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ABSTRACT

This is a 1971 reprint of the chemistry syllabus printed in 1966 for the state of New York. This course of study presents a modern view of chemistry suitable for pupils with a wide range of skills and abilities. The outline of topics provides the unifying principles of chemistry together with related facts. The principles included in the outline are basic to man's understanding of his environment. The topical outline is divided into nine major units: Matter and Energy, Atomic Structure, Bonding, Periodic Table, Mathematics of Chemistry, Kinetics and Equilibrium, Acid-Base Theories, Redox and Electrochemistry, and Organic Chemistry. A tenth unit, Application of Principles of Reaction, is to be related to, and integrated in, other units as it applies. Each unit is subdivided into specific topics, each of which has additional material associated with it. Understandings and Fundamental Concepts outlines the basic concepts, while Supplementary Information provides amplification and explanation of the basic concepts. Two appendices deal with an energy level diagram used to depict electron configuration and with some mathematical concepts used in the course. (Author/TS)

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CHEMISTRY

A Secondary School Course With Major Emphasis on Fundamental Concepts

1971 Reprint

The University of the State of New York
The State Education Department
Bureau of Secondary Curriculum Development
Albany, 1966

THE UNIVERSITY OF THE STATE OF NEW YORK

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(with years when terms expire)

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FOREWORD

In the spring of 1963, the State Education Department appointed a Science Advisory Committee for the purpose of initiating revisions in its courses in science. The committee membership included representatives from the secondary and collegiate levels, from industry, and from research institutions. The function of this committee was to establish guidelines to aid the specific syllabus revision committees in their task of updating the various syllabuses in the light of recent developments in society, science, and science education.

Major recommendations of the Science Advisory Committee included: (1) that the present science courses be brought up to date in the light of recent developments in the field of science, (2) that a greater emphasis be placed on the understandings and concepts involved in the particular subject matter areas, and (3) that attention be given to coordinating the laboratory work with the content aspects of each course.

Shortly after the Science Advisory Committee meeting, a chemistry syllabus committee was appointed by the Department to develop a State course of study with updated content and which would incorporate recent trends in chemistry education.

The members of the Revision Committee were:

Jacob Brodkin, Plainview High School
Joseph F. Castka, Martin VanBuren High School, N.Y. City
Edgar M. Clemens, Ithaca High School
Henry Dorin, Boys High School, N.Y. City
Fred Riebesell, State University College at Oneonta
Anne Sperry, Jamesville-DeWitt High School
Henry Weisman, North Shore High School

Under the joint supervision of the Bureau of Science Education and the Bureau of Secondary Curriculum Development, the original draft of the new syllabus was written during the summer of 1964 and tested in 27 selected schools during the 1964-65 school year. A revision of the original draft was made during t
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George H. Murphy
Director, Divisi

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A revision of the original draft
was made during the summer of 1965 based on the experience and reactions of the cooperating schools. In addition a copy of the 1965 revision was sent to all the schools in the State in order that chemistry teachers might evaluate the scope, depth, and direction of the new course.

The new course of study was received favorably by teachers in the field thus making it possible to prepare the final revision during the summer of 1966. Examinations which were administered and analyzed supplied further evidence as to the applicability of the materials.

A writing team composed of W. Allister Crandall, Pulaski High School; Seymour Kopilow, Farmingdale High School; and Benedict Varco, Eden Central School, prepared the original draft as consultants to the Bureau of Secondary Curriculum Development. The modifications and improvements included in the 1965 revision were incorporated by Mr. Crandall and Sol Medoff of Westbury High School. The final draft was prepared by Mr. Crandall and Kenneth Schnobrich of the State University College at Buffalo, formerly at Clarence Central School.

John V. Favitta, Associate in Science Education, has guided the syllabus development since its inception, and Hugh Templeton, Chief of the Bureau of Science Education, acted as consultant throughout the project. Robert G. MacGregor, formerly Associate in Science Education and now Supervisor of Secondary Education, reviewed the manuscript and made valuable suggestions.

A review of content accuracy was made by John Lanese of the Chemistry Department of Union College. Robert F. Zimmerman, Associate in Secondary Curriculum, prepared the final copy for publication.

Gordon E. Van Hooft,
Chief, Bureau of Secondary Curriculum Development

George H. Murphy
Director, Division of School Supervision

Introduction to Teachers

Content

This course of study presents a modern view of chemistry suitable for pupils with a wide range of skills and abilities. The outline of topics provides the unifying principles of chemistry together with their related facts. The principles included in the outline are basic to man's understanding of his environment.

Sequence

This syllabus does not prescribe a specific sequence that must be followed. *The teacher is at liberty to use any sequence of topics and such teaching techniques as are appropriate for his students.*

The teacher should not hesitate to introduce briefly any of the subject matter as the occasion demands, even though it will be treated in depth later in the course. The teaching should be an interlocking process so that the students get both a foretaste of what is to come and a review of the material previously covered.

Teachers may find it desirable to incorporate the material outlined in Unit 5, "The Mathematics of Chemistry," in appropriate places in the course, rather than as a separate unit.

The processes outlined in Unit 10, "Application of Principles of Reaction," are intended to serve as illustrations, and should be related to other units as they apply. *Details of the processes will not be subject to examination.*

Prerequisites

In order to be able to understand and use the principles of chemistry introduced in this outline,

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Student
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Introduction to Teachers

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lined in Unit 10, "Application of ," are intended to serve as uld be related to other units as *f the processes will not be subject*

ble to understand and use the ry introduced in this outline, students should be familiar with the use of standard notation of numbers, significant figures, metric system of units, heat units, dimensional analysis (inclusion of units in mathematical computations), and an understanding of direct and inverse relationships.

Students enrolling in the Regents course in chemistry must have completed Ninth Year Mathematics, Course I - (Algebra). They also should have completed or, at least, be currently enrolled in Tenth Year Mathematics. While very little of the content of Mathematics 10 is used directly in chemistry, the experience in setting up and solving problems, and the analytical thinking developed in this course are most useful in chemistry.

Time Requirement

The minimum time required for this course is six 45-minute periods per week, although seven periods are recommended. This time allotment should include at least one double period for laboratory work each week.

State Diploma Credit

This course may be used as one unit of the Group II major science sequence or for Group III credit as an elective toward a State Diploma.

Laboratory

Laboratory work should be designed to encourage students to search for relationships. To do this, some laboratory exercises must be quantitative in nature. It is suggested that individual and class results of laboratory data be analyzed from time to time so that students can develop concepts of precision and accuracy. The type of experiment sometimes referred to as "cook-book chemistry" offers little toward the development of understanding.

A chemistry laboratory exercise is defined here as the laboratory work done by the student during one school period. In addition, a satisfactory written report of this work is required. The minimum laboratory requirement can be met by performing 30 individual exercises requiring 30 laboratory periods. The minimum requirement may also be met by 30 periods involving a smaller number of experiments which may require two or more periods for completion. This is not intended to permit the student to spread what is ordinarily a one-period exercise over two or more periods in order to earn additional time credit for this exercise.

Technology

The technological impact of chemistry should not be the only driving motivation for the study of chemistry. Students should also be made aware of the total effect of the application of chemical principles on our lives. Teachers are encouraged to use those applications which they deem important and which have local or intrinsic significance.

Organization of Syllabus

The material in the syllabus is organized under three headings:

Topics - This colu
outline. Sections mark
NOT subject to examinat
Understandings and
column outlines the bas
the course.
Supplementary Info
amplification and expla
with examples to illust
intended.
Statements which d
examination, and specif
printed in italics in t
Appendix I is a sc
that may be used in dep
Appendix II deals
of significant figures,
analysis, and the use o

Changes in Syllabus

Corrections or min
may become necessary wi
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In addition, a satisfactory written report
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yllabus

in the syllabus is organized under

Topics - This column contains the topical outline. Sections marked "†" are optional, and are NOT subject to examination.

Understandings and Fundamental Concepts - This column outlines the basic concepts to be introduced in the course.

Supplementary Information - This column includes amplification and explanations of the basic concepts, with examples to illustrate the depth of treatment intended.

Statements which delimit the material subject to examination, and specific suggestions to teachers are printed in italics in this column.

Appendix I is a schematic energy level diagram that may be used in depicting electron configuration.

Appendix II deals with the use and manipulation of significant figures, standard notation, dimensional analysis, and the use of graphs.

Changes in Syllabus

Corrections or minor changes in the Syllabus that may become necessary will be brought to the attention of school principals by means of a supervisory letter from the Department.

Topical Outline

Page

UNIT 1 - MATTER AND ENERGY

UNIT 2 - ATOMIC STRUC

Sections marked "†" are optional and not subject to examination

Topical Outline

UNIT 2 - ATOMIC STRUCTURE

nal and not subject to examination

UNIT 3 - BONDING

UNIT 4 - PERIODI

† Optional

UNIT 4 - PERIODIC TABLE

Page

UNIT 5 - MATHEMATICS OF CHEMISTRY*

UNIT 6 - KINET

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* *This material need not be taught as a unit, but may be incorporated in varic discretion of the teacher.*

† Optional

CS OF CHEMISTRY*

UNIT 6 - KINETICS AND EQUILIBRIUM

eed not be taught as a unit, but may be incorporated in various places in the syllabus at the he teacher.

UNIT 7 - ACID-BASE THEORIES

UNIT 8 - REDOX

UNIT 9 - ORGAN

† Optional

Page

UNIT 8 - REDOX AND ELECTROCHEMISTRY

UNIT 9 - ORGANIC CHEMISTRY

** *The material included under this heading should not be taught as a unit itsel be related to, and integrated in, other units as it applies.*

† Optional

UNIT 10 - APPLICATION OF PRINCIPLES OF REACTION**

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included under this heading should not be taught as a unit itself. Rather, this material should , and integrated in, other units as it applies.

Unit I - Matter and Energy

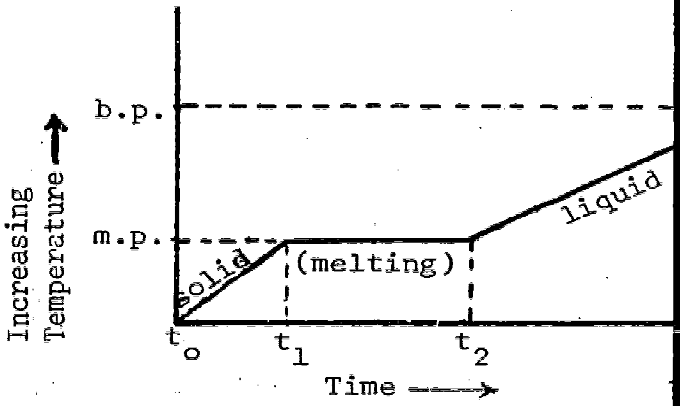
Topics	Understandings and Fundamental Concepts	Supplemen
I. Matter		*The concepts shown i* *Topics I and II of th* *to some degree in ear*
A. Substances	A substance is any variety of matter, all specimens of which have identical properties and composition. A substance is homogeneous.	All samples of a part heat of vaporization, and other properties be used for identific
1. Elements	An element is a substance which cannot be decomposed by a chemical change.	All samples of an ele the same atomic numbe
2. Compounds	A compound is a substance which can be decomposed by a chemical change. A compound is composed of two or more different elements.	All samples of a comp
B. Mixtures	A mixture consists of two or more distinct substances differing in properties and composition. The composition of a mixture can be varied.	Mixtures may be homog mixtures of gases)or of iron and sulfur).
II. Energy		
A. Forms of energy	Heat, light, and electricity are forms of energy.	
	Energy may be converted from one form to another but is never destroyed in a change.	Examples should inclu of forms of energy.
B. Energy and chemical change	Energy is either given off or absorbed in any chemical change.	Energy absorbed by th bonds which hold ator liberated when strong cept is developed in Although it is someti between physical char

Unit I - Matter and Energy

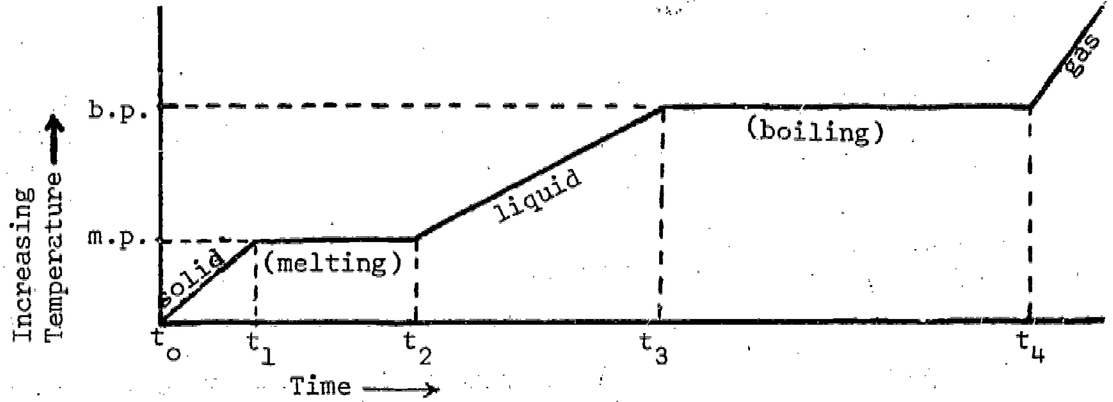
nderstandings and Fundamental Concepts	Supplementary Information
	The concepts shown in the column at the left for Topics I and II of this unit will have been covered to some degree in earlier science courses.
A substance is any variety of matter, all specimens of which have identical properties and composition.	All samples of a particular substance have the same heat of vaporization, melting point, boiling point, and other properties related to composition which can be used for identification.
A substance is homogeneous.	
An element is a substance which cannot be decomposed by a chemical change.	All samples of an element are composed of atoms with the same atomic number.
A compound is a substance which can be decomposed by a chemical change. A compound is composed of two or more different elements.	All samples of a compound have identical composition.
A mixture consists of two or more distinct substances differing in properties and composition. The composition of a mixture can be varied.	Mixtures may be homogeneous(e.g., solutions, or mixtures of gases)or heterogeneous(e.g., a mixture of iron and sulfur).
Heat, light, and electricity are forms of energy.	
Energy may be converted from one form to another but is never destroyed in a change.	Examples should include reactions involving a variety of forms of energy.
Energy is either given off or absorbed in any chemical change.	Energy absorbed by the reactants may be stored in the bonds which hold atoms together. This energy may be liberated when stronger bonds are formed. This concept is developed in Unit 6.
	Although it is sometimes convenient to distinguish between physical changes and chemical changes the

Topics	Understandings and Fundamental Concepts	
		distinction is of little
1. Exothermic reaction	An exothermic reaction releases energy.	The role of pp. 44-45.
2. Endothermic reaction	An endothermic reaction absorbs energy.	
3. Activation energy	Activation energy is the minimum energy required to initiate a reaction.	
C. Measurement of energy	Because energy in various forms may be converted to heat, the chemist uses heat units (calories or kilocalories) to measure the energies involved in chemical reactions.	
	One calorie is the amount of heat required to raise the temperature of one gram of water one Celsius degree.	*For scientif used.*
	One kilocalorie is equivalent to 1000 calories.	
1. Thermometry	Temperatures are indicators of the direction in which heat will flow. Heat flows spontaneously from a body at higher temperature to a body at a lower temperature.	The temperat kinetic ener At the same the particle
a. Fixed points on a thermometer	The fixed points on a thermometer are the ice-water equilibrium temperature at 1 atmosphere pressure, and the steam-water equilibrium temperature at 1 atmosphere pressure.	The temperat equilibria pressure eq kinetic the pages 4-5.
	On the Celsius scale the ice-water equilibrium temperature occurs at 0°C., and the steam-water equilibrium temperature occurs at 100°C.	Another sca (Absolute) temperature equilibrium

Understandings and Fundamental Concepts	Supplementary Information
	distinction is not clear cut, and to most scientists is of little importance.
An exothermic reaction releases energy.	The role of energy in reactions is developed in Unit 6, pp. 44-45.
An endothermic reaction absorbs energy.	
Activation energy is the minimum energy required to initiate a reaction.	
Because energy in various forms may be converted to heat, the chemist uses heat units (calories or kilocalories) to measure the energies involved in chemical reactions.	
One calorie is the amount of heat required to raise the temperature of one gram of water one Celsius degree.	*For scientific uniformity the term "Celsius" will be used.*
One kilocalorie is equivalent to 1000 calories.	
Temperatures are indicators of the direction in which heat will flow. Heat flows spontaneously from a body at higher temperature to a body at a lower temperature.	The temperature of a body is a measure of the average kinetic energy of its particles. At the same temperature, the average kinetic energy of the particles of all bodies is the same.
The fixed points on a thermometer are the ice-water equilibrium temperature at 1 atmosphere pressure, and the steam-water equilibrium temperature at 1 atmosphere pressure.	The temperature of the ice-water and steam-water equilibria can be defined more accurately as vapor pressure equilibria after the introduction of the kinetic theory in Section III, A, 7 of this unit, pages 4-5.
On the Celsius scale the ice-water equilibrium temperature occurs at 0°C., and the steam-water equilibrium temperature occurs at 100°C.	Another scale frequently used in science is the Kelvin (Absolute) scale on which the ice-water equilibrium temperature occurs at 273°K. and the steam-water equilibrium temperature occurs at 373°K.

Topics	Understandings and Fundamental Concepts	Su
2. Calorimetry	One method of measuring the heat absorbed or released in a reaction is by using a calorimeter.	*Illustrations should* *data. Attention shou* *figures in all calcul*
III. Phases of Matter	The term, "phase," is used to refer to the gas, liquid or solid form of matter.	The term, "phase," is avoid confusion with equilibrium."
	Change of phase of a substance is accompanied by the absorption or release of heat.	*Heating curves should* *temperature is plotte* *heat is added to a su* *Students should be ab* 
A. Gases	Gases take the shape and volume of the container.	
1. Boyle's law	At constant temperature, the volume of a given mass of a gas varies inversely with the pressure exerted on it.	It is useful to intro introduction to const *Students should be ab* *the change in volume* *at constant temperatu* *limited to simple mat*
2. Charles' law	At constant pressure, the volume of a given mass of a gas varies directly with the Kelvin (Absolute) temperature.	The Kelvin (Absolute) point at -273°C. with as on the Celsius sca

Understandings and Fundamental Concepts	Supplementary Information
One method of measuring the heat absorbed or released in a reaction is by using a calorimeter.	*Illustrations should be presented from laboratory data. Attention should be given to significant figures in all calculations. See Appendix II, p. 86.*
The term, "phase," is used to refer to the gas, liquid or solid form of matter.	The term, "phase," is used instead of "state" to avoid confusion with other conditions such as "state of equilibrium."
Change of phase of a substance is accompanied by the absorption or release of heat.	*Heating curves should be constructed in which the temperature is plotted against the time during which heat is added to a substance at a constant rate. Students should be able to interpret this type of curve:*



Understandings and Fundamental Concepts	Supplementary Information
Gases take the shape and volume of the container.	
At constant temperature, the volume of a given mass of a gas varies inversely with the pressure exerted on it.	It is useful to introduce the value PV = k as an introduction to constants. *Students should be able to predict the direction of the change in volume with a specified change in pressure at constant temperature. Minimum requirements will be limited to simple mathematical relationships.*
At constant pressure, the volume of a given mass of a gas varies directly with the Kelvin (Absolute) temperature.	The Kelvin (Absolute) temperature scale has its zero point at -273°C. with the size of the degrees the same as on the Celsius scale.

Topics	Understandings and Fundamental Concepts	
		The volume of at 0°C. for eac pressure remair While the basis could be extrap Law experiments absolute zero h liquify before *Students shoul* *change in volu* *at constant pr* *limited to sim*
3. Standard Temperature and Pressure (S.T.P.)	Standard temperature and pressure (S.T.P.) are defined as 0°C. (273°K.) and 760 mm. of mercury pressure (1 atmosphere).	Since the volu change in temp usually calcul
†4. Combined gas laws	$\frac{PV}{T}$ = constant	Since changes occur simultan two equations
†5. Graham's Law	Under similar conditions of temperature and pressure, gases diffuse at a rate inversely proportional to the square roots of their densities.	The density of at S.T.P.
†6. Partial pressures	The pressure exerted by each of the gases in a gas mixture is called the partial pressure of that gas.	The total pres sum of the ind comprising the
7. Kinetic theory	Study of gas behavior has led to a model based on the following assumptions:	*In explaining* *often convenie* *ture," a mathe* *Such models ca* *of gases. It* *an approximati* *predict behavi* *construction.*

Sections marked "†" are optional and not subject to examination.

Understandings and Fundamental Concepts	Supplementary Information
	The volume of a gas decreases by 1/273 of its volume at 0°C. for each decrease of 1°C., provided the pressure remains constant.
	While the basis for the Kelvin (Absolute) scale could be extrapolated from data relating to Charles' Law experiments, it should be pointed out that absolute zero has not been reached, and that all gases liquify before that point.
	Students should be able to predict the direction of the change in volume with a specified change in temperature at constant pressure. Minimum requirements will be limited to simple mathematical relationships.
Standard temperature and pressure (S.T.P.) are defined as 0°C. (273°K.) and 760 mm. of mercury pressure (1 atmosphere).	Since the volume of a given mass of a gas varies with change in temperature and pressure, gas volumes are usually calculated to an arbitrary standard (S.T.P.).
$\frac{PV}{T}$ = constant	Since changes in volume, pressure, and temperature often occur simultaneously, it is convenient to combine the two equations of Boyle and Charles into a single equation
Under similar conditions of temperature and pressure, gases diffuse at a rate inversely proportional to the square roots of their densities.	The density of gases is usually expressed in grams/liter at S.T.P.
The pressure exerted by each of the gases in a gas mixture is called the partial pressure of that gas.	The total pressure of a gas mixture is equal to the sum of the individual partial pressures of the gases comprising the mixture.
Study of gas behavior has led to a model based on the following assumptions:	*In explaining and interpreting observed behavior it is often convenient to use a model, which may be a "picture," a mathematical expression, or other mechanism. Such models can be useful in the study of the behavior of gases. It should be emphasized that the model is only an approximation, and is only as good as its ability to predict behavior in conditions not used in its original construction.*

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"+" are optional and not subject to

Topics

Understandings and Fundamental Concepts

- A gas is composed of individual particles which are in continuous, random, straight-line motion.
- Collisions between gas particles may result in a transfer of energy between particles, but the net total energy of the system remains constant.
- The volume of the gas particles themselves is ignored in comparison with the volume of the space in which they are contained.
- Gas particles are considered as having no attraction for each other.

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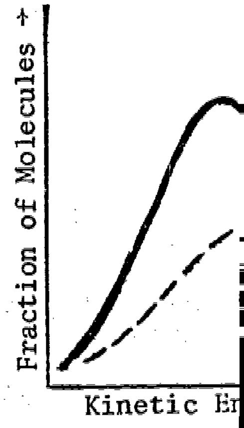
be shown graphi



8. Deviations from the gas laws

Deviations from the gas laws occur because the model is not perfect, and particularly because the gas particles do have volume and do exert some attraction for each other.

These factors become significant when the space between gas particles is reduced, as in conditions of relatively high pressure and low temperature.

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Understandings and Fundamental Concepts

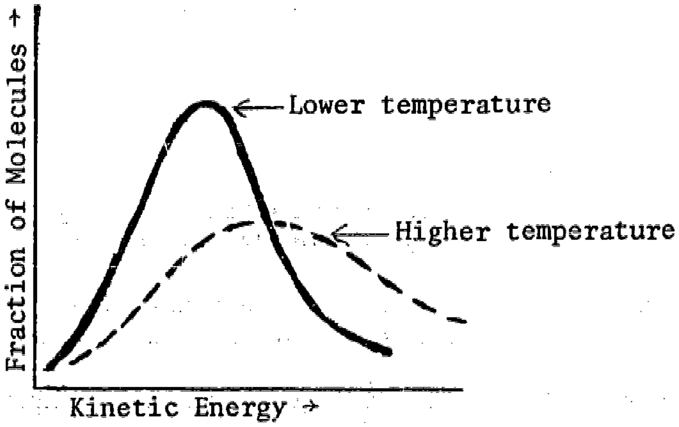
- A gas is composed of individual particles which are in continuous, random, straight-line motion.
- Collisions between gas particles may result in a transfer of energy between particles, but the net total energy of the system remains constant.
- The volume of the gas particles themselves is ignored in comparison with the volume of the space in which they are contained.
- Gas particles are considered as having no attraction for each other.

Deviations from the gas laws occur because the model is not perfect, and particularly because the gas particles do have volume and do exert some attraction for each other.

These factors become significant when the space between gas particles is reduced, as in conditions of relatively high pressure and low temperature.

Supplementary Information

Not all of the particles of a gas have the same kinetic energy, but the average kinetic energy is proportional to the measured Kelvin (Absolute) temperature of the gas. An example of the distribution of kinetic energy among the particles of a gas at two temperatures may be shown graphically.



A gas which would conform strictly to the model would be an ideal gas. However, the model does not exactly represent any gas under all conditions. No real gas is ideal under all conditions of temperature and pressure.

Experimental data (graphs) may be presented to show deviations from predicted values. Students should be encouraged to speculate as to the cause of these deviations. (Refer to Unit 3, Section III C, p. 21.)

Topics	Understandings and Fundamental Concepts	
9. Avogadro's Hypothesis	Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of particles.	For e numbe as th altho and 1
	The mass of matter that contains 6.02 x 10^{23} (Avogadro's number) structural particles is called a mole of matter.	Since struc etc.) struc mater
		This is ca
	A mole of particles of any gas occupies a volume of 22.4 liters at S.T.P.	See a 3, p.
B. Liquids	Liquids have definite volume but take the shape of the container.	Parti are i
1. Vapor pressure	When a liquid substance changes to a gas the process is called evaporation. Evaporation tends to take place at all temperatures.	The t gas p or sc
	In a closed system the vapor (gas) produced exerts a pressure which increases as the temperature of the liquid is raised and is specific for each substance and temperature.	*Demon at va*
2. Boiling point	A liquid will boil at the temperature at which the vapor pressure equals the pressure on the liquid.	
	The normal boiling point is the temperature at which the vapor pressure of the liquid equals one atmosphere.	Usual of a indic

Understandings and Fundamental Concepts	Supplementary Information
Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of particles.	For example, at the same temperature and pressure, the number of particles in 1 liter of hydrogen is the same as the number of particles in 1 liter of oxygen, although the individual particles of oxygen are heavier and larger than the individual particles of hydrogen.
The mass of matter that contains 6.02 x 10^{23} (Avogadro's number) structural particles is called a mole of matter.	Since it is inconvenient to work with individual structural particles (atoms, molecules, ions, electrons, etc.) chemists have chosen a unit containing many structural particles for comparing amounts of different materials.
	This unit contains 6.02 x 10^{23} structural particles and is called a mole.
A mole of particles of any gas occupies a volume of 22.4 liters at S.T.P.	See also Unit 5, Section II C, p. 34 and Section III A, 3, p. 37.
Liquids have definite volume but take the shape of the container.	Particles of a liquid have no regular arrangement and are in constant motion.
When a liquid substance changes to a gas the process is called evaporation. Evaporation tends to take place at all temperatures.	The term "vapor" is frequently used to refer to the gas phase of a substance that is normally a liquid or solid at room temperature.
In a closed system the vapor (gas) produced exerts a pressure which increases as the temperature of the liquid is raised and is specific for each substance and temperature.	*Demonstrations of the vapor pressure of a few liquids at various temperatures should be shown.*
A liquid will boil at the temperature at which the vapor pressure equals the pressure on the liquid.	
The normal boiling point is the temperature at which the vapor pressure of the liquid equals one atmosphere.	Usually when reference is made to the "boiling point" of a substance, it is the normal boiling point that is indicated.

