ABSTRACT
        This student guide is for the second year of the
Portland Project, a three-year integrated secondary school science
curriculum. "Motion and Energy," the first of the two parts in this
volume, begins with the study of motion, going from the quantitative
description of motion to a consideration of what causes motion and a
discussion of Newton's laws. There follows the development of the
laws of conservation of momentum and energy, including a discussion
of energy in biological systems, and culminating with a discussion of
the kinetic molecular theory. The major objective of "Chemical
Reactions," part two, is to build some of those basic chemical
concepts that are necessary for the analytical approach in parts
three and four of the third year course: "The Chemistry of Living
Matter" and "Energy Capture and Growth." Some of the topics
considered in part two include the mole concept, equation writing,
energetics associated with chemical reactions, the dynamic nature of
particles and their interactions, and the application of energy and
equilibrium to chemical systems. The Harvard Physics Project textbook
and the Chemical Educational Materials Study (CHEMS) textbook are
used for most of the reading assignments in the second year course. A
review of the development of the three-year program, its rationale
and content, and a three-year course subject outline are included.
(Also see SE 013 702, SE 013 703, SE 013 705, and SE 013 706.)
(Author/PR)

Year Two

MOTION AND ENERGY
CHEMICAL REACTIONS

PARTS ONE & TWO
OF
AN INTEGRATED SCIENCE SEQUENCE
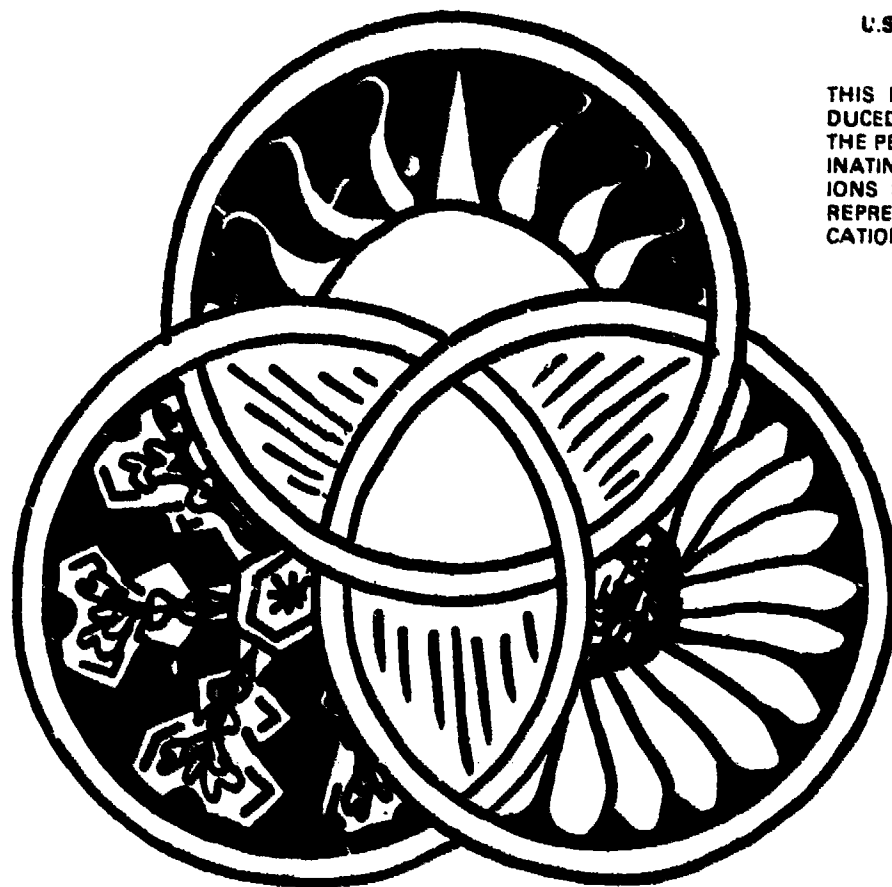
STUDENT GUIDE

1971 EDITION

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STUDENT GUIDE
MOTION & ENERGY
CHEMICAL REACTIONS

AN INTEGRATED SCIENCE
SEQUENCE

1971 EDITION

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Portland, Oregon

Co-Director:

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Portland State University
Portland, Oregon

ERIC

Table of Contents

Dedication

This volume is dedicated to the memory of
Vernon Cheldelin under whose guidance and
leadership integration of the sciences for
Oregon secondary school youth was begun in
1963.

Introduction

The Portland Project was initiated in the fall of 1962 when two
secondary school teachers, one with background in CBA chemistry, the
other having responsibility for PSSC physics, began to note and discuss
the redundancy in their respective courses. Why should students be
subjected to this repetitious and fragmented representation of the physical
sciences? they asked. A Steering Committee met to pursue the problem
further and perhaps enlist the support of a funding organization to
permit its exploration in depth. Under the able and devoted leadership
of Vernon Cheldelin, Dean of the School of Science at Oregon State University (deceased), two proposals prepared for support by the National Science
Foundation were funded in the summers of 1963 and 1964.

Thirty-five scientists and teachers devoted various quantities
of time as writers, consultants, pilot teachers, and evaluators, with
the aim of ascertaining the feasibility and efficiency of the integration
of chemistry and physics. Concurrently and subsequently, other groups in
other parts of the country have carried on studies that are approximately
parallel to this one.[1] Though the conceptual development and points of
emphasis differ, the various groups are satisfied that integration of
science courses is not only feasible but highly desirable.

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[1]Federation for Unified Science (FUSE) was recently organized to
act as a clearinghouse of information on integrated science courses.
Victor Showalter at Ohio University is the chairman of this committee.

Dr. Michael Fiasca of the Education and Science Staffs of Portland

State University conducted an evaluation which revealed that subject

matter achievement in chemistry and physics and critical thinking

abilities are enhanced among students who studied the integrated

courses over tnose who study the separate disciplines of chemistry and

physics. It should be emphasized that though these differences were

apparent, it could not be demonstrated that they were statistically

significant.[2] A concomitant result showed that enrollments in the

two-year integrated courses were dramatically greater than in the

separate courses.

A survey completed April 16, 1967 showed that there were

forty-four schools in twenty states using the Portland Project

integrated chemistry-physics course.[3]

Mounting evidence in the literature from prominent persons

working in science education strongly supported this mode of organiza-

tion. Dr. Jerrold R. Zacharias, the prime instigator of the PSSC

physics program, exemplified the changing attitude of scientists and

educators:

> The division of science at the secondary school level,
> into biology, chemistry, and physics is both unreason-
> able and uneconomical.

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[2]Detailed results of this study may be obtained by writing to
Dr. Fiasca at Portland State University.

[3]Detailed enrollment figures and addresses of people who are
using the Portland Project courses may also be obtained from Dr. Fiasca.

Ideally, a three-year course that covered all three disciplines would be far more suitable than a sequence of courses which pretends to treat them as distinct. Today such a three-year course would be difficult to fit into the educational system, but much of this difficulty might be overcome at once if such a course existed, and it might well be that present tendencies in education would soon overcome the rest.

In any case, a greater coordination of the three subjects is possible even within the existing framework. It is understandable that the groups which developed the existing programs, each of which faced great problems of its own as it worked toward its goals, were reluctant to embark on the larger task of giving coherence to the sum of their efforts. With the programs now complete or approaching completion, it may be that the time has arrived for this necessary step.[4]

Stimulated by the apparent success of their original work towards this kind of integrated course, persons close to the Portland Project began to discuss extension of their work to include biology with chemistry and physics in a three-year sequence. A third proposal was prepared in 1966 and granted support by the National Science Foundation. Dr. Arthur Scott, member of the Chemistry Department at Reed College who has had deep interest in the Portland Project since its inception, graciously offered his talents, energy and time to carry on the project after Dean Cheldelin's death.

A writing conference was conducted on the Portland State

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[4]From page 52 of Innovation and Experiment in Education, a Progress Report of the panel on Educational Research and Development to the U.S. Commissioner of Education, the Director of the National Science Foundation, and the Special Assistant to the President for Science and Technology, March, 1964.

University campus during the summer of 1967 to develop materials such
as teacher and student guides. Eight local pilot schools committed
approximately five hundred students and twelve pilot teachers for
testing and evaluation. Dr. Donald Stotler, Supervisor of Science
for the Portland School District, has had an active part in this and
other phases of this project.

Twenty-six persons whose functions were riting. consulting,
analysis, and editing met on the Portland State campus beginning
June 14, 1967, to begin preliminary work on the integrated course.
Their first task was to formulate an outline that displayed logical
content development utilizing concepts out of biology, chemistry and
physics. Particular attention was paid to matching students'
abilities, interest and maturity level with the sophistication of
concepts as nearly as this was possible to do. Then the committee
perused material developed by the national curriculum groups --
PSSC, Project Physics, CBA, CHEMS, BSCS and IPS -- in search of
material to implement the outline they constructed previously. In
the absence of appropriate materials, major and minor writing projects
were initiated.

The writing committee continued its work in the summers of
1968 and 1969 with Dr. Karl Dittmer, Dean of the Division of Science,
as director. Four major projects were tackled and completed: (1)
extensive revisions were effected in the three-year outline, (2) the
first and second year courses were revised based upon student and

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teacher feedback, (3) the third-year course was developed incorporating Harvard Project Physics materials as a main vehicle, and (4) an evaluation program for the three-year course was developed.

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Portland, Oregon

Pilot Schools

The following schools have used the integrated course during
all or part of the past three academic years:

Adams High School
Portland, Oregon
  Glen Hampshire
  Lloyd Meskimen
  Thomas Miles

Aloha High School
Aloha, Oregon
  Mary Lou Combs
  Elvis Dellinger
  Nelson Doeleman
  Ted Parker

Beaverton High School
  Roger Berg
  Jean Halling
  Lois Helton
  H. Dean Smith

Benson Polytechnic School
Portland, Oregon
  Howard Browning
  W. B. Chase
  W. L. Hoffman
  Michael E. Lopez

Central Catholic High School
Portland, Oregon
  Jacob A. Mosbrucker
  Peter Roerig

Cleveland High School
Portland, Oregon
  John Brown
  Edmund McCollough
  Ralph Parton

Franklin High School
Portland, Oregon
  John Neeley
  Joseph Sklenicka

Jefferson High School
Portland, Oregon
  Ronald Kawamoto
  Leslie Morehead
  Kenneth Starbuck

Parkrose High School
Parkrose, Oregon
  Donald Pearson

Rex Putnam High School
Milwaukie, Oregon
  Dennis Axness
  David Cox
  Jerry Fenton
  Henry Kilmer
  Jack McGoldrick

Roosevelt High School
Portland, Oregon
  Renee Bergman
  Kenneth Fuller
  Sue Storms

Sunset High School
Beaverton, Oregon
  Leonard M. Dooley

Wilson High School
Portland, Oregon
  Curtis D. Guthrie
  Norman Sipple

Woodland High School
Woodland, Washington
  Donald G. Fry
  George L. Stroud

Acknowledgements

Three-Year Course Rationale

The decision to try to develop a three-year integrated science course which would replace the traditional three courses in biology, chemistry, and physics is based on several considerations. Among them are:

(1) A conviction that modern developments have made the division of science under these three headings obsolete;

(2) A recognition that the traditional courses overlap in many areas, resulting in a great deal of duplication and repetition as in the gas laws, atomic and nuclear structure, calorimetry, and the kinetic molecular theory;

(3) A feeling that terminal students, who take no more than one year of science, deserve to get a taste of all of science rather than just one aspect, as they do in the conventional programs; and

(4) A desire to emphasize the unity in the approach to natural phenomena and the similarity in the methods, techniques, and apparatus used by scientists in all fields.

A natural question arises as to what distinguishes this course from a general science course expanded to three years. The answer is that this course does not consist of a number of unrelated topics that might be taken up in any order; rather, it treats science as a structure that proceeds from observation to the development of general principles and then to the application of those principles to more involved problems.

The emphasis in a general science course is on the results of science; the emphasis here is on the methods and reasoning by which scientists have arrived at these results.

The three-year course outline shows that a number of topics such as properties of matter, energy, heat, and certain biological concepts are discussed at the first-year level and again later in the course. This recycling is deliberate. It is intended to introduce students in a semiquantitative way to some of the significant generalizations of science and to show how these generalizations arose. These topics are treated again in the second and third years when greater facility with mathematics on the part of students makes it possible for them to understand and appreciate discussions of these topics in the succinct and precise language of mathematics.

An excessively formal and quantitative approach is avoided in the first year for several reasons. Students at this level do not extract essential meaning from such a presentation of information; furthermore, first encounters with new ideas should proceed from an intuitive, non-quantitative confrontation to one that is more quantitative. Teachers have spoken out against teaching and learning methods which substitute equations, formulas and other quantitative representations for first-hand experience, word descriptions, examples, and illustrations. These criticisms are just as valid for students who are very capable and very interested in science as they are for other students. Moreover, the mathematical

sophistication of students at this level is such that they are unable to follow most mathematical arguments as explanations for natural phenomena.

The typical science experience of most secondary school students consists of one or two years devoted to general science and biology. Few study physics and chemistry. A significant advantage to the course of study described here is that students are given a chance to study physics and chemistry at a level of rigor that is consistent with their ability and their mathematical maturity. Students who terminate their study of science at the end of one year get a significant exposure to the structure of biology, chemistry, and physics as they are presented in the latest curricular developments. Students who might not elect science beyond the first year because of lack of interest in biology may be attracted by the chemistry or physics portions of the course and elect to take an additional year or two of science. Students who are "turned on" by biology may wish to pursue further study of biochemical topics in Years II and III.

First Year Course

After considering these problems and goals, the general course outline for the first year of the course was derived. It consists of four main parts:

(1) Perception and Quantification

(2) Heat, Energy and Order

(3) M'ce and Men

(4)   Environmental Balance?

The year begins with a study of the perceiver, moves on to the perceived, and ends with the interaction of the perceiver with the perceived. The first-year student starts out by gaining a better awareness of the nature of his perception and senses -- the faculties that make him aware of the world around him. With an increased understanding of these perceptual abilities, he can turn to the environment and then relate himself to it. He finds that his perception is limited and that he often needs to call on technological and conceptual extensions and that even these have their limitations.

The importance of organization and classification as parts of perception is emphasized. The physical properties of matter are introduced and studied as aids in organization and classification of chemicals. The identification of unknowns by study of their physical properties and use of organized data on punch cards is the culminating experiment of the Perception unit.

Apart from the great diversity exhibited in nature, which the scientist must organize in order to comprehend, certain unifying principles are essential for deeper understanding. The most powerful of these is the energy concept, which is explored in the "Heat, Energy, and Order" unit in several of its ramifications - physical, chemical, and biological. The discussion begins by developing an experientially important energy form, viz., heat. The macroscopic aspects of heat as embodied in calorimetry are related to the microscopic in terms of

random molecular motion. This builds confidence in the idea of the
atomic nature of matter, which is essential to much of the unit.
Various energy conversions form the vehicle for extending and
generalizing the energy concept. Nuclear energy is developed in suf-
ficient detail to underscore its environmental and social significance.
Finally, the thermodynamic limitations and implications of energy
conversion are explored, ending with a view of life as a supremely
artful organizer in nature, a mechanism powered by energy which
creates wondrous "local order", but always at the expense of influ-
encing its environment.

The growth of a mouse colony carries the thread of the unit
"Mice and Men." As it develops, students learn many things about
the concept of population. The food and water consumed and products
eliminated tie the mouse colony back to the unit "Heat, Energy and
Order" and point ahead to the chapter on communities and to the unit
"Environmental Balance?".

The cell concept is given prime position in this unit and is
used to enter topics on reproduction, embryology and maturation which
are observed in the mice and other organisms. The mice selected for
the original colony are such that an experiment in Mendelian genetics
comes out of the observations students make as the colony develops.
In most of the chapters man is an important organism and receives as
such attention as the mouse although the data are often secondhand.

A rather unpleasant fact that must be faced is that as our

population increases and human activities are directed towards increasing the standard of living for this population, strains are placed upon the environment. As students discovered in "Mice and Men," the size of the community has a relation to both the quantity of the food, water, and energy required and the quantity of waste products produced. To develop the concept of a closed system and point out the necessity for environmental management, an analogy between the earth and a spaceship is made. Students are then introduced by a multi-media approach to the nature of some of our common pollutants (with emphasis upon air, water, heat, noise, and radiation) as well as their effects. Following this students are encouraged to undertake a rather detailed study of a particular type or aspect of pollution. Emphasis here is placed upon student activity, which may take any number of forms. The culminating activity centers around discussion of these special studies together with the complex relations involved within the environment. It is hoped that out of these studies students will become aware of threats which exist to man's future on this planet.

Second Year Course

The second year of the course is considerably more quantitative in its approach than the first. This is the case because (1) the students are one more year along in their mathematical preparation, (2) the students who elect to take a second year of science are more likely to exert the effort to master more difficult topics, and (3) many of the quantitative aspects of physics and chemistry are basic to an

understanding of molecular biology, which is an important part of the following year's work.

The second year consists of two parts:

(1) Motion and Energy

(2) Chemical Reactions

Year II begins with the study of motion, going from the quantitative description of motion to a consideration of what causes motion and a discussion of Newton's laws. There follows the development of the laws of conservation of momentum and energy, including a discussion of energy in biological systems. This section, which is primarily mechanics, culminates with a discussion of kinetic molecular theory.

Due to recent advances in both molecular biology and biochemistry, the descriptive approach to biology has gradually given way to one that is primarily analytical. It is now necessary, even on the high school level, for the serious biology student to have a more thorough understanding of those concepts normally embodied in the "modern" high school physics and chemistry courses. The major objective of "Chemical Reactions" is to build some of those basic chemical concepts that are necessary for an analytical study of "The Chemistry of Living Matter" and "Energy Capture and Growth."

