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ABSTRACT

This course in chemistry for energy technology is one of 16 courses in the Energy Technology Series developed for an Energy Conservation-and-Use Technology curriculum. Intended for use in two-year postsecondary technical institutions to prepare technicians for employment, the courses are also useful in industry for updating employees in company-sponsored training programs. Comprised of five modules, the course is designed with a special emphasis on all aspects of chemistry as it relates to the work of an energy technician. Basic chemical information and techniques are presented. (Chemistry for Energy Technology II is available separately as CE 030 777.) Written by a technical expert and approved by industry representatives, each module contains the following elements: introduction, prerequisites, objectives, subject matter, exercises, laboratory materials, laboratory procedures (experiment section for hands-on portion), data tables (included in most basic courses to help students learn to collect or organize data), references, and glossary. Module titles are Safety in Chemical Operations; Structure of Matter; Chemical Equations and Calculations; Refrigeration, Gases, and Air Pollution; and Solutions. (YLB)

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CHEMISTRY
FOR
ENERGY TECHNOLOGY I

CENTER FOR OCCUPATIONAL RESEARCH AND DEVELOPMENT
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PREFACE

ABOUT ENERGY TECHNOLOGY MODULES

The modules were developed by TERC-SW for use in two-year postsecondary technical institutions to prepare technicians for employment and are useful in industry for updating employees in company-sponsored training programs. The principles, techniques, and skills taught in the modules, based on tasks that energy technicians perform, were obtained from a nationwide advisory committee of employers of energy technicians. Each module was written by a technical expert and approved by representatives from industry.

A module contains the following elements:

Introduction, which identifies the topic and often includes a rationale for studying the material.

Prerequisites, which identify the material a student should be familiar with before studying the module.

Objectives, which clearly identify what the student is expected to know for satisfactory module completion. The objectives stated in terms of action-oriented behaviors, include such action words as operate, measure, calculate, identify and define, rather than words with many interpretations, such as know, understand, learn and appreciate.

Subject Matter, which presents the background theory and techniques supportive to the objectives of the module. Subject matter is written with the technical student in mind.

Exercises, which provide practical problems to which the student can apply this new knowledge.

Laboratory Materials, which identify the equipment required to complete the laboratory procedure.

Laboratory Procedures, which is the experiment section, or "hands-on" portion of the module (including step-by-step instruction) designed to reinforce student learning.

Data Tables, which are included in most modules for the first year (or basic) courses to help the student learn how to collect and organize data.

References, which are included as suggestions for supplementary reading/viewing for the student.

Glossary, which defines and explains terms or words used within the module that are uncommon, technical, or anticipated as being unfamiliar to the student.

CONTENTS

CHEMISTRY FOR ENERGY TECHNOLOGY

INTRODUCTION

Chemistry for Energy Technology is a course designed with a special emphasis on all aspects of chemistry as it relates to the work of an energy technician.

The first module in the course, "Safety in Chemical Operations," is devoted to the hazards of working with dangerous chemicals since the job of an energy technician may involve the use of flammable materials, explosives, and acids. Hazards associated with chemicals are considered, as well as possible hazards involved in the use of laboratory equipment.

Chemistry is the branch of science that deals with composition of matter and changes that take place in matter. Many of these changes, or reactions of matter, consume or produce energy – an important consideration to the energy technician. Module CH-02, "Structure of Matter," discusses chemical bonding, elements, compounds, electrons, protons, neutrons, atoms, and molecules, and it explains the use of symbols and formulas used in chemical equations.

Chemical equations are important since they can be used to calculate such things as the amount of energy involved in chemical reactions – for example, the combustion of natural gas. Another important function of chemical equations is to calculate and report the results of analytical laboratory procedures such as titrations and gravimetric analyses. The writing and balancing of these equations are discussed in Module CH-03, "Chemical Equations and Calculations." In addition, this module investigates some of the quantitative information that can be determined through the use of chemical equations.

Air conditioning and refrigeration systems are important factors for the energy technician to consider since heating and cooling of homes and businesses account for approximately 50% of their total energy consumption. In Module CH-04, "Refrigeration, Gases, and Air Pollution," the fundamentals of refrigeration are studied, as well as proper maintenance and control of heating and cooling systems as a means for conserving energy. This module also discusses gases and air pollution.

Most chemical reactions are conducted in the liquid phase; therefore, Module CH-05, "Solutions," presents a study of unsaturated, saturated, and supersaturated solutions, as well as solubility and rate of dissolution. Properties of liquids, acids, bases, and salts are examined in this module, along with an important laboratory technique called titrimetric analysis.

Corrosion damage to metal objects amounts to a loss of approximately $10 billion each year in the United States. Corrosion and some of the means available to control corrosion are considered in Module CH-06, "Corrosion and Electrochemistry." Energy technicians are expected to encounter many corrosion related problems during their careers.

Module CH-07, "Metals," and Module CH-10, "Plastics, Adhesives, and Lubricants," deal with the three major materials families – metals, plastics, and ceramics. The properties and uses of these materials are presented, as well as those of lubricants and adhesives, which are widely used in the energy field.

The principles of thermodynamics are considered in Module CH-08, "Thermodynamics and Thermochemistry." These principles are important in power plants, refrigeration, and other devices that use heat to produce power or that use engines to add or remove heat. Topics included are heat, specific heat, refrigeration capacity, heats associated with changes of state, and solar energy storage. Thermal chemistry is the study of the heat that accompanies chemical reactions, such as the combustion of fuel. The method of experimentally determining these heats of reaction is presented in this module.

The sources, composition, and processing of solid, liquid, and gaseous fuels are discussed in Module CH-09, "Fuels." In addition to conventional fuels, alternate energy sources must be developed and utilized to some extent in supplying the world's energy needs; therefore, methanol, ethanol, hydrogen, biomass, solar, nuclear, wind, tidal power, hydroelectric, and geothermal sources of energy are included in the module. Much of the work of an energy technician will be closely related to the production and efficient consumption of fuels.

Nuclear reactions are accompanied by energy changes that are several orders of magnitude greater than those associated with ordinary chemical reactions. These large amounts of energy that are evolved from nuclear reactions are the primary reason for their importance. Because nuclear energy has the potential of supplying much of the world's energy needs, Module CH-11, "Nuclear Chemistry," presents a discussion of nuclear reactions. Radioactive isotopes — which are widely used in medicine, industry, and chemistry — is also a topic that is discussed, as well as fission and fusion processes.

In summary, <u>Chemistry for Energy Technology</u> is an energy-specific chemistry course designed for the student interested in pursuing a career in the energy field. The basic chemistry information and techniques presented in the

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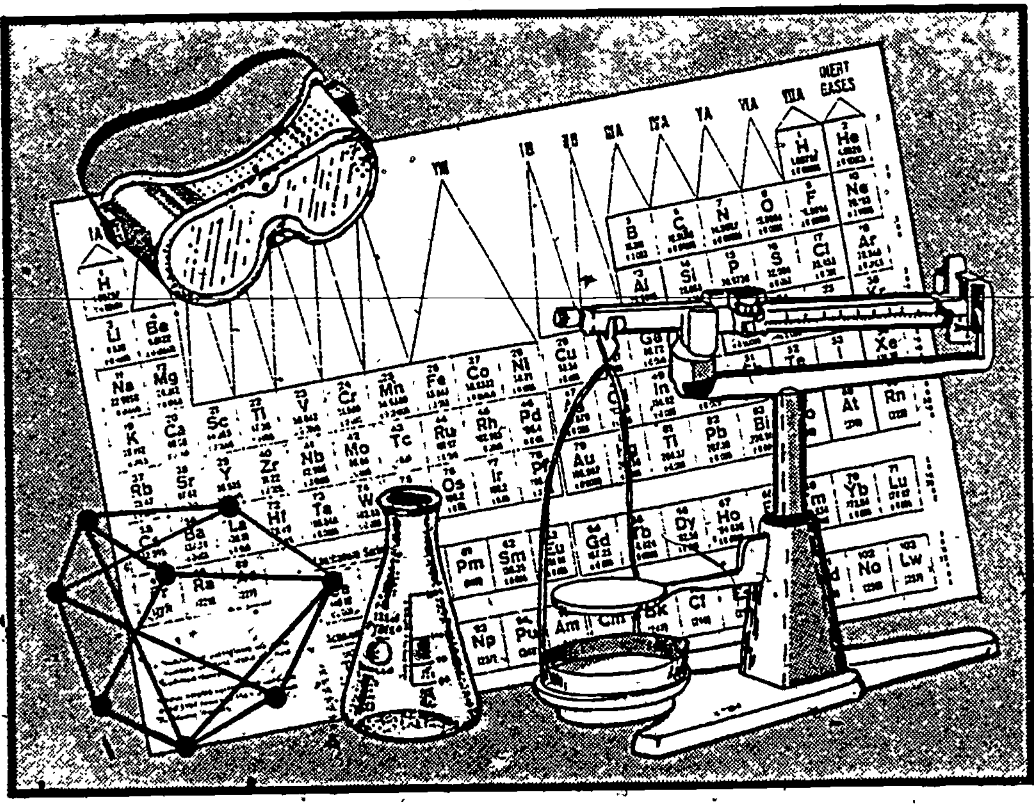
11 modules of this course have been deemed necessary for the applications that will be encountered by the energy technician.

ENERGY TECHNOLOGY

CONSERVATION AND USE

CHEMISTRY FOR ENERGY TECHNOLOGY I



MODULE CH - 01

SAFETY IN CHEMICAL OPERATIONS

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INTRODUCTION

Business and industry demand safety-conscious employees. Good safety practices are important to an employer, whose interest lies in the safety of each employee and the protection of property and equipment. Accidents cost an employer; they cause loss of production, damage to equipment and facilities, and often cause expenses such as medical treatment for injuries.

Because chemical operations involve the use of explosives, flammable materials, and acids – which can cause severe burns – potential hazards are a primary consideration. In spite of these many hazards, chemical laboratories generally remain among the safest places to work. In fact, a chemical laboratory is actually safer than the average home because a lab and its contents have been designed for safe operation. Moreover, the chemical workers have been trained in safe practices.

Each experiment or operation should begin with the anticipation and careful consideration of possible dangers involved. This module discusses some of the hazards of working with chemicals, as well as measures that can be taken to minimize these hazards. Learning laboratory safety habits can be compared to learning to drive an automobile safely: initially one must concentrate on safety, but, once properly learned, safe practices become second-nature and do not impede progress.

PREREQUISITES

None.

OBJECTIVES

Upon completion of this module, the student should be able to:

1. Define the following terms:
   a. Combustion.
   b. Flash point.
   c. Ignition temperature.
   d. Auto-ignition temperature.
   e. Pyrophoric materials.
   f. Acute poisoning.
   g. Chronic poisoning.
2. Identify the hazards associated with the following:
   a. Strong acids.
   b. Alkaline materials.
   c. Toxic chemicals.
   d. Volatile liquids.
   e. Handling and assembling glassware.
   f. Using various heating devices.
   g. Disposal of chemical wastes.
   h. Vacuum apparatus.
   i. Perchloric acid.
   j. Mercury.
   k. Carbon tetrachloride.
   l. Benzene.
   m. Electrical equipment.
   n. Excessive pressure.
   o. Cryogenics.
3. Describe proper procedures and hazards associated with the following aspects of handling compressed gases:
   a. Transporting cylinders.
   b. Selection of regulator.

4. List three sources of ignition of chemicals.
5. Identify three common ways in which toxic chemicals enter the body, and precautions to be taken to prevent ingestion.
6. Describe circumstances in which the following protective equipment should be used:
   a. Safety glasses.
   b. Face shield.
   c. Safety shield.
   d. Safety showers.
   e. Fire extinguisher.
   f. Fume hoods.
   g. Gloves.
   h. Lab coats and aprons.
   i. Safety fountain.

LABORATORY SAFETY RULES

The term "chemicals" includes many substances that are essentially harmless, such as sodium chloride (table salt), sodium bicarbonate (baking soda), 5% acetic acid solution (vinegar), and ascorbic acid (vitamin C). However, this term also includes materials, such as ethers and solvents, that are fire and explosion hazards. Similarly, it includes pesticides so toxic that one drop on the skin can cause instant death, and acids that can cause severe eye or flesh damage. Some materials are hazardous when they are extremely hot or extremely cold — even though they are not dangerous at room temperature. In working with chemicals and laboratory equipment, one must be familiar with the type of hazards that might be encountered. The following is a list of safety measures that should be heeded:

1. Before starting a task or experiment, be aware of the location of safety equipment, such as gas masks, fire blankets, fire extinguishers and hoses, alarm systems, eyewash fountains, safety showers, and emergency exits.
2. Never work in a laboratory alone.
3. Wear protective equipment appropriate to the work being done: safety glasses at all times; protective clothing when corrosive materials are being handled; rubber or latex gloves when handling poisonous materials; asbestos gloves for hot materials; or safety goggles or face shield when the possibility exists of a liquid splattering or an explosion occurring.

4. Because many chemicals are poisonous, never taste one, and avoid smelling them. (If it is necessary to determind the odor of a chemical, a cautious, exploratory whiff may be taken by bringing the odor just slightly toward the nose.)
5. Never eat, drink, or smoke in a laboratory.
6. Properly dispose of waste materials – extreme care should be taken in disposing of any unknown chemicals.
7. Properly label all reagents; do not use unlabeled chemicals.
8. Check for fumes before lighting burners or using electrical apparatus.
9. Before using any unfamiliar chemical, learn something about its possible hazards; consult an instructor, literature, or a fellow worker for information. Read labels on reagents for possible warnings.
10. Be certain to use the exact reagent called for by a procedure. When reading the name of a complex organic compound, be certain that the name (including each letter and number) identifies the exact compound that is wanted. (For example, 2, 2-dimethylhex<u>a</u>ne is very different from 2, 2-dimethylhex<u>e</u>ne, even though the names differ by only one letter.)
11. Perform procedures involving skin-irritating or dangerous fumes in a fume hood.
12. If chemicals come into contact with skin or eyes, wash immediately with large amounts of water for at least 15 minutes (at an eyewash fountain if possible).
13. Never point a test tube being heated at oneself or anyone else; it may erupt like a geyser and splash its contents over a large area.

14. Never perform unauthorized experiments.
15. The laboratory is no place for a "practical joker."
16. Be very careful in handling glassware; broken glassware is a hazard, as may be the contents of a broken vessel.
17. Many commonly used laboratory reagents, such as alcohol, acetone, and ether, are highly flammable; do not use them near an open flame.
18. Observe special precautions that are given for any experiment.
19. Notify the instructor or a co-worker(s) of any accident.
20. Keep the work area clean; proper "housekeeping" can eliminate many possible hazards.

The words "use adequate ventilation" are often seen on labels or in instructions. In general, a chemical so noted should be used only in a fume hood or in a laboratory with an exhaust fan in operation. Most people are aware of the effect of a single overexposure to toxic vapors, but, often, they fail to consider the cumulative effect of daily exposure to smaller quantities of such vapors. Both types of overexposure must be avoided. When working with chlorine, hydrogen sulfide, hydrogen cyanide, and other highly toxic gases, one should always wear a gas mask that has an absorbent specifically designed for the gas being used.

Glassware is used in laboratories because it is (1) inert (does not react with most chemicals), (2) fairly easy to clean, (3) relatively inexpensive, and (4) convenient to use because it is transparent. However, glassware is fragile and does require careful handling. For instance, special care should be taken to avoid breaking the glass when inserting thermometers or glass tubing into a cork or rubber stopper. The result could be a serious cut from the jagged ends. Glassware that is cracked, chipped, or has jagged edges should be discarded.

Glassware designed to be used in operations that require heating over hot plates or open burners is made from "hard" glass. Pyrex and Kimax, which are brands of hard glass, are commonly used to make test tubes, beakers, and flasks that are to be heated. Most reagent bottles, graduated cylinders, and other thick-walled glassware are made of "soft" glass, which will break if heated on a hot plate, a burner, or in an oven.

Care also must be taken in using glassware for vacuum operations, or in operations in which excess pressure may develop as a result of a chemical reaction. In either case, shattering of the glass apparatus may occur, causing broken glass to fly through the air — an obvious hazard to anyone in the laboratory.

In many chemical operations, heat is used because most chemical reactions occur at a faster rate as temperature increases. Gas-fired burners, electric mantles, heating tapes, and water, oil, and sand baths are commonly used in chemical operations. Many of these devices have self-contained, adjustable thermostats that spark as they automatically turn off and on. This spark may ignite the vapors from a volatile substance, causing a serious fire or explosion. A heating operation involving a volatile substance must be performed in a fume hood, which removes vapors as they escape, thereby preventing accumulation of enough vapors to ignite.

When a liquid is heated to its boiling point in a glass container, it may not boil smoothly; often it will superheat

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in spots and suddenly boil over. "Boiling chips" may be added prior to heating a liquid to prevent uneven heating and boil-over. Heat should be applied evenly over a large area of solvent.

If a container of solvent does catch fire, the fire may often be smothered quickly by covering the container with a watch glass, an asbestos pad, or some other object that can block out the air.

Thermal burns and cuts on one's skin – the most common result of laboratory accidents – can be avoided by exercising due care.

When purchased, chemicals often are nearly 100% pure; if not, their impurities are clearly shown on their labels. Through improper handling/storage techniques, these chemical reagents may become contaminated and, when used, unwanted side reactions may occur. Such side reactions may be dangerous or cause wasted time and effort because the experimental results are not reproducible. A chemical should never be returned to its reagent bottle. Only enough chemical needed for the experiment should be removed – any excess must be discarded. Care should be taken that bottle stoppers do not become contaminated. Many chemicals absorb moisture out of the atmosphere; therefore, bottle caps and stoppers should be promptly replaced on their containers. Such moisture (water) can render an expensive chemical useless.

In handling larger reagent bottles, one hand should be placed under the bottle while the other hand is placed on the bottle's neck, as shown in Figure 1. This procedure minimizes any possibility of the neck breaking off when the bottle is picked up. In addition, it keeps the bottle in front of the person, not swinging at the side where it may

Figure 1.
Chemical Carrying.

hit an object and break. Thick rubber bottle carriers are available for safe carrying of acids, alkalies, and solvents.

Health and safety require prompt and proper disposal of waste materials. Discarding many different kinds of chemicals into the same waste container can result in chemical reactions that generate unpleasant or toxic gases, violent reactions, or a boil-over mess.

Special ceramic crock containers should be used for solid-waste disposal. Solids should not be placed in sinks — they will stop them up.

Solvent wastes should be placed in a solvent-waste safety can. A specially designed safety can has a flame arrestor at the mouth of the can to smother any flame, thereby preventing the entire content of the can from burning. Normally, these flammable liquid wastes should not be washed down a sink drain (they could explode in the sewer system); however, small quantities of acids or alkalies may be washed down a sink drain with large amounts of flowing water.

Many chemical laboratories use glass traps in sinks instead of metal ones in order to prevent reactions of corrosive materials with the metal. Federal, state, and local laws, designed to prevent water and air pollution, should be considered when it is necessary to dispose of large quantities of chemicals.

PERSONAL PROTECTIVE EQUIPMENT

Ideally, accidents could be prevented by the proper use of safety precautions; however, even the most cautious person occasionally incurs an unexpected reaction or incident. Laboratory work, by its nature, is experimental, and results cannot always be predicted. "Personal protective equipment" describes any device used to eliminate (or reduce the seriousness of) an accident. Figure 2 illustrates protective clothing and safety devices that can be used in the laboratory.

Figure 2. Protective Clothing and Devices.

Eyes are valuable and irreplaceable; but, unfortunately, they can be injured easily by corrosives, hot chemicals, or flying objects, such as glass from an explosion. Therefore, safety glasses must be worn at all times in the laboratory. Some states have laws requiring that safety glasses be worn in laboratories, and most industrial labs require that they be worn, with violators of safety rules often being subject to dismissal. Contact lenses may be worn if safety glasses are also worn over them.

As previously stated, industry demands that safety precautions be observed. In a laboratory, prescription safety glasses should be worn by anyone who normally wears glasses. However, safety glasses offer minimum eye protection; therefore, in instances where hazards are obvious, additional protection, such as safety goggles or a face shield, should be used.

In instances where explosives are being used, or an explosion is likely, a shatterproof safety shield should be used. Many reactions can be conducted safely while standing behind such a shield since the shield offers protection from an explosion. The apparatus in use can be clearly seen and manipulated – by reaching around the sides of the shield – with minimum personal exposure.

If a chemical gets in an eye, especially an acid, the eye should be washed immediately in an eyewash fountain. This wash should be continued for 15 minutes. Should chemicals touch a large portion of one's body, a safety shower should be used. (The ring in the safety shower is quickly pulled to release large amounts of water.) If the action to wash off the chemicals with water is fast enough, chemical burns can be minimized. Eyes should be flushed with water for at least 15 minutes, then prompt medical attention should be sought. If clothing should catch fire, or if a corrosive chemical is spilled on the skin, every second counts in preventing severe burns. Because one may be confused at such a time, it is important that the exact locations of safety showers and eyewash fountains be noted before work in any laboratory area is begun. As noted previously, one should never work along in a laboratory since a worker may need assistance in an emergency. For example, with a corrosive burning the eyes, it would be difficult to locate the eyewash fountain without help.

Laboratory coats and aprons offer some degree of protection from chemicals for one's clothes and body. Such coats and aprons are available in a wide variety of materials and styles, including resin-coated cloth, rubberized cloth, cotton duck, and vinyl plastic. If protective clothing or regular clothing becomes soaked with a toxic or corrosive material, the clothes should be removed immediately to prevent further exposure to the chemicals. A quick shower in the emergency shower should then be taken.

A variety of gloves are available to protect hands from burns, caustics, and toxic contamination. Asbestos gloves or mittens are used to protect hands when handling hot objects. Latex surgeon-type gloves, which are tight-fitting and flexible, provide protection against chemicals. Neoprene rubber and natural rubber gloves, which resist acids and most chemicals, are widely used in chemical operations.

For work in a pilot plant or production plant location, a "hard hat" (helmet) and hard toe safety shoes are generally required. These items provide protection from falling objects for one's head and feet.

A fume hood (Figure 3) provides a work area from which toxic, flammable, and unpleasant vapors and dusts can be exhausted. Such a hood usually includes a sink, hot and cold water spigots, and gas and electrical services – features which allow the performance of most types of experiments.

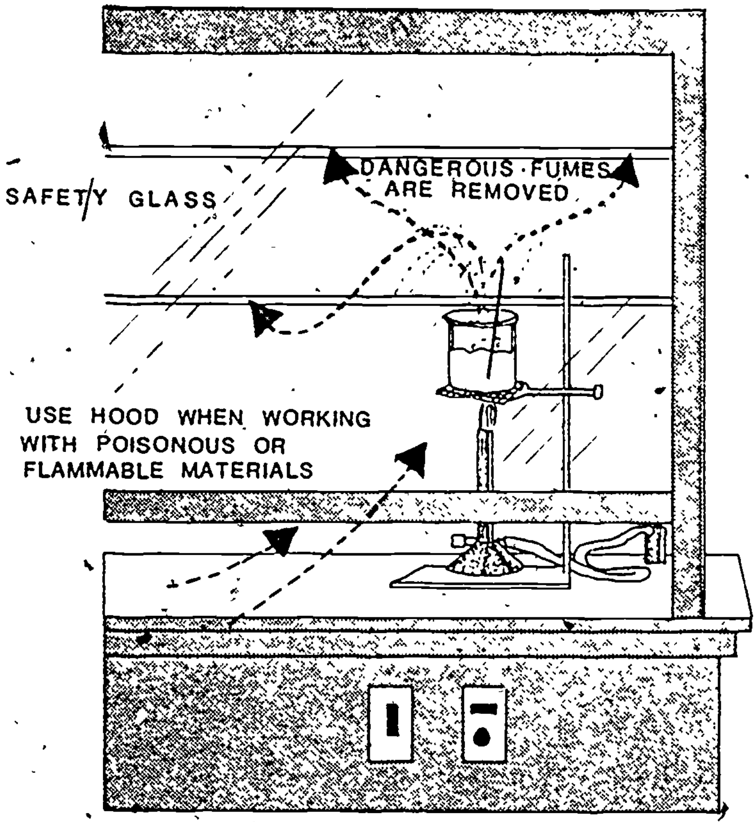


Figure 3. Fume Hood.

Hoods should not be used for the storage of chemicals; vented cabinets are available for such storage.

Prior to using a fume hood, one should be sure the air exhaust system is operating properly. The hood's operation can be checked by turning the exhaust fan on and closing the front shield to within an inch of the bench top. If the hood is operating properly, a strong air flow can be felt.

Some laboratories have exhaust fans for ventilating large work areas. These fans are not substitutes for hoods; the air exchange is inadequate for most operations. If such fans are used, they should not work in opposition to the fume hood(s) by drawing fumes out of the hood(s) into the open work area.

In some instances, fume hoods or exhaust fans do not provide adequate protection. In these instances, respirators having filters and chemicals (in canisters) that filter out vapors and fumes can be used. Care must be exercised to ensure that the proper respirator is chosen for the type of hazard encountered. Most are designed for a specific type of vapor, but some are useful only for dusts and mists. In the event of fire, oxygen content of the air may be too low, in which case a canister-type respirator will be of no use. A self-contained breathing apparatus, with its own air or oxygen supply, should be used for fire fighting or for rescue from fire and smoke-filled areas.

FIRE SAFETY AND EXPLOSIONS

In considering fire safety, one's first concern should be for prevention. Proper storage of chemicals, adequate precautions, and a knowledge of the properties of chemicals being used in chemical operations can prevent fires from starting. However, should a fire start, steps must be taken to prevent injury, loss of life, and property damage.

FIRE PREVENTION

Three factors contribute to combustion:

- Fuel
- Oxygen
- Ignition (such as from the heat of a spark)

Of these three factors, the only one that can be eliminated from a laboratory is ignition.

Combustion (fire) is defined as "rapid oxidation of a substance which produces light and heat." This definition enables one to distinguish between fire and slow oxidation, such as rusting of iron.

If a source of ignition, such as a flame, is brought close to the surface of a flammable liquid in an open container, one of three things may occur:

1. Nothing.
2. A temporary flash, but no continuation of burning.
3. Ignition and continued burning.

Three characteristic actions are defined below:

- Flash point: The lowest temperature at which a substance or mixture in an open vessel gives off enough combustible vapors to produce a momentary flash of fire when a small flame is passed near its surface.
- Ignition temperature: The lowest temperature at which the vapor over a liquid will ignite and continue to burn if an ignition source is applied near the surface.
- Auto-ignition temperature: The lowest temperature at which a vapor will ignite spontaneously without any outside source of ignition.

In the latter instance, heat within the liquid is sufficient to cause it to burn.

The flash points and auto-ignition temperatures of some chemicals are given in Table 1.

TABLE 1. TYPICAL FLASH POINTS AND AUTO-IGNITION TEMPERATURES.

Chemical	Flash Point (°C)	Auto-Ignition Temp. (°C)
Octane	56	428
Benzene	12	1040
Acetone	0	869
Hexane	-7	437
Ethyl ether	-49	320

A lower flash point means a greater potential fire hazard. Knowledge of the exact flash points and auto-ignition temperatures of all common chemicals is not required. These temperatures, however, can be used as a guide for determining whether a high, medium, or low hazard exists. It is best to assume that once the flash point of a chemical is reached, it may catch fire and continue to burn.

A technician working with chemicals should know that an open container of ethyl will readily ignite at room temperature. If the vapors of ethyl ether are allowed to escape into a laboratory, they may travel great distances along the floor to an ignition source, where they may catch fire, explode, or flash back to the source of the vapor. As can be seen from Table 1, octane (a major component of gasoline) is relatively safer than ethyl ether. Portable vapor-detecting meters, or "sniffers," may be used to detect such flammable vapors.

As previously stated, most laboratory chemicals are stored in glass bottles to prevent contamination; however, since bottles break easily, the smallest bottle that is consistent with the amount of a chemical to be used should be utilized. Larger quantities of flammable liquids are usually purchased in 5-gallon metal cans, although the purity of these liquids is not as great as those of liquids in glass bottles.

Fires can be prevented if flammable vapors are kept contained or properly exhausted to the atmosphere. All ignition sources must be kept completely away from flammable substances. Ignition sources include open flames, hot wires, electric sparks from switches or motors, and static electricity. Should a fire occur, one needs to know the location of fire extinguishers and to be able to use them properly. Three methods of extinguishing a fire are shown in Figure 4.

Some materials, called pyrophorics, heat and ignite spontaneously in the presence of air. These materials are identified by their container labels.

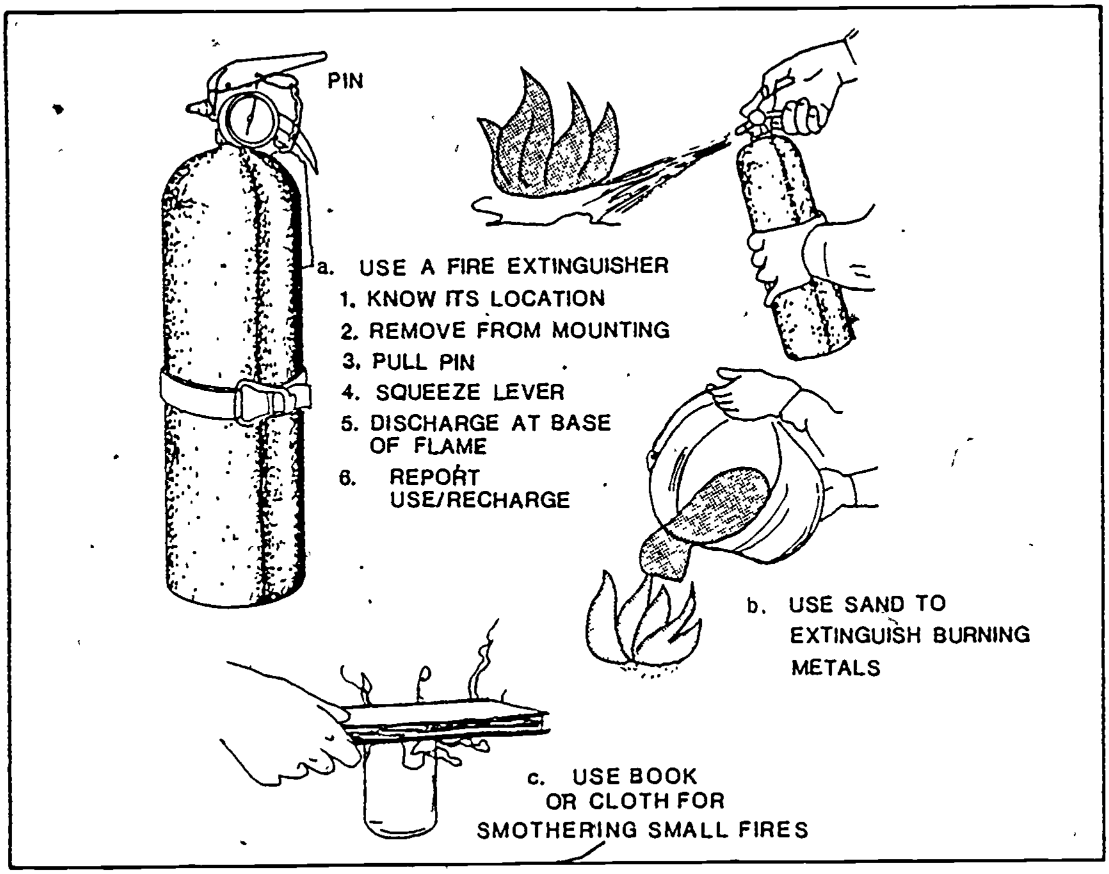


Figure 4. Methods of Extinguishing a Fire.

FIRE FIGHTING

Equipment needed for general fire fighting – automatic sprinkling systems, fire hoses, fire extinguishers, fire blankets, and so forth – should be installed in any building that houses a chemical laboratory. Specialized fire fighting equipment may also be required. The leading causes of laboratory fires include the following:

- Defective electrical wiring
- Sparks from non-explosive-proof electrical apparatus

- Fires in fume hoods
- Carelessness in the use of gas burners
- Ill-advised use of volatile liquids

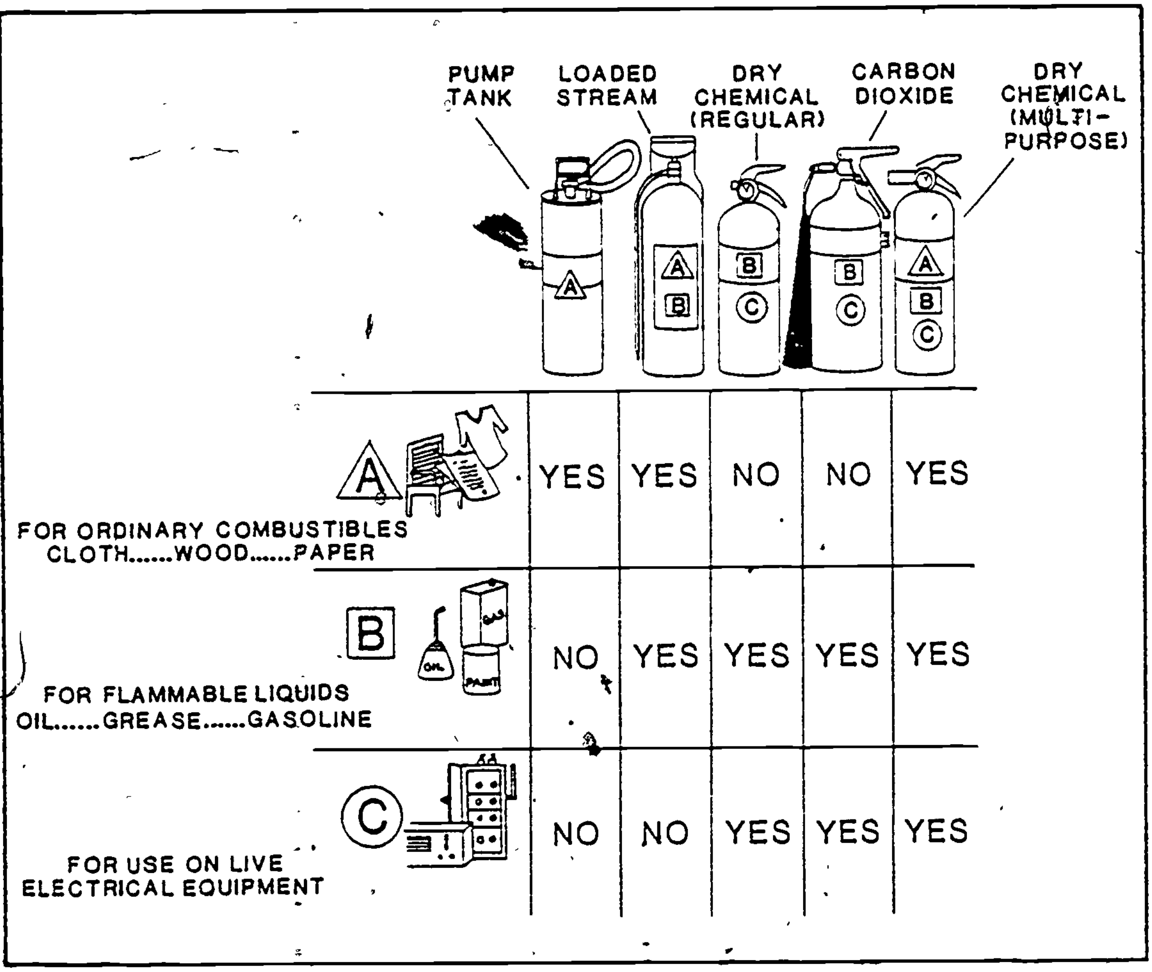
Fire in clothing should always be extinguished by smothering it in a fire blanket or heavy toweling or by beating the flame out. Do not try to douse it out with water. Dousing with water should only be used if other means are not available.

When fighting a fire with a fire extinguisher or water hose, the area immediately surrounding the fire should be cooled first in order to prevent the flames from spreading. Next, the base of the blaze should be extinguished. Finally, the remainder of the fire should be smothered. Trying to extinguish the flames of a fire is practically useless; rather, it is better to attack the base of the fire. Figure 4 shows how to use a fire extinguisher, how to extinguish burning metals, and how to smother small fires.

To extinguish an electrical fire, pull the circuit breaker to stop the electrical current, then use a carbon dioxide (CO_2) extinguisher – not water.

If a poisonous gas is emitted by a fire, and a proper mask is not immediately available, it is best to flee the room entirely. It is important to note, too, that when a fire gets out of control, it is futile to attempt to fight the fire with extinguishers that have limited capacity. If the fire cannot be controlled, then a check should be made to see that (1) no one if left inside the room, (2) the door is shut to prevent drafts, and (3) the fire department is called.

Fires are classified according to the type of material being consumed. Use of the proper type of extinguisher for each class of fire will afford the best control of the situation and will avoid compounding the problem (that is, a Class A fire extinguisher should be used on a Class A fire, and so forth). Figure 5 shows different types of fire extinguishers.



