New and improved demonstrations for use in teaching scientific principles in chemistry

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http://hdl.handle.net/2144/4846
BOSTON UNIVERSITY
SCHOOL OF EDUCATION

Thesis

NEW AND IMPROVED DEMONSTRATIONS
FOR USE IN TEACHING
SCIENTIFIC PRINCIPLES IN CHEMISTRY

Submitted by
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B. S. in Education, Boston University, 1941

In Partial Fulfillment of
the Requirements for the Degree of
Master of Education

1950
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A survey of recent studies shows a definite trend in the teaching of science toward placing the emphasis on the teaching of principles and generalizations and toward stressing the value of demonstrations as a teaching technique. There are now available to the teacher of secondary school science reports of the findings of extensive studies which have been devoted to the selection of scientific principles upon which a satisfactory course in secondary school science could be organized. These studies, in presenting the groundwork for improved courses in secondary school science, courses which will adequately meet the needs of the student of today, have made a valuable contribution to the advancement of the teaching of science. The selection of scientific principles, however, no matter how carefully it may have been done, will prove to be of little value to the students of our secondary schools unless these principles can be presented in such a manner that the student will be able to understand and to apply them. Much has been written on the value of the classroom

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1Chapter II
demonstration as a teaching technique. Yet it has been the experience of the writer, as a classroom teacher, to find that work in the field of developing good classroom demonstrations and assigning them to selected principles has been somewhat neglected. A diligent search through all volumes of the Journal of Chemical Education since 1935 revealed a surprisingly small number of demonstrations which could be used to illustrate scientific principles. Books devoted entirely to demonstrations catalogued according to the principle which they illustrate are rare. Many of the demonstrations described in these books and in available textbooks, laboratory manuals, and journals have failed, when tested, to illustrate adequately a scientific principle. Some demonstrations failed completely.

In this study, therefore, the writer has devoted his efforts to testing and improving existing demonstrations and to developing new ones, devoting a major portion of his time and effort to the latter. It is hoped that this study will aid in directing the attention of other graduate students to the development of better classroom demonstrations and that in the not far distant future more and better demonstrations will be available to the classroom teacher.
STATEMENT OF THE PROBLEM

This study has been undertaken to develop "new" or improved demonstrations which have simplicity of design, forceful presentation of the principle involved, and "purity of concept", and which are catalogued according to the principles which they illustrate.

SCOPE AND LIMITATION OF THE INVESTIGATION

This investigation has been divided into two major sections. In the first, the writer has attempted to develop one "new" or improved "good" demonstration which could be assigned to each of four selected scientific principles in chemistry. The second contains the results of his having followed valuable leads to new demonstrations which were uncovered during the investigation. The four principles were selected from the sixty-six of the two hundred seventy-two physical science principles validated by Wise\(^1\) which were assigned by Blanchet\(^2\) to chemistry. If new demonstrations

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could not be developed, existing ones were improved to satisfy the following criteria which had been established for a "new" demonstration and a "good" demonstration.

The term "demonstration" means any kind of activity in connection with which apparatus or materials are manipulated by the teacher and/or one or more pupils, in an attempt to solve a problem in the field of science, and is for observation by all members of the class.¹

A demonstration was considered a "good" demonstration if a jury composed of five experts in the field and five teachers of secondary school science unanimously agreed that it satisfied the following: (1) it had "purity of concept", (2) the concept could be recognized easily by students of the secondary school level, (3) it could be performed in forty minutes or less, and (4) it employed only "accessible material".

A demonstration was considered to have "purity of concept" when it illustrated one and only one principle or when, if all other principles could not be eliminated, they did not detract from the principle for which the demonstration was performed or lead to misconception of it.

The term "accessible material" referred to equipment and chemicals which could be found in the average high school laboratory, could be purchased locally, or could be made from inexpensive materials locally purchased.

RESEARCH PROCEDURES AND TECHNIQUES USED

Having analyzed the problem and defined specific terms as given above, the writer proceeded to select the four scientific principles and the five reference sources which would serve as a basis for this study. In order to be fair in his selection of the principles and not to take the "cream of the crop", the writer wrote each of the sixty-six of the two hundred seventy-two physical science principles validated by Wise\(^1\) and assigned by Blanchet\(^2\) to chemistry on a slip of paper. After these sixty-six slips of paper had been placed in a box, the writer drew from the box the following four principles:

2. Every pure sample of any substance, whether simple or compound, under the same conditions will show the same physical properties and the same chemical behavior.

58. Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

\(^1\)Wise, Loc. cit.

\(^2\)Blanchet, Loc. cit.
67. Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

188. The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

This study was limited to four principles because an extensive investigation of the entire sixty-six would be a task much too great for one person.

Because of their completeness and because they are considered by experts to be an authoritative source, the following references were chosen as a basis for judgment in determining whether or not a demonstration was "new" and as source material for demonstrations to be improved.


4. Journal of Chemical Education, all issues since 1935.

5. The Science Teacher, all issues since 1935.

The procedure followed for developing "new" demonstrations and improving existing ones may be divided into six phases: search of the selected reference sources to assign demonstrations to principles, testing of demonstrations found, improvement of demonstrations found, the
development of "new" demonstrations, and the validation of "new" and improved demonstrations.

After writing each principle at the top of a regular 8½" x 11" sheet of paper, the writer searched the literature, page by page, for demonstrations that would illustrate the principle. When such a demonstration was found, instructions for performing it were copied with complete details, including a diagram of apparatus whenever necessary. If no demonstration was found, the sheet was saved for recording in detail the new one developed.

The writer then performed and studied each demonstration to determine whether or not it was a "good" demonstration according to the criteria established for this study. Comments and, whenever possible, suggestions for improving the demonstration to make it a "good" demonstration were recorded on the page. An attempt was then made to develop a "new" demonstration for each of the principles selected as a basis for this study. All of the demonstrations were tested before a class.

The final, tested, improved or "new" demonstrations with comments and suggestions were sent to a selected jury consisting of five experts in the field and five teachers of secondary school science. The judges were provided with the criteria for judging a demonstration as "new" and "good". If the members of the jury unanimously agreed that a demon-
stration satisfied the criteria established for this study, the demonstration was considered "new" or improved and "good". If they did not agree, further search for demonstrations was made.
CHAPTER II

SURVEY OF RELATED LITERATURE

As early as the turn of the century, leaders in the teaching of science became aware of the fact that the real value to be derived from the study of science lay in the mastering of principles and generalizations. Ostwald's book, The Principles of Inorganic Chemistry, published in 1900, marked the beginning of this new pedagogical era. Since that time, many authors have adopted the use of the word "principle" in the titles of their texts. In some cases, however, as recent studies have shown\(^1\), the word applies in title only. Some of the latest textbooks, upon careful examination, were found to devote more space to unrelated descriptions and isolated facts than to the development of principles and generalizations. Wise\(^2\) has made a valuable contribution to the advancement of the teaching of principles and generalizations by his selection and establishment of the principles suitable for science courses in general education.

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\(^2\)Wise, Loc. cit.
Fowles\textsuperscript{1}, in his explanation of the modern method of teaching chemistry says:

Physical principles are the basis of chemical actions; consequently, by mastering these principles the study of chemistry is facilitated and the unity of the subject promoted. Those who adopt the modern method, while treating topics in any chosen order expedient for presentation and study endeavor to show that chemistry is a series of related phenomena and not a mass of isolated facts.

Alexander Smith, who was appointed in 1911 to succeed Professor Chandler as head of the department of chemistry at Columbia University, inaugurated, in 1906, a revolution in methods of instruction in America, comparable only with that effected in Germany by Ostwald. "Instead of a hodgepodge of isolated facts, the student found in these volumes a clear and logical presentation of the fundamental principles of science."	extsuperscript{2}

To the ultra-conservative, the critic who always decries change and prefers the status quo to progress but who at last has been forced to accept grudgingly the inductive method, may say, "I told you so. After all the clamor for the teaching of simple, true to life experiences, you now return to the teaching of the abstract." The progressive

\textsuperscript{1}Fowles, \textit{Op. cit.}, p. 518.

too may question whether or not the stress on the teaching of principles and generalizations does not conflict with the inductive method of developing ideas through concrete examples and experiences closely related to the everyday living of the student.

The answer to both these critics is the same. The teaching of principles and generalizations does not conflict with the inductive method of developing ideas through concrete, everyday examples. It simply organizes and clarifies the wealth of ideas which could be so taught. It stresses the teaching of fundamental truths, or principles, which will be basic in innumerable experiences of everyday living.

Elliot Rowland Downing is obviously a strong advocate of science closely related to the everyday life of the student, as shown in the following statement.