The following subtopics of this section help in the realization of the major objective: Some of the topics discussed are the mole concept, equation writing, energetics associated with chemical reactions,

the dynamic nature of particles and their interactions and the applica-
tion of energy and equilibrium to chemical systems.

Third Year Course

Year III consists of four parts:

(1) Waves and Particles

(2) The Orbital Atom

(3) Chemistry of Living Matter

(4) Energy Capture and Growth

The underlying rationale of the third year is a study of energy
and its importance to life. The first thrust is to build the orbital
model of the atom using, as background, waves, electromagnetism and
historical models of the atom. Once the orbital model is established
as a representation of the localization and directionalization of
electronic energy, structural models are built to show how biopolymers
are spatially arranged and experiments are done to give evidence of
energy relationships. With shape, size and energy relationships of
molecules established, the DNA molecule is introduced. The culmination
of this work comes in the final section when photosynthesis is considered.
With this topic, much that has gone before is brought to a logical focus.

These topics are most appropriately placed in the third year of
the integrated sequence after students have developed some facility with
basic ideas from chemistry and physics - e.g., quantitative knowledge
about energy, mechanisms of chemical reactions, equilibrium, rate of
reaction, the photon and wave nature of light, electrical phenomena,

and kinetic molecular theory. They should not now simply parrot bio-chemical processes such as photosynthesis and cell respiration but should understand the many chemical and physical principles which underlie these processes.

Time is allotted at the conclusion of Year III for individual investigation and studies.

Three-Year Course Outline

TOPIC	REFERENCE
First Year	
Part One: Perception and Quantification	
I. Sensing and Perceiving	PP*
II. Measurement, Distribution, Organization and Communication	PP
Part Two: Heat, Energy and Order	
I. Heat	PP
II. Temperature and Chaos	PP
III. Energy	PP
IV. Nuclear Energy and Radioactivity	PP
V. Trends in Nature	PP
Part Three: Mice and Men	
I. Reproduction and Development	PP
II. Genetics	PP
III. Genetics and Change	PP
IV. Populations	PP
V. Ecology	PP
Part Four: Environmental Balance?	PP

*PP designation signifies materials produced by the Portland Project.

TOPIC	REFERENCE

Second Year

Part One: Motion and Energy

	REFERENCE
I. Motion	HP*
II. Newton Explains	HP
III. Multi-Dimensional Motion	HP
IV. Conservation	HP
V. Energy - Work	HP
VI. Kinetic Theory of Gases	HP

Part Two: Chemical Reactions

	REFERENCE
I. The Mole as a Counting Unit	PP
II. Combinations of Gases	PP
III. A Useful Form of P=kDT	PP
IV. Chemical Equations	PP
V. Electrical Nature of Matter	CHEMS+
VI. Basic Particles	CHEMS
VII. Energy Effects in Chemical Reactions	CHEMS
VIII. Rates of Reactions	CHEMS
IX. Equilibrium	CHEMS
X. Solubility	CHEMS
XI. Acid-Base	CHEMS

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*HP designates Harvard Project Physics material.
+CHEMS designates material derived from the Chemical Educational
Materials Study.

MOTION AND ENERGY

A Note to Teachers

This Guide is in no way meant to supplant other Teacher's Guides. It has been produced to help bridge gaps which result from the integration. It is imperative that the participating teachers become familiar with the material of all guides.

The experienced classroom teacher knows that time is always a vital factor. One objective of this integration is time. The teacher should search for that correct emphasis which permits him to cover each subject in depth and yet not belabor it such that future materials are not reached. A "development section" and a "teacher organization sheet" are included to help the teacher find that correct emphasis. The teacher should be particularly alert to the tendency to "bog down" early in the course at the expense of materials to come later.

Symbolism has been used in the teacher's guide. The following is a list of the symbols and their meanings.

PSSC	Physical Science Study Committee
CHEMS	The Chemical Education Materials Study
PP	Portland Project
HP	Harvard Project Physics
TG	Teacher's Guide
CEx.	CHEMS Exercise
CExpt.	CHEMS Experiment
HPEx.	Harvard Project Exercise
HPExpt.	Harvard Project Experiment
R	Harvard Project Reader

A PHILOSOPHY FOR USING HARVARD PROJECT PHYSICS MATERIALS

Designed for each topic are many different activities, such as text material, demonstrations, transparencies, student activities, film loops, experiments, films, problems and reader articles. To utilize most effectively these various options a class may do one of three things: (1) select those options which are most practical or most effective for the class; (2) use nearly all options in carefully planned and prepared class meetings; or (3) let different students be responsible for different options. Each of these three methods may be used alternatively or any one might always be used or a compromise to fit the occasion and alternatives available could be arranged. Each of the three methods has drawbacks, however: (1) If only certain options are chosen the time is more flexible but the opportunities for approach to a student on a given topic through various media is narrowed. (2) If all options are available and used, time for class discussion is limited and a strain is placed on the teacher for planing and pre-class set up. (3) If each option is used by teams of students, time must be made available for reports to the rest of the class which is a problem with periods of less than 50 minutes.

We have therefore attempted to give an idea of priorities as we see them, based on activities which will carry the concept of the section. Additional activities are always helpful, and specific teacher background will help determine which activities may be chosen. As with any of the media, reading in the text may be considered as a resource, to be used when other activities don't convey information or as a supplement to the other activities. A possible priority for activities might be

    laboratory experience
    demonstration
    student activity
    film loops
    film
    transparency
    text
    home desk lab problems
    reader articles
    lecture

A two week time cushion has been planned in addition to time alloted for the material in the course. This cushion may be used for post lab, pre lab, or seminar discussions and testing during the fourteen week period.

Suggestions for Laboratory Procedures

A laboratory is a place where scientists look at phenomena under controlled conditions. It is a place for serious work. Always prepare for an experiment by reading the directions in the manual before you come to the lab. Make a special effort to know all precautions.

Do only the experiments approved by your teacher. If you wish to do an extension (this is encouraged), check with your teacher. This general rule is for the safety of you and your fellow students. Laboratory safety is as much an attitude as a set of rules. The lab will become a safe place for investigation if the student continually uses common sense about his safety and the safety of others. If any accident does occur, report to your teacher. What seems a minor injury may have severe consequences.

You will be asked to write laboratory reports. Opinions concerning the content of these reports vary greatly. It follows that teacher judgment will determine the type of laboratory reports you are asked to write. The following ways to improve laboratory reports are to be taken as suggestions only.

(1) Mistakes should not be erased. If there is room for the correction, the mistake should be crossed out without obliterating it and the correction made. If there is insufficient room, an extra piece of paper should be added.

(2) Spelling and punctuation are important. Sentence fragments should be avoided.

(3) The report should be carefully planned. It is best to know what type of observations should be sensed and, if possible, what regularities can be found. Planning will lead to the placement of items in a logical sequence in the report.

(4) The name of the experiment should be included.

(5) The date on which the experiment was done should be included.

(6) The names of all participants should be included and the name of the person who actually prepared the report should be designated.

(7) Some reports should include a simple statement or schematic diagram of the apparatus used in the investigation.

(8) Some reports will require a brief explanation of purpose and procedure. If these are given in the laboratory manual, they should not be included in the report. Copying items is "busy work."

(9) Nearly all experiments require taking measurements and subsequent collection of data. This must be carefully tabulated. If it is possible for you to make data tables before coming to the laboratory, you will have more time for observation, which is a major part of any laboratory experience.

(10) If computations are required to interpret results, they should be included in the report. However, if several computations of a similar nature are needed, they should be illustrated with a typical example. Mathematical equations, not arithmetical operations, should be shown.

(11) If the investigation could be altered to get better results, a statement to this effect should be included.

(12) If the investigation suggests extensions, these should be described.

(13) Reading professional reports from magazines such as The Journal of Chemical Education and Scientific American should result in better reports.

(14) Many times the most significant information about the experiment is to be found by graphing results. Whenever appropriate, graphs should be included in the report; they give a picture from which regularities can be sought. You will find the following suggestions very helpful.

(a) Always use a full sheet of graph paper.

(b) Position the ordinate and abscissa far enough from the edge of the paper to allow proper labeling.

(c) Assuming a relationship exists, the abscissa should represent the independent variable; the ordinate, the dependent variable. As an example: The distance of the gas pedal from the floorboard in an automobile would be the independent variable, plotted on the x axis, while the speed of the car would be the dependent variable, plotted on the y axis.

(d) Each axis must show units - e.g., cm/sec.

(e) Labeling of each axis should run parallel to the axis.

(f) The scale of each axis should be chosen such that the functional plot covers most of the graph paper.

(g) The name of the graph, the name of the experiment and the date of the experiment should be suitably placed on the graph.

(h) When plotting data, draw a circle around each point to indicate the uncertainty associated with the measurements.

(i) Draw the smoothest possible curve suggested by your data.

MOTION AND ENERGY

Outline:

CHEMICAL REACTIONS

GRAPHICALLY SPEAKING . . .

A. GRAPHS AND EQUATIONS

You may recall from your math classes the meaning of the words "variable," "functional relationship," "graphical analysis," etc. The purpose of the exercises you are about to perform is to help you understand what these concepts have to do with the world of experimental science, i.e., the world outside of mathematics in which you observe, measure and predict the behavior of real objects.

First let us review some of these ideas from the viewpoint of a scientist. In some experiments we measure more than one quantity simultaneously. If there is some sort of necessary connection between the varying quantities, that is, if there is some sort of law of nature or mechanism which relates changes in one quantity to changes in another, the quantities are called <u>variables</u>. For

instance, in Year I you observed the motion of a pendulum. The position of a pendulum bob varies as time goes on. The distance of the bob measured along a ruler can be considered one variable (call it "d"), while the time observed on a stopwatch can be considered the other variable (call it "t").
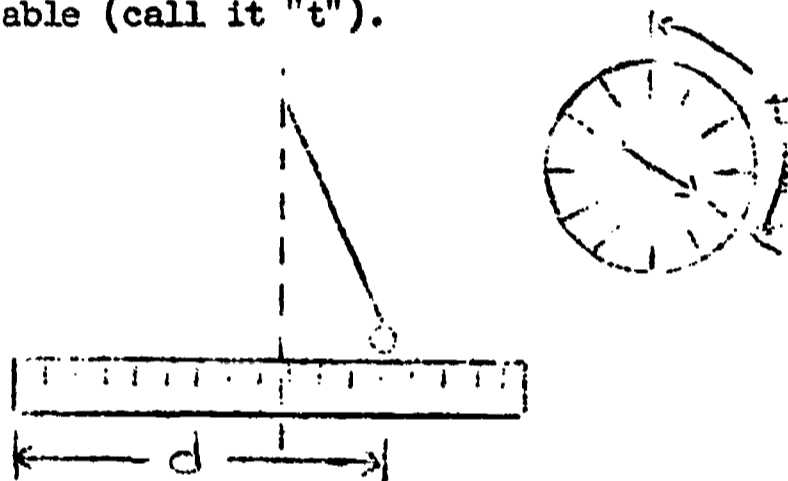
The fact that the two measured quantities have some well determined connection to one another is often stated by saying that one variable is a function of the other, i.e., that they are functionally related. (In the above example d is a function of t.) There are several common ways of expressing this relationship: (1) a table can be made which shows values of one variable for many

corresponding values of the other; (2) a
graph can be drawn in which values of one
variable are plotted on one axis and values
of the other are plotted on the other axis;
and (3) an equation can often be found
which simultaneously contains the two vari-
ables. When a functional relationship has
been expressed in one of these three forms,
it is often possible to transcribe this
information directly into the other forms.
In these exercises we will concentrate
on graphs and equations. In particular
we shall show how to transform the infor-
mation contained (in some simple graphs
such as might be obtained from experi-
ments) into equation form.

In plotting graphs we consider one
of the quantities to be the independent
variable and plot it on the horizontal
axis. The independent variable is
usually thought of as the quantity whose
value is imposed on the experiment from
the outside. For instance, in measuring
the performance of a racing car it might

be the amount of gas which the driver de-

cides to feed the engine. In the case of

the pendulum it would be the time $\underline{t}$. The

other quantity is thought of as represent-

ing the resultant behavior of the system

and is called the dependent variable. In

the case of the racing car it might be

speed; in the pendulum it would be $\underline{d}$. The

dependent variable is plotted on the ver-

tical axis. A graph of the pendulum

motion would appear as shown below:
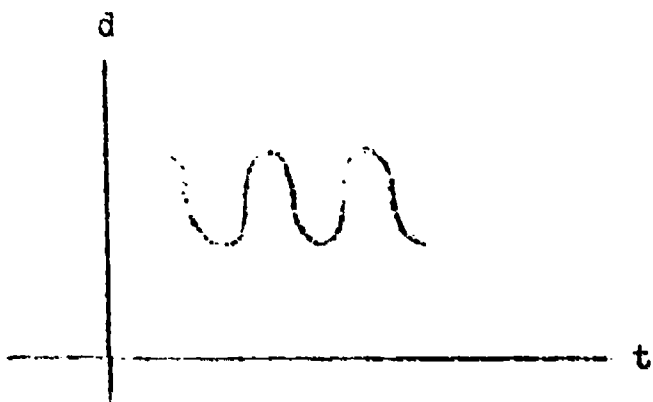
When a functional relation is ex-

pressed in equation form you all know how

to obtain a corresponding graph. For

practice you should try plotting the

following equations: $y = 2x$; $y = 6/x$;

and $y = x^2$. Consider $\underline{y}$ to be the depen-

dent variable, and $\underline{x}$ to be the indepen-

dent variable. Note the characteristic shapes of these graphs. They are of the type you shall be most concerned with in this exercise. While the equation for the pendulum motion is a little more complicated, we shall also do that before the course is over.

The reverse process, i.e., obtaining an equation when you already have the graph, may not be familiar to you. However, it is extremely important in experimental science. Our main task now is to learn how to do this in some important simple cases.

B. FROM DATA TO AN EQUATION

When experimental data is plotted directly on a graph, it may be possible to write the equation for the resulting curve. In the cases we shall be studying the process is two-fold. The first step is to discover the proportional relationship, if any, that exists between the variables. The second step is to de-

termine the value of a proportionality
constant which enables us to transform
the statement of proportion into a bona
fide equation.

B.1 PROPORTIONAL RELATIONSHIPS

The appearance of some typical
curves suggest certain proportional re-
lationships between variables. Several
common general cases are illustrated
below.
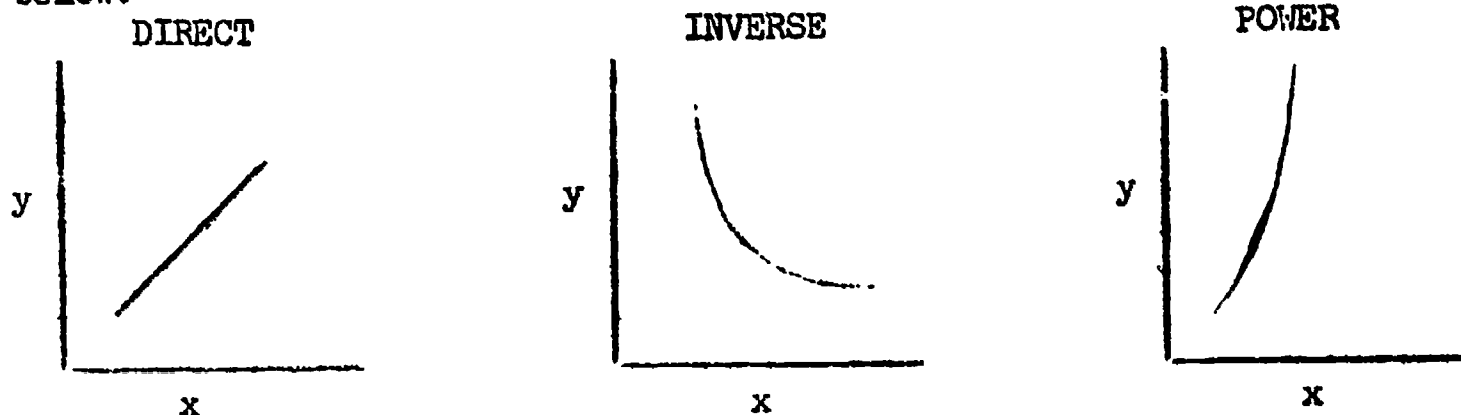
In the first graph y is directly pro-
portional to x. Symbolically this is
written $y \propto x$. The plot is a straight
line. (Therefore this sort of relation-
ship is often called a "linear" relation-
ship.)

The second graph illustrates a case
where y is in inverse proportion to some

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power of x. The case you are most famil-
iar with is $y \propto 1/x$, but proportionality
to $1/x^2$, $1/x^{\frac{1}{2}}$, etc. would yield curves of
much the same shape. Notice how the curve
bows in towards the origin and approaches
the axes at the extreme ends.

The third curve is obtained in the
case where y is <u>proportional to a power</u>
of x. This may be written $y \propto x^n$, where
the exponent n may have values such as
2, 3, etc., depending on the particular
situation.* The curve approaches the
origin for small values of x, but rises
very steeply for large values.

B.2 THE PROPORTIONALITY CONSTANT

You probably already know that two

---

*You may have realized that the linear
relationship illustrated in the first
graph is actually a special case of the
power relationship, where n = 1. How-
ever, this is such an important case,
and yields such a distinctive curve, that
it is listed as a case all by itself.

quantities which are proportional to one
another may be linked together to form
a correct equation by multiplying one
side of the proportion by a __constant of
proportionality__. If, for instance,
$y \sim 1/x^2$, the corresponding equation
could be written $y = k/x^2$. The graph
would have the same shape, but it would
pass through somewhat different points.
When evaluated and expressed in proper
units, it has special physical meaning
in many experiments. In the graphs you
plotted as an exercise above, the value
of k for the linear relationship was 2,
for the inverse relationship was 6, and
for the power relationship was 1. Try
plotting those equations with other
values of k, say 1, 3, and $\frac{1}{2}$ respective-
ly, in order to see what bearing changing
the constant has on the curves.