	PUMP TANK	LOADED STREAM	DRY CHEMICAL (REGULAR)	CARBON DIOXIDE	DRY CHEMICAL (MULTI-PURPOSE)
A FOR ORDINARY COMBUSTIBLES CLOTH......WOOD......PAPER	YES	YES	NO	NO	YES
B FOR FLAMMABLE LIQUIDS OIL......GREASE......GASOLINE	NO	YES	YES	YES	YES
C FOR USE ON LIVE ELECTRICAL EQUIPMENT	NO	NO	YES	YES	YES

Figure 5. Types of Fire Extinguishers.

- Class A Fires: For wood, paper, textiles, and similar materials, use foam, water, or almost any type of extinguisher.

- Class B Fires: For grease, oil, paint, or related materials, use foam, dry chemical, or vaporizing liquid extinguishers.
- Class C Fires: For fires in electrical equipment or in areas where live electricity is present, use carbon dioxide, dry chemical, or vaporizing liquid extinguishers.
- Class D Fires: For fires involving magnesium, aluminum, zinc, zirconium, lithium, sodium, potassium, or metal hydrides, smother with fine dry soda ash, sodium chloride (salt), sand, or graphite (Figure 4b).

Underwriters' Laboratories (UL), Inc., has established ratings for portable fire extinguishers. These ratings appear on the labels of UL-approved extinguishers. A letter on the label refers to the class of fire for which the extinguisher is most effective. A number-letter symbol indicates the relative fire-extinguishing potential of the extinguisher, as well as the approximate square-foot area an average operator can extinguish. For example, an extinguisher rated 10-B is used on grease, oil, paint, and related materials; and it can extinguish a 10-square-foot area. Moreover, a 5-C extinguisher is used for electrical fires (carbon dioxide); and it can extinguish approximately a 5-square-foot area.

CHEMICAL TOXICITY

Many chemicals are toxic or poisonous; therefore, precautions should be taken in working with these chemicals. For instance, various cleaners, paints, and solvents used in the home are toxic; and products used on lawns and gardens, including fungicides, insecticides, fumigants, and herbicides, are often toxic. Such products should be used with great care.

In the laboratory, even greater care is required because many of these chemicals are used in concentrated form.

Poisoning may be "acute" (a single or sudden exposure) or it may be "chronic" (a repeated or prolonged exposure). For example, sudden exposure to chlorine gas may cause acute poisoning, and a gradual exposure to mercury or lead in the environment may cause chronic poisoning. Often the symptoms of chronic poisoning are difficult to detect. Both types of poisoning should be of concern to the laboratory worker.

The three most common ways by which toxic chemicals enter the body are inhalation (breathing), ingestion (swallowing), and absorption through the skin. Of these three, the most common is absorption; but the entry method causing the highest percentage of fatalities is inhalation. Table 2 shows the disabling work injuries from poisoning that occurred in California in one year.

TABLE 2. DISABLING WORK INJURIES.

Contact Method	Total	Fatal	Non-Fatal
Inhalation	1203	23	1180
Absorption	4123	3	4120
Ingestion	145	2	143

PHYSIOLOGICAL EFFECTS OF CHEMICALS

The extent of a human body's reaction upon exposure to toxic chemicals is determined by many factors, the most important of which include the following:

- The concentration of toxic material in the blood system, or in a specific, sensitive organ.
- The body's tolerance for a particular toxic chemical (tolerance varies widely from one individual to another).
- The rate at which the toxic material can be changed or eliminated by the body before damage is done.

The level of hazard depends upon (1) the concentration of the chemical in the environment, (2) the length of exposure, (3) the general health of the individual exposed, (4) the individual's tolerance, and (5) the speed with which curative action is taken after the exposure is recognized. No matter how poison enters the body, responses that may develop include the following:

- A change in blood composition
- A change in blood pressure
- A change in blood circulation rate
- A rapid cell destruction rate
- A change in the breathing rate

These changes may alter the functioning of organs such as the heart, kidneys, liver, brain, or lungs.

Boundary limits have been set for toxic materials to prevent atmospheric concentrations that can cause bodily harm. These limits, called threshold limit values (TLV), have been established for toxic gases, vapors, mists, dusts, and fumes. They also may be called maximum allowable concentrations (MAC). Gases and vapors usually have their TLV expressed in parts per million (ppm). The TLV of a toxic dust or mist usually is expressed in milligrams per cubic meter of air.

Some vapors or gases, such as ammonia, mercaptains (skunk-like odor), or tear gases, give warning of their presence by their odor or by an immediate irritating effect on the eyes, nose, or throat. Some, such as carbon monoxide, have no color, odor, or other warning effects. Still others, such as hydrogen sulfide, initially have an easily detected warning odor; however, this odor soon becomes unnoticeable, and the exposed person is unaware of the continued presence of the chemical, even in deadly concentrations.

Precautions and action must be taken at the first indication of excessive exposure to poisonous vapors. One should not wait until physical symptoms occur since recovery is often impossible at this stage. Frequently, the first symptoms to appear are blurred vision and malfunction of muscles. In many cases, a combination of exposures is much worse than single-chemical exposures. For example, exposure to carbon tetrachloride during a day's work at the lab, followed by an alcoholic drink at home, can lead to kidney failure. Of course, an excessive exposure to carbon tetrachloride alone can also lead to kidney failure, but somehow the combination is more severe.

Skin contact is the most frequent method of exposure to most toxic chemicals, but it does not cause as many fatalities or major poisoning as does inhalation. Some chemicals, such as organic solvents, penetrate the skin rapidly, are rapidly absorbed by the blood, and are then transported throughout the body. Contact with these chemicals should be avoided.

The common practice of washing oil or grease from the hands with solvents or gasoline ~~also~~ should be avoided. Some skin contact, over a period of time, can cause redness, roughness, cracking, and other irritations leading to skin dermatitis. Contact with strong acids and bases can cause skin and flesh destruction. Skin contact damage may be seriously increased by continued contact with clothing that has been wetted by the chemical.

Wet clothing prolongs the skin's exposure to a high concentration; therefore, contaminated clothing should be removed, and the affected area should be flushed with water immediately — regardless of considerations of modesty.

The treatment of skin that has been in contact with a poison or an irritating substance consists of (1) quickly washing the affected area with large amounts of water for at least 15 minutes, then (2) gently cleansing the affected skin area with soap and water. If the material cannot be removed in this manner, a medicinal oil (for example, paraffin oil) usually will safely remove the contaminant.

Poisoning by swallowing is rather rare for laboratory workers because they are generally well aware of the potential hazards. The practice of using the mouth to provide the suction for pipetting must be avoided. A mouthful of concentrated sulfuric acid can cause painful thermal and chemical burns. Various rubber bulbs for pipetting are available and should be used. Licking labels, moistening one's finger to flip pages in a notebook, and smoking, eating, or drinking in the laboratory are all possible sources of poisoning. Chemicals should never be tasted for identification or any other reason. The practice of storing food and chemicals together in a refrigerator should be avoided. Labware should never be used for making coffee, cooking soup, or as eating or drinking utensils. Hands should be thoroughly washed before eating, drinking, or smoking. If a toxic substance is swallowed, it must be removed or made harmless as quickly as possible. Try to determine the source of poisoning because this information is important for treatment of the victim.

DANGERS OF ACIDS

A knowledge of chemical and physical properties of some of the more common chemicals, as well as hazards involved, will enhance safety in the laboratory.

Concentrated acids commonly found in chemical operations include the following:

- Hydrochloric acid
- Nitric acid
- Sulfuric acid
- Hydrofluoric acid
- Perchloric acid

Any of the preceding acids can cause severe burns; therefore, eye protectors should be worn when using them.

In mixing water with acids, it is important that the acid always be added slowly to the water. Adding water to acid is dangerous because the heat that develops at the point of addition may cause the acid to boil and splatter out of the container.

Hydrochloric acid, commonly called muriatic acid, is widely used in industrial applications, in laboratories to dissolve metals, and in acid-base titrations. It is a strong acid and can cause severe damage through chemical burns.

Nitric acid may not cause a burning sensation immediately upon contact, thus the urgency to wash it off may not seem great; however, it must be washed off immediately upon contact because damage does start immediately. Splashes of this acid on the skin will turn the skin yellow in a few minutes. After a few days, the skin will peel off. Nitric acid is used to dissolve most metals that cannot be dissolved by hydrochloric acid.

Sulfuric acid, which incorporates all the hazards of hydrochloric and nitric acids, also harbors additional hazards. The density of sulfuric acid is high; therefore, a bottle of this acid is surprisingly heavy. A given volume of sulfuric acid weighs nearly twice as much as an equal volume of water. (Many serious accidents have occurred when laboratory workers have tried to pick up a bottle and then let it drop.) This acid is a powerful dehydrating agent; it reacts with water or moisture on the skin and turns it into steam, thereby causing thermal burns in addition to chemical burns. More serious accidents are caused by this acid than by any other — because of its properties and because it is used frequently in chemical operations.

Hydrofluoric acid dissolves silicon oxide, a major component of glass; therefore, hydrofluoric acid cannot be stored in glass bottles because it dissolves the glass. Plastic or rubber gloves must always be used when handling hydrofluoric acid. Burns from this acid are very painful and may last for several weeks. The burns continue until the acid is neutralized by calcium in the bones. (A physician can inject a calcium compound solution to neutralize the acid.)

Perchloric acid, when cold and dilute, is similar to any other strong acid; however, when this acid is hot and concentrated, it is a very powerful oxidizing agent. When perchloric acid is in contact with wood, tile, and other organic materials, it can cause explosions through rapid reaction and subsequent release of gases. A specially designed fume hood must be used with perchloric acid. Since air accumulation of its acid mist on dirt particles in the hood can cause a severe explosion, the special hood provides for a flow of water to continuously wash down the perchloric fumes from the sides of the hood.

Spills of acids can be neutralized with sodium bicarbonate (baking soda). A paste of sodium bicarbonate can be applied to the skin to neutralize acid on the skin. The most immediate treatment, however, is to immediately flush the affected area with water. Greasy ointments should not be applied to chemical burns.

A strong base, such as sodium hydroxide (lye) or potassium hydroxide, is usually dissolved in water and used as a solution. Such a strong alkaline material can cause burns in either its solid or liquid form. An immediate, prolonged water wash is the most effective treatment. Boric acid, which is very weak, can be used to treat burns caused by bases.

VAPOR POISONING

Carbon monoxide, which is very poisonous, can cause death if inhaled for just a few minutes. Carbon monoxide gives no warning by smell or taste. A victim's first symptoms are headache, dizziness, and general weakness — by which time irreversible damage may have occurred. At an atmospheric concentration of 0.05%, a human can lose consciousness in a matter of minutes. Carbon monoxide, which causes more accidents than all other toxic gases combined, is produced whenever incomplete combustion of carbon-containing substances takes place. It is produced by internal combustion engines and luminous (yellow) flames.

Poisoning from mercury vapor also occurs. Mercury is widely used to measure pressures, and the glass instruments in which it is normally contained are too easily broken. Mercury spills are virtually impossible to clean up. Powdered sulfur should be spread over the mercury, and then the mercury may be swept up and disposed of. Even so, small balls of mercury roll into inaccessible cracks and crevices, making it virtually impossible to remove them all. The mercury can then give off vapor continuously for years. Extra precautions are in order because mercury has a cumulative effect. The sum of many apparently insignificant exposures can lead to serious poisoning. The symptoms of mercury poisoning are emotional and physical instability.

Hydrogen sulfide has a TLV of only 10 ppm, but it is 10 times more poisonous than carbon monoxide. It has a detectable and characteristic odor, but the sense of smell is quickly dulled, leading the victim to believe that the hazard no longer exists. Exposure to hydrogen sulfide causes headache, nausea, dizziness, shallow breathing, and lowered blood pressure.

Carbon tetrachloride has been widely used in the past as a spot cleaner for clothing. Its use has been banned in most household products. Because of its excellent solvent powers, however, it finds wide use in chemical laboratories. Symptoms of carbon tetrachloride poisoning appear after the damage has been done. Commonly, the victim of "carbon tet" poisoning has kidney failure, often resulting in death. Only three thimblefuls of carbon tet can saturate the air to the danger point in any ordinary size, unventilated laboratory.

Benzene is another popular solvent. Its odor is faint and not disagreeable. Symptoms of low exposure include headache, weakness, and bleeding at the nose and mouth. Brief inhalation of high concentrations may result in unconsciousness and death.

MERCK INDEX

The Merck Index is a valuable resource for determining the possible hazards associated with chemicals. This encyclopedia of chemicals and drugs provides concise descriptions of the preparation and general properties of compounds and gives their structure and trade name, as well as their toxic properties. The Merck Index should be consulted prior to using any unfamiliar chemical. For example, prior to working with perchloric acid, it would be important to know the following information found in the Merck Index:

> Perchloric acid decomposes with explosive violence when distilled. Aqueous solutions of the acid are caustic and will explode when they contact organic substances. Perchloric acid is used in the preparation of explosives, as solvent for inorganic substances, and in plating metals.

Referencing the properties of chemicals or consulting with co-workers regarding potential hazards should be a standard procedure for everyone.

ELECTRICAL HAZARDS

Electrical appliances and apparatuses are so common that one often fails to respect electricity and its hazards. High voltage is often thought to be the greatest danger from electricity; however, it is current (and how long it flows) that is hazardous – not voltage. The normal 110-volt household current can readily cause death.

Safety fuses or circuit breakers are used in most circuits to interrupt the flow of electricity if the current exceeds a certain safe maximum – usually 15 amperes. But safe maximum means "a safe current for the equipment and wiring system." (A 15-ampere shock to a human generally is instantly fatal.) The chance of receiving severe electrical shock increases with a greater flow of electrical current through the body. Therefore, to minimize the potential, the hands should always be clean and dry when using an electrical apparatus.

The following data illustrates the effect on resistance to electrical current caused by water and salt:

	Condition	Resistance (ohms)
1.	One dry finger of each hand on an electrode.	100,000
2.	One moist finger from each hand on an electrode.	40,000
3.	Same as 2, except fingers wet with salt solution.	16,000
4.	Tight grip, dry hands, one on each electrode.	1,200
5.	Same as 4, except immersed in a salt solution.	700

It should be obvious from the preceding experimental data that the hands should be kept dry and free of chemicals when any possibility exists of contact with an electrical current.

To prevent electrical shock, follow the rules listed below:

- Be sure equipment is in proper working condition.
- Be sure wires are not frayed, loose, or broken.
- Be sure equipment power switches are in OFF position before plugging unit into an outlet; this prevents "sparking" at the plug.
- Be sure hands are dry and clean before handling any electrical device.
- Be sure provisions have been made for proper grounding.
- Be sure all electrical work is done by a qualified electrician.
- Be sure electrical equipment is inspected periodically.
- If an extension cord is used, be sure that it is heavy enough to carry the power safely.
- Be sure not to overload an electrical circuit.

SPECIAL HAZARDS

Compressed gases are widely used in laboratory and industrial applications. These gases are stored in heavy metal "cylinders" of various sizes. Gas cylinders are filled to very high pressures (up to 3000 psi). Each cylinder is color-coded to identify its contents. Gas is released from the cylinder by a high-pressure valve. The cylinder must be handled with great care (Figure 6), even if the gas is as harmless as helium or air. When transporting a cylinder, one should be sure that a protective dome is properly fitted over the valve. The cylinder should be strapped to a wheeled cart for transporting, and it should be chained or strapped to a work bench or wall for use. This prevents it from

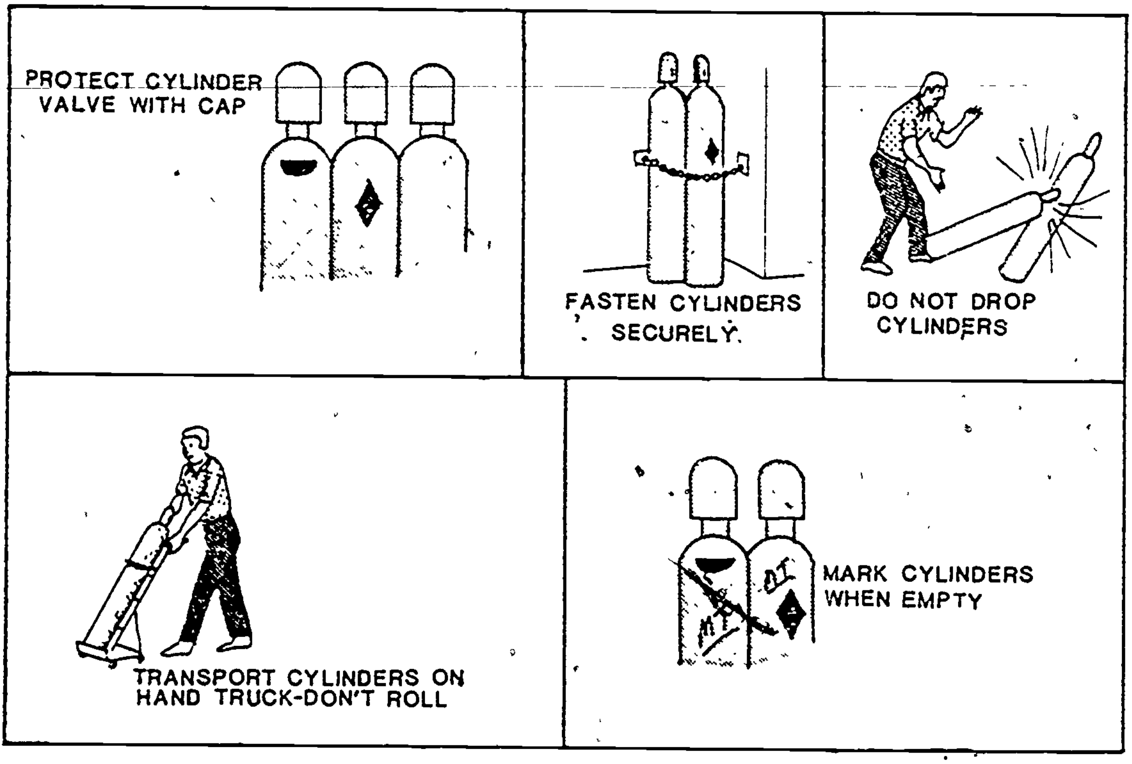


Figure 6. Gas Cylinders.

falling over and possibly knocking the valve off, which could cause the cylinder to become a deadly "jet" missile.

A gas regulator is used to control the flow of gas from a cylinder. A regulator that has been used to regulate some other gas should not be used as a regulator on an oxygen cylinder. The high-pressure oxygen will react with any oil in the regulator, which could lead to a serious explosion or fire. (The oil comes from the manufacturing process, during which the gas is "oil pumped.")

Working with materials at very low temperatures poses special hazards. Cryogenics are gases that have been liquefied. Spills of these liquefied gases onto human tissue can cause severe "burns" similar to those caused by hot objects. Another hazard arises when liquid gases are warmed, or when they come into contact with a warm object; an enormous increase in volume and consequent generation of high pressure occurs.

Many fires and accidents are caused by unattended operations. Experiments should not be left unattended – certainly not overnight. Heating units may fail, the flow of cooling water may be interrupted, equipment may break, and other problems may develop which could be corrected by the operator if he or she were present.

Conventional refrigerators have inside electrical connections and contacts which could spark and set off an explosion if used for storage of chemicals such as ethyl ether. Laboratory refrigerators are specially designed to eliminate this explosion hazard; but these refrigerators should never be used for food storage, and the ice cubes should never be used to cool beverages for consumption.

A vacuum can cause an implosion (the collapse of a vacuum desiccator, flask, or vacuum distillation apparatus). Such equipment should be wrapped with a special adhesive tape, which may not prevent collapse but can minimize the amount of flying glass caused by an implosion. In addition, wire-mesh cages may be used to provide protective shielding.

The buildup of excessive pressure in a glass apparatus can result in a serious explosion. Equipment must be carefully set up to allow gases to escape or to be trapped without the buildup of such excessive pressures.

Peroxides are a special source of hazard in ethers and ether storage. Many explosions have been caused by ethers that have been stored too long. (Over a period of weeks or months, ethers form peroxides, especially in the presence of air and in glass containers. These peroxides decompose upon being heated and, when the decomposition reaction starts, an explosion results.) Sometimes an ether container, in which peroxides have formed, will explode simply from being jarred or struck, even at room temperature.

Ethyl ether, which is widely used in laboratories as an extraction solvent, should be tested for peroxides prior to use. Ethers should be labeled with the date they are received, ~~and~~ unused ether should be discarded periodically.

RADIATION HAZARDS

An increasing number of modern laboratories use radioactive isotopes, or equipment (for example, an X-ray diffraction instrument) that produces radiation; therefore, the technician must be aware of some of the precautions to be observed when working with or around radiation.

Radiation generally falls into one of the following three categories:

1. Alpha particles are streams of heavy, positively-charged particles. These particles travel about an inch in air and have very little penetrating effect. A piece of paper or the human skin stops them.
2. Beta particles are streams of lighter, negatively-charged particles. These particles have a range of up to 30 feet in air and can penetrate the skin. However, almost all such particles are stopped by a one-eighth inch aluminum sheet.
3. Gamma and X rays are electromagnetic waves with high frequencies and great penetrating power. Lead and concrete are often used to shield these rays.

Radiation produces biological injury by damaging cells through their ionizing properties. In a laboratory, contamination may be encountered on tools, glassware, working surfaces, clothing, hands, in wastes, and in the air.

Special instruments must be used to detect radiation. Radiation safety primarily involves techniques mentioned previously regarding toxic chemicals. Hoods should be used to prevent the ingestion of radioactive dust. Protective gloves and shoe covers are necessary. Radioactive materials must be clearly marked. Smoking, eating, and drinking in the laboratory are strictly forbidden. All equipment and personnel must be monitored and decontaminated upon leaving a restricted work area. Proper waste disposal is extremely important.

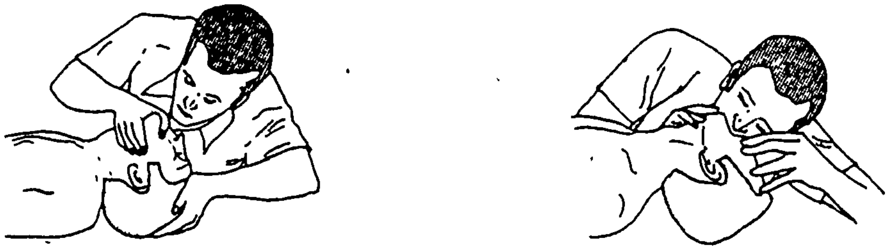
Because the cumulative effect of many small radiation exposures may cause serious health problems, it is desirable to maintain complete records of personal exposure (as shown by pocket monitoring devices). Yearly, or more frequent, medical examinations for all exposed personnel also are recommended.

LABORATORY FIRST AID

First aid training is a must for a laboratory worker. This training enables the worker to render life-saving aid should it be required. Most larger companies have in-house safety programs that include first aid training. It should be remembered that first aid is the immediate assistance given to the victim of an accident before professional treatment by a doctor can be arranged. Do not attempt to give full treatment, just emergency relief. Stop bleeding, prevent shock, then treat the wound - in that order. The primary rule is to keep calm; proper aid cannot be given by anyone who is not in control of himself or herself.

The following is a list of the basic principles of first aid:

1. Call an ambulance, and state the type of accident, its location, and the approximate number of people injured.
2. Keep calm and keep crowds away; give the victim fresh air. Do nothing else unless the proper procedure is known.
3. Stop any bleeding.
4. Prevent physical shock.
5. Restore breathing by artificial respiration if required. The correct technique for artificial respiration is shown in Figure 7.



1. Lift victim's neck with one hand and tilt his or her head back by holding the top of his or her head with the other hand.
2. Pull the victim's chin up with the hand that was lifting the neck so that the tongue does not fall back to block the air passage.
3. Take a deep breath and place mouth over the victim's mouth, making a leak-proof seal, and at the same time pinch the victim's nose closed.
4. Blow breath into the victim's mouth or nose until the chest rises.
5. Remove mouth and let the victim exhale while taking another deep breath.
6. As soon as the victim breathes out, replace mouth over his or her mouth or nose and repeat procedure 15 times per minute.

Figure 7. Artificial Respiration.

6. Remove the victim(s) from hazardous environments (spills of chemicals or high concentrations of gases, vapors, or fumes). Wear proper respiratory equipment and protective clothing to avoid exposing oneself to hazards.
7. Never give liquids to an unconscious person.
8. Do not move a person with possible broken bones or possible head or internal injuries unless necessitated by fire or fumes.

A slight thermal burn can be effectively treated by immersing it in cold water to relieve pain. Do not use ointments on any burns; ointments cover a wound and hold in the heat. Treat the victim for shock.

CHEMICAL BURNS

Treat chemical burns as follows:

1. Remove the contaminated clothing and the source of the contamination. Care should be taken not to contaminate oneself.
2. Flush the contaminated skin area with large quantities of water for at least 15 minutes.
3. Do not use oils, fats, salves, or ointments; soap may be used. The victim should be taken immediately to a hospital for further treatment.

WOUNDS AND FRACTURES

Treat wounds and fractures as follows:

1. Bleeding must be stopped before other aid can be given. Apply a large compress over the wound, with direct pressure on the wound. If the wound is on an extremity, pressure can be applied to one of the pressure points.
2. If the wound is slight and bleeding is not profuse, remove all foreign material projecting from the wound. This removal may best be accomplished by careful washing with soap and water.
3. All wounds should be securely, but not tightly, bandaged.
4. In the case of a puncture wound (from a broken thermometer, glass tubing, etc.), the victim should be taken to a hospital. Puncture wounds are difficult to treat.
5. In the case of possible bone fracture, do not move the victim unless there is additional danger present (such as fumes or fire).

6. Treat for bleeding and shock, but leave splinting to a professional. When it is necessary to transport the victim to treatment, improvise a splint support to prevent further injury.

SHOCK

Shock, which occurs to some extent in all injuries, can cause death. Symptoms of shock include paleness, cold and moist skin, nausea, shallow breathing, and trembling. Place the victim in a reclining position, with the head lower than the body. Control any bleeding. Wrap the victim with blankets. If there is no bleeding, rub the victim's extremities briskly to restore circulation. Reassure the victim and remain calm.

Treat electrical shock as follows:

1. Shut off the current, or cautiously remove the electrical contact from the victim (use an insulator such as a rope or stick).
2. To restore breathing, start artificial respiration immediately.
3. The victim may be rigid or stiff; even so, artificial respiration should be continued until death is certified by a physician.
4. Keep the victim warm, using blankets or hot water bottles against the victim's body.

POISONS (SWALLOWED)

Treat swallowed poisons as follows:

1. If the victim is conscious, give two to four glasses of water immediately.
2. Call an ambulance immediately.
3. Induce vomiting, except when the poison is a strong acid, strong base, cyanide, gasoline, kerosene, or other hydrocarbon.
4. When the poison is a strong acid or base, give the victim egg whites, a glass of milk, or one teaspoon of dried albumin in a glass of cold water.
5. For poisons in general, give one heaping teaspoonful of Universal Antidote in half a glass of warm water. The ingredients are as follows:

Universal Antidote

2 parts activated charcoal
1 part magnesium oxide
1 part tannic acid
(Keep ingredients dry until used).

6. Prevent shock by keeping the victim warm.

POISONS (INHALED)

Treat inhaled poisons as follows:

1. Call ambulance.
2. Wear respiratory equipment and protective clothing.

Move the victim to fresh air immediately and give the victim oxygen if available.

3. Allow the victim to rest. At any sign that breathing has stopped, begin artificial respiration immediately.
4. Treat for shock by keeping the victim warm.

HEAT STROKE/HEAT EXHAUSTION

Neither heat stroke nor heat exhaustion is apt to occur in a laboratory, but it can occur in a plant or outdoors. The victim of heat stroke is very hot and dry; therefore, body temperature must be reduced immediately. Remove the victim from the source of heat and use ice and cold water to reduce temperature. A victim of heat exhaustion feels cool to the touch, and he or she should be treated as if in physical shock. Keep the victim warm and make sure the victim's head is lower than the body.

LABORATORY MATERIALS

Personal protective equipment:

- Safety glasses
- Face shield
- Explosion shield
- Fire extinguisher
- Eyewash fountain
- Safety shower

Major laboratory equipment:

- Vacuum pump
- Gas cylinder
- Gas regulator
- Hot plate
- Fume hood
- Waste disposal

LABORATORY PROCEDURES

1. Assign locker/laboratory equipment; discuss notebook and laboratory reporting procedures.
2. Demonstrate use of personal protective equipment and major laboratory equipment.
3. Sketch floor plan of laboratory; identify on drawing all major safety equipment, such as fume hood, eyewash fountain, safety shower, fire extinguishers, emergency exits, and places for solid- and liquid-waste disposal.
4. View 16 mm Safety film - "Safety in the Chemical Laboratory" (see Reference Section).

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FILM

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GLOSSARY

Acute: A single or sudden event.

Auto-ignition temperature: The lowest temperature at which a vapor will ignite spontaneously without any outside source of ignition.

Chronic: A repeated or prolonged event.

Combustion: The rapid oxidation of a substance which produces light and heat.

Cryogenics: Gases that have been liquified and therefore are very cold.

Flash point: The lowest temperature at which a substance in an open vessel gives off enough combustible vapors to produce a momentary flash of fire.

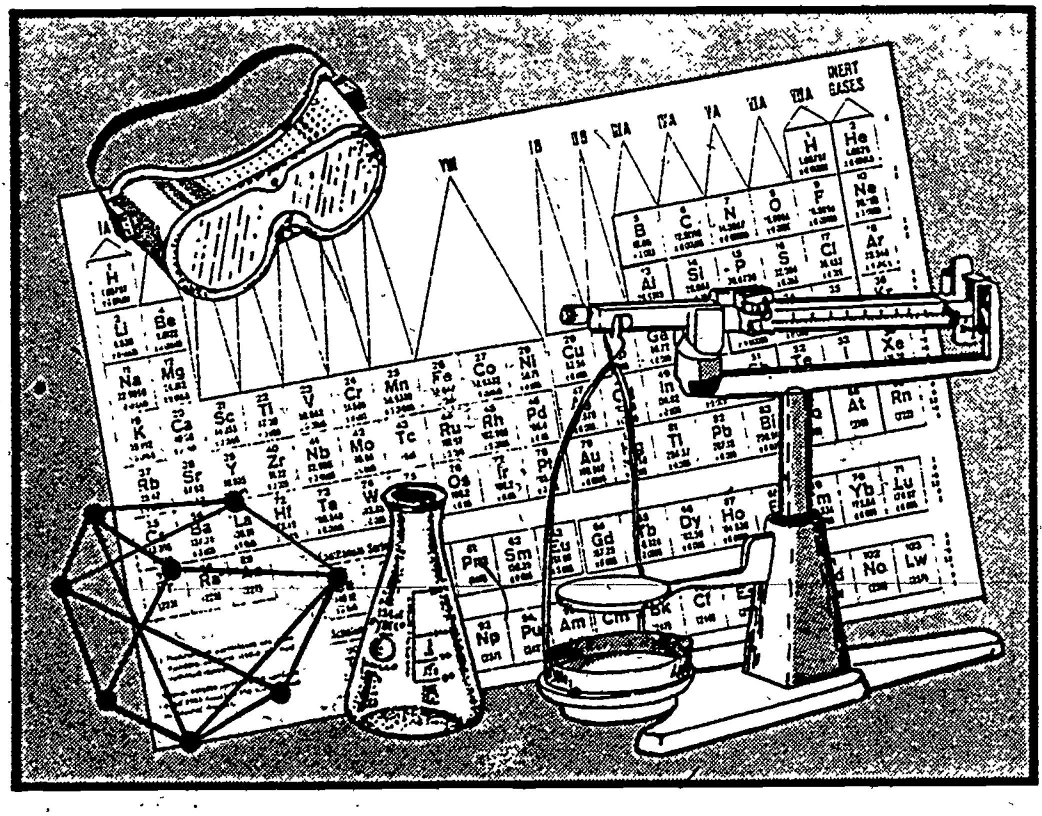
Ignition temperature: The lowest temperature at which the vapor over a liquid will ignite and continue to burn if an ignition source is applied near the surface.

ecut

ENERGY TECHNOLOGY

CONSERVATION AND USE

CHEMISTRY FOR ENERGY TECHNOLOGY I



MODULE CH-02

STRUCTURE OF MATTER

CORD CENTER FOR OCCUPATIONAL RESEARCH AND DEVELOPMENT

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INTRODUCTION

Chemistry is the branch of science that deals with composition of matter and changes that take place in matter. Many of these changes, or reactions of matter, consume or produce energy and, therefore, are important to the energy technician. Most matter is a complex mixture of elements and compounds, although all matter is composed of a relatively small number of fundamental particles or building blocks. These particles, electrons, protons, and neutrons combine in various ways to form atoms which, in turn, combine to form molecules. The forces which hold these molecules together are called chemical bonds.

Chemical bonding, elements, compounds, electrons, protons, neutrons, atoms, and molecules are all considered in this module to help explain the structure of matter. In addition, the student will learn accepted symbols and formulas – the "shorthand" of chemistry – that condense the volume of information, making it easier to learn and utilize.

PREREQUISITES

The student should have completed one year of high school algebra and Module CH-01 of Chemistry for Energy Technology I.

OBJECTIVES

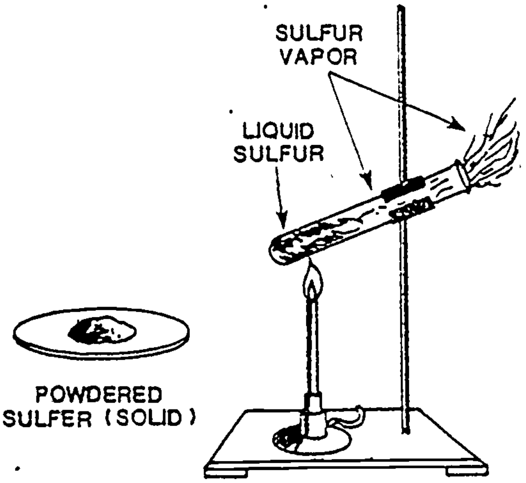
Upon completion of this module, the student should be able to:

1. Define the following terms:
    a. Element.
    b. Atom.
    c. Molecule.
    d. Homogeneous.
    e. Heterogeneous.
    f. Mixture.
    g. Compound.
    h. Ion.
    i. Ionic bond.
    j. Covalent bond.
    k. Hydrogen bond.
    l. Metallic bond.
2. Write basic chemical formulas.
3. Write the name of a compound, given its formula; or write the formula, given its name.
4. Match the name of an element with its symbol.
5. Identify changes as chemical or physical changes.
6. Write the electrical charge and relative weight of electrons, protons, and neutrons.
7. Draw a pictorial representation of simple atoms, giving the location of electrons, protons, and neutrons.
8. List characteristics that distinguish ionic substances from covalent substances.
9. Identify conductors from a list of compounds.

STATES OF MATTER

Matter, anything which occupies space and has weight, exists in three states: gas (or vapor), liquid, or solid. A gas has neither shape of its own nor a fixed volume. It takes the shape and volume of any container into which it is placed. It can be compressed readily to fit a small container, or it will expand to occupy a large one. Air, helium, and carbon dioxide are examples of some fairly common gases. A liquid has no specific shape; it assumes the shape of the container that it occupies. A liquid does not expand to fill the entire container; it has a specific volume. Liquids are only slightly compressible. Water, gasoline, and milk are common liquids. A solid has a firmness that is not associated with either gases or liquids; it has a fixed shape and volume. Like liquids, solids are only slightly compressible. The three states of matter are shown in Figure 1.

Figure 1. The Three States of Matter.



CHEMICAL AND PHYSICAL CHANGES

The state of a substance depends upon its temperature. For example, above 100°C, water exists as a gas; between 0°C and 100°C, it exists as a liquid; and below 0°C, it exists as a solid (ice). The changes of state, such as a change from solid to liquid or liquid to gas, are examples of physical changes. These changes often consume (or give off) large amounts of heat and, therefore, are of importance to an energy technician. For example, the absorption and release of heat during changes of state (also called phase changes) are the basic processes in air conditioning and refrigeration.

A physical change is one that does not involve creation of new substances — just a change in state or physical appearance. The following are examples of physical changes: melting of ice, changes in particle size (in grinding for example), cutting of wood, condensation of steam, and melting of iron. Chemical changes, also called chemical reactions, involve the conversion of one substance into another. Burning of paper, souring of milk, rusting of iron, and burning of gasoline are all examples of chemical changes. The product of these reactions is a different substance from the original material.

CLASSIFICATION OF MATTER

All matter can be classified either as pure substances or as mixtures of two or more substances. A pure substance is a material that is homogeneous (alike throughout) and has a distinct set of properties. Substances are of two types:

compound substances and elementary substances. These are usually referred to as "compounds" and "elements." Most substances are *compounds*; they are composed of two or more elements united chemically in definite proportions by weight. *Elements* are the simplest substances; they cannot be decomposed or changed into simpler substances. Figure 2 shows the classification of matter. *Mixtures* have a variable composition, and they usually can be separated into their components by physical means. The components of a mixture retain their individual properties. Mixtures usually are *heterogeneous* (unlike throughout), but solutions are an exception. Solutions are homogeneous mixtures in which one or more substances (solutes) are dispersed in another (solvent). Sugar in water is an example of a solution (homogeneous mixture). Sand in water is an example of a heterogeneous mixture.