The science of the high school needs to be taught as a preparation for life-consumer science rather than producer science. If the public schools, both elementary and secondary, could concentrate on the mastery of those principles of science that are most often needed in the solution of those problems that arise in the life of Mr. Everyman and on the clarification of the apprehensive mass needed to understand them, then the pupils would be prepared for life in the field of science.

Yet he forcefully stresses the importance of teaching fundamental principles and generalizations rather than the

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solutions to specific, unrelated problems.  

While the number of specific things of a scientific nature to be done in any community are exceedingly numerous, the scientific principles involved in them are relatively few. Only a dozen principles underlie the 399 operations in the care of the automobile. Let the pupil understand the few simple principles involved in wise food selection and he will apply them in China as well as in his home town. The principles of science taught in his school days enable the old person to understand now the action of all these new appliances excepting the radio, which involves some new principles. Principles are nearly universal. Specific problems are local in both place and time.

It seems wise, then, to give a pupil such understanding of the important principles of science and so much drill in solving problems occurring in life-situations that he will be skillful in the recall of such principles and in applying them when needed to solve his specific life problem."

Preston also emphasizes the value of teaching principles and generalizations rather than specific details.

We have not time to teach our child all the details, nor would it be advisable to do so if we could, for many of them may have become obsolete ten years hence. But we can select as our teaching goals the mastery of principles so fundamental that the changes which occur will be changes merely in their applications. And these principles we can show applied in so many ways that our pupils, through imitation, will form the habit of looking for further opportunities to recognize and use them.

It is because the truth of this is recognized that more and more our instruction is being changed from the plan of teaching isolated facts, or even facts topically grouped, to that of


teaching carefully selected principles, the mastery of each of which brings some new adjustment of the pupil to his environment or opens his eyes to a set of previously invisible relationships.

Regarding the functional understanding of principles, the Thirty-First Yearbook committee\(^1\) stated:

In a program of general education which is for the most part the program of the public schools, the work of instruction will be directed toward increasing the understanding of those principles and generalizations of science that have had largest application in moulding the character of our society and of those which have within them the potentialities for influence in the future.... The objectives may be stated as the principles and generalizations that are functional for the individual, in that they enable him to interpret the experiences of living. The objectives may be formulated (1) as statements that function directly in thinking, (2) as statements that describe attitudes toward products of thought and toward methods of thinking.

In the Forty-Sixth Yearbook\(^2\) the functional understanding of principles is given as one type of objective for science teaching.

The functional understanding, for example, of such a principle as "energy can be changed from one form to another" is basic to an understanding of physical science and of many human relationships. It has far-reaching social and economic relationships. As the learner's understanding of such a principle grows, he will attain a more intelligent and sympathetic appreciation of many problems of


the world in which he lives.

If we are to teach principles and generalizations, then, we must consider carefully the possible methods of instruction which we could employ. The demonstration is recommended by authorities in both science and education as a most effective means to this end. Elder states that if demonstrations are not used to illustrate principles, "... the teacher is losing an opportunity to make his lectures far more interesting than those given by the teachers of non-science courses". Elder also says that "the demonstration must beware of certain pitfalls which it is easy to drop into and hard to get out of. There is the danger of trying to demonstrate too many principles in one period."

Arthur's comments on the value of the demonstration as an effective technique in the teaching of science are well worth consideration.

Experience has shown teachers that "Seeing is believing" is a fundamental maxim of practically every individual. Students vary as to the type of teaching to which they respond most readily. Some remember best the things they read; others, the things they hear; still others, the things they see. Demonstrations enable the students to see and hear at the same time, and textbooks enable them to read the material covered. Thus the teacher who makes use of demonstrations is taking advantage of

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all the types of learning instead of crippling his efforts by using only a few of the avenues open to him.

In encouraging the increased use of the demonstration as a teaching technique, the writer is in no conflict with those authorities who advocate pupil experiments and the theory of "learning by doing". He is heartily in accord with the following opinion expressed by Joel H. Hildebrand\(^1\) in his article, "A Philosophy of Teaching".

... the approach I advocate is the "psychological". It is based upon two fundamentals of the psychology of learning, first, that it is easier to interest students in the more immediate, live problems than in the more remote, dead ones. The story of a scientific idea is far more exciting if it leads up to the present, into explorations still going on, with unsolved elements inviting immediate hypotheses and possible experiments. The discovery and isolation of plutonium has, I think, far greater pedagogic value than the story of the development of the air pump. The second fundamental is that "ability to do something is developed not just by hearing it expounded but by first-hand effort and practice. The role of the learner must be an active one. If the goal of 'understanding science' on the part of laymen is to bring a little more reasonableness into human affairs, is not the most appropriate method the encouragement of a little scientific thinking and experimentation on the part of students, even if necessarily on a rather elementary plane, rather than reliance solely upon descriptions of scientific efforts by others? This principle is well understood in the realm of physical education, where muscular control is developed by performance, with repetition directed towards ever higher levels of difficulty and skill, and only incidentally by lectures and demonstrations. Athletes, including amateurs, are developed on the field, not on the bleachers.\(^8\)

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The Forty-Sixth Yearbook\(^1\) contains a similar opinion of the value of pupil experiments in the teaching of science.

Is the observation of a demonstration experiment as effective and valuable to a pupil as his performance of that experiment? ... researches indicated that in certain important respects the individual method is superior to the demonstration method.

Since experimentation involves "learning by doing", there can be no substitute for it. Pupil experimentation is an essential part of a good science education. In every course of science offered at any level, therefore, opportunities should be provided for the pupils to perform experiments.

The writer believes, however, that because of limited time and facilities, the teacher of secondary school science can accomplish more by using both demonstrations and pupil experiments than he could by relying upon pupil experiments alone. Moreover, the demonstrations presented in this study lend themselves easily to pupil participation and could be used, if desired, as pupil experiments. They offer excellent material for project work by superior students and so may also be of service to the teacher who is trying conscientiously to provide for individual differences.

The Forty-Sixth Yearbook\(^2\) committee has summarized its comparative evaluation of the individual and the demonstration methods as follows:

\(^1\) The Forty-Sixth Yearbook, p. 53.

\(^2\) Idem.
The discussion of the individual and the demonstration methods in the Thirty-First Yearbook is believed by this committee still to be sound and practicable. The summary of it is, therefore, included here with only minor modifications:

a. Each method possesses certain unique values. Hence, both are necessary, each to supplement the other, in every science course at every level.

b. In the interest of economy both of time and of money, it is desirable to perform more laboratory exercises by the demonstration than by the individual method.

c. At the start of every laboratory course there should be sufficient use of the demonstration method to acquaint the pupils with apparatus and with accepted methods of experimentation. Following this period of orientation, the pupil should be allowed to perform some exercises individually in order to acquire, early in the course, desirable manipulatory skills and laboratory techniques and habits.

d. The time saved in each course by the use of the demonstration method should be used for other types of learning activities, which may include additional laboratory exercises, reading projects, individual investigations, observations, and drill upon essentials. The money saved by purchasing single sets of apparatus for demonstration certain experiments, instead of multiple sets for individual performance, should be expended for books, visual aids, additional laboratory equipment, and other means of enriching and improving science instruction.

There is widespread confusion among teachers with respect to what constitutes individual pupil experimentation.... Individual laboratory work can mean only experiments performed by a single pupil.

Research evidence indicates that pupils working in pairs derive less benefit from an experiment than do pupils working individually. The difference, however, is not great enough to justify the added expense of providing individual sets of equipment for all experiments.

The fact that the demonstrations presented in this study lend themselves readily to pupil participation has been previously mentioned. Concerning the comparative values
of teacher participation and pupil participation in demonstrations, the Forty-Sixth Yearbook\(^1\) contains the following opinion:

The primary purpose of a demonstration is to provide vicarious experience; it is not intended to provide a means by which pupils secure firsthand experience in manipulating apparatus and in conducting experiments. It is good practice, however, to have the pupils assist the teacher during demonstrations, but such participation should be distributed as widely as possible throughout the class so that all may profit by whatever values such participation provides.

A survey of the related literature, then, reveals the following trends in the teaching of science: the use of principles and generalizations as a basis upon which to build courses in science, and the use of a combination of demonstrations and individual pupil experiments to most effectively present those principles. Beauchamp\(^2\) makes the following conclusions as the result of his investigation of the teaching of science in the secondary schools:

The analysis of courses of study and classroom observation indicate certain practices which may be considered as innovating and hence should be carefully examined and evaluated by those who are engaged in curriculum building and who are desirous of improving instruction in science. These practices are as follows:

a. A shift has occurred from the organization

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\(^1\)Forty-Sixth Yearbook, pp. 54-55.

of courses in terms of topics and subtopics to organization around certain major ideas or concepts. These ideas may be generalizations or principles of science or important ideas underlying the understanding and control of certain phases of man's environment.

b. A shift has taken place from the topical method of developing a topic to the problem method of development. Each problem is focused on some important idea or generalization of science....

h. Teacher or pupil demonstration has replaced individual experimentation to a marked degree in the junior high school. A great increase in the use of the demonstration is also observed in the specialized science courses....