C. A SIMPLE LINEAR CASE: DENSITY OF METALS

The physical interpretation of a
constant of proportionality can be

easily illustrated in the case of a lab-
oratory exercise with which you are al-
ready quite familiar. Suppose you had a
number of pieces of metal, say of gold,
silver, and aluminum, of various shapes
and sizes. If you measured the volume
and mass of each piece and plotted these
quantities versus each other on graph
paper, you would probably obtain straight
lines similar to those shown below.

GOLD	SILVER	ALUMINUM
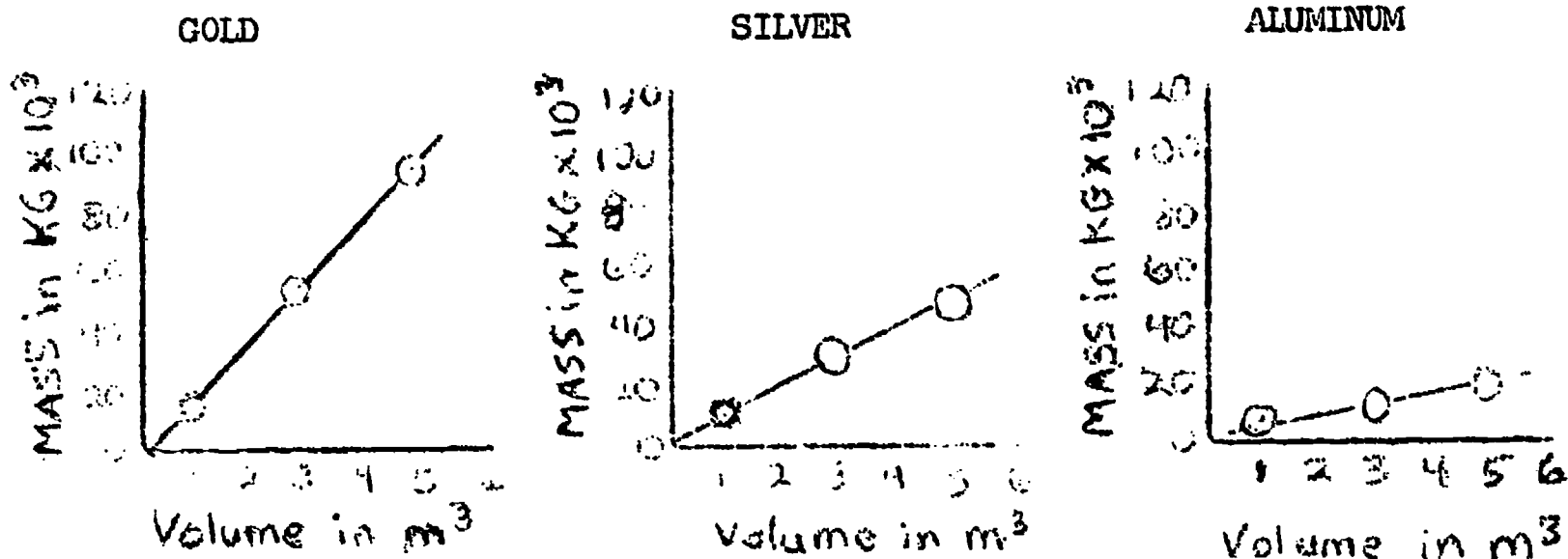
The basic meaning of the graphs is the
same in all three cases, viz., the bigger
the volume of material, the greater the
mass of material, in direct proportion.
This is of course exactly what we expect,
and it is just as likely for one sub-
stance as another. But, although the

graphs are all straight lines, they differ in one important respect. The lines slant at different angles to the axis. They have different slopes. The mass of gold increases at a faster rate when the volume of material is increased than does the mass of aluminum or silver. The constant of proportionality expresses this difference for us; its value is, in fact, the slope of the graph.

The slope tells us the ratio of the change in the variable plotted on the vertical axis and the corresponding change in the variable plotted along the horizontal axis. The slope is sometimes referred to as rise/run, or designated as $\frac{\Delta y}{\Delta x}$ (where y refers to the vertical axis and x to the horizontal). Note the calculation for gold.

$$k(\text{gold}) = \text{Slope} = \frac{\Delta y}{\Delta x} = \frac{20 \times 10^3 \text{kg}}{1\text{m}^3}$$

$$= \frac{20 \times 10^3 \text{kg}}{\text{m}^3}$$

GOLD

Graph: MASS IN KG ×10³ (y-axis) vs Volume in m³ (x-axis). Slope indicators: $\Delta y = 20 \times 10^3 kg$, $\Delta x = 1 m^3$.
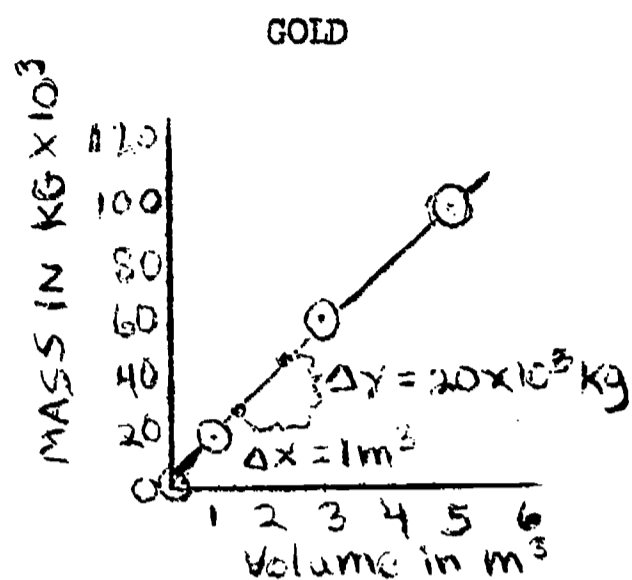
Check yourself in this skill by finding

the slope of the lines in the silver and

aluminum graphs and then comparing your

answers to those given below.

$$k(\text{silver}) \cong 10 \times 10^3 \frac{kg}{m^3}$$

$$k(\text{aluminum}) \cong 3 \times 10^3 \frac{kg}{m^3}$$

For the examples we have been

using in this discussion, the constant

of proportionality is given a special

name--density. As you probably remem-

ber from Year I, density is a character-

istic property of matter that tells us

how the mass of any particular sample of

matter is related to its volume.

In summary then, the step by step

analysis of our graphical data in these

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examples leads us from a statement of pro-
portionality to a mathematical equation,
and thence to an evaluation of density.
For gold:

$m \propto v$

$m = kv$

where k, the density =

$$20 \times 10^3 \frac{kg}{m^3}$$

D. ANOTHER EXAMPLE: ELECTRICAL HEATING

D.1 GETTING THE DATA

In Year I you performed experiments
in which the temperature of water was
raised using electrical immersion heaters.
Electrical energy was put into a water
bath and the consequent rise in temper-
ature was observed. You might recall
that some of the ideas you were concerned
with were conservation of heat, meaning
of a calorie, and conversion of energy.
We will now go back and look at some data
from these experiments and further ana-
lyze them in terms of graphs and equa-
tions.

In the Year I experiments styrofoam
tanks were used to hold the water and the
immersion heaters were operated from a
D.C. power supply with provisions for
varying the voltage and current. We
varied the mass of the water, the length
of the heating time, and the current
flowing through the immersion heater,
while measuring the temperature change in
the water. Time was measured with an
accuracy of $\pm$ 5 sec. using the sweep
second hand on a wrist watch. The mass
of the water was found indirectly from
a volume measurement. The temperature
change was measured with a Centigrade
thermometer, the accuracy being $\pm$ 0.2°C.
The current flowing through the heating
element of the immersion heater was
measured with an ammeter to within 0.01
ampere. Perhaps you would like to re-
peat this same experiment with equipment
readily available in your laboratory.
Some typical results are given below
in tabular form.

Temperature Change (°C) upon Heating
Using 0.40 Amp Current

Mass(kg) Time(sec)	0.300	0.250	0.200	0.150
60	0.4°	0.6°	0.7°	1.1°
120	1.1°	1.3°	1.7°	2.2°
240	2.1°	2.7°	3.3°	4.4°
360	3.1°	4.0°	5.0°	6.5°
480	4.2°	5.2°	6.5°	8.4°
600	5.3°	6.4°	8.0°	10.4°

Table D.1

Temperature Change (°C) of 0.200kg of
Water Heated for 300 Seconds

Current	Temperature Change
0.1 amps	0.3°C
0.2 amps	1.0°C
0.3 amps	2.5°C
0.4 amps	4.0°C

Table D.2

Notice that in the course of our experiments we were concerned with four variables (change in temperature, mass of water, heating time, and heater current),

but for the present time we will examine
them only two at a time. Finally we will
use all of our information to obtain a
combined result.

D.2 TEMPERATURE VS. TIME

One relationship that we are inter-
ested in is that which may exist between
the temperature change in the water and
the length of heating time. However,
we must be sure to keep the other two
variables constant during this search
so that our two-variable analysis will
be valid. We have already learned that
the first step in graphical analysis is
the construction of a properly labeled
graph. Plot the temperature change
($\triangle T$) against the length of heating time
(t) for a mass of 0.200 kilograms of
water. Remember to plot the dependent
variable (temperature change in this
case) on the vertical axis and the inde-
pendent variable (heating time) on the
horizontal axis. After plotting your

points, construct a smooth curve which

fits the points as best you can estimate.

Does your graph appear indicative of any

of the proportional relationships that

we examined earlier? If so, you may be

able to state the proportional relation-

ship that exists between the variables.

As we stated earlier, to proceed from a

statement of proportionality to an equa-

tion it is necessary to determine a con-

stant of proportionality. If you have

obtained a straight line, find the con-

stant of proportionality from your graph

by measuring the slope of the graph. Be

sure to indicate units in your result.

Now write the equation relating the two

variables you have been examining

You may be curious as to whether or

not this same relation applies to heat-

ing other masses of water. On the same

sheet of graph paper, plot the data for

each of the other masses of water. Do

all the masses show the same proportion-

al relationship? Do all of the lines
have the same equation? What can you
conclude?

D.3 TEMPERATURE AND MASS

Now let's take a look at the rela-
tionship that may exist between the
change in temperature of the water and
the mass of water being heated. The time
of heating and current through the heat-
ing element is to be maintained at the
same value for all the measurements.
Plot a graph of temperature change ($\triangle T$)
versus mass of water (m) for a heating
time of 120 seconds. Does your graph
appear to have the same general shape as
one of those already shown? Do we have
a technique for finding its slope? If
not, we need to try something different
before we can proceed with the task of
writing an equation.

Your graph probably has a shape re-
sembling the inverse proportion curve
discussed earlier. If, in fact, the re-

lationship is a simple inverse propor-
tion $\Delta T \propto 1/m$, it is relatively easy
to show this. Simply plot ΔT versus $1/m$
and a straight line should be obtained,
from which a slope can be found. Try it.
You should now be able to write the equa-
tion describing the relationship that
exists between temperature change of the
water and mass of water for a heating
time of 120 seconds. You may have a sus-
picion that the same proportional rela-
tionship holds true for other masses of
water. If you wish, go ahead and graph
data for the other masses of water to
find the answer for yourself.

D.4 TEMPERATURE AND CURRENT

Now that we have been reinforced by
our success with the first two relation-
ships, let's proceed on to the third.
Here we are interested in the relationship
that may exist between the temperature
change in the water being heated (ΔT)
and the quantity of current flowing

through the heating element (I). Construct your graph, plotting ΔT against I. Something new should be discovered with this graph, since it seems indicative of a power curve.

Before we can even write the correct proportional relationship between the variables, however, we need to know the value of the unknown exponent. There was actually a similar problem in the case of the inverse relation discussed in Section D.3. There we assumed the simplest possibility, viz., a simple inverse of the form $y \propto 1/x$, and came out all right. In the general case, however, there are several ways to proceed. We could look to the theory of the experiment and make an <u>educated guess</u> as to the power relation. Before the course is over, we will know enough about electricity to be able to do this, but for the present we will not go about it this way. Another approach uses slightly sophisticated mathematical techniques involving log-

arithms, but this will not be necessary
here. The approach we will take, an
approach which is used surprisingly often
in scientific research, is that of trial
and error. This simply means that we
try plotting ΔT against different powers
of I until hopefully we find a power
that will give us a straight line. Plot-
ting ΔT against I^2 is probably a good
way to start. Try it. If you obtained
a straight line you can now write the
proportional relationship and equation.
If you didn't obtain a straight line,
try plotting ΔT versus I^3 or $I^{\frac{1}{2}}$, and
see if that does the job. Once you have
obtained a straight line, you can pro-
ceed.

D.5 A GRAND FINALE

By now you should have found the re-
lationship between temperature change and
time of heating, temperature change and
mass of water, and temperature change and
current in the heating element. It is

perfectly possible that in a real situation all three of the independent variables could be varying during the experiment. How then can we determine the relationship between temperature change and each of them at the same time? It's really quite easy. If the temperature change is directly proportional to the time of heating (t), the reciprocal of the mass $\left(\frac{1}{m}\right)$, and the square of the current (I^2), then it is directly proportional to the product of the three.

$$T \varpropto t\ I^2\left(\frac{1}{m}\right) \quad \text{or} \quad T \varpropto \frac{t\ I^2}{m}$$

Writing an equation from this statement of proportionality involves nothing beyond the techniques we have already utilized.

Choose any one of your three completed graphs ($\triangle T$ versus t, $\triangle T$ versus $\left(\frac{1}{m}\right)$, or $\triangle T$ versus I^2). In each case, note that two of the independent variables were held constant and one varied. Therefore you can read a temperature

change corresponding to three known values
of the independent variables at any point
on the graph. Using this technique, plot
a graph of $\triangle T$ versus $\dfrac{t\ I^2}{m}$ and write the
equation for the line.

E. CONCLUSION

You have now accomplished an impor-
tant task. The ability to express in the
form of an equation the results of lab-
oratory research is the essence of gen-
eralization in science. This type of
process has led us to many of the best
known theories and principles in science,
including many of those you will be
studying in your first work in physics
--concepts of motion.

MOTION AND ENERGY

II. NEWTON EXPLAINS MOTION

A. MASS AND FORCE
     Read section 3.1 HP

B. VECTORS
     Read section 3.2 HP

C. NEWTON'S LAW
     Read sections 3.3 - 3.7 HP

C.1 Experiment: NEWTON'S SECOND LAW
      (HP Expt.)

D. NEWTON's LAWS
     Read sections 3.8 - 3.11 HP

III. MULTI-DIMENSIONAL MOTION

    A. CURVED PATHS
       Read sections 4.1 - 4.4 HP

    A.1 Experiment: TRAJECTORIES - I

    B. CIRCULAR MOTION
       Read sections 4.5 - 4.7 HP

Additional Problems - Chapter 4,
        Sec. 4.2 & 4.3

1. A boy standing on a bridge above
the water throws a stone horizontally
outward. He times the stone as taking
6 seconds to strike the water.

    a) How high is the bridge?

    b) How far did the stone travel
        horizontally before striking the
        water if thrown with a speed of
        50 m/sec.?

2. a) What was the magnitude of dis-
placement of the stone in the above
problem after only 2 seconds?

    b) What is the direction of the
displacement? (Express as a ratio of
vertical/horizontal distances).

3. a) What is the speed of the stone
in problem 1 along the x & y axis
after 2 seconds?

    b) What is the magnitude of the
velocity of the stone after 2 seconds?

    c) What is the direction of the
velocity of the stone after 2 seconds?

IV. CONSERVATION

    A.   CONSERVATION OF MASS
         Read section 9.1 HP

    B.   COLLISIONS
         Read section 9.2 HP

    B.1   Experiment: COLLISIONS IN ONE
                    DIMENSION
         (HPExpt. 22)

    C.   CONSERVATION OF MOMENTUM
         Read sections 9.3 - 9.7 HP

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V. ENERGY - WORK

    A. WORK AND KINETIC ENERGY
        Read section 10.1 HP

    B. POTENTIAL ENERGY
        Read section 10.2 HP

    C. CONSERVATION OF MECHANICAL
        ENERGY
        Read section 10.3 HP

    D. ENERGY IN BIOLOGICAL SYSTEMS
        Read section 10.8 HP

    E. CONSERVATION OF ENERGY
        Read section 10.9 HP

VI.   KINETIC THEORY OF GASES

    A.   KINETIC THEORY
        Read sections 11.1 - 11.5 HP

    A.1 Experiment: BEHAVIOR OF GASES
        (HPExpt. 29)

    B   SECOND IAW OF THERMODYNAMICS
        Read sections 11.6 - 11.8

CHEMICAL REACTIONS

I. THE MOLE AS A COUNTING UNIT

A. COUNTING THINGS

   1. Exercise: COUNTING "BB'S"

B. COUNTING UNITS

C. THE MOLE

D. ATOMIC WEIGHT

   1. Experiment: COPPER WIRE, SILVER NITRATE AND MOLES

   2. Experiment: FROM MOLES OF SILVER TO MOLES OF SILVER CHLORIDE

E. FORMULAS FOR MOLECULES

F. MOLECULAR WEIGHTS

G. RELATIVE WEIGHTS OF MOLECULES

Chapter I: The Mole as a Counting Unit

A. COUNTING THINGS

In previous chapters a model for the
gaseous state has been examined. Speeds
and sizes of molecules were discussed
and the ideal gas law was derived by
considering topics such as pressure, mo-
mentum and temperature of gases. (P=kDT)

If man wants to communicate about
particles or things, he must be able to
describe them. What is the shape of a
gas particle? Is it spherical? How
many are there in a given volume? How
does one count these extremely small
particles that go dashing randomly about
in a container? Let's first think about
counting these gas particles. How does
man usually count things and what are
some alternative ways of counting things?