Topics	Understandings and Fundamental Concepts	
3. Heat of vaporization	The energy required to vaporize a unit mass of liquid at constant temperature is called its heat of vaporization.	The energy involve required to overco and does not incre Thus there is no phase change. *Teachers may wish* *vaporization expe* *of vaporization f* *subject to examin*
C. Solids	Solids have definite shape and volume.	
1. Crystals	All true solids have a crystalline structure.	Certain materials super-cooled liqu
	Crystals contain particles arranged in a regular geometric pattern.	
	Particles are constantly vibrating even in the solid phase.	In solids, althou do not change the geometric pattern
2. Melting point	Melting point is the temperature at which a solid substance changes to a liquid.	Melting points ca which are obtaine page 3, if read f cooling curve.
		A melting point m at which the soli equilibrium.
3. Heat of fusion	The energy required to change a unit mass of a solid to a liquid at constant temperature is called its heat of fusion.	*Teachers may wish* *fusion experiment* *from experimental* *examination.*

Understandings and Fundamental Concepts	Supplementary Information
The energy required to vaporize a unit mass of liquid at constant temperature is called its heat of vaporization.	The energy involved in the change of phase is required to overcome binding forces between particles and does not increase their average kinetic energy. Thus there is no increase in temperature during the phase change. *Teachers may wish to have students determine heats of vaporization experimentally. Calculation of heats of vaporization from experimental data will not be subject to examination.*
Solids have definite shape and volume.	
All true solids have a crystalline structure.	Certain materials often considered solids are really super-cooled liquids; e.g., glass, some plastics.
Crystals contain particles arranged in a regular geometric pattern.	
Particles are constantly vibrating even in the solid phase.	In solids, although the particles are vibrating, they do not change their relative positions in the regular geometric pattern.
Melting point is the temperature at which a solid substance changes to a liquid.	Melting points can be determined from cooling curves which are obtained experimentally. The curve shown on page 3, if read from right to left, would illustrate a cooling curve. A melting point may also be defined as the temperature at which the solid and liquid phases can exist in equilibrium.
The energy required to change a unit mass of a solid to a liquid at constant temperature is called its heat of fusion.	*Teachers may wish to have students determine heats of fusion experimentally. Calculation of heats of fusion from experimental data will not be subject to examination.*

Topics	Understandings and Fundamental Concepts
4. Sublimation	Sublimation is a change from the solid phase directly to the gas phase without passing through an apparent liquid phase.
IV. Definition of Chemistry	Chemistry is the study of the composition, structure and properties of matter, the changes which matter undergoes, and the energy accompanying these changes.

Understandings and Fundamental Concepts

Sublimation is a change from the solid phase directly to the gas phase without passing through an apparent liquid phase.

Chemistry is the study of the composition, structure and properties of matter, the changes which matter undergoes, and the energy accompanying these changes.

Supplementary Information

Unit 2 - Atomic Structure

Topics	Understandings and Fundamental Concepts	
I. Atoms		
A. Introduction to atomic structure	Man's concept of the nature of the atom has undergone change and will probably continue to do so.	*Teachers sh history of structure.*
B. Fundamental particles	The atom is a complex unit of various particles.	*This materi Only a brie*
1. Electrons	An electron has a mass of 1/1836 of a proton and a unit negative charge.	
2. Nucleons	The particles which compose the nucleus are called nucleons.	
a. Protons	A proton has a mass of approximately one atomic mass unit and a unit positive charge.	An atomic m of the ^{12}C
b. Neutrons	A neutron has a mass of approximately one atomic mass unit and zero charge.	
†c. Other particles	Although protons and neutrons are the only nuclear particles that have been identified in an intact nucleus, other particles have been identified among the break-down products of certain nuclear disintegrations.	The relati and stabil current re
C. Structure of atoms	Atoms differ in the number of protons and neutrons in the nucleus and in the configuration of electrons surrounding the nucleus.	

† Optional

Unit 2 - Atomic Structure

Understandings and Fundamental Concepts	Supplementary Information
Man's concept of the nature of the atom has undergone change and will probably continue to do so.	*Teachers should acquaint pupils with at least a brief history of the development of the theory of atomic structure.*
The atom is a complex unit of various particles.	*This material is treated in the Science 7-8-9 program. Only a brief review should be necessary.*
An electron has a mass of 1/1836 of a proton and a unit negative charge.	
The particles which compose the nucleus are called nucleons.	
A proton has a mass of approximately one atomic mass unit and a unit positive charge.	An atomic mass unit is defined as exactly 1/12 the mass of the ^{12}C atom.
A neutron has a mass of approximately one atomic mass unit and zero charge.	
Although protons and neutrons are the only nuclear particles that have been identified in an intact nucleus, other particles have been identified among the break-down products of certain nuclear disintegrations.	The relationship of these particles to the structure and stability of the nucleus is the subject of much current research.
Atoms differ in the number of protons and neutrons in the nucleus and in the configuration of electrons surrounding the nucleus.	

Topics	Understandings and Fundamental Concepts	
1. "Empty space" concept	Most of the atom consists of empty space.	*Teachers should poin gold foil experiment mostly empty space.*
2. Nucleus	The mass of the atom is concentrated almost entirely in the nucleus.	The nature of the fo together is not adeq of much current rese
a. Atomic number	The atomic number indicates the number of protons in the nucleus.	*The teacher may wish X-ray spectra in rel atomic numbers. Mos charge on the nucleu which is interpreted nucleus.*
b. Isotopes	Isotopes are atoms with the same atomic number but a different number of neutrons. For a given element the number of protons in the nucleus remains constant, but the number of neutrons may vary.	The atomic number id in the number of neu
c. Mass number	The mass number indicates the total of the number of protons and neutrons.	Since the masses of approximately one, t of the isotope. The be calculated by sub the mass number.
d. Atomic mass	The atomic mass of an element is the weighted average mass of the naturally occurring isotopes of that element. This average is weighted according to the proportions in which the isotopes occur. Atomic mass is measured in atomic mass units based on ^{12}C equal to 12.000 atomic mass units.	Most elements occur This accounts for fr reference tables. I most abundant isotop by rounding off the nearest whole number

rstandings and Fundamental Concepts	Supplementary Information
of the atom consists of empty e.	*Teachers should point out the significance of Rutherford's gold foil experiments which indicated the atom to be mostly empty space.*
mass of the atom is concentrated st entirely in the nucleus.	The nature of the forces holding nuclear particles together is not adequately understood and is the subject of much current research.
atomic number indicates the number rotons in the nucleus.	*The teacher may wish to discuss Moseley's work on the X-ray spectra in relation to the determination of the atomic numbers. Moseley's experiments measured the charge on the nucleus in units of elementary charge, which is interpreted as the number of protons in the nucleus.*
opes are atoms with the same ic number but a different number eutrons.	The atomic number identifies the element. The difference in the number of neutrons affects the mass of the atom.
a given element the number of ons in the nucleus remains tant, but the number of neutrons vary.	
mass number indicates the total ne number of protons and neutrons.	Since the masses of the protons and neutrons are each approximately one, the mass number approximates the mass of the isotope. The number of neutrons in an atom can be calculated by subtracting the atomic number from the mass number.
atomic mass of an element is the nted average mass of the rally occurring isotopes of that ent. This average is weighted rding to the proportions in which isotopes occur.	Most elements occur naturally as mixtures of isotopes. This accounts for fractional atomic masses found in reference tables. In general, the mass number of the most abundant isotope of an element can be determined by rounding off the atomic mass of the element to the nearest whole number.
ic mass is measured in atomic units based on ^{12}C equal to 00 atomic mass units.	

Topics	Understandings and Fundamental Concepts	Su
	The gram atomic mass of an element is the mass in grams of Avogadro's number of atoms of that element as it occurs naturally.	The gram atomic mass i atomic mass. See also
3. Electrons	The electrons are outside the nucleus at various energy levels.	
	In a neutral atom the total number of electrons is equal to the number of protons in the nucleus.	
D. Bohr model of the atom	In the Bohr model, electrons are considered to revolve around the nucleus in one of several concentric circular orbits or shells.	*Models should be used should be pointed out are approximations and Students should be abl of atoms and ions and using the model.*
1. Principal energy levels	The orbits, or shells, are called principal energy levels, and can be denoted by the letters K, L, M, N, O, P, Q, or by the numbers 1, 2, 3, 4, 5, 6, 7.	
	Electrons in orbits near the nucleus are at lower energy levels than those in orbits farther from the nucleus.	When the electrons are levels the atom is sai
	When atoms absorb energy, electrons may shift to a higher energy level.	When electrons have ab energy levels, the ato
	The excited state is unstable, and the electrons fall back to lower energy levels.	
2. Quanta	Electrons can absorb energy only in discrete amounts called quanta.	
	When electrons return to a lower energy level, energy is emitted in quanta.	

nderstandings and Fundamental Concepts	Supplementary Information
e gram atomic mass of an element s the mass in grams of Avogadro's mber of atoms of that element as occurs naturally.	The gram atomic mass is numerically equal to the atomic mass. See also Unit 5, Section II A, p. 34.
e electrons are outside the cleus at various energy levels.	
a neutral atom the total number f electrons is equal to the number f protons in the nucleus.	
the Bohr model, electrons are onsidered to revolve around the ucleus in one of several concentric ircular orbits or shells.	*Models should be used to represent atomic structure. It should be pointed out to the student that these models are approximations and do not picture the actual atom. Students should be able to represent probable structure of atoms and ions and to indicate electronic changes using the model.*
e orbits, or shells, are called rincipal energy levels, and can be enoted by the letters K, L, M, N, O, , Q, or by the numbers 1, 2, 3, 4, 5, , 7.	
lectrons in orbits near the nucleus re at lower energy levels than those n orbits farther from the nucleus.	When the electrons are in the lowest available energy levels the atom is said to be in the "ground state."
hen atoms absorb energy, electrons ay shift to a higher energy level.	When electrons have absorbed energy and shifted to higher energy levels, the atom is said to be in an "excited state."
e excited state is unstable, and the lectrons fall back to lower energy evels.	
lectrons can absorb energy only in iscrete amounts called quanta.	
hen electrons return to a lower nergy level, energy is emitted in ua	

Topics	Understandings and Fundamental Concepts	Supple
3. Spectral lines	When electrons in an atom in the excited state return to lower energy levels, the energy is emitted as radiant energy of specific frequency, producing characteristic spectral lines which can be used to identify the element.	The study of spectral lin evidence regarding energy
E. Orbital model of the atom	The orbital model differs from the Bohr model in that it does not represent electrons as moving in planetary orbits around the nucleus.	Although Bohr's model acc hydrogen spectrum, it did of heavier and more compl
	The term orbital refers to the average region traversed by an electron. Electrons occupy orbitals that may differ in size, shape, or space orientation.	
1. Energy levels	The energy levels of electrons within an atom are represented by quantum numbers.	
a. Principal quantum number	The principal quantum number (n) represents the principal energy level.	The principal quantum num of the principal energy l Bohr atom and is the same periodic table.
b. Sublevels	The principal energy levels may be divided into sublevels.	Additional spectral lines atoms heavier than hydrog assuming that the princip into sublevels.
	The total number of possible sublevels for each principal quantum number is equal to the principal quantum number. These sublevels are designated by the letters s, p, d, and f.	The principal quantum num energy level called 1s; c up of two sublevels calle (n=3) is made up of three 3d; quantum number 4 (n=4 called 4s, 4p, 4d, and 4f energy level the lowest s

rstandings and Fundamental Concepts	Supplementary Information
electrons in an atom in the ted state return to lower energy ls, the energy is emitted as ant energy of specific frequency, ucing characteristic spectral s which can be used to identify element.	The study of spectral lines has provided much of the evidence regarding energy levels within the atom.
orbital model differs from the model in that it does not esent electrons as moving in etary orbits around the nucleus.	Although Bohr's model accounted for the lines of the hydrogen spectrum, it did not account for the spectra of heavier and more complicated atoms.
term orbital refers to the age region traversed by an tron. Electrons occupy orbitals may differ in size, shape, or e orientation.	
energy levels of electrons within tom are represented by quantum ers.	
principal quantum number (n) esents the principal energy l.	The principal quantum number (n) is equal to the number of the principal energy level as referred to under the Bohr atom and is the same as the period number on the periodic table.
principal energy levels may be ded into sublevels.	Additional spectral lines appearing in the spectrum of atoms heavier than hydrogen can be explained only by assuming that the principal energy levels are divided into sublevels.
total number of possible sublevels each principal quantum number is l to the principal quantum number. e sublevels are designated by the ers s, p, d, and f.	The principal quantum number, when n=1, comprises one energy level called 1s; quantum number 2 (n=2) is made up of two sublevels called 2s and 2p; quantum number 3 (n=3) is made up of three sublevels called 3s, 3p, and 3d; quantum number 4 (n=4) is made up of four sublevels called 4s, 4p, 4d, and 4f. Within a given principal energy level the lowest sublevel is the "s" sublevel

Topics

Understandings and Fundamental Concepts

and the highest is t
principal quantum nu
d, or f are used to d
in a particular subl
sublevels does not e
than 4.)

c. Orbitals

Each sublevel may consist of one or more orbitals with each orbital having a different spatial orientation.

Each electron occupies an orbital. An orbital can hold no more than two electrons. The number of orbitals within the same principal quantum number (n) is equal to n^2.

The s sublevel consists of 1 orbital; the p sublevel consists of 3 orbitals; the d sublevel consists of 5 orbitals; the f sublevel consists of 7 orbitals.

The maximum number o
energy levels and th
following table.

Principal quantum number (n)	Number of orbitals (n^2)
1	1
2	4
3	9
4	16

2. Electron configuration

Electron configuration of the atoms in order of their atomic numbers starting with hydrogen can be built up by adding one electron at a time according to the following rules:

- No more than two electrons can be accommodated in any orbital.
- The added electron is placed in the unfilled orbital of lowest energy.
- In a given sublevel, a second electron is *not* added to an orbital until each orbital in the sublevel contains one electron.

For the purposes of
will be limited to
number 1 - 20.

For teachers wishing
rations of elements
in Appendix I, page
levels of electrons

Understandings and Fundamental Concepts

Each sublevel may consist of one or more orbitals with each orbital having a different spatial orientation.

Each electron occupies an orbital. An orbital can hold no more than two electrons. The number of orbitals within the same principal quantum number (n) is equal to n^2.

The s sublevel consists of 1 orbital; the p sublevel consists of 3 orbitals; the d sublevel consists of 5 orbitals; the f sublevel consists of 7 orbitals.

Electron configuration of the atoms in order of their atomic numbers starting with hydrogen can be built up by adding one electron at a time according to the following rules:

- No more than two electrons can be accommodated in any orbital.
- The added electron is placed in the unfilled orbital of lowest energy.
- In a given sublevel, a second electron is *not* added to an orbital until each orbital in the sublevel contains one electron.

Supplementary Information

and the highest is the "f" sublevel. Both the principal quantum number and one of the letters s, p, d, or f are used to describe the energy of an electron in a particular sublevel. (The number of occupied sublevels does not exceed four even when n is greater than 4.)

The maximum number of electrons possible in the various energy levels and their distribution is shown in the following table.

Principal quantum number (n)	Number of orbitals (n^2)	s orbitals	p orbitals	d orbitals	f orbitals	Maximum number of electrons ($2n^2$)
1	1	1				2
2	4	1	3			8
3	9	1	3	5		18
4	16	1	3	5	7	32

For the purposes of examination, electron configurations will be limited to atoms of elements of atomic number 1 - 20.

For teachers wishing to discuss the electron configurations of elements of higher atomic numbers, the table in Appendix I, page 84 indicates the relative energy levels of electrons in the various sublevels.

Topics	Understandings and Fundamental Concepts	
	• No more than four orbitals are occupied in the outermost principal energy level of any atom.	In writ electro followi calcium- $3p^6$, 4s
3. Valence electrons	The electrons in the outermost principal energy level of an atom are referred to as the valence electrons.	The che valence The ter atom ex consist valence The val dot sym represe valence $\dot{\mathrm{Na}}$
†F. Probability model of the atom	It is not possible to determine both the exact location and the velocity of an electron within an atom.	This is Any exp electro The imp electro approac probabi
	It is possible to calculate mathematically the probability of finding an electron in a given region.	*Calcula involve mathema*
	An orbital is a region described by the probability distribution of an electron.	The orb cloud."

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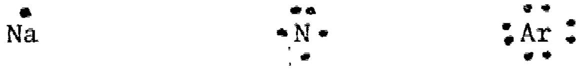
† Optional

Understandings and Fundamental Concepts	Supplementary Information
• No more than four orbitals are occupied in the outermost principal energy level of any atom.	In writing electron configurations, the number of electrons in a sublevel is indicated by a superscript following the designation of the sublevel. For example, calcium would be represented as $1s^2$ $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$; oxygen would be represented as $1s^2$, $2s^2$, $2p^4$.
The electrons in the outermost principal energy level of an atom are referred to as the valence electrons.	The chemical properties of an atom are related to the valence electrons. The term "kernel" is sometimes used to refer to an atom exclusive of the valence electrons. The kernel consists of the nucleus and all electrons except the valence electrons. The valence electrons may be represented by electron dot symbols in which the kernel of the atom is represented by the letter symbol for the element, and the valence electrons are represented by dots. For example: 
It is not possible to determine both the exact location and the velocity of an electron within an atom.	This is known as the Heisenberg Uncertainty Principle. Any experiment devised to determine the position of an electron alters both its position and its velocity. The impossiblity of determining the actual path of an electron has led to the discarding of the deterministic approach to atomic structure, and the adoption of a probability approach based on wave mechanics.
It is possible to calculate mathematically the probability of finding an electron in a given region.	*Calculations of probabilities based on wave mechanics involve concepts beyond the scope of high school mathematics.*
An orbital is a region described by the probability distribution of an electron.	The orbital is sometimes represented as an "electron cloud."

Topics	Understandings and Fundamental Concepts	
†G. Quantum numbers	Four quantum numbers are required to describe accurately the energy and most probable location of any electron of an atom. • Principal quantum number. This number (n) denotes the major axis of the orbital of an electron and represents the average distance from the nucleus. • Orbital quantum number. This number (ℓ) indicates the shape of the orbital. • Magnetic quantum number. This number (m) indicates how an orbital is oriented in space. • Spin quantum number. This number (s) describes the spin of the electron.	The orbital quantum as the azimuthal qu
H. Ionization energy	Ionization energy is the amount of energy required to remove the most loosely bound electron from an atom in the gaseous phase.	Ionization energy r electron. The seco removal of the seco Each successive ion previous one. In Elements in the *Ref* ionization energies Ionization energies properties of eleme D, pp. 26-27)
I. Electron affinity	Electron affinity is a measure of the energy released when an extra electron is added to an atom to form a negative ion.	Electron affinities been determined for elements.
J. Electro-negativity	Electronegativity is the ability of an atom to attract the electrons that form a bond between it and another atom.	Electronegativity v scale on which fluo element, is assigne

† Optional

Understandings and Fundamental Concepts	Supplementary Information
Four quantum numbers are required to describe accurately the energy and most probable location of any electron of an atom.	
• Principal quantum number. This number (n) denotes the major axis of the orbital of an electron and represents the average distance from the nucleus.	
• Orbital quantum number. This number (ℓ) indicates the shape of the orbital.	The orbital quantum number is sometimes referred to as the azimuthal quantum number.
• Magnetic quantum number. This number (m) indicates how an orbital is oriented in space.	
• Spin quantum number. This number (s) describes the spin of the electron.	
Ionization energy is the amount of energy required to remove the most loosely bound electron from an atom in the gaseous phase.	Ionization energy refers to the removal of the first electron. The second ionization energy refers to the removal of the second most loosely bound electron. Each successive ionization energy is greater than the previous one. In the table of "Representative" Elements in the *Reference Tables for Chemistry*, ionization energies are expressed in electron volts. Ionization energies are used to compare chemical properties of elements. (See Unit 4, Section II C and D, pp. 26-27.)
Electron affinity is a measure of the energy released when an extra electron is added to an atom to form a negative ion.	Electron affinities are difficult to measure and have been determined for only a relatively small number of elements.
Electronegativity is the ability of an atom to attract the electrons that form a bond between it and another atom.	Electronegativity values are based on an arbitrary scale on which fluorine, the most electronegative element, is assigned a value of 4.0.

Topics	Understandings and Fundamental Concepts	
		The electronegativi the reactivity of t electronegativity b predicting the natu between the two ele II B, pp. 19-20.)
II. Natural Radioactivity		*This topic is also* *Block L - Living wi*
	Radioactivity is the spontaneous disintegration of the nucleus of an atom with the emission of particles and rays.	When one element is a result of a chang is called transmuta
	Some naturally occurring elements are radioactive.	There are no stable numbers greater tha
A. Differences in emanations	Emanations differ from each other in mass, charge, penetrating power, and ionizing power.	*Students should be* *equations.*
	Nuclear disintegration of naturally radioactive atoms produces alpha particles, beta particles, and gamma radiations.	
1. Alpha decay	When an alpha particle is given off as the result of nuclear disintegration the reaction is called alpha decay.	An atom which emits alpha emitter. When an atom emits number is reduced b reduced by 4. Example: $^{226}_{88}\text{Ra} \longrightarrow$ 2
	Alpha particles are helium nuclei.	

Understandings and Fundamental Concepts

Radioactivity is the spontaneous disintegration of the nucleus of an atom with the emission of particles and rays.

Some naturally occurring elements are radioactive.

Emanations differ from each other in ass, charge, penetrating power, and onizing power.

uclear disintegration of naturally adioactive atoms produces alpha articles, beta particles, and amma radiations.

hen an alpha particle is given off s the result of nuclear disintegration he reaction is called alpha decay.

pha particles are helium nuclei.

Supplementary Information

The electronegativity of an element does not measure the reactivity of the element. The difference in electronegativity between two elements is used in predicting the nature of the bond that may be formed between the two elements. (See also Unit 3, Section II B, pp. 19-20.)

This topic is also treated in the Science 7, 8, 9 - Block L - Living with the Atom.

When one element is changed to another element as a result of a change in the nucleus, the change is called transmutation.

There are no stable isotopes known with atomic numbers greater than 83.

Students should be able to balance given nuclear equations.

An atom which emits an alpha particle is called an alpha emitter.

When an atom emits an alpha particle, the atomic number is reduced by 2 and the mass number is reduced by 4.

Example:

$$^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$$

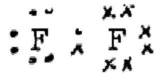
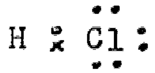
Topics	Understandings and Fundamental Concepts	
2. Beta decay	When a beta particle is given off as the result of nuclear disintegration, the reaction is called beta decay.	An atom which emits emitter. When an atom emits is increased by 1 same. Exampl e: $^{234}_{90}\mathrm{Th} \longrightarrow$
	In natural radioactivity beta particles are high speed electrons.	
3. Gamma radiation	Gamma rays are similar to high energy X-rays.	Gamma rays are not charge.
B. Separating emanations	Separation of emanations is possible by an electric field or magnetic field.	In an electric fie toward the negativ toward the positiv affected by the fi
C. Half-life	The half-life of a radioactive isotope is the time required for one-half of the nuclei of any given sample of that isotope to disintegrate.	*Simple examples sh* *mass of an isotope* *calculated.* A knowledge of the radioactive isotop calculate the age

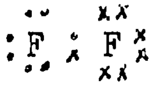
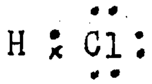
ndings and Fundamental Concepts	Supplementary Information
eta particle is given off as lt of nuclear disintegration, tion is called beta decay.	An atom which emits a beta particle is called a beta emitter. When an atom emits a beta particle the atomic number is increased by 1 and the mass number remains the same. Example: $^{234}_{90}\text{Th} \longrightarrow {}^{234}_{91}\text{Pa} \quad {}^{0}_{-1}\text{e}$
al radioactivity beta s are high speed electrons.	
ys are similar to high energy	Gamma rays are not particles and do not have mass or charge.
on of emanations is possible ectric field or magnetic	In an electric field, alpha particles are deflected toward the negative electrode, beta particles toward the positive electrode, while gamma rays are not affected by the field.
f-life of a radioactive isotope ime required for one-half of lei of any given sample of that to disintegrate.	*Simple examples should be used to demonstrate that the mass of an isotope remaining after a given time can be calculated.*
	A knowledge of the half-life and concentrations of a radioactive isotope in a substance enables one to calculate the age of the substance.

Topics	Understandings and Fundamental Concepts	
I. The Nature of Chemical Bonding	A chemical bond results from the simultaneous attraction of electrons to two nuclei.	
A. Energy stored in bonds	Chemical energy is potential energy. The energy is stored in chemical bonds.	*The* *ene* *the*
B. Energy changes in bonding	When a chemical bond is formed, energy is released. When a chemical bond is broken, energy is absorbed.	Whe the sep
C. Bonding and stability	Generally, systems at low energy levels are more stable than systems at high energy levels.	
	Generally, chemical changes will occur among atoms if the changes lead to a lower energy condition, and hence a more stable structure.	For nea it res lev exo dec dec exa amo the com
II. Bonds Between Atoms	The electrons involved in bond formation may be transferred from one atom to another, or may be shared equally or unequally between two atoms.	Whe the mor max pla max
A. Ionic	An ionic bond is formed by the transfer of one or more electrons between atoms.	The ion fer

Unit 3 - Bonding

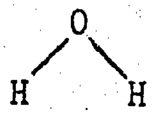
Understandings and Fundamental Concepts	Supplementary Information
A chemical bond results from the simultaneous attraction of electrons to two nuclei.	
Chemical energy is potential energy. The energy is stored in chemical bonds.	*The teacher may wish to use the analogy of potential energy stored in an object when it is lifted against the force of gravity.*
When a chemical bond is formed, energy is released. When a chemical bond is broken, energy is absorbed.	When two atoms are held together by a chemical bond, they are at a lower energy condition than when they are separated.
Generally, systems at low energy levels are more stable than systems at high energy levels.	
Generally, chemical changes will occur among atoms if the changes lead to a lower energy condition, and hence a more stable structure.	For example, when an element such as fluorine, with a nearly filled outer shell, reacts with another element it usually releases energy in the process because the resulting state of the compound is at a lower energy level. Such a reaction would be an example of an exothermic reaction. When the compound thus formed is decomposed, energy must be put into the compound to decompose or break it up. Therefore this would be an example of an endothermic reaction. Exactly the same amount of energy would be required to force this endothermic reaction to proceed as was released when the compound was formed.
The electrons involved in bond formation may be transferred from one atom to another, or may be shared equally or unequally between two atoms.	When atoms of the elements enter into chemical reaction, they do so in a manner that results in their becoming more like inert gas atoms in that they contain their maximum complement of valence electrons. This would place them in a condition of lowest energy content and maximum stability.
An ionic bond is formed by the transfer of one or more electrons between atoms.	The transfer of electrons results in the formation of ions. In ionic bonding the number of electrons transferred is such that the atoms involved achieve an inert

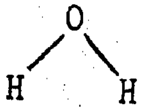
Topics	Understandings and Fundamental Concepts	Supplementary In
		gas configuration. Since the ion ha electron configuration than the atom the ion differ from those of the ato
	Ionic solids have high melting points, but as solids they do not conduct electricity.	In the geometric structure of the so ions form the crystal lattice and ar relatively fixed positions by electr When melted, vaporized, or dissolved lattice is destroyed and the ions mo of ionic solids are sodium chloride
B. Covalent	A covalent bond is formed when two atoms share electrons.	
	When electrons are shared between atoms of equal electronegativity, they are shared equally and the resulting bond is nonpolar.	An example of a nonpolar covalent bo fluorine molecule: 
	When electrons are shared between atoms of unequal electronegativity, they are shared unequally and the resulting bond is polar.	An example of a polar covalent bond hydrogen chloride molecule: 
	When the two shared electrons forming a covalent bond are both donated by one of the atoms, this bond is called a coordinate covalent bond.	A coordinate covalent bond, once for from an ordinary covalent bond. The the source of the electrons involved The coordinate covalent bond is freq the bonding within radicals and poly The coordinate covalent bond is of i acid-base theory. (See Unit 7, Secti
	The ionic or covalent character of the bond can be estimated from differences in electronegativity of the reacting species.	Electronegativity differences give character of the bond. Electronegat of 1.7 or greater indicate a bond th ionic in character. Differences les that the bond is predominantly coval exceptions to this may be found. Fo

standings and Fundamental Concepts	Supplementary Information
	gas configuration. Since the ion has a different electron configuration than the atom, the properties of the ion differ from those of the atom.
solids have high melting points, s solids they do not conduct ricity.	In the geometric structure of the solid ionic crystal, ions form the crystal lattice and are held in relatively fixed positions by electrostatic attraction. When melted, vaporized, or dissolved in water the crystal lattice is destroyed and the ions move freely. Examples of ionic solids are sodium chloride and magnesium oxide.
alent bond is formed when two share electrons.	
electrons are shared between atoms ual electronegativity, they are d equally and the resulting bond npolar.	An example of a nonpolar covalent bond is found in the fluorine molecule: 
electrons are shared between atoms equal electronegativity, they are d unequally and the resulting is polar.	An example of a polar covalent bond is found in the hydrogen chloride molecule: 
the two shared electrons forming valent bond are both donated by one e atoms, this bond is called a dinate covalent bond.	A coordinate covalent bond, once formed, is no different from an ordinary covalent bond. The difference lies in the source of the electrons involved in the bond. The coordinate covalent bond is frequently involved in the bonding within radicals and polyatomic ions. The coordinate covalent bond is of importance in modern acid-base theory. (See Unit 7, Section III B, p. 60)
ionic or covalent character of the can be estimated from differences lectronegativity of the reacting ies.	Electronegativity differences give an *indication* of the character of the bond. Electronegativity differences of 1.7 or greater indicate a bond that is predominantly ionic in character. Differences less than 1.7 indicate that the bond is predominantly covalent. (Some exceptions to this may be found. For example, the metal

Topics	Understandings and Fundamental Concepts	Supple
		hydrides, with an electro less than 1.7, are predomi
1. Molecular substances	A molecule may be defined as a discrete particle formed by covalently bonded atoms.	A molecule has also been particle of an element or pendent motion.
		Stable molecules usually atom has the electron con when the shared electrons counted as belonging to b covalent bond.
		Examples of molecules are $C_6H_{12}O_6$.
		The teacher may wish to p exceptions to the octet r
	Molecular substances may exist as gases, liquids, or solids, depending on the attraction that exists between the molecules.	The forces of attraction general properties of the developed below under Sec 21-22.
	Generally, molecular solids are soft, electrical insulators, poor heat conductors, and have low melting points.	
2. Network solids	Certain solids consist of covalently bonded atoms linked in a network which extends throughout the sample with an absence of simple discrete particles. Such a substance is said to be a network solid, sometimes called a macro molecule.	
	Generally, network solids are hard, electrical insulators, poor conductors of heat, and have high melting points.	Examples of network solid diamonds, silicon carbide (SiO_2).