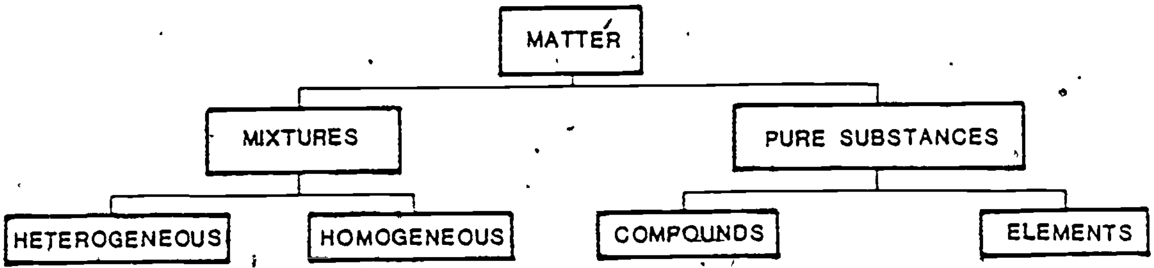


Figure 2. Classification of Matter.

ATOMIC STRUCTURE

As can be seen in Figure 2, pure substances are either compounds or elements. Elements contain only one kind of particle, called an *atom*. An atom is defined as "the smallest particle of an element which can take part in a chemical change." The element iron, for example, is composed of only iron atoms. A sample of iron can be subdivided over and over again until, theoretically, the last remaining particle will be an atom identical to all other iron atoms. Atoms are the basic building blocks of matter; they are the smallest units of an element that can combine with other elements.

Several atoms can combine to form molecules, which are electrically neutral particles composed of combinations of tightly bound atoms. Compounds are groups of molecules, just as elements are groups of atoms.

John Dalton, in 1808, proposed an atomic theory to explain the structure of matter. Dalton considered the atom to be an indivisible object; however, data has slowly accumulated to indicate that the atom is not a single particle, but actually is composed of several particles. Only three of the subatomic particles are of interest to this study of chemistry: the proton, neutron, and electron.

The electrical nature of matter was first studied by Benjamin Franklin, who discovered that there are two types of electrical charge, which he called "positive" and "negative." The electron is a fundamental particle of atoms, and it has a negative electrical charge. The size of the electron is extremely small; its mass is approximately 1/2000 the mass of the lightest atom, which is the element hydrogen.

The proton has an equal charge with the electron, but is opposite in sign; that is, the proton has a positive charge. The proton has a mass of one atomic mass unit (amu). The atomic mass unit is used to relate the weights of small particles such as atoms, protons, and electrons. The overall charge of the atom is neutral; therefore, it follows that atoms contain equal numbers of electrons and protons. For example, helium contains two electrons and two protons, and the resultant helium atom is neutral.

The electron and proton can account for the charges of an atom but leave unanswered the question of mass of the atom. For example, the mass of the helium atom is 4. Two mass units are due to protons, and two mass units are due to a neutral particle equal in weight to the proton. This neutral particle is the neutron. The properties of the three fundamental particles of an atom – electron, proton, and neutron – are given in Table 1.

TABLE 1: PROPERTIES OF THE THREE FUNDAMENTAL PARTICLES.

Particle	Charge	Mass (in amu)	Location in Atom
Electron	Negative, -1	0.000548597	Surrounding nucleus
Proton	Positive, +1	1.00727663	Inside nucleus
Neutron	Neutral, 0	1.0086654	Inside nucleus

THE ATOM

All atoms are composed of the same basic building blocks: the electron, neutron, and proton. The primary difference between atoms of different elements is the number of these building blocks that are present. Niels Bohr described the atom as being much like the solar system, but on a very small scale. There is a very small core, known

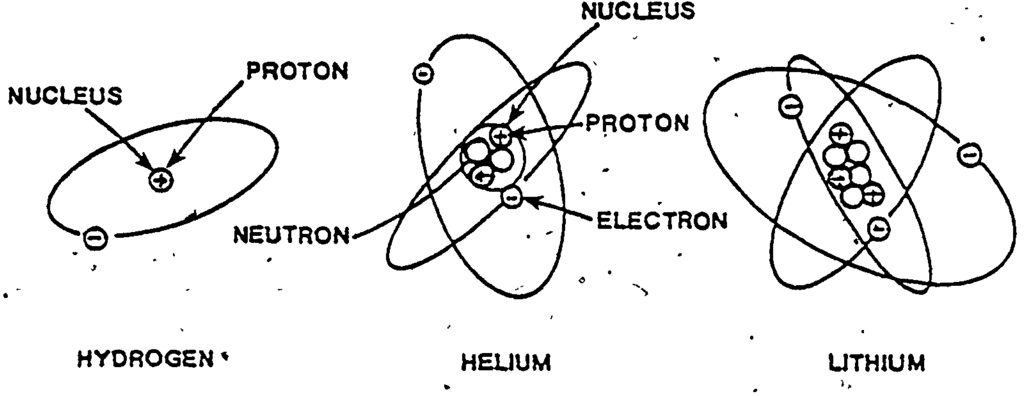


Figure 3. Structure of Atoms.

as the nucleus of the atom, around which particles spin at tremendous speeds (Figure 3). These spinning particles are negatively charged electrons. Electrons that spin around the nucleus are very far apart, just as the planets and the sun in the solar system are. The electrons' speed varies from 10,000 to 100,000 miles per second, which is approaching the speed of light. The nucleus is composed of neutrons and protons in sufficient number to make up the total weight of the atom (electrons have a negligible weight). Since the number of positive protons in the nucleus is the same as the number of negative electrons, the overall charge of the atom is neutral.

COMPOSITION OF THE ATOM

Each atom is designated by an atomic number, which is the number of protons in the atom and also the number of electrons in the atom. These numbers range from 1 (for hydrogen) to 103 (for lawrencium). That is, there are 103 known elements. (Scientists in Russia have claimed that they have made or discovered elements number 104 and 105, but these discoveries have not been fully verified.) The atomic weight of an element is the sum of the weights of the protons, neutrons, and electrons. As already indicated, for all practical purposes the weight of the electrons can be ignored since it is so small compared to the weights of protons and neutrons. Since the weights of both protons and neutrons are very nearly equal to 1, atomic weights of many elements are nearly a whole number. Atomic weights, when rounded off, give the number of nucleons (total number of particles in the nucleus). This number is called the mass number of the element. Table 2 gives the composition of the first 10 atoms. For example, boron (with the shorthand symbol of B) has a mass number of 11. Its atomic number, number of protons, and number of electrons are 5, and the number of neutrons is 6.

TABLE 2. ATOMIC COMPOSITIONS.

Element	Symbol	Atomic Weight	Mass Number	Atomic Number	Number Protons	Number Electrons	Number Neutrons
Hydrogen	H	1.00797	1	1	1	1	0
Helium	He	4.0026	4	2	2	2	2
Lithium	Li	6.939	7	3	3	3	4
Beryllium	Be	9.0122	9	4	4	4	5
Boron	B	10.811	11	5	5	5	6
Carbon	C	12.01115	12	6	6	6	6
Nitrogen	N	14.0067	14	7	7	7	7
Oxygen	O	15.9994	16	8	8	8	8
Fluorine	F	18.9984	19	9	9	9	10
Neon	Ne	20.183	20	10	10	10	10

In summary: Atomic number = number of protons = number of electrons.
Mass number = number of protons + number of neutrons.

The structure of the first three atoms, hydrogen, helium, and lithium, is shown in Figure 2. This representation of the atoms is an attempt to show their three-dimensional nature. In more complex atoms it is impossible to show all of the protons, electrons, and neutrons. For example, lawrencium has 103 protons, 103 electrons, and 154 neutrons and would be impossible to depict in a drawing. Therefore, the following representations will be used to indicate atomic structures.

- Hydrogen: The hydrogen atom (Figure 4) — consists of one proton (p^+) and one electron (e^-). The single electron moves around the proton and is in the first shell.

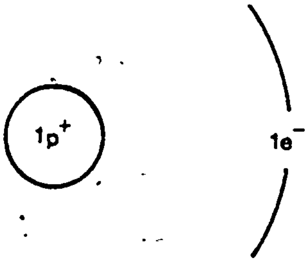


Figure 4. Hydrogen Atom.

- Helium: Another simple atom (Figure 5) in terms of atomic structure is gas helium, which has two protons, two neutrons (n), and two electrons. The two electrons in the first shell completely fill it.

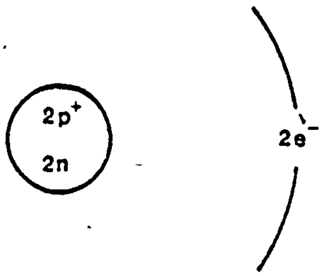


Figure 5. Helium Atom.

The maximum number of electrons possible in the first five shells is given in Table 3.

TABLE 3. MAXIMUM NUMBER OF ELECTRONS POSSIBLE IN FIRST FIVE SHELLS.

Shell	Maximum Number of Electrons
1	2
2	8
3	18
4	32
5	50

Lithium: The next atom (Figure 6) has an atomic number of three and an atomic weight of 7. Lithium has three electrons, three protons, and four neutrons. Two electrons are in the first shell, and one is in the second shell.

Figure 6. Lithium Atom.

- Sodium: The sodium atom (Figure 7) has 11 electrons, 11 protons, and 12 neutrons. Since the first and second shells are filled with 2 and 8 electrons respectively, the eleventh electron is in the third shell.

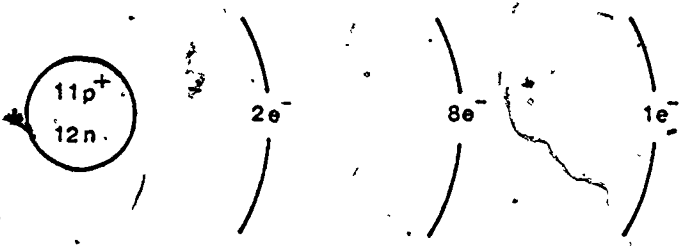


Figure 7. Sodium Atom.

The preceding examples show the orderly arrangement of the atoms. Each succeeding atomic structure is simply "built-up" by adding additional electrons, protons, and neutrons.

SYMBOLS

Chemical symbols are often used instead of writing the full name of an element. The symbol consists of one or two letters of the element's English or Latin name. The elements using a symbol based on their Latin name have been known for many centuries. These include the following: iron (Fe, for ferrum); copper (Cu, for cuprum); gold (Au, for aurum); silver (Ag, for argentum); tin (Sn, for stannum); and lead (Pb, for plumbum). Symbols for some important elements are given in Table 4.

TABLE 4. SYMBOLS OF IMPORTANT ELEMENTS.

Element	Symbol	Element	Symbol
Aluminum	Al	Lead	Pb
Antimony	Sb	Lithium	Li
Argon	Ar	Magnesium	Mg
Arsenic	As	Manganese	Mn
Barium	Ba	Mercury	Hg
Beryllium	Be	Neon	Ne
Bismuth	Bi	Nickel	Ni
Boron	B	Nitrogen	N
Bromine	Br	Oxygen	O
Cadmium	Cd	Phosphorus	P
Calcium	Ca	Platinum	Pt
Carbon	C	Potassium	K
Chlorine	Cl	Radium	Ra
Chromium	Cr	Silicon	Si
Cobalt	Co	Silver	Ag
Copper	Cu	Sodium	Na
Fluorine	F	Strontium	Sr
Gold	Au	Sulfur	S
Helium	He	Tin	Sn
Hydrogen	H	Tungsten	W
Iodine	I	Uranium	U
Iron	Fe	Zinc	Zn

NAMING OF COMPOUNDS

Distinct, known compounds number in the millions. Because of this, a systematic way of naming compounds is of utmost importance. The situation would be hopeless if every compound had a name totally independent of all others; therefore, some relatively simple rules are used for naming compounds. Some chemical compounds are still identified by their traditional names, such as H_2O (water) and NH_3 (ammonia). It is very important that careful attention be given to chemical names. For example, substituting perchloric for hydrochloric acid in a procedure could result in a serious explosion. A similar accident could occur if potassium chlorate is used for potassium chloride.

The simplest compounds are those that contain only two elements. They are called binary compounds. If the compound contains a metal and a nonmetal, the rule is as follows: The metal is named first, and the suffix -ide is made part of the nonmetal, as shown below:

$NaCl$	Sodium chloride
CaO	Calcium oxide
Na_2S	Sodium sulfide
Mg_3N_2	Magnesium nitride
BF_3	Boron fluoride

Nonmetals often combine with other nonmetals to form compounds. Because the same nonmetals can form several compounds, it is necessary to use prefixes to indicate the number of atoms of an element that is in the compound (Table 5).

TABLE 5. COMPOUND PREFIXES.

Prefix	Number	Example.
mono-	one	CO = carbon monoxide
di-	two	CO_2 = carbon dioxide
tri-	three	NCl_3 = nitrogen trichloride
tetra-	four	CCl_4 = carbon tetrachloride
penta-	five	PCl_5 = phosphorus pentachloride
hexa-	six	SF_6 = sulfur hexafluoride

Sometimes it is necessary to use the prefix to give the number of atoms for both elements in a binary compound. This is particularly true when an entire series of different compounds can be formed between two elements:

N_2O	Dinitrogen oxide
NO	Nitrogen oxide
N_2O_3	Dinitrogen trioxide
N_2O_5	Dinitrogen pentoxide

Some metals have variable combining capacities (often called <u>valences</u>) when combining with a nonmetal. There are two ways in use to handle the preceding situation: in one, the suffix -ous is used for the lower valence; and in the other, the suffix -ic is used for the higher valence. Another system uses a Roman numeral with the English name of the element to indicate the valence of the first element being named:

		Valence
$CuCl$	= Cuprous chloride (copper [I] chloride)	+1
$CuCl_2$	= Cupric chloride (copper [II] chloride)	+2
FeO	= Ferrous oxide (iron [II] oxide)	+2
Fe_2O_3	= Ferric oxide (iron [III] oxide)	+3

The simple acids that are water solutions of compounds of hydrogen with nonmetals are called hydro_____ic acids. The following are examples:

HCl = hydrochloric
HBr = hydrobromic
H_2S = hydrosulfuric
H_2Se = hydroselenic

All the salts of these acids are -ide compounds, such as the following:

$NaCl$ = sodium chloride
KBr = potassium bromide
MgS = magnesium sulfide
$CaSe$ = calcium selenide

Oxygen-containing acids are named according to the nonmetal and the number of oxygens contained in the compound. The suffix -ous is used for the lesser amount of oxygen, and -ic is used for the larger amount, as follows:

H_2SO_3 = sulfurous acid
H_2SO_4 = sulfuric acid
H_3PO_3 = phosphorous acid
H_3PO_4 = phosphoric acid

Some nonmetals form an even greater variety of oxygen-containing acids. In the following examples of chlorine-containing acids, the prefix hypo- means "less than," and the prefix per- means "more than." The names of salts are dervied from the names of the corresponding acids. In salts, the suffix -ite corresponds to the -ous acids; the suffix -ate corresponds to the -ic acids. If the acid name includes a prefix, this is retained in naming the salt, as shown by the following:

Acid	Salt
$HClO$ = hypochlorous acid	$NaClO$ = sodium hypochlorite
$HClO_2$ = chlorous acid	$NaClO_2$ = sodium chlorite
$HClO_3$ = chloric acid	$NaClO_3$ = sodium chlorate
$HClO_4$ = perchloric acid	$NaClO_4$ = sodium perchlorate
H_2SO_3 = sulfurous acid	Na_2SO_3 = sodium sulfite
H_2SO_4 = sulfuric acid	Na_2SO_4 = sodium sulfate

FORMULA WRITING

In chemistry, symbols are used for elements, formulas are used for compounds, and equations are used for expressing chemical reactions. These devices not only save time, they also make possible a great deal of mathematical calculations.

Valence has been defined as "the combining capacity of an atom or ion radical." Many people prefer to call these numbers oxidation numbers, rather than valences; both terms are commonly used. The valence, or oxidation number, is closely related to the structure of the atom. Hydrogen, with one electron in its outer structure, has an oxidation number of +1; calcium, with two electrons, has an oxidation number of +2; and aluminum, with three electrons in its outer structure, has an oxidation number of +3. It can be said that hydrogen has a valence of +1, calcium a valence of +2, and aluminum a valence of +3. Common valences, or oxidation numbers, of some elements are given in Table 6.

TABLE 6. COMMON VALENCES OR OXIDATION NUMBERS.

Element	Oxidation Number	Element	Oxidation Number
Al	+3	Fe	+3, +2
Sb	+3, +5	Pb	+2, +4
As	+3, +5	Li	+1
Ba	+2	Mg	+2
Be	+2	Mn	+2, +4, +6, +7
Bi	+3	Hg	+2, +1
B	+3	Ni	+2
Br	-1	N	+5, +3, -3
Cd	+2	O	-2
Ca	+2	P	+3, +5, -3
C	+4, -4	K	+1
Cl	-1	Ra	+2
Cr	+2, +3, +6	Si	+4, -4
Co	+2	Ag	+1
Cu	+2, +1	Na	+1
F	-1	Sr	+2
He	0	S	+6, +4, -2
H	+1	Sn	+2, +4
I	-1	Zn	+2

Some elements have several valences — even positive and negative valences. This simply means that these atoms have variable combining capacity. In some compounds, they may be the most positive element; and in some compounds, they may be the most negative element.

Rules for formula writing include the following:

- Elements of positive valences or oxidation numbers can unite only with elements of negative valences or oxidation numbers.
- The symbol of the element with the positive valence is normally placed first.
- The number of atoms of a given element in a formula is represented by a subscript number to the right of the symbol, unless it is 1. If no number appears, 1 atom is assumed for that element. $NaHSO_4$ indicates that there is 1 sodium, 1 hydrogen, 1 sulfur, and 4 oxygen atoms in the molecule.
- The ratio of the numbers of atoms used in a formula is inverse to the ratio of the valences of the two elements. The following is an example:

  A compound contains iron, with an oxidation number of +3, and oxygen, with an oxidation number of -2; the correct formula for this compound is Fe_2O_3. The following step-by-step procedure can be used to find the correct formula. Write down the elements involved and their valences:

  $$Fe^{+3}\ O^{-2}$$

  The lowest common multiple of the two charges, +3 and -2, is 6. If the charge of each atom is divided into 6, one finds that 2 iron atoms, with a +3 charge each, are needed to balance 3 oxygen atoms, with a -2 charge each. Then the following may be written:

  $$Fe^{+3}_{2}\ O^{-2}_{3}$$

The plus 6 of the iron balances the minus 6 of the oxygen. Leaving off the charges now gives the correct formula as follows:

$$Fe_2O_3$$

An ion is an atom or group of atoms which has gained or lost one or more electrons, thus giving the ion a positive or negative charge. For example, the sodium atom (Na) can lose an electron, forming the sodium ion (Na^+). Groups of ions often behave as if they were single ions in chemical reactions. The valences of some of the common radical ions are in Table 7.

TABLE 7. VALENCES OF RADICAL IONS.

Ion	Name	Valence
(NH_4)	Ammonium	+1
(OH)	Hydroxide	-1
(NO_3)	Nitrate	-1
(ClO_3)	Chlorate	-1
(SO_4)	Sulfate	-2
(CO_3)	Carbonate	-2
(PO_4)	Phosphate	-3
(CN)	Cyanide	-1

In writing formulas containing radical ions, the same rules apply as though the ion consisted of a single element. Thus, the compound of sodium (+1) with the carbonate ion (-2) gives the compound Na_2CO_3. If two or more radical ions are used in a formula, parentheses must be used around the radical ion. An example of this is in the formula for barium nitrate: $Ba(NO_3)_2$. This formula indicates the compound contains 1 barium atom and 2 nitrate ions, each containing 1 nitrogen and 3 oxygen atoms.

CHEMICAL BONDING

Molecules are composed of two or more atoms held together by chemical bonds. The chemical bond is a force which is electrical in nature. Ionic bonding, covalent bonding, hydrogen bonding, and metallic bonding will be considered.

IONIC BONDING

Ionic bonds are formed by the electrical force of attraction between a positive and a negative ion. For example, in sodium chloride (NaCl) the positive sodium ion (Na^+) is attracted to the negative chlorine ion (Cl^-). The ions form a three-dimensional crystal, as shown in Figure 8. In general, ionic compounds are hard, brittle, and have relatively high melting points. The strong force caused by the attraction of the oppositely charged ions accounts for the high melting point and hardness. When ionic compounds are dissolved in water, the ions are free to move, and they will conduct an electric current.

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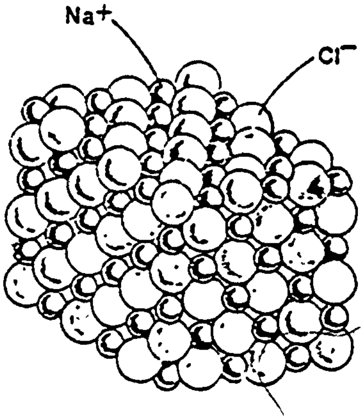

Figure 8. Crystal Structure of Sodium Chloride.

COVALENT BONDING

In some compounds the atoms share electrons to form a chemical bond which is called a covalent bond. For example, hydrogen (H_2) is bonded together through the sharing of an electron from each hydrogen atom to form an electron pair which is the covalent bond. Covalently bonded molecules generally have a relatively weak attraction for each other. In contrast to ionic crystals, covalent compounds are soft, have low melting points, and do not form ions in solution.

HYDROGEN BONDING

Hydrogen bonding is a weak force of attraction between certain molecules rather than a force within the molecules.

This type of bonding occurs in water. The structure of water is shown in Figure 9. The hydrogen end (positive) of one water molecule attracts the oxygen end (negative) of another water molecule. The result is a cluster of water molecules. This type of attraction, of hydrogen for a negative atom of an adjacent molecule, is called hydrogen bonding.

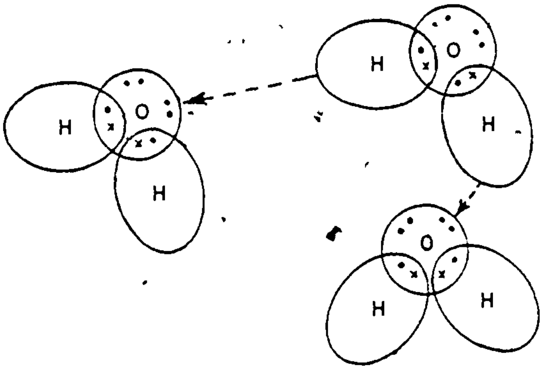

Figure 9. Structure of Water.

Hydrogen bonding is a very weak type of bond compared to ionic or covalent, but it greatly changes the expected properties of molecules. A graphic illustration of the effect of hydrogen bonding is found in Figure 10. The abnormally high boiling points for H_2O and HF are accounted for by hydrogen bonding that takes place in these compounds.

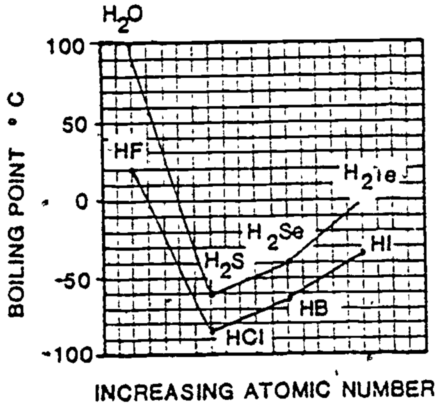

Figure 10. Effect of Hydrogen Bonding on Boiling Points.

METALLIC BONDING

Properties of metals, such as high electrical and thermal conductivity, luster, and high reflectivity are thought to be due to an electron-sea in which the electrons are mobil and move rather freely around fixed metallic ions. The force that holds the whole structure together is an attraction of positively charged metallic ions for the sea of negatively charged electrons. This force of attraction is called metallic bonding. An electron-sea model for a singly charged metal ion such as sodium, and for a doubly charged metal ion such as magnesium, is shown in Figure 11.

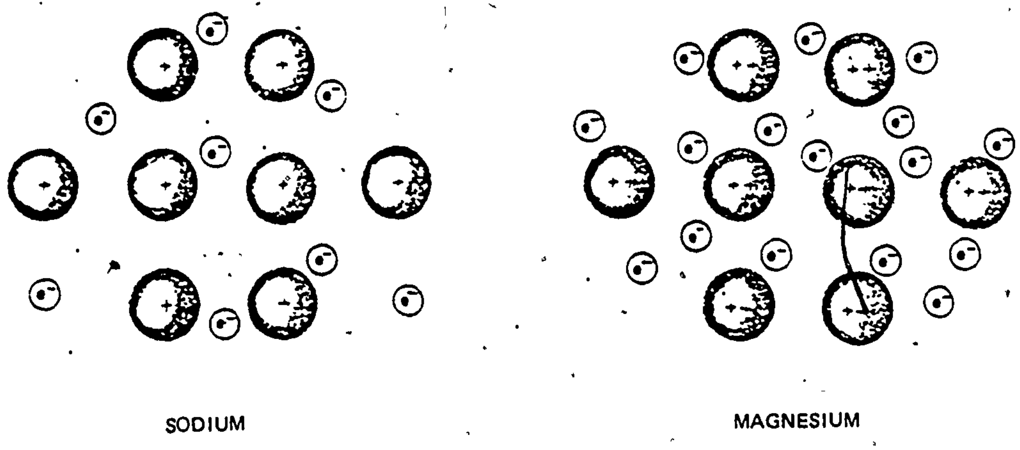

Figure 11. Electron-Sea Model for Metallic Bonding.

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EXERCISES

1. Chlorine is a greenish-yellow gas. It can be changed to a liquid by cooling to -34.6°C; and it reacts explosively with sodium to form sodium chloride (table salt). Which of these properties are physical properties and which are chemical?

	Physical Property	Chemical Property
Greenish-yellow	______	______
Gas	______	______
Changes to a liquid by cooling to -34.6°C	______	______
Reacts explosively with sodium to form sodium chloride	______	______

2. Identify the following as an element, compound, or mixture by placing an E (element), C (compound), or M (mixture) in the blank space.

Milk ____
Gold ____
Ink ____
Water ____
Zinc ____
Sodium chloride ____

3. Use the rules given in the Subject Matter to name the following compounds:

CaO ______________________________
$MgCl_2$ ______________________________
KOH ______________________________
CO ______________________________
PCl_5 ______________________________
KCN ______________________________
$MgCO_3$ ______________________________
$CuCl$ ______________________________
$CuCl_2$ ______________________________

NaH_2PO_4 ____________________

$PbCl_2$ ____________________

$PbCl_4$ ____________________

4. Give the valences of the first element in each of the following formulas:

$FeCl_2$ ______

Fe_2O_3 ______

$Cu(OH)_2$ ______

Cu_2S ______

PbO ______

PbO_2 ______

$HgNO_3$ ______

$HgSO_4$ ______

5. Write the correct formulas for the following compounds:

Iron (II) hydroxide ________________

Zinc cyanide ________________

Tin (IV) oxide ________________

Boron trifluoride ________________

Ammonium nitrate ________________

Potassium iodide ________________

Carbon disulfide ________________

Silicon tetrafluoride ________________

6. Complete the following table:

	Radical				
Ion	Cl^-	OH^-	NO_3^-	SO_4^{-2}	PO_4^{-3}
Ammonium, NH_4^+					
Sodium, Na^+	$NaCl$				
Calcium, Ca^{+2}		$Ca(OH)_2$			
Aluminum, Al^{+3}				$Al_2(SO_4)_3$	

7. Calcium has an atomic number of 20 and a mass number of 40. Supply the following information about calcium:

Number of electrons = __________

Number of protons = __________

Number of neutrons = __________

Number of electrons in the first shell = __________

Number of electrons in the second shell = __________

Number of electrons in the third shell = __________

Number of electrons in the fourth shell = __________

LABORATORY MATERIALS

Laboratory 1

Sodium bicarbonate, $NaHCO_3$
Magnesium sulfate, $MgSO_4 \cdot 7\ H_2O$ (epsom salt)
4 small crucibles with lids
Meker or Bunsen burner
2 ring stands
2 small iron rings
2 clay triangles
1 pair crucible tongs
1 desiccator
Balance

Laboratory 2

Electrical conductance apparatus
8 – 50 ml beakers
Distilled water
Methyl alcholol
Kerosene or gasoline
Acetone
Carbon tetrachloride
Potassium bromide
Potassium chloride
Barium chloride
Potassium nitrate
Sugar
Sodium chloride
Copper sulfate
Sodium phosphate

LABORATORY PROCEDURES

LABORATORY 1. DECOMPOSITION OF COMPOUNDS.

PROCEDURE

Carefully record all data in Data Table 1.

1. Clean and dry four crucibles with covers.
2. Weight each crucible, plus its lid, to the nearest 0.1 g.
3. Place approximately 2 g of sodium bicarbonate in one of the crucibles, then cover the crucible and weigh it to the nearest 0.1 g.

4. Repeat Step 3 with the second crucible, using a sample of sodium bicarbonate that weighs approximately 4 g.
5. Support one of the covered crucibles on the clay triangle, then place the lid off center and gently heat it with a burner (Figure 12). The crucible should be heated to a dull red color for 15 minutes. Heat the second crucible the same way.

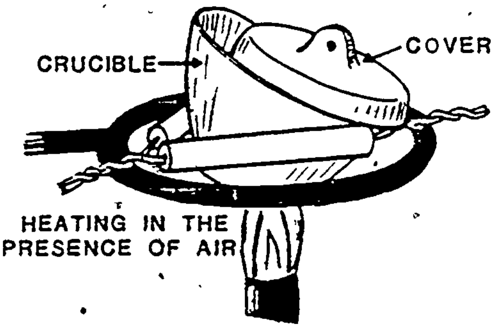

Figure 12. Heating a Crucible.

6. While the sodium bicarbonate is heating, weigh the second set of crucibles as in Step 2. Weigh out samples of epsom salt as described in Steps 3 and 4.
7. After the samples of sodium bicarbonate have been heated for about 15 minutes, remove the burners, center the lids, and allow them to cool slightly. Then place the crucible in a desiccator to cool to room temperature. Use crucible tongs to handle the hot crucibles.
8. While the crucibles containing sodium bicarbonate are cooling, heat the two crucibles containing the samples of epsom salt, using the procedure given in Step 5.
9. Weigh the crucibles and lids containing the residue from the decomposition of sodium bicarbonate as soon as they are cool.

10. After the samples of epsom salt have been heated to a dull red glow for 15 minutes, remove the burners, cover the crucibles, and allow them to cool slightly. Then place the samples in a desiccator and allow them to cool to room temperature. Weigh the cooled crucibles and lids.
11. After the experiments are completed, dispose of the residues and clean the crucibles with soap and water.
12. Perform the calculations as outlined in Data Table 1 to calculate the percent of residue for both samples of sodium bicarbonate and magnesium sulfate. If one substance has changed into another, a chemical reaction has occurred. In the experiment, the two reactions can be indicated by the following equations:

Sodium bicarbonate → Sodium carbonate + Carbon dioxide + Water   Equation 1

Epsom salt → Magnesium sulfate + Water   Equation 2.

In the first reaction, carbon dioxide and water are driven off into the atmosphere, and the residue is sodium carbonate.

In the second reaction, water is driven off from the epsom salt, leaving a residue of magnesium sulfate. In both of these reactions, the same percentage of residues for the 2- and 4-gram samples should be obtained, within experimental error. No matter what the source of sodium bicarbonate, if it is pure, it can always be decomposed to give the same percentage of sodium carbonate as every other pure sample. The same is true for other pure compounds, including epsom salt. All pure samples of epsom salt have the same percentage composition.

LABORATORY 2. CONDUCTIVITY OF ELECTROLYTES AND NONELECTROLYTES.

Electrolytes are substances that conduct an electric current when put into solution. Nonelectrolytes, on the other hand, are nonconductors of electricity. Substances that are ionically bonded form "charged" ions in solution which can "carry the electrical current." Acids, bases, and salts, generally are found to be electrolytes. Substances that remain as molecules in solution will not conduct electrical current. These substances are typically carbon containing or organic chemicals, and they are covalently bonded. In this experiment the student will differentiate between electrolytes and nonelectrolytes by measuring their conductivity.

A variety of electrical conductivity measuring devices are available. A common one utilizes 110 volt a.c. circuit with a light bulb in the circuit. When an electrical conductor comes in contact with the two electrodes, the bulb lights up. Some indication of the strength of the electrolyte can be determined from the intensity of the light. Some conductivity devices use milliammeters to enable one to measure the amount of current flowing. A diagram of two simple conductivity apparatus is shown in Figure 13.

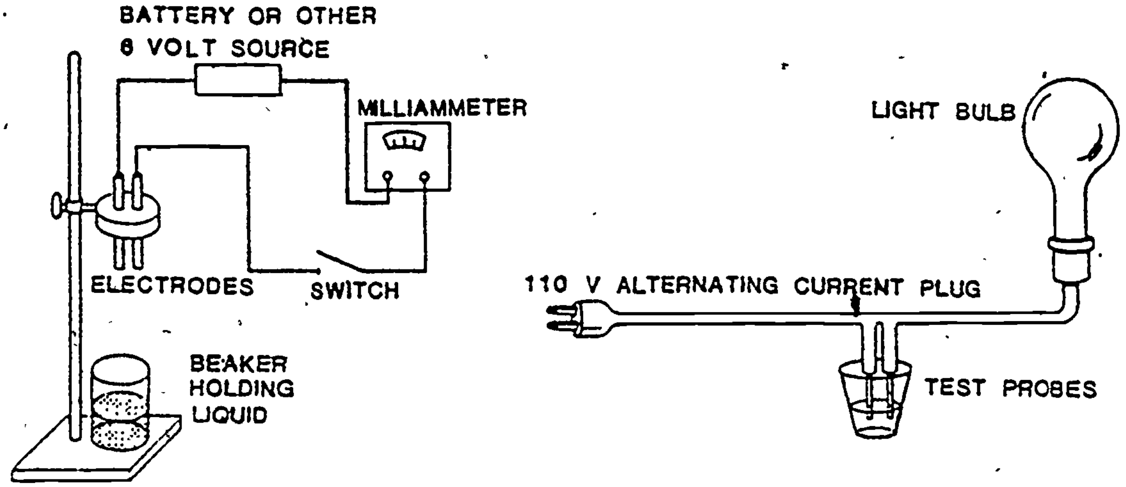

Figure 13. Assemblies to Test the Conductance of Electrolytes.

PROCEDURE

Carefully record all data in Data Table 2.

1. Be certain that the electrodes of the apparatus are clean. Rinse them prior to use and following each solution, using distilled water.
2. Place enough of the liquid to be investigated in a beaker and lower electrodes into place.
3. Use liquids shown in Data Table 2 and indicate whether solution is a nonconductor or conductor; if a conductor, indicate as weak, medium, or strong.

DATA TABLES

DATA TABLE 1. DECOMPOSITION OF COMPOUNDS.

DATA FOR DECOMPOSITION OF SODIUM BICARBONATE		
	Sample 1	Sample 2
Weight of crucible + lid + sample before heating		
Weight of crucible + lid + sample after heating		
Weight of crucible + lid		
Weight of sample (1 – 3)		
Weight of residue (2 – 3)		
Weight of gaseous product (1 – 2)		
% of residue (5/4) x 100%		

DATA FOR DECOMPOSITION OF MAGNESIUM SULFATE		
	Sample 1	Sample 2
Weight of crucible + lid + sample before heating		
Weight of crucible + lid + sample after heating		
Weight of crucible + lid		
Weight of sample (1 – 3)		
Weight of residue (2 – 3)		
Weight of gaseous product (1 – 2)		
% of residue (5/4) x 100%		

DATA TABLE 2. CONDUCTIVITY OF ELECTROLYTES AND NONELECTROLYTES.

COMPOUND	CONDUCTANCE			
	None	Weak	Medium	Strong
Distilled water				
Methyl alcohol				
NaCl solution				
Kerosene				
Acetone				
KBr solution				
KI solution				
Carbon tetrachloride				
Water and methyl alcohol				
$MgCl_2$ solution				
$BaCl_2$ solution				
Water and acetone				
Water and sugar				
Na_3PO_4 solution				
$CuSO_4$ solution				
KNO_3 solution				

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GLOSSARY

Atom: Smallest particle of an element.

Atomic mass unit: Unit of weight used to relate the weights of small particles such as electrons, protons, and neutrons.

Atomic number: A number equal to the number of electrons around the nucleus of an atom; also equal to the number of protons in the nucleus.

Atomic weight: The sum of the weights of the protons, neutrons, and electrons in an atom.

Binary compounds: Compounds containing two kinds of atoms.

Chemical bond: The force holding atoms together.

Chemical change: Involves the conversion of one substance into another.