A.1 - Exercise: COUNTING "BB'S"

Although "BB's" are enormously
large when compared to gas particles,
the practice in devising an alternative

method of counting these fairly small

particles will be helpful.

Count out ten BB's; don't laugh--

do it. How did you count this small

number of particles? Now count out one

thousand BB's. Think about alternative

ways of counting. What is the easiest

way to count one thousand BB's? Do it.

What would be the easiest way of count-

ing all the BB's in the classroom?

Then counting particles is an easy

task. All that is needed is the mass

of a number of particles. But, you say,

"ten gas particles mass so little

(about 10^{-21} grams) that our laboratory

balances cannot mass a unit of ten."

What we really need is an appropriate

counting unit for gas particles.

B. - COUNTING UNITS

One of the most used counting units

is the dozen. We can think of the mass

of twelve cows or people, but the mass

of twelve atoms is very little differ-

ent from the mass of ten. Obviously

the dozen is not a good counting unit
to use when working with gas particles.
What about the gross (144 things)?
No, since a gross of gas particles masses
about 10^{-20} grams. We need a "really
big" counting unit.

C. - THE MOLE

Such a unit already exists. It is
called the mole. It has nothing to do
with little animals that live in the
ground, unless you wish to count them.
The mole is 6.02×10^{23} things, just
as the dozen is 12 things. One can talk
about a mole of cows or mosquitoes;
however the use of such a large count-
ing unit for this purpose is ridicu-
lous. One mole of cows would mass
about 10^{26} Kg! This **is approximately**
sixteen times greater than the mass of
the earth!! Gas particles. however,
have such small masses that a mole of
gas can be massed on a typical labora-
tory balance.

D. - ATOMIC WEIGHT

If our gas particle happens to be

mono-atomic (composed of just one atom),

then the weight in grams of one mole

of the gas is its atomic weight. The

weight of one mole of helium gas, for

example, is four grams.

Find a table of atomic weights and

determine the weight of one mole of

neon, krypton, and zenon. The idea of

atomic weight does not stop with just

gases. The weight in grams of one mole

of any mono-atomic element is the atomic

weight of that element. Notice the term

"weight" is used in the definition of

ATOMIC WEIGHT. We know the definition

should read, "the mass in grams of one

mole of atoms is the atomic ~~weight~~ mass".

Traditionally chemists have used the

term weight when they mean mass. Al-

though tradition can be confusing,

most areas of study are bound by it; and

we must simply remember the chemist us-

ually means mass when he says "weight".

D.1 - Experiment: COPPER WIRE, SILVER
                   NITRATE AND MOLES

In this experiment you will pre-
pare a solution of silver nitrate.
You will also weigh a length of copper
wire, and after it has remained in the
silver nitrate solution over night, you
will weigh it again. In this and the
following experiment you will count, by
weighing, the number of moles of part-
icles involved.

Copper wire is composed of copper
particles. It is an element; it con-
tains only one kind of atom. The sym-
bol for copper is Cu. Silver nitrate
is a compound which has the formula
$AgNO_3$. You will find your laboratory
work even more exciting if you prepare
a data table before you come to lab.

Your data should include the following:

Weight of Cu wire before immersion

Weight of Cu wire after immersion

Weight change of Cu wire

Weight of vial plus $AgNO_3$

Weight of vial

Weight of AgNO$_3$

Weight of beaker plus residue

Weight of beaker

Weight of residue

Obtain a 30 cm. piece of copper wire. Clean it with steel wool. Then form a coil by wrapping the wire around a small test tube. Leave about 5 cm. for a handle.

Carefully weigh the copper wire. Also weigh a clean dry 250 ml beaker. Weigh the vial of silver nitrate provided by the instructor. Do not use tap water because of impurities. Fill the beaker about two-fifths full with de-ionized water. Add the solid nitrate (AgNO$_3$) to the water. Stir gently with a glass rod until all the silver nitrate (AgNO$_3$) crystals have dissolved. Weigh the empty vial. Silver nitrate (AgNO$_3$) solution or solid will stain your skin and clothing.

Cu wire
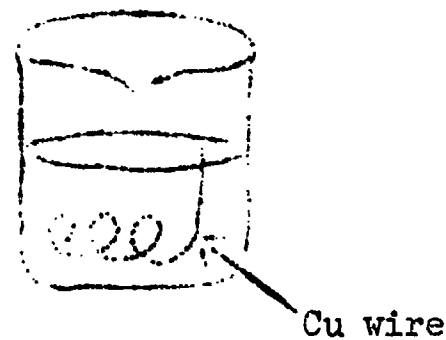
watch glass to te placed on top of beaker to keep out dust

Fig. D.1

It will be easy for your instructor
to determine who is the sloppy exper-
imenter.

Now place the copper wire in the
beaker with the coil immersed in the
$AgNO_3$ solution. Observe and record any
changes that take place for at least
five minutes. Then place the beaker,
covered with a watch glass, in the de-
signated place until the next laboratory
period.

When you return to the laboratory,
be very careful not to disrupt the
contents of the beaker until you have
had a chance to observe the contents.
Record your observations.

Shake the crystals off the coil
and lift the coil out of the solution.

Use a wash bottle to rinse into the
beaker any crystals which tend to ad-
here to the coil. Now set the coil
aside to dry after you have rinsed it in
acetone. Weigh the coil when it
is dry.

Let the crystals settle in the
beaker and carefully decant the solu-
tion. Decant means to pour off the
liquid and leave the solid behind.

See Figure D.2.A. and B.

Next add 5 ml. of dilute $AgNO_3$
solution and stir gently until any Cu
flecks left in the residue in the ori-
ginal beaker disappear. Carefully de-
cant again. Wash the residue with 10 ml.
of water and carefully decant again.
You may neglect the few particles which
may float over with the wash water be-
cause their weight is usually insignif-
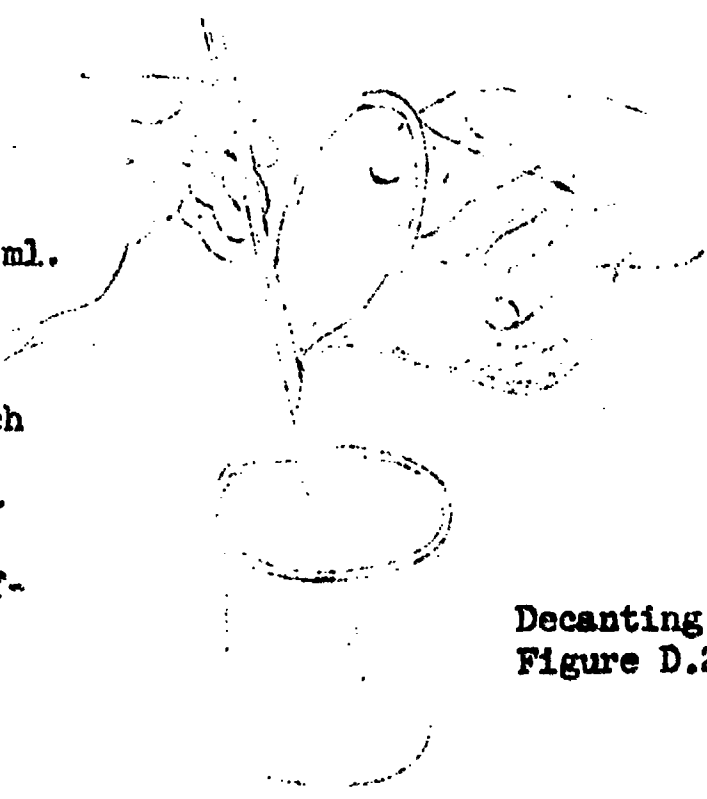icant. (Uncertainty in measurement?)

Now the residue must be dried.
Your instructor will suggest a suitable
method. This usually takes less than

Rinsing with a Wash Bottle

Figure D.2.A.

Decanting
Figure D.2.B

Drying of the sample may
be carried out with a heat
lamp or an oven.

73

24 hours.

When you return to the laboratory allow the beaker and its contents to cool before weighing. Should you use the same balance you used previously? Why? Record the weight. How can you be certain the contents are dry?

COVER THE BEAKER AND CONTENTS WITH A WATCHGLASS AND SAVE IT FOR THE NEXT EXPERIMENT.

Calculations:

1. Calculate the part of a mole of copper that went from the wire into the solution.

2. The residue is silver (Ag). Calculate the part of a mole of silver you obtained.

3. What is the ratio of moles of silver to moles of copper? Express your answer using scientific notation.

4. What you have observed can be partially described by the following: 1 mole of copper (solid) •

_____ mole(s) of sil-

ver nitrate (in water) ——→

____ mole(s) of silver (solid)

↓ ___ mole(s) of copper nitrate

(in water).

Using your results, write the

proper whole numbers in the

blanks.

5. How many atoms of Cu were lost

by the wire?

6. How many atoms of silver were

left as the residue?

7. What is the ratio of atoms of

silver to atoms of copper
(from your calculations in
#5 & #6)?

8. Your instructor or you may want

to do a frequency distribution

for the results of this exper-

iment. (How often did stu-

dents get the same answer for

the ratio of copper to silver;)

If you do the frequency distri-

bution, please consider how the

range of values obtained com-

pa.. s with the uncertainty you

consider justifiable from your

measurements.

What do you think causes the blue color

of the solution?

D.2 - Experiment: FROM MOLES OF SILVER
TO MOLES OF SILVER
CHLORIDE

In your next experiment the silver

produced in the previous experiment is

used to form a water solution of $AgNO_3$.

This is done by allowing the silver to

react with <u>nitric</u> <u>acid</u> (HNO_3). A water

solution of <u>sodium</u> <u>chloride</u> (NaCl) is

then added to the silver nitrate solution

and silver chloride is produced.

Many techniques are reviewed in

this experiment and a few new ones are

learned. Learn as much as you can of

these techniques--careful weighing,

decanting, washing, drying and filter-

ing--because they <u>will</u> be used often

in future experiments.

Remember to record your data care-
fully. Include the uncertainty and
units in each measurement. You should
have a data table before you come to
the laboratory.

Your data should include the following:

Weight of silver from previous
    experiment

Weight of silver nitrate used in
    previous experiment

Weight of beaker #1

Weight of beaker #1 and solid
    $AgNO_3$

Weight of beaker #1, filter paper,
    and solid AgCl

Weight of beaker #2

Weight of beaker #2 and solid
    NaCl

Weight of filter paper

Weight of beaker #2 and solid res-
    idue

Record the weight of the beaker
used in the last experiment. (The one
with the solid silver in it.) Label

this beaker #1. Also record the weight
of silver nitrate used and the weight
of silver produced.

Add 15 ml. of nitric acid (HNO_3)
labeled 6 M HNO_3, to the beaker con-
taining the silver. Don't inhale any of
the poisonous reddish-brown fumes of
nitrogen dioxide (NO_2) which form as
the silver dissolves. Put an identi-
fying mark on your beaker (#1) and leave
it in the fume hood or near an open
window to dry for about 24 hours.

When you return to the laboratory,
weigh the beaker which now contains
$AgNO_3$. (Does this material stain
clothing?)

Add 15 ml. of de-ionized water to
the $AgNO_3$ in beaker #1. Stir until no
more change takes place.

Pour out about $\frac{1}{2}$ teaspoonful of
sodium chloride (NaCl) as shown below:

Rotate the bottle to
obtain the approximate
amount of solid
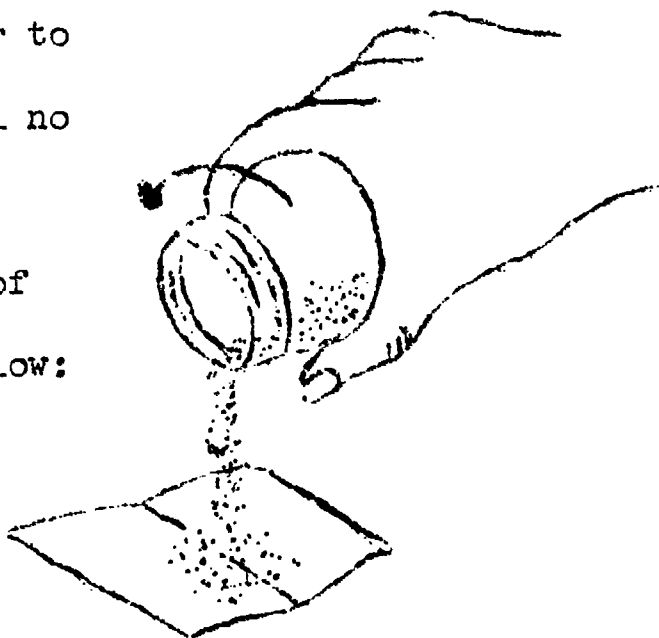
Now label a clean dry 100 ml. beaker as #2 and weigh it to the nearest 0.01 g.

Adjust your balance so it reads about 2.0 to 2.5 grams greater than beaker #2. Carefully add the NaCl from the paper to beaker #2 until the balance "trips." Then carefully weigh the beaker with the NaCl and record this weight on your data table.

Add about 15 ml. of de-ionized water to the NaCl in beaker #2. Stir until no more change is observed.

Add the NaCl solution slowly to the $AgNO_3$ solution while rapidly stirring the $AgNO_3$ solution. The white solid produced is the compound silver chloride, (AgCl). Rinse beaker #2 with about 5 ml. of de-ionized water.

Rinsing a beaker with a
wash bottle

By using two different size beakers there is less chance that the students will confuse the two solutions.
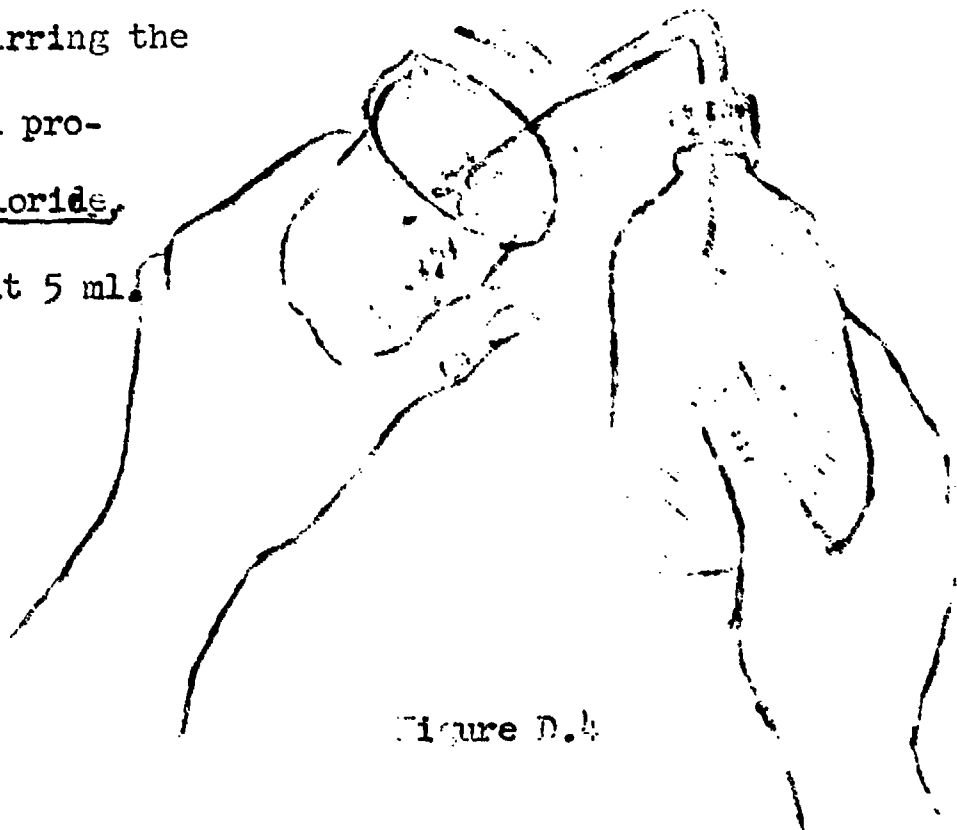
Figure D.4

Add the rinse water to the mixture in beaker #1. Rinse beaker #2 again with de-ionized water and discard the rinse water. Beaker #2 will be used later.

Now heat beaker #1 to boiling for about 2 minutes. The solution should become reasonably clear as the precipitate settles. If you place a stirring rod in the beaker, it will help prevent unsteady boiling (bumping).

Weigh a piece of filter paper. Then fold it as shown below.
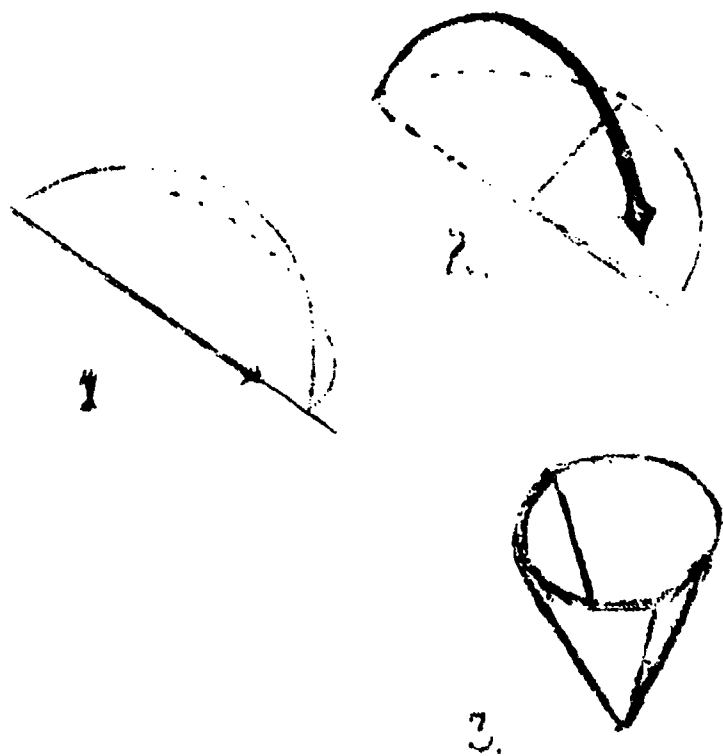
Place beaker #2 under the funnel to catch the filtrate (the liquid that comes through the filter paper). Decant the liquid from beaker #1 into the funnel. A small amount of precipitate may go into the filter paper, but try to keep most of it in beaker #1 for washing.