Understandings and Fundamental Concepts	Supplementary Information
	hydrides, with an electronegativity difference of less than 1.7, are predominantly ionic.)
A molecule may be defined as a discrete particle formed by covalently bonded atoms.	A molecule has also been defined as the smallest particle of an element or compound capable of independent motion.
	Stable molecules usually have structures such that each atom has the electron configuration of an inert gas atom when the shared electrons forming each covalent bond are counted as belonging to both atoms connected by that covalent bond.
	Examples of molecules are H_2, NH_3, H_2O, HCl, CCl_4, S_8, $C_6H_{12}O_6$.
	The teacher may wish to point out that there are exceptions to the octet rule in certain molecular species.
Molecular substances may exist as gases, liquids, or solids, depending on the attraction that exists between the molecules.	The forces of attraction between molecules and the general properties of the resulting substances are developed below under Section III of this unit, pp. 21-22.
Generally, molecular solids are soft, electrical insulators, poor heat conductors, and have low melting points.	
Certain solids consist of covalently bonded atoms linked in a network which extends throughout the sample with an absence of simple discrete particles. Such a substance is said to be a network solid, sometimes called a macro molecule.	
Generally, network solids are hard, electrical insulators, poor conductors of heat, and have high melting points.	Examples of network solids are asbestos, graphite, diamonds, silicon carbide (SiC), and silicon dioxide (SiO_2).

Topics	Understandings and Fundamental Concepts	Supplement
C. Metallic	Metallic bonding occurs between atoms which have vacant valence orbitals and low ionization energies.	A metal consists of an arrange which are located at the cryst are immersed in a "sea" of mob mobile electrons can be consid whole crystal rather than to i mobility of electrons distingu from an ionic or covalent bond
III. Molecular Attraction	Groups of atoms covalently bonded in a molecule may in turn be attracted to similar molecules or to ions.	
A. Dipoles	The asymmetric distribution of electrical charge in a molecule gives rise to a molecule which is polar in nature and is referred to as a dipole.	A molecule composed of only tw the bond between the atoms is Molecules composed of more tha nonpolar even though the indiv the shape of the molecule is s distribution of charge results nonpolar molecule that can be while H_2O is a polar molecule
	Dipoles attract one another with electrostatic forces.	
B. Hydrogen bonding	Hydrogen bonds are formed between molecules in which hydrogen is covalently bonded to an element of small atomic radius and high electronegativity.	When a hydrogen atom is bonded negative atom, the hydrogen ha the electron pair that it is a As such it can be attracted to of an adjacent molecule. This boiling point of H_2O as compar of H_2S. Hydrogen bonding is hydrogen with fluorine, oxygen
C. Van der Waals' forces	Even in the absence of dipole attraction and hydrogen bonding, as in nonpolar molecules, weak attractive forces exist between molecules. These forces are called van der Waal's forces.	Van der Waal's forces make it small nonpolar molecules (such oxygen, etc.) to exist in the under conditions of low temper

nderstandings and Fundamental Concepts	Supplementary Information
etallic bonding occurs between atoms hich have vacant valence orbitals nd low ionization energies.	A metal consists of an arrangement of positive ions which are located at the crystal lattice sites, and are immersed in a "sea" of mobile electrons. These mobile electrons can be considered as belonging to the whole crystal rather than to individual atoms. This mobility of electrons distinguishes the metallic bond from an ionic or covalent bond.
roups of atoms covalently bonded in a olecule may in turn be attracted to imilar molecules or to ions.	
e asymmetric distribution of lectrical charge in a molecule gives ise to a molecule which is polar in ture and is referred to as a dipole.	A molecule composed of only two atoms will be a dipole if the bond between the atoms is polar. Molecules composed of more than two atoms ma[illegible]be nonpolar even though the individual bonds ar[illegible]polar, if the shape of the molecule is such that symme[illegible]c distribution of charge results. For exampl[illegible]O_2 is a nonpolar molecule that can be represented a[illegible] O = C = O while H_2O is a polar molecule that may be r[illegible]resented as
poles attract one another with ectrostatic forces.	
drogen bonds are formed between lecules in which hydrogen is valently bonded to an element of all atomic radius and high ectronegativity.	When a hydrogen atom is bonded to a highly electronegative atom, the hydrogen has such a small share of the electron pair that it is almost like a bare proton. As such it can be attracted to the electronegative atom of an adjacent molecule. This accounts for the high boiling point of H_2O as compared with the boiling point of H_2S. Hydrogen bonding is important in compounds of hydrogen with fluorine, oxygen, or nitrogen.
en in the absence of dipole traction and hydrogen bonding, as in npolar molecules, weak attractive ces exist between molecules. These ces are called van der Waal's forces.	Van der Waal's forces make it possible for species of small nonpolar molecules (such as hydrogen, helium, oxygen, etc.) to exist in the liquid and solid phase under conditions of low temperature and high pressure.

Topics	Understandings and Fundamental Concepts	S
	Van der Waal's forces increase with an increasing number of electrons (hence with increasing molecular size and molecular mass) and with decreasing distance between the molecules.	The effect of m van der Waal's boiling points as the alkane s molecular mass. halogens with i explained in te
		Van der Waal's distribution of dipole attracti
D. Molecule-ion attraction	Polar covalent compounds, when interacting with ionic compounds, attract ions from these compounds and form a solution.	Ionic compounds as water, alcoh compounds. The dissolved is at adjacent polar attracted to th Water is the po dissolve these compound is dis destroyed and w hydrated ions. this attraction positive or neg molecules aroun This process is
IV. Directional Nature of Covalent Bonds	Generally the geometric structure of covalent substances which results from the directional nature of the covalent bond helps to explain properties of the resulting molecule.	The polarity of the molecule.
V. Chemical Formula	A chemical formula is both a qualitative and a quantitative expression of the composition of an element or a compound.	
A. Symbol	A symbol may represent one atom or one mole of atoms of an element.	One mole of at (6.02×10^{23}).

ics	Understandings and Fundamental Concepts	Supplementary Information
	Van der Waal's forces increase with an increasing number of electrons (hence with increasing molecular size and molecular mass) and with decreasing distance between the molecules.	The effect of molecular size on the magnitude of the van der Waal's forces accounts for the increasing boiling points of a series of similar compounds (such as the alkane series of hydrocarbons) with increasing molecular mass. The increasing boiling points of the halogens with increasing molecular mass can also be explained in terms of increasing van der Waal's forces. Van der Waal's forces appear to be due to chance distribution of electrons resulting in momentary dipole attractions.
e-ion ion	Polar covalent compounds, when interacting with ionic compounds, attract ions from these compounds and form a solution.	Ionic compounds are generally soluble in liquids such as water, alcohol, and liquid ammonia, which are polar compounds. The negative ion of the substance being dissolved is attracted to the positive end of the adjacent polar molecules, while the positive ion is attracted to the negative end of the polar molecules. Water is the polar substance most commonly used to dissolve these ionic compounds. When an ionic compound is dissolved in water, its crystal lattice is destroyed and water molecules surround each ion forming hydrated ions. It is because water is a dipole that this attraction between the water molecules and the positive or negative ion exists. The orienting of water molecules around ions is called hydration of the ions. This process is important in aqueous chemistry.
al Nature ent Bonds	Generally the geometric structure of covalent substances which results from the directional nature of the covalent bond helps to explain properties of the resulting molecule.	The polarity of water is explained by the shape of the molecule.
Formula	A chemical formula is both a qualitative and a quantitative expression of the composition of an element or a compound.	
	A symbol may represent one atom or one mole of atoms of an element.	One mole of atoms is Avogadro's number of atoms (6.02×10^{23}).

Topics	Understandings and Fundamental Concepts	Suppleme
B. Formula	A formula is a statement in chemical symbols which represents the composition of a substance.	*Where necessary, teachers sh subscripts, brackets, and co* <u>*Science 7, 8, 9, Block J.*</u>
1. Molecular	A molecular formula indicates the total number of atoms of each element needed to form the molecule.	
2. Empirical	An empirical formula represents the ratio in which the atoms combine to form a compound.	Empirical formulas are used which do not exist as discre
C. Nomenclature	The chemical name of a compound generally indicates the chemical composition of the substance.	*Students should be familiar nomenclature used in naming and salts.*

Students should know the rel of an acid and the name of t For example:

Name of Acid	*N*
Hydrochloric	*S*
Hypochlorous	*S*
Chlorous	*S*
Chloric	*S*
Perchloric	*S*

In naming the salts of metal one oxidation number the Sto In this system, Roman numera number of the metal ion. Fo iron (II) oxide, and Fe_2O_3 i

The Stock system should be e compounds of two nonmetals.

nitrogen (I) oxide, NO is na and NO_2 is named nitrogen (I standings and Fundamental Concepts

mula is a statement in chemical ls which represents the compo- n of a substance.

ecular formula indicates the total r of atoms of each element needed rm the molecule.

pirical formula represents the in which the atoms combine to a compound.

hemical name of a compound gener- indicates the chemical composition e substance.

Supplementary Information

Where necessary, teachers should review the use of subscripts, brackets, and coefficients as treated in *Science 7, 8, 9, Block J.*

Empirical formulas are used to represent ionic compounds which do not exist as discrete molecular entities.

Students should be familiar with the traditional nomenclature used in naming the common acids, bases, and salts.

Students should know the relationship between the name of an acid and the name of the corresponding salt. For example:

Name of Acid	*Name of Sodium Salt*
Hydrochloric	*Sodium chloride*
Hypochlorous	*Sodium hypochlorite*
Chlorous	*Sodium chlorite*
Chloric	*Sodium chlorate*
Perchloric	*Sodium perchlorate*

In naming the salts of metals which may have more than one oxidation number the Stock system should be used. In this system, Roman numerals indicate the oxidation number of the metal ion. For example, FeO is named iron (II) oxide, and Fe_2O_3 is named iron (III) oxide.

The Stock system should be extended to the naming of compounds of two nonmetals. For example, N_2O is named nitrogen (I) oxide, NO is named nitrogen (II) oxide, and NO_2 is named nitrogen (IV) oxide.

Topics	Understandings and Fundamental Concepts	Supplem
VI. Chemical Equations	An equation represents the qualitative and quantitative changes in bonding and energy that take place in a chemical reaction.	*Simple equations that should be introduced;* $2H_2 + O_2 \rightarrow 2H_2O$ *Balancing redox react Section III, pp. 69-7*
	Equations must conform to the laws of conservation of mass and charge.	Energy is usually omi are not concerned wit In an equation, it is phase of reactants an using the symbols (s) gas; and (aq) - in aq shown above could be $2H_2(g) + O_2(g) \rightarrow 2H_2O$

	Understandings and Fundamental Concepts	Supplementary Information
ions	An equation represents the qualitative and quantitative changes in bonding and energy that take place in a chemical reaction.	*Simple equations that can be balanced by inspection should be introduced; e.g.,* $2H_2 + O_2 \rightarrow 2H_2O + heat$ *Balancing redox reactions is taken up in Unit 8, Section III, pp. 69-70.*
	Equations must conform to the laws of conservation of mass and charge.	Energy is usually omitted from the equation when we are not concerned with it. In an equation, it is often desirable to indicate the phase of reactants and products. This may be done by using the symbols (s) - solid; (ℓ) - liquid; (g) - gas; and (aq) - in aqueous solution. The equation shown above could be written: $2H_2(g) + O_2(g) \rightarrow 2H_2O\ (\ell) + heat$

Unit 4 - Periodic Table

Topics	Understandings and Fundamental Concepts	
I. Development of Periodic Table	The periodic table of the elements has passed through many stages of development evolving into the present form.	Observed regularitie Mendeleev and others to be functions of t lished that properti functions of the ato
	The atomic number is the basis of the arrangement in the present form of the periodic table.	
	The properties of the elements depend on the structure of the atom, and vary with the atomic number in a systematic	
II. Properties of Elements in the Periodic Table	[illegible]zontal rows of the periodic table are called periods or rows. The properties of elements change systematically through a period.	
	The vertical columns of the periodic table are called groups or families. The elements of a group exhibit similar or related properties.	
A. Atomic radii	The atomic radius is one-half the measured inter-nuclear distance in the solid phase. Atomic radius is a periodic property of the elements.	The relation between be interpreted in te in the orbitals of a
	Within a single period of the periodic table, the atomic radius generally decreases as the atomic number increases.	Within any one perio orbitals are arrange same number of fille left to right in the due to the increasin electrons more tight increased attraction between the added el thus the atomic radi

Unit 4 - Periodic Table

Understandings and Fundamental Concepts	Supplementary Information
The periodic table of the elements has passed through many stages of development evolving into the present form.	Observed regularities in properties of elements led Mendeleev and others to consider these regularities to be functions of the atomic mass. Moseley established that properties of elements are periodic functions of the atomic number.
The atomic number is the basis of the arrangement in the present form of the periodic table.	
The properties of the elements depend on the structure of the atom, and vary with the atomic number in a systematic way.	
The horizontal rows of the periodic table are called periods or rows. The properties of elements change systematically through a period.	
The vertical columns of the periodic table are called groups or families. The elements of a group exhibit similar or related properties.	
The atomic radius is one-half the measured inter-nuclear distance in the solid phase. Atomic radius is a periodic property of the elements.	The relation between atomic radius and atomic number can be interpreted in terms of the arrangement of electrons in the orbitals of atoms and in terms of nuclear charge.
Within a single period of the periodic table, the atomic radius generally decreases as the atomic number increases.	Within any one period the electrons in the outer orbitals are arranged around a kernel containing the same number of filled levels. As one proceeds from left to right in the period, increase in nuclear charge due to the increasing number of protons pulls the electrons more tightly around the nucleus. This increased attraction more than balances the repulsion between the added electron and other electrons, and thus the atomic radius is reduced.

Topics	Understandings and Fundamental Concepts	
	The members of any group in the periodic table generally show an increase in atomic radius with an increase in atomic number.	For a membe levels are fa increa increa
		Measu *and d* *result* *unders* *differ*
B. Ionic radii	A loss or gain of electrons by an atom causes a corresponding change in size.	Metal ions. corres
		Nonme become the c
		Atomi (Å) u
C. Metals	Metal atoms possess relatively low ionization energies.	More t
	Metal atoms tend to lose electrons to form positive ions when combining with other elements.	Metal eleme
	Metals usually possess the properties of high thermal and electrical conductivity, metallic luster, malleability, and ductility.	All me

Understandings and Fundamental Concepts	Supplementary Information
The members of any group in the periodic table generally show an increase in atomic radius with an increase in atomic number.	For a group of elements, the atoms of each successive member have a larger kernel containing more filled levels. Hence the electrons in the unfilled orbitals are farther from the nucleus. This results in an increase in atomic radius as the atomic number increases among the elements in a group. *Measurement of inter-nuclear distances is difficult, and different methods may give slightly different results. For purposes of this course the important understanding is the relationship of atomic number to differences in atomic size.*
A loss or gain of electrons by an atom causes a corresponding change in size.	Metal atoms lose one or more electrons when they form ions. Ionic radii of metals are smaller than the corresponding atomic radii. Nonmetal atoms gain one or more electrons when they become ions. Ionic radii of nonmetals are larger than the corresponding atomic radii. Atomic and ionic radii are usually measured in Angstrom ($\mathring{A}$) units (1$\mathring{A}$ = 10^{-8} cm.)
Metal atoms possess relatively low ionization energies.	More than two-thirds of the elements are metals.
Metal atoms tend to lose electrons to form positive ions when combining with other elements.	Metallic properties are most pronounced in those elements on the lower left side of the periodic table.
Metals usually possess the properties of high thermal and electrical conductivity, metallic luster, malleability, and ductility.	All metals except mercury are solids at room temperature.

Topics	Understandings and Fundamental Concepts	Suppl
D. Nonmetals	Nonmetal atoms possess high ionization energies and high electron affinities, and thus have high electronegativities.	Nonmetallic propert elements in the upp table.
	Nonmetal atoms tend to gain electrons when in combination with metals, or to share electrons when in combination with other elements.	Nonmetals tend to b network solids. (Th volatile liquid at
	Nonmetals in the solid phase tend to be brittle, to have low thermal and electrical conductivities, and to lack metallic luster.	
E. Metalloids	Metalloids are those elements which have some properties characteristic of metals and other properties characteristic of nonmetals.	Examples of metallo and tellurium.
III. Chemistry of a Group (Family)		*Minimum requirement* *individual groups a* *sections of this se* *will discuss the pr* *important compounds* *the* <u>*principles*</u> *outl* *properties of the c* *to uses in home and* *Unit 10 of this syl* *point. Technologic* *in home and industr*
	The elements in each group exhibit related chemical properties. Similarities in chemical properties within a group are associated with similarity in the number of valence electrons.	Related chemical pr similarity in the t of a group. For ex chlorides having th represents any memb IIA form chlorides

lerstandings and Fundamental Concepts	Supplementary Information
metal atoms possess high ionization rgies and high electron affinities, thus have high electronegativities.	Nonmetallic properties are most pronounced in those elements in the upper right corner of the periodic table.
metal atoms tend to gain electrons n in combination with metals, or to re electrons when in combination with er elements.	Nonmetals tend to be gases, molecular solids, or network solids. (The exception is bromine which is a volatile liquid at room temperature.)
metals in the solid phase tend to be ttle, to have low thermal and ctrical conductivities, and to lack allic luster.	
alloids are those elements which e some properties characteristic metals and other properties racteristic of nonmetals.	Examples of metalloids are boron, silicon, arsenic, and tellurium.
	Minimum requirements for the detailed study of individual groups are indicated in the following subsections of this section. It is expected that teachers will discuss the properties of the elements and their important compounds in sufficient detail to illustrate the <u>*principles*</u> *outlined in these subsections. The properties of the compounds studied should be related to uses in home and industry. Some material from Unit 10 of this syllabus may be incorporated at this point. Technological details of preparations, and uses in home and industry will* <u>*not*</u> *be subject to examination.*
elements in each group exhibit ated chemical properties. Similar- es in chemical properties within group are associated with similarity the number of valence electrons.	Related chemical properties are illustrated by the similarity in the type of compound formed by the members of a group. For example, the elements in Group IA form chlorides having the general formula MCl where M represents any member of the group. Elements in Group IIA form chlorides having the general formula MCl_2.

Topics	Understandings and Fundamental Concepts	
	In general, the properties of elements in a group change progressively as the atomic number increases.	Prope expla atomi that group does group the e close the r elect
	In the "A" groups of the periodic table, as the atomic number increases,	
	• the radius of the atom increases.	
	• the ionization energy of the element generally decreases.	The i numbe elect incre are a
	• the electronegativity of the element generally decreases.	It sh the a react elect from
	• the elements tend to have more metallic properties.	
A. Groups IA and IIA	Groups IA and IIA include the most reactive metals.	Eleme Eleme metal Becau occur

Understandings and Fundamental Concepts	Supplementary Information
In general, the properties of elements in a group change progressively as the atomic number increases.	Properties of the members of a group should be explained in terms of bonding, electronegativity, atomic size, and so forth. It should be pointed out that anomalies in the properties of elements within a group do occur. For example, in Group IIIA - boron does not form a +3 ion as do other members of the group. These anomalies occur most frequently among the elements in Period 2 because of the relative closeness of the valence electrons to the nucleus, and the relatively small shielding effect of the two electrons in the K shell.
In the "A" groups of the periodic table, as the atomic number increases,	
• the radius of the atom increases.	
• the ionization energy of the element generally decreases.	The ionization energy decreases with increasing atomic number because of the increased distance of the valence electron(s) from the nucleus and also because of the increased shielding effect as new occupied energy levels are added to the atom.
• the electronegativity of the element generally decreases.	It should be pointed out that the relative tendency of the atoms to form compounds (sometimes called the reactivity of the element) cannot be predicted from the electronegativity of the atom, but should be ascertained from the Standard Oxidation Potential Tables.
• the elements tend to have more metallic properties.	
Groups IA and IIA include the most reactive metals.	Elements in Group IA are called the alkali metals. Elements in Group IIA are called the alkaline earth metals. Because of their reactivity Group IA and IIA elements occur in nature only in compounds.

Topics	Understandings and Fundamental Concepts	Sup
	The elements in both groups have relatively low ionization energies and electronegativities. They lose electrons readily to form ionic compounds that are relatively stable.	The elements in these g gas structure by covale compounds.
	Generally, the reactivity within both groups tends to increase with an increase in atomic number.	Exceptions to this occu reactivity of lithium i lithium ion and its cor
	In the same period, each Group IA metal is more reactive than the corresponding Group IIA metal.	
	The elements in both groups are usually reduced to their free state by the electrolysis of their fused compounds.	
B. Group VA or VIA		*Minimum requirements ar VA or VIA, but not both*
1. Group VA	The elements in Group VA show a marked progression from nonmetallic to metallic properties with increasing atomic number.	Nitrogen and phosphorus and antimony are classi metallic in both appear
	The element nitrogen is relatively inactive at room temperature. The element phosphorus is more reactive than nitrogen at room temperature.	In general, the reacti group decreases with i and phosphorus are an e exists as a diatomic m between the two atoms. larger atomic size, do molecule at room tempe molecule, P_4. The high triple N≡N bond expla nitrogen.
	Nitrogen compounds are essential constitutents of all living matter.	*Students who have stud should have some knowl*

Understandings and Fundamental Concepts	Supplementary Information
The elements in both groups have relatively low ionization energies and electronegativities. They lose electrons readily to form ionic compounds that are relatively stable.	The elements in these groups cannot achieve an inert gas structure by covalent bonding, and form only ionic compounds.
Generally, the reactivity within both groups tends to increase with an increase in atomic number.	Exceptions to this occur in Group IA. The high reactivity of lithium is due to the small size of the lithium ion and its corresponding high hydration energy.
In the same period, each Group IA metal is more reactive than the corresponding Group IIA metal.	
The elements in both groups are usually reduced to their free state by the electrolysis of their fused compounds.	
	Minimum requirements are limited to the study of Group VA or VIA, but not both.
The elements in Group VA show a marked progression from nonmetallic to metallic properties with increasing atomic number.	Nitrogen and phosphorus are typical nonmetals; arsenic and antimony are classified as metalloids; bismuth is metallic in both appearance and properties.
The element nitrogen is relatively inactive at room temperature.	In general, the reactivity of nonmetals in the same group decreases with increasing atomic number. Nitrogen and phosphorus are an exception to this rule. Nitrogen exists as a diatomic molecule with a triple bond between the two atoms. Phosphorus, because of its larger atomic size, does not exist as a diatomic molecule at room temperature, but exists as a tetratomic molecule, P_4. The high energy required to break the triple $N\equiv N$ bond explains the relative inactivity of nitrogen.
The element phosphorus is more reactive than nitrogen at room temperature.	
Nitrogen compounds are essential constitutents of all living matter.	*Students who have studied biology or science 7, 8, 9 should have some knowledge of the nitrogen cycle.*

Topics	Understandings and Fundamental Concepts	
	Generally, nitrogen compounds are relatively unstable.	The instab useful as
	Phosphorus compounds are essential constituents of all living matter.	*Students u with calci teeth, an and RNA.*
2. Group VIA	The elements in Group VIA show a marked progression from nonmetallic to metallic properties with an increase in atomic number.	Oxygen an tellurium metallic p
	The element oxygen is an active nonmetal.	Oxygen fo existence high react productio Because o pounds alw when combi
	Sulfur is less reactive than oxygen.	Sulfur in oxidation
	Selenium and tellurium are rare elements.	Selenium a when combi they show
	Polonium is a radioactive element.	Polonium i Polonium i
C. Group VIIA	The elements in Group VIIA are typical nonmetals.	Group VII Although t atomic nu metal. (As life. It propertie

Understandings and Fundamental Concepts	Supplementary Information
Generally, nitrogen compounds are relatively unstable.	The instability of many nitrogen compounds makes them useful as explosives.
Phosphorus compounds are essential constituents of all living matter.	*Students who have studied biology will be familiar with calcium phosphate as a constituent of bones and teeth, and with the phosphate linkage present in DNA and RNA.*
The elements in Group VIA show a marked progression from nonmetallic to metallic properties with an increase in atomic number.	Oxygen and sulfur are typical nonmetals; selenium and tellurium are classified as metalloids; polonium shows metallic properties.
The element oxygen is an active nonmetal.	Oxygen forms compounds with most elements. The existence of oxygen in its free state, in spite of its high reactivity, is explained by the continuous production of oxygen by plants during photosynthesis. Because of its high electronegativity, oxygen in compounds always shows a negative oxidation state except when combined with fluorine.
Sulfur is less reactive than oxygen.	Sulfur in compounds shows both negative and positive oxidation states.
Selenium and tellurium are rare elements.	Selenium and tellurium show a negative oxidation state when combined with hydrogen. In most other compounds they show positive oxidation states.
Polonium is a radioactive element.	Polonium is an alpha emitter. Polonium is a degradation product of uranium.
The elements in Group VIIA are typical nonmetals.	Group VIIA is known as the halogen family.
	Although the metallic character increases with increasing atomic number, none of the elements in the group is a metal. (Astatine is radioactive with a very short half-life. It has not been found in nature and its properties are not well known.)

Topics	Understandings and Fundamental Concepts	Supplemen
	The elements in Group VIIA have relatively high electronegativities.	Fluorine has the highes element, and in compoun oxidation state. The other elements of t oxidation states, in co negative elements. The tion states of the halo increasing atomic numbe
	The physical form of the free element, at room temperature, varies with increasing atomic number.	At room temperature flu bromine is a liquid and The change in physical van der Waal's forces.
	The elements are usually prepared from the corresponding halide ion by removing one of the electrons from the ion.	Because of their high r nature only in compound Since fluorine is the m is no chemical oxidizin fluoride ion to fluorin electrolysis of its fus Chlorine, bromine, and i chemical methods.
D. Transition elements	Transition elements are those elements in which the two outermost shells may be involved in a chemical reaction.	*It is recommended that trate the behavior of t elements in period 4. transition elements are differently in various* The transition elements Group VIII of the peri
	Transition elements generally exhibit multiple oxidation states.	
	The ions of transition elements are usually colored, both in solid compounds and in solution.	

rstandings and Fundamental Concepts	Supplementary Information
elements in Group VIIA have rela- ly high electronegativities.	Fluorine has the highest electronegativity of any element, and in compounds can show only a negative oxidation state. The other elements of the group may exhibit positive oxidation states, in combination with more electronegative elements. The ease with which positive oxidation states of the halogens are formed increases with increasing atomic number.
physical form of the free element, oom temperature, varies with easing atomic number.	At room temperature fluorine and chlorine are gases, bromine is a liquid and iodine a solid. The change in physical form is due to the increase in van der Waal's forces. (See Unit 3, Section III C, p. 21)
elements are usually prepared from corresponding halide ion by ving one of the electrons from the	Because of their high reactivity the halogens occur in nature only in compounds. Since fluorine is the most electronegative element, there is no chemical oxidizing agent that can oxidize the fluoride ion to fluorine. Fluorine is prepared by the electrolysis of its fused compounds. Chlorine, bromine, and iodine can be prepared by various chemical methods.
sition elements are those elements hich the two outermost shells may involved in a chemical reaction.	*It is recommended that the choice of elements to illustrate the behavior of transition elements be limited to elements in period 4. The electron structures of some transition elements are in doubt, and will be predicted differently in various references.* The transition elements are found in the "B" groups and Group VIII of the periodic table.
sition elements generally exhibit tiple oxidation states.	
ions of transition elements are ally colored, both in solid pounds and in solution.	

