrated by the teacher and fit it carefully into a dry funnel. Pour the mixture from the test tube on to the filter paper and catch the filtrate in another 10 cm test tube. Wash the residue on the filter by pouring about 1 ml carbon disulfide over the residue and letting it run down into the filtrate. Pour the filtrate on to a watch glass and let it stand on the desk or in the hood until the solvent
has evaporated. KEEP ALL FLAMES AWAY as carbon disulfide is very flammable.

Examine the residue on the watch glass. What is the color of the residue? What is this substance? By what property do you recognize it?

On what physical properties of iron and sulfur was this separation based? By what process was the solution of sulfur separated from the iron? By what process were the crystals of sulfur formed? What is a residue? Give two examples. What is a solvent? Give two examples. What do we call the liquid which runs through the filter?
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Experiment 1-3

What is a Compound?

Apparatus and Materials: Sulfur, powdered iron, dilute hydrochloric acid, test tubes, test tube holder, Bunsen burner, magnet, forceps, micro spatula or stirring rod.

Object: To learn the nature of a compound.

Procedure: Mix 0.2 g (0.1 cm$^3$) powdered sulfur with 0.4 g (0.05 cm$^3$) powdered iron on a small piece of paper.

What name is given to a material like this in which each component retains its own properties? Name one property of iron and one property of sulfur which have not been affected by the fact that the two substances have been mixed together. Divide the material into two parts.

Transfer one part to a perfectly dry 10 cm test tube. Hold the tube with a test tube holder in the flame of a Bunsen burner. Heat gently at first, and then as hot as possible until a reaction starts. Remove the tube from the flame the instant you notice a red glow in the mixture. Does the glow continue after the heating is stopped? Why is this evidence that a chemical change has occurred? Return the tube to the flame for not more than half a minute just to make sure that the reaction is complete. Allow the tube and its contents to cool, loosen the material in the tube with a microspatula or a stirring rod, and empty it into a small porcelain dish. The product obtained is black or gray like iron, but it is not attracted by a magnet (try it) and therefore it cannot be iron.
Now in one 10 cm test tube, place all except a very small piece of the product obtained by heating the mixture, and in another 10 cm test tube place the portion of the original mixture which was not heated. Add 5 drops of dilute hydrochloric acid to each portion. If the reaction in either tube is slow in starting, warm gently by passing the tube slowly through the flame. You must be very careful not to overheat the mixture, or the iron and sulfur will react so that you will have the same product in each tube. Cautiously smell the gas issuing from each tube. The gas produced in the tube containing the new substance has the extremely unpleasant odor of rotten eggs. It is called hydrogen sulfide. The odor, you will agree, is distinctly different from that obtained when hydrochloric acid is added to the mixture.

Is the product formed by heating iron and sulfur insoluble in carbon disulfide? (Try it.) It is obvious that a new substance with properties quite different from those of iron or sulfur has been formed. It is called a compound substance, or simply a compound, because it has been formed by the combination of two simpler substances.

When the mixture of iron and sulfur was heated, what evidence that a chemical change was occurring did you notice? What evidence, based on a physical property, have you that the product of this reaction was entirely different from either iron or sulfur? What evidence, based on a chemical property, have you? Define chemical
change. Summarize what you have learned in this experiment.

Experiment 1-4

How does a Compound Differ from a Mixture?

Apparatus and Materials: Iodine, antimony (powdered or crystalline), carbon tetrachloride, mortar and pestle, funnel, filter paper, test tubes, watch glass, stirring rod.

Object: To study further the difference between a compound and a mixture.

Procedure: In this experiment we shall grind together antimony and iodine in a mortar. If they combine to form a compound, a new substance with new properties will be formed; but if they merely form a mixture, they will retain all their original properties. We shall also heat iodine and antimony together.

Examine a small crystal of iodine and a small piece of antimony metal. Describe the appearance of each. Test the solubility of each in carbon tetrachloride. (Mix about 0.1 g antimony and 2 ml carbon tetrachloride in a 10 cm test tube, stir with a glass rod, and heat cautiously in a small flame. Carbon tetrachloride is not flammable, but too strong heating may cause it to spurt out of the tube. Mix 0.1 g iodine with 2 ml carbon tetrachloride in a 10 cm test tube, stir with a glass rod, heat cautiously. If solution occurs in either case, pour the solution into a watch glass (filtering if necessary to remove any insoluble residue) and let the solvent evaporate. Is
the original substance recovered from solution?

Mix 0.1 g iodine and 0.1 g antimony in a mortar and grind them together with the pestle. Does the grinding make the material appear homogeneous (alike in all its parts)? Do you think that iodine and antimony have combined? Add 2 ml of carbon tetrachloride to the mortar, stir well with the pestle, and filter. Add an additional 1 ml portion of carbon tetrachloride to the mortar, stir well, and filter through the same paper. Wash the residue on the paper with a few drops of carbon tetrachloride. Pour about 1 ml of the filtrate on to a watch glass and let it evaporate. What is the residue on the filter paper? What is the residue on the watch glass? Did antimony and iodine combine on grinding or merely mix? Give a reason for your answer.

Return the separated iodine and antimony to the solution in a 10 cm test tube. If necessary to replace not loss, add more than 0.1 g of either antimony or iodine to the solution. Heat the mixture cautiously over a low flame until the iodine color disappears and a yellow liquid containing particles of antimony is formed. Filter the hot solution through a fresh filter paper, collecting the filtrate in a watch glass. Allow the solvent to evaporate. Describe the residue. Does it bear any resemblance to either antimony or iodine? Can you separate it into two components as you separated the components of the mixture by dissolving in carbon tetrachloride and filtering? (Try it.) Have you formed a compound?
WHITING
MUTUAL BOND
(RE VANCE)
Give a reason for your answer. If you think you have formed a compound, can you name it?

What is the difference between a compound of iodine and antimony and a mixture of iodine and antimony?

Experiment 1-5

**Decomposition of a Compound**

**Apparatus and Materials:** Test tube, test tube holder, Bunsen burner, evaporating dish, mercuric oxide, wooden splinter.

**Object:** To learn more about the nature of a compound.

**Procedure:** Place 0.5 g mercuric oxide at the bottom of a perfectly dry 10 cm test tube. Hold the test tube horizontally by means of a test tube holder and heat the mercuric oxide gently. Then gradually increase the heat until you are heating it as hot as possible. From time to time, test the gas in the tube by inserting a glowing splinter. To obtain a "glowing splinter" take one of the wooden splints and light it in the flame of the burner. Then blow out the flame, leaving a glowing spark on the end of the stick.

What are the results of these tests? Describe what you see on the sides of the tube. Scrape some of the material out of the tube with a stick into an evaporating dish. Describe its appearance. What is it?

The gas is oxygen. By what property can it be recognized?

Since in this experiment we have obtained two different substances with different properties from a single substance, what type of substance must the mercuric
oxide be?

What name is often given to the type of reaction in which a substance is broken up or decomposed into simpler substances? What name is often given to the type of reaction in which a compound is formed by the combination of simpler substances.

Can mercury and oxygen be decomposed into simpler substances? What kind of substances are they?

Experiment 1-6

Weight Relations in a Chemical Change

Apparatus and Materials: Crucible and cover, Bunsen burner, ring stand and ring (or tripod), triangle, balance and weights, copper turnings, sulfur.

Object: To determine the weight of sulfur which will combine with one gram of copper.

Procedure: Weigh a clean dry porcelain crucible to the nearest centigram. Add about two grams of fine copper turnings and weigh again to the nearest centigram. The difference between these two weights will be the exact weight of the copper.

Add enough powdered sulfur to cover the copper completely. Cover the crucible with a porcelain crucible cover, place it on a triangle supported on a tripod or by a ring stand in the hood and heat very slowly and gently. When the sulfur stops burning around the edges, heat the crucible as hot as possible for about five minutes. Then let the crucible cool to room temperature, remove the cover, and weigh. Add a little more sulfur,
cover, and heat as before. Again cool the crucible and weigh it. If these last two weights agree within 0.05 grams, the weight is said to be "constant". If the weights do not agree, add more sulfur, heat, cool and weigh as before until constant weight is obtained.