Wash the precipitate in beaker #1 with about 15 ml. of de-ionized water, stirring with a glass rod to aid the washing. Decant the wash water into the funnel. Repeat the washing procedure and again decant the wash water into the funnel.

After the filtration is complete place the filter paper, partially opened, in beaker #1.

Now place both beakers, #1 and #2, in a place designated by your teacher for drying. Be certain you can recognize your beakers.

When you return to the laboratory, weigh both dry samples and record the weights. (How can you be certain

the samples are dry?) Save the silver
chloride as directed by your teacher.

Calculations:

Fill in the following table:

	Wt.-gms.	Moles
Ag(previous exp.)		
$AgNO_3$(previous exp.)		
$AgNO_3$ produced this exp.		
NaCl added		
AgCl (in beaker #1)		

1. How does the number of particles
   of $AgNO_3$ produced in this exper-
   iment compare with the number of
   particles of $AgNO_3$ used in the
   previous experiment? How do
   you account for any similarity
   or difference?

2. Compare the sum of the weights
   of the $AgNO_3$ and NaCl used with
   the sum of the weights of the
   AgCl and the residue in beaker
   #2. What is the significance

<u>of these results</u>? Perhaps

some class data would be help-

ful.

3. Check on the moles of silver

atoms involved in both exper-

iments. Considering the un-

certainty associated with your

measurements, <u>what can you con-</u>

<u>clude about the number of moles</u>

<u>of silver atoms involved in this</u>

<u>series of chemical changes</u>?

4. Pure $AgNO_3$ is a white solid.

How do you account for any color

which may be present in the pre-

pared sample?

If you have time, you may wish to

devise an experiment to determine the

composition of the residue in beaker #2.

Check with your teacher.

E. FORMULAS FOR MOLECULES

Molecules are made of atoms in def-

inite arrangements and definite numbers.

Symbolism aids us in demonstrating the

number of kinds of atoms in a molecule.
For example, H_2 means two atoms of hydrogen make up a molecule of hydrogen gas. A molecule of water is composed of two atoms of hydrogen and one atom of oxygen. The formula for water, H_2O, conveys all this information. The molecular formula for ammonia is NH_3. What does the formula represent? It is possible to find tables of the symbols commonly used for elements. The Handbook of Chemistry and Physics is a good place to look for molecular formules.

It is often desirable to translate molecular formulas from digits to words. This is done by using the appropriate Greek prefix from the list below.

mono . . . . one	hexa . . . . six
di . . . . . two	hepta. . . . seven
tri. . . . . three	octa . . . . eight
tetra. . . . four	nona . . . . nine
penta. . . . five	deca . . . . ten

Often no prefix is used and in these cases "mono" is understood.

Examples:

molecular formula	translation in words
CO_2	carbon dioxide
PBr_3	phosphorus tri-bromide
CCl_4	carbon tetra-chloride
PCl_5	phosphorus penta-chloride
N_2O_5	dinitrogen pent-oxide
NaCl	sodium chloride

F. MOLECULAR WEIGHTS

If a mole of atoms has a definite weight, then a mole of molecules which is composed of atoms in definite numbers must also have a definite weight. One mole of H_2O must contain two moles of hydrogen atoms and one mole of oxygen atoms. Therefore the molecular weight of H_2O is $1 + 1 + 16 = 18$. The gram molecular weight is the weight in grams of a mole of molecules. What is the

gram molecular weight of H_2? of NH_3?

G. RELATIVE WEIGHTS OF MOLECULES

Imagine that in your left hand you
have a dozen identical red marbles that
have a total weight of thirty-two grams
and in your right hand a dozen iden-
tical white marbles that have a total
weight of sixteen grams. The ratio of
the total weight in your left hand to
the total weight in your right is thirty-
two to sixteen or two to one. But--can
we determine the ratio of weight of one
red to one white marble from this same
information? Of course; the ratio of
total weights is the same as the ratio
of individual weights. Therefore, if
the ratio of molecular weights is known,
the ratio of weights of individual mol-
ecules is also known, if we know there
is the same number of each. For example,
one mole of oxygen weighs 32 grams and
one mole of hydrogen weighs two grams.
The ratio of their molecular weights

is 32 to 2 or 16 to 1. The ratio by
weight of an individual oxygen molecule
(O_2) to an individual hydrogen molecule
(H_2) must also be 16 to 1. This com-
parison of the weight of a mole of one
substance to the weight of a mole of a
second substance yields a meaningful
ratio because a mole of any substance
contains the same number of molecules
(6.02×10^{23}) as a mole of any other
substance.

Exercises for Home, Desk and Lab (HDL)

1. A stone about the size of a
    grapefruit weighs roughly a
    kilogram. How many moles of
    such stones would be needed to
    account for the entire mass of
    the earth, which is about 6 x
    10^{24} kg?

2. If we had one mole of dollars
    to divide among all the people
    in the world, how many dollars
    would each of the three billion
    inhabitants receive?

3. The gram atomic weight of gold
    is 197. Determine the weight
    in grams of one gold atom.

4. The most delicate balance can
    detect a change of about 10^{-8}
    grams. How many atoms of silver
    would be in a sample of that
    weight? The gram atomic weight
    of silver is 108.

5. How many moles of atoms are in
    (a) 9.0 grams of aluminum
    (b) 0.83 grams of iron?

6. What do the following symbols represent? K, Ca, Co, Pb, Sc.

7. Write formulas for:

   (a) silicon dioxide

   (b) sulfur dichloride

   (c) nitrogen trifluoride

   (d) aluminum tribromide

   (e) dinitrogen diiodide

8. (a) What does the molecular formula CCl_4 mean?

   (b) What information is added by the following structural formula? Guess!

   $$Cl - \overset{\displaystyle Cl}{\underset{\displaystyle Cl}{\overset{|}{\underset{|}{C}}}} - Cl$$
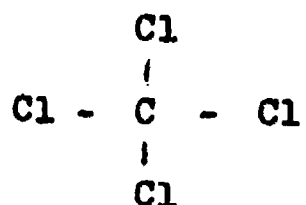
9. How many moles of oxygen atoms are in one mole of perchloric acid molecules ($HClO_4$)?

10. Consider the following data:

Element	Atomic Weight
X	12.0
Y	19.0

X and Y combine to form a new substance Z. If one mole of

X combines with four moles of
Y atoms to give one mole of Z,
(a) what is the formula of Z
in terms of X and Y? (b) what
is the weight of one mole of Z?

11. A chemist weighs out 10.0
grams of carbon dioxide,
10.0 grams of water and 10.0
grams of magnesium. How many
moles of each substance does
he have?

12. Write the molecular formulas
for the following compounds and
give the weight of one mole of
each: carbon monoxide, sili-
con dioxide, phosphorus tri-
bromide, titanium tetrachloride,
and uranium hexafluoride.

13. Calculate the molecular weight
of each of the following:

(a) SiF_4

(b) HF

(c) F_2

(d) Xe

(e) NO_2

14. The ratio of the weight of a liter of chlorine gas (Cl_2) to the weight of a liter of oxygen gas (O_2) is 2.22 (both measured at room temperature and pressure). Calculate the molecular weight of chlorine.

15. Gaseous uranium hexafluoride is important in the preparation of uranium as a source of atomic energy. A flask is filled with gaseous uranium hexafluoride. The uranium hexafluoride, UF_6, is found to weigh 7.04 grams. The same flask is filled with oxygen and, under similar conditions, the oxygen weighs 0.64 grams. What is the ratio of the weight of one uranium hexafluoride molecule to the weight of one oxygen molecule?

16. Determine the weight ratio of an oxygen atom to a carbon atom.

17. Oxygen (O_2) is found to weigh
8 times as much as an unknown
gas. What is the molecular
weight of the unknown gas:

II. COMBINATIONS OF GASES

A. AVOGADRO'S HYPOTHESIS

A.1 Experiment: THE WEIGHTS CF
                 EQUAL VOLUMES OF
                 GASES AT THE SAME
                 TEMPERATURE AND
                 PRESSURE

A.2 THE EFFECT OF BUOYANCY ON
     WEIGHING

B.   THE VOLUME OF A MOLE OF GAS

C.   S. T. P.

D.   TOTAL PRESSURE AND PARTIAL
     PRESSURE

D.1 Experiment: A QUANTITATIVE IN-
                  VESTIGATION OF THE
                  REACTION OF MAGNE-
                  SIUM WITH HYDRO-
                  CHLORIC ACID

Amadeo Avogadro, an Italian scientist who lived during the 18th and 19th century, noticed that gases combine in small whole number ratios. For example:

<u>Two</u> liters of the gas nitric oxide	combines with	<u>one</u> liter of the gas oxygen	to form	two liters of the gas nitrogen dioxide
<u>One</u> liter of the gas oxygen	combines with	<u>two</u> liters of the gas hydrogen	to form	<u>two</u> liters of the gas water

These combinations in small whole numbers occurred only if the volumes were measured at the same temperature and pressure. What do these observations have to do with particles and counting them?

A. <u>Avogadro's Hypothesis</u>

<u>Equal volumes of gases, measured at the same temperature and pressure, contain equal numbers of molecules.</u>

This hypothesis, first made in 1811, is still accepted today. That equal volumes contain equal numbers of particles at similar conditions allows the direct determination of relative weights of gases.

A.1 - Experiment: THE WEIGHTS OF EQUAL
VOLUMES OF GASES AT
THE SAME TEMPERATURE
AND PRESSURE

In this experiment you will weigh
equal volumes of carbon dioxide and oxy-
gen and perhaps another gas. As you
know, the volume of a gas varies with
temperature and pressure. Precautions
must be taken to keep these variables
the same as the gases are weighed. Pre-
cautions must also be taken in weigh-
ing, because the weight of the gas sam-
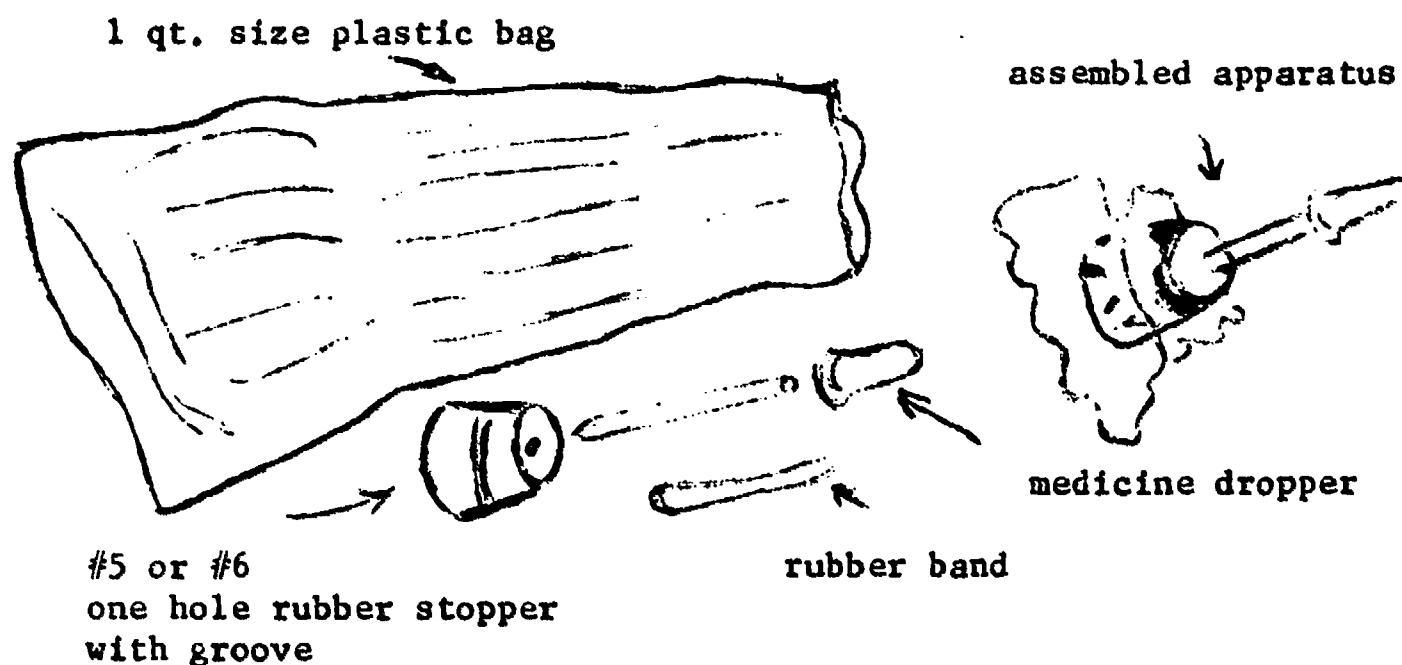ples will be at most a few grams.
With such small weights, the uncertain-
ty of measurement could destroy the sig-
nificance of your data. A few greasy
finger prints or a moisture droplet
are enough to destroy your results. In
other words "cleanliness is next to
good results." ·

Before coming to the laboratory
read the experiment (pages 42 - 50)
and organize a data table.

Part I. <u>Weighing carbon dioxide gas</u> $(CO_2)_{(g)}$

Obtain the apparatus pictured below
from your teacher and put it together
as pictured. Figure A. 1.

Good quality bags are required.
"Glad Bags" work.

1 qt. size plastic bag

assembled apparatus

medicine dropper

#5 or #6
one hole rubber stopper
with groove

rubber band

Be certain the large end of the
stopper is inside the bag. If you fold
the bag around the stopper in small pleats,
there will be less chance of leakage. Be
careful when you force the medicine drop-
per into the stopper. Experience shows
it is less painful if there is toweling
between glass and your hand if the glass
breaks. Ask your teacher for some help
if you do not already know the proper

Save the stopper assembly
for future years.

technique for handling glass tubing.

Press out any air in the bag by smoothing it flat. Then replace the rubber cap on the medicine dropper and weigh the complete apparatus. Don't forget to record the uncertainty.

Remove the rubber cap and connect the apparatus to the $CO_{2(g)}$ source supplied by your teacher. Allow the bag to be fully inflated. Hold the apparatus by the stopper and disconnect the rubber tubing from the medicine dropper. Don't squeeze the bag; but let the excess gas escape so that the bag will be at room pressure. Now replace the rubber cap while still holding the bag.

Weigh the bag apparatus with the CO_2 gas and record, the uncertainty in the measurement. You may wish to repeat the weighing of the $CO_{2(g)}$ in order to check your work.

Part II. Weighing oxygen gas $((O_2)_g)$

Empty the bag and stopper apparatus. Fill it with $O_{2(g)}$ using the source supplied by your teacher. Remember to allow the excess $O_{2(g)}$ out of the bag. Weigh the bag apparatus with the $O_{2(g)}$ and record with the uncertainty. Remember-- the idea is to have equal volumes of gases at the same <u>temperature</u> <u>and</u> <u>pressure</u>. Save the apparatus (bag filled with oxygen) for the determination of volume that follows.

If another gas is available and you have time, you may want to obtain the weight of the third bag of gas.