Chemical symbols: The one or two letter abbreviation of an element.

Compound: A chemical substance containing more than one kind of atom.

Covalent bond: The bond formed by the sharing of electrons of atoms.

Electron: Fundamental particle with a charge of -1 and negligible weight.

Element: The basic substances which cannot be decomposed into simpler substances.

Heterogeneous: Having nonuniform composition.

Homogeneous: Having uniform composition.

Hydrogen bond: A weak bond formed by the attraction of hydrogen for the negative atom of an adjacent molecule.

Ion: An atom or group of atoms which has gained or lost one or more electrons.

Ionic bond: The bond formed by the electrical force of attraction between a positive and a negative ion.

Mass number: The number of protons and neutrons in the nucleus (same as nucleons).

Matter: Anything which occupies space and has weight.

Metallic bond: The force of attraction in a metal in which the positively charged metallic ions are attracted to a negative "sea" of electrons.

Mixture: Combinations of compounds or elements with a variable composition.

Molecule: Combination of atoms.

Neutron: Fundamental particle which is neutral and has a relative weight of 1 amu.

Nucleons: The total number of particles in the nucleus.

Nucleus: The small, dense, positively charged region at the center of an atom.

Physical change: A change that does not involve creation of new substances.

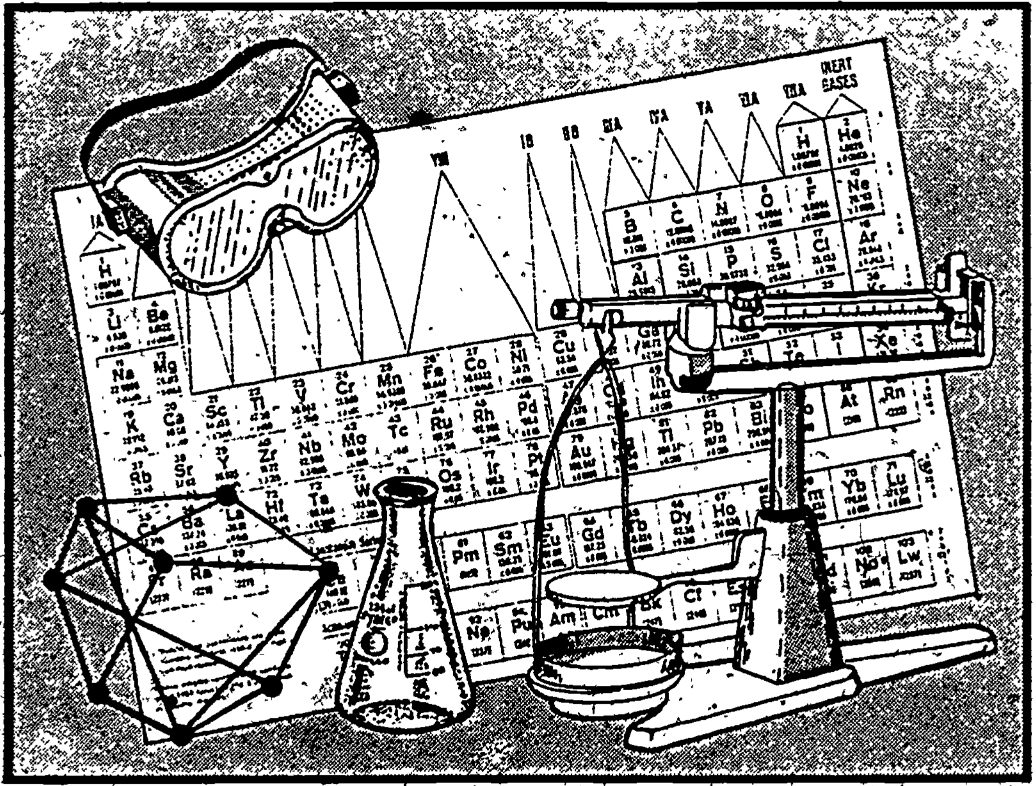
Proton: Fundamental particle with a charge of +1 and a weight of approximately 1 amu.

Valence: The combining capacity of an atom.

ENERGY TECHNOLOGY

CONSERVATION AND USE

CHEMISTRY FOR ENERGY TECHNOLOGY I



MODULE CH - 03

CHEMICAL EQUATIONS AND CALCULATIONS

CORD CENTER FOR OCCUPATIONAL RESEARCH AND DEVELOPMENT

INTRODUCTION

Most chemical work involves quantitative measurements and calculations which require an understanding of mathematics as it relates to chemistry. Scientific notation, significant figures, and the metric system are considered in this module, as well as a problem solving technique called dimensional analysis. The knowledge gleaned from the previous module, "Structure of Matter," is expanded with the study of atomic weights, isotopes, molecular weights, and mole concepts.

This module goes beyond writing and balancing chemical equations and investigates some of the quantitative information that can be determined through the use of chemical equations. The analytical method of gravimetric analysis is also presented.

PREREQUISITES

The student should have completed one year of high school algebra and Modules CH-01 and CH-02 of Chemistry for Energy Technology I.

OBJECTIVES

Upon completion of this module, the student should be able to:

1. Express conventional numbers in exponential notation and convert exponential numbers into conventional numbers.
2. Determine the correct number of significant figures in answers to mathematical problems.

3. Convert Fahrenheit to Celsius temperatures and convert Celsius to Fahrenheit temperatures.
4. Use the technique of dimensional analysis in solving problems.
5. Write and balance chemical equations.
6. Determine atomic and molecular weights of elements and compounds.
7. Calculate the number of moles in a given quantity of an element or compound.
8. Calculate the number of grams in a sample, given the number of moles in the sample.
9. Write the nuclear isotope symbols for an element, given the mass numbers.
10. Calculate amounts of reactants and products from chemical equations.

CHEMICAL CALCULATIONS

Most chemical work is quantitative in nature, that is, it deals with specific amounts of substances; and the results of chemical investigations are reported as numbers. Scientific notation is used to report figures which are unusually small or unusually large. Through the use of significant figures one can indicate the degree of accuracy in laboratory measurements. Moreover, chemical calculations are greatly simplified through the use of the metric system since it is based upon the decimal system. Scientific notation, significant figures, and the metric system, along with a problem solving technique called dimensional analysis, are all aids in performing chemical calculations and are considered in this module.

SCIENTIFIC NOTATION

Numbers used in chemistry are often very large or very small. For example, the number of molecules in 18.02 grams of water is 602,250,000,000,000,000,000,000, and the mass of an individual hydrogen atom is 0.00000000000000000000000167 grams. Obviously, these numbers would be cumbersome to handle in calculations; but expressed in exponential form, they are easier to use. Expressed exponentially, the first number above becomes 6.0225×10^{23}, and the second number becomes 1.67×10^{-25}. Table 1 shows general relationships between conventional numbers and exponential notation.

The exponent is used as a substitute for the number of places the decimal point must be moved to the right or left to give the long form of the number.

TABLE 1. EXPONENTIAL NOTATION.

$$1000 = 1 \times 10 \times 10 \times 10 = 1 \times 10^{3}$$

$$100 = 1 \times 10 \times 10 = 1 \times 10^{2}$$

$$10 = 1 \times 10 = 1 \times 10^{1}$$

$$1 = 1 = 1 \times 10^{0}$$

$$0.1 = \frac{1}{10} = \frac{1}{10^{1}} = 1 \times 10^{-1}$$

$$0.01 = \frac{1}{10 \times 10} = \frac{1}{10^{2}} = 1 \times 10^{-2}$$

$$0.001 = \frac{1}{10 \times 10 \times 10} = \frac{1}{10^{3}} = 10^{-3}$$

In order to add or subtract numbers expressed in exponential notation, the powers of 10 must be the same. For example, the following numbers may be added as follows:

$$(4.16 \times 10^{4}) + (2.14 \times 10^{2}) =$$

$$416 \times 10^{2} + 2.14 \times 10^{2} =$$

$$418.14 \times 10^{2} .$$

(addition)

When numbers expressed in exponential notation are multiplied, the exponents are added; but when numbers expressed in exponential notation are divided, the exponent of the divisor is subtracted from the exponent of the dividend. Note the following examples:

$(6.3 \times 10^4)(1.1 \times 10^8) =$

$(6.3 \times 1.1) \times 10^{4+8} =$ (multiplication)

6.93×10^{12}

$\frac{8.4 \times 10^5}{4.2 \times 10^3} =$

$\frac{8.4}{4.2} \times 10^{5-3} =$ (division)

$2.0 \times 10^2 = 200$

SIGNIFICANT FIGURES

In every experimental measurement there is a degree of uncertainty which depends upon the type of measuring device used and/or the skill of the person conducting the measurement. For example, 6 ml of liquid can be measured with a 100-ml graduated cylinder and appear to be 5 or 7 ml. In this case, an error of 1 ml is possible. However, the accuracy of measuring 6 ml of liquid with a 10-ml graduated cylinder can be expected to be within 0.1 ml of the 6 ml, or 5.9 to 6.1 ml. Finally, 6 ml of liquid can be measured with a buret with an uncertainty of only 0.01 ml. In this instance volume measurement could range from 5.99 to 6.01 ml. On the following page is a list of accuracies arising from these three measuring methods:

6 ml	(large graduated cylinder)
6.0 ml	(small graduated cylinder)
6.00 ml	(buret)

The accuracies listed above are expressed in terms of significant figures. There is one significant figure in 6 ml, two significant figures in 6.0, and three significant figures in 6.00 ml. When experimental quantities are multiplied or divided, the number of significant figures in the result cannot exceed the number contained in the least accurate measurement. Or, this concept may be expressed as follows: An answer in an experimental determination is not better than the least accurate measurement in the experiment.

METRIC SYSTEM

The metric system of weights and measures is used throughout the world – except in the United States. However, this country is gradually moving toward use of the metric system. The major advantage of the metric system is the fact that it is based upon a decimal system; therefore, many calculations involve only moving the decimal point. The metric system employs a series of prefixes to indicate decimal fractions of basic measurements. Some of these prefixes are given in Table 2.

Metric units are also known as International System units (SI units).

TABLE 2. METRIC SYSTEM PREFIXES.

Prefix	Abbreviation	Value	Example
Mega-	M	10^6	megawatt = 1 million watts
Kilo-	k	10^3	kilometer = 1000 meters
Deci-	d	10^{-1}	decimeter = 0.1 meter
Centi-	c	10^{-2}	centimeter = 0.01 meter
Milli-	m	10^{-3}	millimeter = 0.001 meter

The basic metric unit of length is the meter. As can be seen from the comparison between metric and English system units shown in Table 3, the meter is slightly longer than 1 yard.

TABLE 3. COMPARISON OF ENGLISH AND METRIC UNITS.

	Metric System	English System
Length	1 meter = 1.094 yd	2.54 cm = 1 in
Mass	1 kilogram = 2.205 lb	453.6 g = 1 lb
Volume	1 liter = 1.06 quarts	1 cubic ft = 28.32 liter

The basic metric unit of mass is the kilogram (kg). Although it is common practice to use the terms "mass" and "weight" interchangeably, this practice is not correct. Mass,

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which is a measure of the amount of material in an object, does not depend upon the attractive force of gravity. In other words, mass is constant; whereas, weight is dependent upon gravitational forces. The unit of mass most often used in chemistry is the gram (g), which is one one-thousandth of a kilogram.

The liter is the common metric measurement for volume. A liter is the volume occupied by a cubic decimeter; therefore, there are 1000 cubic centimeters (cc) in a liter. The milliliter (ml) is one one-thousandth of a liter and is equivalent to a cubic centimeter. In a laboratory, most chemical operations are conducted in the liquid phase; therefore measurement of volume is important.

Density, a quantity that relates mass to volume, is widely used to identify substances. Density is defined as "the ratio of mass to volume."

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

EXAMPLE A: DETERMINATION OF DENSITY.

Given: The information that 790 grams (g) occupy a volume of 100 cm^3 (cubic centimeters).

Find: Calculate the density of iron, using the above information.

Solution:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{790\ g}{100\ cm^3} = 7.90\ g/cm^3.$$

The common temperature scale used in the United States is Fahrenheit. In scientific studies, Celsius (centigrade) and Kelvin scales are employed. The Celsius scale is based on the assignment of 0° to the freezing point of water and 100°C to the boiling point of water (at sea level). Of course, the corresponding points on the Fahrenheit scale are 32°F and 212°F. The mathematical relationship between Fahrenheit and Celsius scales is as follows:

$$°C = \frac{5}{9}\,(°F - 32°F)$$

The Kelvin scale is related to the properties of gases. Zero on this scale corresponds to -273.15°C. Kelvin and Celsius scales are related as follows:

$$K = °C + 273.15°$$

The Kelvin, Celsius, and Fahrenheit temperature scales are compared in Figure 1.

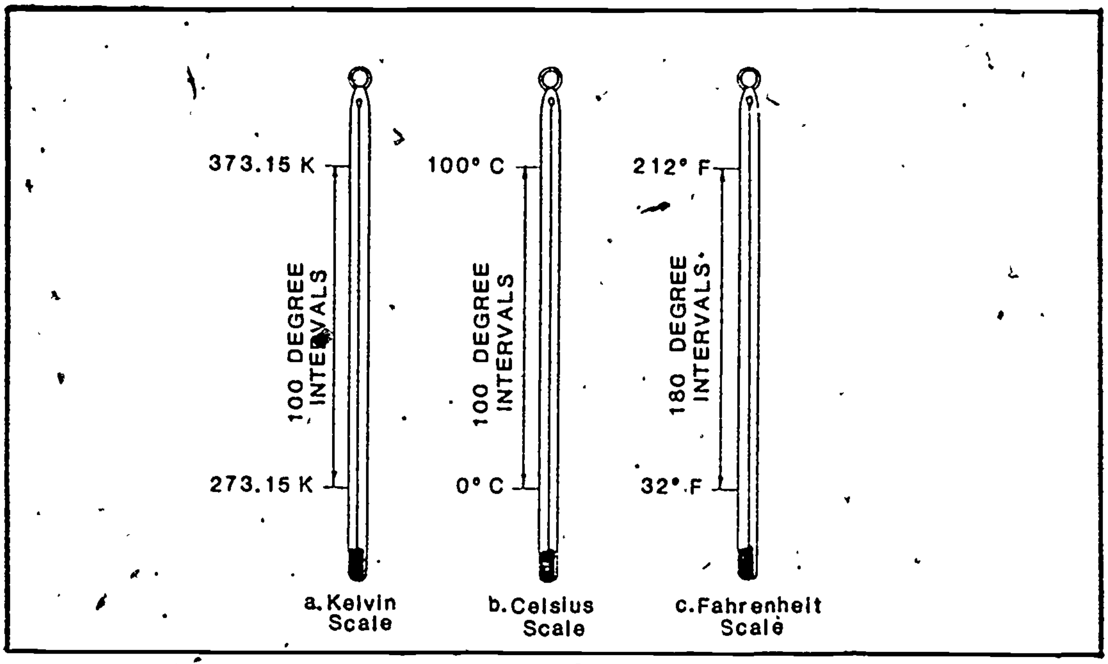

Figure 1. Comparison of the Kelvin, Celsius, and Fahrenheit Temperature Scales.

PROBLEM SOLVING

In recording experimental data or answers to mathematical problems it is important to indicate the units of measurement. A number reported for a measured quantity is meaningless unless its units are known. An approach which makes use of units in problem solving is called <u>dimensional analysis</u>, or the factor method. Consider the following example of this method of problem solving:

EXAMPLE B: DIMENSIONAL ANALYSIS.

Given: A trafic sign that indicates the speed limit is 40 kilometers (km) per hr.

Find: The corresponding speed in miles per hour.

Solution:

$$\frac{mi}{hr} = 40\left(\frac{\cancel{km}}{hr}\right)\left(\frac{1000\ \cancel{m}}{1\ \cancel{km}}\right)\left(\frac{1.094\ \cancel{yd}}{1\ \cancel{m}}\right)\left(\frac{1\ mi}{1760\ \cancel{yd}}\right)$$

$$= 25\ \frac{mi}{hr}.$$

Notice that the units are canceled, leaving only the desired units mi/hr. The units should be carried through all calculations to make sure that they cancel properly. The magnitude of answers should always be considered to determine if they are reasonable. For example, if the decimal point were misplaced in the above problem, the answer could be 2.5 or 250 mi/hr. Obviously, neither of these answers would be reasonable for a posted speed limit.

CHEMICAL EQUATIONS

The symbol of a chemical element is the shorthand notation for the element. The symbol represents an atom of the element. The molecular or ionic formula of a compound is a combination of symbols for the individual atoms of the compound. Symbols and formulas are both used in writing chemical equations. Chemical equations are used to indicate chemical reactions. A chemical equation is, therefore, both a qualitative and a quantitative statement of a chemical change. As will be seen later in this module a balanced chemical equation can be used to calculate the amount of product of a given reaction, and for other important calculations. The chemical equation and the atomic and molecular weight relationships of the atoms and molecules involved in a chemical reaction are used in an important analytical technique called gravimetric analysis.

ATOMIC WEIGHT

In Module CH-02 the atomic weight of an element was defined as "the sum of the weights of the protons, neutrons, and electrons in the atom." The actual weight of hydrogen, the lightest atom, has been determined to be 1.7 x 10^{-24} grams. This method of expressing the actual weights of atoms is inconvenient, and since all elements are of the same small order of magnitude, it is more convenient to use relative weights than actual weights. The relative weights of atoms of different elements are known as atomic weights. The atomic weights of all elements, along with their symbols and atomic numbers, are given in Table 4.

On the atomic weight scale, the weight of hydrogen is 1.00797; that of carbon is 12.01115; and that of oxygen is 15.9994. Therefore, carbon atoms weigh approximately twelve times as much as hydrogren atoms, and oxygen atoms are approximately sixteen times heavier than hydrogen atoms. Sulfur, which weighs 32.064, is approximately twice as heavy as oxygen. (In the development of the atomic weight scale, a particular kind of carbon atom, called carbon-12 isotope, was given a weight of 12.00000.) Isotopes will be considered in the next section of this module.

TABLE 4. TABLE OF ATOMIC WEIGHTS.

	Symbol	Atomic No.	Atomic Weight		Symbol	Atomic No.	Atomic Weight
Actinium	Ac	89	227	Mercury	Hg	80	200.59
Aluminum	Al	13	26.9815	Molybdenum	Mo	42	95.94
Americium	Am	95	[243]*	Neodymium	Nd	60	144.24
Antimony	Sb	51	121.75	Neon	Ne	10	20.183
Argon	Ar	18	39.948	Neptunium	Np	93	[237]
Arsenic	As	33	74.9216	Nickel	Ni	28	58.71
Astatine	At	85	[210]	Niobium	Nb	41	92.906
Barium	Ba	56	137.34	Nitrogen	N	7	14.0067
Berkelium	Bk	97	[249]	Nobelium	No	102	[253]
Beryllium	Be	4	9.0122	Osmium	Os	76	190.2
Bismuth	Bi	83	208.980	Oxygen	O	8	15.9994
Boron	B	5	10.811	Palladium	Pd	46	106.4
Bromine	Br	35	79.909	Phosphorus	P	15	30.9738
Cadmium	Cd	48	112.40	Platinum	Pt	78	195.09
Calcium	Ca	20	40.08	Plutonium	Pu	94	[242]
Californium	Cf	98	[251]	Polonium	Po	84	210
Carbon	C	6	12.01115	Potassium	K	19	39.102
Cerium	Ce	58	140.12	Praseodymium	Pr	59	140.907
Cesium	Cs	55	132.905	Promethium	Pm	61	[145]
Chlorine	Cl	17	35.453	Protactinium	Pa	91	231
Chromium	Cr	24	51.996	Radium	Ra	88	226.05
Cobalt	Co	27	58.9332	Radon	Rn	86	222
Copper	Cu	29	63.54	Rhenium	Re	75	186.2
Curium	Cm	96	[247]	Rhodium	Rh	45	102.905
Dysprosium	Dy	66	162.50	Rubidium	Rb	37	85.47
Einsteinium	Es	99	[254]	Ruthenium	Ru	44	101.07
Erbium	Er	68	167.26	Samarium	Sm	62	150.35
Europium	Eu	63	151.96	Scandium	Sc	21	44956
Fermium	Fm	100	[253]	Selenium	Se	34	78.96
Fluorine	F	9	18.9984	Silicon	Si	14	28.086
Francium	Fr	87	[223]	Silver	Ag	47	107.870
Gadolinium	Gd	64	157.25	Sodium	Na	11	22.9898
Gallium	Ga	31	69.72	Strontium	Sr	38	87.62
Germanium	Ge	32	72.59	Sulfur	S	16	32.064
Gold	Au	79	196.967	Tantalum	Ta	73	180.948
Hafnium	Hf	72	178.49	Technetium	Tc	43	[99]
Helium	He	2	4.0026	Tellurium	Te	52	127.60
Holmium	Ho	67	164.930	Terbium	Tb	65	158.924
Hydrogen	H	1	1.00797	Thallium	Tl	81	204.37
Indium	In	49	114.82	Thorium	Th	90	232.038
Iodine	I	53	126.9044	Thulium	Tm	69	168.934
Iridium	Ir	77	192.2	Tin	Sn	50	118.69
Iron	Fe	26	55.847	Titanium	Ti	22	47.90
Krypton	Kr	36	83.80	Tungsten	W	74	183.85
Lanthanum	La	57	138.91	Uranium	U	92	238.03
Lawrencium	Lw	103	[257]	Vanadium	V	23	50.942
Lead	Pb	82	207.19	Xenon	Xe	54	131.30
Lithium	Li	3	6.939	Ytterbium	Yb	70	173.04
Lutetium	Lu	71	174.97	Yttrium	Y	39	88.905
Magnesium	Mg	12	24.312	Zinc	Zn	30	65.37
Manganese	Mn	25	54.9380	Zirconium	Zr	40	91.22
Mendelevium	Md	101	[256]				

*A value given in brackets denotes the mass number of longest-lived or best-known isotope.

Since atomic weights are relative, they can be assigned any units (for example, grams, ounces, pounds, and so forth). Most chemical work is done using the gram; therefore, atomic weights of elements are often called <u>gram-atomic weight</u>, which is the atomic weight expressed in grams. Thus, one gram-atomic weight of oxygen is 15.9994 grams, one gram-atomic weight of hydrogen is 1.00797 grams, and so forth. The term "atomic mass unit" (amu) is used to indicate the weight of one atom. For example, one atom of sulfur weighs 32.064 amu. Almost all calculations in chemistry involve the use of atomic weights.

ISOTOPES

In Module CH-02 the mass number of an element was defined as "the number of particles in the nucleus" – that is, the number of protons plus the number of neutrons. The atomic weights of elements are often very close to being whole numbers and can be rounded off to obtain the mass numbers. However, consider the fact that many atomic weights – such as chlorine's 35.45 – are far from being whole numbers. This led to the discovery that most elements are mixtures of two or more kinds of atoms having different atomic masses but similar chemical properties. Chlorine, with an atomic weight of 35.45, exists as two kinds of chlorine atoms that have masses very close to the whole numbers 35 and 37. Both types of chlorine have an atomic number of 17, meaning they both have 17 protons and 17 electrons. The difference lies in the number of neutrons in the nuclei of the different types of atoms: chlorine-35 has 18 neutrons and chlorine-37 has 20 neutrons. Atoms of the same atomic number and different atomic weight (different numbers of neutrons) are called isotopes. Chlorine contains 75.6% isotope 35 and 24.4% isotope 37. The average weight of this mixture of isotopes is 35.45.

The existence of isotopes is widespread among the elements. Some elements have only one isotope, some have a few, and some have many. Tin is composed of a mixture of nine isotopes that vary in mass from 112 to 124.

A symbolism has been devised to distinguish between isotopes of an element. The two isotopes of chlorine, for example, are designated by the symbols $^{35}_{17}Cl$ and $^{37}_{17}Cl$ to indicate chlorine-35 and chlorine-37 respectively. Note that the mass number is made the superscript and the atomic number, the subscript.

Hydrogen is composed of the three isotopes shown in Figure 2. More than 99.9% of hydrogen is normal hydrogen — that is, $^{1}_{1}H$, with one proton and one electron that weigh one.

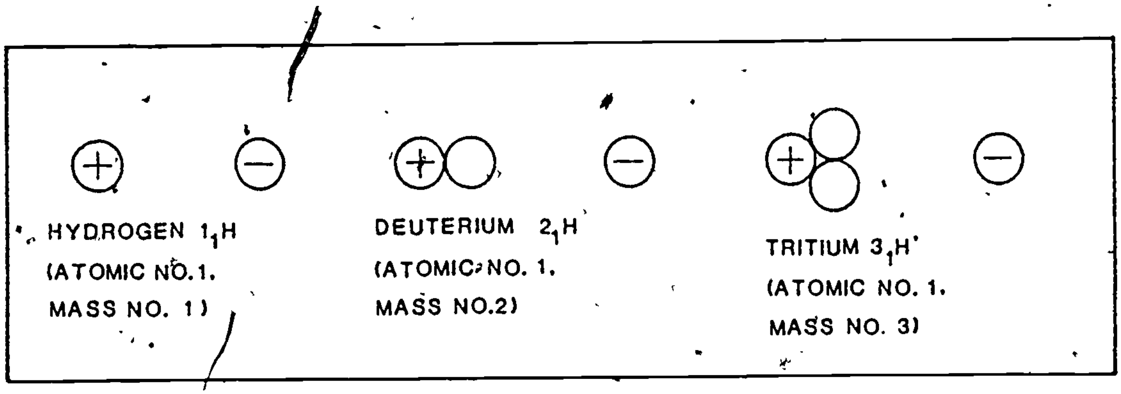

Figure 2. Isotopes of Hydrogen.

About one atom out of every 7000 atoms is heavy hydrogen (also called deuterium), which is a hydrogen atom whose nucleus contains one proton and one neutron. The symbol for heavy hydrogen is $^{2}_{1}H$. Another isotope of hydrogen, tritium, is extremely rare. Tritium ($^{3}_{1}H$) has two neutrons, one proton, and one electron. Tritium finds some use in the fusion reaction used in the hydrogen bomb. Since the majority of hydrogen is hydrogen-1, it is not surprising that its atomic weight is close to one (1.0079). The reason so many atomic weights are almost whole numbers is that there is usually one isotope present in very high concentration, just as in hydrogen.

Neon consists of three isotopes: $^{20}_{10}Ne$, $^{21}_{10}Ne$, and $^{22}_{10}Ne$. The natural abundances of these isotopes are 90.9%, 0.3%, and 8.8%, respectively. As would be expected from these numbers, the atomic weight of neon is close to 20 (20.179).

Since the electronic structure of all isotopes of an element are identical, the chemical properties are the same. Some of the physical properties (freezing point, melting point, or density) may be slightly different since these properties are somewhat dependent upon the total weight of the atom.

MOLECULAR WEIGHTS

The molecular weight of a compound is the sum of the atomic weights of all the atoms shown in its formula. The term "formula weight" is preferred by some people, since some ionic compounds such as NaCl do not exist as molecules. For most purposes, the terms "molecular weight" and "formula weight" are interchangeable.

What is the molecular weight of the following compounds?

NaOH	# Atoms		Atomic weight	
Na	1	x	22.99	= 22.99
O	1	x	15.99	= 15.99
H	1	x	1.01	= 1.01
			TOTAL	39.89 grams

H_2SO_4	# Atoms,		Atomic Weight	
H	2	x	1.01	= 2.02
S	1	x	32.06	= 32.06
O	4	x	15.99	= 63.96
			TOTAL	98.07 grams

THE MOLE

Atomic weights are related to the number of atoms in a very definite way – which is explained by the mole concept. A mole of any element is defined as "the amount of that element that contains the same number of atoms" (as exactly 12 grams of carbon-12). It has been determined experimentally that this number is 6.02×10^{23} atoms. This number is called Avogadro's number in honor of the Italian physicist, Ameaeo Avogadro. A mole of ions, atoms, or molecules contains Avogadro's number of ions, atoms, or molecules respectively.

1 mole of C atoms = 6.02×10^{23} C atoms
1 mole of H_2O molecules = 6.02×10^{23} H_2O molecules
1 mole of Cl^- ions = 6.02×10^{23} Cl^- ions

A mole of atoms of any element has a weight in grams equal to the atomic weight of the element, as follows:

1 mole of C weighs 12 grams
1 mole of S weighs 32 grams
1 mole of Al weighs 27 grams

Because atoms and molecules react with each other on a particle-to-particle level and not a gram-to-gram level, the mole concept is very important; it allows the technician to weigh materials (in grams for example) and to relate these weights to chemical reactions. Several mole relationships are given in Table 5.

TABLE 5. MOLE RELATIONSHIPS.

Name	Formula	Formula Weight (amu)	Weight of 1 mole (g)	Number and kind of Particles in 1 mole
Atomic oxygen	O	16	16	6.02×10^{23} O atoms
Molecular oxygen	O_2	32	32	6.02×10^{23} O_2 molecules $2(6.02 \times 10^{23})$O atoms
Silver	Ag	108	108	6.02×10^{23} Ag atoms
Silver ions	Ag^+	108	108	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208	208	6.02×10^{23} $BaCl_2$ molecules 6.02×10^{23} Ba^{+2} ions $2(6.02 \times 10^{23})Cl^-$ ions

EXAMPLE C: MOLE CALCULATIONS.

Given: The quantity of 80.3 grams of sulfur.

Find: Gram-atomic weights (moles).

Solution:

$$\text{Gram-atoms of S} = \frac{\text{Grams of S}}{\text{Gram-atomic weight of S}}$$

$$= \frac{80.3\ g}{32.1\ g} = 2.50.$$

EXAMPLE D: MOLE CALCULATIONS

Given: The quantity of 138 grams of ethyl alcohol (C_2H_5OH).

Find: Number of moles.

Solution:

$$2\ C = 2 \times 12 = 24\ g$$
$$6\ H = 6 \times 1 = 6\ g$$
$$1\ O = 1 \times 16 = \underline{16\ g}$$
$$\text{Molecular weight of ethyl alcohol} = 46\ g$$

$$\text{Moles of ethyl alcohol} = \frac{\text{Grams of ethyl alcohol}}{\text{Molecular weight of ethyl alcohol}}$$

$$= \frac{138\ g}{46\ g} = 3.$$

EXAMPLE E: MOLE CALCULATIONS.

Given: Three moles of sodium hydroxide (NaOH).

Find: Quantity (weight) of NaOH for an experiment.

Solution:

$$\text{Grams of NaOH} = \text{Moles of NaOH} \times \text{Molecular weight of NaOH}$$

$$= 3\ \text{moles} \times (23 + 16 + 1)\ g/\text{mole}.$$

$$\text{Grams of NaOH} = 3\ \text{moles} \times 40\ g/\text{mole}$$

$$= 120\ \text{grams}.$$

WRITING AND BALANCING EQUATIONS

The use of symbols and chemical formulas as a condensed method of identifying elements and compounds was discussed in Module CH-02. Now, symbols and formulas will be used in writing chemical equations. A chemical equation is both a qualitative and a quantitative statement of a chemical change; and it expresses a number of laws and facts. The following are examples:

- The law of conservation of mass – a law that states that matter is neither created nor destroyed in a chemical reaction. Therefore, in a chemical equation which represents what happens in a reaction, the weight and number of each kind of atom is the same before and after the chemical reaction. That is, the total mass of the products is equal to the total mass of the reactants.
- The law of constant composition – each formula expresses the fact that the compound it represents has a constant composition. Each symbol and formula involved must be written correctly, using knowledge of valences. One cannot change a formula to achieve a balanced equation.
- Each equation indicates the relative weights of the reacting materials and the products formed. Weight relationships of chemical equations will be illustrated later.

To write a chemical equation, one must know the following:

- That the reaction actually can take place. This information is determined experimentally. If one writes equations but reactions do not occur, obviously, these equations are meaningless and misleading.

- The correct symbols and formula of the reacting substances (reactants). These formulas are obtrained from experimental data.
- The correct symbols and formulas of the products.
- Number of moles needed to balance atoms on each side of the equation. (The number of moles are the coefficients placed before the formulas.)

The word "equation" is not used in a mathematical sense. In mathematics, equation indicates that both sides of an expression are identical. Chemical equations are used to indicate chemical reactions – which, by definition, result in new and different substances on the right from the original substances on the left. The reactants are not "equal" to the products. The arrow in a chemical equation means "yields" or "produces." The "+" sign means "reacts with."

Examples of writing and balancing equations follow, Word equations for chemical reactions can be written as follows: Zinc reacts with hydrochloric acid to produce hydrogen and zinc chloride. This word equation can be written then as a chemical equation as follows:

$$Zn + HCl \longrightarrow H_2 + ZnCl_2$$

The following information is needed before the preceding equation can be written.

- The symbol for metallic zinc is Zn.
- The formula for hydrochloric acid is HCl.
- Zinc and hydrochloric acid will actually react.
- Hydrogen gas is produced in the reaction and is a diatomic gas, H_2.
- Zinc chloride is also a product, with a formula of $ZnCl_2$.

Note that 1 Zn, 1 H, and 1 Cl are on the left and that 1 Zn, 2H, and 2 Cl are on the right. The equation is unbalanced in this case; but the reaction can be balanced by placing a "2" in front of the HCl. Notice that the formula cannot be changed to H_2Cl_2 because the formula of hydrochloric acid is HCl – not H_2Cl_2.

$$Zn + 2\ HCl \longrightarrow H_2 + ZnCl_2$$

This balanced equation now states that one atom of zinc will react with two molecules of hydrochloric acid to yield one molecule of hydrogen gas and one molecule of zinc chloride. This equation also shows that one mole of zinc will react with two moles of hydrochloric acid to yield one mole of hydrogen gas and one mole of zinc chloride.

Magnesium reacts with sulfuric acid to produce hydrogen and magnesium sulfate as follows:

$$Mg + H_2SO_4 \longrightarrow H_2 + MgSO_4$$
$$1\ Mg,\ 2\ H,\ 1\ S,\ 4\ O \longrightarrow 1\ Mg,\ 2\ H,\ 1\ S,\ 4\ O$$

This equation as written is already balanced, and no coefficients are needed. Aluminum reacts with hydrochloric acid to produce hydrogen and aluminum chloride as follows:

$$Al + HCl \longrightarrow H_2 + AlCl_3$$

To balance this equation, one may use trial and error to place coefficients in front of the symbols and formula until a balanced equation results. For example, try placing a 3 before HCl to get 3 Cl on each side, giving 3 H on the left and 2 on the right. An even coefficient that is also divisible by 3 placed in front of HCl will balance the Cl on the right. The following is an example:

$$Al + 6\ HCl \longrightarrow H_2 + AlCl_3$$

To balance the H and Cl, a 3 before H_2 and a 2 before $AlCl_3$ is needed.

$$Al + 6\ HCl \longrightarrow 3\ H_2 + 2\ AlCl_3$$

Then, balance the Al as follows:

$$2\ Al + 6\ HCl \longrightarrow 3\ H_2 + 2\ AlCl_3 \text{ (balanced)}$$
$$2\ Al,\ 6\ H,\ 6\ Cl \qquad 2\ Al,\ 6\ H,\ 6\ Cl$$

When the products of a reaction are not known, it is much more difficult to write and balance chemical equations. Fortunately, one does not have to memorize every possible chemical reaction that can occur. Instead, learning types of chemical reactions and knowing how substances react makes it possible to predict how similar substances would also react. The examples given involved metals reacting with hydrochloric acid; and, in each case, the products were hydrogen gas and a salt of the metal. It can be predicted that most metals will react with different types of acids to form hydrogen gas and the salt of the metal - which, in fact, is the case. Now consider the four major types of chemical reactions.

TYPES OF CHEMICAL REACTIONS.

Combination: When two substances unite to form a more complicated substance, the reaction is known as a combination. In the following reactions there are two or more reactants, but only one product:

$$Fe + S \longrightarrow FeS$$
$$2\ Hg + O_2 \longrightarrow 2\ HgO$$
$$2\ H_2 + O_2 \longrightarrow 2\ H_2O$$
$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Decomposition: When a substance breaks down into two or more products, the reaction is called decomposition.

$$CaCO_3 \longrightarrow CaO + CO_2$$
$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2$$
$$2\ HgO \longrightarrow 2\ Hg + O_2$$

Single Replacement: In these reactions an element and a compound interact to form a different element and a new compound.

$Zn + H_2SO_4 \longrightarrow$ ~~$ZnSO_4 + H_2$~~

$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

Double Replacement: In these reactions the atoms or radicals from two compounds interchange to form two new compounds.

$$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$$
$$NaOH + HCl \longrightarrow NaCl + H_2O$$
$$Ba(NO_3)_2 + 2\ KCl \longrightarrow 2\ KNO_3 + BaCl_2$$

CALCULATIONS BASED UPON CHEMICAL EQUATIONS

A balanced chemical equation can be used for calculating the amount of product of a given reaction, and for other important calculations. Several of these calculations will be illustrated.