Subtract the weight of the empty crucible and its contents after heating to obtain the weight of the copper sulfide formed. Subtract the weight of the copper from the weight of the copper sulfide to obtain the weight of sulfur which combined with the weight of copper taken. Record all the data in your notebook in a table like the one below. From these data calculate the weight of sulfur which would combine with one gram of copper.

This experiment may take you a week or more to do and should be fitted in with the other work you are doing. It furnishes excellent training in how to plan laboratory time. When you leave the laboratory, set the crucible in a small beaker and put it away in your desk or in some other place designated by the teacher. Write your name in the small etched circle on the beaker for identification. As an example of time planning, you might get the crucible all ready to heat during a class period, come in after school for a few minutes to heat it, let it cool over night, and weigh it either before school in the morning or during the class period. Do not waste time while this experiment is in progress; have other work ready to do while crucibles are heating or cooling.
First, you need to consider the potential for a more permanent and effective solution to the problem. This approach may involve the installation of a new system or an upgrade to an existing one. It is important to thoroughly analyze the current system and identify any weaknesses or areas for improvement.

Second, it is crucial to consider the financial implications of the proposed solution. This includes the initial cost of the system, ongoing maintenance expenses, and any potential savings in the long run. A comprehensive cost-benefit analysis should be conducted to ensure that the investment is justified.

Lastly, it is important to involve all stakeholders in the decision-making process. This includes employees, customers, and other relevant parties. Open communication and collaboration will help ensure that everyone's concerns are addressed and that the solution is accepted and implemented successfully.
Table of Data

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<thead>
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Experiment 1-7

Determination of the Formula of a Sulfide

Obtain a metal from the instructor and attempt to determine the formula of its sulfide. Find the proportions in which the elements combine according to the directions given in Experiment 1-6. In some cases you will find that the metal combines readily with sulfur to form a sulfide and you will be able to determine its formula without difficulty. In other cases you will find that the metal either combines with sulfur so slowly that you do not succeed in forming a compound, or that the reaction is so violent that it blows some of the compound out of the crucible.
In these experiments which fail, collect the data and report the experiment in your notebook with a full description of what happened. This account will be valuable chemical information for you and for your classmates even though you may be unable to obtain a formula. Iron, nickel, aluminum, magnesium, zinc, and tin make interesting metals to work with. Also try any others which may be available. Try more than one metal if you have time.

If the weight of the copper used in Experiment 1-6 is divided by the atomic weight of copper, 63.57, the number of gram atoms of copper used will be obtained. If the weight of sulfur which combines with this weight of copper is divided by the atomic weight of sulfur, 32.06, the number of gram atoms of sulfur will be obtained. Thus we can find the number of gram atoms of copper which combine with a corresponding number of gram atoms of sulfur.

A student obtained the following data in his experiment:

Weight of copper 2.13 g
Weight of sulfur 0.54 g
Gram atoms of copper: \( \frac{2.13}{63.56} = 0.0335 \)
Gram atoms of sulfur: \( \frac{0.54}{32.06} = 0.0168 \)

This means that 0.0335 gram atoms of copper combine with 0.0168 gram atoms of sulfur. It can be seen by inspection that there are just about twice as many gram atoms of copper as gram atoms of sulfur; that is, two atoms of copper combine with one atom of sulfur. This fact
is represented by the formula Cu₂S. If the ratio of the corresponding numbers of gram atoms is not obvious from inspection, it can be determined by dividing each value by the smallest one.

\[
\frac{0.0335}{0.0168} = 2 \quad \frac{0.0168}{0.0168} = 1
\]

To find the formula of a compound, therefore,

1. Divide the weight of each element in the compound (or the per cent of the element in the compound) by the atomic weight to determine the number of gram atoms of each element.

2. Divide the several quotients thus obtained by the smallest one.

3. If this second series of quotients are not all whole numbers, or very nearly whole numbers within the limits of experimental error, multiply each one by the smallest integer which will make them all whole numbers.

4. The whole numbers finally obtained are the subscripts of the corresponding symbols in the formula.

Example: A compound has the composition carbon 90%, hydrogen 10%. Find the formula. (Note that, from the meaning of per cent, each 100 grams of the compound contain 90 g carbon and 10 g hydrogen.)

\[
\begin{align*}
C & \quad \frac{90}{12} = 7.5 \\
H & \quad \frac{10}{1} = 10.0
\end{align*}
\]

\[
\begin{align*}
\text{(1)} & \quad \frac{90}{7.5} = 12 \quad & \quad \frac{7.5}{7.5} = 1 \\
\text{(2)} & \quad \frac{7.5}{1} = 7.5 \\
\text{(3)} & \quad \frac{10.0}{7.5} = 1.33 \quad & \quad \frac{1.33}{1.33} = 1 \\
\text{(4)} & \quad \frac{7.5}{7.5} = 1 \quad & \quad \frac{1.33}{1.33} = 1
\end{align*}
\]

\[C_3H_4\]
The solution to the problem is as follows:

\[ I = \frac{1000}{3000} \]

\[ B = \frac{0.01}{1.5} \]

To find the value of the constant (equation [3]), we substitute the expression:

\[ B = \frac{0.01}{1.5} \]

The value of the constant is then used to determine the solution to the problem.
Experiment 1-8

Determination of the Formula of Tin Oxide

Apparatus and Materials: Crucible, ring stand and ring (or tripod), Bunsen burner, triangle, balance and weights, tin, nitric acid.

Object: To determine the formula of tin oxide.

Procedure: Weigh a clean dry crucible without cover to the nearest centigram. Add about 1 gram of pure tin and weigh again. The difference between these two weights will be the weight of the tin. Place the crucible on a triangle supported on a tripod or on a ring stand in the hood and add slowly 5 ml of nitric acid diluted with an equal volume of water. Heat the crucible very gently with a very low flame. Hold the burner in the hand, moving the flame back and forth under the crucible, and remove the flame instantly if the material shows any tendency to spatter. When the reaction seems to be over and the material is dry, gradually increase the heat. Finally heat the crucible as hot as possible for 10 minutes. Let the crucible cool to room temperature and weigh it. Add 5 drops of concentrated nitric acid and repeat the process until the weights are constant within 0.05 grams. Subtract the weight of the empty crucible from the weight of the crucible and contents after heating to find the weight of tin oxide formed. Subtract the weight of the tin from the weight of the tin oxide to obtain the weight of the oxygen by their respective atomic weights to obtain the number of gram atoms of each element which entered into combination. From these data, calculate the formula of tin oxide. Record
all the data in your notebook in a table similar to that in Experiment 1-7.

Experiment 2-1

**Filtration**

Obtain about 5 ml of muddy water in a beaker and filter half of it into a small test tube. Does the water come through clear? The answer to this question will depend on the size of the pores in the filter paper and on the size of the particles of solid in the muddy water. Large particles will be retained on the filter, but very fine particles will pass through into the filtrate.

To the remainder of the muddy water, add a small crystal of alum (potassium aluminum sulfate) and stir until it is dissolved. Pour the mixture into a small test tube and let it stand for a few minutes. Does the mud settle to the bottom? Shake up the mixture again and filter it. Does the filtrate come through clear or clearer than before? Alum is usually added to the water in municipal purification plants to assist in clearing the water. The alum reacts with the water to form a gelatinous precipitate of aluminum hydroxide, \( \text{Al(OH)}_3 \), which traps and holds the small pieces of mud. A reaction between a salt and water is called hydrolysis and the salt is said to be hydrolyzed.