Topics	Understandings and Fundamental Concepts	Sup
E. Group 0	Group 0 elements are monatomic gases. The atoms of these elements have complete outer shells, which results in an electron configuration that is stable.	Group 0 is referred to rare gases, inert gases
		The term "inert" is no group, since it is poss krypton, xenon, and rad However, the term is st electron configuration the "inert gas structur
IV. Chemistry of a Period	As a period is observed from left to right, a study of the elements leads to certain generalizations.	*It is suggested that th* *period 3.*
	In each period, as the atomic number increases, •the radius of the atom generally decreases. •the ionization energy of the element generally increases. •the electronegativity of the element generally increases. •the elements generally change from very active metals to less active metals to metalloids to less active nonmetals to very active nonmetals, and finally to an inert gas. •there is a transition from positive to negative oxidation states. •the metallic characteristics of the "A" group elements decrease.	Elements near the cente positive and negative o

Understandings and Fundamental Concepts	Supplementary Information
Group 0 elements are monatomic gases. The atoms of these elements have complete outer shells, which results in an electron configuration that is stable.	Group 0 is referred to by a variety of terms including rare gases, inert gases, and noble gases.
	The term "inert" is no longer strictly applicable to this group, since it is possible to form compounds of krypton, xenon, and radon with fluorine and oxygen. However, the term is still in general use, and the electron configuration is quite generally referred to as the "inert gas structure."
As a period is observed from left to right, a study of the elements leads to certain generalizations.	*It is suggested that the teacher limit examples to period 3.*
In each period, as the atomic number increases, • the radius of the atom generally decreases. • the ionization energy of the element generally increases. • the electronegativity of the element generally increases. • the elements generally change from very active metals to less active metals to metalloids to less active nonmetals to very active nonmetals, and finally to an inert gas.	
• there is a transition from positive to negative oxidation states.	Elements near the center of the period may exhibit both positive and negative oxidation states.
• the metallic characteristics of the "A" group elements decrease.	

Topics	Understandings and Fundamental Concepts	
†V. Lanthanide Series—Actinide Series	Elements with atomic numbers 57-71 comprise the lanthanide series.	In these series th from each other in
	Elements with atomic numbers 89-103 comprise the actinide series.	The actinide serie (above 92) which a importance of this properties of the
	These series represent other types of transition elements which differ from the regular transition elements by having the three outermost shells incomplete.	

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† Optional

derstandings and Fundamental Concepts	Supplementary Information
ements with atomic numbers 57-71 mprise the lanthanide series.	In these series the elements do not differ markedly from each other in chemical properties.
ements with atomic numbers 89-103 mprise the actinide series.	The actinide series includes the transuranium elements (above 92) which are "synthetic" elements. The importance of this series lies mainly in the nuclear properties of the elements.
ese series represent other types of ansition elements which differ from e regular transition elements by ving the three outermost shells complete.	

Unit 5 - Mathematics of Chemis

Topics	Understandings and Fundamental Concepts	
	This material need not be taught as a unit, but may be i in various places in the syllabus at the discretion of t	
I. Mole Interpretation	A mole is Avogadro's number of particles.	See also Unit
II. Use of the Mole Concept	The mole may be used in calculations involving the number of particles (atoms, molecules, ions, electrons, or other particles) involved in chemical reactions, the mass of elements or compounds, or the volume relationships in gases.	
A. Gram atomic mass (gram-atom)	Gram atomic mass (gram-atom) of an element represents the mass in grams of Avogadro's number of atoms of the element.	See also Unit The gram atomi mass as shown
B. Gram molecular mass	The gram molecular mass (mole mass) is the sum of the gram atomic masses of the atoms that make up a particular molecule.	
	The gram formula mass (mole mass) of a substance is the sum of the gram atomic masses of the atoms that make up a particular empirical formula.	The gram formu formula is use solids, since
C. Mole volume of a gas	A mole (Avogadro's number) of molecules of any gas occupies a volume of 22.4 liters at S.T.P. It has a mass equal to the molecular mass expressed in grams.	See also Unit

Unit 5 - Mathematics of Chemistry

nderstandings and Fundamental Concepts	Supplementary Information
his material need not be taught as a unit, but may be incorporated n various places in the syllabus at the discretion of the teacher.	
mole is Avogadro's number of articles.	See also Unit 1, Section III A 9, p. 6.
he mole may be used in calculations nvolving the number of particles atoms, molecules, ions, electrons, r other particles) involved in hemical reactions, the mass of elements r compounds, or the volume relation-hips in gases.	
ram atomic mass (gram-atom) of an lement represents the mass in grams f Avogadro's number of atoms of the lement.	See also Unit 2, Section I C, 2d, p. 10. The gram atomic mass is numerically equal to the atomic mass as shown in the periodic table.
he gram molecular mass (mole mass) is he sum of the gram atomic masses of he atoms that make up a particular olecule.	
he gram formula mass (mole mass) of a ubstance is the sum of the gram atomic asses of the atoms that make up a articular empirical formula.	The gram formula mass calculated from the empirical formula is used for ionic substances and network solids, since they are not molecular substances.
mole (Avogadro's number) of olecules of any gas occupies a olume of 22.4 liters at S.T.P. It as a mass equal to the molecular ass expressed in grams.	See also Unit I, Section III A, 9, p. 6.

Topics	Understandings and Fundamental Concepts	Supp
III. Stoichiometry	Stoichiometry is the study of the quantitative relationships implied by chemical formulas and by chemical equations.	In stoichiometry it is f the mole interpretation solving of problems.
A. Problems involving formulas		
1. Percentage composition	The percentage by mass of an element in a compound can be calculated from the ratio of the mass of that element represented in the formula to the mass of the mole.	Sample problems: a. Calculate the percent water (H_2O). Solution: From formula, H_2O Mass of hydrogen, 2 Mass of oxygen, 1 Mass of water % hydrogen = mass mass $= \frac{2.0g.}{18\ g.}$ = 11% % oxygen = mass of mass of $= \frac{16\ g.}{18\ g.}$ x = 89% b. Calculate the percen mass in gypsum (CaSO

Understandings and Fundamental Concepts

Stoichiometry is the study of the quantitative relationships implied by chemical formulas and by chemical equations.

The percentage by mass of an element in a compound can be calculated from the ratio of the mass of that element represented in the formula to the mass of the mole.

Supplementary Information

In stoichiometry it is frequently convenient to use the mole interpretation and mole relationships in the solving of problems.

Sample problems:

a. Calculate the percentage composition by mass of water (H_2O).

Solution:
From formula, H_2O

$$\begin{array}{lll} \text{Mass of hydrogen,} & 2 \times 1.0 \text{ g.} & = 2.0\text{g.} \\ \text{Mass of oxygen,} & 1 \times 16 \text{ g.} & = \underline{16 \text{ g.}} \\ \text{Mass of water} & & = 18 \text{ g.} \end{array}$$

$$\% \text{ hydrogen} = \frac{\text{mass of hydrogen}}{\text{mass of water}} \times 100 = \frac{2.0\text{g.}}{18 \text{ g.}} \times 100 = 11\%$$

$$\% \text{ oxygen} = \frac{\text{mass of oxygen}}{\text{mass of water}} \times 100 = \frac{16 \text{ g.}}{18 \text{ g.}} \times 100 = 89\%$$

b. Calculate the percentage of water of hydration by mass in gypsum ($CaSO_4 \quad 2H_2O$).

Topics	Understandings and Fundamental Concepts	Su
		Solution: From formula, $CaSO_4$ Mass of calcium, Mass of sulfur, Mass of oxygen, Mass of water, Total mass of gypsu $\% \text{ water} = \frac{\text{mass of}}{\text{total mas}}$ $= \frac{36 \text{ g.}}{172 \text{ g.}}$ $= 21\%$
2. Formula from percentage composition	The empirical formula of a compound can be determined from the percentage composition of the compound and the atomic masses of the elements. If the compound is molecular, the molecular formula can also be determined if the molecular mass is known.	Sample problem: A compound was foun carbon and 20.% hyd empirical formula? The gram molecular What is the molecu Solution: Rel. No. atoms C = Rel. No. atoms H = Ratio of C:H = 6.7

Solution:

From formula, $CaSO_4 \cdot 2H_2O$

Mass of calcium, 1 x 40.g. = 40.g.
Mass of sulfur, 1 x 32 g. = 32 g.
Mass of oxygen, 4 x 16 g. = 64 g.
Mass of water, 2 x 18 g. = 36 g.
Total mass of gypsum, =172 g.

$$\% \text{ water} = \frac{\text{mass of water}}{\text{total mass of gypsum}} \times 100$$

$$= \frac{36 \text{ g.}}{172 \text{ g.}} \times 100$$

$$= 21\%$$

The empirical formula of a compound can be determined from the percentage composition of the compound and the atomic masses of the elements. If the compound is molecular, the molecular formula can also be determined if the molecular mass is known.

Sample problem:

A compound was found by analysis to consist of 80.% carbon and 20.% hydrogen by mass. What is the empirical formula?

The gram molecular mass was found to be 30. grams. What is the molecular formula?

Solution:

$$\text{Rel. No. atoms C} = \frac{80.\text{g.}}{12\text{g./g-atom}} = 6.7\text{g-atoms}$$

$$\text{Rel. No. atoms H} = \frac{20.\text{g.}}{1.0\text{g./g-atom}} = 20.\text{g-atoms}$$

Ratio of C:H = 6.7:20

Topics	Understandings and Fundamental Concepts	*Supplementar
		To simplify the ratio, divide by $C = \frac{6.7\text{ g-atoms}}{6.7\text{ g-atoms}} = 1$ $H = \frac{20.\text{ g-atoms}}{6.7\text{ g-atoms}} = 3$ (approx.) Empirical formula is CH_3. Formula mass is 15g./formula uni Molecular mass is 30.g./mole Molecule contains $\frac{30.\text{g./mole}}{15\text{g./formula u}}$ = 2 formula uni Molecular formula is $(CH_3)_2$ or C
3. Molecular mass from gas density	The molecular mass of a gas can be determined from the gas density (mass per unit volume, usually expressed in grams/liter) by using the relationship between mole volume and mass in gases.	Sample problem: The density of a gas is 1.35 Calculate the gram molecular Solution: Gram molecular mass = density = 1.35 g./ℓ. = 30.2 g./mo *It should be emphasized that, i* *involving units, it is important* *carried through all calculations* *with the problem.*

ecular mass of a gas can be ned from the gas density (mass t volume, usually expressed in iter) by using the relationship mole volume and mass in gases.

To simplify the ratio, divide by the smaller number.

$$C = \frac{6.7 \text{ g-atoms}}{6.7 \text{ g-atoms}} = 1$$

$$H = \frac{20. \text{ g-atoms}}{6.7 \text{ g-atoms}} = 3 \text{ (approx.)}$$

Empirical formula is CH_3.

Formula mass is 15g./formula unit

Molecular mass is 30.g./mole

$$\text{Molecule contains } \frac{30.\text{g./mole}}{15\text{g./formula unit}}$$

$$= 2 \text{ formula units/mole}$$

Molecular formula is $(CH_3)_2$ or C_2H_6

Sample problem:

The density of a gas is 1.35 grams/liter at S.T.P. Calculate the gram molecular mass of the gas.

Solution:

Gram molecular mass = density x mole volume.

$$= 1.35 \text{ g./}\ell. \times 22.4\ \ell\text{/mole}$$

$$= 30.2 \text{ g./mole}$$

It should be emphasized that, in solving problems involving units, it is important that all units be carried through all calculations, and solved algebraically with the problem.

Topics	Understandings and Fundamental Concepts	
†4. Atomic mass from electrolysis (Faraday)		*Introduction of thi at this point excep preferably those wi*
B. Problems involving equations	The coefficients used in balancing a chemical equation represent moles. In stoichiometric problems involving equations it is assumed that the reaction is a single reaction (with no side reactions), that the reaction goes to completion, and that the reactants are completely reacted.	*In solving many pro convenient to use m methods may be used*
1. Mass - mass problems	A balanced equation shows the mole proportions of products to reactants. It is possible to determine the mass of one substance that reacts with, or is produced from, a given mass of another.	Sample problem: How many grams of s 355 grams of chlori Solution: Equation: 2NaCl — Moles: 2 Step I: $\frac{\text{g. Cl}}{\text{mol. m}}$ $\frac{355\text{g.}}{71.0\text{g.}}$ Step II: (moles Cl_2) (5.00 moles

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† Optional

Understandings and Fundamental Concepts

The coefficients used in balancing a chemical equation represent moles. In stoichiometric problems involving equations it is assumed that the reaction is a single reaction (with no side reactions), that the reaction goes to completion, and that the reactants are completely reacted.

A balanced equation shows the mole proportions of products to reactants.

It is possible to determine the mass of one substance that reacts with, or is produced from, a given mass of another.

Supplementary Information

Introduction of this type of problem is not recommended at this point except for select groups of students, preferably those with a background in physics.

In solving many problems involving equations it is convenient to use mole relationships. However, other methods may be used.

Sample problem:

How many grams of sodium chloride are needed to produce 355 grams of chlorine gas?

Solution:

Equation: $2NaCl \longrightarrow 2Na + Cl_2$

Moles: 2 2 1

Step I: $$\frac{\text{g. } Cl_2}{\text{mol. mass } Cl_2} = \text{moles of } Cl_2$$

$$\frac{355\text{g.}}{71.0\text{g./mole}} = 5.00 \text{ moles } Cl_2$$

Step II:

$$(\text{moles } Cl_2) \cdot \left(\frac{\text{moles NaCl}}{\text{moles } Cl_2}\right) = \text{moles NaCl}$$

$$(5.00 \text{ moles } Cl_2) \cdot \left(\frac{2 \text{ moles NaCl}}{1 \text{ mole } Cl_2}\right) = 10.0 \text{ moles NaCl}$$

Topics | Understandings and Fundamental Concepts | Supplement

Step III:

(moles NaCl) . (mol. ma

(10.0 moles) . (58.5 g.

2. Mass - volume problems

Since, in the balanced equation, the mole unit serves to relate the quantities of products and reactants, it is possible to determine quantitative results in desired units, which may not necessarily be the same as the original units. At S.T.P., one mole of gas occupies 22.4 liters.

Sample problem:

How many grams of zinc metal a
hydrochloric acid to produce
gas at S.T.P.?

Solution:

Equation: $Zn + 2HCl \longrightarrow Z$

Moles: 1 2

Step I:

$$\text{moles } H_2 = \frac{\text{vol. of } H_2 \text{ in li}}{\text{mole volum}}$$

$$= \frac{11.2\ell}{\frac{22.4\ell}{\text{mole}}}$$

$$= 0.500 \text{ mole } H_2$$

Step II:

$$(\text{moles } H_2) \cdot \left(\frac{\text{moles Zn}}{\text{moles } H_2}\right.$$

$$(0.500 \text{ mole } H_2) \cdot \left(\frac{1 \text{ mol}}{1 \text{ mol}}\right.$$

Step III:

(moles Zn) . (mole mass
(0.500 mole). (65.4g/mol

Understandings and Fundamental Concepts

Since, in the balanced equation, the mole unit serves to relate the quantities of products and reactants, it is possible to determine quantitative results in desired units, which may not necessarily be the same as the original units. At S.T.P., one mole of gas occupies 22.4 liters.

Supplementary Information

Step III:

(moles NaCl) . (mol. mass NaCl) = grams NaCl

(10.0 moles) . (58.5 g./mole) = 585 g. NaCl needed.

Sample problem:

How many grams of zinc metal are needed to react with hydrochloric acid to produce 11.2 liters of hydrogen gas at S.T.P.?

Solution:

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

Moles: 1 2 1 1

Step I:

$$\text{moles } H_2 = \frac{\text{vol. of } H_2 \text{ in liters (S.T.P.)}}{\text{mole volume (S.T.P.)}}$$

$$= \frac{11.2\ell}{\frac{22.4\ell}{\text{mole}}}$$

$$= 0.500 \text{ mole } H_2$$

Step II:

$$(\text{moles } H_2) \;.\; \left(\frac{\text{moles Zn}}{\text{moles } H_2}\right) = \text{moles Zn}$$

$$(0.500 \text{ mole } H_2) \;.\; \left(\frac{1 \text{ mole Zn}}{1 \text{ mole } H_2}\right) = 0.500 \text{ mole Zn}$$

Step III:

(moles Zn) . (mole mass Zn) = grams Zn
(0.500 mole) . (65.4g/mole) = 32.7g. Zn

Topics	Understandings and Fundamental Concepts	Suppl
3. Volume - volume problems	Since one mole of any gas occupies the same volume as one mole of any other gas, at the same temperature and pressure, the volumes of gases involved in a reaction are proportional to the number of moles indicated by the numerical coefficients in a balanced equation.	Sample problem: In the reaction: N_2 + 3H calculate the volume of 100. liters of NH_3. Solution: Equation: $N_2 + 3H_2 \longrightarrow 2NH$ Moles: 1 3 2 When reacting gases are pressure, the volume rat (Avogadro's Hypothesis). Since the molar ratio bet volumes of gases are in t Therefore: volume of H_2 needed $= \left(\frac{\text{moles H}}{\text{moles N}}\right.$ $= \left(\frac{\text{3 moles}}{\text{2 moles}}\right.$ = 150. lit
IV. Solutions	A solution is a homogeneous mixture of two or more substances, the composition of which may vary within limits.	*Factors affecting solubi* *7-8-9, Block J, and shou*

Understandings and Fundamental Concepts

Since one mole of any gas occupies the same volume as one mole of any other gas, at the same temperature and pressure, the volumes of gases involved in a reaction are proportional to the number of moles indicated by the numerical coefficients in a balanced equation.

A solution is a homogeneous mixture of two or more substances, the composition of which may vary within limits.

Supplementary Information

Sample problem:

In the reaction: $N_2 + 3H_2 \longrightarrow 2NH_3$,

calculate the volume of hydrogen required to form 100. liters of NH_3.

Solution:

Equation:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

Moles:

1 3 2

When reacting gases are at the same temperature and pressure, the volume ratio is equal to the mole ratio (Avogadro's Hypothesis).

Since the molar ratio between H_2 and NH_3 is 3:2, the volumes of gases are in the same ratio.

Therefore:

$$\begin{aligned}\text{volume of } H_2 \text{ needed} &= \left(\frac{\text{moles } H_2}{\text{moles } NH_3}\right)\left(\text{volume of } NH_3 \text{ produced}\right)\\ &= \left(\frac{3 \text{ moles } H_2}{2 \text{ moles } NH_3}\right)(100.\ \text{liters } NH_3)\\ &= 150.\ \text{liters } H_2\end{aligned}$$

Factors affecting solubility are discussed in *__Science 7-8-9, Block J__*, *and should be reviewed as necessary.*

Topics	Understandings and Fundamental Concepts	Supplementary
	The component of a solution which is present in larger amount is called the solvent, while the other component is called the solute.	Most solutions dealt with in beg chemistry are aqueous solutions.
A. Methods of indicating concentrations	Concentration of solutions may be indicated in a variety of ways.	*Students should be familiar wit "concentrated"; "saturated", "un "supersaturated"; and with the u the solubility charts in the* *<u>Ref Chemistry.</u>*
1. Molarity	The molarity (M) of a solution is the number of moles of solute contained in a liter of solution.	A two molar (2M) solution contai per liter of solution and a 0.1 contains 0.1 mole of solute per concentration in moles per lite volume in liters equals the numb in the solution. The mass in grams of solute can b multiplying the number of moles mass. Sample problems: a. How many moles of NaOH are c 0.1M solution of NaOH? How contained in this solution? Solution: Moles of solute = molarity x = 0.1 mole/ℓ = 0.02 mole Grams of solute = no. of mol = 0.02 mole = 0.8g. NaOH

ngs and Fundamental Concepts	Supplementary Information
nt of a solution which is larger amount is called , while the other component he solute.	Most solutions dealt with in beginning courses in chemistry are aqueous solutions.
on of solutions may be n a variety of ways.	*Students should be familiar with the terms "dilute" and "concentrated"; "saturated", "unsaturated", and "supersaturated"; and with the use and interpretation of the solubility charts in the* ***Reference Tables for Chemistry.***
y (M) of a solution is the oles of solute contained of solution.	A two molar (2M) solution contains 2 moles of solute per liter of solution and a 0.1 molar solution (0.1M) contains 0.1 mole of solute per liter of solution. The concentration in moles per liter multiplied by the volume in liters equals the number of moles of solute in the solution. The mass in grams of solute can be determined by multiplying the number of moles of solute by the mole mass.

Sample problems:

a. How many moles of NaOH are contained in 200 ml. of a 0.1M solution of NaOH? How many grams of NaOH are contained in this solution?

Solution:

Moles of solute = molarity x volume (in liters)

$= 0.1 \text{ mole}/\ell \times 0.2\ell$

$= 0.02$ mole NaOH

Grams of solute = no. of moles x mole mass

$= 0.02 \text{ mole} \times 40\text{g./mole}$

$= 0.8$g. NaOH

Topics	Understandings and Fundamental Concepts	
		b. What is the m of the soluti Solution: No. moles of : Molarity = = =
2. Molality	The molality (m) of a solution is the number of moles of a solute dissolved in 1000 grams of solvent.	Molality is used between moles of is related to th *Problems involvi not be subject t*
B. Effect of solute on solvent	The presence of dissolved particles affects some properties of the solvent.	
	Properties which depend on the relative number of particles rather than on the nature of the particles are called colligative properties.	Colligative prop changes in boili and osmotic pres *Changes in vapor be subject to ex*
	The presence of a solute raises the boiling point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of a no raises the boili

Understandings and Fundamental Concepts	Supplementary Information
	b. What is the molarity of a solution of NaCl if 500 ml. of the solution contains 11.6 grams of NaCl? Solution: $\text{No. moles of solute} = \frac{\text{mass of solute}}{\text{mole mass}}$ $= \frac{11.6\text{g.}}{58\text{g./mole}} = 0.2 \text{ mole}$ $\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume(in liters)}}$ $= \frac{0.2 \text{ mole}}{0.5 \text{ liter}}$ $= 0.4 \text{ mole/liter or } 0.4M$
The molality (m) of a solution is the number of moles of a solute dissolved in 1000 grams of solvent.	Molality is used where one is interested in the relation between moles of solute and moles of solvent. (This is related to the study of colligative properties.) *Problems involving preparation of molal solutions will not be subject to examination.*
The presence of dissolved particles affects some properties of the solvent.	
Properties which depend on the relative number of particles rather than on the nature of the particles are called colligative properties.	Colligative properties as related to solutions include changes in boiling point, freezing point, vapor pressure, and osmotic pressure. *Changes in vapor pressure and osmotic pressure will not be subject to examination.*
The presence of a solute raises the boiling point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of a nonelectrolyte per 1000 grams of water raises the boiling point of the water 0.52°C.

Topics	Understandings and Fundamental Concepts	Supplementar
	The presence of a solute lowers the freezing point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of a nonelectrolyte per lowers the freezing point of the *Students are not required to mem (0.52°C. and 1.86°C.) which are solvents and which may vary at h solute.* *Teachers may wish to point out t boiling point elevation or freez of a solvent can be used to dete of the solute.*
1. Abnormal behavior of electrolytes	Electrolytes in solution cause greater changes in properties of the solvent than do nonelectrolytes.	This behavior of electrolytes in credence to the existence of ion *Some teachers may wish to introd electrolytes under Unit 7 where treated in greater detail.*

Understandings and Fundamental Concepts	Supplementary Information
The presence of a solute lowers the freezing point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of a nonelectrolyte per 1000 grams of water lowers the freezing point of the water 1.86°C. *Students are not required to memorize the constants (0.52°C. and 1.86°C.) which are different for different solvents and which may vary at high concentrations of solute.* *Teachers may wish to point out that measurement of the boiling point elevation or freezing point depression of a solvent can be used to determine the molecular mass of the solute.*
Electrolytes in solution cause greater changes in properties of the solvent than do nonelectrolytes.	This behavior of electrolytes in solution gives credence to the existence of ionic particles. *Some teachers may wish to introduce this property of electrolytes under Unit 7 where electrolytes are treated in greater detail.*

Topics	Understandings and Fundamental Concepts	
I. Kinetics	Chemical kinetics is the branch of chemistry concerned with the rate of chemical reactions and the mechanisms by which chemical reactions occur.	The rate of a ch of the number of of product forme
	The mechanism of a chemical reaction is the sequence of stepwise reactions by which the over all change occurs.	For most reactio The net reaction changes that occ
A. Role of energy in reaction rates	Energy is required to initiate a chemical reaction. Energy may be released or absorbed in a chemical reaction.	By graphing the activation energ products against energies involve
1. Activation energy	Activation energy is the minimum energy required to initiate a reaction.	
2. Heat of reaction	Heat of reaction (ΔH) is the heat energy released or absorbed in the formation of the products. It represents the difference in heat content between the products and the reactants. $\Delta H = H_{products} - H_{reactants}$	
	In an exothermic reaction, the products have a lower potential energy than the reactants, and the sign of ΔH is negative.	The sign that ma chemical equatio sign for ΔH. Fo of hydrogen and $H_2(g) + \frac{1}{2}O_2(g$
	In an endothermic reaction, the products have a higher potential energy then the reactants, and the sign of ΔH is positive.	Since this is a negative and its In equations tha species should b

Unit 6 - Kinetics and Equilibrium

Understandings and Fundamental Concepts	Supplementary Information
Chemical kinetics is the branch of chemistry concerned with the rate of chemical reactions and the mechanisms by which chemical reactions occur.	The rate of a chemical reaction is measured in terms of the number of moles of reactant used up (or moles of product formed) per unit volume in a unit of time.
The mechanism of a chemical reaction is the sequence of stepwise reactions by which the over all change occurs.	For most reactions only the net reaction is observable. The net reaction represents a summation of all the changes that occur.
Energy is required to initiate a chemical reaction.	By graphing the potential energy of the reactants, the activation energy, and the potential energy of the products against a time sequence, one may describe the energies involved in a chemical reaction.
Energy may be released or absorbed in a chemical reaction.	
Activation energy is the minimum energy required to initiate a reaction.	
Heat of reaction (ΔH) is the heat energy released or absorbed in the formation of the products. It represents the difference in heat content between the products and the reactants. $\Delta H = H_{products} - H_{reactants}$	
In an exothermic reaction, the products have a lower potential energy than the reactants, and the sign of ΔH is negative.	The sign that may be used when energy is included in a chemical equation should not be confused with the sign for ΔH. For example, the equation for the reaction of hydrogen and oxygen to form water may be written: $H_2(g) + \frac{1}{2}O_2(g) = H_2O\ (\ell) + 68.3$ kcal.
In an endothermic reaction, the products have a higher potential energy then the reactants, and the sign of ΔH is positive.	Since this is an exothermic reaction, the sign of ΔH is negative and its value is -68.3 kcal. In equations that include heat, the phase of each species should be specified.