This list of innovating practices does not necessarily indicate the changes which should be brought about in the teaching of science. The trends do, however, disclose the variations from the dominant practice which may well result in progress and which, as already stated, should be given careful consideration by those who are undertaking the improvement of instruction in science at the secondary level.
CHAPTER III

DEMONSTRATIONS TO ILLUSTRATE SELECTED PRINCIPLES

The search for new demonstrations has been a task both interesting and challenging. In carrying out the investigation to find a new demonstration to illustrate each of the principles in the selected list, the writer uncovered several leads which, when followed through, resulted in more new demonstrations. Since these seemed too valuable to be disregarded, although they illustrate principles other than the ones chosen for this study, they have been included in chapter four.

This chapter includes three types of demonstrations: (1) demonstrations found in the source material which illustrate the selected principles, (2) improved demonstrations, and (3) new demonstrations developed by the writer.

When there was doubt in the minds of the judges as to the value of a demonstration, the writer made comments and suggestions immediately following the demonstration.

If the same demonstration was found in more than one of the reference sources, or if several demonstrations very similar in nature were found therein, the writer did not copy duplicate demonstrations. Those included in this chapter, therefore, as "demonstrations found in the source material" are, in the opinion of the writer, representative
of the best examples included in the selected references.

All of the demonstrations included in this chapter and in chapter four have been tested both in the laboratory and before a class. The new demonstrations have been performed many times to prove that they are reproducible.
Every pure sample of any substance, whether simple or compound, under the same conditions will show the same physical properties and the same chemical behavior.

Demonstrations to illustrate this principle are far too numerous to be included in this study. In order to illustrate this principle, one needs only to refer to any good textbook of qualitative analysis, a handbook of chemistry, or any reliable textbook of general chemistry for hundreds of good illustrations.

The writer suggests, therefore, that this principle can be more practically presented by emphasizing it each time that a specific test is given for a particular substance, or whenever melting point, boiling point, freezing point, or other physical or chemical properties are discussed, than it could be by the use of any single demonstration or set of demonstrations.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

According to Bennett's *Standard Chemical and Technical Dictionary*, ignition temperature is defined as "the lowest temperature at which chemical reaction in a given gaseous medium becomes self sustained". Brisco¹ defines kindling temperature as "the temperature at which a substance bursts into flame and combustion proceeds without further application of heat". Kindling temperature is, however, a very misleading term. It gives the impression that there is a definite temperature at which combustion will start. The temperature is only one of the conditions which produce combustion. Other conditions which affect combustion are the physical condition of the solid, the pressure of the gas or vapor, the presence or absence of a catalyst, the nature of the catalyst, and the concentration of oxygen in the gas in contact with the substance. No definite kindling temperature can be given for a substance unless these other conditions are specified also. Kindling conditions involve several variables, of which the temperature is only one. These facts should be kept in mind whenever a demonstration is performed.

to illustrate this principle.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles——Exp. 362 p. 287

Procedure
Silica heated with excess of magnesium forms magnesium silicide.

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

Magnesium silicide interacts with hydrochloric acid to form silicon hydride and magnesium chloride.

\[ \text{Mg}_2\text{Si} + 4\text{HCl} \rightarrow \text{SiH}_4 + 2\text{MgCl}_2 \]

Finally, the spontaneously inflammable silicon hydride reacts with the oxygen of the air.

\[ \text{SiH}_4 + 2\text{O}_2 \rightarrow \text{SiO}_2 + 2\text{H}_2\text{O} \]

Place in a small test-tube of soft glass a mixture of 4 gm. of magnesium powder and 2.5 gm. of dry silica. Clamp the tube, with its mouth pointing in a safe direction, on a retort-stand and place an iron tray under the tube. Heat the mixture; reaction takes place with a dazzling flash of light. Break the tube, place the product in a small flask, and add dilute hydrochloric acid. Silicon hydride comes off and as it meets the air bursts into flame with a sharp pop.
The silica appears as a white smoke.

Comments

All silanes are extremely sensitive to oxidation. As oxidation continues, the temperature of the system is greatly increased until the ignition temperature is reached and the silicon hydride ignites spontaneously. This is a spectacular demonstration and a good one to illustrate spontaneous combustion. The writer would hesitate, however, to use it to illustrate the above principle for the following reasons:

1. Since considerable oxidation precedes the ignition, it would be extremely difficult even to estimate the ignition temperature.

2. It would be unfair to allow the students to believe that the ignition took place at room temperature.

3. Since the kindling temperature of the hydride cannot be determined by this method, comparison with the kindling temperatures of other substances cannot be made.

If this demonstration is used, the writer suggests that the quantities of materials be reduced to 0.4 gm. of magnesium and 0.25 gm. of silicon dioxide to make the demonstration less hazardous and easier to perform on the open lecture bench.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Fig. 1.---Preparation of Phosphine
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles-----Exp. 378 p. 296

Procedure

Fit up an apparatus as shown in the diagram on the preceding page. Half-fill the small flask with strong caustic soda solution, 30-40 gm. of sodium hydroxide per 100 c.c. of water. Add 2 or 3 pieces of yellow phosphorous of the size of green peas. Place warm water in the trough. Send coal gas through for a minute to expel air from the apparatus; whereon close the rubber tubing with a clip. Now heat the solution and, as soon as the gas begins to come over, adjust the flame so that the rate of evolution is about 2 bubbles a second. As soon as the coal gas is expelled the bubbles of phosphine take fire on meeting the air and form vortex rings of smoke. When the phenomenon has been sufficiently seen, remove the flame and expel the phosphine still in the apparatus by once more passing coal gas into the flask.

Comments

Fowles does not ask that this demonstration be performed to illustrate kindling temperature but rather suggests that the student recall the facts of the demonstration presuppos-
ing that it had been performed previously. The writer finds two main criticisms of this demonstration: (1) although it is an excellent demonstration to illustrate that some substances have a very low kindling temperature, the instructor must take care that the students are not misled to think that the phosphine catches fire. The kindling temperature of phosphine is in the vicinity of 100° C, but when it is prepared in this way, an appreciable amount of the liquid hydride (\( \text{P}_2\text{H}_4 \)) is formed. This is spontaneously inflammable in contact with air at room temperature. Pure phosphine is not spontaneously inflammable.\(^1\) (2) This demonstration is somewhat dangerous since phosphine is highly poisonous. It never should be performed in the open.

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Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles----Exp. 524 p. 479

Procedure
Dry a piece of phosphorus and touch it with a warm rod, not warmer than can just be held with the hand. Phosphorus ignites at 60° C.

Comments
This demonstration illustrates a difference in kindling temperature between the hand, which does not ignite, and the phosphorus, which does.

It is of interest to note that Fowles states the kindling temperature of phosphorus to be 60° C, while it is given in Smith's College Chemistry and in Mellor's Modern Inorganic Chemistry as 30° C. Other books give values ranging between these two temperatures. Are the books wrong? Probably not. The difference is because of the methods used to make the determination or the condition of the substance.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles----Exp. 524 p. 479

Procedure

Stuff a tuft of cotton wool loosely in the mouth of a test-tube and clamp the tube in an upright position. Have ready, in a corked tube or flask, about 1 c.c. of carbon disulfide. Heat a glass rod in a Bunsen flame to dull redness. Pour the carbon disulfide on the cotton wool, turn out the flame, and when the rod has cooled just short of dull redness hold it about \( \frac{1}{2} \) inch above the cotton wool. The carbon disulfide vapour ignites and burns with a blue luminous flame.

Carry out a second experiment in the same way, using dry ether in place of carbon disulfide. A rod heated to dull redness fails to ignite the vapour. Now heat the end of the rod until the glass begins to soften. The hot rod will now ignite the vapour. Try familiar solvents and cleansing liquids, such as petrol, alcohol, westersol, and carbon tetrachloride in the same way.

A flame is necessary to kindle the vapour of alcohol. Boil a few c.c. of alcohol in a small flask and ignite the
vapour with a burning taper.

Comment

This is an excellent demonstration.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles-----Exp. 525 p. 480

Procedure
The flame of a candle cooled below the ignition temperature is extinguished.

Make a helix, of thick copper wire, sufficiently large to envelop the candle flame without touching it. Place the helix around the flame of the candle. The flame is soon extinguished. Now make the helix hot, relight the candle, and once again envelop it in the helix. The candle remains alight.