Experiment 2-2

**Distillation**

In this experiment either a Liebig condenser or an air condenser can be used. A distilling flask of 25 ml,
null
50 ml, 100 ml, or even larger capacity may be used. The top is closed with a cork stopper and the side arm connected to the condenser through a bored cork stopper. If a regular distilling flask is not available, use a 50 ml Erlenmeyer flask with an air condenser as shown in the diagram. The air condenser is made by bending a 40 cm piece of glass tubing twice at right angles so that the short arm is about 6 cm long and the long arm is about 25 cm.

\[\text{Dissolve a small crystal of a highly colored salt (copper sulfate, potassium permanganate, potassium dichromate, etc.) in about } 15 \text{ ml of water in a beaker and introduce the solution into the distilling flask. (Adjust the volume to the capacity of the flask; the flask should be not over } \frac{1}{3} \text{ full.)}\]

Distill over about 3 ml of liquid, catching the distillate in a small test tube. Be careful that none of the solution itself boils over - it is the vapor which must pass over and be condensed. What evidence have you that the water has been purified?

Repeat the procedure using 15 ml of water and about 0.1 g sodium chloride. Add one drop of silver nitrate solution to 3 ml of the distillate and one drop of silver nitrate solution to the liquid remaining in the flask. Can salt water be purified by distillation? This test depends on the fact that soluble silver nitrate reacts with soluble
sodium chloride to form insoluble silver chloride. An insoluble substance produced in this way is called a precipitate.

Mix 15 ml of water and 1 ml of ammonia solution in a beaker. Pour 1 ml of the mixture into a small test tube and set it aside. Pour the rest into the distilling flask and distill over about 3 ml of the distillate as before. Ammonia is a very volatile substance; that is, it is easily changed to the vapor or gaseous state. The presence of ammonia vapor is easily detected by the odor.

Add one drop of phenolphthalein solution to the reserved portion of the original mixture and one drop to the distillate. Were you able to separate water from ammonia by distillation? This test depends on the fact that ammonia will turn phenolphthalein red or pink.

Experiment 2-3

Reaction of water with metals

Potassium and sodium when pure are soft, silver colored metals. They corrode so easily that they must be kept away from air either in a vacuum or under kerosene. DO NOT TOUCH EITHER METAL WITH THE FINGERS. Calcium is a hard, silver colored metal which is so much less corrosive than potassium or sodium that it can be handled. It should be protected from the air, however, in a tightly stoppered bottle. Potassium and sodium are lighter than water; calcium is heavier than water.

Obtain a small piece of potassium about the size of a
grain of rice. Hold it at arm's length with forceps and drop it into 10 ml of water in a 100 ml beaker. Stand four or five feet away from the beaker.

The potassium displaces half of the hydrogen from the water and combines with the oxygen and the other half of the hydrogen to form a compound known as potassium hydroxide.

\[ 2 \text{K} + 2 \text{H}_2\text{O} \rightarrow 2 \text{KOH} + \text{H}_2 \]

This reaction generates so much heat that the liberated hydrogen almost always catches fire and burns. Often the end of the reaction is characterized by an explosion. Pour the solution which now contains the potassium hydroxide into a labeled test tube.

Repeat the procedure with a small piece of sodium. A similar but somewhat less violent reaction occurs, with the formation of sodium hydroxide and hydrogen.

\[ 2 \text{Na} + 2 \text{H}_2\text{O} \rightarrow 2 \text{NaOH} + \text{H}_2 \]

The hydrogen may or may not catch fire from the heat of the reaction. Save the solution of sodium hydroxide.

Repeat the procedure with a small piece of calcium. Since the laboratory supply of calcium is often badly corroded, it may be advisable to place a large piece of calcium in a 600 ml beaker for a few minutes and then divide the solution among the members of the class. The calcium sinks to the bottom of the water and reacts to form calcium hydroxide and hydrogen.

\[ \text{Ca} + 2 \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2 \]
It is perfectly safe to put a piece of calcium as big as your fist into water; but the reaction of either potassium or sodium is so violent that only small pieces may be used. **Do not forget this.**

Lay a strip of red litmus paper on a clean dry watch glass. Dip a clean stirring rod in the solution of potassium hydroxide saved from the first part of the experiment above and transfer a drop of the solution to one end of the strip of litmus paper. Wash off the rod and in the same way transfer a drop of the sodium hydroxide solution to the middle of the strip of litmus. Again clean the rod and transfer a drop of the calcium hydroxide solution to the other end of the strip. Use only one strip of litmus. Describe the changes in color which occur.

Another substance which can be used to recognize hydroxides is phenolphthalein. Add one drop of a solution of phenolphthalein to each of the hydroxide solutions. What is the color?

Substances like litmus and phenolphthalein are called indicators. Solutions which turn litmus paper blue and phenolphthalein red are called basic or alkaline solutions; those which turn litmus paper red and phenolphthalein colorless are acidic.

**Experiment 2-4**

**Reaction of Water with Oxides**

(a) Obtain a very small piece of sodium about the size of a grain of rice and place it in a clean perfectly dry 20 cm test tube. Clamp the test tube in a horizontal
position. Direct the flame of a Bunsen burner held at arms length against the bottom of the test tube. The sodium will first melt (melting point 97.5°C) and then catch fire and burn to form sodium oxide. As soon as the sodium starts to burn, withdraw the flame. You will be wise to choose the smallest piece of sodium you can find because a large piece may burn explosively and spatter the hot corrosive substance out of the test tube. Make sure that the mouth of the test tube does not point at any one. Now add 10 drops of water to the sodium oxide and test the resulting solution with litmus and phenolphthalein. Result?

(When sodium burns in a limited amount of air, as in a test tube, sodium oxide, Na₂O, is formed. The sodium oxide will react violently with water to form sodium hydroxide.

\[ 4 \text{Na} + \text{O}_2 \rightarrow 2 \text{Na}_2\text{O} \]

\[ \text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{NaOH} \]

When sodium burns in an excess of air, as in an open evaporating dish, sodium peroxide, Na₂O₂, is formed.

\[ 2 \text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}_2 \]

The formula is written Na₂O₂ instead of NaO₂ because the "O₂" in this compound is a radical. The same radical occurs in hydrogen peroxide H₂O₂. Sodium peroxide is manufactured commercially by passing purified air over metallic sodium in aluminum trays at 300°C. The compound is sold under the trade name of "Oxone". It is sometimes used in the
laboratory to prepare oxygen.

\[ 2 \text{Na}_2\text{O}_2 + 2 \text{H}_2\text{O} \rightarrow 4 \text{NaOH} + \text{O}_2 \]

(b) Obtain a piece of magnesium ribbon about 2 cm long. Using forceps, hold the magnesium in the flame until it is ignited and then hold it over a clean evaporating dish containing 2 ml of water. Allow the white oxide to drop into the water. Warm the mixture slightly (ring stand, ring, wire gauze, Bunsen burner) for a moment and stir. Is magnesium oxide more or less soluble in water than sodium oxide? What compound is formed when magnesium oxide reacts with water? Test with litmus and phenolphthalein. Is this compound acidic or basic?

\[ 2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO} \]

\[ \text{MgO} + \text{H}_2\text{O} \rightarrow \text{Mg(OH)}_2 \]

(c) Place about 0.05 g of powdered black copper oxide in an evaporating dish and add 2 ml of water. Heat and stir just as you did in the case of the magnesium oxide. Does copper oxide seem to dissolve in water? Filter the solution into a clean 10 cm test tube and test the filtrate with litmus and phenolphthalein. Do you conclude that copper hydroxide, Cu(OH)_2, was formed or not? Why? Would you say that the reaction \( \text{Cu} + \text{H}_2\text{O} \rightarrow \text{Cu(OH)}_2 \) is complete, partially complete, or practically incomplete?

(d) Place 0.5 g calcium oxide in an evaporating dish, add 10 drops of water, and stir. Describe the reaction. The common name for calcium oxide is quicklime, or simply lime. Its reaction with water is known as slaking and the product, calcium hydroxide, is slaked lime.
\[(\text{H}_2 + \text{H}_2 \text{O}) \xrightarrow{\text{Catalyst}} \text{H}_2 \text{O}_2 + \text{H}_2 \text{O} \]

This reaction can be represented as:

\[\text{H}_2 + \text{H}_2 \text{O} \xrightarrow{\text{Catalyst}} \text{H}_2 \text{O}_2 + \text{H}_2 \text{O} \]

The catalyst used in this reaction is often referred to as a "link" or a "bridge" in the chemical context. The reaction occurs under specific conditions, such as temperature and pressure, to ensure the desired outcome. The products of this reaction can be further utilized in various applications, including the production of hydrogen peroxide and the conversion of water into hydrogen and oxygen. The catalyst plays a crucial role in accelerating the reaction rate and improving the efficiency of the process.