Topics

Understandings and Fundamental Concepts

Supplement

3. Potential energy diagram

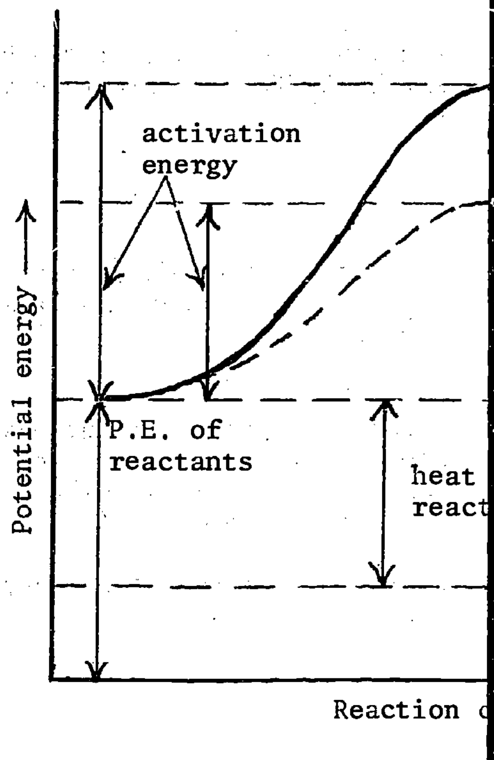
The relationship between activation energy and heat of reaction for a given reaction can be shown graphically in a potential energy diagram by plotting potential energy against a reaction coordinate representing the progress of the reaction.

The difference in potential energy between the final products and the initial reactants represents the heat of reaction.

If the potential energy of the products is lower than the potential energy of the reactants, energy has been liberated (exothermic reaction). If the potential energy of the products is higher than the potential energy of the reactants, heat has been absorbed (endothermic reaction).

The highest point in the curve represents the potential energy of the intermediate products (activated complexes). The difference between this point and the initial potential energy of the reactants represents the activation energy of the reaction.

Students should be able to id
reaction and activation energ
between exothermic and endoth
potential energy diagram.



The graph shown above, when r
represents an exothermic reac
at a lower potential energy t
the graph is reversed, or rea
would represent an endothermi

B. Factors affecting rate of reaction

Generally chemical reactions depend on collisions between the reacting particles, atoms, molecules, or ions.

In considering the factors af
reaction it is useful to use
particles.

Understandings and Fundamental Concepts

The relationship between activation energy and heat of reaction for a given reaction can be shown graphically in a potential energy diagram by plotting potential energy against a reaction coordinate representing the progress of the reaction.

The difference in potential energy between the final products and the initial reactants represents the heat of reaction.

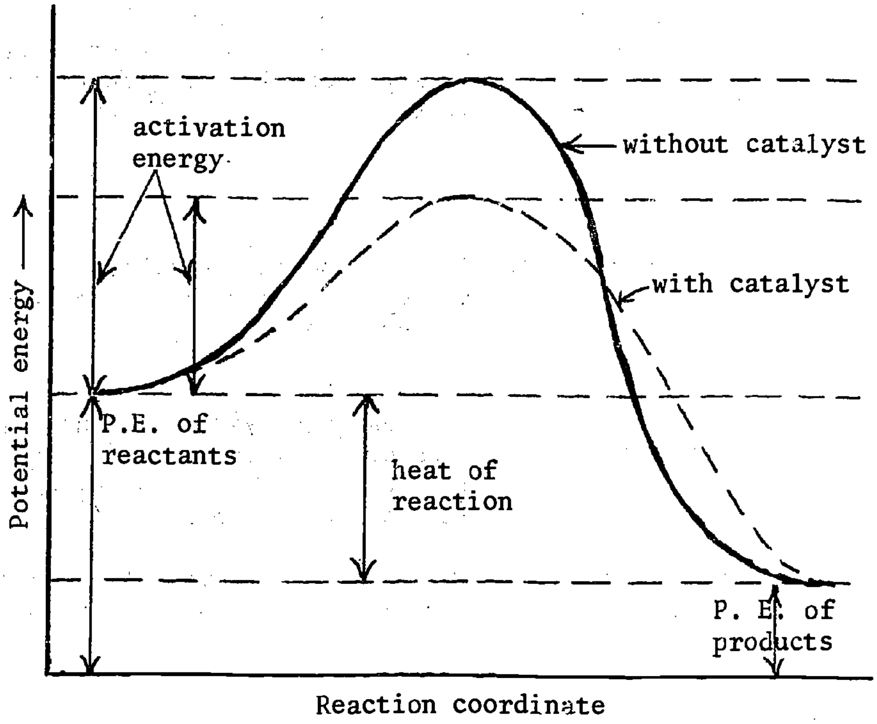
If the potential energy of the products is lower than the potential energy of the reactants, energy has been liberated (exothermic reaction). If the potential energy of the products is higher than the potential energy of the reactants, heat has been absorbed (endothermic reaction).

The highest point in the curve represents the potential energy of the intermediate products (activated complexes). The difference between this point and the initial potential energy of the reactants represents the activation energy of the reaction.

Generally chemical reactions depend on collisions between the reacting particles, atoms, molecules, or ions.

Supplementary Information

Students should be able to identify the heat of reaction and activation energy, and to distinguish between exothermic and endothermic reactions from a potential energy diagram.



The graph shown above, when read from left to right, represents an exothermic reaction, with the products at a lower potential energy than the reactants. If the graph is reversed, or read from right to left, it would represent an endothermic reaction.

In considering the factors affecting the rate of a reaction it is useful to use a model of colliding particles.

Topics	Understandings and Fundamental Concepts	
	The rate of reaction is affected by the number of collisions occurring and the fraction of these collisions that are effective.	*Examples of t of principles in Unit 10, S*
1. Nature of the reactants	Since bonds may be broken or formed in a reaction, the nature of the bond is an important factor affecting reaction rates.	Reactions tha usually rapic of ionic subs Reactions tha slow at room between hydro
2. Concentration	An increase in the concentration of one or more reactants increases the rate of reaction.	An increase i the frequency Concentration
	In a gaseous system, an increase in pressure will result in an increase in concentration, and thus an increase in the rate of reaction.	The effect of be confused w equilibrium.
3. Temperature	An increase in temperature increases the rate of all chemical reactions.	An increase i the kinetic e only the numb greater impo *A graph show distribution be found in t*
4. Reaction mechanism	The reaction mechanism affects the rate of the reaction.	The mechanism net equation reaction, $N_2 + 3H_2$ — it is unlike hydrogen mol tion probabl two-particle
	Most chemical reactions take place as a result of a series of simpler steps.	

Understandings and Fundamental Concepts	Supplementary Information
The rate of reaction is affected by the number of collisions occurring and the fraction of these collisions that are effective.	*Examples of the application to commercial processes of principles affecting rates of reaction may be found in Unit 10, Section II A, pp. 80-82.*
Since bonds may be broken or formed in a reaction, the nature of the bond is an important factor affecting reaction rates.	Reactions that involve negligible bond rearrangement are usually rapid at room temperature, for example, reactions of ionic substances in aqueous solutions. Reactions that involve the breaking of bonds tend to be slow at room temperature, for example, the reaction between hydrogen and oxygen.
An increase in the concentration of one or more reactants increases the rate of reaction.	An increase in the concentration of a reactant increases the frequency of collisions. Concentrations are usually measured in moles per liter.
In a gaseous system, an increase in pressure will result in an increase in concentration, and thus an increase in the rate of reaction.	The effect of pressure on the rate of reaction should not be confused with the effect of pressure on chemical equilibrium. (See Section II, C, 1b of this unit, p.50)
An increase in temperature increases the rate of all chemical reactions.	An increase in temperature increases the speed (and thus the kinetic energy) of the particles, and increases not only the number of collisions per unit time, but of greater importance, the effectiveness of the collisions. *A graph showing the effect of temperature on the distribution of kinetic energy in the gaseous phase may be found in Unit I, Section III A, p. 5.*
The reaction mechanism affects the rate of the reaction. Most chemical reactions take place as a result of a series of simpler steps.	The mechanism of a reaction cannot be deduced from the net equation for the reaction. For example, in the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$, it is unlikely that one nitrogen molecule and three hydrogen molecules collide simultaneously. Such a reaction probably takes place as the result of a series of two-particle collisions, with the formation of extremely

Topics	Understandings and Fundamental Concepts	Supplementa
		transitory intermediate product complexes. There is no indicat such reactions in the net equat
5. Catalysts	Catalysts change the activation energy required and thus change the rate of reaction. A catalyst does not initiate a chemical reaction.	A catalyst changes the mechanis involving less activation energ over all process.
II. Equilibrium	Most reactions are reversible. Equilibrium is a state of balance between two opposite reactions (physical or chemical) occurring at the same rate.	Equilibrium is dynamic and only appearance of the system. It d activity of individual particle implies motion, and dynamic equ condition in which the interact the reactants in one direction action of the particles of the direction. Although the reacti opposing reactions are equal, a may exist in which the quantiti products are not equal. Thus e when only a small quantity of t formed or when only a small qua remains. For a system in equilibrium, a (such as temperature, concentra result in a change in the equil Because the reactions in an equ it follows that equilibrium may the forward or the reverse reac

erstandings and Fundamental Concepts	Supplementary Information
	transitory intermediate products called activated complexes. There is no indication of the sequence of such reactions in the net equation.
alysts change the activation rgy required and thus change the e of reaction.	A catalyst changes the mechanism of a reaction to one involving less activation energy, but does not change the over all process.
atalyst does not initiate a mical reaction.	
t reactions are reversible.	Equilibrium is dynamic and only describes the over all appearance of the system. It does not describe the activity of individual particles. The word dynamic implies motion, and dynamic equilibrium is that condition in which the interaction of the particles of the reactants in one direction is balanced by the interaction of the particles of the products in the opposite direction. Although the reaction rates for the opposing reactions are equal, a state of equilibrium may exist in which the quantities of reactants and products are not equal. Thus equilibrium may be reached when only a small quantity of the products has been formed or when only a small quantity of reactants remains. For a system in equilibrium, a change in conditions (such as temperature, concentration, or pressure) may result in a change in the equilibrium point. Because the reactions in an equilibrium are reversible, it follows that equilibrium may be attained either from the forward or the reverse reaction.
ilibrium is a state of balance ween two opposite reactions ysical or chemical) occurring at same rate.	

Topics	Understandings and Fundamental Concepts	Supple
A. Phase equilibrium	In general, phase changes (solid to liquid, or liquid to gas) are reversible, and, in a closed system, equilibrium may be attained.	In general, if a solid o closed container, eventua particles in the vapor ph is equal to the rate of e equilibrium results in wh vapor pressure characteri liquid.
B. Solution equilibrium	The solubility of a solute is defined as the mass of that solute dissolved in a given volume of solvent at equilibrium under specified conditions.	
1. Gases in liquids	In a closed system equilibrium may exist between a gas dissolved in a liquid and the undissolved gas above the liquid.	
	The equilibrium between dissolved and undissolved gas is affected by temperature and pressure.	Increased temperature de in liquids. Increased pressure increa liquids.
2. Solids in liquids	A solution equilibrium exists when the opposing processes of dissolving and crystallizing of a solute occur at equal rates.	
	A solution exhibiting equilibrium between the dissolved and undissolved solute is known as a saturated solution.	Solubility may be define solute in a saturated so
C. Chemical equilibrium	Chemical equilibrium is attained when the concentration of the reactants and products remains constant.	When observable changes temperature) no longer o system, the system has r At this point the forwar reaction are occurring a
		Refer to Unit 10, Sectio

andings and Fundamental Concepts	Supplementary Information
eral, phase changes (solid to , or liquid to gas) are ible, and, in a closed system, brium may be attained.	In general, if a solid or a liquid is confined in a closed container, eventually there will be enough particles in the vapor phase so that the rate of return is equal to the rate of escape. Thus, a dynamic equilibrium results in which there is an equilibrium vapor pressure characteristic of the solid or the liquid.
lubility of a solute is defined mass of that solute dissolved iven volume of solvent at brium under specified conditions.	
losed system equilibrium may between a gas dissolved in a and the undissolved gas above quid.	
uilibrium between dissolved and olved gas is affected by ature and pressure.	Increased temperature decreases the solubility of gases in liquids. Increased pressure increases the solubility of gases in liquids.
tion equilibrium exists when posing processes of dissolving ystallizing of a solute occur al rates.	
tion exhibiting equilibrium n the dissolved and undissolved is known as a saturated solution.	Solubility may be defined as the concentration of solute in a saturated solution.
al equilibrium is attained when ncentration of the reactants and ts remains constant.	When observable changes (such as color, pressure, and temperature) no longer occur in a reacting chemical system, the system has reached a state of equilibrium. At this point the forward reaction and the reverse reaction are occurring at equal rates.

Refer to Unit 10, Section II, A, pp. 80-82.

Topics	Understandings and Fundamental Concepts	Suppl
1. Le Chatelier's principle		*Minimum requirements will treatment of the applicat to systems in equilibrium* *Application of Le Chateli processes may be found in pp. 80-82.*
	If a stress, such as a change in concentration, pressure, or temperature, is applied to a system at equilibrium, the equilibrium is shifted in a way that tends to relieve the effects of the stress.	When a chemical system at chemical reaction occurs at a different point (i. reactants and products).
a. Effect of concentration	Increasing the concentration of one substance in a reaction will cause the reaction to go in such a direction as to consume the increase. Eventually equilibrium will be reestablished at a new equilibrium point.	For example, in the Haber increasing the concentrat hydrogen will increase th If the system remains clo of ammonia that results w decomposition of ammonia, will be established.
	Removal of one of the products of a reaction results in a decrease in its concentration, and will cause the reaction to go in such a direction as to increase the concentration of the products.	Removal of a product tend to go more nearly to comp the product may destroy t removing all of that reac reaction. Products may be removed f part, by the formation of insoluble product (precip by the formation of an es such as water.

Understandings and Fundamental Concepts	Supplementary Information
	Minimum requirements will be limited to a qualitative treatment of the application of Le Chatelier's principle to systems in equilibrium.
	Application of Le Chatelier's principle to commercial processes may be found in Unit 10, Section II, A, pp. 80-82.
If a stress, such as a change in concentration, pressure, or temperature, is applied to a system at equilibrium, the equilibrium is shifted in a way that tends to relieve the effects of the stress.	When a chemical system at equilibrium is disturbed, chemical reaction occurs and equilibrium is reestablished at a different point (i.e., with new concentrations of reactants and products).
Increasing the concentration of one substance in a reaction will cause the reaction to go in such a direction as to consume the increase. Eventually equilibrium will be reestablished at a new equilibrium point.	For example, in the Haber process, $N_2 + 3H_2 \rightleftarrows 2NH_3$, increasing the concentration of either nitrogen or hydrogen will increase the rate of ammonia formation. If the system remains closed, the increased concentration of ammonia that results will increase the rate of decomposition of ammonia, and a new equilibrium point will be established.
Removal of one of the products of a reaction results in a decrease in its concentration, and will cause the reaction to go in such a direction as to increase the concentration of the products.	Removal of a product tends to cause the forward reaction to go more nearly to completion. Continuous removal of the product may destroy the equilibrium system by removing all of that reactant necessary for the reverse reaction.
	Products may be removed from a reaction, wholly or in part, by the formation of a gas, the formation of an insoluble product (precipitate) or, in an ionic reaction, by the formation of an essentially un-ionized product such as water.

Topics	Understandings and Fundamental Concepts	
b. Effect of pressure	A change in pressure can only affect chemical equilibria in which gases are involved. An increase in pressure will displace the point of equilibrium in the direction that favors the formation of smaller volume. If no change in volume is involved in the reaction, a change of pressure has no effect on the equilibrium.	For example, four moles of Since molar v temperature a in a decrease will result i However, in t change in pre
c. Effect of temperature	When the temperature of a system in equilibrium is raised, the equilibrium is displaced in such a way that heat is absorbed.	All chemical or the absorp equilibrium, are taking pl reaction is f exothermic re should be poi both endother increase in t unit, p. 46.) increased une equilibrium. $N_2 + 3H_2 \rightleftarrows$ favors the de
d. Effect of catalyst	In a system in equilibrium, a catalyst increases the rate of both the forward and reverse reactions equally, and produces no net change in the equilibrium concentrations.	A catalyst ma quickly, but

Understandings and Fundamental Concepts	Supplementary Information
A change in pressure can only affect chemical equilibria in which gases are involved. An increase in pressure will displace the point of equilibrium in the direction that favors the formation of smaller volume. If no change in volume is involved in the reaction, a change of pressure has no effect on the equilibrium.	For example, in the Haber process, $N_2 + 3H_2 \rightleftarrows 2NH_3$. four moles of reactants form two moles of products. Since molar volumes of gases at the same pressure and temperature are equal, the forward reaction results in a decrease in volume. Thus an increase in pressure will result in an increased production of ammonia. However, in the reaction $H_2 + Cl_2 \rightleftarrows 2HCl$, a change in pressure does not affect the equilibrium.
When the temperature of a system in equilibrium is raised, the equilibrium is displaced in such a way that heat is absorbed.	All chemical changes involve either the evolution or the absorption of energy. In every system in equilibrium, an endothermic and an exothermic reaction are taking place simultaneously. The endothermic reaction is favored by an increase in temperature, the exothermic reaction by a decrease in temperature. It should be pointed out that the rates of all reactions, both endothermic and exothermic, are increased by an increase in temperature. (See Section I, B, 3 of this unit, p. 46.) However, the opposing reactions are increased unequally, resulting in a displacement of the equilibrium. For example, in the Haber process, $N_2 + 3H_2 \rightleftarrows 2NH_3 + 22kcal.$, raising the temperature favors the decomposition of ammonia.
In a system in equilibrium, a catalyst increases the rate of both the forward and reverse reactions equally, and produces no net change in the equilibrium concentrations.	A catalyst may cause equilibrium to be reached more quickly, but does not affect the equilibrium reached.

Topics	Understandings and Fundamental Concepts	Supplement
2. Law of chemical equilibrium	When a reversible reaction has attained equilibrium at a given temperature, the product of the molar concentrations of the substances to the right of the equation, divided by the product of the molar concentrations of the substances to the left (each concentration raised to the power equal to the number of moles of that substance appearing in the equation) is a constant. For the reaction: $aM + bN = cP + dQ$ $\frac{[P]^c \times [Q]^d}{[M]^a \times [N]^b} = K =$ a constant at constant temperature This constant is called the equilibrium constant.	*Minimum requirements will be l qualitatively the effects of c in simple reactions, and to th significance of the values of* *The derivation of the law of c not be subject to examination.* In the mathematical expression brackets are used to indicate in moles per liter." By conve equilibrium constant, the conc on the right of the chemical e numerator and the concentratio form the denominator. In equa reversible reaction at equilib reactions proceeding at equal may be used instead of double
	The magnitude of K is used by chemists to predict the extent of chemical reactions. A large value of K indicates the favoring of products, that is, the equilibrium mixture consists largely of products. A small value of K indicates that reactants are favored.	The equilibrium constant has a given chemical reaction at a p This value remains constant ev of the substances involved may For example, in the reaction g concentration of P would cause the left, thus decreasing the increasing the concentrations of K would remain constant.
	The equilibrium constant does change with a change of temperature.	In a reversible reaction the r ward and reverse reactions are by change in temperature.

erstandings and Fundamental Concepts

n a reversible reaction has ained equilibrium at a given perature, the product of the molar centrations of the substances to right of the equation, divided by product of the molar concentrations the substances to the left (each centration raised to the power equal the number of moles of that stance appearing in the equation) a constant.

the reaction:

$$+ bN = cP + dQ$$

$$\frac{{}^{c}\ \ x\ \ [Q]^{d}}{{}^{a}\ \ x\ \ [N]^{b}} = K = \text{a constant at constant temperature}$$

s constant is called the equilibrium stant.

magnitude of K is used by chemists predict the extent of chemical ctions.

arge value of K indicates the oring of products, that is, the ilibrium mixture consists largely products.

mall value of K indicates that ctants are favored.

equilibrium constant does change n a change of temperature.

Supplementary Information

Minimum requirements will be limited to predicting qualitatively the effects of concentration changes in simple reactions, and to the interpretation of the significance of the values of K.

The derivation of the law of chemical equilibrium will not be subject to examination.

In the mathematical expression of this law, square brackets are used to indicate "concentrations measured in moles per liter." By convention, in calculating the equilibrium constant, the concentrations of products on the right of the chemical equation form the numerator and the concentrations of products on the left form the denominator. In equations representing a reversible reaction at equilibrium, implying two reactions proceeding at equal rates, the equal sign may be used instead of double arrows.

The equilibrium constant has a numerical value for any given chemical reaction at a particular temperature. This value remains constant even though the concentrations of the substances involved may increase or decrease. For example, in the reaction given, an increase in the concentration of P would cause the reaction to go to the left, thus decreasing the concentration of Q and increasing the concentrations of M and N. The value of K would remain constant.

In a reversible reaction the reaction rates of the forward and reverse reactions are not affected equally by change in temperature.

Topics	Understandings and Fundamental Concepts	
a. Ionization constant (K_A)	Weak electrolytes in aqueous solution attain an equilibrium between ions and the undissociated compound. The equilibrium constant for such systems is called the ionization constant, K_A, (sometimes called the dissociation constant). For the reaction: $HB\ (aq) = H^+(aq) + B^-(aq)$, when temperature is constant, $K_A = \frac{[H^+]\ .\ [B^-]}{[HB]}$	Ionization a of K_A are de page 61.
b. Solubility product constant (K_{sp})	In a saturated solution of an ionic solid, an equilibrium is established between the ions of the saturated solution and the excess solid phase. For the reaction, $AB\ (s) = A^+(aq) + B^-(aq)$ when the temperature is constant $K_{sp} = [A^+]\ .\ [B^-]$	*Minimum requ* *to the quali* *of relative* The concentr essentially equilibrium the concentr included in equation. T equilibrium, equilibrium $K = \frac{[A^+]\ .}{[AB]}$ Since $[AB]$ i $[A^+]\ .\ [B$ The magnitud

Weak electrolytes in aqueous solution attain an equilibrium between ions and the undissociated compound. The equilibrium constant for such systems is called the ionization constant, K_A, (sometimes called the dissociation constant).

For the reaction:

$HB\ (aq) = H^+(aq) + B^-(aq)$,

when temperature is constant,

$$K_A = \frac{[H^+]\ .\ [B^-]}{[HB]}$$

In a saturated solution of an ionic solid, an equilibrium is established between the ions of the saturated solution and the excess solid phase. For the reaction,

$AB\ (s) = A^+(aq)\ +\ B^-(aq)$

when the temperature is constant

$$K_{sp} = [A^+]\ .\ [B^-]$$

Ionization and the significance of differ[illegible] values of K_A are developed in more detail in Unit 7, Section IV, page 61.

Minimum requirements for examination will be limited to the qualitative interpretation of the significance of relative K_{sp} values.

The concentration (mass/unit volume) of a solid is essentially constant. In the expression of the equilibrium constant for any reaction involving a solid, the concentration of the solid can therefore be included in the constant and does not appear in the equation. Thus, in the case of the solubility equilibrium, application of the law of chemical equilibrium would give

$$K = \frac{[A^+]\ .\ [B^-]}{[AB]}$$

Since $[AB]$ itself is constant, we may say

$$[A^+]\ .\ [B^-] = K\ [AB] = K_{sp}$$

The magnitude of K_{sp} is used in comparing the solubilities

Topics	Understandings and Fundamental Concepts	Supplementary
		of slightly soluble salts. For example, at room temperature K_{sp} $CaSO_4$ = 2.4×10^{-5} K_{sp} $BaSO_4$ = 1.6×10^{-9} Thus $BaSO_4$ is less soluble than C precipitated at a lower concentra Section II A, 4, 5, pp. 81-82.)
III. Spontaneous Reactions	Spontaneous reactions depend on the balance between two fundamental tendencies in nature, (1) toward a lower energy state, and (2) toward randomness.	*Minimum requirements for examinat a concept of the meaning of entro*
A. Energy changes	At constant temperature and pressure, a system tends to undergo a reaction so that, in its final state, it has lower energy than in its initial state.	A system tends to change from a one of low energy. For chem is represented mathematically as A, 2 of this unit, p. 44.) This favors the exothermic reaction, negative.
B. Entropy changes	Entropy is a measure of the disorder, randomness, or lack of organization of a system.	The solid phase, in regular crys more organized than the liquid ph is more organized than the gaseo
	Entropy is so defined that the more random a system is, the higher the entropy.	An increase in entropy during a a system means that in its final more disordered (random) than in
	High entropy (randomness) is favored by high temperatures.	High temperatures increase the r particles, and thus tend to incr

derstandings and Fundamental Concepts	Supplementary Information
	of slightly soluble salts. For example, at room temperature $K_{sp}\ CaSO_4 = 2.4 \times 10^{-5}$ $K_{sp}\ BaSO_4 = 1.6 \times 10^{-9}$ Thus $BaSO_4$ is less soluble than $CaSO_4$, and would be precipitated at a lower concentration. (See Unit 10, Section II A, 4, 5, pp. 81-82.)
ntaneous reactions depend on the lance between two fundamental ndencies in nature, (1) toward a wer energy state, and (2) toward ndomness.	*Minimum requirements for examination will be limited to a concept of the meaning of entropy and free energy.*
constant temperature and pressure, system tends to undergo a reaction that, in its final state, it has wer energy than in its initial ate.	A system tends to change from a state of high energy to one of low energy. For chemical systems this energy is represented mathematically as ΔH. (See Section I, A, 2 of this unit, p. 44.) This tendency in nature favors the exothermic reaction, in which ΔH is negative.
tropy is a measure of the disorder, ndomness, or lack of organization a system.	The solid phase, in regular crystalline arrangement, is more organized than the liquid phase; the liquid phase is more organized than the gaseous phase.
tropy is so defined that the more ndom a system is, the higher the tropy.	An increase in entropy during a change in the state of a system means that in its final state the system is more disordered (random) than in its initial state.
gh entropy (randomness) is favored high temperatures.	High temperatures increase the rate of motion of the particles, and thus tend to increase randomness.

Topics	Understandings and Fundamental Concepts	Supplement
	At constant temperature, a system tends to undergo a reaction so that in its final state it has higher entropy (greater randomness) than in its initial state.	A system tends to change from a state of less order. For chemical systems this chan by $T\Delta S$, where T represents the ΔS the change in entropy.
C. Free energy change	The difference between energy change and entropy change is the free energy change (ΔG). $\Delta G = \Delta H - T\Delta S$	*Minimum requirements for exami to the interpretation of the s values in the equation.*
D. Predicting spontaneous reactions	For a spontaneous change to occur in a system, the free energy change (ΔG) must be negative. (In a system at equilibrium, the free energy change is zero.)	
	When the two factors, tendency toward lower energy content and tendency toward higher entropy in a system, cannot be satisfied simultaneously, the spontaneous change that may take place will be determined by the factor that is dominant at the temperature of the system.	From energy changes alone, exo always be expected to occur sp thermic reactions would never spontaneously. Exceptions to may occur when a change in ent reaction, or favors an endothe example: (1) the change in pha exothermic reaction, and, from energy only, water might be ex taneously at any temperature. toward higher entropy favors t water. At temperatures below energy change is dominant and taneously. At temperatures ab the entropy change becomes the ice melts. (2) the reaction, is an endothermic reaction, ar would oppose a spontaneous rea action results in an increase

Understandings and Fundamental Concepts	Supplementary Information
At constant temperature, a system tends to undergo a reaction so that in its final state it has higher entropy (greater randomness) than in its initial state.	A system tends to change from a state of great order to a state of less order. For chemical systems this change in order is described by $T\Delta S$, where T represents the Kelvin temperature, and ΔS the change in entropy.
The difference between energy change and entropy change is the free energy change (ΔG). $\Delta G = \Delta H - T\Delta S$	*Minimum requirements for examination will be limited to the interpretation of the significance of the values in the equation.*
For a spontaneous change to occur in a system, the free energy change (ΔG) must be negative. (In a system at equilibrium, the free energy change is zero.)	
When the two factors, tendency toward lower energy content and tendency toward higher entropy in a system, cannot be satisfied simultaneously, the spontaneous change that may take place will be determined by the factor that is dominant at the temperature of the system.	From energy changes alone, exothermic reactions would always be expected to occur spontaneously, and endothermic reactions would never be expected to occur spontaneously. Exceptions to both of these predictions may occur when a change in entropy opposes an exothermic reaction, or favors an endothermic reaction. For example: (1) the change in phase from water to ice is an exothermic reaction, and, from the consideration of energy only, water might be expected to freeze spontaneously at any temperature. However, the tendency toward higher entropy favors the reaction from ice to water. At temperatures below the freezing point the energy change is dominant and water will freeze spontaneously. At temperatures above the freezing point the entropy change becomes the dominant factor, and ice melts. (2) the reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$ is an endothermic reaction, and the energy change would oppose a spontaneous reaction. However, the reaction results in an increase in entropy, due to the

formation of a gas and solid simple than the original solid. At high t effect of the entropy change become overcome the effect of the energy c the reaction takes place. If the t not high enough the reaction will n formation of a gas and solid simpler in organization than the original solid. At high temperatures the effect of the entropy change becomes sufficient to overcome the effect of the energy change, and the reaction takes place. If the temperature is not high enough the reaction will not take place.