Consider the reaction that takes place in a rocket motor that used hydrazine as a fuel and dinitrogen tetroxide as an oxidizer. It has been established experimentally that molecular formulas of these two compounds are N_2H_4 and N_2O_4, respectively. Analysis of the mixture coming out of the rocket exhaust shows that it consists of gaseous elementary nitrogen (N_2) and water. Recall that to write a chemical equation one must know the formulas of the reactants and products, and that a reaction will actually take place. With these factors established, the equation can be written as follows:

$$N_2H_4 + N_2O_4 \longrightarrow N_2 + H_2O \quad \text{(unbalanced)}$$

Notice that this equation is unbalanced since the same number of atoms of each element is needed on both sides of the equation. To begin the balancing process, observe that there are four oxygen atoms on the left and one on the right. Placing a 4 in front of H_2O gives the following:

$$N_2H_4 + N_2O_4 \longrightarrow N_2 + 4\,H_2O \quad \text{(unbalanced)}$$

Notice that there are eight hydrogen atoms on the right; so placing a 2 in front of N_2H_4 will give the following:

$$2\ N_2H_4 + N_2O_4 \longrightarrow N_2 + 4\ H_2O \quad \text{(unbalanced)}$$

Finally, there are six nitrogen atoms on the left; so placing a 3 before the N_2 gives the following balanced equation:

$$2\ N_2H_4 + N_2O_4 \longrightarrow 3\ N_2 + 4\ H_2O \quad \text{(balanced)}$$

The coefficients can be considered to represent molecules; therefore, the above equation indicates that 2 molecules of N_2H will react with 1 molecule of N_2O_4 to produce 3 molecules of N_2 and 4 molecules of H_2O. Or, it can mean that 2 moles of N_2H_4 will react with 1 mole N_2O_4 to produce 3 moles of N_2 and 4 moles of H_2O. The gram molecular weights of these compounds are as follows:

$N_2H_4 = (2 \times 14) + (4 \times 1) = 28 + 4 = 32$ g
$N_2O_4 = (2 \times 14) + (4 \times 16) = 28 + 64 = 92$ g
$N_2 = 2 \times 14 = 28$ g
$H_2O = (2 \times 1) + 16 = 2 + 16 = 18$ g

The balanced equation also indicates that 64 (2 x 32) grams of N_2H_4 will react with 92 (1 x 92) grams of N_2O_4 to produce

84 (3 x 28) grams of N_2 and 72 (4 x 18) grams of H_2O. Notice that 92 + 64 = 84 + 72 = 156; therefore, the law of conservations of mass is observed.

Suppose it is necessary to determine the amount of water produced when 100 grams of N_2O_4 reacts with hydrazine. The following are steps in the calculation:

Step 1: Write the balanced equation:

$$2\ N_2H_4 + N_2O_4 \longrightarrow 3\ N_2 + 4\ H_2O$$

Step 2: Write known and unknown quantities above the formulas in the balanced equation:

$$2\ N_2H_4 + \overset{100\ g}{N_2O_4} \longrightarrow 3\ N_2 + \overset{x\ g}{4\ H_2O}$$

Step 3: Write the molecular weights below the formulas involved, remembering to multiply by the number of moles involved:

$$2\ N_2H_4 + \underset{1\ x\ 92}{\overset{100\ g}{N_2O_4}} \longrightarrow 3\ N_2 + \underset{4\ x\ 18}{\overset{x\ g}{4\ H_2O}}$$

Step 4: Set up a ratio and proportion between information above the equation and that below the equation, and solve for unknown quantity:

$$\frac{100\ g}{1 \times 92\ g} = \frac{X\ g}{4 \times 18\ g}$$

$$\frac{100\ g}{92\ g} = \frac{X\ g}{72\ g}$$

$$92\ X = 100 \times 72 = 7200$$

$$X = \frac{7200}{92} = 72\ g$$

Therefore, when 100 g of dinitrogen tetroxide react with hydrazine, 78 g of water is produced.

Another example is to calculate the weight of carbon dioxide produced by complete combustion of 10.0 lb of ethane (C_2H_6) in air. The steps are as follows:

Step 1: Write the balanced equation:

$$2\ C_2H_6 + 7\ O_2 \longrightarrow 4\ CO_2 + 6\ H_2O \quad \text{(balanced)}$$

In combustion reactions, fuel reacts with oxygen to produce carbon dioxide and water. Knowing this information allows the student to write and balance the equation above.

Step 2: Write known and unknow quantities above the formulas in the balanced equation:

$$\overset{10\ lb}{2\ C_2H_6} + 7\ O_2 \longrightarrow \overset{X\ lb}{4\ CO_2} + 6\ H_2O$$

Step 3: Write the molecular weight below the formulas involved:

$$\begin{array}{lcl} 10\ \text{lb} & & X\ \text{lb} \\ 2\ C_2H_6 + 7\ O_2 & \longrightarrow & 4\ CO_2 + H_2O \\ 2 \times 30\ \text{g} & & 4 \times 44\ \text{g} \end{array}$$

Step 4: Set up a ratio and proportion and solve for the unknown quantity:

$$\frac{10\ \text{lb}}{2 \times 30\ \text{g}} = \frac{X\ \text{lb}}{4 \times 44\ \text{g}}$$

$$60 \cdot X = 176 \times 10 = 1760$$

$$X = \frac{1760}{60} = 29.3\ \text{lb}$$

Therefore, when 10 pounds of ethane burn in air, 29.3 pounds of carbon dioxide are produced. Notice that in solving problems of this type, units do not matter as long as they are the same. Problems can be solved for grams, pounds, tons, and so forth.

GRAVIMETRIC ANALYSIS

Much work in the laboratory is directed at determining composition of materials. It may be necessary to find out what elements or compounds are present. Determining what is in a material is called <u>qualitative analysis</u>; determining how much of a given substance is present is called <u>quantitative analysis</u>. However, quantitative analysis is much more difficult and exacting than this simple description. Gravimetric analysis, volumetric analysis, and instrumental analysis are three general categories of quantitative analysis. Volumetric analysis is a widely used technique in which the exact volume of a reagent required to react with an unknown is determined by a method known as "titration." In a later module the student will conduct some acid-base titrations and consider several techniques in detail. Instrumental methods of analysis are also important. Performing gravimetric analysis is time-consuming, whereas performing volumetric and instrumental analyses is relatively fast. However, instrumental analysis methods are the least accurate, followed by volumetric and gravimetric analysis methods. In terms of cost of equipment needed, gravimetric methods are lowest and instrumental methods are highest. Since labor costs are high, initial investment in instruments may be justifiable. Many determinations, such as that of chloride, may be conducted by all three methods. In general, the degree of accuracy required determines the method to be used. A gravimetric analysis usually involves the following steps:

- Drying and then accurately weighing representative samples of material to be analyzed.
- Dissolving samples in a suitable solvent.
- Precipitating the constituent in the form of a substance of known composition by adding a suitable reagent.
- Separating the precipitate from liquid by filtration.
- Washing the precipitate to free it of contaminants.
- Drying the precipitate to a constant weight.
- Calculating percentage of the desired substance from the weight of the precipitate and sample.

As an illustration, consider the determination of chloride, as conducted in Laboratory 2 of this module. First of all, a representative sample must be obtained. A representative sample contains a true cross-section of material being sampled. (For example, to sample a carload of coal it would be easier just to take a few lumps from the top of the car; however, the dust and smaller pieces that have settled to the bottom of the car may have a different chemical composition from the hard lumps at the top. The offical method for obtaining a representative sample of coal involves 34 steps.) After a representative chloride sample is obtained, it must be dried since the moisture content can be variable and, therefore, lead to inaccurate results. Next, after weighing, the sample must be dissolved. Depending on composition, either water, acids, or bases are used to dissolve the sample. Chloride samples readily dissolve in water. Next, a reagent is added to produce a precipitate of definite composition. In the chloride determination, silver ion is added to precipitate silver chloride, as follows:

$$Ag^+ + Cl^- \longrightarrow AgCl$$

This type of equation is called an ionic equation since only ions actually involved in the reaction are shown. Nitrate or sodium ions, which may also be present, need not be shown since they are not involved in the reaction. The precipitate is separated from the liquid by filtering. Often a special type of filter paper which can be burned off in a crucible can be used. This paper, called ashless, does not leave a residue that can be weighed. In determination of chloride, special filtering crucibles must be used since silver chloride decomposes when heated in the presence of carbon (filter paper).

Suppose a chloride sample weighs 0.3230 g and that 0.5230 g of AgCl is recovered. Then the following may be written:

$$Ag^+ + \underset{35.5\text{ g}}{\overset{X\text{ g}}{Cl^-}} \longrightarrow \underset{143.5\text{ g}}{\overset{0.5230\text{ g}}{AgCl}} \qquad \text{(balanced)}$$

Step 1: $$\frac{X\text{ g}}{35.5\text{ g}} = \frac{0.5230\text{ g}}{143.5\text{ g}}$$

$$143.5\ X = 35.5 \times 0.5230$$

$$X = \frac{35.5 \times 0.5230}{143.5} = 0.129\text{ g of }Cl^-$$

Step 2: $$\%\ Cl^- = \frac{\text{g } Cl^-}{\text{g sample}} \times 100 = \frac{0.129\text{ g}}{0.3230} \times 100 = 40.0$$

One can now see how calculations based upon balanced chemical equations are used in gravimetric analysis. Steps 1 and 2 in the above calculations can be combined to give the equation used in Laboratory 2:

$$\%\ Cl^- = \frac{\frac{\text{Atomic weight Cl}}{\text{Molecular weight AgCl}} \times \text{wt AgCl}}{\text{Weight of sample}} \times 100$$

EXAMPLE F: GRAVIMETRIC ANALYSIS.

Given: The quantity of 0.6045 g of sulfate gives 0.4231 g of $BaSO_4$.

Find: The percentage of sulfate.

Solution:

$$Ba^{++} + \overset{X\ g}{SO_4^{-2}} \rightarrow \underset{233\ g}{\overset{0.4231\ g}{BaSO_4}} \qquad \text{(balanced)}$$

$$\frac{X\ g}{96\ g} = \frac{0.4231\ g}{233\ g}$$

$$X = \frac{0.4231 \times 96}{233} = 0.174\ g\ SO_4^{-2}$$

$$\%\ SO_4^{-2} = \frac{g\ SO_4^{-2}}{g\ \text{of sample}} \times 100 = \frac{0.174}{0.6045} \times 100$$

$$= 28.8\%.$$

In summary, gravimetric analysis is a widely used, accurate method of analysis, although it is time-consuming.

EXERCISES

1. Work the following problem and express the answer with the correct number of significant figures.

$$\frac{1.23}{5.043}$$

   a. 2.43
   b. 0.2439
   c. 0.244
   d. 0.240

2. Express the following number in exponential notation.

189,000

   a. 1.89×10^{5}
   b. 1.89×10^{5}
   c. 18.9×10^{5}
   d. 1.89×10^{-5}

3. Express the following number in exponential notation.

0.000453

   a. 4.53×10^{4}
   b. 4.53×10^{-4}
   c. 45.3×10^{3}
   d. 0.453×10^{-3}

4. Express the following in standard form.

4.5639×10^4

a. 4,563,900
b. 45639
c. 0.00045639
d. 456390

5. Add the following numbers and express the answer with the correct number of significant figures.

```
124.78
 56.2597
 12.1
```

a. 193.1397
b. 193.1
c. 193.14
d. 193.139

6. Using the mathematical relationship between the Fahrenheit and Celsius temperature scales, convert 30°C to degrees F.

a. 68°
b. 49°
c. 86°
d. 303°

7. An automobile engine has a displacement of 350 cubic inches. What is its displacement in liters? (1 cu ft = 28.3 liters, 1728 cu in = 1 cu ft.)

a. 12.4

b. 2.4

c. 1780

d. 5.73

8. Using the following equation, find how many grams of hydrogen sulfide will combine with 60 grams of oxygen.

$$H_2S + O_2 \longrightarrow H_2O + SO_2$$

a. 42.5

b. 63.8

c. 120

d. 23.3

9. Calculate the molecular weight of the following compound:

$$ZnCl_2$$

a. 143.5

b. 136.3

c. 100

d. 94.7

10. Calculate the molecular weight of the following compound:

$Al_2(SO_4)_3$

a. 384
b. 298
c. 368
d. 342

11. Calculate the molecular weight of the following compound:

$NaOH$

a. 23
b. 17
c. 40
d. 34

12. Determine the number of moles of selenium in 118.5 grams of selenium.

a. 118.5
b. 1.5
c. 3
d. 40

13. Calculate the gram-molecular weight of sucrose, $C_{12}H_{22}O_{11}$.

a. 320
b. 45
c. 432
d. 342

14. Calculate the number of moles in 210 grams of $MgCO_3$.

a. 2.5
b. 1
c. 3
d. 5

15. Calculate the weight of 0.6 moles of $(NH_4)_2CO_3$.

a. 44.4
b. 67.6
c. 57.6
d. 63.4

16. Balance the following chemical equations:

a. $Na + H_2O \longrightarrow NaOH + H_2$
b. $K_2SO_4 + Ba(OH)_2 \longrightarrow BaSO_4 + KOH$
c. $H_2 + N_2 \longrightarrow NH_3$

LABORATORY MATERIALS

Laboratory 1

Bunsen burner
7 test tubes
Test tube rack
Aluminum metal
Copper metal
Lead metal
Magnesium
Tin
Zinc

Laboratory 2

2-250 ml beakers
2 sintered glass crucibles
2 crucible holders
Vacuum tubing
Vacuum-filter flask
25 ml graduated cylinder
Stirring rod
Rubber policeman
Bunsen burner

Laboratory 1 (Continued)

Iron
6 M HCl
3 M H_2SO_4
3 M NaOH
3 M HNO_3

Laboratory 2 (Continued)

Wire gauze
Oven
Weighing paper
Wash bottle
50 ml beaker
Ring stand
Iron ring
Distilled water
Chloride free 6 M HNO_3
0.2 M $AgNO_3$
Unknown chloride sample

LABORATORY PROCEDURES

LABORATORY 1. REACTIVITY OF METALS.

In this experiment the student will compare the reactivity of seven metals in three different acids and in a base.

PROCEDURE

Carefully record all data in Data Table 1.

1. Obtain three small samples of each of the following metals: magnesium, aluminum, zinc, iron, tin, lead, and copper. If necessary, use sandpaper to produce a shiny, clean surface on each piece of metal.

2. Place about 5 ml of 6 M HCl in seven different test tubes in a test tube rack. Add a clean sample of each of the seven metals – one to each test tube. CAUTION! SOME MAY REACT VIGOROUSLY.
3. After each addition, watch carefully for any changes that may occur. If there is no immediate reaction, leave the tube for a few minutes while going on to the next. If there is no reaction, warm the tube gently with a Bunsen burner. Record changes in the Data Table and write the equation for the reaction.
4. Repeat the procedure for 3 M H_2SO_4 and for 3 M NaOH. Record observations in the Data Table.
5. Obtain a sample of each of the metals that did not react with the 3 M H_2SO_4 above. Repeat the procedure, using 3 M HNO_3.

LABORATORY 2. GRAVIMETRIC DETERMINATION OF CHLORIDE

An introduction to gravimetric procedures is included in the discussion section of this module. This introduction should be reviewed prior to conducting this experiment. In a gravimetric determination of chloride the following steps are involved:

- Solid chloride unknown is accurately weighed.
- Chloride sample is dissolved.
- Chloride is precipitated as solid silver chloride, which is very insoluble (only 0.0019 grams of AgCl will dissolve in one liter of water).

$$Ag^+ + Cl^- \longrightarrow AgCl$$

- Silver chloride solid is separated from the solution by filtering.
- Precipitate is washed and dried.
- Amount of AgCl produced is determined by weighing.
- Percent of chloride in the sample is calculated as follows:

$$\% \ Cl^- = \frac{\frac{\text{Atomic weight Cl}}{\text{Molecular weight AgCl}} \times \text{wt AgCl}}{\text{Weight of sample}} \times 100$$

Gravimetric determination of chloride is an accurate means of analyzing for chloride – if the experiment is carefully conducted.

PROCEDURE

Carefully record all data in Data Table 2.

1. Clean two sintered glass crucibles and place them in an oven to dry, numbering them 1 and 2. Weigh the crucibles.
2. Clean two 250-ml beakers and rinse them with distilled water. Distilled water must be used as a final rinse for all glassware used in this experiment, since tap water contains a small amount of chloride. Number the beakers 1 and 2.
3. Obtain an unknown Cl^- sample from the instructor. This sample should have been dried previously for 2 hours at 110°C. Accurately (to four decimal places – 0.0001 g) weigh 0.30 to 0.35 g samples and place them in beakers number 1 and number 2.

4. Dissolve each of the samples in 100 ml of distilled water. Avoid any splashing.
5. Add 1 ml (about 20 drops) of Cl^- free 6 M HNO_3 to each sample.
6. Transfer about 30 ml of 0.2 M $AgNO_3$ solution to a 50-ml beaker.
7. Heat sample number 1 to about 80°C. Exact temperature is not important – but keep it below boiling. Gradually add $AgNO_3$ solution to the hot solution while stirring constantly with a stirring rod. Continue to keep solution hot until the milky, cloudy solution clears, producing a lumpy precipitate. Repeat with sample number 2. Do not interchange any of the equipment (such as the stirring rod) from one sample to the other.
8. Filter precipitate while it is hot. Use filtering apparatus set-up, as shown in Figure 3. Apply suction to the flask and pour the liquid that is above the precipitate into the crucible (as much as possible). Pour liquid down the glass stirring rod (as shown in Figure 3) to prevent splashing. Gently wash precipitate into the crucible with distilled water from a wash bottle. Particles remaining on walls of the beaker should be removed by rubbing gently with a rubber policeman. During the transfer process, enough water is used to wash precipitate of ions that are present in solution.
9. Dry the filtering crucibles in an oven at 110°C for half an hour.
10. Place the crucibles in a desiccator to cool.
11. Weigh the crucibles.
12. Using Data Table 2, calculate the percentage of chloride in each sample as follows:

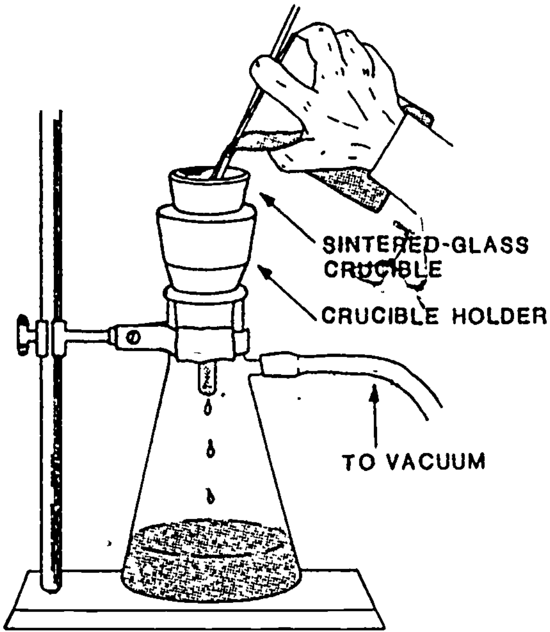

Figure 3. Experimental Set-up for Filtering Silver Chloride.

Calculate weight of precipitate in each crucible by subtracting figure (3) from figure (4). Record this figure (5). Calculate the percentage of Cl^- in each sample, using the following equation:

$$\% \ Cl^- = \frac{\text{Grams AgCl (4)} \times \frac{35.5 \text{ grams}}{143.5 \text{ grams}}}{\text{Sample weight (1)}} \times 100$$

Average the % Cl^- of the two determinations and record the answer.

An example of this calculation is as follows:

Data: 0.3230 g unknown sample

0.5230 g AgCl obtained

$$\% \text{ Cl}^- = \frac{0.5230 \times \frac{35.5}{143.5}}{0.3230} \times 100$$

$$= 40 \% \text{ Cl}^-$$

DATA TABLES

DATA TABLE 1. REACTIVITY OF METALS.

Acid	Metal	Nature of Action	Equation for Action
6 M HCl	Mg		
	Al		
	Zn		
	Fe		
	Sn		
	Pb		
	Cu	(repeat for 3 M H_2SO_4, 3 M NaOH, 3 M HNO_3)	

DATA TABLE 2. GRAVIMETRIC DETERMINATION OF CHLORIDE.

	Sample #1	Sample #2
Weight of sample (1)		
Weight of crucible + precipitate (4)		
Weight of crucible (3)		
Weight of precipitate (5)		
Percent Cl^- in sample		
Average percent Cl^- (average of sample #1 and sample #2)		

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GLOSSARY

Chemical equation: An expression utilizing formulas of elements and compounds that describe chemical reactions.

Dimensional analysis: An approach to problem solving which makes use of units.

Gram-atomic weight: The atomic weight of an element expressed in grams.

Isotopes: Atoms of the same element which have the same number of electrons and protons but different numbers of neutrons.

Mole: The amount of an element that contains the same number of atoms (as exactly 12 grams of carbon-12).

Molecular weight: The sum of the atomic weights of all the atoms shown in the formula of a compound.

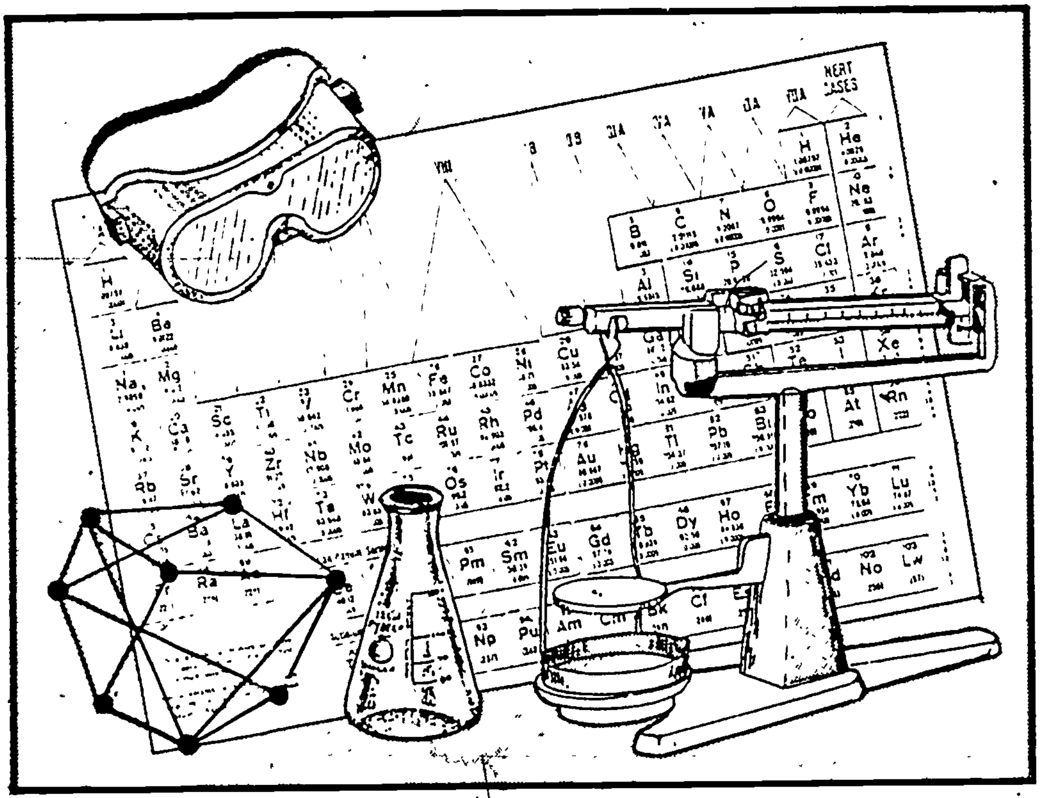
Qualitative analysis: The process of determining what is in a material.

Quantitative analysis: The process of determining how much of a given substance is present in a material.

ENERGY TECHNOLOGY

CONSERVATION AND USE

CHEMISTRY FOR ENERGY TECHNOLOGY I



MODULE CH - 04

REFRIGERATION, GASES, AND AIR POLLUTION

CORD CENTER FOR OCCUPATIONAL RESEARCH AND DEVELOPMENT

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INTRODUCTION

Air conditioning and refrigeration systems are important to the energy technician since approximately 15% of the energy consumed in homes and businesses is used for cooling. If heating is added to this figure, the consumption for space conditioning rises to approximately 50%. Thus, it is easy to see that proper maintenance and control of these systems are needed to conserve energy.

Refrigeration is the removal of heat from an object or space. A refrigeration system may be regarded as a heat pump that pumps heat from a low-temperature level to a higher-temperature level. To understand refrigeration, one must have a thorough understanding of heat and its effects. Heat and energy relationships, called thermodynamics, will be considered in a later module.

Other considerations important to the understanding of refrigeration are the properties of gases. Refrigeration systems use gases that are readily liquified when compressed. Gases have three main variables which describe their behavior: temperature, pressure, and volume. These variables, as well as other characteristics of gases and pressure measurement, are considered in this module.

Also included in this module are examples of major air pollutants and methods of sampling gases for these pollutants. Such sampling techniques are useful in monitoring and controlling gases produced in combustion of fuels.

PREREQUISITES

The student should have completed one year of high school algebra and Modules CH-01, CH-02, and CH-03 of Chemistry for Energy Technology I.

OBJECTIVES

Upon completion of this module, the student should be able to:

1. Identify general characteristics of gases, as compared with solids and liquids.
2. Use gas laws to calculate relationships among pressure, volume, and temperature.
3. Calculate absolute pressure, given the gauge pressure.
4. Define the following terms:
    a. Diffusion.
    b. Barometer.
    c. STP.
    d. Absolute temperature.
    e. ppm.
    f. Smog.
    g. Photochemical.
    h. Torr.
    i. Refrigeration.
    j. Air conditioning.
    k. Absolute pressure.
    l. psig.
    m. Relative humidity.
5. Identify important pollutants present in the atmosphere.
6. Describe a method of sampling gases.
7. List the steps in the cooling cycle of a refrigerator.

8. Identify two types of pressure gauges commonly used in the refrigeration industry.
9. Describe the determination of relative humidity, using a psychrometer.

REFRIGERATION

Refrigeration is the process of extraction of heat from a substance to lower its temperature below that of its surroundings. Refrigeration may be accomplished by cooling with water or ice, by the evaporation of a volat[illegible] liquid, or by the expansion of a compressed gas. Modern-day refrigeration systems depend upon the expansion of a compressed gas to absorb heat from an object. Refrigeration equipment can be classified in several different ways. A common classification divides the field into domestic, commercial, and industrial refrigeration. The term "air conditioning" refers to the refrigeration of air for human comfort.

The most common piece of equipment used in domestic refrigeration is the home refrigerator. Figure 1 shows the refrigeration cycle used in a home refrigerator. A gas that liquifies easily is chosen for this application, usually "Freon-12" (dichlorodifluoromethane, CCl_2F_2), which is odorless and nontoxic. The gas is pressurized by means of a compression pump powered by an electric motor. The liquification process causes the gas to heat up, so it must be cooled by means of a fan. Pressure is suddenly released when the gas is allowed to expand through a small opening (expansion valve). The liquid evaporates quickly as the pressure drops. As the liquid ~~is~~ converted back to a gas, it must absorb its heat of vaporization. This heat is drawn from the content of the refrigerator, making the content cool. The gas is then removed from the refrigerator box and is recycled by being compressed again to a liquid.

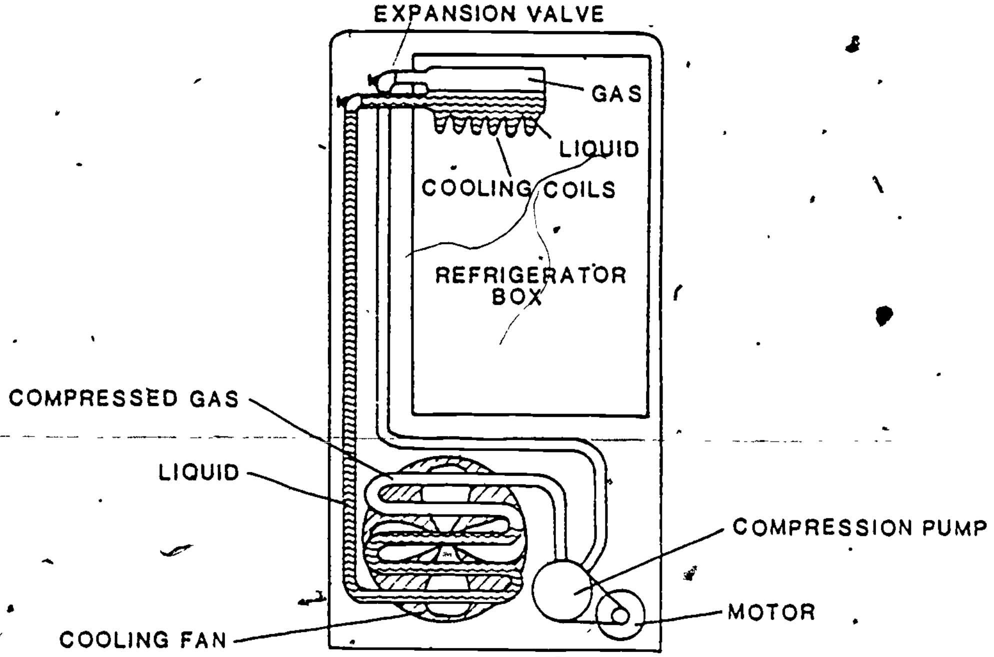


Figure 1. Refrigeration Cycle of a Home Refrigerator.

Large installations use water instead of fans to cool the gas. For this application, these "cooling towers," which allow water to run over the pipes containing the gas, are more effective than using air.

CHARACTERISTICS OF GAS

The relationships among pressure, volume, and temperature are important in the refrigeration process. These relationships and other characteristics of gases will now be considered.

Matter can be classified as solid, liquid, or gas. The molecules that make up these physical states of matter contain different physical properties. Figure 2 illustrates and Table 1 summarizes ~~the~~ following characteristics of matter:

- Solids have definite shape and volume, and they are only slightly compressible due to their tightly packed and adhering molecules.
- Liquids take the shape of their container, have a definite volume, and are slightly compressible due to the mobility of their molecules.
- Gases expand to fill the shape of their container, and they have an indefinite volume and high compressibility due to the independence and relative distance between molecules.

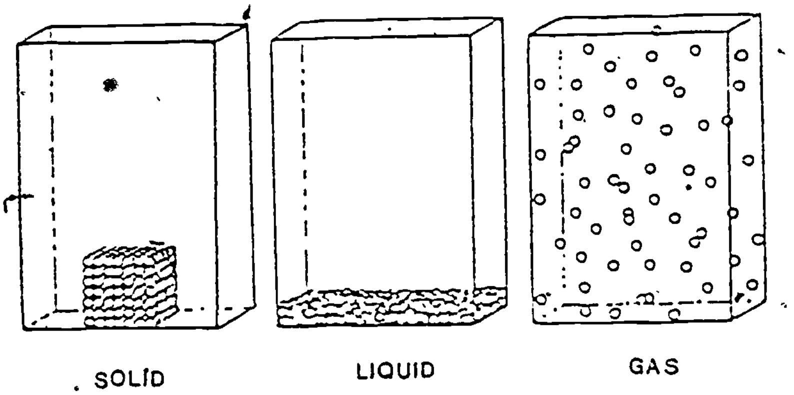


Figure 2. States of Matter.

TABLE 1. PHYSICAL PROPERTIES OF SOLIDS, LIQUIDS, AND GASES.

State	Shape	Volume	Particles	Compressibility
Solid	Definite	Definite	Closely adhering, tightly packed	Very slight
Liquid	Indefinite	Definite	Mobile, adhering	Slight
Gas	Indefinite	Indefinite	Mobile, independent, relatively far apart	High

Gas characteristics which differ from those of solids and liquids are compressibility, expansibility, and diffusibility. Gas in a container with a piston or cylinder (as in an automobile engine) may be compressed readily; however, solids and liquids are practically incompressible. Gases also expand rapidly to fill an evacuated container; solids and liquids do not. For an example of gas diffusion, consider the following: If a bottle containing mercaptain (skunk oil) is opened, the odor spreads rapidly as the mercaptain vaporizes, and the gas moves (diffuses) through the air to reach all corners of the room.

PRESSURE

The molecules of a gas are in constant motion, moving in a straight line until they collide with other molecules or the walls of the container. The collisions with the walls of the container cause or exert pressure. Measurement of pressure is important in the refrigeration industry. Pressure gauges are used to measure gas pressure inside a closed container. Two types of pressure gauges are commonly used in the refrigeration industry: (1) manometer and (2) bourdon gauge.

The manometer type gauge uses a column of liquid to measure pressure; the height of the liquid indicates the magnitude of the pressure. Mercury and water are the liquids usually used in industry. A manometer is a simple U-shaped glass tube that is open at both ends and partially filled with water or mercury. When in use, one side of the tube is connected to the vessel for which the pressure is to be measured.

The vessel pressure that is acting on one leg of the tube is opposed by the atmospheric pressure being exerted on the open leg of the tube. The difference in the heights of the two liquid columns is a measure of the difference in pressure between the pressure of the gas in the container and the pressure of the atmosphere. Figure 3 shows a manometer with a vessel pressure exceeding the atmospheric pressure by 4 inches of mercury.

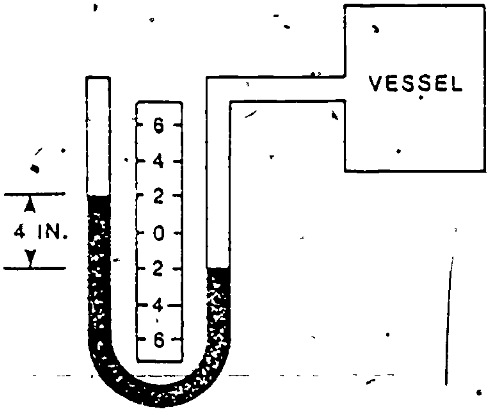


Figure 3. Simple Mercury Manometer.

Pressures below atmospheric are usually called vacuum pressures; these may also be determined with a manometer. Water manometers are especially useful for measuring small pressures. Atmospheric pressure, which will support a column of mercury 29.921 inches high, will lift a column of water to a distance of about 34 feet. A pressure of only 0.036 pounds per square inch is enough to support a column of water 1 inch high. Thus, a water manometer is much more accurate for measuring small pressures than is a mercury manometer.

Because of the excessive length of the tube required, manometer type gauges are not practical for measuring pressures above 15 pounds per square inch (psi). Manometers are normally used for measuring relatively small pressures such as those found in air ducts. Bourdon type gauges are widely used to measure the higher pressures encountered in other refrigeration work. The bourdon tube is a curved metallic tube which tends to straighten as pressure in the tube increases and to curl tighter as pressure decreases. The movement of the tube is transmitted to a pointer to indicate the magnitude of pressure. A typical bourdon type gauge is shown in Figure 4. These gauges are rugged and will measure pressures either above or below atmospheric pressure; and they are called pressure gauges or vacuum gauges, depending upon what they are designed to measure.

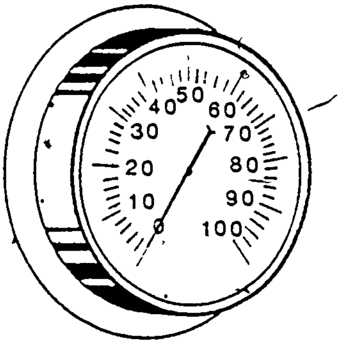


Figure 4. Bourdon Gauge.

The actual or true pressure of a gas is called the absolute pressure, whereas gauge pressure is the pressure indicated by a gauge. Gauge pressures are indicated in units of pounds per square inch gauge (psig). When gas pressure is greater than atmospheric pressure, the absolute pressure is obtained by adding atmospheric pressure to gauge pressure. When gas pressure is less than atmospheric, the absolute pressure is obtained by subtracting gauge pressure from atmospheric pressure.

EXAMPLE A: ABSOLUTE PRESSURE.

Given: A pressure gauge on a refrigerant condenser reads 110 psi.