In summary, the reaction described above is a fundamental process in chemical engineering and has significant applications in different fields, including energy production and environmental management.
\[ \text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \]

Continue to add water drop by drop until no further reaction occurs or until you are sure that the lime has been slaked. Add 2 ml of water and stir. The thick milky mixture is known as "milk of lime". Add 10 ml of water and filter. The filtrate, a solution of calcium hydroxide, is limewater. Test the filtrate with litmus and phenolphthalein. Result?

(e) Place 30 ml of water in a 60 ml bottle. Ignite a little sulfur in a semi-micro deflagration spoon and lower the burning sulfur into the bottle. Do not let the hot spoon touch the water. Cover the bottle as well as possible with a glass plate to confine the fumes of burning sulfur. Remove the spoon, cover the bottle tightly with a glass plate, and shake the bottle well to try to dissolve the sulfur dioxide. Describe the burning of sulfur. Does the solution give an acidic or an alkaline reaction with indicators?

Sulfur burns in air to form sulfur dioxide and the oxide reacts with water to form sulfurous acid, \( \text{H}_2\text{SO}_3 \).

\[
\text{S} + \text{O}_2 \rightarrow \text{SO}_2
\]

\[
\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3
\]

(f) Place a very small piece of red phosphorus (certainly no larger than a grain of rice) in a 20 cm Pyrex test tube supported vertically by a ring stand and clamp and cover the test tube with a small watch glass. Heat the phosphorus cautiously with a small flame until it catches fire. Allow the tube to cool, add 5 ml of water,
\[ g(x) = \frac{e^x}{x} \]

\[ \log x = \ln x \]

\[ \text{contemplate to the nature with the mind of man} \]

\[ \text{interpret according to the nature of life and the mind of man} \]

\[ \text{render to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]

\[ \text{interpret to the nature of life and the mind of man} \]
and shake gently. Does the solution give an acidic or an alkaline reaction with indicators?

**Phosphorus burns in air to form phosphorus pentoxide** $P_2O_5$. The oxide reacts with water to form phosphoric acid.

$$4 \text{ P} + 5 \text{ O}_2 \rightarrow 2 \text{ P}_2\text{O}_5$$

$$\text{P}_2\text{O}_5 + 3 \text{ H}_2\text{O} \rightarrow 2 \text{ H}_3\text{PO}_4$$

**Experiment 2-5**

**Experimental study of hydrates**

Place a small, clear, glassy crystal of washing soda on a dry watch glass, label it with the formula $\text{Na}_2\text{CO}_3 \cdot 10 \text{ H}_2\text{O}$ and your name, and set it aside. The next day examine the crystal to see if there is any indication that it has lost water of crystallization by efflorescence. Describe any changes you observe.

Follow the same procedure with a small crystal of gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, a small piece of anhydrous calcium chloride, $\text{CaCl}_2$, and a small pellet of sodium hydroxide, $\text{NaOH}$. Describe any changes you observe. Which of these substances are efflorescent, hygroscopic, or deliquescent? Which one of these substances seems to have neither lost or gained water?

Place a small clear crystal of washing soda in a 10 cm test tube and heat it gently. What evidence have you that water was driven off by the heat? This part of the experiment indicates that if a substance is a hydrate, it will decompose on heating. The crystal will crumble to a powder, and water will condense at the
top of the tube. Be sure not to heat the top of the tube. Why not? A mere trace of water does not indicate a hydrate, however, for even an anhydrous substance may contain a slight amount of water as an impurity. Such water is said to be "mechanically held".

In a similar way, examine several other crystals to determine whether or not they are hydrates. Heat them in separate dry test tubes and observe the changes which occur. Record the results in your notebook in tabular form. Suitable crystals for examination are potassium chromate, potassium chlorate, potassium nitrate, potassium alum, sodium sulfate, copper sulfate, zinc sulfate, gypsum, etc. Report the amount of water as "little", "much", or "a trace". Record the data in tabular form.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Appearance</th>
<th>Amount</th>
<th>Was substance heated</th>
<th>of residue of water a hydrate?</th>
</tr>
</thead>
</table>

Add five drops of water to the anhydrous copper sulfate prepared as above. Notice the change in color which occurs. Hydrated copper sulfate is blue; the anhydrous salt is nearly white. The addition of water to the dehydrated product will reform the hydrate, which can be recognized by its blue color. This color change is often used as a test for water. Do you think the test would show whether the water were pure or not? If a certain liquid is said to be 100% alcohol, it can be tested for water with anhydrous copper sulfate. If the alcohol contains even a little water, the blue color will be obtained; but if the alcohol is anhydrous, there will
be no change in the color of the copper sulfate. Obtain a sample of alcohol (or some other liquid) and determine the presence or absence of water. If water is present, is there much or only a little?

Experiment 2-6

Water of Hydration in Crystallized Barium Chloride

Weigh a clean dry porcelain crucible to the nearest centigram. Add about one gram of barium chloride crystals and weigh again. Record these weights immediately in your notebook. Place the crucible on a triangle supported by a tripod or ring stand and ring and heat it gently with a low flame for about ten minutes. The heat must be very low at first to prevent the loss of small particles by decrepitation (the violent decomposition of the crystals). The full heat of the Bunsen burner is not needed, but the temperature should be at least $150^\circ C$. If an oven which can be heated to $150^\circ C$ is available, more satisfactory results will be obtained if the crucible is heated in the oven for about two hours. (A longer time will do no harm.)

Let the crucible cool to room temperature and weigh. Heat for ten minutes more over the open flame (or for another hour in the oven) and weigh again. The two weights must agree within 0.05 g. Continue heating until constant weight is obtained. Record all weights in your notebook.

Table of Data

| Weight of crucible and barium chloride | _______ g |
| Weight of empty crucible              | _______ g |
Weight of barium chloride \[\_\_\_\_\_\_\_g\]

Weight of crucible and contents after \[\_\_\_\_\_\_\_g\]

first heating, after second heating, etc.

Loss of weight on heating = weight of \[\_\_\_\_\_\_\_g\]

water of hydration

Percent water of hydration \[\_\_\_\_\_\_\_g\]

Other hydrates may be used in this experiment.

Gypsum can be dehydrated at 163°C. Copper sulfate crystals can be dehydrated at 250°C, but they must not be heated above 650°C because the CuSO\(_4\) will begin to decompose into CuO at that temperature.

Experiment 2-7

Saturated and Supersaturated Solutions

In this experiment we shall prepare solutions of sodium thiosulfate, Na\(_2\)S\(_2\)O\(_3\). This salt is widely used in photography under the name of "hypo".

Place 10 drops of water in a 10 cm test tube and add small crystals of hypo, one by one, with shaking or stirring, until no more will dissolve and the solution is saturated. How can you tell when a solution is saturated?

Now warm the solution gently and add more hypo, crystal by crystal, with frequent shaking or stirring. Continue to heat and add hypo until a hot saturated solution is obtained or until the volume of the solution nearly fills the test tube. The temperature of this solution should be about 80-90 degrees. Although it is not necessary to measure this temperature, it could be done by placing the test tube in a beaker of water and warming this water over a flame.
(ring stand, ring, wire gauze). From time to time, remove the flame, stir the water in the beaker with a rod, and measure its temperature with a thermometer. Why stir the water before inserting the thermometer?

Fit a circle of filter paper to a 32 or 40 mm funnel in the usual way and run hot water through it to heat it and to make sure that it is filtering well. Loosen the paper for a moment so that the water in the stem of the funnel will run out, then readjust the paper. Now filter half the solution into one dry test tube and the other half of the solution into another dry test tube. Cool these two tubes by standing them in a 100 ml beaker containing cold water. Since these solutions were saturated at a high temperature, they will be supersaturated when they cool.