Topics	Understandings and Fundamental Concepts	
I. Electrolytes	An electrolyte is a substance that dissolves in water to form a solution that will conduct an electric current.	Electrolytes i acids.
	The ability of a solution to conduct an electric current is due to the presence of ions that are free to move.	Arrhenius prov behavior of el
	The abnormal effect of an electrolyte on the boiling point and freezing point of the solvent is explained by the presence of ions in the solution.	Since these ef properties of than the natur electrolyte wi particles (mol solution.
II. Acids and Bases	Acids and bases may be defined in terms of operational definitions.	An operational observations, apply when a t
	Conceptual definitions of acids and bases have been extended as understanding of acid-base reactions has grown, and as principles of these reactions have been applied to reactions not in aqueous solutions.	A conceptual d of observed fa
A. Acids	Acids may be defined in terms of their characteristic properties.	These properti form the basis acid.
	• Aqueous solutions of acids conduct electricity.	Acids conduct their ionizati in aqueous so acids) while c are weak elect *Tables for Che*

Unit 7 - Acid-Base Theories

Understandings and Fundamental Concepts	Supplementary Information
An electrolyte is a substance that dissolves in water to form a solution that will conduct an electric current.	Electrolytes include all ionic compounds and all acids.
The ability of a solution to conduct an electric current is due to the presence of ions that are free to move.	Arrhenius provided the first explanation of the behavior of electrolytes in aqueous solution.
The abnormal effect of an electrolyte on the boiling point and freezing point of the solvent is explained by the presence of ions in the solution.	Since these effects are due to the colligative properties of the solution (number of particles rather than the nature of the particles), the effect of the electrolyte will depend on the number of individual particles (molecules and ions) present in the solution.
Acids and bases may be defined in terms of operational definitions.	An operational definition is one based on experimental observations, and includes a set of conditions to apply when a term is used in a particular situation.
Conceptual definitions of acids and bases have been extended as understanding of acid-base reactions has grown, and as principles of these reactions have been applied to reactions not in aqueous solutions.	A conceptual definition is one based on interpretation of observed facts.
Acids may be defined in terms of their characteristic properties.	These properties can be observed experimentally, and form the basis of the operational definition of an acid.
• Aqueous solutions of acids conduct electricity.	Acids conduct electricity in relation to the degree of their ionization. A few acids ionize almost completely in aqueous solution and are strong electrolytes (strong acids) while others ionize only to a slight degree and are weak electrolytes (weak acids). (See the *Reference Tables for Chemistry.*)

Topics	Understandings and Fundamental Concepts	Supplementary
	• Acids will react with certain metals to liberate hydrogen gas.	Acids will react with those metal in a Table of Standard Oxidation Some acids, in addition to their have strong oxidizing ability and very dilute solution, do not re on reaction with metals. For exa concentrated sulfuric acid have s properties.
	• Acids cause color changes in acid-base indicators.	Acid-base indicators are substanc different colors in acid and basi example, litmus is red in acid so basic solution. Different indicators change color concentrations of hydrogen ion. The mechanism of indicator action optional Section IV, C of this un
	• Acids react with hydroxides to form water and a salt.	When hydrogen ions react with hyd formed. This reaction is called
	• Dilute aqueous solutions of acids have a sour taste.	*Teachers should not permit studen chemicals, but should point out t many common foods is due to the p*
1. Arrhenius theory	An acid is a substance that yields hydrogen ions as the only positive ions in aqueous solution. The characteristic properties of acids in aqueous solution are due to an excess of hydrogen ions.	This conceptual definition is ade reactions in aqueous solutions. mechanism of chemical reactions h inclusive definitions have been a

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Understandings and Fundamental Concepts	Supplementary Information
• Acids will react with certain metals to liberate hydrogen gas.	Acids will react with those metals above hydrogen in a Table of Standard Oxidation Potentials. Some acids, in addition to their acid properties, have strong oxidizing ability and thus, except in very dilute solution, do not release hydrogen gas on reaction with metals. For example, nitric acid and concentrated sulfuric acid have strong oxidizing properties.
• Acids cause color changes in acid-base indicators.	Acid-base indicators are substances which have different colors in acid and basic solutions. For example, litmus is red in acid solution and blue in basic solution. Different indicators change color at different concentrations of hydrogen ion. The mechanism of indicator action is developed in the optional Section IV, C of this unit, p. 62.
• Acids react with hydroxides to form water and a salt.	When hydrogen ions react with hydroxide ions, water is formed. This reaction is called neutralization.
• Dilute aqueous solutions of acids have a sour taste.	*Teachers should not permit students to taste laboratory chemicals, but should point out that the sour taste of many common foods is due to the presence of acids.*
An acid is a substance that yields hydrogen ions as the only positive ions in aqueous solution.	This conceptual definition is adequate when considering reactions in aqueous solutions. As knowledge of the mechanism of chemical reactions has increased, more inclusive definitions have been advanced.
The characteristic properties of acids in aqueous solution are due to an excess of hydrogen ions.	

Topics	Understandings and Fundamental Concepts	
2. Brönsted-Lowry theory	An acid is any species (molecule or ion) that can donate a proton to another species.	The Brö Arrheni Lowry d that ar In addi as acid not aci the rea NH_3 + H donates an acid
B. Bases	Bases may be defined in terms of their characteristic properties.	These p definit
	• Aqueous solutions of bases conduct electricity.	Some ex shown i
	• Bases cause color changes in acid-base indicators.	Two com solutio which i solutio
	• Bases react with acids to form water and a salt.	
	• Aqueous solutions of bases feel slippery.	Strong
1. Arrhenius theory	A base is a substance that yields hydroxide ions as the only negative ions in aqueous solution.	
	The characteristic properties of bases in aqueous solution are due to the hydroxide ion.	Accord are hy

Understandings and Fundamental Concepts	Supplementary Information
An acid is any species (molecule or ion) that can donate a proton to another species.	The Brönsted-Lowry theory does not replace the Arrhenius theory, but rather extends it. The Brönsted-Lowry definition of an acid includes all substances that are acids according to the Arrhenius definition. In addition, some molecules and ions are classified as acids under the Brönsted-Lowry definition that are not acids in the Arrhenius sense. For example, in the reaction, $NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$, the water molecule donates a proton to the ammonia, and is here considered an acid in the Brönsted-Lowry sense.
Bases may be defined in terms of their characteristic properties.	These properties form the basis of the operational definition of a base.
• Aqueous solutions of bases conduct electricity.	Some examples of the relative degree of ionization are shown in the *Reference Tables for Chemistry.*
• Bases cause color changes in acid-base indicators.	Two common indicators are litmus which is blue in basic solution and red in acid solution, and phenolphthalein which is pink in basic solution and colorless in acid solution.
• Bases react with acids to form water and a salt.	
• Aqueous solutions of bases feel slippery.	Strong bases have a caustic action on the skin.
A base is a substance that yields hydroxide ions as the only negative ions in aqueous solution.	
The characteristic properties of bases in aqueous solution are due to the hydroxide ion.	According to the Arrhenius definition, the only bases are hydroxides.

Topics	Understandings and Fundamental Concepts	Supplementary
2. Brönsted-Lowry theory	A base is any species (molecule or ion) that can combine with a proton. All bases have at least one pair of unshared electrons.	The Brönsted-Lowry definition ex definition to include many speci the OH^- that can accept a proton the reaction, $H_2O + HCl \rightleftharpoons H_3O^+ +$ C combines with a proton to form t is here considered a base in the
C. Amphoteric substances	An amphoteric (amphiprotic) substance is one that can act either as an acid or as a base, depending on its chemical environment.	The hydroxides of some metals; (e lead, chromium) are amphoteric. a strong base they behave as aci of a strong acid they behave as Lowry sense, species (e.g., H_2O, donate or accept a proton are al
III. Acid-base Reactions		
A. Neutralization	Acid-base neutralization pertains to the reaction which occurs when equivalent quantities of an acid and a hydroxide are mixed.	One mole of hydrogen ions will r hydroxide ions to form water. $H^+ + OH^- \longrightarrow H_2O$
1. Acid-base titration	The molarity of an acid(or base) of unknown molarity can be determined by slowly combining it with a base (or acid) of known molarity (standard solution) until neutralization occurs. This process is called titration. The end point of the titration may be determined by the use of appropriate indicators or by temperature changes or electrode potential changes.	The molarity of the unknown solu calculated from an understanding relationships involved. *Minimum requirements for examina to simple calculations involving neutralization reactions.*

rstandings and Fundamental Concepts	Supplementary Information
se is any species (molecule or that can combine with a proton. bases have at least one pair of ared electrons.	The Brönsted-Lowry definition extends the Arrhenius definition to include many species in addition to the OH^- that can accept a proton. For example, in the reaction, $H_2O + HCl \rightleftharpoons H_3O^+ + Cl^-$, the water molecule combines with a proton to form the hydronium ion, and is here considered a base in the Brönsted-Lowry sense.
nphoteric (amphiprotic) substance ne that can act either as an or as a base, depending on its ical environment.	The hydroxides of some metals; (e.g., aluminum, zinc, lead, chromium) are amphoteric. In the presence of a strong base they behave as acids and in the presence of a strong acid they behave as bases. In the Brönsted-Lowry sense, species (e.g., H_2O, HSO_4^-) which can either donate or accept a proton are also amphiprotic.
-base neutralization pertains to reaction which occurs when valent quantities of an acid and droxide are mixed.	One mole of hydrogen ions will react with one mole of hydroxide ions to form water. $H^+ + OH^- \longrightarrow H_2O$
molarity of an acid(or base) of own molarity can be determined by ly combining it with a base (or) of known molarity (standard tion) until neutralization occurs. process is called titration. end point of the titration may etermined by the use of opriate indicators or by erature changes or electrode ntial changes.	The molarity of the unknown solution can be calculated from an understanding of the molar relationships involved. *Minimum requirements for examination will be limited to simple calculations involving acid-base neutralization reactions.*

Topics	Understandings and Fundamental Concepts	Supplemen
2. Salts	A salt is an ionic compound containing positive ions other than hydrogen and negative ions other than hydroxide.	All salts are strong electro to be completely dissociated
	Some salts in aqueous solution react with the water to form solutions that are acidic or basic. This process is called hydrolysis.	Salts can be considered as h the neutralization of an aci formed from strong acids and aqueous solutions that are a from weak acids and strong b solutions that are basic. S acids and strong bases do no
B. Conjugate pair	According to the Brönsted-Lowry theory, acid-base reactions involve a transfer of protons from the acid to the base.	In order to accept a proton, least one pair of unshared e without any electrons will s belonging to the base, formi bond. (See Unit 3, Section I
	Acid-base reactions are reversible.	
	In an acid-base reaction an acid transfers a proton to become a base, making an acid-base pair. A base gains a proton to become an acid, making a second acid-base pair. Each pair, made up of an acid and a base related by the transfer of a proton, is called a conjugate acid-base pair.	In the following reactions t (acid_1-base_1 and acid_2-base_2 subscripts. $\text{Base}_1 + \text{Acid}_2 \rightleftarrows$ Ac $H_2O + HCl \rightleftarrows H_3$ $NH_3 + H_2O \rightleftarrows$ NH
	The strongest acids have the weakest conjugate bases and the strongest bases have the weakest conjugate acids.	Refer to the *Reference Table*

Understandings and Fundamental Concepts	Supplementary Information
A salt is an ionic compound containing positive ions other than hydrogen and negative ions other than hydroxide.	All salts are strong electrolytes and are considered to be completely dissociated in aqueous solution.
Some salts in aqueous solution react with the water to form solutions that are acidic or basic. This process is called hydrolysis.	Salts can be considered as having been derived from the neutralization of an acid and a base. Salts formed from strong acids and weak bases will give aqueous solutions that are acidic. Salts formed from weak acids and strong bases will give aqueous solutions that are basic. Salts formed from strong acids and strong bases do not hydrolyze in solution.
According to the Brönsted-Lowry theory, acid-base reactions involve a transfer of protons from the acid to the base.	In order to accept a proton, a base must have at least one pair of unshared electrons. The proton without any electrons will share a pair of electrons belonging to the base, forming a coordinate covalent bond. (See Unit 3, Section II, B, p. 19.)
Acid-base reactions are reversible.	
In an acid-base reaction an acid transfers a proton to become a base, making an acid-base pair. A base gains a proton to become an acid, making a second acid-base pair. Each pair, made up of an acid and a base related by the transfer of a proton, is called a conjugate acid-base pair.	In the following reactions the two conjugate pairs (acid$_1$-base$_1$ and acid$_2$-base$_2$) are identified by subscripts. $\text{Base}_1 + \text{Acid}_2 \rightleftarrows \text{Acid}_1 + \text{Base}_2$ $H_2O + HCl \rightleftarrows H_3O^+ + Cl^-$ $NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$
The strongest acids have the weakest conjugate bases and the strongest bases have the weakest conjugate acids.	Refer to the *Reference Tables for Chemistry*.

Topics	Understandings and Fundamental Concepts	S
†C. Buffers	Buffer solutions are solutions which maintain an approximately constant pH on addition of hydrogen ions or hydroxide ions. Buffer solutions contain a weak acid and a salt of that acid (or a weak base and a salt of that base) which act as a conjugate acid-base pair.	A solution which con acetate is an exampl acetic acid acts as ion as the conjugate
IV. Ionization Constant	The equilibrium constant for the ionization of acids (K_A) is a convenient method for comparing the relative strength of acids. For the reaction, $HB = H^+ + B^-$ $K_A = \frac{[H^+] \cdot [B^-]}{[HB]}$	*Minimum requirements to the interpretatio values of K_A.* Ionization constants that are not complet are completely disso The denominator [HB] infinity. Ionization constants *Reference Tables for* the relative strengt acid with a K_A = 1.8 stronger than an aci
A. K_W	In water and aqueous solutions the product of the hydrogen ion concentration and the hydroxide ion concentration is a constant at constant temperature. This constant, K_W, is useful in problems involving hydrogen ion and hydroxide ion concentrations. $K_W = [H^+] \cdot [OH^-]$	*Students should be c the relationship:* $K_W = [H^+] \cdot [OH$ Since the concentrat stant for all reacti value can be include $[H^+] \quad [OH^-] = K_A$ The numerical value In pure water, $[H^+]$

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† Optional

Understandings and Fundamental Concepts

Buffer solutions are solutions which maintain an approximately constant pH on addition of hydrogen ions or hydroxide ions. Buffer solutions contain a weak acid and a salt of that acid (or a weak base and a salt of that base) which act as a conjugate acid-base pair.

The equilibrium constant for the ionization of acids (K_A) is a convenient method for comparing the relative strength of acids.
For the reaction,

$HB = H^+ + B^-$

$$K_A = \frac{[H^+] \cdot [B^-]}{[HB]}$$

In water and aqueous solutions the product of the hydrogen ion concentration and the hydroxide ion concentration is a constant at constant temperature. This constant, K_W, is useful in problems involving hydrogen ion and hydroxide ion concentrations.

$$K_W = [H^+] \cdot [OH^-]$$

Supplementary Information

A solution which contains acetic acid and sodium acetate is an example of a buffer solution. Here the acetic acid acts as the conjugate acid and the acetate ion as the conjugate base.

Minimum requirements for examination will be limited to the interpretation of the significance of the values of K_A.

Ionization constants can be calculated for all acids that are not completely dissociated. For acids that are completely dissociated there is no equilibrium. The denominator [HB] approaches zero and K_A approaches infinity.

Ionization constants such as those given in the *Reference Tables for Chemistry* can be used in comparing the relative strength of acids. For example, an acid with a $K_A = 1.8 \times 10^{-5}$ (although a weak acid) is stronger than an acid with $K_A = 5.8 \times 10^{-10}$.

Students should be able to solve problems based on the relationship:

$$K_W = [H^+] \cdot [OH^-] = 1.0 \times 10^{-14}$$

Since the concentration of water is essentially constant for all reactions in aqueous solution, its value can be included in the ionization constant giving:

$$[H^+] \quad [OH^-] = K_A \cdot [H_2O] = K_W$$

The numerical value of K_W at 25°C. is 1.0×10^{-14}.

In pure water, $[H^+] = [OH^-] = 1.0 \times 10^{-7}$.

Topics	Understandings and Fundamental Concepts	
B. pH	The pH of a solution indicates the concentration of hydrogen ions (acid strength) in a solution. A pH of 7 is neutral. A pH of less than 7 is acidic. A pH greater than 7 is basic.	*Minimum require* *to the calculat* *from pH values* *molar hydrogen* *involving whole* pH is the logar logarithm) of t $pH = log$ $[H^+] = 10$ Pure water has
†C. Indicators	An indicator is a weak acid with different colors for the undissociated form (HIn) and the negative ion (In^-). Color 1 Color 2 $HIn = H^+ + In^-$	$K_A = \frac{[H^+] \cdot}{[HIn}$ The $[H^+]$ at whi its end point. $[In^-] = [HIn]$

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† Optional

Understandings and Fundamental Concepts	Supplementary Information
The pH of a solution indicates the concentration of hydrogen ions (acid strength) in a solution. A pH of 7 is neutral. A pH of less than 7 is acidic. A pH greater than 7 is basic.	*Minimum requirements for examination will be limited to the calculation of molar hydrogen ion concentrations from pH values and calculation of pH values from molar hydrogen ion concentration. Only calculations involving whole number pH values will be used.* pH is the logarithm of the reciprocal (negative logarithm) of the hydrogen ion concentration. $pH = \log \frac{1}{[H^+]}$ $[H^+] = 10^{-pH}$ Pure water has a pH of 7.
An indicator is a weak acid with different colors for the undissociated form (HIn) and the negative ion (In^-). Color 1 Color 2 $HIn = H^+ + In^-$	$K_A = \frac{[H^+] \cdot [In^-]}{[HIn]}$ The $[H^+]$ at which an indicator turns color is called its end point. At the end point, $[In^-] = [HIn]$, or $K_A = [H^+]$

Unit 8 - Redox and Electrochemistry

Topics	Understandings and Fundamental Concepts	Supp
I. Redox	Redox reactions result from the competition for electrons between atoms.	Redox is a term used fo *Examples of the applica of principles involving in Unit 10, Section II,*
A. Oxidation number	The oxidation number (oxidation state) of an atom is the charge which an atom has, or appears to have, when electrons are counted according to certain arbitrary rules.	The oxidation number, a convenient notation for electrons involved in a
	In assigning oxidation numbers electrons shared between two unlike atoms are counted as belonging to the more electronegative atom; electrons shared between two like atoms are divided equally between the sharing atoms.	Application of the gene following operational r number: • In the free elements number of zero. For in Na, and sulfur in numbers of zero. • In simple ions (ions oxidation number is ion. For example, i oxidation number of $FeCl_2$ has an oxidati it has an oxidation • All metals in Group oxidation number is • All metals in Group oxidation number is • Oxygen has an oxidat compounds EXCEPT in it is -1, and in com when it may be +1 or oxygen has an oxidat

Unit 8 - Redox and Electrochemistry

Understandings and Fundamental Concepts	Supplementary Information
Redox reactions result from the competition for electrons between atoms.	Redox is a term used for oxidation-reduction. *Examples of the application to commercial processes of principles involving redox reactions may be found in Unit 10, Section II, B, pp. 82-83.*
The oxidation number (oxidation state) of an atom is the charge which an atom has, or appears to have, when electrons are counted according to certain arbitrary rules.	The oxidation number, although arbitrary, is a convenient notation for keeping track of the number of electrons involved in a chemical reaction.
In assigning oxidation numbers electrons shared between two unlike atoms are counted as belonging to the more electronegative atom; electrons shared between two like atoms are divided equally between the sharing atoms.	Application of the general rules results in the following operational rules for determining oxidation number: • In the free elements, each atom has an oxidation number of zero. For example, hydrogen in H_2, sodium in Na, and sulfur in S_8 all have oxidation numbers of zero. • In simple ions (ions containing one atom) the oxidation number is equal to the charge on the ion. For example, in $CaCl_2$, calcium has an oxidation number of +2 and chlorine -1; iron in $FeCl_2$ has an oxidation number of +2 while in $FeCl_3$ it has an oxidation number of +3. • All metals in Group IA form only +1 ions and their oxidation number is +1 in all compounds. • All metals in Group IIA form only +2 ions and their oxidation number is +2 in all compounds. • Oxygen has an oxidation number of -2 in all its compounds EXCEPT in peroxides (such as H_2O_2) when it is -1, and in compounds with fluorine when it may be +1 or +2. For example, in H_2SO_4 oxygen has an oxidation number of -2.

Topics	Understandings and Fundamental Concepts	
		• Hydrogen compound CaH_2, et hydrogen
	All oxidation numbers must be consistent with the conservation of charge.	For neutral atoms must two hydroge oxygens con must contri polyatomic than one atoms must in $SO_4^=$, the number of - an oxidatio
B. Oxidation	Oxidation refers to any chemical change in which there is an increase in oxidation number. The particle that increases in oxidation number is said to be oxidized. Oxidation represents a loss, or an apparent loss, of electrons. The particle that is oxidized acts as a reducing agent.	
C. Reduction	Reduction refers to any chemical change in which there is a decrease in oxidation number. The particle that decreases in oxidation number is said to be reduced. Reduction represents a gain, or an apparent gain, of electrons. The particle that is reduced acts as an oxidizing agent.	

Understandings and Fundamental Concepts

All oxidation numbers must be consistent with the conservation of charge.

Oxidation refers to any chemical change in which there is an increase in oxidation number. The particle that increases in oxidation number is said to be oxidized.

Oxidation represents a loss, or an apparent loss, of electrons.

The particle that is oxidized acts as a reducing agent.

Reduction refers to any chemical change in which there is a decrease in oxidation number. The particle that decreases in oxidation number is said to be reduced.

Reduction represents a gain, or an apparent gain, of electrons.

The particle that is reduced acts as an oxidizing agent.

Supplementary Information

- Hydrogen has an oxidation number of +1 in all its compounds EXCEPT in the metal hydrides (such as LiH, CaH_2, etc.) when it is -1. For example, in H_2SO_4, hydrogen has an oxidation number of +1.

For neutral molecules the oxidation number of all the atoms must add up to zero. For example, in H_2SO_4 the two hydrogens contribute a total of +2, the four oxygens contribute a total of -8. Therefore, the sulfur must contribute an oxidation number of +6. For polyatomic ions (charged particles that contain more than one atom) the oxidation numbers of all the atoms must add up to the charge on the ion. For example, in $SO_4^=$, the four oxygens contribute a total oxidation number of -8. Therefore, the sulfur must contribute an oxidation number of +6 to give the ion a charge of -2.

Topics	Understandings and Fundamental Concepts	Supplemen
D. Redox reactions	Oxidation and reduction occur simultaneously; one cannot occur without the other.	In oxidation and reduction th decrease of oxidation number electrons. The only way by wh shifted away from an atom (ox be pulled toward another atom
	There is a conservation of charge as well as a conservation of mass in a redox reaction.	
II. Electrochemistry		
A. Half-reactions	A redox reaction may be considered in two parts, one representing a loss of electrons (oxidation), and the other representing a gain of electrons (reduction). Each reaction is known as a half-reaction.	A separate equation showing g (electronic equation) can be half-reaction. For example: $Mg + Cl_2 \longrightarrow MgC$ $Mg \longrightarrow Mg^{++} + 2e^-$ (oxidati $Cl_2 + 2e^- \longrightarrow 2Cl^-$ (reduct
B. Half-cells	It is possible to set up reactions so that each half of a redox reaction takes place in a separate vessel, provided that the vessels are connected by a salt bridge or porous cup which permits the migration of ions but does not allow the solutions to mix.	
1. Half-cell potential	Comparison of the driving force of a half-reaction with that of the hydrogen standard establishes a scale of voltages.	It is impossible to measure t tendency of a half-reaction e an attendant half-reaction. the relative oxidizing tenden has been found convenient to for the oxidation of hydrogen as an arbitrary standard.

rstandings and Fundamental Concepts	Supplementary Information
ation and reduction occur ltaneously; one cannot occur out the other.	In oxidation and reduction the increase and decrease of oxidation number results from a shift of electrons. The only way by which electrons can be shifted away from an atom (oxidation) is for them to be pulled toward another atom (reduction).
e is a conservation of charge ell as a conservation of mass in dox reaction.	
dox reaction may be considered wo parts, one representing a loss lectrons (oxidation), and the r representing a gain of electrons duction). Each reaction is known a half-reaction.	A separate equation showing gain or loss of electrons (electronic equation) can be written for each half-reaction. For example: $Mg + Cl_2 \longrightarrow MgCl_2$ can be represented as $Mg \longrightarrow Mg^{++} + 2e^-$ (oxidation) $Cl_2 + 2e^- \longrightarrow 2Cl^-$ (reduction)
is possible to set up reactions that each half of a redox reaction es place in a separate vessel, vided that the vessels are nected by a salt bridge or porous which permits the migration of s but does not allow the solutions mix.	
parison of the driving force of alf-reaction with that of the rogen standard establishes a le of voltages.	It is impossible to measure the absolute oxidizing tendency of a half-reaction except by comparison with an attendant half-reaction. For purposes of measuring the relative oxidizing tendency of a half-reaction, it has been found convenient to adopt the half-reaction for the oxidation of hydrogen, $H_2 \rightarrow 2H^+ + 2e^-$ as an arbitrary standard.

Topics	Understandings and Fundamental Concepts
	When each half-reaction is compared to the standard under specified conditions of concentration, temperature, and pressure, standard oxidation potentials can be developed.
	The standard oxidation potential (E°) gives the potential difference in volts between the specified half-reaction and the hydrogen half-reaction.
2. Use of standard oxidation potentials	Oxidation potentials are useful in determining whether a specific redox reaction will take place.
	Any pair of half-reactions can be combined to give the complete reaction for a cell whose potential difference can be calculated by adding the appropriate half-cell potentials.

A Table
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Mg $\longrightarrow$

Cl_2 +

Mg +

In comb
must be

2Na

the two

2(Na$\rightarrow$ N

Cl_2 +

2Na +

Understandings and Fundamental Concepts

When each half-reaction is compared to the standard under specified conditions of concentration, temperature, and pressure, standard oxidation potentials can be developed.

The standard oxidation potential (E°) gives the potential difference in volts between the specified half-reaction and the hydrogen half-reaction.

Oxidation potentials are useful in determining whether a specific redox reaction will take place.

Any pair of half-reactions can be combined to give the complete reaction for a cell whose potential difference can be calculated by adding the appropriate half-cell potentials.

Supplementary Information

A Table of Standard Oxidation Potentials is included in the Reference Tables for Chemistry.

In combining half-reactions it must be remembered that, in any redox reaction, there must be an oxidation half-reaction and a reduction half-reaction. In tables of oxidation potentials all half-reactions are written as oxidations. To obtain the equation for the reduction half-reaction, the equation as written in the table must be reversed, and the sign of E° changed. For example, in the reaction,

$$Mg + Cl_2 \longrightarrow MgCl_2$$

the magnesium is oxidized and the chlorine is reduced. The two half-reactions may be combined:

$$Mg \longrightarrow Mg^{++} + 2e^- \qquad E^\circ = +2.37$$

$$Cl_2 + 2e^- \longrightarrow 2Cl^- \qquad E^\circ = +1.36$$

$$Mg + Cl_2 \longrightarrow Mg^{++} + 2Cl^- \quad E^\circ = +3.73$$

In combining half reactions, the electron transfer must be balanced. For example, in the reaction,

$$2Na + Cl_2 \longrightarrow 2NaCl$$

the two half-reactions would be combined as follows:

$$2(Na \rightarrow Na^+ + e^-) \qquad E^\circ = +2.71$$

$$Cl_2 + 2e^- \longrightarrow 2Cl^- \qquad E^\circ = +1.36$$

$$2Na + Cl_2 \rightarrow 2Na^+ + 2Cl^- \quad E^\circ = +4.07$$

Topics

Understandings and Fundamental Concepts

Suppleme

In combining half-reactions, if the potential (E°) for the over all reaction is positive, the reaction is spontaneous.

In combining half-reactions, if the potential (E°) for the over all reaction is negative no reaction will take place.