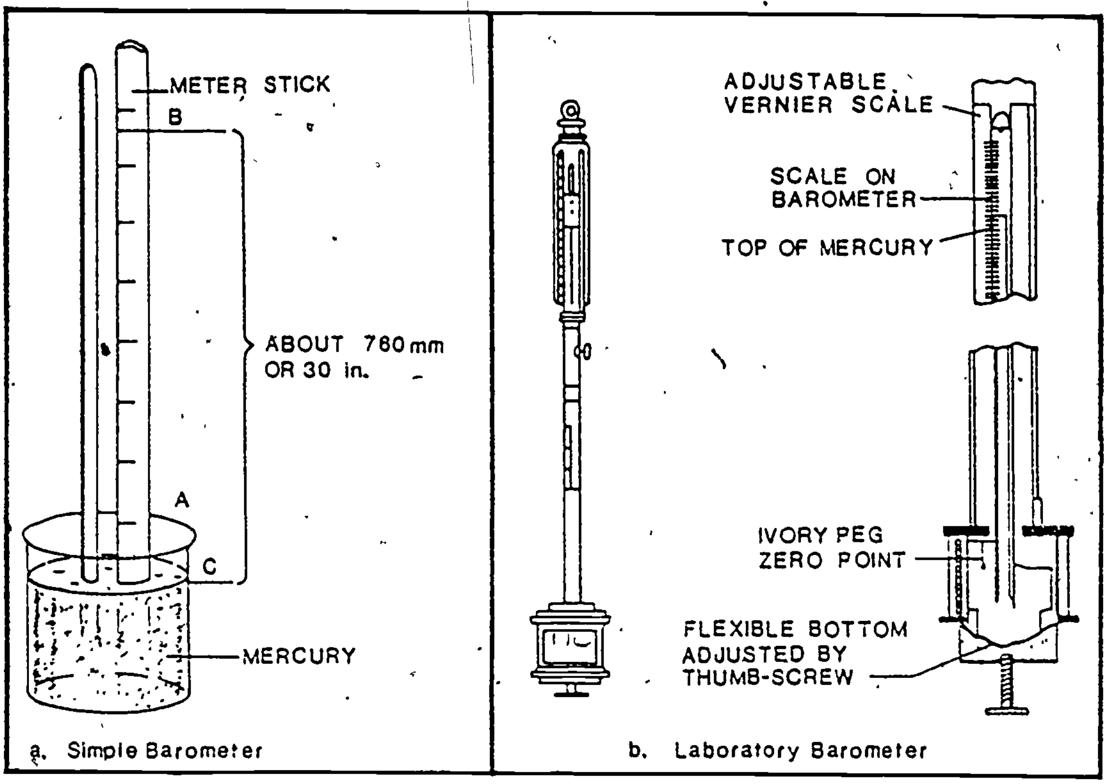
Find: The absolute pressure of the refrigerant in the condenser.

Solution:
110.0 gauge pressure, in psi
+ 14.7 atmospheric pressure, in psi
124.7 absolute pressure of refrigerant.

In addition to psi, pressure can be reported in length: millimeter (mm) and centimeter (cm). A unit called atmosphere (atm) is used for reporting high pressures. The term "torr" is also used in place of mm. The relationship of these units for atmospheric pressure is shown below:

14.7 psi = 76 cm = 760 mm = 760 torr = 1 atm = 29.92 in

Atmospheric pressure is commonly measured with a mercury barometer. Figure 5 shows two types of mercury barometers. Figure 5a is a simple barometer that can be made with the apparatus shown, and Figure 5b is a typical laboratory barometer.



a. Simple Barometer

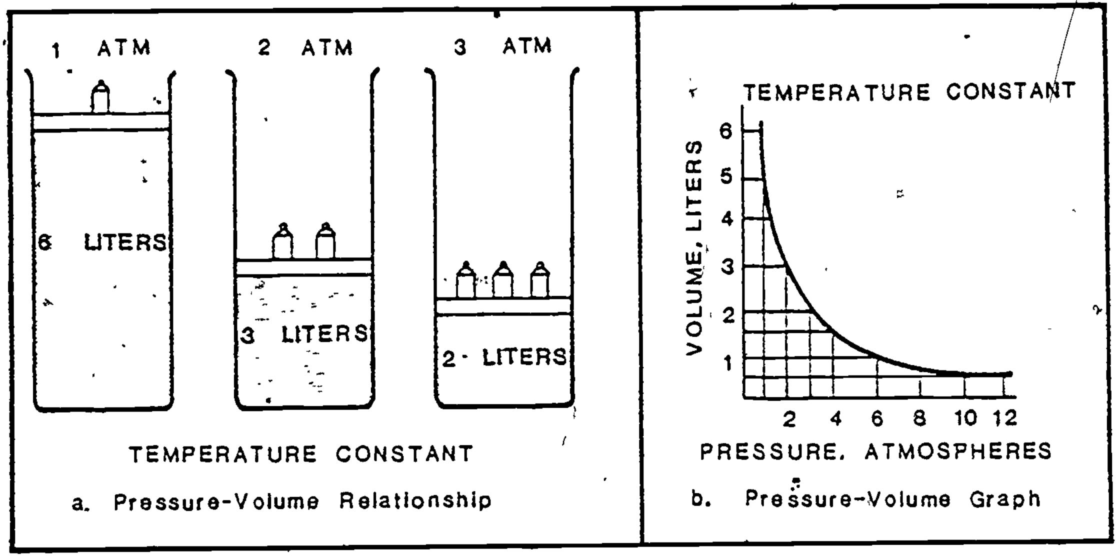
b. Laboratory Barometer

Figure 5. Mercury Barometers.

PRESSURE-VOLUME RELATIONSHIP

In 1660, Robert Boyle established the fact that the product of the pressure and the volume of a fixed quantity of gas is constant at a given temperature. The volume of gas decreases as the pressure increases, or the gas expands as pressure decreases. Boyle's law states the following: If the temperature remains constant, the volume of a given mass of gas is inversely proportional to the pressure; that is, as pressure increases, volume decreases, and as pressure decreases, volume increases.

This relationship is illustrated in Figure 6. The cylinder on the left (Figure 6a) shows the volume of gas under the pressure of 1 atm as being 6 liters. When the pressure increases to 2 atm (doubles) in the center diagram, the volume of gas decreases to 3 liters (halves). In the cylinder on the right of Figure 6a, the pressure increases further to 3 atm, and the volume decreases to 2 liters. This type of relationship – a proportional decrease in a quantity as there is an increase in another quantity – is called an inverse relationship. Figure 6b is a graph that indicates the pressure-volume relationship of a gas. According to Boyle's law, the volume can be determined if the pressure is changed but the temperature remains constant. If V_1 equals the volume at a pressure P_1, and V_2 equals the



a. Pressure-Volume Relationship

b. Pressure-Volume Graph

Figure 6. Relationship of Pressure to Volume – Boyle's Law.

volume at some other pressure P_2, then V_2 can be determined as follows:

$$\frac{P_1}{P_2} = \frac{V_2}{V_1} \qquad \text{Equation 1}$$

where:

P_1 = Pressure 1.
P_2 = Pressure 2.
V_1 = Volume 1.
V_2 = Volume 2.

EXAMPLE B: PRESSURE-VOLUME RELATIONSHIP.

Given: A quantity of gas has a volume of 93 ml when the pressure is 640 mm.

Find: The volume when the pressure has increased to 820 mm, with the temperature remaining constant.

P_1 = 640 mm $\quad V_1$ = 93 ml
P_2 = 820 mm $\quad V_2$ = ?

Solution:

$$\frac{P_1}{P_2} = \frac{V_2}{V_1}$$

$$\frac{640\text{ mm}}{820\text{ mm}} = \frac{V_2}{93\text{ ml}}$$

$$V_2 = \frac{640\text{ mm} \times 93\text{ ml}}{820\text{ mm}}$$

$$V_2 = 72.6\text{ ml.}$$

In this type of problem, any units of pressure or volume may be used as long as both pressure units and volume units are the same. For example, pressure can be measured in psi, torr, mm, cm, or atm; and volume can be measured in quarts, liters, or milliliters. Boyle's law is a mathematical expression of a common experience with gases: Pressure added to a gas compresses it.

VOLUME-TEMPERATURE RELATIONSHIP

The volume-temperature relationships given by Charles' law state the following: The volume of a gas varies directly as the absolute temperature, provided the pressure remains constant (Equation 2).

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \qquad \text{Equation 2}$$

where:

V_1 = Volume 1.
V_2 = Volume 2.
T_1 = Temperature 1.
T_2 = Temperature 2.

The temperature must be absolute temperature. <u>Absolute temperature</u> is obtained by adding 273° to the Celsium temperature, as follows:

$$°K = °C + 273$$

where:

K = Kelvin.

C = Celsius.

A comparison of the Celsius (Centigrade) and absolute (Kelvin) temperature scales is given in Figure 7. The boiling point of water is 100°C, or 373°K; 0°C is 273°K; -273°C is 0°K; and so forth.

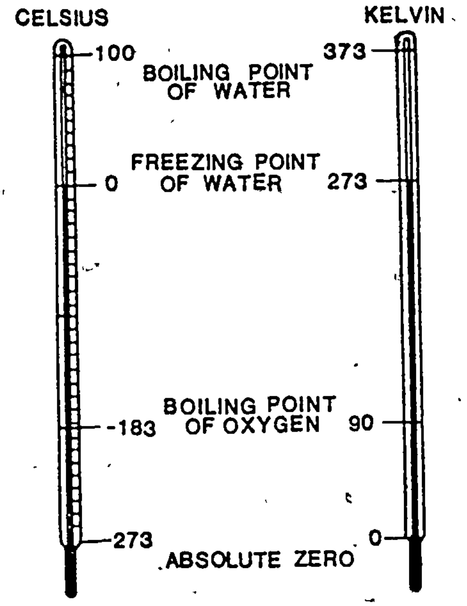


Figure 7. Comparison of Celsius and Kelvin Temperature Scales.

Charles' law is illustrated in Figure 8 as follows: A liter of gas at 100°K is heated to 200°K (double the temperature), and the volume increases to 2 liters (doubles). When the temperature is tripled, the volume triples. Charles' law is a mathematical expression of the facts that (1) gases expand when heated and (2) volume can be determined if the temperature is changed but the pressure remains constant.

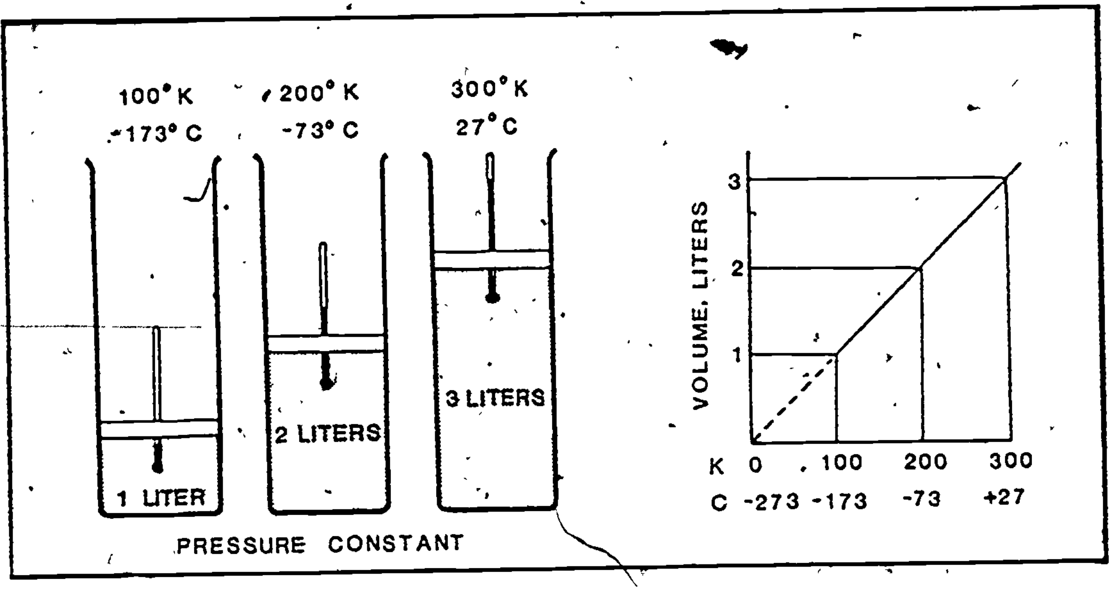


Figure 8. Relationship of Volume to Temperature — Charles' Law.

EXAMPLE C: VOLUME-TEMPERATURE RELATIONSHIP.

Given: A quantity of helium measures 100 liters at a temperature of 27°C.

Find: The volume at 100°C, assuming the pressure remains constant.

$V_1 = 100$ liters $\quad t_1 = 27°C \quad V_2 = ?$

$t_2 = 100°C \quad T_1 = 27° + 273° = 300°K$

$T_2 = 100° + 273° = 373°K$

Solution:

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}$$

$$\frac{100}{V_2} = \frac{300°K}{373°K}$$

$$V_2 = \frac{100 \text{ liters} \times 373°K}{300°K}$$

$$V_2 = 124 \text{ liters.}$$

Example C. Continued.

In this example, as the temperature is increased, the volume increases from 100 liters to 124 liters. The volume may be in any unit as long as both units are the same; but the temperature must be expressed in degrees Kelvin.

STANDARD TEMPERATURE AND PRESSURE

Because gases have an indefinite volume that varies with temperature and pressure, they must be measured and corrected to standard conditions of temperature and pressure (STP) for comparison purposes. Standard temperature is 0° centigrade, or 273° absolute; standard pressure is 760 milligrams (1 atm) of mercury.

PRESSURE-TEMPERATURE RELATIONSHIP

In the eighteenth century, the French chemist, Gay-Lussac, discovered a pressure-temperature relationship with gases. Gay-Lussac's law is stated as follows: When the volume of a fixed weight of gas is kept constant, the pressure of the gas will vary directly with the absolute temperature. The mathematical form of this law is shown in Equation 3 and illustrated in Example D.

$$\frac{P_1}{P_2} = \frac{T_1}{T_2} \qquad \text{Equation 3}$$

EXAMPLE D: PRESSURE-TEMPERATURE RELATIONSHIP.

Given: A set of automobile tires charged with 30 psi of air pressure when the temperature of the tires is 23°C. Later, the car is drive at high speeds on a hot day, and the temperature of the tires climbs to 43°C.

Find: The pressure of the tires, assuming no increase in volume.

$P_1 = 30$ psi $\qquad t_1 = 32°C$

$t_2 = 43°C \qquad P_2 = ?$

$T_1 = 23° + 273° = 296°$

$T_2 = 43° + 273° = 316°$

Solution:

$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$

$$\frac{30 \text{ psi}}{P_2 \text{ psi}} = \frac{296°}{316°}$$

$$P_2 = \frac{30 \text{ psi} \times 316°}{296°}$$

$$P_2 = 32 \text{ psi.}$$

In this example, the pressure is shown to increase from 30 to 32 psi as the temperature increases. Any pressure unit may be used, but the temperature must be degrees Kelvin.

THE COMBINED FORMULA

To this point, this module has introduced the pressure-volume relationship of Boyle's law, the volume-temperature relationship of Charles' law, and the pressure-temperature relationship of Gay-Lussac's law. Simultaneous corrections for temperature and pressure may be carried out in one mathematical operation by combining these laws into a single equation, as shown in Equation 4.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{Equation 4}$$

The combined formula applies to a given weight of gas when all three factors – pressure, volume, and temperature – change. When any five factors in the preceding equation are known, the sixth can be calculated.

EXAMPLE E: DETERMINATION OF VOLUME – COMBINED FORMULA:

Given: A gas that occupies 1000 ml at 30°C and 730 mm pressure.

Find: The volume that this same amount of gas would occupy at STP.

$P_1 = 730\ \text{mm} \qquad V_1 = 1000\ \text{ml}$

$T_1 = 30° + 273° = 303°$

$P_2 = 760\ \text{mm} \qquad V_2 = ?$

$T_2 = 0° + 273° = 273°$

Solution:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{730\ \text{mm} \times 1000\ \text{ml}}{303°} = \frac{760\ \text{mm} \times V_2}{273°}$$

$$V_2 = \frac{730 \times 1000\ \text{ml} \times 273°}{760\ \text{mm} \times 303°}$$

$$V_2 = 866\ \text{ml.}$$

IDEAL-GAS LAW

The ideal-gas law is called the "equation of state" for gases; it relates pressure, volume, temperature, molecular weight, and weight of gas in the following equation:

$$PV = \frac{gRT}{MW} \qquad \text{Equation 5}$$

where:

P = Pressure, in atmosphere.
V = Volume, in liters.
T = Absolute temperature.
g = Weight of gas, in grams.
MW = Molecular weight of the gas, in grams.
R = Gas constant equal to 0.082 liter (atm per degree per mole).

When any four variables in Equation 5 are known, the fifth can be calculated.

EXAMPLE F: DETERMINATION OF PRESSURE — IDEAL-GAS LAW.	
Given:	Sixty-four grams of oxygen contained in a one-liter container at 0°C.
Find:	The pressure in atmospher.
	$P = ?$, $V = 1$ liter $g = 64$ g
	$R = 0.082$ l-atm mole^{-1} deg^{-1}
	$MW = 32$ (oxygen)
	$T = 273° + 0° = 273°$

Example F. Continued.

Solution:

$$PV = \frac{gRT}{MW}$$

$$P = \frac{64 \times 0.082 \times 273}{32 \times 1}$$

$$P = 44.8 \text{ atm.}$$

DALTON'S LAW OF PARTIAL PRESSURE

John Dalton is responsible for a law concerning mixed gases, or an "ideal-gas" mixture, which is stated as follows: The pressure which a given gas exerts in a mixture is independent of the other gases present. The total pressure in a gaseous mixture is the sum of the partial pressures of the individual gases (Equation 6).

$$P_{Total} = P_1 + P_2 + P_3 + \ldots \qquad \text{Equation 6}$$

where:

$P_1, P_2, P_3, \ldots$ = The pressures of the individual gases.

Applications of Dalton's law are important in air conditioning, especially in connection with humidity. Moisture content (humidity) of air is closely connected to the level of comfort. In the desert, where humidity is low, relatively high temperatures can be tolerated; yet, a mild temperature, such as 80°F, can be uncomfortable when the humidity is high. Proper control of humidity in both heating and cooling systems can drastically reduce energy requirements.

It has been determined experimentally that the absolute pressure of water vapor over water (when it is in a closed container at a temperature of 60°F) is 0.26 psi. According to Dalton's law, the air pressure would be 14.44 psi. If this air pressure figure were added to the 0.26 psi (due to the vapor pressure of water), the total pressure, or atmospheric pressure, would be 14.70. If the air were not saturated with water vapor and the vapor pressure were measured (for example, 0.13 psi), the air would be said to be 50% saturated, or to have a relative humidity of 50%. Relative humidity is defined as "the ratio of the weight of water vapor in air compared with the weight of water vapor that would saturate the air." Relative humidity is temperature dependent.

The most common method of determining the relative humidity of air is to use a psychrometer. The psychrometer consists of two thermometers, one with a wet wick over the bulb (called the wet bulb) and one a bare bulb (called the dry bulb). The difference between the dry bulb and wet bulb temperatures is called the wet bulb depression. In saturated air, there is no evaporation from the wet bulb (no cooling effect) and, therefore, no wet bulb depression. The drier the air is, the greater the evaporation and wet bulb depression are. Tables are available which relate the wet bulb depression to relative humidity.

GRAHAM'S LAW OF GASEOUS DIFFUSION

Diffusion is the movement of a gas through another gas or through a porous barrier. Graham's law states that the rates of diffusion of gases are inversely proportional to the square roots of their molecular weights (or densities), as shown in Equation 7.

$$\frac{r_1}{r_2} = \sqrt{\frac{MW_2}{MW_1}} \qquad \text{Equation 7}$$

where:

r_1 = Rate of diffusion of Gas 1.
r_2 = Rate of diffusion of Gas 2.
MW_1 = Molecule weight of Gas 1.
MW_2 = Molecule weight of Gas 2.

EXAMPLE G: DETERMINATION OF DIFFUSION RATES — GRAHAM'S LAW.

Given: In the development of the atomic bomb, it was necessary to separate the low-abundant uranium isotope ^{235}U (0.7%) from the more abundant ^{238}U (99.3%). This was done by converting the uranium into a volatile compound, UF_6 (uranium hexafluoride), which boils at 56°C. The gaseous UF_6 was allowed to diffuse from one chamber into a second through a porous barrier. According to Graham's law, the gas initially appearing in the second chamber should be slightly enriched in the lighter molecule. The diffusion process was repeated thousands of times, leading to an almost complete separation of the two isotopes of uranium.

Example G. Continued.

Find: The relative diffusion rates of the uranium hexafluoride molecules.

Solution:

$$\frac{r_1}{r_2} = \sqrt{\frac{MW_2}{MW_1}}$$

$$\frac{r_{235}\text{ isotope}}{r_{238}\text{ isotope}} = \sqrt{\frac{MW_{238}\text{ isotope}}{MW_{235}\text{ isotope}}}$$

$$r = \sqrt{\frac{352}{349}}$$

$$r = 1.0043.$$

The lighter ^{235}U isotope diffuses 1.0043 times faster than the ^{238}U isotope. Hydrogen, the lightest of all gases, diffuses much faster than all other gases; for example, it diffuses four times faster than oxygen.

KINETIC THEORY OF GASES

This module has covered several characteristics and laws which apply to gases. Because molecules cannot be seen, however, the kinetic theory of gases must be relied upon to explain the behavior of gases. The elements of this theory are as follows:

1. Gases are composed of small particles called molecules.
2. Molecules are far apart; that is, the space between molecules is much greater than the size of the molecules themselves. This means that the total volume of all the molecules in a container is negligible compared to the volume of the container itself.

3. Molecules are in constant, rapid motion. It has been calculated that hydrogen gas molecules travel at a speed of about one mile per second between collisions.
4. Molecules move in a straight line until they collide with another gas molecule or the walls of the container.
5. Collisions of molecules are perfectly elastic; that is, molecules rebound with the same total energy they possessed before collision. The energy lost by one molecule in a collision is gained by the other molecule.
6. The attractive forces between molecules are very small.
7. There is a wide range of speeds in any collision of gas molecules, and increasing the temperature increases the average speed.
8. The kinetic energy is the same for all gases at a given temperature. Kinetic energy (KE) is given as follows:

$$KE = 1/2\ mv^2 \qquad \text{Equation 8}$$

where:

m = Mass of molecules.
v = Velocity of molecules.

Since kinetic energy must be the same for all gases, Equation 8 shows that small molecules must move faster than large molecules.

The kinetic theory of gases explains much of what is known about gases, such as the following:

- Expansibility: Molecules are in rapid motion and move until they collide with each other or the sides of the container. Molecules expand to fill the container.

- Compressibility: There are large distances between molecules. During compression, the molecules of a gas are pressed closer together.
- Gas Pressure: Energy imparted by the bombardment of molecules against the walls of a container causes pressure. As the temperature increases, the molecules move faster; this causes even more molecules to collide with the walls, thereby resulting in a greater increase in pressure.
- Dalton's Law: A law that is explained by the theory that molecules of one gas, while in a mixture of gases, will bombard the walls of the container just as frequently in the presence of other gas molecules as in their absence. Therefore, the total pressure of a mixture of gases will be the sum of the pressures of the individual gases.

DEVIATIONS FROM THE GAS LAWS

The gas laws studied thus far in this module are applicable for many gases at lower pressures and moderate-to-high temperatures. At high pressures and at temperatures close to the liquifying point (low temperatures) the attractive forces between the molecules become large enough that the laws are not applicable.

According to Boyle's law, for example, 2 liters of oxygen at ordinary temperatures, under a pressure of 1 atm, would be expected to occupy 1 liter at 2 atm pressure, if the temperature remains the same. However, the actual volume is 0.9991 liter at 2 atm, since the attractive forces cause the volume to be slightly lower than predicted by Boyle's law.

At high pressures, the volume of the molecules themselves becomes significant as the molecules are compressed so closely that the spaces between them disappear. For example, if 200 liters of nitrogen gas at 1 atm pressure are put under a pressure of 200 atm, the volume should become 1 liter, according to Boyle's law. In this instance, however, the volume actually becomes 1.039 liters at this given pressure. Although some formulas have been developed which take into consideration the attractive forces between molecules and the actual sizes of the molecules themselves, it can be remembered that most laboratory work is conducted at 1 atm pressure and that the gas laws are reasonably accurate without modification. However, the student should understand that gases – as they normally occur in the refrigeration cycle – are close to the saturation point and, therefore, do not follow the ideal gas laws exactly. In working with vapors in refrigeration, it is necessary to use values which have been predetermined experimentally and are tabulated in saturated and superheated vapor tables.

AIR POLLUTION

An air pollutant can be defined as "a substance added to the atmosphere in such quantities as to adversely affect humans, animals, vegetation, or material." Much pollution is of "natural" origin, such as particles from dust storms, gases from marshes, and ash and sulfur oxides from volcanoes and forest fires. Normally, the air is purified through mechanisms such as falling rain; however, even clean air contains many gases which, in larger concentrations, would be considered pollutants. Some of these gases — carbon dioxide, carbon monoxide, nitrous oxide, ozone, nitrogen dioxide, and sulfur dioxide — will be discussed later in the module.

The atmosphere is a mixture of an estimated 5500 trillion tons of gases. Approximately 99% of this total mass is below an altitude of 19 miles. The region containing most of the oxygen and moisture, called the troposphere, extends to an altitude of seven miles. Pollutants litter this area in vertical and horizontal layers, and the action of sunlight on these pollutants causes frequent chemical reactions in the atmosphere.

The Environmental Protection Agency has estimated that the cost of pollution to the United States is $25 billion each year. This is an average of $100 for each person living in this country. Although long-range effects of air pollution on materials and on the health of plants, animals, and humans are not fully known, some indications can be cited by offering the following two examples: (1) The rate of lung cancer in large metropolitan areas is twice as great as the rate in rural areas, and (2) the giant Ponderosa pine is dying from pollution coming from the Los Angeles basin more than 100 miles away.

SOURCES OF AIR POLLUTION

The major sources of air pollutants from man-controlled processes are automobiles, industry, and electric power plants. The major natural sources of air pollutants are volcanic action, forest fires, and dust storms. Table 2 shows sources and quantities of air pollutants that exist in the United States.

TABLE 2. SOURCES OF AIR POLLUTANTS IN THE UNITED STATES (1970, millions of tons).

	Carbon Monoxide	Sulfur Oxides	Hydro-carbons	Nitrogen Oxides	Partic-ulates	Total	% of Total
Transportation	111.0	1.0	19.5	11.7	0.8	144.0	51
Fuel combustion for electricity and heat	0.8	26.4	0.6	10.0	6.7	44.5	17
Industry	11.4	6.4	5.5	0.2	13.3	36.8	11
Solid Waste Disposal	7.2	0.1	2.0	0.4	1.4	11.1	1
Miscellaneous	18.3	0.2	7.3	0.5	4.0	30.3	11
	148.7	34.1	34.9	22.8	26.2	266.7	100.

The major pollutant source, transportation, causes 144 million tons of pollutants to be emitted into the atmosphere each year. The automobile accounts for 95% of this total. In fact, the automobile causes more than 50% of the total pollution in the United States.

PARTICULATES

Not all air pollutants are in the gaseous state. Small, solid particles and liquid droplets, called particulates, also are present in the air. These particulates can pose serious pollution problems. For instance, a wide range of chemical compositions found in the emissions (fly ash) of coal-burning units consist of carbon (principally) and compounds of iron, magnesium, calcium, aluminum, sulfur, silicon, phosphorous, potassium, and sodium.

Table 2 shows that more than 26 million tons of particulates enter the air in the United States each year. These particles present an additional danger in that they possess a large surface area that absorbs moisture and toxic gas, providing a reaction site for a variety of chemical reactions. Sulfur dioxide, nitrogen oxides, hydrocarbons, and carbon monoxide do their greatest damage when concentrated on the surface of particulates.

Particles are removed naturally from the atmosphere by gravitational settling and through absorption by rain and snow. Industrial emissions can be prevented from entering the atmosphere by means of filtration, centrifugal separation, spraying, electrostatic precipitation, and ultrasonic vibration. Electrostatic precipitators, which are better than 98% effective in removing particulates from exhaust gases, use high voltage (about 50,000 volts) to attract particulates. The basis of this lies in the fact that particles that are moving through a region of high electrostatic potential tend to become charged electrically and to become ~~attracted~~ then to an oppositely charged area where they can be collected. The precipitated solid falls to the bottom of the collector and can be removed. A typical electrostatic precipitator is illustrated in Figure 9.

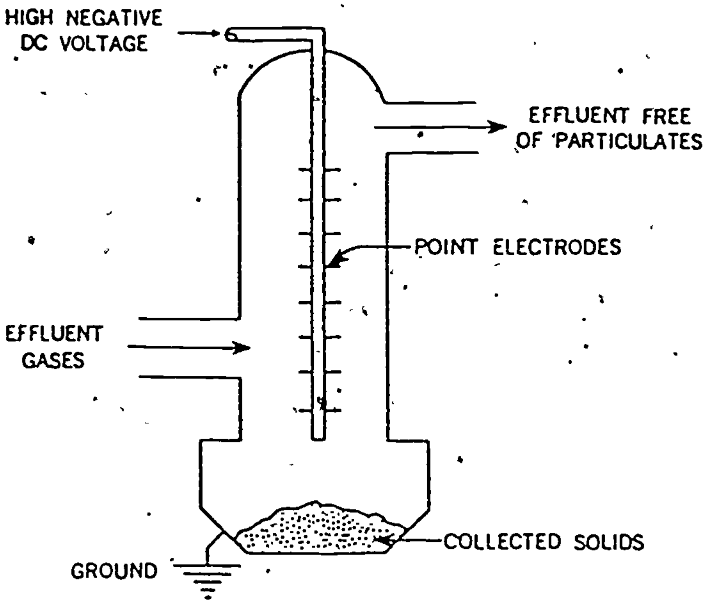


Figure 9. Electrostatic Precipitator.

The following is a description of other methods used in removing contaminants: (1) Wet scrubber devices assist in removing solid, liquid, and gaseous contaminants by spraying water into the gas stream; (2) glass fiber, or silicon treated textile bags, are used to filter hot exhaust gases; (3) filtration uses felts, wool, acrylics, nylon, and cotton for cold air streams; (4) centrifugal separators whirl gases and sling particles against a wall where they collect for removal; and (5) ultrasonic devices produce vibrations which cause increased collisions of particles and allow them to adhere to one another and form large particles that settle out of an air stream by gravity.

SMOG

Smog, a poisonous mixture of smoke, fog, air, and other chemicals, was first referred to in a 1911 report on the London air disaster as the cause of 1150 deaths.

Two general types of smog have been identified as follows: (1) The London type, which results from the combustion of coal, and (2) the chemically oxidizing type, which is typical of Los Angeles. The London type of smog contains sulfur dioxide that is mixed with soot, fly ash, smoke, and partially oxidized organic chemicals. Although this type of smog has been decreasing in seriousness, the increased burning of coal (especially the high sulfur-contaminating kind) has caused it to become a potentially serious problem again. The second type of smog, called photochemical smog because of the importance of sunlight in the photochemical process of its formation, has very little sulfur dioxide; but it contains large amounts of ozone, nitrogen oxides, and hydrocarbons.

Geographical and weather conditions are important to the formation of smog. For example, a period of windlessness allows pollutants to collect and smog generation to accelerate. Mountains also hinder the horizontal movement of air and cause smog generation.

SULFUR OXIDES

The combustion of any sulfur-containing material will produce sulfur dioxides according to the following reaction:

$$S + O_2 \longrightarrow SO_2$$

Most sulfur dioxide in the atmosphere comes from sulfur-containing coal, oil burned in electrical-power generating plants, and smelting plants treating metal sulfide ores. The average sulfur content of coal in the United States is about 2%, although most coal has either less than 1% or more than 3% sulfur. Sulfur dioxide can be oxidized in the air to SO_3; and this can combine then with moisture to form sulfuric acid, as follows:

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

The oxides of sulfur, along with sulfuric acid, are responsible for costly corrosion of metals, vegetation damage, and respiratory problems in animals and humans. Two-thirds of the sulfur-oxide pollution is from natural sources such as volcanoes. This natural pollution causes few problems since it is so widespread. The majority of volcanic action occurs in the South Pacific, which is sparsely populated.

In contrast, man-made pollution through sulfur oxides is highly concentrated in the industrial areas of the world. For example, more than 50% of the SO_2 production in the United States occurs in only the following seven states: New York, Pennsylvania, Michigan, Illinois, Indiana, Ohio, and Kentucky. As shown in Table 2, three-fourths of the total sulfur dioxide emitted (26.4 million tons) is emitted during fuel combustion for producing electricity and heat.

NITROGEN OXIDES

Of the eight nitrogen oxides present in the atmosphere, dinitrogen oxide (N_2O) and nitrogen dioxide (NO_2) are the most important. In combustion processes, the nitrogen in air combines with oxygen to form nitrogen oxide:

$$N_2 + O_2 \text{ heat} \longrightarrow 2\ NO$$

Nitrogen oxide is formed in a similar manner during electrical storms. In the atmosphere, NO reacts with atmospheric oxygen to produce nitrogen dioxide:

$$2\ NO + O_2 \longrightarrow 2\ NO_2$$

Smog formation is initiated by the photochemical decomposition of nitrogen dioxide:

$$NO_2 \longrightarrow NO + O$$

The oxygen that atoms produce can react with oxygen to produce ozone, a major component of photochemical smog:

$$O + O_2 \longrightarrow O_3$$

Ozone molecules are reactive and can attack hydrocarbon molecules to produce formaldehyde, acrolein, and organic nitrates. These compounds produce the eye irritation associated with photochemical smog.

CARBON MONOXIDE

The most abundant and widely distributed air pollutant, carbon monoxide, is produced by the combustion of carbon-containing compounds in an insufficient amount of oxygen. The principal source of carbon monoxide is the combustion of fuels in automobiles. Octane, a major component of gasoline, produces carbon dioxide and water when burned with sufficient amounts of oxygen:

$$2\ C_8H_{18} + 25\ O_2 \longrightarrow 16\ CO_2 + 18\ H_2O$$

However, if the fuel mixture is too "rich" (contains too much fuel and too little air), carbon monoxide may be formed:

$$2\ C_8H_{18} + 23\ O_2 \longrightarrow 12\ CO_2 + 4\ CO + 18\ H_2O$$

Carbon monoxide, when taken into the lungs, reduces the ability of the blood to transport oxygen throughout the body. An exposure to less than 10 ppm (parts per million) can cause a reduction in mental and physical reaction times in humans —

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which is a probable cause for many automobile accidents. A short exposure at 150 ppm can produce fatal results.

HYDROCARBONS

Hydrocarbons are compounds which contain carbon and hydrogen. In addition, they may contain oxygen, nitrogen, halogen, or sulfur. The hydrocarbons come in all shapes and sizes, beginning with methane, CH_4, and ranging to molecules containing many carbon atoms. As with most pollutants, much of the world's contribution of hydrocarbons to the atmosphere comes from natural sources. These sources – including emissions from trees and plants in the form of turpentine and pine oil, and methane from the bacterial decomposition of organic matter – account for 85% of the total emission of hydrocarbons.

The 15% of hydrocarbon emission that people contribute is concentrated in urban areas and comes from incomplete combustion of coal, petroleum, and gasoline. The automobile accounts for more than two-thirds of the man-made emissions of hydrocarbons.

The incidence of lung cancer in non-smokers closely parallel the concentration of hydrocarbons in the air. For example, the concentration of hydrocarbons in urban areas is 10 times that in rural areas, and the incidence of lung cancer is nine times higher in urban areas. As indicated earlier, hydrocarbons can react with nitrogen oxide or with ozone to produce the eye irritants found in smog.

OZONE

Ozone is a natural component of the upper atmosphere. It forms a protective layer which absorbs much of the ultraviolet radiation that strikes the earth's atmosphere. Ozone is formed naturally by the combination of an oxygen atom and an oxygen molecule in the presence of an electrical discharge such as lightning:

$$O_2 + O \xrightarrow{\text{Electrical discharge}} O_3$$

Ozone is a powerful, oxidizing agent. It causes rubber to harden and crack, and, by shortening the life of automobile tires and other rubber products, it is responsible for the loss of millions of dollars each year. Ozone also causes extensive damage to vegetation: tobacco and tomatoes are particularly susceptible.

When high levels of ozone are present, vigorous physical activity is dangerous. Public schools in Los Angeles cancel all outdoor recreation classes when the ozone concentration in that area reaches 0.35 ppm. At this level, ozone also is responsible for several physical symptoms occurring in humans, such as respiration irritation, difficulty in maintaining proper body temperature, and vision impairment (especially night vision).

The major pollutants are classified in Table 3.

TABLE 3. CLASSIFICATION OF POLLUTANTS.

Major Classes	Sub-Classes	Typical Members
Particulate	Solid particulates	Dust, smoke, ash
	Liquid particulates	Spray, mist
Organic Gases	Hydrocarbons	Ethylene, methane, hexane
	Aldehydes and ketones	Acetone, formaldehyde
	Other organics	Alcohols, chlorinated hydrocarbons
Inorganic Gases	Oxides of carbon	Carbon monoxide, carbon dioxide
	Oxides of sulfur	Sulfur dioxide, sulfur trioxide
	Oxides of nitrogen	Nitrogen oxide, nitrogen dioxide
	Other inorganics	Ammonia, chlorine, hydrogen sulfide
Oxidants		Ozone

SAMPLING OF POLLUTANTS

Most work in chemical laboratories is done with solid or liquid samples because of the convenience in handling. Sampling of pollutants is more difficult because pollutants generally are found in gaseous ~~form~~ and in small concentrations. There are several techniques, however, that can be used in sampling for particulates: the dustfall bucket, the high volume sampler, and the tape sampler.