Stir the solution in the first test tube. Drop a very small crystal of hypo into the second test tube. Observe carefully all that happens in each case. Describe and explain the results.

To prevent waste, all the crystals of hypo obtained and all the hypo solutions remaining should be emptied into a large beaker provided by the teacher for this purpose. The teacher should then see that the water is evaporated from this beaker and that the resulting crystals may then be dried in an oven or over a low flame. The dried crystals may then be returned to the stock bottle.

Repeat the entire experiment using crystals of some other salt suggested by the teacher. Sodium acetate, potassium nitrate, sodium sulfate, potassium chlorate,
lead nitrate, and copper sulfate are suitable. Do these salts form supersaturated solutions, or does the excess solute crystallize gradually as the solution cools?

Experiment 2-8

Factors Affecting the Rate of Solution

Select five crystals of copper sulfate about the size of a grain of rice and all about the same size. Add 2 ml of water to each of five test tubes. Cut a strip of paper about 5 mm wide and about 15 cm long, fold it in the shape of a letter "M" and place it over the first test tube so that the middle "V" part of the "M" dips just below the surface of the water in the test tube.

Powder one of the crystals in a mortar or obtain a small quantity of copper sulfate already powdered. Note the time by the clock. Carefully place a crystal of copper sulfate on the paper shelf in the first test tube. Drop another crystal into a second tube and let it lie on the bottom. Introduce the powdered copper sulfate into the third test tube and let this tube stand quietly. Drop the next crystal into the fourth test tube, place your thumb over the mouth of the tube, and shake it until complete solution is obtained. Note the time required to effect complete solution in each case. Drop another crystal into the fifth tube and heat gently. Note the time required.
If you wish to dissolve a given amount of a substance as quickly as possible, what method or combination of methods would you select. Take another crystal of the same size as before and treat it by the procedure you recommend. Compare the time required with the times previously noted.

Experiment 2-9

Solubility in Various Solvents

To determine the solubility of a substance, add a very small crystal of a solid substance (about 0.1 g) or about 4 drops of a liquid (0.2 ml) to 2 ml of the solvent in a 10 cm test tube. In the case of solids, powder, heat, or shake to effect solution. Do not heat organic solvents, however, because of the danger of fire. Ether is an especially flammable substance and must not be used on the same laboratory bench with a lighted burner. Ether vapor has caused many serious fires. These solvents can be heated by placing the test tubes containing them in a beaker of previously heated water, but the water must not be heated anywhere near the solvents themselves.

Determine the solubility of several substances such as fat, cottonseed oil sugar, iodine, naphthalene, calcium chloride, alcohol, ether, acetone, aniline, carbon tetra-chloride, etc. in such solvents as water, alcohol, ether, acetone, and carbon tetrachloride. Record the results in the form of a table. Describe the solubility as "insoluble", "very soluble", or "slightly soluble".
Table of Data

<table>
<thead>
<tr>
<th>Solutes</th>
<th>Water</th>
<th>Alcohol</th>
<th>Acetone</th>
<th>Etc.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fats</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cottonseed oil</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sugar</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Etc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Experiment 2-10

Miscible and Immiscible Liquids

Liquids which will dissolve completely in one another are called miscible liquids; liquids which will not dissolve completely in each other are called immiscible.

To determine the miscibility of two liquids, place two drops of each on a clean, dry microscope slide and try to blend them with a clean, dry, slender stirring rod or with a piece of nichrome wire. Try several combinations such as oil and water, benzene and alcohol, water and alcohol, water and ether, turpentine and kerosene, etc. If possible, examine the mixture under the microscope. Do not repeat combinations already investigated in the previous experiment.

Instead of trying to blend the two liquids on a slide, you may try to mix 1 ml of each liquid in a small test tube. Do the liquids mix or do they form two separate layers?

Arrange your results in the form of a table,
<table>
<thead>
<tr>
<th>Date</th>
<th>Name</th>
<th>Reason</th>
<th>Note</th>
</tr>
</thead>
<tbody>
<tr>
<td>1937</td>
<td>Jc1</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Resolutions**

The board of directors hereby resolves to approve the following resolutions:

1. That the company be authorized to enter into an agreement with [Company Name] for the supply of [Product/Service].
2. That the terms of the contract be as follows: [Insert terms].
3. That the execution of the contract be overseen by [Name].

The resolutions shall be submitted to the shareholders for ratification at the next annual meeting.
Experiment 2-11  
**Solubility of Salts in Water**

Place 3 ml of distilled water in a 10 cm test tube and add powdered calcium sulfate (gypsum) until a saturated solution is obtained. Does it appear that much calcium sulfate had dissolved? Filter the solution, catching the filtrate in a perfectly clean evaporating dish. Evaporate the filtrate to dryness over a small flame. Is there a residue? Why should distilled water and not tap water be used?

Repeat the experiment, using magnesium sulfate (Epsom salts). Do you obtain a residue when the filtrate is evaporated?

Repeat the experiment, using sodium chloride or some other salt.

Compare the solubility of the salts used.

Experiment 2-12  
**Determination of the Solubility of a Salt in Water**

Warm 20 ml of water in a 100 ml beaker and add with constant stirring a little more than enough potassium nitrate (or some other convenient salt) to saturate the solution. Allow the solution to cool to room temperature. How can you be sure that the solution is not supersaturated? Filter the solution into an evaporating dish and determine the temperature with a thermometer.
Weight the dish and the solution carefully and evaporate the solution just to dryness, taking care that it does not spatter. Do not heat potassium nitrate after it is dry or it will decompose. Weigh the dish and the dry salt. Empty all the potassium nitrate into a beaker provided for the purpose; then clean and weigh the evaporating dish.

Since the solubility is defined as the weight of a substance dissolved in 100 grams of water at a definite temperature, we can calculate the solubility of the potassium nitrate from our data by the use of the proportion

\[
\frac{\text{weight of water used}}{\text{weight of salt used}} = \frac{100}{X}
\]

### Table of Data

<table>
<thead>
<tr>
<th>Description</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Weight of dish + solution</td>
<td></td>
</tr>
<tr>
<td>Weight of dish + dry salt</td>
<td></td>
</tr>
<tr>
<td>Weight of empty dish</td>
<td></td>
</tr>
<tr>
<td>Weight of solution</td>
<td></td>
</tr>
<tr>
<td>Weight of solution</td>
<td></td>
</tr>
<tr>
<td>Weight of salt</td>
<td></td>
</tr>
<tr>
<td>Weight of water used to dissolve this weight of salt</td>
<td></td>
</tr>
<tr>
<td>Temperature of the solution</td>
<td></td>
</tr>
<tr>
<td>Solubility of potassium nitrate at _____ °C</td>
<td></td>
</tr>
</tbody>
</table>
In the given problem, we are asked to solve for the value of the unknown variable $x$ in the equation:

$$\text{Percentage of Error} = \frac{\text{Observed Value} - \text{Expected Value}}{\text{Expected Value}}$$

We are given the following data:

<table>
<thead>
<tr>
<th>Variable</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>(1)</td>
</tr>
<tr>
<td>B</td>
<td>(2)</td>
</tr>
<tr>
<td>C</td>
<td>(3)</td>
</tr>
<tr>
<td>D</td>
<td>(4) = (D_1 - D_2)</td>
</tr>
<tr>
<td>E</td>
<td>(5) = (E_1 - E_2)</td>
</tr>
<tr>
<td>F</td>
<td>(6) - (C_1)</td>
</tr>
</tbody>
</table>

The equation to find $x$ is:

$$x = \frac{(100 - x) + (20 - x) + (30 - x)}{x}$$

Solving for $x$, we get:

$$x = \frac{150 - 3x}{2}$$

Rearranging, we get:

$$2x = 150 - 3x$$

$$5x = 150$$

$$x = 30$$
Appendix B

Since students in chemistry are expected to be more or less familiar with the metric system and certain topics from physics, this material was prepared in a supplementary chapter and made available to the class. Very little was done with this chapter with this particular class, but it would be very important in a college preparatory class. This is the material referred to in Optional Related Activities for Unit I, Item 5.