NOTE: the oxidation poten multiplied by the coeffici for the reaction.

Metals with positive oxida hydrogen from an acid.

For example, if magnesium solution of an acid the ne could be represented by th

$Mg + 2H^+ \longrightarrow Mg^{++} +$

Combining the half-reactio

$Mg \longrightarrow Mg^{++} + 2e^-$

$2H^+ + 2e^- \longrightarrow H_2$

$Mg + 2H^+ \longrightarrow Mg^{++} +$

Since this value is positi take place spontaneously.

Metals with negative oxida replace hydrogen from an a

For example, if copper met an acid the reaction, if c by the net equation:

$Cu + 2H^+ \longrightarrow Cu^{++} +$

Combining the half-reactic

$Cu \longrightarrow Cu^{++} + 2e^-$

$2H^+ + 2e^- \longrightarrow H_2$

$Cu + 2H^+ \longrightarrow Cu^{++} +$

tandings and Fundamental Concepts

mbining half-reactions, if the
tial (E°) for the over all
ion is positive, the reaction
ontaneous.

combining half-reactions, if the
ential (E°) for the over all
ction is negative no reaction will
e place.

Supplementary Information

NOTE: the oxidation potentials (E°) are Not *multiplied by the coefficients in calculating the E° for the reaction.*

Metals with positive oxidation potentials will replace hydrogen from an acid.

For example, if magnesium metal is added to a solution of an acid the net reaction, if one occurs, could be represented by the equation:

$$Mg + 2H^{+} \longrightarrow Mg^{++} + H_2$$

Combining the half-reactions:

$$\begin{array}{ll} Mg \longrightarrow Mg^{++} + 2e^{-} & E^{\circ} = +2.37 \\ 2H^{+} + 2e^{-} \longrightarrow H_2 & E^{\circ} = 0.00 \\ \hline Mg + 2H^{+} \longrightarrow Mg^{++} + H_2 & E^{\circ} = +2.37 \end{array}$$

Since this value is positive we know the reaction will take place spontaneously.

Metals with negative oxidation potentials will not replace hydrogen from an acid.

For example, if copper metal is added to a solution of an acid the reaction, if one occurs, could be represented by the net equation:

$$Cu + 2H^{+} \longrightarrow Cu^{++} + H_2$$

Combining the half-reactions:

$$\begin{array}{ll} Cu \longrightarrow Cu^{++} + 2e^{-} & E^{\circ} = -0.34 \\ 2H^{+} + 2e^{-} \longrightarrow H_2 & E^{\circ} = 0.00 \\ \hline Cu + 2H^{+} \longrightarrow Cu^{++} + H_2 & E^{\circ} = -0.34 \end{array}$$

Topics	Understandings and Fundamental Concepts	
		Since this valu will not take p
C. Chemical cells	Redox reactions that occur spontaneously may be employed to provide a source of electrical energy.	*Details of comm* *required.*
	When the two half-cells of a redox reaction are connected by an external conductor, a flow of electrons (electric current) is produced.	In a chemical c produce an elec
1. Electrode	In a chemical cell, oxidation takes place at the negative electrode and reduction takes place at the positive electrode.	
2. Equilibrium	Equilibrium is attained in chemical cells when the voltage measured is equal to zero.	The E° values g Potentials are reaction procee the measured va measured voltag
D. Electrolytic cells	Redox reactions that do not occur spontaneously can be forced to take place by supplying energy with an externally applied electric current.	Electroplating, of brine, and e examples of the a chemical reac *3, p. 83.)*
	The use of an electric current to bring about a chemical reaction is called electrolysis.	
	In an electrolytic cell, oxidation takes place at the positive electrode, and reduction takes place at the negative electrode.	

Understandings and Fundamental Concepts	Supplementary Information
	Since this value is negative, we know the reaction will not take place.
Redox reactions that occur spontaneously may be employed to provide a source of electrical energy.	*Details of commercial cells and batteries are not required.*
When the two half-cells of a redox reaction are connected by an external conductor, a flow of electrons (electric current) is produced.	In a chemical cell, a chemical reaction is used to produce an electric current.
In a chemical cell, oxidation takes place at the negative electrode and reduction takes place at the positive electrode.	
Equilibrium is attained in chemical cells when the voltage measured is equal to zero.	The E° values given in Tables of Standard Oxidation Potentials are for definite concentrations. As a reaction proceeds these concentrations change and the measured value falls off until at equilibrium the measured voltage is equal to zero.
Redox reactions that do not occur spontaneously can be forced to take place by supplying energy with an externally applied electric current.	Electroplating, electrolysis of water, electrolysis of brine, and electrolysis of molten salts are examples of the use of an electric current to produce a chemical reaction. *(Refer to Unit 10, Section II, B, 3, p. 83.)*
The use of an electric current to bring about a chemical reaction is called electrolysis.	
In an electrolytic cell, oxidation takes place at the positive electrode, and reduction takes place at the negative electrode.	

Topics	Understandings and Fundamental Concepts	Supp
III. Balancing Redox Equations	In any reaction the loss of electrons by the species oxidized must be equal to the gain of electrons by the species reduced. There is a conservation of charge as well as a conservation of mass in a redox reaction.	

Several methods have bee
reactions. Teachers may

Two methods are illustra
reaction between copper
METHOD 1. Change in oxid
electrons): T
the reactants a

$Cu + HNO_3 \rightarrow Cu$

Step 1. Assign
element.

0 +1 +5 -2
$Cu + H\ N\ O_3$

Step 2. Determ
number (transf
elements that
(loss of 2 ele
(gain of 3 ele
that some of t
change. Only
oxidation numb
transfer.)
Step 3. Balanc
writing approp
materials oxid

3 x (-2e$^-$)

$Cu + HNO_3 \rightarrow 3C$

Step 4. Inse
with the cons
inspection).

$3Cu + 8HNO_3 \rightarrow$

y reaction the loss of electrons e species oxidized must be equal e gain of electrons by the es reduced.

is a conservation of charge as as a conservation of mass in a reaction.

Several methods have been used to balance redox reactions. Teachers may use the method they prefer.

Two methods are illustrated below for balancing the reaction between copper and dilute nitric acid:

METHOD 1. Change in oxidation number (transfer of electrons): The unbalanced equation shows the reactants and products

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$$

Step 1. Assign oxidation numbers to each element.

$$\overset{0}{Cu} + \overset{+1}{H}\ \overset{+5}{N}\ \overset{-2}{O_3} \rightarrow \overset{+2}{Cu}\ (\overset{+5}{N}\ \overset{-2}{O_3})_2 + \overset{+2}{N}\ \overset{-2}{O} + \overset{+1}{H_2}\ \overset{-2}{O}$$

Step 2. Determine the change in oxidation number (transfer of electrons) of those elements that change. For copper, 0 → +2 (loss of 2 electrons). For nitrogen, +5 → +2 (gain of 3 electrons). (It should be noted that some of the nitrogen does not show this change. Only the nitrogen that changes oxidation number is involved in the electron transfer.)

Step 3. Balance electron gain and loss by writing appropriate coefficients for the materials oxidized and reduced.

$$3 \times (-2e^-)$$

$$Cu + HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + H_2O$$

$$2 \times (+3e^-)$$

Step 4. Insert other coefficients consistent with the conservation of matter (balance by inspection).

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

It should
coefficie
oxidized
electrons
be change

METHOD 2. Balancing
are multi
coefficie

$3(\mathrm{Cu} \longrightarrow$

$2(\mathrm{NO}_3^- +$

$3\mathrm{Cu} + 2$

This equa
not inclu
in the re
a teacher
ions, the

$3\mathrm{Cu} + 8\mathrm{HN}$

It should be pointed out that, when the coefficients are assigned to the materials oxidized and reduced (thus balancing electrons lost and gained), this ratio cannot be changed.

METHOD 2. Balancing half-reactions: The half-reactions are multiplied by the appropriate coefficients to balance the electrons.

$$3(Cu \longrightarrow Cu^{++} + 2e^{-})$$

$$\underline{2(NO_3^{-} + 4H^{+} + 3e^{-} \longrightarrow NO + 2H_2O)}$$

$$3Cu + 2NO_3^{-} + 8H^{+} \longrightarrow 3Cu^{++} + 2NO + 4H_2O$$

This equation is a net equation, and does not include the ions that are not involved in the redox reaction (spectator ions). If a teacher wishes to include the spectator ions, the equation becomes

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

Unit 9 - Organic Chemistry

Topics	Understandings and Fundamental Concepts	Suppleme
I. Definitions	Organic chemistry is the chemistry of the compounds of carbon.	Organic compounds occur exte All living things are compos organic compounds.
	Carbon is able to form four covalent bonds not only with other kinds of atoms but also indefinitely with other carbon atoms. This makes possible a very large number of compounds.	Organic compounds are more n compounds. The major sources of raw mat chemicals are obtained are p other plant products, and an
II. Characteristics of Organic Compounds	Organic compounds are generally insoluble in water and soluble in nonaqueous solvents.	Organic compounds are genera to dissolve in nonpolar solv compounds that are somewhat acid, are soluble in water.
	Organic compounds are generally nonelectrolytes.	Organic acids are weak elect
	Organic compounds generally have low melting points.	Since most organic compounds the intermolecular forces ar compounds have relatively lo 300°C.).
	Reactions involving organic compounds are generally slower than those involving inorganic compounds.	Because of strong covalent b organic compounds do not rea complexes (intermediates), a place slowly. The activatio organic reactions is general
A. Bonding	The carbon atom usually forms compounds by covalent bonding.	
	The carbon atom has four valence electrons and can form four covalent bonds.	The four bonds of the carbon directed toward the corners
	The carbon atom can share electrons with other carbon atoms.	Two adjacent carbon atoms ca pairs of electrons.

Unit 9 - Organic Chemistry

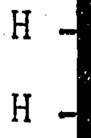
derstandings and Fundamental Concepts	Supplementary Information
ganic chemistry is the chemistry the compounds of carbon.	Organic compounds occur extensively in nature. All living things are composed predominantly of organic compounds.
rbon is able to form four covalent ids not only with other kinds of oms but also indefinitely with other rbon atoms. This makes possible a ry large number of compounds.	Organic compounds are more numerous than inorganic compounds. The major sources of raw materials from which organic chemicals are obtained are petroleum, coal, wood and other plant products, and animal sources.
ganic compounds are generally soluble in water and soluble in aqueous solvents.	Organic compounds are generally nonpolar, and tend to dissolve in nonpolar solvents. Those organic compounds that are somewhat polar, such as acetic acid, are soluble in water.
ganic compounds are generally ielectrolytes.	Organic acids are weak electrolytes.
ganic compounds generally have melting points.	Since most organic compounds are essentially nonpolar, the intermolecular forces are weak. Thus the compounds have relatively low melting points (under 300°C.).
actions involving organic compounds generally slower than those volving inorganic compounds.	Because of strong covalent bonding within the molecule, organic compounds do not readily form activated complexes (intermediates), and thus reactions take place slowly. The activation energy required for organic reactions is generally high.
carbon atom usually forms npounds by covalent bonding.	
carbon atom has four valence ectrons and can form four covalent ids.	The four bonds of the carbon atom are spatially directed toward the corners of a regular tetrahedron.
carbon atom can share electrons h other carbon atoms.	Two adjacent carbon atoms can share one, two, or three pairs of electrons.

Topics	Understandings and Fundamental Concepts	
	The covalent bonding results in compounds that are molecular in structure.	
B. Structural formulas	The covalent bond is usually represented by a short line (or dash) representing one pair of shared electrons. A formula showing the bonding in this manner is known as a structural formula.	It is importan compounds are *Teachers are v three-dimensio*
C. Isomers	Compounds which have the same molecular formula but different structures are called isomers.	For example,th $CH_3\cdot CO\cdot CH_3$ (molecular fo
	Isomers have different physical and chemical properties.	*to name these recognize comp* As the number possibilities number of iso
D. Saturated and unsaturated compounds	Organic compounds in which carbon atoms are bonded by the sharing of a single pair of electrons are said to be saturated compounds.	A bond formed one pair of e
	Organic compounds containing two adjacent carbon atoms bonded by the sharing of more than one pair of electrons are said to be unsaturated compounds.	A bond formed two pairs of bond. A bond sharing of th a triple bond
III. Homologous Series of Hydrocarbons	The study of organic chemistry is simplified by the fact that organic compounds can be classified into groups having related structures and properties. Such groups are called homologous series.	Each member o one before it As the member mass, the boi due to the in Unit 3, Secti

Understandings and Fundamental Concepts	Supplementary Information
The covalent bonding results in compounds that are molecular in structure.	
The covalent bond is usually represented by a short line (or dash) representing one pair of shared electrons. A formula showing the bonding in this manner is known as a structural formula.	It is important to remember that molecules of organic compounds are three-dimensional in nature. *Teachers are urged to use models to represent the three-dimensional structure of organic compounds.*
Compounds which have the same molecular formula but different structures are called isomers.	For example, the compounds $CH_3 \cdot CH_2 \cdot CHO$ (propanal) and $CH_3 \cdot CO \cdot CH_3$ (acetone) are isomers, both having the molecular formula C_3H_6O. *(Students are not required to name these compounds, but would be expected to recognize compounds as isomers.)*
Isomers have different physical and chemical properties.	As the number of atoms in the molecule increases, the possibilities of more spatial arrangements (thus, the number of isomers) increases.
Organic compounds in which carbon atoms are bonded by the sharing of a single pair of electrons are said to be saturated compounds.	A bond formed between carbon atoms by the sharing of one pair of electrons is referred to as a single bond.
Organic compounds containing two adjacent carbon atoms bonded by the sharing of more than one pair of electrons are said to be unsaturated compounds.	A bond formed between carbon atoms by the sharing of two pairs of electrons is referred to as a double bond. A bond formed between carbon atoms by the sharing of three pairs of electrons is referred to as a triple bond.
The study of organic chemistry is simplified by the fact that organic compounds can be classified into groups having related structures and properties. Such groups are called homologous series.	Each member of an homologous series differs from the one before it by a common increment. As the members of the series increase in molecular mass, the boiling point and freezing point increase due to the increase in the van der Waal forces. (See Unit 3, Section III, C, p. 21.)

Topics	Understandings and Fundamental Concepts	Supple
	Compounds containing only carbon and hydrogen are known as hydrocarbons. Most carbon compounds are named from, and can be considered as related to, corresponding hydrocarbons.	
A. Alkanes	The series of saturated hydrocarbons having the general formula C_nH_{2n+2} is called the alkane series.	*Students should be able formulas of the first f methane, ethane, propane* The alkane series is al or the paraffin series. The alkane series show fourth member (butane, *In naming isomers the I should be followed. On butane and pentane will*
B. Alkenes	The series of unsaturated hydrocarbons containing one double bond and having the general formula C_nH_{2n} is called the alkene series. In the I.U.C. (International Union of Chemists) system of nomenclature the alkenes are named from the corresponding alkane by changing the ending "-ane" to "-ene".	*Students should be able and formulas of the fir ethene, propene, butene* The alkene series is al series or the olefin se There are series of hyd than one double bond, s however, are not member
C. Alkynes	The series of unsaturated hydrocarbons containing one triple bond and having the general formula C_nH_{2n-2} is called the alkyne series. In the I.U.C. system of nomenclature the alkynes are named from the corres-	*Ethyne (acetylene) is t series that will be sub* The common name of the series, "acetylene", C_2 and should be familiar

erstandings and Fundamental Concepts	Supplementary Information
pounds containing only carbon and rogen are known as hydrocarbons. t carbon compounds are named from, can be considered as related to, responding hydrocarbons.	
series of saturated hydrocarbons ing the general formula C_nH_{2n+2} called the alkane series.	*Students should be able to recognize the names and formulas of the first five members of this series, methane, ethane, propane, butane, and pentane.* The alkane series is also called the methane series or the paraffin series. The alkane series show isomerism beginning with the fourth member (butane, C_4H_{10}). *In naming isomers the I.U.C. rules of nomenclature should be followed. Only the names of isomers of butane and pentane will be subject to examination.*
series of unsaturated hydro- bons containing one double bond having the general formula H_{2n} is called the alkene series.	*Students should be able to recognize the names and formulas of the first four members of this series, ethene, propene, butene, and pentene.*
the I.U.C. (International Union Chemists) system of nomenclature alkenes are named from the responding alkane by changing ending "-ane" to "-ene".	The alkene series is also called the ethylene series or the olefin series. There are series of hydrocarbons containing more than one double bond, such as the dienes. These, however, are not members of the alkene series.
series of unsaturated hydro- bons containing one triple bond and ing the general formula C_nH_{2n-2} called the alkyne series. In the .C. system of nomenclature the ynes are named from the corres-	*Ethyne (acetylene) is the only member of this series that will be subject to examination.* The common name of the first member of this series, "acetylene", C_2H_2, is still in general use and should be familiar to students.

Topics	Understandings and Fundamental Concepts	
	ponding alkane by changing the ending "-ane" to "-yne".	The seri
D. Benzene	The benzene series is a series of cyclic hydrocarbons having the general formula C_nH_{2n-6}.	*Benz* *seri* The C_6H_6
	All of the carbon-carbon bonds in the benzene ring are the same, and they have structure and properties intermediate between simple single bonds and simple double bonds.	The "sup bond
	In many of its reactions benzene behaves more like a saturated hydrocarbon than an unsaturated hydrocarbon.	 For of t repr 
IV. Other Organic Compounds	Homologous series of organic compounds occur in which one or more hydrogen atoms of a hydrocarbon have been replaced by other elements. These compounds are usually named from their corresponding hydrocarbons, but are not necessarily prepared directly from the hydrocarbon.	

ponding alkane by changing the ending "-ane" to "-yne".

The benzene series is a series of cyclic hydrocarbons having the general formula C_nH_{2n-6}.

All of the carbon-carbon bonds in the benzene ring are the same, and they have structure and properties intermediate between simple single bonds and simple double bonds.

In many of its reactions benzene behaves more like a saturated hydrocarbon than an unsaturated hydrocarbon.

Homologous series of organic compounds occur in which one or more hydrogen atoms of a hydrocarbon have been replaced by other elements.

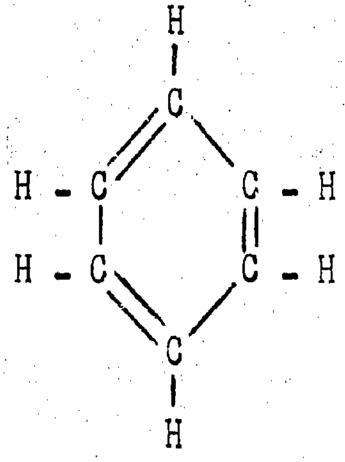
These compounds are usually named from their corresponding hydrocarbons, but are not necessarily prepared directly from the hydrocarbon.

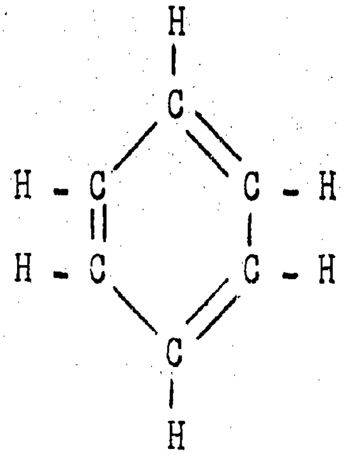
The alkyne series is also called the acetylene series.

Benzene and toluene are the only members of this series that will be subject to examination.

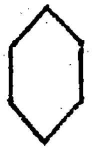
The simplest member of the benzene series is benzene, C_6H_6. The second member is toluene, C_7H_8, ($C_6H_5CH_3$).

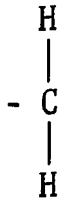
The structure of benzene is often represented as the "super position" or "average" of single and double bonds as shown below:

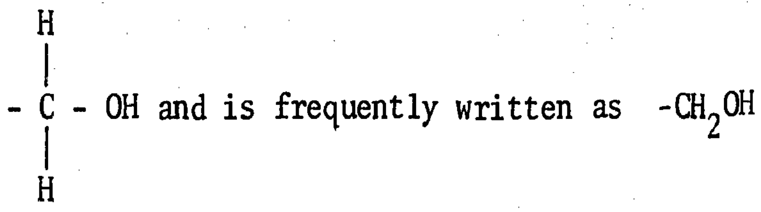


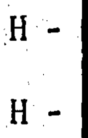


For simplicity the chemist often uses either one of the structures shown above. Other short-hand representations of the benzene ring that are used are:



Topics	Understandings and Fundamental Concepts	Supplem
	A functional group is a particular arrangement of a few atoms which gives characteristic properties to an organic molecule.	*Students should be able to alcohols, dihydroxy alcohol and organic acids by their*
	Organic compounds can often be considered as being composed of one or more functional groups attached to a hydrocarbon radical.	
A. Alcohols	In alcohols, one or more hydrogens of a hydrocarbon have been replaced by an -OH group.	No more than one -OH group carbon atom. The alcohols are not bases. alcohol does not form a hyd solution.
1. Primary alcohols	In primary alcohols, one -OH group is attached to the end carbon of a hydrocarbon. Primary alcohols contain the functional group $-CH_2OH$.	Since the functional group any hydrocarbon, the typica represented as R-OH, where of the molecule. The end g has the structural formula  $-\overset{H}{\underset{H}{C}}-OH$ and is frequent
	In the I.U.C. system of nomenclature, primary alcohols are named from the corresponding hydrocarbon by replacing the final "-e" with the ending "-ol."	*Students should know the na alcohols, methanol, ethanol pentanol.* The common names of the alc derived from the name of th carbon by changing the endi adding the class name. Thu called methyl alcohol. *It is recommended that the*

derstandings and Fundamental Concepts	Supplementary Information
functional group is a particular rangement of a few atoms which ves characteristic properties to organic molecule.	*Students should be able to recognize primary alcohols, dihydroxy alcohols, trihydroxy alcohols, and organic acids by their functional groups.*
ganic compounds can often be considered as being composed of one or more ctional groups attached to a rocarbon radical.	
alcohols, one or more hydrogens a hydrocarbon have been replaced an -OH group.	No more than one -OH group can be attached to one carbon atom. The alcohols are not bases. The -OH group of an alcohol does not form a hydroxide ion in aqueous solution.
primary alcohols, one -OH group attached to the end carbon of a drocarbon.	Since the functional group can be the end group of any hydrocarbon, the typical alcohol is frequently represented as R-OH, where "R" represents the rest of the molecule. The end group of a primary alcohol has the structural formula
imary alcohols contain the ctional group $-CH_2OH$.	 $-\overset{H}{\underset{H}{C}}-OH$ and is frequently written as $-CH_2OH$
the I.U.C. system of nomenclature, imary alcohols are named from the rresponding hydrocarbon by placing the final "-e" with the ding "-ol."	*Students should know the names of the five primary alcohols, methanol, ethanol, propanol, butanol and pentanol.* The common names of the alcohols were formerly derived from the name of the corresponding hydrocarbon by changing the ending "-ane" to "-yl" and adding the class name. Thus CH_3OH, methanol, was called methyl alcohol. *It is recommended that the I.U.C. nomenclature be used.*

Topics	Understandings and Fundamental Concepts	
2. Dihydroxy alcohols	Compounds containing two alcohol groups are known as dihydroxy (dihydric) alcohols or glycols.	*Students will n* *alcohols.* The most import called ethylene 
3. Trihydroxy alcohols	Compounds containing three alcohol groups are known as trihydroxy (trihydric) alcohols.	The most import having the stru  *Students should formula of glyc* Students who ha glycerol and it
†B. Aldehydes	Aldehydes contain the functional group -CHO.	Aldehydes are r The structural group is 

† Optional

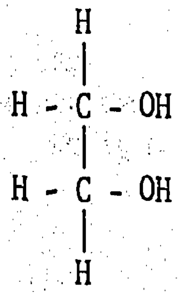
Understandings and Fundamental Concepts

Supplementary Information

Compounds containing two alcohol groups are known as dihydroxy (dihydric) alcohols or glycols.

Students will not be responsible for naming dihydroxy alcohols.

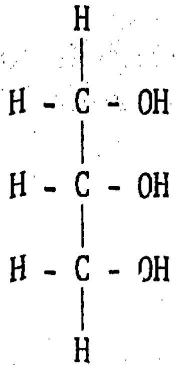
The most important glycol is dihydroxyethane, commonly called ethylene glycol. It has the structural formula:



$HOCH_2CH_2OH$

Compounds containing three alcohol groups are known as trihydroxy (trihydric) alcohols.

The most important trihydroxy alcohol is glycerol, having the structural formula:

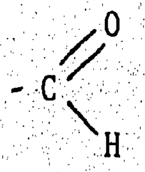


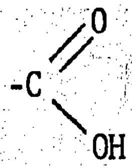
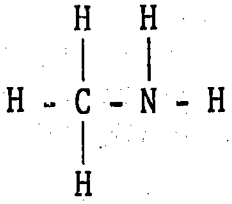
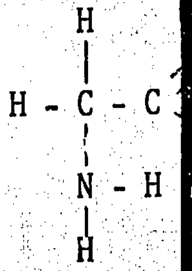
$HOCH_2CH(OH)CH_2OH$

Students should be able to recognize the name and formula of glycerol.

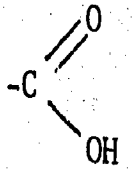
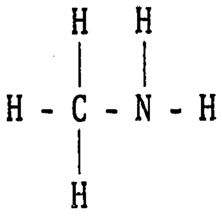
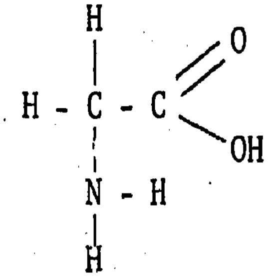
Students who have studied biology will be familiar with glycerol and its relationship to fats.

Aldehydes contain the functional group -CHO.

Aldehydes are represented by the general formula R-CHO. The structural formula of the aldehyde group is  $-C(=O)H$

Topics	Understandings and Fundamental Concepts	Suppleme
	In the I.U.C. system of nomenclature aldehydes are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-al."	The aldehyde of primary im which is generally referre formaldehyde.
C. Organic acids	Organic acids contain the functional group - COOH.	Acids are represented by t The structural formula of 
	In the I.U.C. system of nomenclature organic acids are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-oic" and adding the class name "acid."	The first two members of t HCOOH, and ethanoic acid, familiarly known by their and acetic acid. *Students should know both common name of the first (formic) acid and ethanoi*
†D. Amines	Amines are organic derivatives of ammonia. The functional group of an amine is R - NH_2.	An example of an amine is the structural formula:  $\begin{array}{ccccccc} & & H & & H & & \\ & & \vert & & \vert & & \\ H & - & C & - & N & - & H \\ & & \vert & & & & \\ & & H & & & & \end{array}$
†E. Amino acids	An amino acid is an organic compound containing both an amine group and an organic acid group.	An example of an amino ac with the structural formu 

ional

Understandings and Fundamental Concepts	Supplementary Information
In the I.U.C. system of nomenclature aldehydes are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-al."	The aldehyde of primary importance is methanal, HCHO, which is generally referred to by its common name, formaldehyde.
Organic acids contain the functional group - COOH.	Acids are represented by the general formula R-COOH. The structural formula of the acid group is: 
In the I.U.C. system of nomenclature organic acids are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-oic" and adding the class name "acid."	The first two members of this series, methanoic acid, HCOOH, and ethanoic acid, CH_3COOH, will be more familiarly known by their common names, formic acid and acetic acid. *Students should know both the I.U.C. name and the common name of the first two members, methanoic (formic) acid and ethanoic (acetic) acid.*
Amines are organic derivatives of ammonia. The functional group of an amine is $R - NH_2$.	An example of an amine is CH_3NH_2 (aminomethane) with the structural formula: 
An amino acid is an organic compound containing both an amine group and an organic acid group.	An example of an amino acid is CH_2NH_2COOH (glycine) with the structural formula: 

Topics	Understandings and Fundamental Concepts	Supplementa
†F. Proteins	Proteins are complex nitrogen compounds built up from amino acids.	All living organisms, vegetal contain proteins which are e processes.
V. Organic Reactions	Organic reactions generally take place more slowly than inorganic reactions.	
	Organic reactions frequently involve only the functional groups of the reacting species, leaving the greater part of the reacting molecules relatively unchanged during the course of the reaction.	
A. Substitution	Substitution means replacement of one kind of atom or group by another kind of atom or group. For saturated hydrocarbons, reactions (except for combustion and thermal decomposition) necessarily involve replacement of one or more hydrogen atoms.	The hydrogen atoms of satura replaced by halogen family a for these substitution react the products are called halo *In naming halogen derivative the I.U.C. rules of nomencla Only halogen derivatives of hydrocarbons will be subject*
B. Addition	Addition usually involves adding one or more atoms at the double bond resulting in saturation of the bond. Addition is characteristic of unsaturated compounds.	Some addition reactions are reactions between ions.
	Because addition reactions take place more easily than substitution reactions, unsaturated compounds tend to be more reactive than saturated compounds.	The addition of hydrogen to is called hydrogenation. Th requires the presence of a c temperature. The addition of chlorine and add) takes place at room tem formed are also called halog

† Optional

erstandings and Fundamental Concepts	Supplementary Information
teins are complex nitrogen pounds built up from amino acids.	All living organisms, vegetable and animal, contain proteins which are essential to their life processes.
anic reactions generally take place e slowly than inorganic reactions.	
anic reactions frequently involve y the functional groups of the cting species, leaving the greater t of the reacting molecules atively unchanged during the course the reaction.	
stitution means replacement of kind of atom or group by another d of atom or group. For saturated rocarbons, reactions (except for bustion and thermal decomposition) essarily involve replacement of or more hydrogen atoms.	The hydrogen atoms of saturated hydrocarbons can be replaced by halogen family atoms. The general term for these substitution reactions is halogenation and the products are called halogen derivatives. *In naming halogen derivatives of the hydrocarbons the I.U.C. rules of nomenclature should be followed. Only halogen derivatives of the first five alkane hydrocarbons will be subject to examination.*
ition usually involves adding one more atoms at the double bond ulting in saturation of the bond. ition is characteristic of aturated compounds.	Some addition reactions are about as fast as the reactions between ions.
ause addition reactions take place e easily than substitution ctions, unsaturated compounds tend be more reactive than saturated pounds.	The addition of hydrogen to an unsaturated substance is called hydrogenation. This reaction usually requires the presence of a catalyst and a raised temperature. The addition of chlorine and bromine (iodine doesn't add) takes place at room temperature. The compounds formed are also called halogen derivatives.