Part III. Volume determination

Measure the volume of oxygen in the bag by the method shown in the picture on the following page. (Figure A.2.)
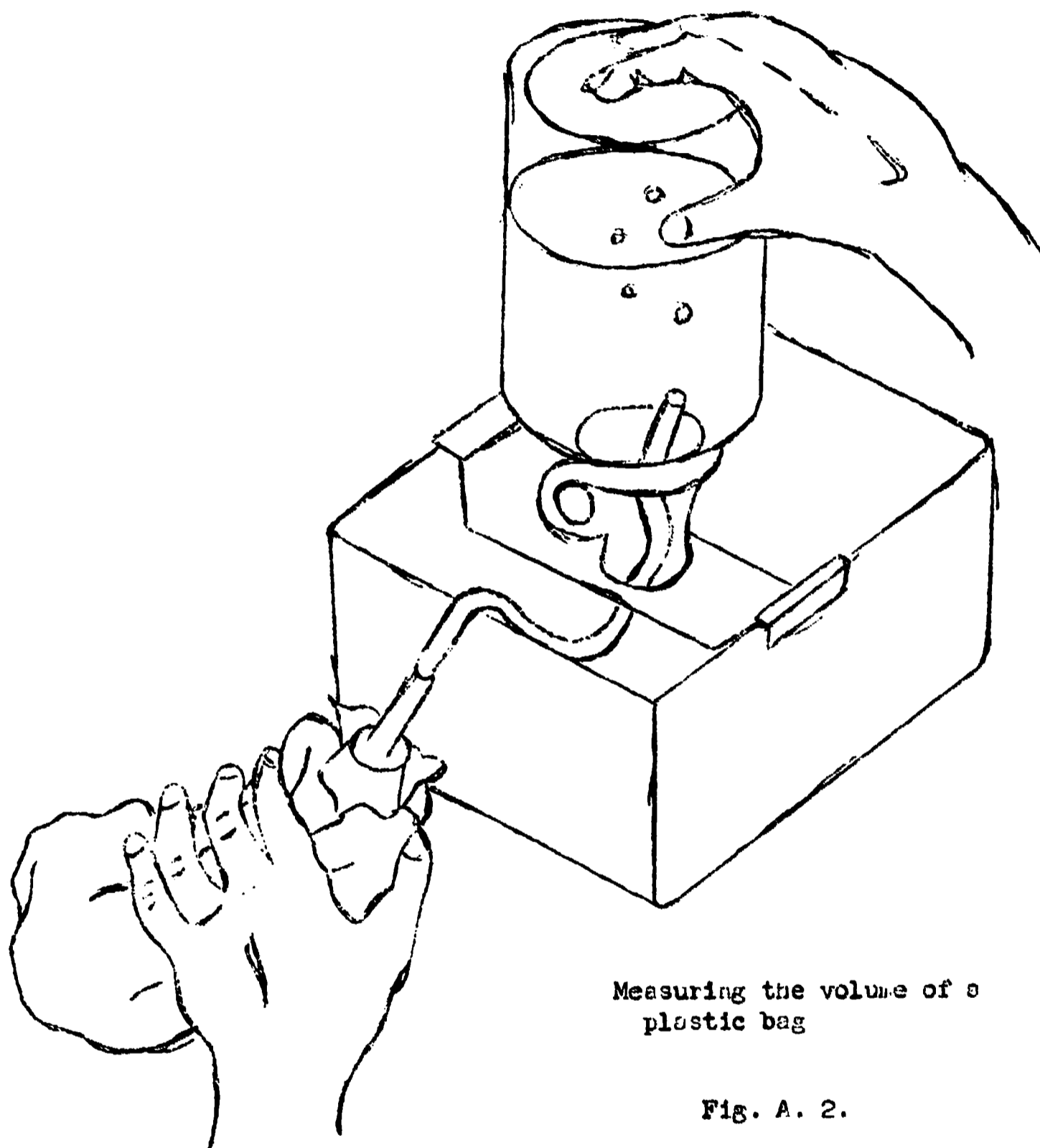
Measuring the volume of a
plastic bag

Fig. A. 2.

Completely fill a large bottle or jug
(about 2 liters) with tap water. Place
a glass square over the mouth of the jug
and invert it over the trough of water.
Remove the glass slide under water.

Shove one end of a suitable piece of
rubber tubing into the jug. Push it
up within one centimeter of the bottom
of the jug.

Now remove the cap from the med-
icine dropper and attach it to the
other end of the rubber tubing.
Gently press on the bag so the gas will
displace the water in the bottle. Fin-
ally smooth out the bag to remove all
the gas.

Quickly pinch the tubing so that
the back pressure will not force oxygen
(or water if the end of the tube is
below the water level) back into the
bag. Pull the tube out of the jug and
set the jug upright on the table.

Measure the volume of water needed
to refill the bottle. Record the vol-

ume of the bag. What is the uncertain-
ty of the measurement? Why is the vol-
ume needed to refill the bottle the vol-
ume of the bag full of gas? Record room
temperature and pressure. If you haven't
read a barometer before, ask your teach-
er or someone who has to show you
how.

A.2 - THE EFFECT OF BUOYANCY ON WEIGHINGS

Because we live in a sea of air, we
usually do not consider its effects on
measurements. When any object is weighed
in air, it is buoyed up by a force equal
to the weight of the volume of air it
displaces. When you step on your bath-
room scales, you read a weight that is
less than your true weight. Your vol-
ume is about 70 liters. At normal con-
ditions of temperature and pressure
seventy liters of air is about 84 grams
or three ounces.

(70 liters)(1.2 g/liter) = 84 grams.
Unless you have an unusual bathroom scale,
the uncertainty of the instrument is
much greater than three ounces. We

usually neglect the buoyancy effect
because it is relatively small.

Suppose that instead of the air a-
round us we were immersed in water
and we had the proper scuba equipment.
As you probably know, objects are easier
to lift when they are submerged in water.
The water buoys up a submerged object
with a force equal to the weight of the
volume of water it displaces. For example,
an object with a volume of one liter
(1000 ml.) when submerged in water dis-
places one liter of water. Therefore,
it is buoyed up with a force of 1000 g.,
the weight of a liter of water. The
same object in air is buoyed up with
a force of about 1.2 grams   (the weight
of the liter of air it displaces).

In this experiment the apparent
weight of the gas is not the true weight
of the gas. We must add the weight of
the volume of air displaced to the
apparent weight of the gas in order to
find the true weight of the gas.

The weight of the volume of air
displaced can be found by using the

following table.

WEIGHT OF A LITER OF AIR IN GRAMS PER LITER, ± 0.01 g, AT VARIOUS
TEMPERATURES AND PRESSURES.

Pressure (mm Hg)	Temperature			
	15°C	20°C	25°C	30°C
700	1.13	1.11	1.09	1.07
710	1.14	1.12	1.10	1.09
720	1.16	1.14	1.12	1.10
730	1.18	1.16	1.14	1.12
740	1.19	1.17	1.15	1.13
750	1.21	1.19	1.17	1.15
760	1.23	1.21	1.19	1.16
770	1.24	1.22	1.20	1.18

Table A.1.

Calculations

1. What is the apparent weight
   of oxygen that was in the bag?

2. What is the weight of air dis-
   placed by the bag?

3. What is the true weight of
   the oxygen that was in the bag?

4. What is the true weight of the
   carbon dioxide that was in the
   bag?

5. What is the ratio by weight of
   the volume of $CO_2(g)$ to the
   volume of $O_2(g)$.

6. What is the ratio by weight of
   an individual $CO_2(g)$ particle to
   an individual $O_2(g)$ particle?

7. Use a table of atomic weights
   to calculate the ratio of a
   mole of $CO_2(g)$ to a mole of
   $O_2(g)$.

8. Keeping uncertainty in mind, how
   do your answers to 5 and 6 com-
   pare with the answer to # 7?

B. THE VOLUME OF A MOLE OF GAS

One mole of most gases will occupy
22.4 liters if the temperature
is 0° Celsius and the pressure is one
atmosphere (760 mm Hg).
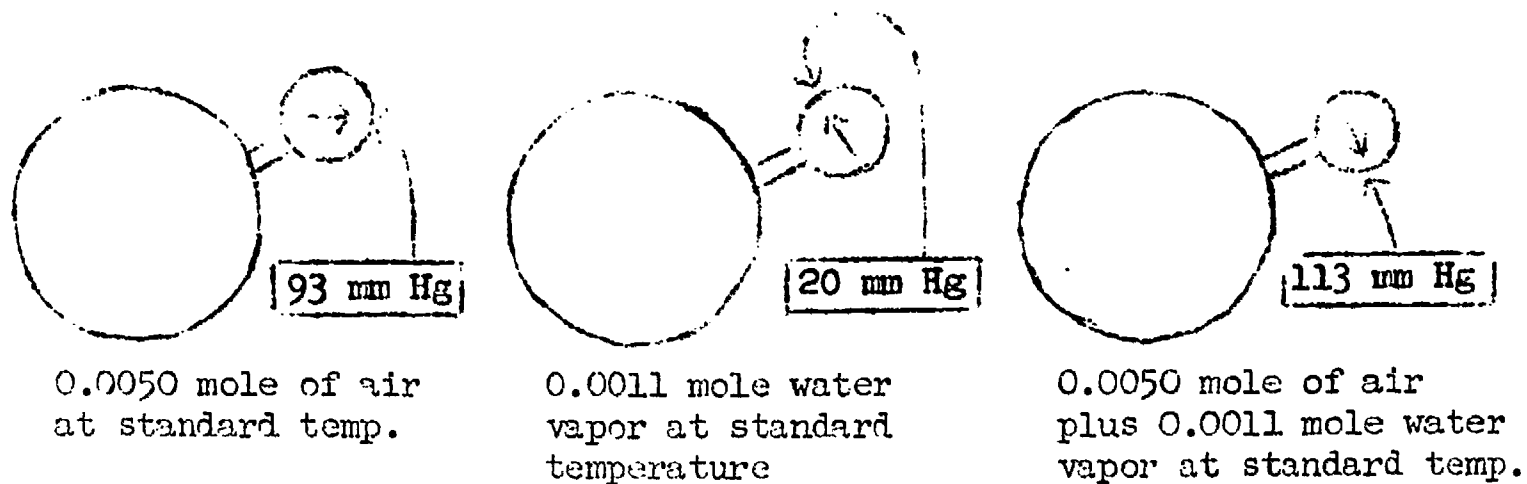
C. STANDARD TEMPERATURE AND PRESSURE (S. T. P.)

By arbitrary agreement among scientists throughout the world, standard temperature is 0° C and standard pressure is one atmosphere (760 mm Hg).

D. TOTAL PRESSURE AND PARTIAL PRESSURE

The figure below shows three bulbs of equal volume. The first bulb contains 0.0050 mole of air. The pressure gauge shows the pressure to be 93 mm Hg. The second bulb contains 0.0011 mole of water vapor. The pressure in the second bulb is 20 mm Hg. The third bulb contains 0.0050 mole of air and 0.0011 mole of water vapor. The pressure in the third bulb is 113 mm Hg. This indicates that the pressure exerted by the mixture of gases is the sum of the pressures of the separate components of the mixture.

113 mm Hg = 93 mm Hg + 20 mm Hg.

Because similar experiments yield the
same results, we can state: The total
pressure is the sum of the pressures
exerted by each of the components of
the gas mixture. The pressure exerted
by each of the gases in a gas mixture
is called the partial pressure of that
gas. In other words, the partial
pressure is the pressure each gas
would exert if it were all by itself in
the container.

| 93 mm Hg | | 20 mm Hg | | 113 mm Hg |

0.0050 mole of air
at standard temp.

0.0011 mole water
vapor at standard
temperature

0.0050 mole of air
plus 0.0011 mole water
vapor at standard temp.

If the 0.0050 mole of air contains
0.0040 mole of N_2 and 0.0010 mole
of O_2, what is the partial pressure of
O_2 in the first bulb; What is the
partial pressure of O_2 in the third
bulb? Don't read further until you
have answered these two questions.

D.1 - Experiment: A QUANTITATIVE
                   INVESTIGATION OF THE
                   REACTION OF MAGNESIUM
                   WITH HYDROCHLORIC ACID

This experiment gives you the op-
portunity to determine the volume of
hydrogen gas produced when a small
amount of magnesium reacts with hydro-
chloric acid. The hydrogen gas is col-
lected over water at room temperature
and pressure. The major purpose of this
experiment is to determine the volume of
dry hydrogen (that is, $H_2(g)$ without any
$H_2O(g)$) that would be produced by one
mole of magnesium.

Obtain a piece of magnesium ribbon
and carefully cut off a piece 5.00 cm
long. Your teacher will give you the
weight of 1.000 meter of Mg ribbon.
Since the ribbon is of uniform thickness
and width, you can calculate the weight
of the short piece of Mg you will use.
Record your data promptly (see p. 56).

Wad the Mg ribbon so that it can
be encased in a cage made of fine
copper wire. Leave about 5 cm of copper
wire to serve as a handle.

Position a ring stand and utility
clamp to hold a 50 ml gas measuring tube
which has been fitted with a one-hole
rubber stopper. Have a 400 ml beaker
about two thirds full of tap water handy.

Now tip the gas measuring tube
slightly to one side and pour in 10 ml
of the hydrochloric acid which is label-
ed 6 M HCl. With the tube in the same
position, carefully fill the tube with
water from your wash bottle. While
pouring, rinse any acid that may be on
the sides of the tube. This will help
make certain the liquid in the top of
the tube contains very little acid at
the very beginning. Bubbles clinging
to the sides of the tube can be dis-
lodged by gently tapping the tube.

Hold the copper coil by the handle
and insert the cage with the Mg ribbon
about 3 cm down the tube. Bend the
remaining copper wire over the edge
of the tube and clamp it there by in-
serting the rubber stopper. The tube
should be completely filled so that the

stopper displaces some water and some
water spurts out the hole.

Cover the hole in the stopper with
your finger and invert the tube in the
400 ml beaker. Clamp the tube in place.
The hydrochloric acid (HCl) is more
dense than water and it will diffuse
down and react with the metal.

After the reaction stops, wait for
about 5 minutes so that the apparatus
will come to room temperature. Tap
the tube gently to dislodge any bubbles
clinging to the side of the tube.

Now cover the hole in the stopper
with your finger and transfer the tube
to a large cylinder which is almost
full of water at room temperature.
The acid solution is dilute enough so
that it will not hurt you if you rinse
your hand with water after the transfer.
Raise or lower the tube until the level
of the solution inside the tube matches
the water level on the outside of the
tube. This permits you to measure the
volume of the gases (hydrogen and water
vapor) at room pressure. The pressure

inside the tube equals the pressure out-
side the tube (room pressure) when the
levels are the same. Why':

Read and record the volume with
the appropriate uncertainty.

Then remove the gas measuring tube
and clean the apparatus.

Finally, record the room pressure.
You may want to repeat the experiment
if you have enough time.

Your data table should include the
following:

Weight of 1.000 meters of Mg ribbon____

Weight of Mg sample used____

Volume of gases:_____

Temperature of water (room temperature):___

Barometer reading (room pressure):____

Partial pressure due to water vapor

at room temperature·_____

The following table will help you
determine the partial pressure due to
water vapor at room temperature:

VAPOR PRESSURE OF WATER AT VARIOUS TEMP-

ERATURES (To three significant figure)

Temperature (^{O}C)	Pressure (mm Hg)
20	17.5
21	18.6
22	19.8
23	21.0
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8

Calculations

1. Find the weight of the Mg you used from the grams per meter relation-ship and the length of your strip of Mg.

2. How many moles of Mg did you use?

3. Assuming the only two gases involved are $H_2(g)$ and $H_2O(g)$, determine the part of the pressure due to hydrogen gas.

4. What volume (V') would the $H_2(g)$ occupy if it were all by itself and the pressure were 760 mm Hg? (standard pressure)

5. What volume (V'') would the $H_2(g)$ occupy if the pressure were 760 mm Hg, and

the temperature were 0° Celsius?

(standard pressure and temperature).

6. According to your findings, how many liters of $H_2(g)$ at S. T. P. would be produced by one mole of magnesium?

Exercises for Home, Desk and Lab (HDL)

1. One volume of nitrogen gas will combine with three volumes of hydrogen gas to produce two volumes of ammonia gas. Use these data and Avogadro's Hypothesis to determine how many mol cules of hydrogen combine with one molecule of nitrogen. How many molecules of ammonia are produced from one molecule of nitrogen?

2. There are a number of gas mixtures that have the following combining volume ratios: one Vol. A + one B $\rightarrow$ 2 vol. C

If a = number of atoms in a molecule of A,

   b = number of atoms in a molecule of B,

   c = number of atoms in a molecule of C,

(a) Show that if a is even, b

must be even.

(b) Show that if a is odd, b

must be odd.

3. One volume of hydrogen gas combines

with one volume of fluorine gas to

produce two volumes of hydrogen

fluoride gas (as measured at the

same temperature and pressure). A

variety of other types of evidence

suggests that hydrogen is an ele-

ment and its molecules are diatomic

(two atoms per molecule).

(a) Which one of the following pos-

sible molecular formulas for the

substance fluorine is <u>not</u> consis-

tent with the volumes that com-

bine? (Use only the data given

here; do not presume the molecular

formula of hydrogen fluoride.)

(1) F_2 (3) H_2F_2 (5) H_2F_4
(2) F_4^2 (4) HF_2

(b) For each formula in part (a) that

<u>is</u> consistent with the combining

volumes data and the formula H_2

for hydrogen, calculate the

molecular weight indicated by

that formula.

(c) For each acceptable formula in

part (a), predict the molecular

formula for the substance hydro-

gen fluoride.

4. Four volumes of hydrogen fluoride

gas combine with two volumes of the gas

dinitrogen difluoride to form four

volumes of gas X.

(a) According to Avogadro's Hypoth-

esis how many molecules of X

are produced from one moelcule

of dinitrogen difluoride?

(b) If a = number of atoms in a

molecule of hydrogen fluoride,

b = number of atoms in a mole-

cule of dinitrogen difluoride,

c = number of atoms in a mole-

cule of X, write the relation

among a, b and c appropriate

to the combining volumes given.

(c) For each of the following possible

values of a and b, calculate the

required value of c.

If $\underline{a}$ is and $\underline{b}$ is then $\underline{c}$ is _____.

$\frac{a}{1}$	$\frac{b}{2}$
1	4
2	2
2	4

(d) No odd value of b is suggested in question (c). Prove that b must be an even integer.

5. Oxygen has a density of 1.429 g/L at S. T. P. and its molecular weight is 32.00. From these data, calculate the volume occupied by one mole of oxygen at S. T. P.

6. (a) What volume will a molecular weight of any gas occupy at S. T. P.?

(b) How many particles of any gas are in a molecular weight at S. T. P.?

7. By what factor will one mole of a gas change in volume if the pressure is reduced to 740 mm Hg?

8. By what factor will one mole of a gas change in volume if the temperature is decreased to 260 degrees Kelvin?

9. Suppose that 200 ml of hydrogen is collected over water at a temperature of 26° and a pressure of 750 mm Hg. The pressure of water vapor at 26° C is 25 mm Hg. What is the pressure of the hydrogen in the dry state at 26° C?

10. Suppose that 500 ml of oxygen is collected over water at a temperature of 23° C and a pressure of 800 mm Hg. Calculate the pressure of dry oxygen.

III. A USEFUL FORM OF P=kDT

    A.   VOLUME, TEMPERATURE RELATIONS

    B.   PRESSURE, VOLUME RELATIONS

    C.   COMBINED GAS LAWS

Chapter III: A USEFUL FORM OF P=kDT

In Chapter Eleven of Harvard Project
Physics, an ideal gas law, P=kDT was con-
structed. P=kDT can be further modified
and hence more easily applied to problems
involving volume, temperature, and pres-
sure of gases.