<table>
<thead>
<tr>
<th>UNIT</th>
<th>MILLIMETER</th>
<th>CENTIMETER</th>
<th>MILLIMETER</th>
<th>MILLIMETER</th>
</tr>
</thead>
<tbody>
<tr>
<td>10</td>
<td>10</td>
<td>100</td>
<td>100</td>
<td>1000</td>
</tr>
<tr>
<td>100</td>
<td>1000</td>
<td>10000</td>
<td>10000</td>
<td>100000</td>
</tr>
<tr>
<td>1 dekameter = 10 meters</td>
<td>1 meter = 10 centimeters</td>
<td>1 centimeter = 10 millimeters</td>
<td>1 millimeter = 10 thousandths</td>
<td></td>
</tr>
</tbody>
</table>
Experiments in one region of the economy and some successful economic policies were followed by a decline in the other. Such a process may continue and eventually lead to a collapse of the entire economic system.
The Metric System:

The system of weights and measures used in all scientific work is the metric system. Its outstanding advantage is that it is a decimal system like the currency system of the United States.

The unit of length is the meter. It is defined as the distance at zero degrees Centigrade (the temperature of melting ice) between two hair lines on a platinum-iridium bar kept at the Bureau of International Weights and Measures near Paris, France. The United States Government has a certified copy of this standard bar.

The meter is 39.37 inches long. The meter is divided into 10 decimeters (dm); each decimeter is divided into 10 centimeters (cm); and each centimeter is divided into 10 millimeters (mm).

\[1 \text{ m} = 10 \text{ dm} = 100 \text{ cm} = 1000 \text{ mm}\]

It is useful to remember that 1 inch equals 2.54 centimeters. The following table may help to make these relationships clear.

<table>
<thead>
<tr>
<th>Meter</th>
<th>Dollar</th>
</tr>
</thead>
<tbody>
<tr>
<td>10 decimeters</td>
<td>10 dimes</td>
</tr>
<tr>
<td>100 centimeters</td>
<td>100 cents</td>
</tr>
<tr>
<td>1000 millimeters</td>
<td>1000 mills</td>
</tr>
<tr>
<td>1 decimeter = 10 centimeters</td>
<td>1 dime = 10 cents</td>
</tr>
<tr>
<td>1 centimeter = 10 millimeters</td>
<td>1 cent = 10 mills</td>
</tr>
</tbody>
</table>
The table below shows the data that is relevant to this study. The percentages are calculated based on the total number of cases. The following data can be used for future reference.
The unit of mass (weight) is a mass of platinum iridium kept at the International Bureau. The United States has a certified copy. This mass is called a kilogram, (kg). The kilogram equals 2.2 pounds. The kilogram is first divided into 1000 grams. Then the gram (g) is divided into 10 decigrams (dg); each decigram is divided into 10 centigrams (cg); and each centigram is divided into 10 milligrams (mg).

\[1 \text{ kg} = 1000 \text{ g}; \quad 1 \text{ g} = 10 \text{ dg} = 100 \text{ cg} = 1000 \text{ mg}\]

The term "mass" is a little more exact than the term "weight". In general, "mass" means the quantity of matter in a body and "weight" means the pull of gravity on this mass. Obviously the quantity of matter in a body will always be the same, but the pull of gravity on a body varies slightly from place to place.

The unit of volume is the liter (l). The liter is the volume of one kilogram of water at 4°C. One liter equals 1.06 quarts. The liter is divided into 1000 milliliters (ml). Another set of volume units is derived from the units of length: the cubic meter (m³), the cubic decimeter (dm³), the cubic centimeter (cm³ or cc), and the cubic millimeter (mm³).

A liter is almost exactly equal to a cubic decimeter. Since a cubic decimeter is a cube measuring 10 centimeters on each edge, its volume is 1000 cubic centimeters. A milliliter, therefore, is almost exactly equal to one cubic centimeter (1 ml = 1.000027 cc). The volumes of liquids and gases in liters and milliliters.
Units of Temperature and Heat

The unit of temperature is the degree Centigrade. On a Centigrade thermometer, water freezes at 0°C and boils at 100°C. On the familiar Fahrenheit thermometer scale, water freezes at 32° and boils at 212°. Temperatures on the two scales are related by the equation

\[ F° - 32 = 1.8 \ C° \]

Standard room temperature is 68°F. To determine the Centigrade equivalent, substitute 68 for "F" in the above equation and solve for C.

\[ 68 - 32 = 1.8 \ C \]
\[ 36 = 1.8 \ c \]
\[ C = 20° \]

Chemical glassware is graduated to contain the volume designated at 20°C. Many chemical measurements are also made at 25°C. What would this temperature be on the Fahrenheit scale?

The unit of heat is the calorie. The calorie is the amount of heat required to raise the temperature of one gram of water 1°C. The number of calories required to raise the temperature of a substance 1°C is called its specific heat. The specific heat of water is one calorie; the specific heat of most other substances is less than one.
Table of Specific Heats

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>1 calorie</td>
</tr>
<tr>
<td>Ice</td>
<td>0.50 calories</td>
</tr>
<tr>
<td>Steam</td>
<td>0.50 calories</td>
</tr>
<tr>
<td>Air</td>
<td>0.24 calories</td>
</tr>
<tr>
<td>Aluminum</td>
<td>0.22 calories</td>
</tr>
<tr>
<td>Iron</td>
<td>0.11 calories</td>
</tr>
<tr>
<td>Zinc</td>
<td>0.094 calories</td>
</tr>
<tr>
<td>Copper</td>
<td>0.093 calories</td>
</tr>
<tr>
<td>Silver</td>
<td>0.056 calories</td>
</tr>
<tr>
<td>Mercury</td>
<td>0.033 calories</td>
</tr>
<tr>
<td>Lead</td>
<td>0.031 calories</td>
</tr>
<tr>
<td>Glass</td>
<td>0.20 calories</td>
</tr>
</tbody>
</table>

Example: How much heat would be required to raise the temperature of 40 grams of aluminum from 17°C to 53°C?

Rule (from the definition): Heat in calories = mass x change in temperature x specific heat; or, cal = m t s.

\[ x = 40 \times (53 - 17) \times 0.22 \]
\[ x = 316.8 \text{ calories}. \]

Density and Specific Gravity

An important property of every substance is its density. Density is defined as the mass of a unit volume:

\[ D = \frac{M}{V} \]

The density of water is 1 g/ml (one gram per milliliter) or almost exactly 1 g/cc. In the English system of measurement, the density of water is 62.4 lbs/ft³. The density of sulfur is 2 g/cc; the density of air is 1.293 g/l; the density of copper is 8.92 g/cc; and the density of ether is 0.708 g/ml. (Explain the reason for the difference in the labels.) To obtain the density of a block of wood, you might measure the length, the width, and the thickness. The product of these three dimensions is the volume. Then weigh the block and divide the weight by the volume.
**Example:** A block of nickel is 15 centimeters long, 8 centimeters wide, and 2 centimeters thick. It weighs 2140 grams. What is the density of nickel?

\[ D = \frac{2140}{15 \times 8 \times 2} = 8.9 \text{ g/cc} \]

The term "specific gravity" is often used to mean "times as heavy as water". To get the specific gravity of any substance, divide its density by the density of water. In the metric system, density and specific gravity will be **numerically the same** (Why?); in the English system, the specific gravity will be numerically equal to the density divided by 62.4. Note that density must always have a label (g/cc, g/ml, g/l, lbs/ft³, etc.), but that specific gravity is just a number.

**Experiment 1-21**

**Used of the Metric System in Chemistry**

**Apparatus and Materials:** Wooden block, porcelain crucible, water, metric ruler, platform balance, pan balance, small test tube, graduated cylinder, Centigrade thermometer, set of weights.

**Object:** To become acquainted with some of the units of measurement used in chemistry.

**Procedure:** Follow the directions given below. Record your procedure in theme form in your notebook. Record all data obtained in tabular form.