Comments
In this demonstration there are two critical factors: namely, the size of the wire and the diameter of the helix. In trying out the demonstration, the writer found it difficult to prepare a helix which would give good results. He made eleven different spirals before finding one which was satisfactory. When he performed the demonstration before several students and three instructors in chemistry, a question arose as to whether the candle flame was smothered or cooled below the kindling temperature by the helix.
To clarify this, he made a helix of glass rod, approximating as closely as possible the dimensions of the copper helix. The fact that the helix of glass rod had very little effect upon the flame proved more conclusively that the flame had been cooled below its kindling temperature by the copper.

Instead of giving specific directions regarding the size of the wire to be used or the diameter of the helix to be made, the writer suggests that anyone wishing to use this demonstration make and test a helix beforehand to insure satisfactory results.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material

Fowles-----Exp. 528 p. 481

Procedure

Platinum black, slightly warm and supported on a piece of asbestos paper, is held in a stream of coal gas. Probably adsorption and then catalytic oxidation occur, for the platinum first becomes red-hot and soon after ignites the gas.

Tear off a small piece, say 1 x 2 inches, of asbestos paper, and moisten the frayed edge with a few drops of platinic chloride. Holding the paper with tongs, heat it in the flame until all fumes are gone and only platinum is left. Now turn out the gas, and immediately -- that is, before the paper has completely cooled -- turn it on again. Hold the frayed edge of paper in the edge of the gas stream, the gas is ignited.

Comments

This demonstration is very effective if coal gas is used. If bottled gas or some other hydrocarbon gas is used, however, the demonstration will not work satisfactorily. When the writer performed the demonstration using several different kinds of bottled gas and gasoline vapor under
pressure, he could not obtain satisfactory results. The platinized asbestos became red-hot but would not relight the flame. When coal gas is used, heat is produced in two ways: (1) by oxidation on the surface of the platinum and (2) by adsorption of hydrogen by the platinum. Thus enough heat is produced to ignite the gas.

Platinum chloride costs $59.04 an ounce. This is a very expensive reagent and consequently will not be found in many high school laboratories.

Modification

Collect three or four bottles of hydrogen. Heat the platinized asbestos in the Bunsen flame and then hold it near the mouth of one of the bottles of hydrogen. The asbestos becomes red-hot and finally the hydrogen oxygen mixture in the bottle explodes with a sharp retort. Now, without reheating, hold the platinized asbestos near each of the remaining three bottles in succession and note the results. This gives a very spectacular demonstration.

When this demonstration was performed using eight bottles of hydrogen placed in a row, mouth upward and uncovered, it was noted that the sharpness of the report gradually increased from the first through the fourth bottle. Then it decreased gradually. This is an interesting and convenient way of showing that the effectiveness and extent of an explosion depend upon the relative amounts of the reacting
gases in the mixture. Diffusion caused a gradual increase in the amount of oxygen in each bottle and a decrease in the amount of hydrogen. The bottle giving the loudest report contained the most effective mixture.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles----Exp. 526 p. 480

Procedure
A suspension of coal dust in oxygen is ignited.

Grind some coal in a mortar to an impalpable powder. Place a 1/8-inch layer of the coal dust in a perfectly dry gas jar. Pass dry oxygen into the jar until it is full, then close the jar with a dry, ungreased cover-glass (so that no dust shall adhere to the glass). Shake the jar very vigorously so as to get the dust in suspension. Remove the cover and immediately apply a lighted taper to the mouth of the jar. The mixture ignites, the vigor of the action depending on the quality of the coal and the fineness of the dust. If a suitable coal is not at hand use powdered wood charcoal.

The danger of using an open light in a coal mine free from combustible gas is manifest from this experiment, likewise the necessity of adequately damping the air before firing. Furthermore, nearly all finely divided metallic and organic substances form mixtures with air which explode when ignited. Unfortunately, this danger is not generally recog-
nized and dust explosions occur in factories and warehouses at far too frequent intervals.

Comments

This is a good demonstration but rather dirty. Lycopodium powder will give the explosion directly without the addition of oxygen.¹ The use of lycopodium powder will give a demonstration much cleaner and easier to set up.

Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material

Fowles-----Exp. 533 p. 483

Procedure

Burning petrol is poured into a recepticle covered with iron gauze; the petrol is extinguished.

Bend a square of ordinary iron gauze over the mouth of a dry 2-lb. glass jam jar or other convenient recepticle. Place 20 - 50 c.c. of petrol in a small saucepan and ignite the petrol. Pour the burning petrol, from a height of 1 - 2 feet, in a steady stream through the gauze into the jar. The petrol is extinguished as it passes through the gauze. The small amount which adhered to the upper side of the gauze soon burns away. Remove the gauze and ignite the petrol in the jar.

Comment

This is an excellent demonstration.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles----Exp. 534 p. 484

Procedure
(a) Hold a piece of wire gauze horizontally with tongs and lower it slowly into a Bunsen flame. The reacting gases, rendered visible as flame, are pressed down and completely cut off, as though the gauze were a sheet of solid metal.

(b) Next make the gauze red-hot, then lower it into the hollow cone. The gases, no longer cooled, pass through the mesh and burn on the upper side of the gauze.

(c) Repeat (a) above, taking care to depress the gauze as far as the zone of unburnt gas. The gas, passing through the gauze, is cooled below its ignition temperature. Bring a lighted taper just above the gauze. The gas now burns, and two distinct flames are seen.

(d) Stand a sheet of gauze on the top of an unlit Bunsen, with the air inlet open. Turn on the gas fully and ignite it above the gauze. Raise the gauze slowly; the flame ascends with the gauze to a height of 4 - 5 inches.

Comment
This is an excellent, simple demonstration.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

Demonstration found in the source material
Fowles----Exp. 535  p. 484

Procedure

A small Davy lamp is constructed, lit, and placed in an explosive mixture. The lamp is withdrawn and the mixture exploded with a naked flame.

Take a cork 3 - 4 cm. in diameter and a piece of fine brass wire gauze 12 cm. square. Roll the gauze into a cylinder of such diameter that the cork will fit tightly into one end of it. Cut out of the same wire gauze a circle, a little larger in diameter than the cylinder. Draw out four or five of the bent wires at the top of the cylinder, so that the straight wires are left projecting free for 5 to 6 mm. Cover the top of the cylinder with the circle, passing the straight wires through the gauze near the edge of the circle, and fix them by bending them over. Now fix a small candle to the cork, by melting the wax at the bottom, and push a stout wire into the cork to serve as a handle. Light the candle, and slip the wire-gauze cylinder over the cork, thus completing the lamp.

Pass a little coal gas into a wide-mouthed bottle so as
to form an explosive mixture. Pass the lamp into the jar and notice what happens. Withdraw the lamp and apply a naked flame to the jar.

If procurable, old and recent designs of Davy lamps should be shown and examined.

Comment

This is an excellent demonstration.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

A new demonstration

Procedure

Make six or eight spiral turns in a platinum test wire. Heat the wire in a Bunsen flame. Then suspend the wire over 20 c.c. of methyl alcohol contained in a small bottle. The heated spiral of platinum glows in the mixture of alcohol vapor and air (oxygen) which exists above the alcohol. The wire gets red hot, and when the kindling temperature of the alcohol is reached, the alcohol ignites with a small flash. When this flash occurs and the alcohol burns, the wire becomes cool. As soon as more air diffuses into the bottle, the wire again becomes red hot. This action is continuous and may be observed for the length of the entire period. The heat is developed by the interaction between the alcohol and the oxygen which takes place with great speed at the surface of the platinum. Note the odor of formaldehyde.

Caution--- If this demonstration is allowed to run for any great length of time, the bottle should be placed in a

---This demonstration is "new" only in the sense that it was not found in the literature. It was suggested to the writer by Professor P. D. Baker, Head of the Chemistry Department, Norwich University.
shallow pan large enough to hold the alcohol, since the heat
developed may crack and break the bottle.

This demonstration could be used to show also the cata-
lytic oxidation of an alcohol.
Each combustible substance has a kindling temperature which varies with its condition but may be greater or less than the kindling temperature of some other substance.