The dustfall bucket is a cylinder container used to collect dust and large particulates for a standard period of 30 days. Water is usually added to help retain the dust. For analysis, the dust and water are transferred to a beaker, the sample is dried, and the total weight of the material collected is determined. Results are usually reported as tons per square mile, per 30 days, or as milligrams per square centimeter, per 30 days. Determinations also may be conducted for carbon, sulfates, chlorides, ammonia, metals, nitrates, phosphates, and insoluble ash.

A high-volume sampler is used on particulates too small to settle out of the air. As a high volume of air is pulled through a filter, the particulates are trapped in the filter. The results are determined by measuring the volume of air and the gain in weight of the filter. The results are reported as micrograms per cubic meter of air. Particulates on the fiberglass filter may be analyzed for carbon, sulfates, chlorides, metals, and so forth.

The tape sampler consists of a vacuum pump, an automatic timer, a cellulose tape filter, and an optical sensor that measures the visibility quality of the ambient atmosphere. As many as 600 sample spots may be collected automatically on 100-foot rolls of filter tape with a minimum of operator attention. The tape is advanced automatically by the tape sampler, which has a fast response time and continually monitors the atmosphere.

Sampling methods for collecting gaseous pollutants include absorption in a liquid, absorption on various solids, condensing or freezing pollutants, and grab sampling. In the first method, gas is drawn through a solution which reacts with the gaseous contaminant to form a nongaseous compound. Then the gas is separated from the liquid and analyzed. The absorption devices use fritted-glass absorbers, or impingers, to disburse the gas evenly through the absorbing liquid.

When solid granular materials are used to absorb pollutant gases, the gas is pulled through a tube containing the granular absorbant. The quantity that can be absorbed by a given amount of absorbant depends upon the surface area of the absorbant, the temperature of the system, the chemical reactivity of the pollutant, and the size and shape of the pores of the absorbing medium. Common absorbants have the capacity to absorb 10 to 40 percent of their weight.

In the freezing or condensation method of collecting gaseous pollutants, the gas stream is passed through a U-tube which is immersed in a cold-bath. As the pollutant comes in contact with the cold surface of the U-tube, it is frozen to the surface. The efficiency of collection can be improved by reducing the flow rate. Some typical cold-baths are listed in Table 4.

TABLE 4. TYPICAL COLD-BATHS.

Substance	Temperature
Ice and salt	-16°C
Dry ice and acetone	-80°C
Liquid nitrogen	-195°C

Grab samples are taken at a particular time — usually a few seconds to a minute of time is involved. An evacuated round-bottom flask is connected to a probe that has been inserted in the area to be sampled (such as an exhaust duct). The stopcock on the flask is opened to collect the sample, and the sample rushes into the vacuum in the flask. A flexible bag made of Mylar is used for collecting combustion gases. For this process, a sampling train (Figure 10) is set up to take the sample. The air is evacuated from the grap sample box, causing the flexible bag to inflate; the sample is then drawn out of the duct. Reversing the flow of air from the pump forces air into the box. One can expel the sample from the bag for analysis.

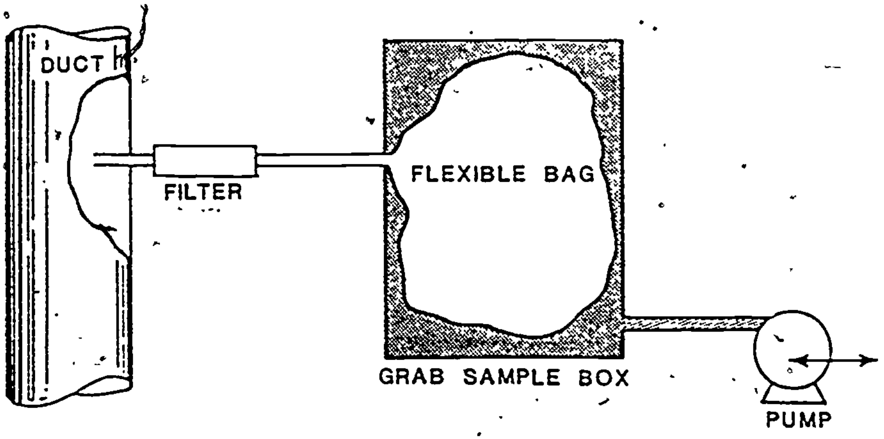

Figure 10. Grab Sampling with a Flexible Bag.

EXERCISES

1. A gas has a volume of 1000 ml at a pressure of 1 atm. What pressure would be required to reduce the volume of the gas to 750 ml without any change in temperature?
   a. 0.75
   b. 1.33
   c. 1.01
   d. None of the above
2. If 100 ml of gas at 23°C were cooled to 0°C, what would be the new volume?
   a. 100 ml
   b. 78
   c. 92
   d. None of the above
3. Gas is placed in a storage tank at a pressure of 50 atm at 23°C. There is a small safety plug in the tank made of metal alloy that melts at an elevated temperature. At this temperature, the gas pressure has reached 75 atm. At what temperature is the metal plug designed to melt?
   a. 171°C
   b. 444°C
   c. 258°C
   d. None of the above
4. What would be the weight of 5 liters of oxygen at 20°C and 740 mm pressure?
   a. 20.7
   b. 6.5
   c. 11.4
   d. None of the above

5. A mixture of 20% oxygen, 30% nitrogen, and 50% helium exerts a total pressure of 700 mm. What is the partial pressure exerted by the nitrogen?
   a. 210
   b. 140
   c. 350
   d. None of the above

LABORATORY MATERIALS

Laboratory 1

400 ml beaker
#4 one-holed rubber stopper
125 ml Erlenmeyer flask
Pneumatic trough
Ice
Thermometer
5-cm length of glass tubing
50 ml graduated cylinder
Tongs
Burner

Laboratory 2

125 ml Erlenmeyer flask
25 cm square of aluminum foil
400 ml beaker
Burner
Ring stand and ring
Wire gauze

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20 ml of an unknown organic liquid – any one of the following: benzene, carbon tetrachloride, ethyl acetate, cyclohexane, or di-chloroethene

Analytical balance

LABORATORY PROCEDURES

LABORATORY 1. DETERMINATION OF ABSOLUTE ZERO.

This experiment investigates the relationship between gas volume and temperature. The student graphs the volume-temperature data obtained and estimates absolute zero by the method of extrapolation. This determination of absolute zero is based on the direct relationship between temperature and volume: as the temperature is reduced, the volume is reduced, and, theoretically, the volume at absolute zero is zero.

PROCEDURE

Carefully record all data in Data Table 1.

1. Fit a 125 ml Erlenmeyer flask with a #4 one-holed stopper containing a short length of glass tubing.
2. Determine the exact volume of the flask with the stopper by filling it with water and then measuring the water in a graduated cylinder. Record this volume in Data Table 1.

3. Dry the flask thoroughly by carefully holding it with tongs and rotating it over the burner flame. It is important for the flask to be absolutely dry to avoid having water vapor in the air.
4. Place the flask in a 400 ml beaker of boiling water for five minutes to allow the temperature of the water and the air in the flask to become equal. Record the temperature of the water in Data Table 1.
5. Place a finger tightly over the glass tubing and transfer the flask to a pneumatic trough containing an ice and water mixture. Release the finger while the tube is under water, and allow the water to rush into the flask. Allow three minutes for equilibrium to be established.
6. Measure the volume of water entering the flask with a graduated cylinder.
7. Calculate the final volume of air by subtracting the volume of water from the total volume of the flask found in Step 2. Record the final volume in Data Table 1.
8. Measure and record the temperature of the ice bath in Data Table 1.
9. Conduct another run with the ice-water bath; then conduct two runs each with water baths of 20°C, 40°C, and 60°C. Be certain to dry the flask thoroughly between each run.
10. Graph all eight data points on the graph in Data Table 1. Draw a single straight line through the data points until the line crosses the temperature axis of the graph.

11. Read the value obtained for absolute zero (the temperature where the extrapolated line crosses the temperature axis). Record this value in Data Table 1.

LABORATORY 2. MOLECULAR WEIGHT OF A VAPOR.

The ideal-gas law relates the pressure, volume, temperature, molecular weight, and weight of a gas. For convenience, the ideal-gas law formula (Equation ~~5) is~~ repeated below as follows:

$$PV = \frac{gRT}{MW}$$

where:

P = Pressure, in atmosphere.
V = Volume, in liters.
T = Absolute temperature.
g = Weight of gas, in grams.
MW = Molecular weight of the gas, in grams.
R = Gas constant equal to 0.082 liter (atm per degree per mole).

Molecular weight is given as follows:

$$MW = \frac{gRT}{PV}$$

In this experiment, a small amount of the unknown liquid is put into a weighed flask. The flask is then placed in boiling water where the liquid will vaporize completely, driving out the air and filling the flask with vapor only. The barometric pressure and temperature of the boiling water are measured. The flask is cooled, thereby condensing the vapor. Then the flask is weighed (plus the condensed vapor) to determine the weight of the vapor. The molecular weight of the vapor can be calculated at this point. The experimental set-up is shown in Figure 11.

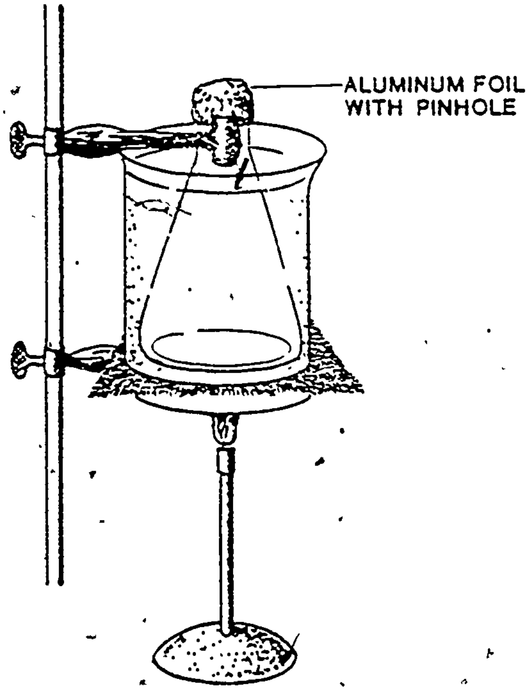


Figure 11. Apparatus for Determining Molecular Weight of Vapor.

PROCEDURE

Carefully record all data in Data Table 2.

1. Make a cap for the 125 ml flask from aluminum foil. Using a pin, poke a small hole (as small as possible) in the center of the cap.
2. On an analytical balance, determine the combined weight of the clean, dry flask and its cap to the nearest 0.0001 gram. Record the weight in Data Table 2.
3. Pour approximately 2 ml of the unknown liquid into the flask and replace the cap securely. Clamp the flask and suspend it in a 600 ml beaker, as shown in Figure 11.
4. Fill the beaker almost full of water and heat it to boiling.
5. Measure the temperature of the boiling water and the barometric pressure. Record these measurements in Data Table 2.
6. As soon as the liquid has evaporated, remove the flask and allow it to cool.
7. Weigh the flask, cap, and condensed unknown liquid.
8. Fill the flask completely full of water and measure its volume by pouring the water into a graduated cylinder.
9. Record data in the calculations section of Data Table 2 by converting the experimental data to the appropriate units.
10. Substitute data to the equation shown in Data Table 2 and calculate the molecular weight of the unknown vapor. Compare this result with the actual molecular weight given by the instructor.

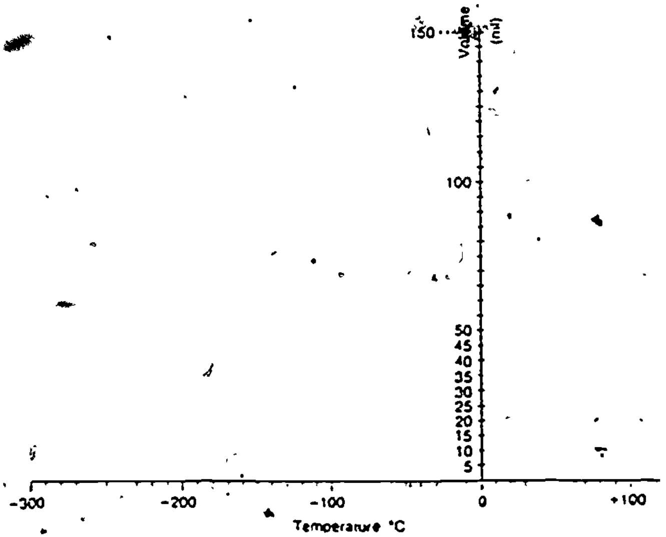
DATA TABLES

DATA TABLE 1. DETERMINATION OF ABSOLUTE ZERO.

Total volume of flask ______ ml

Initial temperature of air (boiling water) ______ °C

	Vol of water (ml) entering flask	Final volume of air (ml)	Final temp of air °C
First run (ice water bath)			
Second run (ice water bath)			
Third run (20°C water bath)			
Fourth run (20°C water bath)			
Fifth run (40°C water bath)			
Sixth run (40°C water bath)			
Seventh run (60°C water bath)			
Eighth run (60°C water bath)			
ABSOLUTE ZERO			



DATA TABLE 2. MOLECULAR WEIGHT OF A VAPOR.

	Trial 1	Trial 2
Weight of flask and aluminum foil cap	g	g
Temperature of boiling water bath	°C	°C
Barometric pressure	torr	torr
Weight of flask, cap, and condensed vapor	g	g
Volume of flask	ml	ml

CALCULATIONS			
Pressure of vapor	P	atm	atm
Volume of vapor (flask)	V	liter	liter
Temperature of vapor	T	°K	°K
Weight of vapor	g	g	g
Gram molecular weight of unknown	MW	g	g

$$MW = \frac{gRT}{PV}$$

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GLOSSARY

Air conditioning: The refrigeration of air for human comfort.

Air pollutant: A substance added to the atmosphere in such quantities as to adversely affect humans, animals, vegetation, or material.

Absolute pressure: The actual or true pressure of a gas.

Absolute temperature scale: A temperature scale obtained by adding 273 degrees to the Celsius temperature.

Boyle's law: At constant temperature, the volume of a given mass of gas is inversely proportional to the pressure.

Charles' law: At constant pressure, the volume of a gas varies directly as the absolute temperature.

Diffusion: The movement of a gas through another gas or through a porous barrier.

Electrostatic precipitator: A device which removes particulates from exhaust gases through electrostatic attraction.

Gauge pressure: The pressure of a gas as indicated by a gauge.

Gay-Lussac's law: At constant volume, the pressure of a fixed amount of gas will vary directly with the absolute temperature.

Particulates: Small, solid particles or liquid droplets which, if present in air, can cause pollution.

Photochemical smog: A poisonous mixture of smoke, fog, air, and other chemicals containing large amounts of ozone, nitrogen oxides, and hydrocarbons, formed by the action of sunlight.

Pressure: The force caused by the collision of gas molecules on the walls of a container.

Refrigeration: The process of extraction of heat from a substance to lower its temperature from that of its surroundings.

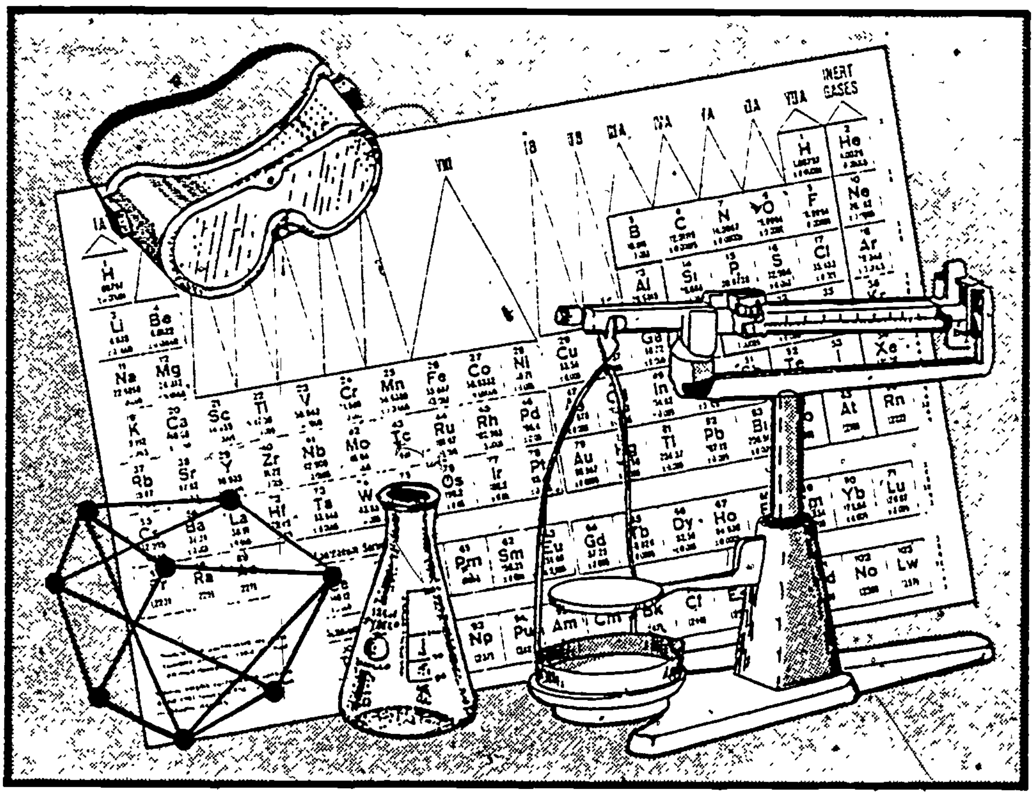
Relative humidity: The ratio of the weight of water vapor in air compared with the weight of water vapor that would saturate the air.

Smog: A poisonous mixture of smoke, fog, air, and other chemicals.

ENERGY TECHNOLOGY

CONSERVATION AND USE

CHEMISTRY FOR ENERGY TECHNOLOGY I



MODULE CH - 05

SOLUTIONS

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INTRODUCTION

Most reactions are conducted in the liquid phase. Generally, the liquids involved are solutions, which are homogeneous mixtures of two or more substances. Of primary importance are those solutions formed when a solid dissolves in water. This module presents a study of unsaturated, saturated, and supersaturated solutions, as well as solubility and rate of dissolution. Considered, also, is a variety of methods for expressing concentration of solutions — namely, percent concentration, molarity, normality, and density. The procedure for making dilutions is also presented.

A number of basic properties of liquids are discussed in this module, including evaporation, condensation, vapor pressure, boiling and freezing points, and viscosity. Properties of acids, bases, and salts are examined, as well as an important laboratory technique called titrimetric analysis. A related topic, the measurement of hydrogen ion concentration (pH) and the measurement of the concentration of other ions also is included.

PREREQUISITES

The student should have completed one year of high school algebra and Modules CH-01, CH-02, CH-03, and CH-04 of Chemistry for Energy Technology I.

OBJECTIVES

Upon completion of this module, the student should be able to:

1. Define the following terms:
    a. Boiling point.
    b. Vapor pressure.
    c. Molarity.
    d. Normality.
    e. Solvent.
    f. Solute.
    g. Unsaturated solution.
    h. Saturated solution.
    i. Supersaturated solution.
2. Distinguish between dilute, concentrated, weak, and strong solutions.
3. Predict the effect of solute concentration on vapor pressure, freezing point, and boiling point.
4. Discuss the effects of pressure and temperature on solubilities.
5. Solve problems concerning the concentration of solutions.
6. Distinguish between suspensions, colloids, and solutions.
7. Define acid and base, and give an example of each.
8. Describe the method of conducting a titrimetric analysis.
9. Explain what is meant by pH, and give the mathematical expression for pH.

THE LIQUID STATE

In a previous module, "Refrigeration, Gases, and Air Pollution," it was learned that molecules of a substance in the gaseous state are in constant, rapid motion, and spaces between the molecules are large compared to the size of the molecules themselves. The kinetic molecular theory of gases, which applies to the behavior of gases, also applies to liquids in the following manner:

As molecules of a gas are brought closer together by application of pressure, the attractive forces between the molecules become stronger. Moreover, as temperature decreases, velocity of the molecules decreases, allowing even more attraction between the molecules. If the pressure is high enough and temperature low enough, attractive forces then become sufficient to cause the gas to condense to a liquid. In the liquid state, molecules are free to flow and take the shape of a container; but the molecules are somewhat close together and only slightly compressible.

EVAPORATION AND CONDENSATION

When a liquid, such as water, is placed in an open container, the water will gradually disappear or evaporate. Evaporation can be explained in terms of motion of molecules in liquid: At a given time, molecules of a liquid move at various speeds; some move slowly, some move at intermediate speeds, and some move very rapidly. A molecule that is moving rapidly near the surface of the liquid may possess sufficient kinetic energy (energy due to motion) to overcome the attraction of its neighboring molecules; that is, it may escape from the surface, or evaporate.

Evaporation of a liquid is illustrated in Figure 1.

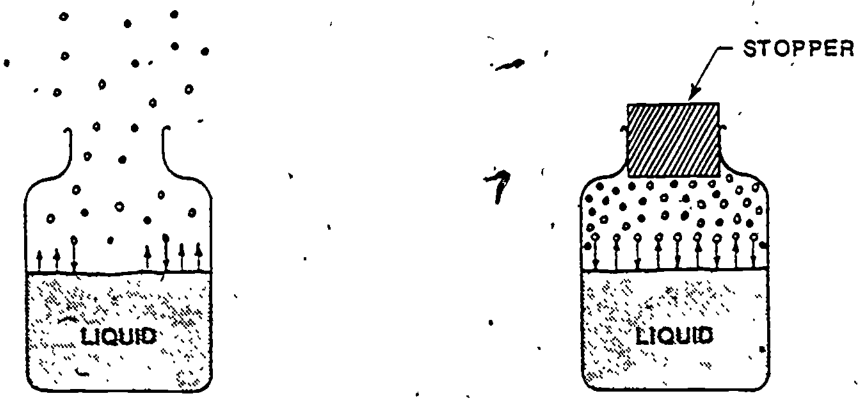


Figure 1. Evaporation of a Liquid.

In the open container on the left, molecules in the liquid gained sufficient energy and speed to escape from the surface of the liquid. In this gaseous phase, the molecules moved about rapidly and found their way out of the container. Only a few gaseous molecules struck the surface of the liquid and were captured and returned to the liquid phase. When a stopper was placed in the bottle, as shown on the right of Figure 1, an equilibrium was established in which the number of molecules leaving the liquid was exactly the same as the number of molecules re-entering the liquid. In other words, the rate of evaporation equaled the rate of condensation. Note that evaporation was not stopped, but that condensation equaled evaporation, and there was no net change. Equilibrium can be shown as follows:

$$\text{Liquid} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftarrows}} \text{Vapor}$$

VAPOR PRESSURE

Molecules in vapor above a liquid in a closed container exert pressure as they strike the sides of the container. Pressure exerted by a vapor in equilibrium, with its liquid at a given temperature, is called <u>vapor pressure</u> of the liquid. This vapor pressure of a liquid varies considerably with changing temperature; as the temperature of the liquid increases, more molecules leave the liquid and enter the gaseous phase, where they may exert an increased pressure. Table 1 lists vapor pressures of water, alcohol, and ether at various temperatures.

TABLE 1. VAPOR PRESSURES OF SOME COMMON SUBSTANCES AT VARIOUS TEMPERATURES (pressures in mm of mercury).

Substance	0°C	20°C	40°C	60°C	80°C	100°C
Water	4.6	17.5	55.0	149.2	355.5	760.0
Alcohol	12.2	43.9	135.3	352.7	812.6	1693.3
Ether	185.3	442.2	921.1	1730.0	2993.6	4859.4

The table shows that the vapor pressure of all three substances increases with increasing temperature. Ether has a much higher vapor pressure than either water or alcohol since water and alcohol molecules attract each other through hydrogen bonding. Thus, water and alcohol molecules cannot easily leave the liquid surface.

When vapor pressure of a liquid becomes equal to atmospheric pressure, the liquid boils. The "normal" boiling point of a liquid is defined as "the temperature at which the vapor pressure becomes equal to 760 mm of mercury." Table 1 shows that the pressure of water at 100°C is 760 mm of mercury; that is, the boiling point of water is 100°C. Likewise, the vapor pressure of alcohol becomes 760 mm at a temperature below 80°C, and the vapor pressure of ether becomes 760 mm at a temperature below 40°C. By plotting vapor pressure versus temperature, one can obtain a vapor pressure curve which can be used to determine the exact boiling point. Vapor pressure curves for water, ether, and alcohol are shown in Figure 2. Ether's normal boiling point is 34.6°C; alcohol's is 78.4°C.

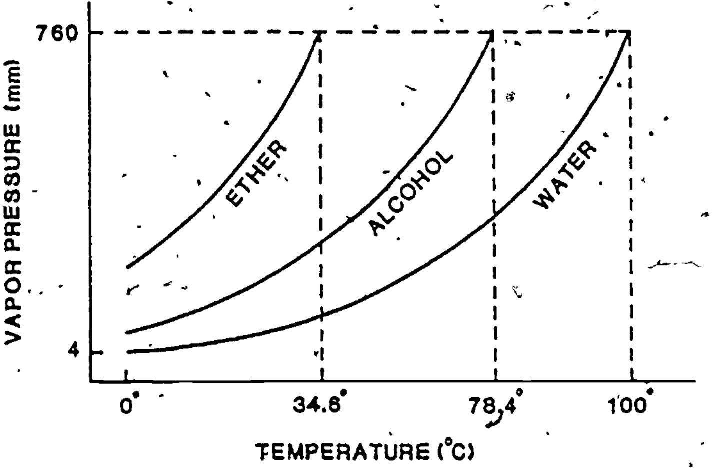


Figure 2. Vapor Pressure Curves for Water, Alcohol, and Ether.

A liquid may boil at temperatures higher than normal under external pressures greater than one atmosphere (atm). On the other hand, the boiling point of a liquid may be lowered below normal by decreasing the pressure on the surface of the liquid below one atmosphere. For example, at high altitudes, where atmospheric pressure is less than 760 mm, water boils at temperatures below 100°C. Food cooked in boiling water cooks more slowly at high altitudes because the temperature of boiling water is lower than it would be nearer sea level. In a pressure cooker, the temperature of boiling water is higher than normal, due to higher vapor pressures; therefore, it is possible to cook foods faster in a pressure cooker than in an open container.

DISTILLATION

Consideration of vapor pressure is important in distillation, which is one of the oldest and easiest methods of separation. In a simple distillation, a liquid is vaporized through the application of heat. The vapor is cooled until it condenses (as a liquid) away from the original liquid, thus effecting a separation. A laboratory setup for a simple distillation is illustrated in Figure 3.

Suppose salt is to be removed from seawater so the water can be used for drinking purposes. The seawater is placed in a distilling flask and heated with a burner. Since water vaporizes at a much lower temperature than salt, the water boils off, leaving a residue of salt in the flask. The water is cooled in the condenser where the vapor is condensed to a liquid. If a mixture containing several volatile components is

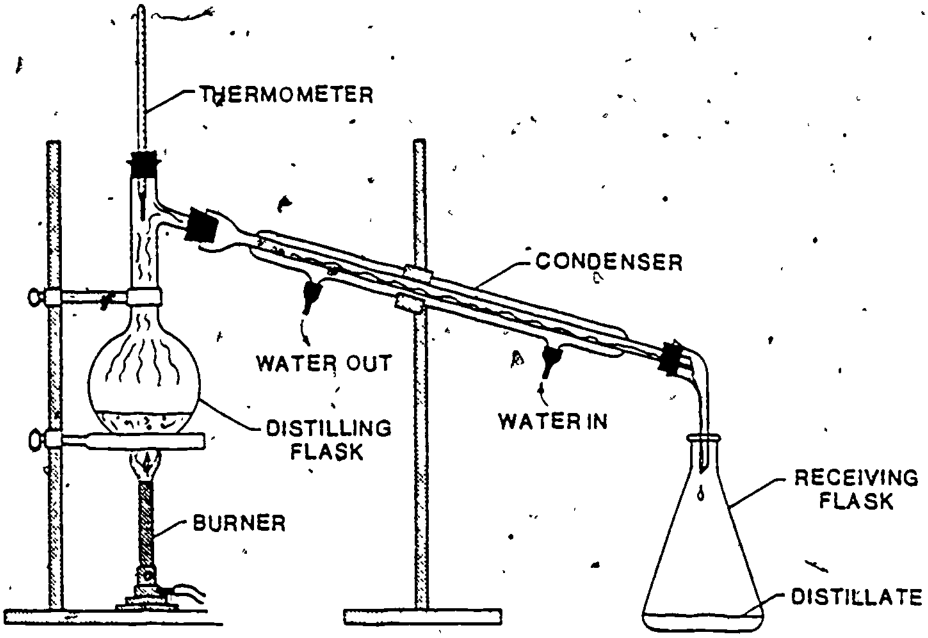


Figure 3. Simple Distillation.

distilled, a complete separation does not occur in one distillation. However, the vapor is richer in the more volatile component. Repeated distillations can eventually effect a separation; however, a fractionating column can be used to separate volatile components in a single operation. Figure 4 illustrates a bubble-cap fractionating column. Each layer of liquid on the plates is equivalent to the boiling liquid in a distillation flask, and the liquid on the plate above it is equivalent to the condenser. Thus, by the time the vapors reach the top of the column, and are condensed, the components will have been separated. In the petroleum industry, fractional distillation is used to separate crude oil into fractions. Gasoline, the more volatile component of crude oil, boils in the range of 60-150°C. Kerosene, diesel, and jet fuel boil in the 150-250°C range; lubricating oils boil in the 250-350°C range. Distillations, which are used in industry to separate liquid mixtures, consume a large amount of energy. Thus, they are important to an energy technician.

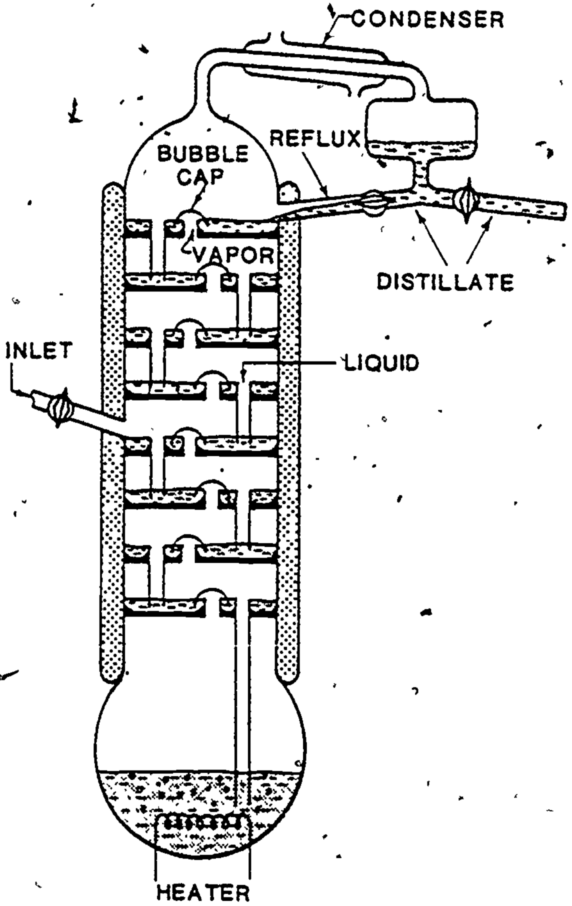


Figure 4. Fractionating Distillation Column.

VISCOSITY

Some liquids – like molasses – flow slowly, while others flow easily. The resistance of liquids to flow is referred to as viscosity. The higher the viscosity is, the slower the flow of the liquid. Molasses, for instance, is a viscous liquid, whereas water is not. Motor oils have various viscosities, depending upon their intended use (that is, for hot or cold weather). Viscosity is related to the ease with which individual molecules can move past other molecules. It depends upon the attractive forces between molecules, as well as the structure of the molecules. One way to determine viscosity is by measuring the time it takes for a given amount of liquid to flow through a small diameter tube under the influence of gravity. Such a device, called an Ostwald viscometer, is shown in Figure 5.

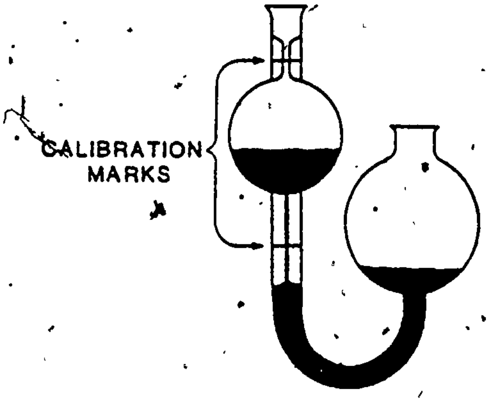


Figure 5. Ostwald Viscometer.

Viscosity is temperature dependent since at higher temperatures the attractive forces between molecules are decreased, thereby allowing the liquid to flow more easily. Hence, careful control of the temperature must be maintained in measuring viscosities. Viscosities of several common liquids at 20°C are given in Table 2. As shown, the viscosity of ether and benzene is less than that of water; whereas the viscosity of ethanol, mercury, and glycerin is greater. Viscosities may be reported in a number of units; those in Table 2 are kilograms of liquid flowing one meter per second. Viscosity, an important property of liquids,

TABLE 2. VISCOSITIES OF LIQUIDS at 20°C.

Substance	Viscosity (kg/m-sec) x 10^{+3}
Ether	0.23
Benzene	0.65
Water	1.00
Ethanol	1.20
Mercury	1.55
Glycerin	1490.00

can be measured by using a variety of laboratory techniques in addition to the Oswald viscometer.

In many manufacturing processes, liquids must be pumped from one location to another, and they must be mixed with other liquids. The ease with which liquids can be pumped or mixed is closely related to the liquids' viscosities. These processes are high energy consumers, and proper control of them can result in significant energy savings.

DISPERSIONS

Water has the ability to dissolve other substances. For example, salt and sugar will dissolve in water to form a solution. To distinguish between the dispersed material in a solution and the dispersing medium, the dispersed material is called the solute, and the medium is called the solvent. Thus, in the above example, salt or sugar is the solute, and water is the solvent. A dispersion is a broader term than solution; it is the result of suspending any substance in a liquid or other medium. Dispersions can be divided into three types: suspensions, colloids, and solutions. The distinction between these types is a matter of the size of the particles dispersed.

SUSPENSIONS

When a dispersion contains particles that are too large to remain suspended indefinitely in the medium, it is called a suspension. Suspensions remain stable for a limited time, until the suspended particles settle out. The dispersed particles that tend to settle out are usually larger than 0.0001 mm in diameter. Examples of suspensions are sand in water, oil in water, and smoke in air.

COLLOIDS

If the size of particles is between approximately 0.0001 and 0.000001 mm, the dispersion tends to be stable for a long time, or even indefinitely. These dispersions are commonly called colloids. Examples of colloids are the following: milk, mayonnaise, hair creams, and gelatin. The familiar "sunbeams" that are seen as sunlight enters a room through a small opening is another example of a colloidal dispersion. This effect is known as the "Tyndall" effect.

SOLUTIONS

If the particles are smaller in diameter than 0.000001 mm, the dispersion is called a solution. Particles of this size are approximately the size of single molecules or ions. Solutions are generally clear, although they may be highly colored. The particles of a solution will never settle out of the solution. Solutions can be characterized as follows: The solute particles (1) are molecular in size; (2) are uniformly distributed throughout the solvent; (3) cannot be seen; (4) do not settle out on standing; and (5) cannot be removed by filtration.

SOLUTION TYPES AND CHARACTERISTICS

Since there are three states of matter (solid, liquid, and gas), nine different types of solutions are possible. The solute may be a gas, liquid, or solid; likewise, the solvent may be a gas, liquid, or solid. The nine possible combinations and examples of each are given in Table 3. The important combinations are solids in a liquid; and of these, the most important are solids in water as a solvent.

Sometimes a substance is said to dissolve when a chemical reaction leads to a homogeneous mixture. For example, magnesium metal can be dissolved by hydrochloric acid as follows:

$$Mg + 2\,HCl \longrightarrow H_2 + MgCl_2$$

TABLE 3. TYPES OF SOLUTIONS.

Solute	Solvent	Example
Gas	Gas	Nitrogen in air
Gas	Liquid	Oxygen in water
Gas	Solid	Hydrogen in steel
Liquid	Gas	Water in air
Liquid	Liquid	Alcohol in water
Liquid	Solid	Mercury in gold
Solid	Gas	Iodine in air
Solid	Liquid	Sugar in water
Solid	Solid	Manganese in iron (alloys)

In a case like this, the chemical form of the substance being dissolved is changed. If the solution is evaporated to dryness, Mg is not recovered; instead, $MgCl_2$ is recovered. In contrast, when NaCl is dissolved in water, it can be recovered unchanged when the water is evaporated. This module is primarily concerned with the type of solution in which the material simply dissolves rather than reacts with the solvent.