The density (D) of a gas is its mass
(m) divided by the volume (V) of the gas:

$$(1) \quad D = \frac{m}{V} \quad \text{where:} \quad \begin{array}{l} D = \text{density} \\ m = \text{mass} \\ V = \text{volume of gas} \end{array}$$

P=kDT can be rewritten as

$$(2) \quad P = (k)\frac{m}{V} T$$

A. VOLUME, TEMPERATURE RELATIONS

Equation (2) can be rearranged to:

$$(3) \quad V = \frac{km}{P} T$$

Note that the mass (m) is directly pro-
portional to the number of particles in-
volved, (recall Avogadro's Hypothesis)
hence assuming the number of particles to be
constant also assumes the mass to be con-
stant, then V becomes directly proportional
to T (the absolute temperature).

$$(4) \quad V \propto T$$

The <u>volume</u> <u>is</u> <u>directly</u> <u>proportional</u> <u>to</u>
<u>the absolute temperature</u> $(V \propto T)$,
<u>a relationship which is referred to as</u>
<u>Charles' Law.</u>

The relation between gas volume and
temperature can be determined experimental-
ly. Table A. 1. on the next page shows
the results of such an experiment. These
results were obtained by holding the
pressure and the mass of a gas constant,
and varying the volume and temperature

It is not possible to obtain data
at extremely low temperatures because
all gases ultimately condense into li-
quids. However, the experimentally
derived curve may be extrapolated to
the point which represents zero volume.
The temperature at this point is $-273^{\circ}C$.
This is designated as <u>absolute zero.</u>
This is, theoretically, the lowest
possible temperature attainable. Any
lower temperature would represent a
negative volume. Scientists have re-
corded temperatures as low as 0.0015

T vs. V for one mole of gas
(P constant at one atmosphere)

Temperature		Volume	V/T
°C	°K	liters	liters/°K
-20	253	20.9	0.0826
0	273	22.4	0.0821
100	373	30.2	0.0810

Table A.1.

When temperatures are plotted against
corresponding volumes a graph such as
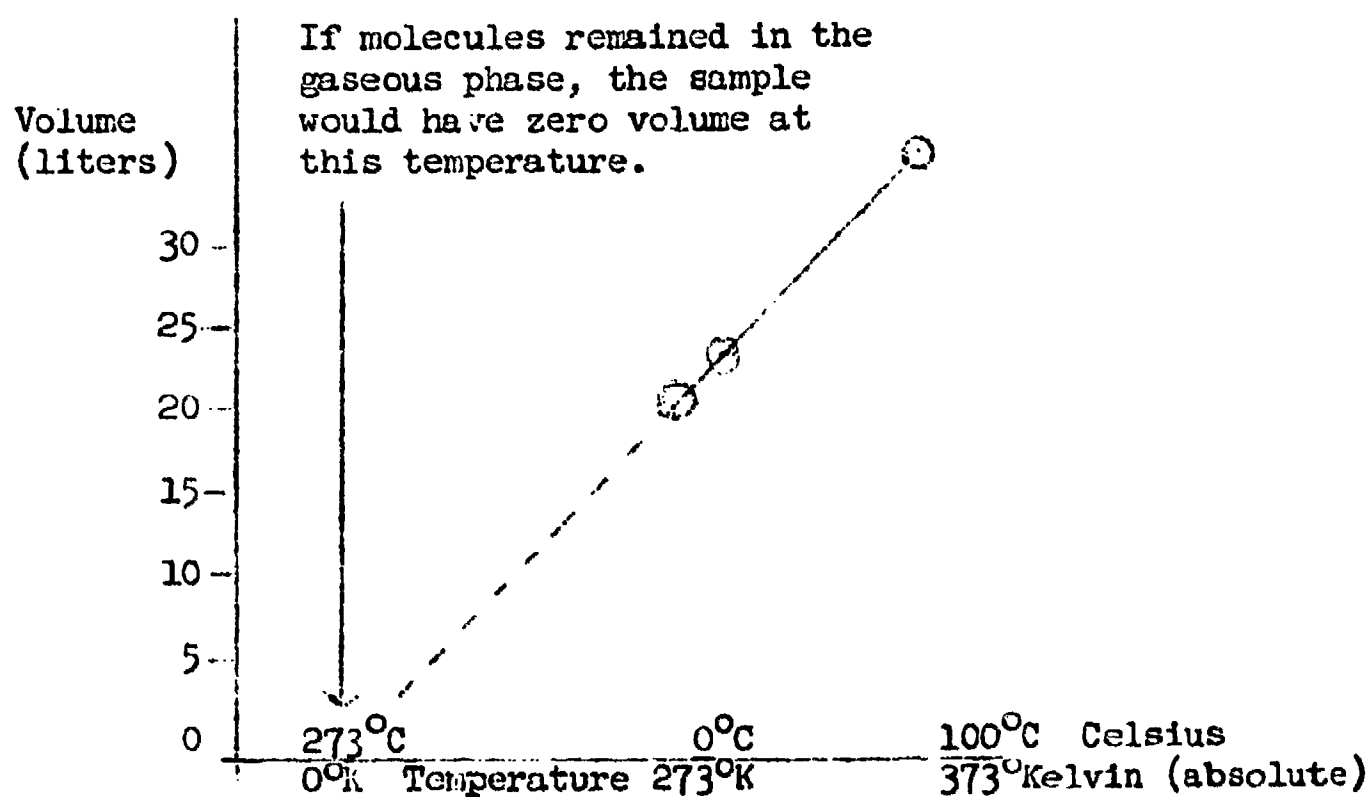that in Figure A.7. is obtained.

Figure A.1. - Volume of Gas vs. Temperature

degrees above absolute zero.

The Kelvin scale (^oK) is a temp-
erature scale based on an absolute zero.
Look at the additional lines marked
Celsius and Kelvin (absolute) below Figure
A.1. How would you convert Celsius temp-
eratures to Kelvin temperatures? You
must add 273° to the ^oC to obtain ^oK.
To change ^oK to ^oC you would subtract
273° from K.

T in degrees Kelvin = 273 + T in degrees

Celcius.

It is important to note that the
mathematical relationship which represents
gas behavior is simplified by using the
Kelvin temperature scale, because it is
easier to work with direct proportions.
Inspection of Figure A.1. shows that if
we use ^oK for the X axis, the volume will
vary by a factor of 1/273 for each degree
change in temperature.

B. PRESSURE, VOLUME RELATIONS

If we hold m and T constant, then

(5) $P \propto 1/V$

The volume is inversely propor-

tional to the pressure. $(P \propto 1/V)$

is often referred to as Boyle's Law.

C. COMBINED GAS LAWS

Now let's assume we have a closed

container with a movable piston (Figure

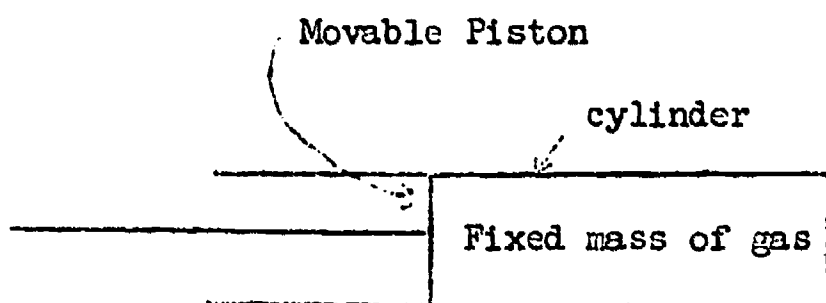A.2.) to which we will supply a definite

volume of gas.

Figure A.2.

The initial conditions will be

room temperature and atmospheric pres-

sure. Our initial conditions can be de-

signated by $\dfrac{PV}{T} = mk$. Now, let's move the

piston (change the volume) and apply heat

(change the temperature), which in each

case would affect the pressure. We now

have a new set of conditions for our

gas sample, and we will designate these new

conditions as $\dfrac{P'V'}{T'} = m'k'$.

When we examine the variables m and m', we observe that in both cases the mass of the gas remains constant. We did not add or subtract molecules, therefore m should be equal to m'. Since k depends on the type of gas, and we did not change the gas in the char er, k should equal k'. It follows that

$$mk = m'k'$$

Since mk = m'k':

$$(6) \quad \frac{PV}{T} = \frac{P'V'}{T'}$$

This combination of both Charles' and Boyle's Laws is called the combined gas law.

Let's use some algebra and rearrange our equation to find V' (the volume under new conditions). Upon rearrangement, we obtain:

$$(7) \quad V' = \left(\frac{P}{P'}\right)\left(\frac{T'}{T}\right)(V)$$

From this we can observe that we have two sets of ratios, one dealing with pressure and one dealing with temperature. Let's apply these ratios and the equation

in a sample problem.

Assuming a sample of O_2 gas in the piston apparatus occupies 100 ml at 27°C and 750 mm Hg Pressure, what volume will it occupy at standard temperature and pressure?

Solution:

1. First, itemize the given data as follows: P = 750 mm Hg

V = 100 ml

T = 27°C

but remember, it is essential to use the Kelvin temperature scale, so:

T_K = 273 + T_C

T_K = 273 + 27

T_K = 300

Where T_K = Temp. in degrees Kelvin

T_C = Temp. in degrees Celcius

Our problem states that the new conditions of the gas will be at standard temperature and pressure. Remember standard temperature is 0° C or 273°K and standard pressure is 760 mm Hg.

Therefore, itemize the new conditions as follows:

P' = 760 mm Hg

T' = 273°K

V' = ?

2. Use our generalizations about
   the behavior of gases to de-
   termine temperature and pressure
   ratios: If pressure increases
   then volume decreases; therefore,
   our $\frac{(P)}{(P')}$ ratio must produce a de-
   crease of volume (V). The pres-
   sure ratio is:

$$\frac{750 \text{ mm Hg}}{760 \text{ mm Hg}}$$

   If temperature decreases
   then volume decreases; there-
   fore, our $\frac{(T')}{(T)}$ ratio must pro-
   duce a decrease of volume. Thus
   the temperature ratio is:

$$\frac{273°K}{300°K}$$

3. To obtain the final volume (V'),
   multiply the initial volume (V)
   by the temperature and pressure
   fractions. The result is:

$$V' = 100 \text{ ml} \frac{(750 \text{ mm Hg})(273° K)}{(760 \text{ mm Hg})(300° K)}$$

$$V' \approx 89.6 \text{ ml}$$

4. Or we may use the combined gas
    law formula, $\dfrac{PV}{T} = \dfrac{P'V'}{T'}$

$$\frac{(750 \text{ mm Hg})(100 \text{ ml})}{(300 {}^{\circ}\text{K})} = \frac{(760 \text{ mm Hg})(V')}{(273 {}^{\circ}\text{K})}$$

Solving for V':

$$V' = \frac{(750 \text{ mm Hg})(100 \text{ ml})(273 {}^{\circ}\text{K})}{(300 {}^{\circ}\text{K})(760 \text{ mm Hg})}$$

$$V' = 89.6 \text{ ml}.$$

Exercises for Home, Desk and Lab (HDL)

1. (a) A thermometer containing hy-
    drogen is at the temperature of
    a carbon dioxide-acetone bath
    (-78° C). The volume of hy-
    drogen at standard atmospheric
    pressure is read. The ther-
    mometer is then moved into a
    bath of boiling water. The
    pressure on the hydrogen gas
    volume is still atmospheric
    pressure. By what factor does
    the volume of hydrogen change?

    (b) If the thermometer had a con-
    stant volume, by what factor

would the pressure of the hy-

drogen gas change?

(c) If we used nitrogen gas instead

of hydrogen in the thermometer,

how would it affect the answers

to (a) and (b)?

2. A student collects 200 ml of dry hy-

drogen at a temperature of 20° C.

What is the volume of dry hydrogen at

standard temperature if the pressure

doesn't change?

3. A student collects 400 ml of oxygen

at a pressure of 750 mm Hg. What is

the volume of oxygen at standard

pressure if the temperature doesn't

change?

4. A 15.0 liter sample of dry air in a

cylinder exerts a pressure of 3.00

atmospheres pressure at a tempera-

ture of 25° C. Without

change in temperature, a piston is

moved in the cylinder until the

pressure in the cylinder is lowered

to 1.00 atmospheres pressure. What

is the new volume of gas in the
cylinder?

5. 100 ml of carbon dioxide gas is col-
lected over water at a temp. of 20°
C and 800 mm Hg pressure. What is
the volume of dry carbon dioxide at
S. T. P?

6. A 300 ml sample of nitrogen at S. T.
P. is brought to 27° C and 800 mm Hg.
Calculate the new volume.

7. A sample of gas occupies 275 ml at
52° C and 720 mm Hg. Correct its
volume to S. T. P.

8. A sample of nitrogen is collected over
water at 23° C. The vapor pressure of
water at 23° is 21 mm Hg. When the
pressure on the sample has been equal-
ized against atmospheric pressure, 750
mm Hg, what is the partial pressure of
nitrogen? What will be the partial
pressure of nitrogen if the volume is
reduced by a factor, 720/760?

9. A cylinder contains oxygen gas and a
small amount of liquid water at a

temperature of 25° C. (The vapor pres-

sure of water at 25° C is 23.8 mm Hg.)

The total pressure is 700.0 mm Hg. A

piston is pushed into the cylinder un-

til the volume is halved. What is the

final total pressure if the temperature

of the gas doesn't change?

10. Would you expect the compressibility

of solids and gases to differ? Would

you expect the compressiblity of gases

and liquids to differ? Why?

IV. CHEMICAL EQUATIONS

    A. WRITING EQUATIONS FOR REACTIONS

    B. TYPES OF CHEMICAL REACTIONS

Chapter IV: CHEMICAL EQUATIONS

A. WRITING EQUATIONS FOR REACTIONS

Scientists have a rapid and concise way of representing the relative amounts of reactants and products involved in chemical reactions. This is done by writing chemical equations. Chemical equations include the usual symbols for the elements or compounds involved and numbers representing the number of moles of each chemical.

In order to predict quantitatively the amounts of reactants needed to produce specific amounts of products, a correctly balanced equation is needed. It is obvious one must know the reactants and products if he is going to balance something involving reactants and products.

A good chemist is one who can predict the products when chemicals (reactants) are put together under given conditions. At this point in the course however, your experience is not

extensive enough for you to be able to predict the products of chemical reactions. Until you become proficient in predicting, you will be provided with the names and formulas of both reactants and products. In the meantime, you should be looking for some definite patterns as we discuss different kinds of reactions.

General "Tasks" for Writing Equations

The general tasks for writing and balancing equations are few but need to be adhered to faithfully.

Task #1 -- The first step in writing an equation for a chemical reaction is to determine the reactants used and the products formed.

Task #2 -- The second step is to write the correct chemical formula for each reactant and each product. Reactants go on the left of the equation and are separated from the products by an arrow which is read "yields."

$$A + B \longrightarrow C + D$$

Reactant "A" plus Reactant "B" yields product "C" plus product "D".

Task #3 -- The law of conservation of mass must be satisfied. In accordance with this law, <u>the number of atoms in the system undergoing change must be constant.</u> A new atom cannot appear on the product side and no atom from the reactant side may disappear. You may feel the urge to change the formulas of reactants and products as you apply task #3. You must resist this urge because in changing them you are admitting you really don't have the correct formulas. You may change only the numbers that precede the formulas.

Example: Solid copper will react with an aqueous solution of silver nitrate to produce solid silver plus an aqueous solution of copper nitrate, and, if you know the respective formulas: Cu, $AgNO_3$, Ag, $Cu(NO_3)_2$, you have satisfied task #1.

Let's satisfy task #2 --

$$Cu + AgNO_3 \longrightarrow Ag + Cu(NO_3)_2$$
(Eq.1)
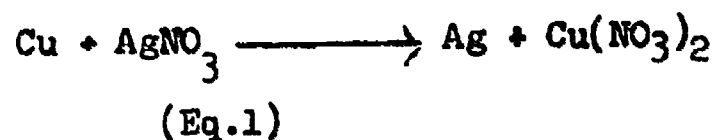
We now must satisfy task #3. We must be certain that our representation is consistent with our belief in the conservation of atoms. We have one mole of "Cu" on the left and one on the right. Good! Let's check the "Ag". We have one mole of "Ag" on the left and one on the right. Fine! Now let's check "N".

COPS!!!

One mole on the left but two on the right. The subscript number after the parenthesis means that everything inside the parentheses is doubled. Hence two moles of nitrogen atoms and six moles of oxygen atoms as well as one mole of copper atoms are represented by $Cu(NO_3)_2$.

Let's put "$\frac{1}{2}$" before the $Cu(NO_3)_2$. Now we have --

$$Cu + AgNO_3 \longrightarrow Ag + \frac{1}{2} Cu(NO_3)_2$$
(Eq. 2)
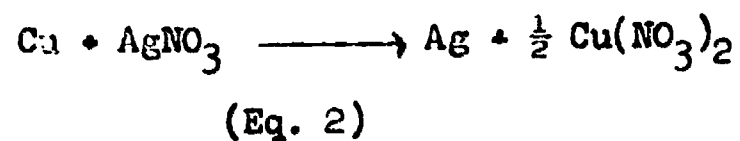
Now the "N" balances with one mole

on each side. A quick check shows three

moles of "O" on each side. <u>But</u> you complain,

we have now represented one mole of "Cu" on

the left and only $\frac{1}{2}$ mole on the right.