--- Glass rod

--- Platinum spiral

--- Methyl alcohol

Fig. 2.---Catalytic Oxidation
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

There are several sorts of chemical behavior which make it possible to place the metals and hydrogen in a relative activity series. They may be summarized as follows:

1. Diminishing rate of reaction with oxygen (under comparable conditions) down through mercury.
2. Diminishing stability of oxide toward heat. The oxides of mercury and the metals below it are decomposed over an ordinary Bunsen flame.
3. Diminishing resistance of the oxide toward reduction by hydrogen (under comparable conditions). The oxides of iron and the metals below it are progressively more easily reduced by hydrogen.
4. Diminishing rate of reaction with cold water, hot water, and finally superheated steam, down through iron.
5. Diminishing rate of reaction with dilute acids down through lead.
6. Any element will displace any succeeding, less active, element from solution.
7. From the voltage of battery-cells in which the metals are anodes. The free energies of reactions, or their tendency to go, is proportional to the voltage developed.
when the reaction is carried out in a cell.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Fowles----Exp. 118 p. 106

The detailed directions for performing this demonstration will not be given here, since this demonstration is concerned mainly with the properties of calcium carbonate and minerals of calcium carbonate. A part of the demonstration, however, illustrates the diminishing stability of carbonates toward heat, which may be used as a measure of the relative activity of the metallic component. Copper carbonate breaks up at a low temperature; lead carbonate, at a moderate temperature; and sodium carbonate, only slowly at white heat.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Fowles----Exp. 124  p. 112

This demonstration, illustrating the reduction of carbon dioxide by zinc or iron, could be used to show the order of activity of metals by repeating the experiment using metals other than zinc and iron. It would be exceedingly difficult, however, to perform with sufficiently good results any such demonstration in forty minutes. For this reason the details of this demonstration will not be included.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Fig. 3.---An electrode-potential cell.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur----Exp. 121 p. 315

Procedure

In a 600-ml. beaker, place a 50 by 100-mm. porous clay cup of the type used for diffusion experiments or in Bunsen cells. Fill the cup three-fourths full of IN sodium sulfate solution, and in the beaker place 0.5N sulfuric acid, making its surface level with the surface of the solution inside the cup. In the cup, place a 15-cm. zinc rod; in the beaker outside the cup, place a platinum electrode or graphite rod. Connect the zinc and platinum to a large demonstration galvanometer, and note the magnitude and direction of the deflection.

Leave the platinum electrode connected to the galvanometer. In place of the zinc rod, use, in succession, rods of magnesium, freshly polished aluminum, soft iron, tin, cadmium, and copper. The direction of the deflection will be the same for all of these metals; the magnitudes will be in the order expected from a displacement chart, the mag-

1See Fig. 3, p. 51.
nesium and aluminum giving the greatest, and the copper, the least deflection.

Using the platinum and zinc as at first, connect the two directly together by a piece of copper wire. Note that the hydrogen is given off at the surface of the platinum. Let several students observe this closely and report to the class.

Comment

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur-----Exp. 122 p. 316

Procedure

a. Fill a 250- by 150- by 50-mm. glass battery jar with plane sides, three-fourths full of 1N HCl. In this solution, place a 30- by 4-cm. strip of freshly polished sheet copper, leaning the strip from one side of the jar to the other. Insert a 30-cm. zinc rod into the solution, and press it firmly against the lower end of the copper strip. Note that hydrogen is given off at the surface of the copper, but not at the zinc. Raise the zinc rod so it no longer touches the copper. The hydrogen now comes from the zinc and not from the copper. In this case, however, the hydrogen is evolved much more slowly than formerly.

b. Try the demonstration described in (a), using copper with iron or tin. Try also iron with zinc and iron with tin. In the case of iron and zinc, the iron gives off the hydrogen; with iron and tin, the tin gives off the hydrogen. Since the metal that is least attacked gives off the hydrogen, it can be seen that zinc is a better protective coating for iron than is tin.
At the end of this demonstration, Arthur gives directions for demonstrating the action of couples by using projection equipment. Although the idea seemed to be an excellent one, since projection equipment was not available, this portion of the demonstration was not tested.

**Comment**

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur-----Exp. 124, part 2 p. 321

Procedure

Fill a 1-liter beaker one-fourth full of mossy tin, and add 500 ml. concentrated HCl. Cover the beaker with a watchglass, and heat the mixture almost to boiling. Set aside until no further reaction takes place, hastening the process by warming the mixture whenever it begins to get cool. When all reaction has ceased, filter the resulting solution of stannous chloride through a cotton plug placed in a large funnel. In the bottom of a 30- by 5-cm. diameter glass cylinder, put 1N HCl to a depth of about 10 cm. Using a long-stemmed funnel, run the solution of stannous chloride under the HCl solution, avoiding mixing as much as possible so as to leave the two liquids with the HCl stratified on a 15-cm. layer of stannous chloride. Support a 30-cm. tin rod of 5 to 6 mm. diameter in the cylinder by means of a rubber stopper. The rod should extend down the center of the cylinder, passing through the HCl and far down into the stannous chloride. Set aside where the cylinder will not be disturbed. In 10 to 15 min., a cluster of long crystals of tin
will appear on the rod at the junction of the two liquids. This is due to the fact that tin dissolves in the dilute HCl solution, and practically an equivalent quantity of tin crystallizes out from the concentrated stannous chloride.

Comment

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur----Exp. 125 p. 322

1. In a 600-ml. beaker, place a 50- by 100-mm. porous cup. Fill the cup three-fourths full of 1N sodium chloride; then fill the beaker outside the cup to the same level with 1N HCl. In the sodium chloride solution, place a strip of zinc 1 by 6 in. or a 6-in. zinc rod; in the HCl, place a platinum electrode of the type shown in Fig. 3, p. 51. Connect the zinc and the platinum electrodes to a galvanometer, and note the deflection of the galvanometer needle.

Connect the zinc and platinum electrodes directly by means of a short piece of No. 18 copper wire. In a few minutes, note the formation of bubbles of hydrogen on the surface of the platinum. The reactions at the two electrodes should be explained. It should be explained also that the reaction here is the same as if zinc had been placed directly in the HCl. However, in this case, the energy released during the reaction is electricity instead of heat.

Repeat, using the same method but using iron, cadmium, or some other metal above hydrogen in the displacement series, instead of zinc.
2. Demonstrate the construction of a Daniell's cell. Use it to operate an electric bell.

The Daniell's cell consists of a glass jar from the edge of which is suspended a zinc crowfoot. In the bottom of the jar is a long strip of sheet copper bent usually in a zigzag shape and lying on edge. This is connected to a wire which passes through a glass tube to the top of the jar and serves to make electrical connection to the copper. Covering the copper is a layer of saturated copper sulfate solution with crystals of the latter lying in the bottom of the jar to replace the cupric ions as they are used. Above the copper sulfate solution is stratified a 5% zinc sulfate solution. The zinc electrode dips into the zinc sulfate solution but must not come in contact with the copper sulfate solution.

If it is necessary to make the cell from ordinary equipment, a 800-ml. beaker may be used for the jar. The zinc crowfoot may be replaced by a heavy piece of sheet zinc supported so it occupies a horizontal position just beneath the surface of the stratified zinc sulfate. The copper electrode on the bottom may be connected to electrical apparatus by a copper wire sealed into a glass tube by means of picein cement. The solutions may be stratified by pouring an 8-cm. layer of 5% zinc sulfate solution in the beaker with the copper electrode in place. Then the saturated copper sulfate solution may be run carefully under the zinc sulfate solution.
by means of a long-stemmed funnel.

Another form of this cell consists of a jar with a porous cup, inside which is a zinc electrode dipping into a solution of zinc sulfate. Outside the cup is a copper sheet bent to fit the outside of the cup. High on the copper sheet is built a porous copper cup in which crystals of copper sulfate are suspended. The jar is filled with saturated copper sulfate solution.

Note that the Daniell's cell makes use of the displacement of cupric ions by zinc to form zinc ions and copper.

3. Prepare a porous cup-beaker cell as in (1) above. Inside the cup, place a IN solution of sodium chloride and a carbon or platinum electrode. Outside the cup, place a IN solution of KI, 10 ml. fresh starch solution, and a second carbon or platinum electrode. Connect both electrodes to a large demonstration galvanometer, then inside the porous cup pour 5 ml. strong chlorine water or bromine water. Note the strong deflection of the needle of the galvanometer. Short-circuit the cell by connecting the electrodes with a short piece of copper wire. After a few minutes, note the blue color around the electrode outside the porous cup. The blue color shows that iodine is being set free around the electrode, although the oxidizing agent responsible is inside the porous cup.

This illustrates the formation of electrical energy by
the displacement of one nonmetal by a more active nonmetal. Chlorine water and KBr, or chlorine, bromine, or iodine solution with a saturated solution of hydrogen sulfide, may be used instead of the above solutions, if desired.

Comment

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur----Exp. 126  p. 324

Procedure

1. Prepare an apparatus like that shown in Fig. 3, p.51. Inside the porous cup, place a zinc rod and some 1N sodium bromide solution; outside the cup, use 1N sodium bromide solution and a platinum or carbon electrode. Connect the zinc and platinum electrodes to the demonstration galvanometer; then around the platinum electrode pour 5 to 10 ml. saturated bromine water. Note the large deflection of the galvanometer needle.