IMPORTANCE OF SOLUTIONS

Many reactions occur only when the reactants are in solution. For example, when powdered silver nitrate is mixed with powdered sodium chloride, no noticeable reaction takes place. However, when solutions of silver nitrate and sodium choride are mixed, an immediate reaction takes place, as indicated by the formation of a precipitate of silver chloride:

$$AgNO_3 + NaCl \longrightarrow \underline{AgCl} + NaNO_3$$

When reactants are in the solid state, reaction is possible only for the molecules or ions in the surface layers of the particles. The mobility of ions and molecules in solution is much greater, allowing for rapid reaction. Almost all reactions conducted in chemical operations involve chemical solutions. Solubilities are an important consideration in working with solutions. In the preceding example, the usefulness of the reactions depends on the insolubility of silver chloride in water as compared to the high solubility of sodium nitrate. The difference in solubilities allows the silver chloride precipitate to be separated from the sodium nitrate by filtration.

SOLUTION PROCESS

When crystalline materials (like sugar) dissolve in water, the strong attraction of the sugar particles for themselves is overcome by the stronger attraction of water molecules for the sugar particles. As each sugar molecule is removed from the crystal, molecules of water cluster around it, forming a solvated molecule. The water molecules "pull" the sugar molecules from the crystal and, in the process, surround the molecule. Often, when molecules are separated from each other (as when substances dissolve) heat is absorbed. Most substances absorb heat on solution. For instances, when a salt such as ammonium nitrate, NH_4NO_3, is dissolved, it will absorb enough heat so that the resulting solution is very cold. Water may form on the outside of a beaker during the dissolution process; or, in some cases, frost may form.

On the other hand, a few substances produce heat when they dissolve in water. Lithium carbonate and calcium acetate are examples of salts that produce hot solutions when dissolved. When concentrated sulfuric acid is added to water, enough heat may be evolved in the process to raise the temperature of the solution to more than 100°C; that is, the solution can actually boil. Obviously, care must be exercised when using concentrated sulfuric acid because of the possibility of thermal, as well as chemical, burns. The heat envolved or absorbed in the process of solution of a substance is known as the <u>heat of solution</u>.

UNSATURATED AND SATURATED SOLUTIONS

A limited amount of solute can be dissolved in a given amount of solvent. For example, 1 gram (g) of table salt will dissolve in 100 g of water at 25°C. If another gram of salt is added to the solution, it, too, will dissolve. This solution is said to be "unsaturated." An unsaturated solution is one in which more solute can dissolve in a given weight of the solvent at a given temperature. However, as more salt is added to the solution, it will be discovered that, after the solution contains 36.5 g of salt, no more salt will appear to dissolve. When this point is reached, the solution is said to be "saturated." After this point, even one additional grain of salt will not appear to dissolve. In reality, however, the salt may continue to dissolve, but an equal amount of salt will crystallize. An equilibrium is established at the saturation point as follows:

$$\text{Solute} + \text{Solvent} \xrightleftharpoons[\text{Crystallize}]{\text{Dissolve}} \text{Solution}$$

A <u>saturated solution</u> is one that is in equilibrium with undissolved solute. A saturated solution may also be defined as "one in which the rate of dissolution equals the rate of crystallization."

SOLUBILITY

Solubility is usually defined as "the number of grams of substance that will dissolve in 100 g of solvent at a given temperature." In the example of table salt given previously, the solubility is 36.5 g/100 g water. Solubilities of substances in water vary over a wide range. Most gases are rather insoluble in water, whereas the solubility of solids and liquids in water is quite variable. Table 4 lists some of the solubilities of substances in water.

TABLE 4. SOLUBILITIES OF SUBSTANCES IN WATER.

Solute	Solubility (g/100 g water at 0°C)
Silver iodide	0.0000003
Barium sulfate	0.00003
Nitrogen	.003
Oxygen	.003
Ether	7.0
Sodium chloride	35.7
Sodium nitrate	73.0
Potassium iodide	127.5

Table 4 shows that more than 127 g of potassium iodide will dissolve in 100 g of water, whereas less than one-millionth of a gram of silver iodide will dissolve in 100 g of water. Compounds like silver iodide and barium sulfate, which have low solubilities, are often referred to as being insoluble, although nothing is truly insoluble.

In general, solution takes place most readily when the structure of the molecules of the solvent and solute is alike. A useful rule of thumb is "like dissolves like." This rule can be used in selecting a solvent for a particular application.

EFFECT OF TEMPERATURE ON SOLUBILITY

Solubility of solids in liquids generally increases with an increase in temperature. However, there are some exceptions to this general rule. For example, the solubility of lithium carbonate, calcium acetate, and calcium hydroxide decreases with an increase in temperature. Figure 6 shows the solubilities at various temperatures of the following three compounds: sodium nitrate, sodium chloride, and cerium sulfate. Sodium nitrate shows a marked temperature dependence in its solubility, increasing from approximately 73 g/100 g water at 0°C to almost 150 g/100 g water at 80°C. Sodium chloride is not very temperature dependent in its solubility since it only increases from 35 g/100 g water at 0°C to approximately 40 g/100 g water at 100°C. Cerium sulfate is an unusual salt; its solubility decreases with increasing temperature, as shown in Figure 6. Solubility of gases in liquids decreases as temperature increases; whereas solubility of liquids in liquids is not very dependent upon temperature.

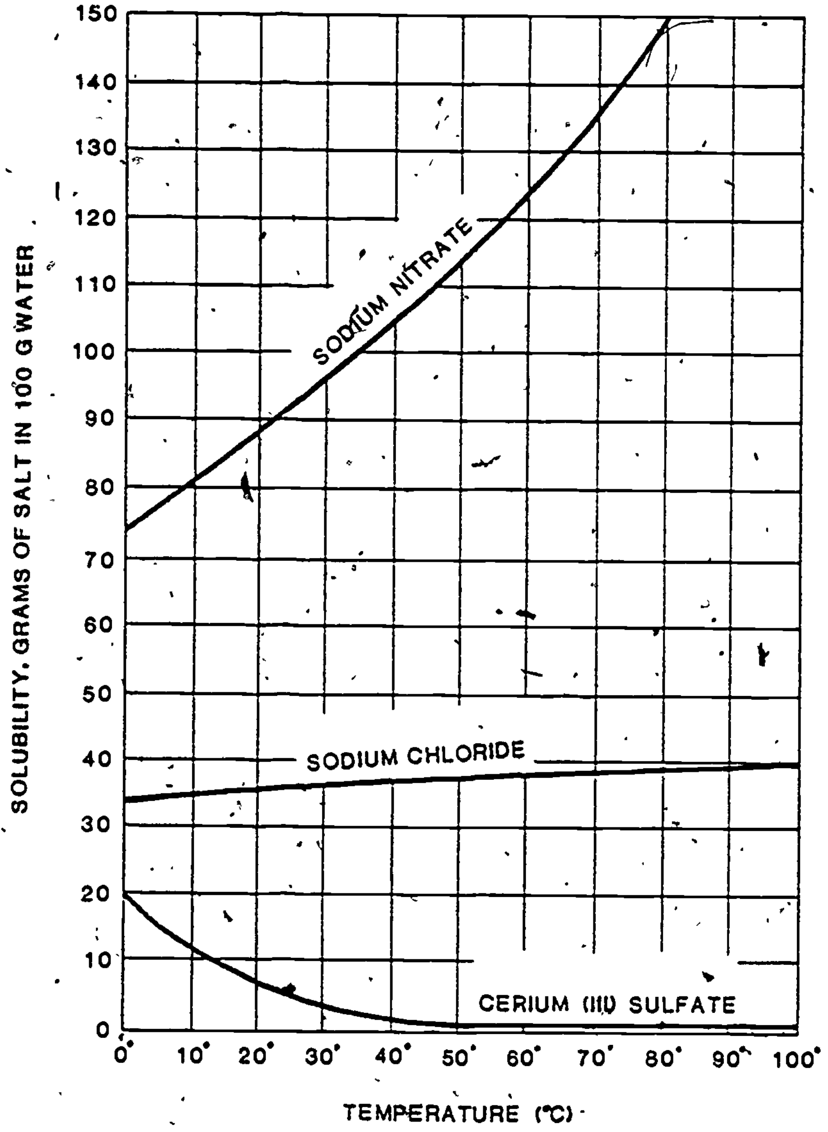


Figure 6. Solubility Curves.

EFFECT OF PRESSURE ON SOLUBILITY

Changes in pressure do not noticeably affect solubility of liquid and solids in liquids. However, the solubility of a gas is very dependent upon its pressure. The weight of a gas that can dissolve in a fixed weight of liquid at a given temperature is directly proportional to the pressure of the gas. For example, oxygen has a solubility of 0.007 g/100 g of water at 0°C and 1 atm of pressure. If the pressure is increased to 2 atm, the solubility of oxygen will increase to 0.014 g/100 g of water. As the pressure is doubled, the solubility of oxygen doubles.

RATE OF SOLUTION

As indicated earlier, solubility depends on the nature of the solute and solvent, and on temperature and pressure. Solubility is the amount of a substance that will dissolve in a solvent. This amount is independent of the speed at which the dissolving takes place. The most important factors in solution rate are as follows:

- Surface area: Pulverizing the solute increases its surface area; since solution occurs at the surface, an increase in surface area will increase the solution rate.
- Agitation: Stirring the solution allows more dilute solution to come into contact with the solid, thereby increasing the rate of solution.
- Heat: In many cases, heating the solution increases the rate of solution since kinetic energy of the molecules is increased.
- Amount of solute present: When the solute is first added to the liquid, the rate of solution is faster than when the solute is already present in the solution. Therefore, an equilibrium is established, and,

as saturation is approached, almost as much solute crystallizes as that which dissolves.

SUPERSATURATED SOLUTIONS

Occasionally, it is possible to have more than the saturation value of solute present at a given temperature and the solution still be perfectly clear. This situation can occur if a saturated solution is made at a high temperature and the solution is cooled without stirring or shaking of any kind. As the temperature decreases, the solute can stay in solution, and the solution is said to be "supersaturated." A simple way to convert a supersaturated solution to a saturated solution is to add a small crystal of the solute, upon which the excess solute in solution will rapidly build. Shaking or stirring also may convert a supersaturated solution to a saturated solution. Often, this conversion is fast and spectacular.

CONCENTRATION OF SOLUTIONS

The concentration of a solution can be expressed either qualitatively or quantitatively. The terms "concentrated" and "dilute" are used to qualitatively describe a solution. A solution with a relatively small amount of solute is said to be dilute, whereas one with a large amount of solute is said to be concentrated. Sometimes the terms "weak" and "strong" are incorrectly used to describe concentration of a solution. A weak solution is one in which the solute is not completely ionized, such as acetic acid. A weak solution may be concentrated, such as glacial acetic; or it may be dilute, such as a 5% solution of acetic acid (vinegar). A strong solution is one in which the solute is highly ionized, such as hyrochloric acid. A strong solution may be concentrated or dilute, such as concentrated or dilute hydrochloric acid. It should be remembered, then, that dilute and concentrated refer to the amount of solute present; whereas weak and strong refer to the state of ionization or dissociation of the solute in the solvent.

One of the simplest quantitative ways of expressing concentrations is in terms of weight percentage (wt %). The weight percentage formula is as follows:

$$\text{Wt \%} = \frac{\text{Mass solute}}{\text{Mass solute + Mass solvent}} \times 100$$

EXAMPLE A: CONCENTRATIONS IN TERMS OF WEIGHT %.

Given: A solution containing 8.3 g of Na_2CO_3/100 g of water.

Find: The wt % of sodium carbonate in this solution.

Solution:

$$\text{Wt \%} = \frac{\text{Mass solute}}{\text{Mass solute + Mass solvent}} \times 100$$

$$= \frac{8.3 \times 100}{8.3 + 100}$$

$$= \frac{8.3}{108.3} \times 100$$

$$\text{Wt \%} = 7.7\%.$$

The wt % of solvent in this solution is as follows: 100 - 7.7 = 92.3%.

MOLAR CONCENTRATION

The most common unit of concentration in chemical work is molarity. The molarity of a solution is the number of moles of solute per liter of solution:

$$\text{Molarity} = \frac{\text{Number of moles of solute}}{\text{Liters of solution}}$$

A one-molar (1 M) solution contains one gram-formula weight, or mole, of the solute in one liter of solution. Note that a liter of solution, rather than a liter of solvent, is specified in this definition. Molar solutions are convenient to use because equal volumes of one-molar solutions will contain the same number of molecules of solute. For example, consider the reaction of sodium hydroxide with hydrochloric acid:

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

One liter of 1 M NaOH will contain the same number of molecules as one liter of 1 M HCl. Likewise, 300 ml of 0.5 M NaOH will contain the same number of molecules as 300 ml of 0.5 M HCl.

To prepare one liter of 0.500 M solution of sugar ($C_{12}H_{22}O_{11}$) in water, the following steps are necessary:

1. The molecular weight (mol wt) of sucrose from the molecular formula is calculated as follows:

12 C 12 x 12.01 = 144.1

12 H 22 x 1.0 = 22.0

11 O 11 x 16.00 = 176.0

Mol wt sucrose = 342.1 g/mole

2. The number of grams in 0.500 mole is calculated as follows:

    # grams = 342.1 g/mole x 0.500 mole = 171

3. One-half mole of sucrose, 171 g, is weighed and placed in a one-liter volumetric flask. A volumetric flask is designed to contain a specific volume of liquid, such as one liter. The flask contains the specific volume when filled to the line inscribed on its neck.
4. Enough water is added to dissolve the sucrose (approximately one-half a flask). The ground-glass stopper is inserted, and the flask is shaken to aid in the dissolution of the sucrose.
5. When all of the sucrose has been dissolved, additional water is added to the flask to bring the volume exactly to the mark on the neck.
6. Then the contents of the flask are thoroughly mixed by inverting the stoppered flask and bringing it to its normal position. Several minutes of shaking are required to thoroughly mix the solution.

As a result of the above process, one liter of a 0.500 M solution of sucrose has been prepared. Any concentration of solution can be made this way. Volumetric flasks, which are commonly used, have volumes of 50 ml, 100 ml, 250 ml, 500 ml, and 1000 ml. In solving problems concerning molarity, the following formula is convenient to use:

Grams of solute = (l of solution) x (M) x (g-mol wt)

EXAMPLE B: MOLAR CONCENTRATION.

Given:	The chemical formula Na_2SO_4.	
Find:	The number of grams of Na_2SO_4 that are required to make 350 ml of 0.50 M Na_2SO_4.	
Solution:	Mol wt of Na_2SO_4	= (2 x 23) + (1 x 32) + (4 x 16)
		= 142 g/mole
	Grams Na_2SO_4	= (0.350 l) x (0.50 M) x (42) g/mole)
	Grams Na_2SO_4	= 24.8 g.

NORMAL CONCENTRATION

In an experiment, it can be shown that sodium hydroxide in one liter of a 1-M solution will neutralize the hydrochloric acid in one liter of a 1-M solution of the acid. However, two liters of 1-M sodium hydroxide are required to neutralize the acid on one liter of 1-M sulfuric acid. Two liters are required because each mole of sulfuric acid supplies two hydrogen ions for each hydroxide ion available from one mole of sodium hydroxide. The reactions are as follows:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

$$H_2SO_4 + 2\ NaOH \longrightarrow Na_2SO_4 + 2\ H_2O$$

It is important to use equivalent weights rather than molecular weights in instances such as the neutralization of sulfuric acid. Gram equivalent weight is defined as "the number of grams of a substance associated with the transfer of one hydrogen ion." In the preceding reactions, the gram equivalent weights are as follows:

HCl:	Gram equivalent weight is equal to the molecular weight, since one hydrogen is transferred in the reaction.
H_2SO_4:	Gram equivalent weight is equal to one-half the molecular weight, since two hydrogen ions are transferred in the reaction.

Molecular weight of sulfuric acid = (2 x 1) + (1 x 32) + (4 x 16)
= 98 grams per mole.

Equivalent weight of sulfuric acid = 98/2
= 49 grams per mole.

<u>Normality</u> (N) of a solution is the number of gram equivalents of solute per liter of solution. A 1-N solution of sulfuric acid contains 49 grams of the acid in one liter. A 1-M solution of sulfuric acid contains 98 grams per liter, making a 1-M solution of the acid equal to 2-N.

EXAMPLE C: MOLARITY AND NORMALITY OF A SOLUTION.

Given: A solution of H_2SO_4.

Find: The molarity and normality of the solution made by dissolving 5 g of H_2SO_4 in enough water to make 200 ml of solution.

Solution: The molecular weight of H_2SO_4 is 98.

$$\text{Molarity} = \frac{5\ \text{g}}{200\ \text{ml} \times 98\ \text{g}} = 0.255\ \text{M}.$$

Because H_2SO_4 has two available hydrogen ions, there are two equivalents in a mole, and the normality is twice the molarity:

$$\text{Normality} = \frac{5\ \text{g}}{200\ \text{ml} \times 49\ \text{g}} = 0.510\ \text{N}.$$

DENSITY

To this point, this module has discussed the following ways of expressing concentration of solutions: weight percentage, molarity, and normality. Another method of expressing concentration is density. Density of a solution is given by its weight divided by its volume:

$$\text{Density} = \frac{\text{Weight}}{\text{Volume}}$$

For liquids and solutions, density is usually given in grams per milliliter (g/ml). If accurately weighed, one ml of water should weigh one gram; that is, the density of water is

1 g/ml. Likewise, the density of sulfuric acid is 1.84 g/ml; that of benzene is 0.879 g/ml; and that of ethyl alcohol is 0.790 g/ml. Because weighing liquids accurately is rather difficult, densities of liquids are commonly determined by using a hydrometer, a sealed glass tube that floats when placed in a liquid since it displaces an amount of liquid equal to its own weight. Hydrometers are calibrated to indicate the exact density of the liquid. They are widely used in measuring the density of sulfuric acid solutions in automobile batteries and the density of anti-freeze solutions in automobile radiators.

DILUTIONS

Quite often it is necessary to use a reagent chemical in a diluted form, rather than the concentrated form in which it is available. The usual procedure for preparing dilute solutions (for example, of acids) is to mix the concentrated reagent with distilled water. The following formula can be used to determine how much reagent is needed to give a final dilute solution of the required volume and known concentration:

$$M_{conc} V_{conc} = M_{dil} V_{dil} \qquad \text{Equation 1}$$

where:

M_{conc} = Molarity of the concentrated reagent.
V_{conc} = Volume of the concentrated reagent.
M_{dil} = Molarity of the dilute solution.
V_{dil} = Volume of the dilute solution.

The use of this equation is illustrated in the following example:

What volume of 12 M HCl should be diluted to make one liter of 3 M HCl?

$$\text{Number of liters conc HCl} = \frac{3\text{ M} \times 1\text{ liter}}{12\text{ M}}$$

$$= 0.25\text{ liter.}$$

Therefore, the required solution is made by adding 0.25 liter of 12 M HCl to enough distilled water to make the final volume of one liter of 3 M acid.

ACID, BASES, AND SALTS

One of the most important classifications of substances involves their acid and base properties. Certain substances, called acids, possess a sour taste and are capable of dissolving metals such as zinc. Acids also cause certain dyes or indicators to turn a characteristic color. For example, litmus turns red on contact with acid. Common acids include citric (found in citrus fruits), ascorbic (vitamin C), sulfuric, nitric, and hydrochloric. Whereas acids turn litmus red, bases turn litmus blue. Bases have a characteristic bitter taste and feel slippery to the touch. Most bases will not dissolve metals; however, aluminum is a notable exception. Two common bases are sodium hydroxide (lye) and potassium hydroxide.

Since all acids have some common properties, and all bases also have common properties, it seems logical to assume that there are some features common to each. It has been determined that acids are substances that contain hydrogen ion (H^+), and bases are substances that contain hydroxide ion (OH^-). Bases and acids react in a process called neutralization, in which the hydrogen ion combines with the hydroxide ion to form water as follows:

$$H^+ + OH^- \longrightarrow H_2O$$

As an example, consider the reaction of hydrochloric acid with sodium hydroxide to form water and sodium chloride:

$$HCl + NaOH \longrightarrow H_2O + NaCl$$

TITRIMETRIC ANALYSIS

Reactions such as the preceding one between hydrochloric acid and sodium hydroxide are commonly encountered in an analytical technique called titrimetric analysis. For example, if a known amount of acid having a known concentration is added to a known volume of base, then the concentration of the base may be calculated. An indicator such as litmus will change color when an equivalent amount of acid and base are present, thus signaling the stopping point of the titration. The following equation applies to titrimetric determinations:

$$ml_{acid} \times N_{acid} = ml_{base} \times N_{base} \qquad \text{Equation 2}$$

As an illustration of the use of this equation, consider the following analysis:

In a titration, 33 ml of hydrochloric acid with a concentration of 0.12 N are required to reach the equivalent (or end point) when titrating a 25 ml sample of sodium hydroxide. The concentration of sodium hydroxide may be found as follows:

$ml_{acid} = 33$

$N_{acid} = 0.12$

$ml_{base} = 25$

$N_{base} = \text{unknown}$

Substituting these quantities in the preceding equation gives the following:

$$N_{base} = \frac{33 \text{ ml} \times 0.12 \text{ N}}{25 \text{ ml}} = 0.16 \qquad \text{Equation 3}$$

Thus, the concentration of the sodium hydroxide is 0.16 N. The actual process of performing a titration is relatively accurate, fast, easy, and inexpensive; therefore, the technique of titrimetric analysis is widely used in industrial laboratories. In Laboratory 2 of this module, the percentage of acetic acid in vinegar will be determined by utilizing titrimetric analysis.

pH

The concentration of hydrogen ion in a solution is a measure of its acidity or basicity. Because of the wide range of hydrogen ion concentrations possible, it has been found convenient to express the concentration of hydrogen ion in terms of the negative logarithm of the hydrogen ion concentration. This quantity, known as "pH" of a solution, is expressed mathematically as follows:

$$pH = -\log [H^+] = \log \frac{1}{[H^+]} \qquad \text{Equation 4}$$

where:

The symbol [ ] = Molar concentration.

The pH scale generally is considered to range from 0 to 14, although pH's of less than 0 and more than 14 are possible. A pH of 7 indicates a neutral solution, such as water, in which the concentrations of hydrogen ion and hydroxide ion are equal. A pH below 7 indicates an acidic solution, and a pH above 7 indicates a basic solution. The approximate pH's of several commonly encountered solutions are given in Table 5.

TABLE 5. pH'S OF COMMON SOLUTIONS.

Solution	pH	
0.1 M HCl	1.0	
Gastric juice	1.8	
Wine	3.3	Increasingly acidic
Coffee	5.0	
Milk	6.7	
Pure water	7.0	
Blood	7.4	
Borax	9.3	
Household ammonia	11.8	Increasingly basic
0.1 M NaOH solution	13.0	

MEASUREMENT OF pH

Several methods are available for the measurement of pH. The simplest method is the use of an indicator, which is a colored substance that can exist in either acid or base form. The two forms have a different color; therefore, by adding a small amount of the indicator to a solution and noting its color, it is possible to determine whether the indicator is in the acid or base form. By knowing the range in which the indicator changes color, it is possible to estimate the pH value. For example, methyl red is red at a pH below 4.2; but it is yellow at a pH above 6.3.

Paper tape can also be impregnated with various indicators for use in approximating pH. However, a more accurate method is to use a pH meter to determine the pH of a solution.

A pH meter consists of a pair of electrodes that are placed in the solution to be measured and a sensitive meter that measures small voltage differences between the two electrodes. (This voltage is dependent upon concentration of hydrogen ion in the solution. The meter is calibrated to read pH.) One of the electrodes is a reference electrode that produces a stable, known voltage; the other electrode, known as a glass electrode, is sensitive to the concentration of hydrogen ion in the solution. The tip of the glass electrode is made of a special thin glass that is permeable to hydrogen ion. The pH meter is one of the most common instruments found in chemical laboratories. Many portable models are available for use in environmental studies.

Recently developed electrodes can measure the concentrations of specific ions, such as Ca^{++}, F^-, Cl^-, and others. The concentrations of the ions in solution can be determined in much the same manner as a pH measurement. These specific ion electrodes can be used for fast, accurate concentration determinations. The primary difficulty lies in possible interferences of other ions in the solution.

SALTS

As indicated previously, acids react with base to form water and salts. Salts are compounds of metals and nonmetals. All salts are crystalline solids which consist of ions. Therefore, when salts are dissolved in water, they conduct electricity — that is, they are electrolytes. Since they are ionic, salts have a relatively high melting point; for example, the melting point of NaCl is 800°C, and the melting point of Na_3PO_4 is 1340°. Many salts are found in nature; others are manufactured by a variety of processes.

EXERCISES

1. A solution contains 10 g of NaCl per 100 g of water. What is the weight percentage concentration of solute in this solution?
   a. 10.0%
   b. 9.1%
   c. 90.0%
   d. None of the above

2. How many grams of K_2CO_3 are needed to prepare 500 ml of a 0.25 M solution?
   a. 17.3
   b. 25.0
   c. 13.8
   d. None of the above

3. What is the molarity of a solution containing 15 g of $BaCl_2$ in 250 ml of solution?
   a. 0.15
   b. 137
   c. 0.29
   d. None of the above

4. How many grams of $AgNO_3$ are needed to prepare one liter of a 0.1 M solution?
   a. 108
   b. 85
   c. 10
   d. None of the above

5. A 10 ml sample of hydrochloric acid solution requires 36.1 ml of 0.14 N KOH for neutralization. What is the concentration (normality) of the acid?
   a. 0.51
   b. 1.97
   c. 5
   d. None of the above

6. Calculate the pH of 0.01 M HCl. (The concentration of hydrogen ion is 0.01, or 10^{-2}. Substitute this value in the expression for pH.)
   a. -2
   b. 12
   c. 2
   d. None of the above

7. If 10 ml of a liquid weighs 13 grams, what is the density of the liquid?
   a. 13
   b. 1.3
   c. 0.77
   d. None of the above

LABORATORY MATERIALS

Laboratory 1

Large test tube
Ring stand
Clamp
Asbestos wire gauze
400-ml beaker

Laboratory 2

0.2 N sodium hydroxide, standardized
10 ml pipette
Distilled water
50 ml buret

Laboratory 1 (Continued)

Thermometer
Glass stirring rod
Pipette, 10 ml, graduated
Balance
Potassium nitrate
Sodium nitrate
Potassium chlorate
Distilled water

Laboratory 2 (Continued)

250 ml Erlenmeyer flask
Phenolphthalein indicator
Buret holder
White table vinegar
(approximately 5%)

LABORATORY PROCEDURES

LABORATORY 1. SOLUBILITY CURVES.

Solubilities of salts generally increase with increasing temperature. In this experiment, the solubilities of potassium nitrate, sodium nitrate, and potassium chlorate are determined at several temperatures. A solubility curve is then constructed for each compound by plotting solubility against temperature.

PROCEDURE

Carefully record all data in Data Table 1.

1. Set up the apparatus, as shown in Figure 7, and bring the water in the beaker to boiling.
2. Weigh exactly 5 grams of potassium nitrate into the test tube.

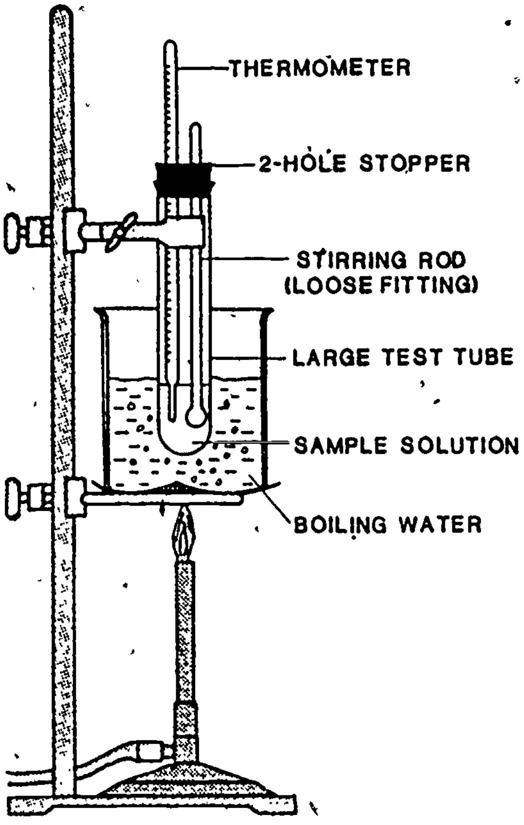


Figure 7. Apparatus for Determining Solubility.

3. Add exactly 3 ml of distilled water to the test tube, using a pipet.
4. Insert the stopper with the thermometer and stirring rod into the test tube and clamp the tube into position in the boiling water.
5. While heating the test tube, stir the solution until all the salt has dissolved. If the salt does not dissolve, add more distilled water 1 ml at a time, keeping a record of the amount of water added.
6. Remove the tube from the hot water as soon as all the salt has dissolved.
7. Record the temperature at which time crystals of salt are first observed in the test tube. Record this saturation temperature, along with the ml of water added, in Data Table 1.
8. Add an additional 1 ml of distilled water. Put the test tube back into the boiling water and repeat the determination of the saturation point under the new concentration condition. Record the temperature and ml of water in Data Table 1.

9. Continue adding 1 ml of water (as in Step 8), until five readings of saturation temperature have been obtained. Record all data in Data Table 1.
10. Repeat the above steps for sodium nitrate and potassium chlorate and record all data in Data Table 1.

For each reading, the concentration data needs to be converted into grams/100 g of water. Since the density of water is approximately 1 g/ml, the number of milliliters of water used is the number of grams of solvent. For each reading, calculate what weight of salt would be dissolved in 100 grams of water. For example, if 5 grams of salt dissolved in 4 grams of water, then calculate the number of grams that would be dissolved in 100 grams of water as follows:

$$\frac{5 \text{ Grams salt}}{4 \text{ Grams water}} = \frac{(x) \text{ Grams salt}}{100 \text{ Grams water}}$$

$$4 \times (X) = 5 \times 100$$

$$(X) = \frac{500}{4}$$

$$= 125 \text{ g/100 Grams water}$$

$$= \text{Solubility}$$

Record the calculated concentrations in the middle column of Data Table 1.

On the graph in Data Table 1, plot the solubility-temperature data and draw a smooth curve through the points for each salt. There should be a solubility curve for each of the three salts. Determine the solubility of each of the salts from the Handbook of Chemistry and Physics (Chemical Rubber Publishing Company) and compare it with the solubility that was determined experimentally. What are some of the sources of error which could affect the results?

LABORATORY 2. PERCENTAGE OF ACETIC ACID IN VINEGAR.

The percentage of acetic acid in vinegar is determined in this experiment by titrating the vinegar with a standardized or known solution of sodium hydroxide. The exact amount of the sodium hydroxide solution that neutralized the acid in the vinegar is determined by using phenolphthalein as the indicator. The percentage of acetic acid in vinegar is determined by using the following equation:

$$\text{\% Acetic Acid in Vinegar} = \frac{\text{Grams of acetic acid}}{\text{Weight of sample of vinegar}} \times 100$$

$$= \frac{ml_{NaOH} \times N_{NaOH} \times 0.06005}{\text{Weight of sample of vinegar}} \times 100 \qquad \text{Equation 5}$$

(The equivalent weight of acetic acid is 60.05 grams, and the milliequivalent weight is 0.06005 g.) A 10 ml sample of vinegar is used, and since the density of vinegar can be assumed to be near 1 g/ml (almost the same as water), the weight of the sample of vinegar is 10 grams. The percentage of acetic acid in most commercially available vinegars is approximately 5%.

PROCEDURE

Carefully record all data in Data Table 2.

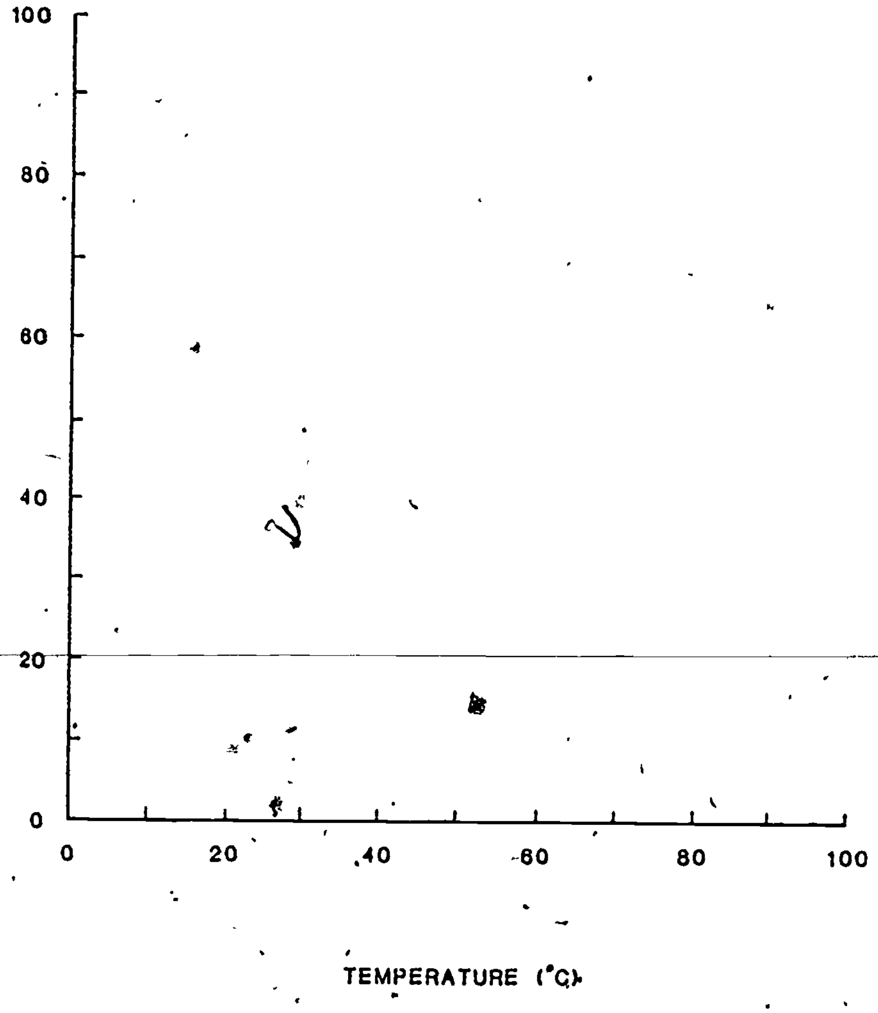
1. Pipette a 10 ml sample of vinegar into a 250 ml Erlenmeyer flask containing 50 ml of distilled water.
2. Add two drops of phenolphthalein to the flask.
3. Fill a buret with a standardized solution of 0.2 *N* NaOH.
4. Slowly add the NaOH to the vinegar until the end point is reached. In Data Table 2, record the ml of NaOH used.
5. Perform the titration twice more, using two additional samples of the same vinegar.
6. Calculate the percentage of acetic acid in the samples, and calculate the average of the three determinations. Record the calculations in Data Table 2.

DATA TABLES

DATA TABLE 1. SOLUBILITY CURVES.

ml Water Added	Calculated Conc (g/100 g water).	Saturation Temperature
Potassium Nitrate:		
1.		
2.		
3.		
4.		
5.		
Sodium Nitrate:		
1.		
2.		
3.		
4.		
5.		
Potassium Chlorate:		
1.		
2.		
3.		
4.		
5.		

Data Table 1. Continued.



DATA TABLE 2. PERCENTAGE OF ACETIC ACID IN VINEGAR.

	1st Run	2nd Run	3rd Run
Initial buret reading			
Final buret reading			
Volume NaOH used			
% Acetic acid			
Average % acetic acid: ____________			

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GLOSSARY

Acid: A substance containing hydrogen ion (H^+).

Atmosphere: A unit of pressure equal to the pressure of the air at sea-level or approximately 14.7 pounds to the square inch.

Base: A substance containing hydroxyl ion (OH^-).

Concentrated solution: One with a relatively large amount of solute.

Colloids: A dispersion which tends to be stable for a long time, but is not a true solution.

Density: Weight per unit volume.

Dilute solution: One with a relative small amount of solute.

Heat of solution: The heat evolved or absorbed in the process of solution of a substance.

Molarity: The number of moles of solute per liter of solution.

Neutralization: The reaction of an acid and base to form water and a salt.

Normality: The number of gram equivalents of solute per liter of solution.

Ostwald viscometer: A device for measuring the viscosity of liquids.

pH: A measure of the hydrogen ion concentration of a solution.

Salts: Crystalline solids composed of metal and nonmetal ions.

Saturated solution: One that is in equilibrium with undissolved solute.

Solute: The material dissolved ~~in~~ a medium.

Solution: A mixture in which the particles will never settle out since the particles are very small.

Solvent: The medium which dissolves a material.

Specific ion electrode: A device for measuring the concentration of ions in solution in much the same way in which pH is determined.

Strong solution: One in which the solute is highly ionized.

Supersaturated solution: A solution which holds more than the saturation value of solute at a given temperature.

Suspension: A dispersion of particles which are too large to remain in suspension indefinitely.

Unsaturated solution: One in which additional solute can dissolve in a given weight of the solvent at a given temperature.

Vapor pressure: The force exerted by a vapor in equilibrium with its liquid at a given temperature.

Viscosity: The resistance of liquids to flow.

Weak solution: One in which the solute is not completely ionized.