Perhaps the equation will balance if we

put "$\frac{1}{2}$" before the "Cu" on the left.

$$\frac{1}{2} Cu + AgNO_3 \longrightarrow Ag + \frac{1}{2} Cu(NO_3)_2$$

(Eq. 3)
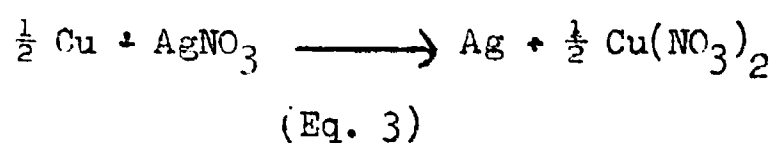
HURRAY !!

Another check shows that we are firm

believers in the conservation of mass

and that we have completed all the tasks.

It should be noted here that we could

have chosen 1 mole of Cu metal as the a-

mount of reactant consumed in this re-

action. We have already seen that we

need twice as many molecules of $AgNO_3$ as

molecules of Cu.

Therefore the reaction --

$$Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2 Ag$$

(Eq. 4)
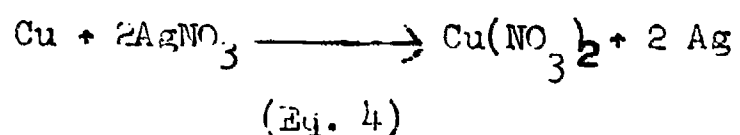
is also a chemical reaction in which

atoms are conserved. It is just as

correct as equation (3). To show this
we can multiply equation (3) by 2 in
order to obtain equation (4). Remember
that all coefficients in an equation
may be multiplied or divided by a common
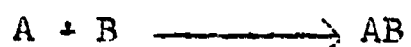factor, resulting in an equally valid
equation.

There are two other tasks that will
need attention as more advanced chemical
reactions are studied. The first of
these is that the number of electrons
gained must equal the number lost in
the reaction and the second is that the
overall charge of the reactants must
equal the overall charge of the products.
These aspects of balancing equations
will be considered when the moon is in
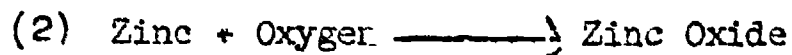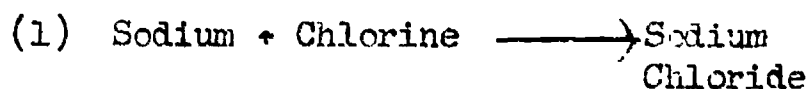the seventh house.

B. TYPES OF CHEMICAL REACTIONS

Chemists have found that most chemi-
cal reactions tend to group themselves into
four general categories or types of reactions.
They are as follows:

Type I: Composition (synthesis) reac-
tions are those in which 2 or more substances
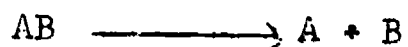react to form a more complex substance.
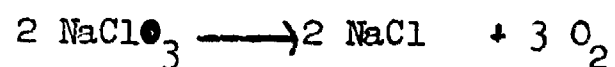Composition reactions have the general form:

A + B ————————→ AB

Examples:

(1) Sodium + Chlorine ————→Sodium
                                    Chloride

2Na + Cl$_2$ ————————→ 2 NaCl

(2) Zinc + Oxygen ————————→ Zinc Oxide
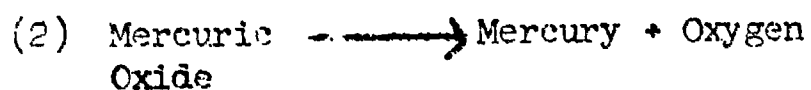
2 Zn + O$_2$ ————————→ 2 ZnO
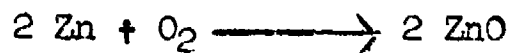
Type II: Decomposition reactions are
those in which a complex substance will
break down or separate into two or more
less complex substances. You will note
that decomposition reactions are the re-
verse of composition (synthesis) reactions.
Decomposition reactions have the general form:

AB ————————→ A + B

Examples:

(1) Sodium ————→Sodium   + Oxygen
     Chlorate   Chloride

2 NaClO$_3$ ————→2 NaCl   + 3 O$_2$

(2) Mercuric —————⟶Mercury + Oxygen
     Oxide

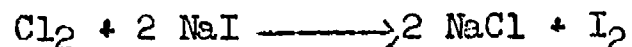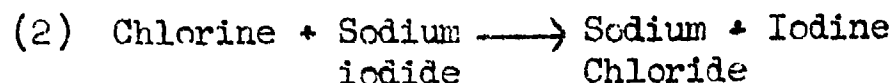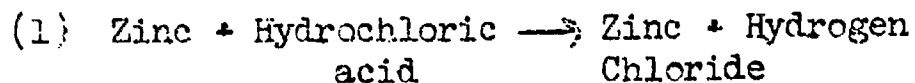   2 HgO —————⟶ 2 Hg + O_2

Type III: Single replacement reac-

tions are those in which one substance is

replaced in a given compound by another

substance. A single replacement reaction
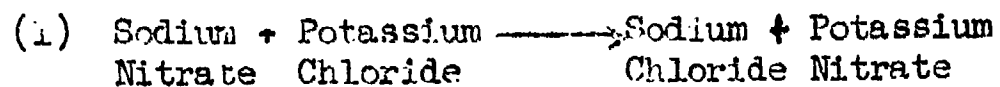
has the general form:

     A + BX —————⟶ AX + B

Examples:

(1) Zinc + Hydrochloric ——⟶ Zinc + Hydrogen
                  acid   Chloride

     Zn + 2 HCl ————⟶$ZnCl_2$ + H_2

(2) Chlorine + Sodium ——⟶ Sodium + Iodine
                  iodide   Chloride

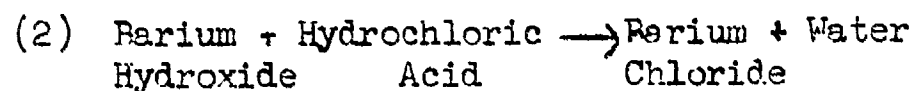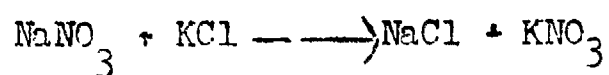 Cl_2 + 2 NaI ————⟶2 NaCl + I_2

Type IV: Double replacement reactions

are those in which two compounds interact

and exchange their component parts to form

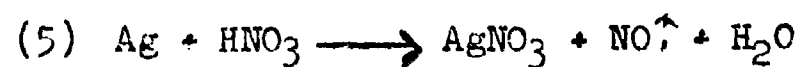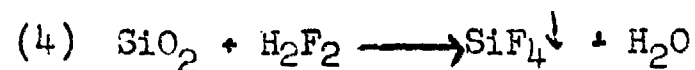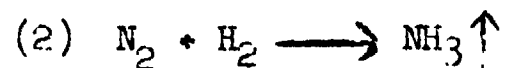two new compounds. A double replacement
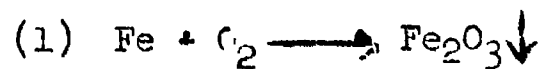
reaction has the general form:

     AX + BY —————⟶ AY + BX

Example:

(1) Sodium + Potassium ———⟶Sodium + Potassium
     Nitrate   Chloride   Chloride Nitrate

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

(2) Barium + Hydrochloric $\longrightarrow$ Barium + Water
    Hydroxide   Acid   Chloride
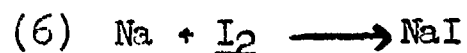
$$Ba(OH)_2 + 2 HCl \longrightarrow BaCl_2 + 2 H_2O$$

Exercises for Home, Desk and Lab (HDL)

Write balanced equations for each of
the following chemical reactions. The
symbol "$\downarrow$" means the element or compound
is a solid precipitate and the symbol "$\uparrow$"
means the element or compound is a gas.
A "$\triangle$" over the "yield" symbol means
the addition of heat energy to the
system.

(1) $Fe + O_2 \longrightarrow Fe_2O_3\downarrow$

(2) $N_2 + H_2 \longrightarrow NH_3\uparrow$

(3) $KClO_3 \xrightarrow{\triangle} KCl + O_2\uparrow$

(4) $SiO_2 + H_2F_2 \longrightarrow SiF_4\downarrow + H_2O$

(5) $Ag + HNO_3 \longrightarrow AgNO_3 + NO\uparrow + H_2O$

Balance the equations for each of the
following reactions. Begin on the basis
of one mole of the substance underscored.

(6) $Na + \underline{I_2} \longrightarrow NaI$

(7) $C_2H_2 + O_2 \longrightarrow H_2O\uparrow + CO_2\uparrow$

(8) $H_2 + O_2 \longrightarrow H_2O\uparrow$

(9) $\underline{Ca} + Cl_2 \longrightarrow CaCl_2$

(10) $NO_2 + H_2O \longrightarrow \underline{HNO_3} + NO\uparrow$
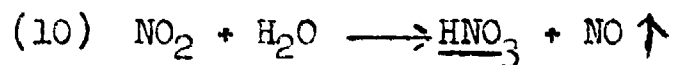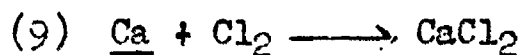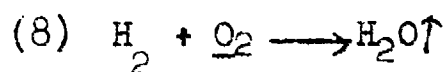
Write a balanced equation and in-
dicate the reaction type (composition, de-
composition, single or double replace-
ment) for each of the following word
equations:

(You may wish to consult the Hand-
book of Physics and Chemistry in order
to obtain the correct formulas for the
reactants and products involved in these
reactions.)

(11) Sodium + Water $\longrightarrow$ Sodium + Hydrogen
                                                   Hydroxide

(12) Calcium hydroxide + Phosphoric acid —

            $\longrightarrow$ Calcium Phosphate + Water

(13) Barium Nitrate + Sodium Sulfate $\longrightarrow$

       Barium Sulfate + Sodium Nitrate

(14) Ammonium + Potassium $\longrightarrow$
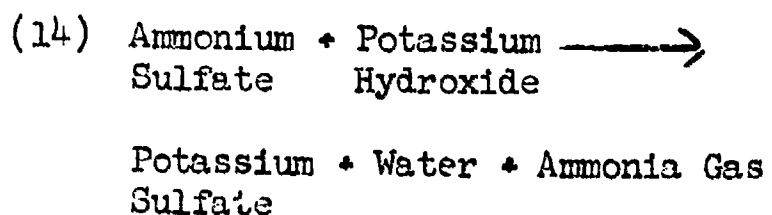   Sulfate   Hydroxide

       Potassium + Water + Ammonia Gas
       Sulfate

(15) Chlorine + Magnesium Bromide $\longrightarrow$

       Magnesium Chloride + Bromine
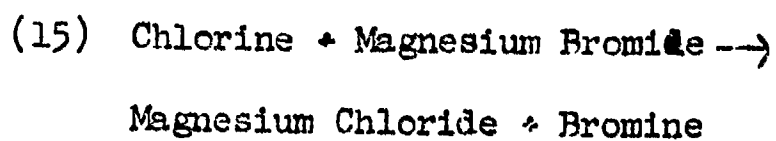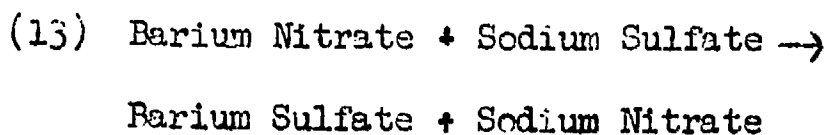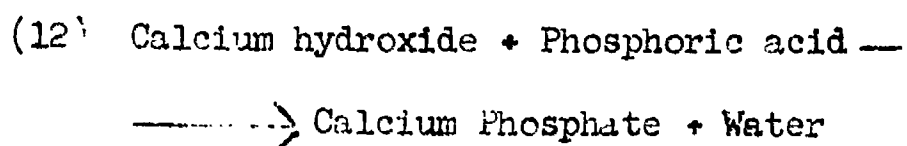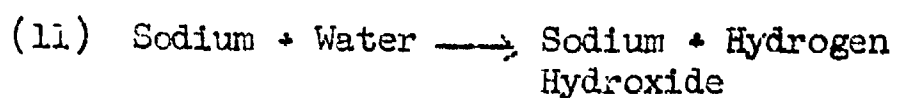
Propane, a commercially bottled gas, has the formula C_3H_8. When it is burned completely the reaction products are carbon dioxide and water vapor.

(16) Write the balanced equation for the combustion of propane.

(17) One mole of propane vapor produces how many moles of water vapor when burned:

(18) How many moles of oxygen would be needed in order to burn 0.2 moles of propane?

(19) How many moles of propane would be needed to produce one mole of carbon dioxide?

(20) How many molecules of oxygen would be needed in order to burn one mole of propane?

V. ELECTRICAL NATURE OF MATTER

    A. PURE SUBSTANCES
        Read sections 5-1 through 5-1.3 CHEMS

    B. SOLUTIONS
        Read sections 5-2 through 5-2.7 CHEMS

    C. ELECTRICAL EFFECTS
        Read sections 5-3.1 through
        5-3.5 CHEMS

    D. ELECTRICAL PROPERTIES OF CONDENSED
        PHASES
        Read sections 5-4.4 CHEMS

  D.1. Experiment: REACTIONS BETWEEN
       IONS IN AQUEOUS SOLUTION
       (CExpt. 11)

VI. BASIC PARTICLES

   A. STRUCTURE OF THE ATOM
       Read sections 6-1 through 6-1.5
       CHEMS

   B. THE SIMPLEST CHEMICAL FAMILY -
       THE INERT GASES
       Read sections 6-2 through 6-2.4
       CHEMS

   C. THE ALKALIES
       Read sections 6-3 through 6-3.2
       CHEMS

   D. THE HALOGENS
       Read sections 6-4 through 6-4.3
       CHEMS

   E. HYDROGEN - A FAMILY BY ITSELF
       Read sections 6-5 through 6-5.2
       CHEMS

   F. THE THIRD ROW ELEMENTS
       Read sections 6-6 through 6-6.2
       CHEMS

   G. THE PERIODIC TABLE
       Read sections 6-7
       CHEMS

VII. ENERGY EFFECTS IN CHEMICAL REACTIONS

    A. HEAT AND CHEMICAL REACTIONS
        Read sections 7-1 through 7-1.4
        CHEMS

    A.1 Experiment: A STUDY OF REACTIONS
        (CExpt. 12)

    A.2 Experiment: THE HEAT OF REACTION
        (CExpt. 13)

    B. THE LAW OF CONSERVATION OF ENERGY
        Read sections 7-2 through 7-2.4
        CHEMS

    C. THE ENERGY STORED IN A MOLECULE
        Read sections 7-3 through 7-3.2
        CHEMS

    D. THE ENERGY STORED IN A NUCLEUS
        Read sections 7-4 through 7-4.1
        CHEMS

VIII. RATES OF REACTIONS

A. FACTORS AFFECTING REACTION RATES
    Read sections 8-1 through 8-1.5
    CHEMS

A.1 Experiment: A STUDY OF REACTION
    RATES
    (CExpt. 14)

B. THE ROLE OF ENERGY IN REACTION
    RATES
    Read sections 8-2 through 8-2.4
    CHEMS

IX. EQUILIBRIUM

    A.   QUALITATIVE ASPECTS OF EQUILIBRIUM
         Read sections 9-1 through 9-1.7
         CHEMS

    A.1 Experiment: CHEMICAL EQUILIBRIUM
         (CExpt. 15)

    B.   QUANTITATIVE ASPECTS OF EQUILIBRIUM
         Read sections 9-2 through 9-2.4
         CHEMS

X. SOLUBILITY

    A.   SOLUBILITY: A CASE OF EQUILIBRIUM
            Read sections 10-1 through 10-1.4
            CHEMS

    B.   AQUEOUS SOLUTIONS
            Read sections 10-2 through 10-2.6
            CHEMS

    B.1   Experiment: DETERMINATION OF THE
            SOLUBILITY PRODUCT CONSTANT OF
            SILVER ACETATE
            (CExpt. 16)

XI. ACID - BASE

    A.   ELECTROLYTES - STRONG OR WEAK
          Read sections 11-1 through 11-1.3
          CHEMS

    A.1 Experiment: THE HEAT OF SOME
          ACID-BASE REACTIONS
          (CExpt. 17)

    B.   EXPERIMENTAL INTRODUCTION TO
          ACIDS AND BASES
          Read sections 11-2 through 11-2.7
          CHEMS

    B.1 Experiment: DETERMINATION OF THE
          HYDROGEN ION CONCENTRATION OF
          SOLUTIONS USING INDICATORS
          (CExpt. 18)

    C.   STRENGTHS OF ACIDS
          Read sections 11-3 through 11-3.5
          CHEMS

    C.1 Experiment: REVERSIBLE CHEMICAL
          REACTIONS
          (CExpt. 19)

XII. OXIDATION - REDUCTION

A. ELECTROCHEMICAL CELLS
    Read sections 12-1 through 12-1.4
    CHEMS

A.1 Experiment: INTRODUCTION TO OXI-
    DATION REACTIONS
    (CExpt. 20)

A.2 Experiment: ELECTROCHEMICAL CELLS
    (CExpt. 21)

B. ELECTRON TRANSFER AND PREDICTING
    REACTIONS
    Read sections 12-2 through 12-2.5
    CHEMS

C. BALANCING OXIDATION-REDUCTION
    REACTIONS
    Read sections 12-3 through 12-3.3
    CHEMS

D. ELECTROLYSIS
    Read section 12-4
    CHEMS

XIII. STOICHIOMETRY

   A.· A PATTERN FOR STOICHIOMETRIC
        CALCULATIONS
        Read section 13-1
        CHEMS

   B.   THE MANUFACTURE OF SULFURIC ACID
        Read sections 13-2 through 13-2.5
        CHEMS

   B.1 Experiment: REACTIONS BETWEEN IONS
        (CExpt. 22, Parts I and II)

   B.2 Experiment: QUANTITATIVE TITRATION
        (CExpt. 23)