Other metals such as iron or copper may be used instead of zinc. Also, sodium chloride solution and chlorine water may be used instead of sodium bromide and bromine water. Point out in each case that the reactions involve combinations just as much as if the metals were placed in direct contact with the nonmetals.

2. Assemble a porous cup-beaker cell as before. Inside the cup, place some 0.1M ferric sulfate solution, and surround the cup with 1N sodium sulfate. Use platinum electrodes in both solutions, and connect the cell to a demon-
stration galvanometer; then to the sodium sulfate solution, add 10 ml. fresh starch solution and 20 ml. 3M potassium iodide solution, and stir. Note the deflection of the galvanometer needle and the appearance of the characteristic blue of the starch-iodide complex.

The reactions involved may be considered either an illustration of the decomposition of ferric iodide to form ferrous iodide and iodine, or a simple reduction of ferric ions by iodide ions.

Similar results are obtained if the ferric sulfate is replaced by copper sulfate. In this case, the decomposition of cupric iodide to form cuprous iodide and iodine is the reaction involved.

Comments

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material
Arthur----Exp. 127  p. 325

Procedure

1. Prepare an apparatus like that shown in Fig. 3, p. 51. Inside the porous cup, place a platinum electrode and some 0.1N sulfuric acid solution. Outside the cup, put a solution made by mixing equal volumes of 1N sulfuric acid and 1N potassium iodide solution. To this mixture, add also 10 ml. fresh starch solution, and insert another platinum electrode into this solution. Connect both electrodes to a large demonstration galvanometer, then into the porous cup pour 10 ml. 1N potassium permanganate solution. Note the strong deflection of the galvanometer needle. After a few minutes, note the blue color of the free iodine-starch mixture around the outer electrode. Show by equations the reaction at each electrode.

2. Repeat, using potassium chromate instead of potassium permanganate. The results are similar.

3. Any oxidation-reduction reaction may be used in place of the foregoing if desired. In this connection, it may be pointed out that combinations, decompositions, and
displacements are really special types of oxidation-reduction reactions.

Comments

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Demonstration found in the source material

Elder----p. 105

Procedure

Into a 10% stannous chloride solution, made up with sufficient HCl to prevent hydrolysis, place copper and carbon electrodes. Attach three dry cells in series and connect the copper to the negative pole of the 3-cell battery and the carbon electrode to the positive pole of the battery. Tin will plate out on the cathode and form strands which look like an inverted tree.

Comments

This is an excellent demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

A new demonstration

Fig. 4
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

A new demonstration

Preparation

1. Prepare 500-ml. of a saturated solution of potassium chloride and 500-ml. of a saturated solution of potassium nitrate. Prepare 500-ml. of a 0.1N solution of magnesium nitrate, aluminum nitrate, zinc nitrate, cadmium nitrate, lead nitrate, copper nitrate, and silver nitrate.

2. Place a 3\(\frac{1}{2}\) inch watchglass upside down in the center of an 8 inch square of blattting paper and trace around it with a pencil. Draw seven arms, each 1\(\frac{1}{2}\) inches wide and 2 inches long, projecting outward from the circumference of the circle. Cut out this pinwheel\(^2\), place it on a 3\(\frac{1}{2}\) inch watchglass, and support the watchglass on a 100-ml. beaker. Around this beaker place 7 100-ml. beakers so arranged that each arm of the pinwheel will dip into a separate beaker. This pinwheel is to serve as a salt bridge connecting two half-cells.

3. Place a small pile (approximately $\frac{1}{2}$ inch square)

\(^1\)Fig. 4, p. 67.
\(^2\)Fig. 5, p. 71.
of mercurous chloride on a glass plate. Add a few drops of the saturated solution of potassium chloride and mix until a thick paste of mercurous chloride is obtained. Mold the mercurous chloride paste into a mound and allow it to harden.

4. Obtain a small piece of each of the following metals: magnesium, aluminum, zinc, cadmium, lead, copper, and silver. The size or shape of the metal is not important so long as it can be gripped easily with a pinch clamp connector and dipped into the solutions contained in the 100-ml. beakers.

5. Make a platinum electrode by sealing a platinum test wire 8 inches long into a piece of glass tubing 3 inches long so that it extends approximately 1 inch from one end and 4 inches from the other. Attach the longer end to a small metal connector supported firmly on a ring stand. To this same connector attach one end of a convenient length of copper wire. Attach the other end of the wire to a voltmeter.

All of the preparation explained above should be made prior to the demonstration.

Procedure

1. Fill the 7 100-ml. beakers with the saturated solution of potassium nitrate and allow each arm of the pinwheel salt bridge to dip into the solution. This saturates the extremities of the salt bridge with nitrate ions and thus prevents the precipitation of insoluble metallic chlorides.
While the potassium nitrate is saturating the arms, saturate the central portion of the pinwheel with the saturated solution of potassium chloride. Make sure that the pinwheel is saturated with salt solution.

2. Add 2 or 3 drops of the saturated solution of potassium chloride to the hardened mound of mercurous chloride. Work the mercurous chloride into a thick paste and transfer it to the center of the pinwheel. Mold the mercurous chloride paste into a mound and make a crater at the top. Using a medicine dropper, place 3 or 4 drops of mercury in the crater.

3. Adjust the ring stand holding the platinum electrode so that the end of the platinum wire dips into the mercury.

4. Empty the potassium nitrate from 1 of the 7 beakers. For convenience, use the one nearest to the voltmeter. Fill this beaker with the 0.1N solution of magnesium nitrate. Attach one end of a convenient length of copper wire to the negative terminal of the voltmeter and the other end to a spring clamp connector. Holding the piece of magnesium with the spring clamp connector, dip it into the magnesium nitrate solution. Observe and record the voltmeter reading. Repeat this procedure using successively each of the other 6 beakers, the solutions of aluminum nitrate, zinc nitrate, cadmium nitrate, lead nitrate, copper nitrate, and silver nitrate, and the pieces of the corresponding metals.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

Fig. 5.---Pinwheel Salt Bridge

Fig. 6.

Mercurous chloride---Mercury
---Gooch crucible
Note

Before the metals are dipped into their respective salt solutions, they should be cleaned to remove any oxide or foreign material from their surfaces.

Better results may be obtained if the more active metals, magnesium and aluminum, are amalgamated slightly by being dipped into a solution of mercuric chloride.

Variation in procedure

Instead of the mound of mercurous chloride in the center of the pinwheel a Gooch crucible may be used to contain the calomel and the mercury.¹

Comments

This demonstration has proved to be very reliable in placing each of the metals tried in its proper order in the activity series.² It allows for a comparison of the activity of several different metals during the same demonstration. It has simplicity, yet it follows closely in principle one of the methods commonly used in assigning electrode potential values to the metals. At the same time it illustrates the use of a commonly used standard reference electrode. The actual demonstration can be completed in approximately 15 minutes.

¹See Fig. 6, p. 71, and picture, p. 74
²See table on page 73 for results obtained by the writer in performing this demonstration.
Metals may be arranged in an activity series according to their tendency to pass into ionic form by losing electrons.

In testing the new demonstration, p. , the writer obtained the results shown in the following table.

<table>
<thead>
<tr>
<th>METAL</th>
<th>VOLTMEtER READING</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1</td>
</tr>
<tr>
<td>Mg → Mg⁺⁺ 2e</td>
<td>1.30</td>
</tr>
<tr>
<td>Al → Al⁺⁺⁺ 3e</td>
<td>0.95</td>
</tr>
<tr>
<td>Zn → Zn⁺⁺ 2e</td>
<td>0.70</td>
</tr>
<tr>
<td>Cd → Cd⁺⁺ 2e</td>
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<tr>
<td>Pb → Pb⁺⁺ 2e</td>
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<tr>
<td>Cu → Cu⁺⁺ 2e</td>
<td>-0.05</td>
</tr>
<tr>
<td>Ag → Ag⁺⁺⁺ e</td>
<td>-0.30</td>
</tr>
</tbody>
</table>
A Demonstration to Illustrate the Activity Series of Metals
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

The influence of pressure on the solubility of gases which are not very soluble in water is summed up in Henry's Law which states that: the amount of a gas dissolved by a given amount of a liquid at a given temperature is proportional to the pressure. It may be expressed algebraically as follows:

\[ c = K_t p \]

where "c" is the concentration of the gas in the solution, "p" is the pressure at which the gas is supplied over the solution, and "K_t" is called the solubility coefficient for the particular temperature "t". It is the concentration when the pressure is unity.