Draw a line about three inches long in your notebook and measure it to the **nearest** hundredth of a centimeter.

Since the end of a ruler may be worn, the zero mark may not
Once the area is known as to avoid a constant, the area is
measured to the nearest 1/1000th of an inch.

\[
\text{Area} = \frac{\text{Width} \times \text{Length}}{1000^2}
\]

• To have a neat "spelling" of the word "mater" in the

• The area of a square with a side length of 1.2 inches is

• The area of a circle with a radius of 3 cm is

\[
\text{Area} = \pi r^2
\]

• If the length of a rectangle is 12 cm and the width is 5 cm,

\[
\text{Area} = \text{Length} \times \text{Width}
\]

• The area of a parallelogram with a base of 10 cm and a

• The area of a triangle with a base of 8 cm and a

\[
\text{Area} = \frac{1}{2} \times \text{base} \times \text{height}
\]

• The area of a trapezoid with bases of 12 cm and 8 cm and

• The area of a kite with diagonals of 10 cm and 8 cm is

\[
\text{Area} = \frac{1}{2} \times \text{D_1} \times \text{D_2}
\]

• The area of a regular polygon with a side length of 5 cm

\[
\text{Area} = \frac{1}{2} \times \text{Perimeter} \times \text{Apothem}
\]

• The area of a sector of a circle with a radius of 10 cm and

\[
\text{Area} = \frac{\theta}{360} \times \pi r^2
\]

• The area of a hemisphere with a radius of 5 cm is

\[
\text{Area} = 2 \pi r^2
\]

• The area of a sphere with a radius of 4 cm is

\[
\text{Area} = 4 \pi r^2
\]
be exact. It is preferable, therefore, to place the "1" mark at the end of the line to be measured. Suppose that the other end of the line falls between the 3.5 and the 3.6 marks on the ruler. We can then estimate the exact position of the end of the line to the nearest hundredth. We may say, for example, that it falls at exactly 3.57. The total length of the line is then 3.57 - 1, or 7.57 centimeters. Do not forget to subtract the 1.

Weigh your chemistry book on the platform balance to the nearest gram. Record this weight both in grams and in kilograms. Weigh the porcelain crucible on the pan balance to the nearest centigram.

Detailed directions for the use of the balance depend to some extent on the kinds of balances available, so the teacher will show you how to use the balances in your laboratory. Pay very careful attention to this instruction and demonstration because the balance is one of the most important tools for the chemist. In fact, a chemist is often judged by the way he uses a balance. In general, however, follow these rules:

1. The pans must always be supported when adding or removing weights or objects.

2. Place the object to be weighed in the left pan and the weights in the right pan. Always start with a weight which is too heavy. Then remove this weight and try the next smaller one. Leave weights that are too light in the pan; remove weights that are too heavy.
be able to explain the reasons for taking any action that may be necessary to ensure the safe and effective use of the equipment. It is important to consult the manual and follow the instructions carefully. In case of doubt, it is advisable to seek assistance from a qualified technician or contacted the manufacturer directly.
3. Never weigh an object while it is hot.
4. Never touch weights with the fingers; always use forceps.
5. Always keep weights in the box except when in use; never leave weights on the balance or on the table.
6. Never make adjustments on the balance; if anything goes wrong, ask the teacher about it.

Measure and record the length of the test tube to the nearest tenth of a centimeter. Fill the test tube with water and determine the volume of this water by measuring it with the graduated cylinder. Pour the water into the cylinder. Note that the surface of the water in the graduate is not quite horizontal but curves upward around the edges to form a meniscus (from a Greek word meaning crescent moon). Read the bottom of the meniscus to the nearest tenth of a milliliter. Measure out one milliliter of water by means of the graduated cylinder and pour it into a test tube. Measure the height to which the water fills the test tube.

Let the faucet water run until it is as cold as possible. Determine its temperature by means of the Centigrade thermometer. Also determine the temperature of the room in °C.

Measure the length, width, and thickness of the wooden block to the nearest tenth of a centimeter. Weigh the block on the platform balance to the nearest tenth of a gram. Compute the density of the block. Find the specific gravity of the block.
The Introduction

The views here presented to the professor may be
never have appeared on the published on the printed.

We have never been instructed on the printed to
in the printed of the printed to

water into the sandy. Have just the same
of the water in the printed to the water. Note the
change in the sand and 5000 to form a

(1) How a water from washing a soapstone can
form a matter. Matter of the washing to the water to make a
matter. 2. The water to make a soapstone can
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become. Resemblance the formulation of the frame of the
Central the formulation. This formulation the frame of the

are look to be of the

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March 20, 1950 to the 1950 paper on a conference. Analyze the
the paper on the 1950 conference to the manner of the
a look at the 1950 conference. Look at the

Editorial Remarks at the Front
<table>
<thead>
<tr>
<th>Table of Data</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length of line drawn in notebook ________ cms.</td>
</tr>
<tr>
<td>Weight of chemistry book ________ g = ________ kg.</td>
</tr>
<tr>
<td>Weight of crucible _____________ g = ___________ cg.</td>
</tr>
<tr>
<td>Length of test tube _____ cms; volume of test tube ____ ml.</td>
</tr>
<tr>
<td>Height to which 1 ml of water fills the test tube ______ cm.</td>
</tr>
<tr>
<td>Temperature of the water ____°C; temperature of the room ______°C.</td>
</tr>
<tr>
<td>Length of the wooden block ____ cm; width _____ cm; thickness _____ cm; volume ____ cm³; weight ______ g.</td>
</tr>
<tr>
<td>Density of the block ___________ g/cm³; specific gravity ______.</td>
</tr>
</tbody>
</table>

**Significant figures**

Suppose that when we measured the wooden block in the above experiment we obtained the following dimensions; length, 12.48 cm; width, 6.31 cm; thickness 5.92 cm. Since the second decimal place is estimated from the position occupied by the end of the block between two of the lines on the ruler, it might possibly be in error by at least one-tenth of the distance between those lines; that is, the length might have been either 12.47 or 12.48 cm; the width might have been either 6.30 or 6.32 cm, and the thickness might have been either 5.91 or 5.93 cm. In the first measurement, the length, we are sure of three figures, but in the other two measurements, the width and the thickness, we are sure of only two. If we compute the volume from the data recorded we obtain

\[ 6.31 \times 12.48 \times 5.92 = 466.192396 \text{ cm}^3 \]

If we stop and think a moment, however, we see that it is absurd to expect that we can obtain a value for the
volume correct to six decimal places from measurements which in two cases might have been in error in the third digit. If we assume that all the dimensions were actually smaller by 0.01 cm, the volume would have been

$$6.30 \times 12.47 \times 5.91 = 464.295510 \text{ cm}^3$$

and if we assume that the dimensions were actually larger by 0.01 cm, the volume would have been

$$6.32 \times 12.49 \times 5.93 = 468.095224 \text{ cm}^3$$

Notice that these three values for the volume agree exactly in the first two figures but that they do not agree in the third figure. Is it not obvious that the decimals are absolutely meaningless? Can you see that the answer is no more accurate than the observed data? We say that we measured the length to four significant figures and the width and thickness to three significant figures. The above calculations indicate that the answer in a problem of this sort is accurate to the same number of significant figures as the least number of significant figures in the data. From the data assumed in this section, we should have recorded the volume as 466 cm$^3$. Even if we had made no errors of measurement greater than 0.01 cm, all we can tell about the volume is that it lies between 464 and 468 cm$^3$.

Experiment 1-22

Density of sulfur

Apparatus and Materials: Lump of roll sulfur, water, overflow can and bucket, platform balance, thread.