An understanding of this principle is important in both applied and theoretical chemistry. A method for separating oxygen from the air, although not practicable, is based on it. This principle explains why carbon dioxide, having a relatively large solubility in water, is not completely washed from the air during a heavy rain storm and the use of decompression chambers for deep sea divers and sand hogs. It is utilized also in the commercial manufacture of carbonated beverages.
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

Demonstration found in the source material
Arthur----Exp. 37 p. 135

Procedure

Part 1 c. Introduce 10 ml. tap water and 10 ml. freshly boiled water into the vacuums of separate barometers. The unboiled water will drop farther than the other, showing (1) air dissolved in the tap water and (2) the decrease in solubility of air with decrease in pressure.

Part 2 b. Allow a bottle of ordinary soda pop to stand until it has reached room temperature, then slightly loosen the cap and call attention to the hissing as the gas escapes. Remove the cap completely, and transfer 100 ml. of the pop to a 250-ml. Florence flask. Fit the latter with a one-hole rubber stopper and delivery tube. Warm the liquid, and pass the gas through 100 ml. limewater in a test glass to show that the gas is carbon dioxide. Let the students taste the pop before and after heating.

Part 2 c. Display some soda water in a seltzer bottle. Operate the bottle to show the pressure of the gas and to show that when the pressure is relieved the gas rapidly escapes from the solution.
Part 4. In a Hulett eudiometer\(^1\) put 60 ml. acetylene prepared by the action of water on calcium carbide. Level the mercury inside and outside the eudiometer, and mark the position of the mercury in the eudiometer. In the overflow bulb at the top of the eudiometer, put some acetone. Lower the leveling bulb so the pressure of the gas in the eudiometer is less than atmospheric. Cautiously open the stopcock, and let about 10 ml. acetone into the eudiometer. Close the stopcock, and level the mercury. It will be found that the volume of the acetylene is much less than before. Raise the leveling bulb well above the level of the mercury inside the tube, and move it up and down for a few seconds. The increased pressure and motion will cause more of the gas to dissolve. Then lower the bulb to reduce the pressure. Vigorous bubbling of the acetone indicates escape of acetylene from the liquid.

By leveling the bulb between each operation, it can be shown that the acetylene is much more soluble at high, than at low, pressures.

\((\text{It is necessary to level the mercury before each reading if one is to distinguish between the effect of pressure upon the volume of the gas and its effect on the solubility of the gas. This procedure is possible owing to the fact})\)

\(^1\text{See Fig. 7, p. 79.}\)
that after a change in conditions several minutes are required before the new equilibrium point is established.)

Comments

This is an excellent demonstration, but since it is unlikely that many high school laboratories would have a Hulett eudiometer equipped with a two-way stopcock, the writer suggests a modification which he has found to be very successful. See Fig. 7, p. 79.
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

Fig. 7.---Modified Hulett Eudiometer
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

Air and carbon dioxide are accessible in every laboratory and make excellent gases with which to demonstrate this principle. Strictly speaking, however, Henry's Law refers to gases which do not act chemically on the solvent. When carbon dioxide dissolves in water, one portion dissolves in the physical sense and one portion enters into chemical combination until a state of equilibrium is reached, which may be represented by the following equation:

\[ \text{CO}_2 \text{(gas)} \rightleftharpoons \text{CO}_2 \text{(dissolved)} + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ \text{HCO}_3^- \rightleftharpoons 2\text{H}^+ \text{CO}_3^2^- \]

Although it is only the carbon dioxide dissolved in the physical sense that comes within the province of Henry's Law, an inspection of the above equilibrium equation will show that an increase in pressure on the system will shift the equilibrium to the right because of the increased amount of dissolved carbon dioxide, and a reduction in pressure will shift the equilibrium to the left because of the decrease in the amount of dissolved carbon dioxide. Since this shift in equilibrium causes a change in the hydrogen ion concentration which can be detected by the visible color change of a suitable indicator, the change in solubility with change in pressure may be demonstrated with carbon dioxide, as shown in the following demonstrations.
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

A new demonstration

Procedure

1. Prepare a saturated solution of calcium bicarbonate by passing carbon dioxide into a saturated solution of calcium hydroxide until a precipitate of calcium carbonate forms and then dissolves. It may be difficult to dissolve completely all of the calcium carbonate. If the solution remains somewhat cloudy after carbon dioxide has been bubbled through it for two hours, filter it and use the filtrate in part two.

2. Place 25 ml. of the bicarbonate solution in a 100-ml. Erlenmeyer flask fitted with a one hole rubber stopper. Insert a three inch piece of glass tubing into the stopper and connect it by means of pressure tubing to a vacuum pump. (An aspirator or hand vacuum pump such as may be found in many of the smaller high school laboratories will serve equally well.) Reduce the pressure over the bicarbonate solution. Note the vigorous bubbling due to the escape of carbon dioxide from the solution. If the bubbling becomes too vigorous, stop the pump for a few minutes before proceeding. Continue the evacuation for about ten minutes; then seal off the tube with a pinch clamp and let stand.
Beside this flask place another similar one containing an equal amount of the bicarbonate solution prepared in part one. At the end of the period, observe and note whether or not a precipitate has formed. If no precipitate is found, let stand overnight. Observe again. After some time the solution which has been subjected to reduced pressure will show a precipitate of calcium carbonate. The other will remain clear. The reaction may be represented by the following reversible equilibrium equation:

\[ \text{CaCO}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{Ca(HCO}_3\text{)}_2 \]

Once the equilibrium is upset, a certain amount of time is required before the equilibrium is reestablished. Evidence of a precipitate may be noticed in 12 hours or less, indicating that the equilibrium is shifting to the left. It may take, however, 24 to 48 hours for a sizable precipitate to form.

The formation of the bicarbonate solution and the reprecipitation of the calcium carbonate illustrates an interesting reversible equilibrium reaction which plays an important part in the formation of hard water. Limestone caves, with their picturesque stalagmites and stalactites, to a large degree owe their formation to a decrease in the solubility of carbon dioxide caused by reduced pressure.
The solubility of a gas in an inert solvent varies directly with the pressure to which the gas is subjected.

Varying amounts of carbon dioxide dissolved in water will produce solutions of different hydrogen ion concentra-
tion. Ordinary distilled water is often supersaturated with carbon dioxide and will usually have a pH below 5 and often as low as 4.5. The pH of a saturated solution of carbon dioxide in water at one atmosphere pressure and at 25°C. is slightly higher than 4. By varying the pressure over water containing carbon dioxide and using a suitable indicator, such as methyl red, bromoresol green, or a universal indi-
cator, it is possible to show by means of a visible color change that the solubility of a gas in water is directly proportional to the pressure. The change in pressure causes a shift in the following equilibrium:

\[ CO_{2}(\text{gas}) \rightleftharpoons CO_{2}(\text{dissolved}) + H_{2}O \rightleftharpoons H_{2}CO_{3} \rightleftharpoons H^{+} + HCO_{3}^{-} \rightleftharpoons 2H^{+} + CO_{3}^{2-} \]

and consequently a change in the hydrogen ion concentration, which is readily detected by any of the indicators recommend-

1To prepare the indicator solution, dissolve 0.2 g. of the indicator in 60 ml. 95% alcohol and dilute to 100 ml. by adding water. See Lange's Handbook of Chemistry for alternative method.

2Gramercy Universal Indicator Solution may be purchased from Fisher Scientific Co., Pittsburgh, Pa., or from Elmer and Amend, New York, New York. For use, follow instructions supplied.
Procedure

1. Place approximately 10 ml. of ordinary distilled water in a test tube and add 2 to 3 drops of methyl red indicator. Shake well and note the color of the solution. This indicator is red at a pH of 4.4 and yellow at a pH of 6. Stopper the test tube with a one hole rubber stopper containing a short piece of glass tubing. Connect to a vacuum pump (an aspirator or hand vacuum pump will serve equally well) and subject the water to a reduced pressure for 2 to 3 minutes shaking the test tube occasionally. Note the escape of carbon dioxide bubbles from the solution and the change in color from red to yellow, which indicates a decrease in the hydrogen ion concentration.

2. Repeat part one using bromcresol green indicator in place of methyl red. Bromcresol green gives a yellow color at a pH of 4 and blue at a pH of 5.6.

3. Repeat part one using the universal indicator in place of methyl red. The universal indicator gives a red color at a pH of 4 and a yellow green at a pH of 6.