Object: To determine the density of sulfur in g/cm$^3$. 
Notice that your terminology for the volume of the cone is incorrect. To calculate the volume of a cone, you need to use the formula $V = \frac{1}{3}\pi r^2 h$. The correct calculation for your cone would be:

$$V = \frac{1}{3}\pi (6)^2 (8) = \frac{1}{3}\pi (36) (8) = 96\pi \text{ cubic units}$$

When calculating volumes, it's important to use the correct formula and units. In this case, the volume of the cone is approximately $301.59$ cubic units.
Procedure: Weigh the lump of sulfur and record its weight in a table of data in your notebook. Place the bucket below the spout of the overflow can and pour water into the can until the water runs out the spout. When the excess water has run out of the can, empty the bucket, wipe it dry, and weigh it. Record the weight of the empty bucket in the table of data. Then place the dry empty bucket back under the spout. Now tie a piece of thread around the lump of sulfur and lower it carefully into the overflow can. Catch in the bucket all the water which runs out of the can. This water should have exactly the same volume as the lump of sulfur because it has been displaced by the sulfur. Now weigh the bucket with the water in it. Record this weight.

To determine the weight of the displaced water, subtract the weight of the empty bucket from the weight of the bucket with the water in it. Since one gram of water has a volume of almost exactly one cubic centimeter, the weight of the water in grams will be numerically equal to the volume of the water in cubic centimeters. This volume is also the volume of the sulfur. Divide the weight of the sulfur by its volume to determine the density. Then divide the density of sulfur by the density of water to obtain the specific gravity. (Remember that density has a label, but that specific gravity, being a ratio, is a pure number.)
Table of Data

<table>
<thead>
<tr>
<th>Description</th>
<th>Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>Weight of sulfur</td>
<td>g</td>
</tr>
<tr>
<td>Weight of bucket and water</td>
<td>g</td>
</tr>
<tr>
<td>Weight of empty bucket</td>
<td>g</td>
</tr>
<tr>
<td>Weight of water displaced</td>
<td>g</td>
</tr>
<tr>
<td>Volume of the displaced water</td>
<td>g</td>
</tr>
<tr>
<td>Volume of the sulfur</td>
<td>g</td>
</tr>
<tr>
<td>Density of sulfur</td>
<td>g</td>
</tr>
<tr>
<td>Specific Gravity of sulfur</td>
<td>g</td>
</tr>
</tbody>
</table>

(Note: If an overflow can and bucket is not available, the volume of the sulfur can be determined by placing the sulfur in a large graduate about half full of water and noting the increase in volume.)

Experiment 1-23

**Specific Gravity of a Liquid**

**Apparatus and Materials:** Balance and weights, pycnometer, water, some liquid such as acetone, alcohol, carbon tetrachloride.

**Object:** To determine the specific gravity of a liquid by means of pycnometer.

**Discussion:** The density or specific gravity of a liquid is an important physical property frequently determined for identification or characterization of a material. The measurement is usually made in an apparatus known as a pycnometer. This is a glass vessel with a capillary neck in which a definite volume of a liquid is weighed. Sometimes the stem contains a thermometer.
Specific gravity can be determined by dividing the weight of a definite volume of a liquid at a given temperature by the weight of the same volume of water at a given temperature. These temperatures are usually, but not necessarily, the same and should be specified. Sp. Gr. 20°/4°C means that the liquid was weighed at 20°C and the water at 4°C. Note that Sp. Gr. 20°C/4°C is numerically equal to density in g/ml because by definition 1 ml of water at 4°C weighs 1 g. The density of water becomes slightly less as the temperature increases, but the change is slight. For example,

\[
\text{Density of a liquid at 20°C} = \text{Sp. Gr. 20°C/4°C} = \text{Sp. Gr. 20°C/20°C} \times 0.9982
\]

Procedure: Clean and dry the pycnometer. Weigh it to the nearest 0.01 g. Fill the pycnometer with freshly boiled and cooled distilled water and insert the stopper. Wipe off all surplus water with a clean cloth and reweigh. The difference in weight is the weight of the water. Note and record the temperature of the water.

Dry the pycnometer carefully and fill it with the liquid, the density of which is to be determined. Avoid the inclusion of air bubbles. Insert the stopper, wipe off the outside of the vessel with a clean cloth, and weigh. Record the temperature of the liquid. (The pycnometer may be dried by rinsing in alcohol, allowing it to drain, and placing it in a warm place or blowing air through it. Do not place it in an oven.)
Table of Data

<table>
<thead>
<tr>
<th>Description</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Weight of pycnometer + water</td>
<td>g</td>
</tr>
<tr>
<td>Weight of pycnometer empty</td>
<td>g</td>
</tr>
<tr>
<td>Weight of water</td>
<td>g</td>
</tr>
<tr>
<td>Temperature of the water</td>
<td>°C</td>
</tr>
<tr>
<td>Weight of pycnometer + liquid</td>
<td>g</td>
</tr>
<tr>
<td>Weight of pycnometer empty</td>
<td>g</td>
</tr>
<tr>
<td>Weight of liquid</td>
<td>g</td>
</tr>
<tr>
<td>Temperature of liquid</td>
<td>°C</td>
</tr>
<tr>
<td>Sp. Gr. at °C/°F</td>
<td></td>
</tr>
</tbody>
</table>
Questions

1. What are the metric standards of length, mass, and volume? Define each and state its relation to a corresponding unit in the English system.

2. Compare the sub-divisions of the metric units with the American dollar, dime, cent, and mill. What are the advantages of a decimal system?

3. A sheet of paper measures 8.0 by 10.5 inches. If the length and width are calculated in centimeters, in what unit would the area be expressed? Find the area in these units.

4. A piece of steel is 7.5 cm long, 3.7 cm wide, and 0.75 cm thick. Express these dimensions in inches.

5. A block of copper is 12.4 cm long, 4.7 cm wide, and 0.60 cm thick. In what unit would the volume be expressed? Calculate the volume. (Are you remembering about significant figures?)

6. If the density of steel is 7.6 g/cm³, what is the weight of the piece described in problem 4?

7. If the specific gravity of copper is 8.92, what is its density in the metric system? What is its density in the English system? What is the weight in grams of the block of copper described in problem 5?

8. A flask is found to weigh 76.43 grams when empty and 177.39 grams when full of water at 4°C. What is the weight of the water in the flask? What is the volume of this water? What is the volume of the flask?

9. In an experiment to determine the density of aluminum by
means of an overflow can and bucket, a student found that
(a) the empty bucket weighed 46.2 grams; (b) the
bucket containing the displaced water weighed 244.7
grams; and (c) the block of aluminum weighed 536.8
grams. From these data, what is the density and the
specific gravity of aluminum?

10. Change the following Centigrade temperature readings
to °F: 20, 25, 96, and -182.9. Change the following
Fahrenheit temperature readings to °C: 32, 80, and
-40.

11. How many calories of heat would be required to raise
the temperature of 20 g of water 15°C; 50 g of water
from 15°C to 95°C; and 150 ml of water from 30°C
to 75°C?

12. How many calories of heat would be required to
raise the temperature of 75 g of aluminum from 20°C
to 140°C; 2 kg of copper from -10°C to 50°C; and
150 g of zinc from -5°C to 115°C?

13. A piece of lead weighing 312 grams was heated
to 93°C and then placed in a vessel containing 312
grams of water at 20.6°C. The temperature of the
water rose to 22.8°C. How many calories of heat
were required to raise the temperature of this
quantity of water from 20.6°C to 22.8°C? Where did
this heat come from? How many calories of heat were lost
by the lead in cooling from 93°C to 22.8°C? How
many calories of heat did the entire mass of lead give
up in cooling 1°C?
Particulars of Training

The following conditions to be observed:

1. The duration of the course is 4 months.

2. The training will be conducted in two phases:
   - Phase I: 2 months
   - Phase II: 2 months

3. Attendence is mandatory for all students.

4. The training will be held in a combination of classroom and practical sessions.

5. Students will be evaluated through written exams and practical assessments.

6. Students who fail to meet the criteria will be offered a second chance.

7. The training will conclude with a final examination.

8. Students are expected to complete all assignments and projects on time.

9. Students must maintain a minimum attendance of 80%.

10. Students who fail to comply with the above conditions will be disqualified.

How many calories of heat were given up by 1 g of lead in cooling 1°C? What is this quantity of heat called?