4. Place in a test tube 10 ml. of ordinary distilled water which has stood in contact with the air over night. Point out to the class that the carbon dioxide dissolved in the water is in equilibrium with the carbon dioxide of the air and under a partial pressure of approximately 0.0004
atmospheres. Add 2 or 3 drops of methyl red and pass carbon dioxide over the water until all the air above the water has been replaced by carbon dioxide. Stopper the test tube. Explain to the class that the partial pressure of carbon dioxide on the water is now approximately one atmosphere. Shake the test tube and note the color change from reddish yellow to a deep red, indicating a shift in equilibrium to the right\(^1\) and an increase in the hydrogen ion concentration due to the increased solubility of carbon dioxide under the increased partial pressure. Now connect the test tube to the vacuum pump and note the escape of carbon dioxide from the solution and the change in color from deep red to yellow. Stopper the test tube and save it for use as directed in part seven.

5. Repeat part four using bromcresol green indicator in place of methyl red.

6. Repeat part four using the universal indicator in place of methyl red.

7. Blow into the test tube containing the yellow solution saved from part four. Shake occasionally. The carbon dioxide in the breath will increase the partial pressure enough to change the color from yellow to red. Allow the test tube to stand for a few minutes with occasional shaking. The color will change back to yellow. This test tube may be

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\(^1\)See equation p. 83.
stopped and saved indefinitely. It may be used for an excellent quick demonstration when explaining equilibrium reactions and when explaining Dalton's Law of partial pressures.

Note

In part four, when displacing the air over the solution with carbon dioxide, be careful that no hydrochloric acid from the carbon dioxide generator is allowed to pass into the solution. Otherwise the color change will not reverse when the pressure is decreased. The fact that the color change is reversed when the pressure is reduced is adequate proof that hydrochloric acid is not being carried over and causing an increase in the hydrogen ion concentration. The hydrochloric acid would react with the water, forming a stable solution that would not be changed appreciably by the change in pressure.

Comment

This demonstration may be very useful in explaining Dalton's Law of partial pressures, which, when applied to a mixture of gases exposed to the action of a solvent, is really a simple extension of Henry's Law. It may be summed up as follows: When a mixture of gases is exposed to the action of a solvent, the amount of each gas which is dissolved by the solvent is proportional to its partial pressure. Each gas behaves as if the others were absent.
A Demonstration to Illustrate the Effect of Pressure on the Solubility of Gases
CHAPTER IV

THE MIGRATION OF IONS

In offering an explanation of the demonstration found on page 68 illustrating the activity series, it will be necessary to discuss the phenomenon of ionic migration. The writer, therefore, is including in this chapter two demonstrations to illustrate this phenomenon. They also may be used to supplement other demonstrations employing a chemical cell.

One of the new demonstrations is really a modification of a commonly used one\(^1\) which makes use of the blue color of the cupric ion and the orange-yellow color of the dichromate ion. The writer, however, believes that it has decided advantages over any demonstration found in the source material. It employs only simple apparatus and may be made ready in less than three minutes. It eliminates the commonly used gelatin dispersion, which takes time to prepare and which often gives a blue color much deeper than that due to the cupric ion.

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\(^1\)This demonstration is described and illustrated in each of the following:

The other new demonstration shows that ions will migrate not only in an aqueous solution but also in a molten substance, in a seemingly solid one (actually a supercooled liquid) such as glass, and in a gas. It is a spectacular demonstration from the viewpoint of the student, since it shows a metallic deposit forming on the inside of a light bulb while the bulb is in operation. It also offers an excellent problem for investigation by superior students.
Fig. 8.—Apparatus to Show the Migration of Ions
The Migration of Ions

A new demonstration

Procedure

1. On a large piece of filter paper place 30 g. of sodium nitrate and 30 g. of sodium nitrite, and thoroughly mix with a spatula\textsuperscript{1}. Transfer the mixture to a 100-ml. beaker supported on an asbestos centered wire gauze over a Bunsen burner\textsuperscript{2}.

2. Obtain a small 110 volt, 15 watt night light bulb and support it on the same ring stand that holds the salt mixture in such position that it can be lowered into and raised out of the mixture easily. Connect one end of a convenient length of copper wire to the filament of the bulb and connect the other end to the negative pole of a six volt storage battery. Melt the salt solution and adjust the flame so that the mixture remains in the fluid state. Support a graphite rod on another ring stand and place the stand so that the graphite rod dips into the molten mixture. Connect the graphite rod to the positive terminal of the six volt battery and lower the light bulb into the mixture so that a portion of the filament is beneath the surface. Light the

\textsuperscript{1}The melting point of sodium nitrate is 308°C. That of sodium nitrite is 284°C. That of the mixture is approximately 225°C.

\textsuperscript{2}See Fig. 8, p. 90, and picture, p. 94.
bulb.

The nitrate ions travel to the cathode and are reduced. The sodium ions migrate to the hot filament inside the bulb and are oxidized to metallic sodium. The hot filament vaporizes the metallic sodium which plates out on the upper, cooler part of the bulb. A good mirror is obtained in 30 - 45 minutes.

Note

The writer confirmed that the metallic substance was sodium by: (a) the ordinary flame test using a platinum test wire, (b) spectrum analysis, and (c) microscopical qualitative analysis, first using uranyl acetate, and second using bismuth sulfate\(^1\). He proved that the sodium came from outside the bulb by weighing the bulb under controlled conditions both before and after plating it. He observed a gain in weight of 4.5 mg. after running the current for 2 hours.

Comments

This demonstration may be performed within the time limit of a regular class period. An explanation may be given to the class and the following questions may be raised:

- Is the metallic deposit sodium?
- Did it come from outside the bulb?

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The instructor may suggest that someone take these questions as the challenge for a problem for further investigation. The following procedure is suggested for this investigation.

1. Wash the bulb with water and wipe it dry with a clean towel. Weigh the bulb to constant weight on an analytical type balance (If an analytical type balance is not available, a pulp balance will serve very well). Record the weight of the bulb. Carry out the plating following the procedure given in parts 1 and 2, p. 91, for 2 hours. Remove the bulb from the molten mixture and allow it to cool. Wash the bulb with water, dry it with a clean towel, and weigh to constant weight. Record the weight and compare it with the weight before plating.

2. Make a deep scratch with a file approximately \( \frac{1}{4} \)" from the tip of the bulb. Heat an iron rod (1/8 to 1/4" in diameter) and place it in the scratch. The glass will crack in a circular direction around the bulb. Follow the crack with the hot rod, heating the rod when necessary. In a short time the crack will extend completely around the bulb and the head of the bulb will drop off.

3. As soon as the top of the bulb is removed, place 5 to 6 drops of water (caution) in the bulb. Note signs of vigorous reaction. Test the solution with litmus paper. What metals, when placed in water, will give this test? Pour the solution onto a watch glass and apply the flame test.
A Demonstration to Illustrate the Migration of Ions
The Migration of Ions

An improved demonstration

Procedure

1. Prepare 200 ml. of N copper sulfate and 200 ml. of N potassium dichromate. Note the color of the solutions.

2. Mix 75 ml. of the N copper sulfate solution with 75 ml. of the N potassium dichromate solution and add a few drops of dilute sulfuric acid to prevent hydrolysis. Note the color of the mixture.

3. Arrange three 100-ml. beakers in a row and fill the two end beakers three-fourths full of dilute sulfuric acid. Support a graphite rod on a ring stand so that it dips into the sulfuric acid in one of the end beakers and connect it by means of a convenient length of copper wire to the positive terminal of a six volt battery. Support a second graphite rod similarly in the other end beaker and connect it to the negative terminal of the battery.

4. Obtain a cotton roll such as are used by dentists.¹ Saturate two inches at each end of the roll by dipping it into dilute sulfuric acid. Leave the center two inch portion dry. Now place the roll over the middle beaker so that one end just dips into the sulfuric acid contained in each of the

¹Johnson & Johnson absorbent cotton rolls No. 2 - 6" by 3/8" may be obtained from any dentist or dental supply house. Many drug stores also carry them. They may be purchased for less than a cent each. Besides serving as an excellent salt bridge, these rolls may find many other uses in the laboratory.
A Demonstration to Illustrate the Migration of Ions in a Salt Bridge
other two beakers\textsuperscript{1}. Using a medicine dropper, saturate the center portion of the roll with the cupric dichromate mixture prepared in part two. Pour the remainder of the solutions prepared in parts one and two into suitable flasks or beakers. Place the cupric dichromate mixture behind the middle portion of the cotton roll. Place the copper sulfate solution next to it on the side nearer the beaker containing the negative electrode and the potassium dichromate solution in a similar position on the side nearer the positive electrode.

In approximately five minutes it is observed that the blue color of the cupric ion is moving toward the negative electrode and the orange-yellow color of the dichromate ion is moving toward the positive electrode. In approximately thirty minutes a complete separation of the colors may be observed. A color comparison may be made between the different sections of the cotton roll and the solutions placed behind it.

\textsuperscript{1}See Fig. 9, p. 98.
Fig. 9. -- The Migration of Ions in Aqueous Solution
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