Topics	Understandings and Fundamental Concepts	Supplem
C. Esterification	Esterification is the reaction of an acid with an alcohol to give an ester and water.	Esterification is not an covalent compounds.
	acid + alcohol $\rightleftarrows$ ester + water	Esters usually have pleas many fruits, flowers, and
	Esterification proceeds slowly and is reversible.	Fats are esters derived f organic acids.
D. Saponification	The hydrolysis of esters by bases is called saponification.	To make soap, fat (a glyc hot alkali. The products organic acid) and glycero
E. Polymerization	Polymerization involves the formation of a large molecule from smaller molecules.	Synthetic rubbers, plasti other large chain molecul polymers.
		In nature polymerization proteins, starches, and organisms.
F. Fermentation	In the fermentation process, enzymes secreted by living organisms act as catalysts for the breakdown of molecules.	A common fermentation pro the fermentation of suga For example: $C_6H_{12}O_6 \xrightarrow[\text{zymase (from yea}]{}$

Understandings and Fundamental Concepts	Supplementary Information
Esterification is the reaction of an acid with an alcohol to give an ester and water.	Esterification is not an ionic reaction. Esters are covalent compounds.
acid + alcohol $\rightleftarrows$ ester + water	Esters usually have pleasant odors. The aromas of many fruits, flowers, and perfumes are due to esters.
Esterification proceeds slowly and is reversible.	Fats are esters derived from glycerol and long-chain organic acids.
The hydrolysis of esters by bases is called saponification.	To make soap, fat (a glycerol ester) is saponified by hot alkali. The products are soap (a salt of an organic acid) and glycerol.
Polymerization involves the formation of a large molecule from smaller molecules.	Synthetic rubbers, plastics such as polyethylene, and other large chain molecules synthesized by man are polymers. In nature polymerization occurs in the production of proteins, starches, and other chemicals by living organisms.
In the fermentation process, enzymes secreted by living organisms act as catalysts for the breakdown of molecules.	A common fermentation product, ethanol, results from the fermentation of sugar. For example: $C_6H_{12}O_6 \xrightarrow[\text{zymase (from yeast)}]{} 2C_2H_5OH + 2CO_2$

Unit 10 - Application of Principles of R

The material included under this heading should not be taught as a material should be related to, and integrated in, other units as i

Topics	Supplementary Informati
I. Chemical Theory and Industry	Pure research is directed toward the acquiring of knowledge consideration of the immediate practical application of tha The scientist's search for truth for its own sake often lea benefit of mankind.
	Industrial processes are concerned with obtaining a maximum products with maximum economic efficiency. An understandin use of the principles involved in an industrial process can affect the quality, yield, and cost of the product.
	It is recommended that teachers take advantage of the oppor arise during the course to relate theory to practical examp
II. Some Principles Employed in Industry	*Details of the processes discussed under this heading will Students are not required to memorize the equations involve*
A. Equilibrium and reaction rates	An application of the factors affecting the rates of reacti conditions make many processes practical. Examples of proc application of factors affecting rate of reaction and/or eq following:
1. Haber process	$N_2 + 3H_2 \rightleftarrows 2NH_3 + 22$ kcal.
	This process has already been used as an example to illustr pressure, temperature, and catalysts. (See Unit 6, Section

Unit 10 - Application of Principles of Reaction

The material included under this heading should not be taught as a unit itself. Rather, this material should be related to, and integrated in, other units as it applies.

Supplementary Information

Pure research is directed toward the acquiring of knowledge without the consideration of the immediate practical application of that knowledge. The scientist's search for truth for its own sake often leads to the benefit of mankind.

Industrial processes are concerned with obtaining a maximum yield of products with maximum economic efficiency. An understanding and proper use of the principles involved in an industrial process can materially affect the quality, yield, and cost of the product.

It is recommended that teachers take advantage of the opportunities that will arise during the course to relate theory to practical examples.

Details of the processes discussed under this heading will not be subject to examination. Students are not required to memorize the equations involved.

An application of the factors affecting the rates of reaction as well as equilibrium conditions make many processes practical. Examples of processes based on the application of factors affecting rate of reaction and/or equilibrium include the following:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3 + 22\text{ kcal.}$$

This process has already been used as an example to illustrate the effect of concentration, pressure, temperature, and catalysts. (See Unit 6, Section II, C, 1, pp. 49-50.)

Topics | Supplementary Information

2. Ostwald process

In the oxidation of ammonia to nitric acid the first reaction inv of nitrogen (II) oxide:

$$4NH_3 + 5O_2 \text{ (from air)} \longrightarrow 4NO + 6H$$

This reaction is favored by the use of a catalyst and by a high c

In one of the intermediate stages involved in this process, nitro oxidized to nitrogen (IV) oxide:

$$2NO + O_2 \longrightarrow 2NO_2 + 28 \text{ kcal.}$$

As can be predicted from the equation, heat does not favor the fo (IV) oxide, hence the nitrogen (II) oxide, formed by the highly oxidation of ammonia, must be cooled to favor the formation of n

This reaction is also favored by increasing the concentration of

3. Contact process

An important reaction in the manufacture of sulfuric acid is the (IV) oxide to sulfur (VI) oxide:

$$2SO_2 + O_2 \rightleftarrows 2SO_3 + 47 \text{ kcal.}$$

The rate of formation of SO_3 is slow at low temperatures, so the be raised to increase the rate of formation. However, since thi reaction, high temperatures decrease the stability of the SO_3. temperature of 400° to 450°C. is selected to secure reasonable reasonable rate of formation. This reaction is favored by the u

4. Solvay process

The Solvay process is used for the preparation of sodium bicarbo Ammonia is combined with carbon(IV)oxide and water to produce am

$$2NH_3 + CO_2 + H_2O \longrightarrow (NH_4)_2CO_3$$

which is converted by an excess of carbon(IV)oxide and water to

$$(NH_4)_2CO_3 + CO_2 + H_2O \longrightarrow 2NH_4HCO_3$$

Supplementary Information

the oxidation of ammonia to nitric acid the first reaction involves the production nitrogen (II) oxide:

$$4NH_3 + 5O_2 \text{ (from air)} \longrightarrow 4NO + 6H_2O$$

is reaction is favored by the use of a catalyst and by a high concentration of air.

one of the intermediate stages involved in this process, nitrogen (II) oxide is idized to nitrogen (IV) oxide:

$$2NO + O_2 \longrightarrow 2NO_2 + 28 \text{ kcal.}$$

can be predicted from the equation, heat does not favor the formation of nitrogen V) oxide, hence the nitrogen (II) oxide, formed by the highly exothermic catalytic idation of ammonia, must be cooled to favor the formation of nitrogen (IV) oxide.

is reaction is also favored by increasing the concentration of oxygen.

important reaction in the manufacture of sulfuric acid is the conversion of sulfur V) oxide to sulfur (VI) oxide:

$$2SO_2 + O_2 \rightleftarrows 2SO_3 + 47 \text{ kcal.}$$

e rate of formation of SO_3 is slow at low temperatures, so the temperature must raised to increase the rate of formation. However, since this is an exothermic action, high temperatures decrease the stability of the SO_3. In practice a compromise mperature of 400° to 450°C. is selected to secure reasonable yields of SO_3 at a asonable rate of formation. This reaction is favored by the use of a catalyst.

e Solvay process is used for the preparation of sodium bicarbonate and sodium carbonate. monia is combined with carbon(IV)oxide and water to produce ammonium carbonate,

$$2NH_3 + CO_2 + H_2O \longrightarrow (NH_4)_2CO_3$$

ich is converted by an excess of carbon(IV)oxide and water to ammonium bicarbonate,

$$(NH_4)_2CO_3 + CO_2 + H_2O \longrightarrow 2NH_4HCO_3$$

A saturated solution of sodium chloride is precipitate sodium bicarbonate, which is on

$$NH_4^+ + HCO_3^- + Na^+ + Cl^- \longrightarrow$$

Sodium carbonate is obtained by the decompo Sodium carbonate cannot be precipitated dir The Solvay process is an excellent example carbon (IV) oxide is prepared by the decomp oxide,

$$CaCO_3 \longrightarrow CaO$$

The lime is used to reclaim the ammonia fro

$$2NH_4Cl + CaO \longrightarrow$$

Thus the only raw materials consumed are li water which are inexpensive, and the only value.

5. Magnesium from sea water

Recovery of magnesium from sea water is ano addition of calcium hydroxide precipitates hydroxide:

$$Mg^{++} + Ca^{++} + 2OH^- \longrightarrow$$

B. Redox

Principles involving the "competition" for processes used by industry. Many basic red elements from compounds, refine them, prote

1. Reduction of metals

In extractive metallurgy, carbon is used ex

Metals that form relatively stable compound relatively stronger reducing agents. For e the iron:

$$Al + Fe^{+++} \longrightarrow Al^{+}$$

Supplementary Information

A saturated solution of sodium chloride is added to the ammonium bicarbonate to precipitate sodium bicarbonate, which is only sparingly soluble.

$$NH_4^+ + HCO_3^- + Na^+ + Cl^- \longrightarrow NH_4^+ + Cl^- + NaHCO_3$$

Sodium carbonate is obtained by the decomposition of the bicarbonate by heat. Sodium carbonate cannot be precipitated directly because of its greater solubility. The Solvay process is an excellent example of efficiency in a chemical process. The carbon (IV) oxide is prepared by the decomposition of limestone into lime and carbon (IV) oxide,

$$CaCO_3 \longrightarrow CaO + CO_2$$

The lime is used to reclaim the ammonia from the ammonium chloride,

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

Thus the only raw materials consumed are limestone, sodium chloride, and water which are inexpensive, and the only byproduct, calcium chloride, has commerical value.

Recovery of magnesium from sea water is another process based on solubilities. The addition of calcium hydroxide precipitates the magnesium ion as insoluble magnesium hydroxide:

$$Mg^{++} + Ca^{++} + 2OH^- \longrightarrow Ca^{++} + Mg(OH)_2$$

Principles involving the "competition" for electrons by atoms help to explain many chemical processes used by industry. Many basic redox reactions are employed to reduce or oxidize elements from compounds, refine them, protect them, and to produce valuable byproducts.

In extractive metallurgy, carbon is used extensively as a reducing agent.

Metals that form relatively stable compounds can be liberated from their compounds by relatively stronger reducing agents. For example, in the Thermit process, aluminum reduces the iron:

$$Al + Fe^{+++} \longrightarrow Al^{+++} + Fe$$

Topics	Supplementary Information
2. Oxidation of nonmetals	Nonmetals can be liberated by relatively stronger oxidizing agents. Fo the replacement of bromine in sea water by chlorine: $2Br^- + Cl_2 \longrightarrow Br_2 + 2Cl^-$
3. Electrolysis	Electrolysis (Unit 8, Section II, D, p.68) is used extensively in the c industry to produce elements from relatively stable compounds. For exa Electrolysis of water $2H_2O \longrightarrow 2H_2 + O_2$ Electrolysis of brine $2\ Cl^- + 2\ H_2O \longrightarrow Cl_2 + H_2 + 2OH^-$ Electrolysis of molten salts $2Na^+ + 2Cl^- \longrightarrow 2Na + Cl_2$
4. Prevention of corrosion	Corrosion of metals is caused by an undesirable redox reaction. As an is oxidized in the presence of water and oxygen. Chemical prin in the prevention of corrosion include: a. isolating the reactants H_2O and O_2 from the metal by various coatings including organic materials such as paints, oils, greases b. protecting with corrosion resistant metals such as chromium by means of electroplating, or protection by zinc by means of galvani
5. Bleaching	Compounds that are colored in their oxidized form can be bleached by a agent; for example, the bleaching of wood pulp by SO_2 in the manufactur Compounds that are colored in their reduced form can be bleached by an for example, the action of chlorine bleach on organic dyes in cloth.

Supplementary Information

Nonmetals can be liberated by relatively stronger oxidizing agents. For example, the replacement of bromine in sea water by chlorine:

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

Electrolysis (Unit 8, Section II, D, p.68) is used extensively in the chemical industry to produce elements from relatively stable compounds. For example:

Electrolysis of water $2H_2O \rightarrow 2H_2 + O_2$

Electrolysis of brine $2\,Cl^- + 2\,H_2O \longrightarrow Cl_2 + H_2 + 2OH^-$

Electrolysis of molten salts $2Na^+ + 2Cl^- \longrightarrow 2Na + Cl_2$

Corrosion of metals is caused by an undesirable redox reaction. As an example, iron is oxidized in the presence of water and oxygen. Chemical principles involved in the prevention of corrosion include:

a. isolating the reactants H_2O and O_2 from the metal by various coatings including organic materials such as paints, oils, greases

b. protecting with corrosion resistant metals such as chromium by means of electroplating, or protection by zinc by means of galvanizing.

Compounds that are colored in their oxidized form can be bleached by a reducing agent; for example, the bleaching of wood pulp by SO_2 in the manufacture of paper.

Compounds that are colored in their reduced form can be bleached by an oxidizing agent; for example, the action of chlorine bleach on organic dyes in cloth.

Appendix I. Schematic Energy Level Diagram

The diagram on the following page indicates the relative order of energy of the various sublevels, and can be used to build up electron configurations of the elements. The y-axis represents increasing energy, but the intervals between levels are not drawn to scale.

In the diagram the square boxes represent orbitals. Each orbital can contain two electrons. The numbers indicate the predicted position of each added electron. For example, the number 12 shows the orbital in which the twelfth electron would be added. An atom with twelve electrons would, of course, have electrons in positions numbered 1 - 11 as well.

It should be pointed out that there are variations from the order of entry of added electrons as predicted by this diagram. For example, if one examines the electron configurations of the transition elements in Period 4 as shown on the periodic table, where electrons are being added to the 3d subshell, it will be seen that the third she
regularly from 8 to 18.

This variation appear
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each d orbital is half-fil
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giving it the electron con
$3p^6$, $3d^8$, $4s^2$, or, as show
2 - 8 - 16 - 2. The atom
number, $_{29}Cu$, would be exp
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sublevel. In order to achi
one electron apparently is
giving the configuration 1s
(completely filled), $4s^1$,
table, 2 - 8 - 18 - 1 inst

Appendix I. Schematic Energy Level Diagram

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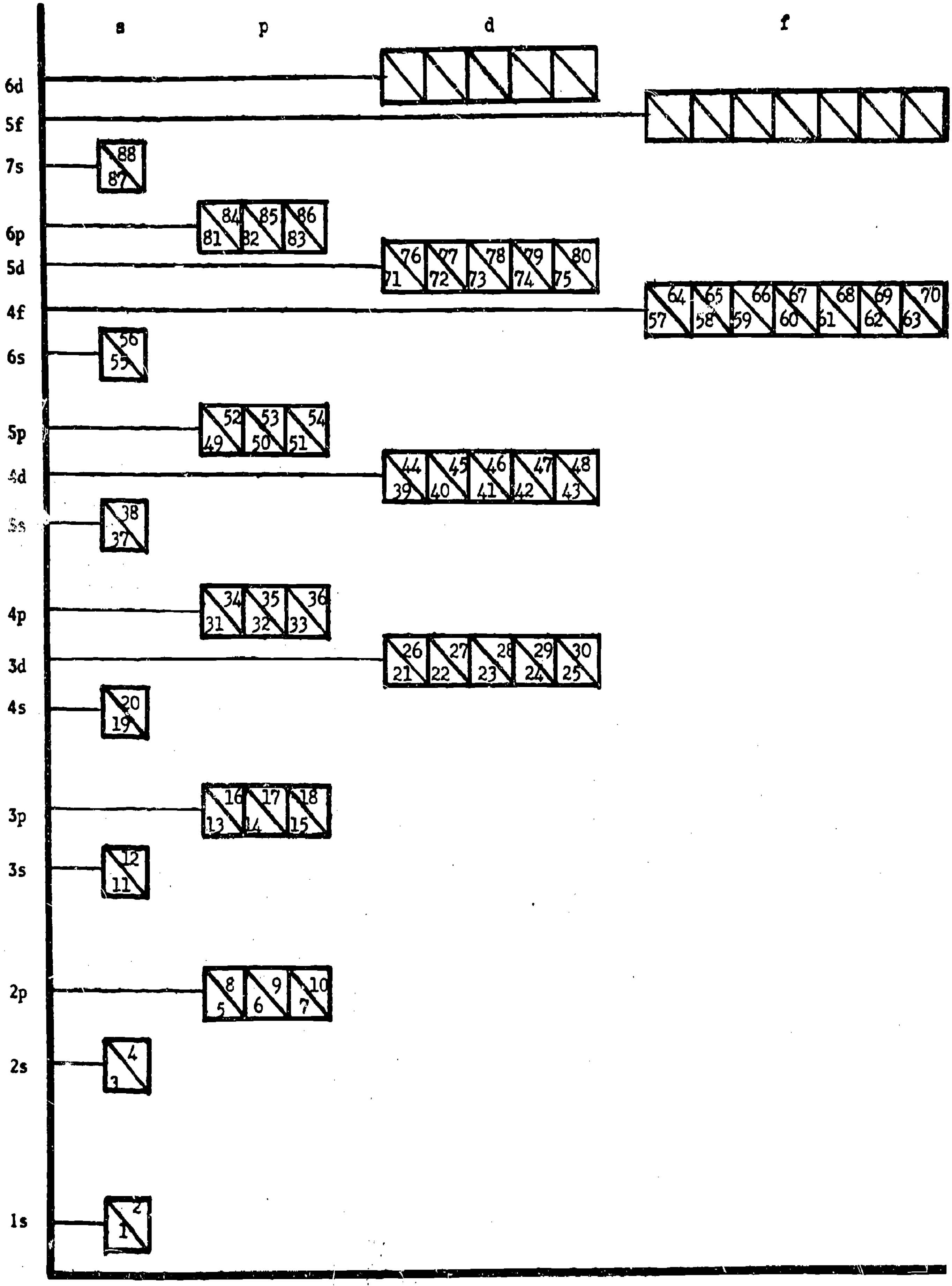
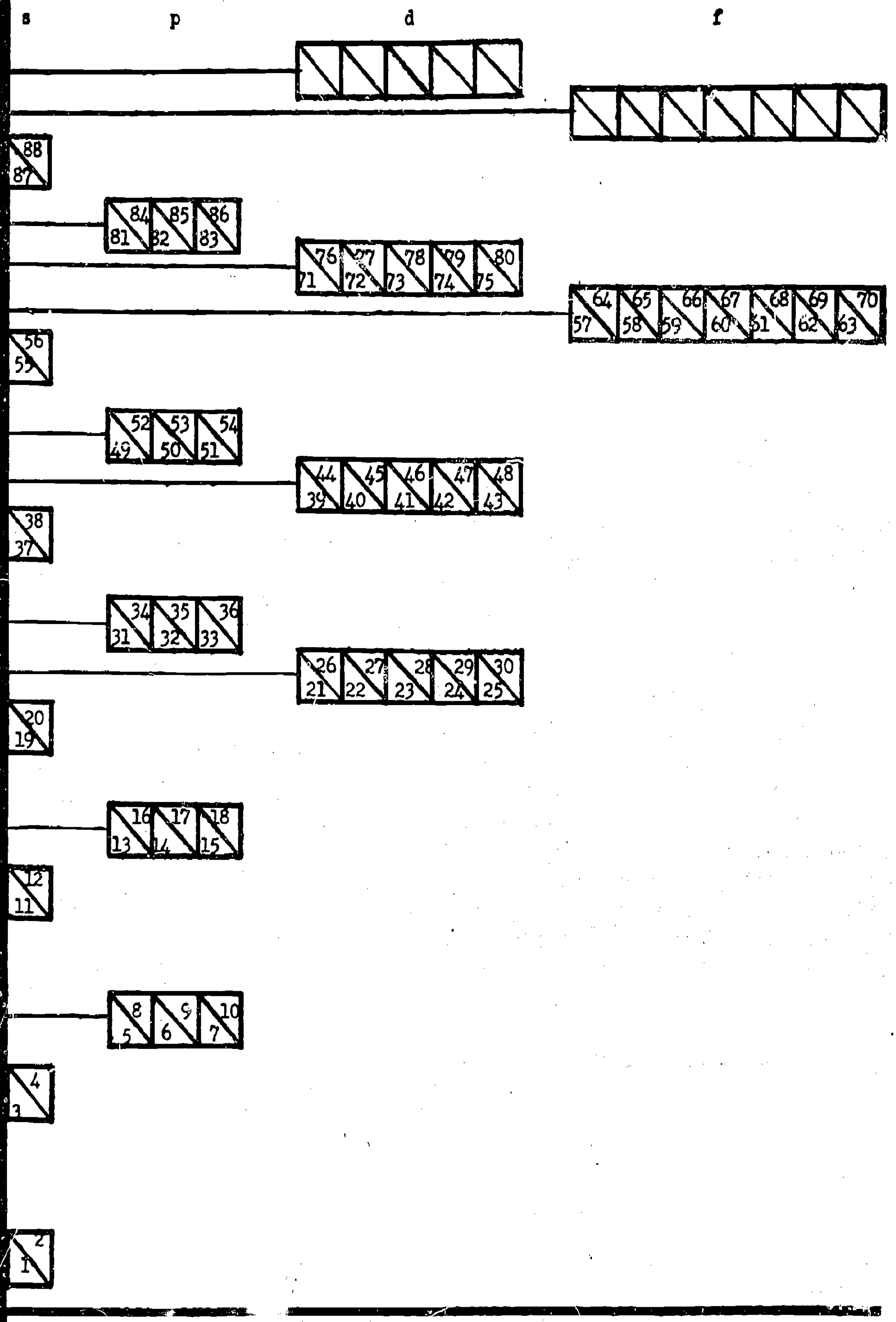
added to the 3d subshell, it will
be seen that the third shell does not increase regularly from 8 to 18.

This variation appears to be due to the peculiar stability of the arrangement in which the d orbitals are completely filled, and the arrangement in which each d orbital is half-filled. Thus $_{28}$Ni has the 28th electron in the orbital as shown on this diagram, giving it the electron configuration $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^8$, $4s^2$, or, as shown on the periodic table, 2 - 8 - 16 - 2. The atom with the next highest atomic number, $_{29}$Cu, would be expected to have nine electrons in the 3d sublevel. This would be just one less than the more stable arrangement of a completely filled 3d sublevel. In order to achieve this added stability, one electron apparently is promoted from the 4s orbital, giving the configuration $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^{10}$ (completely filled), $4s^1$, or as shown on the reference table, 2 - 8 - 18 - 1 instead of 2 - 8 - 17 - 2.



Appendix I Schematic Energy Level Diagram



Appendi I Schematic Energy Level Diagram

Appendix II.

Some Mathematical Concepts

Measurement

Measurement is a comparison of an unknown quantity with a known quantity. All measurements are subject to errors.

Errors may be due to the method used, environmental fluctuations, instrumental limitations, and personal error. Systematic errors tend to be in one direction. Random errors tend to fluctuate in both directions. The random error may be reduced by increasing the number of observations.

Significant Figures

The accuracy of a measurement or calculated result can be indicated by the use of significant figures.

A significant figure is one which is known to be reasonably reliable. In expressing the results of a measurement, one estimated figure is considered significant; for example, in measuring temperature, if the thermometer is calibrated in degrees, the reading may be *estimated* to the tenth of a degree. In this case, in the reading 20.3°, the figure "3" is considered significant.

Zeros which appear *in front* of a number are not significant figures. The number 0.083 contains two significant figures.

Zeros which appear *between* numbers are always significant. The number 803 contains three significant figures.

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Appendix II.

Some Mathematical Concepts

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enth of a degree. In this case,
, the figure "3" is considered

ear in front of a number are not
The number 0.083 contains two

ear between numbers are always
ber 803 contains three significant

Zeros which appear after a number are significant only (1) if followed by a decimal point, or (2), if to the right of a decimal point. The number 1800 contains two significant figures, but the numbers 1800. and 18.00 contain four significant figures.

For the whole numbers ending in two or more zeros there is no way of indicating that some, but not all, of the zeros are significant; for example, the number 186,000 would indicate three significant figures if no decimal point is expressed, and six significant figures if the decimal point is expressed. There is no way of indicating its accuracy to four or five significant figures except by the use of standard notation. (See the next section of this Appendix).

In calculations involving measured values the results must be rounded off to the number of significant figures justified by the accuracy of the measurement. Otherwise, the result of the calculation might appear to be more accurate than the original measurements.

The following rules will assist pupils when rounding off a number:

When the number dropped is less than 5, the preceding number remains unchanged; for example, 5.3634 to three significant figures becomes 5.36.

When the number dropped is 5 or more, the preceding number is increased by 1; for example, 2.4179 to three significant figures becomes 2.42.

97

When adding or subtracting, the answer should be rounded off to contain the least accurately known figure as the final one; for example,

Add

32.6
431.33
6144.212
6608.142 = 6608.1

Subtract

531.46
86.3
445.16 = 445.2

When multiplying or dividing, the answer should be rounded off to contain only as many significant figures as are contained in the least accurate number; for example,

Multiply

1.36
4.2
272
544
5.712 = 5.7

Divide

5.1 by 2.13

2.39 = 2.4
2.13/5.1000
4.26
840
639
2010

When adding, subtracting, multiplying, or dividing, numbers may be rounded off to one more than the number of significant figures to be carried in the answer before the manipulation is carried out; for example, 2.7468 x 3.2 = 2.75 x 3.2 = 8.8.

Standard Notation (scient

Standard notation sh
number of significant fig
mathematical operations w

In standard notation
form A x 10^n, where A is
the left of the decimal p
All the digits in A are s
into standard notation mo
only one non-zero digit i
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For example, 186,000 beco
becomes 5.20 x 10^{-5}. In
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Multiplication and d
To multiply or divide nu
multiply or divide the s
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Adjust the decimal point
than one non-zero digit
point. Examples:

$2.2 \times 10^4 \times 3.01 \times 10^2 =$

$2.2 \times 10^{-4} \times 3.01 \times 10^2$

$6.0 \times 10^3 \times 3.01 \times 10^4 =$

$6.0 \times 10^5 \div 3.0 \times 10^2 =$

$6.0 \times 10^5 \div 3.0 \times 10^{-2}$

$3.0 \times 10^2 \div 6.0 \times 10^5 =$

racting, the answer should be
the least accurately known
; for example,

Subtract

531.46
86.3
445.16 = 445.2

r dividing, the answer should
in only as many significant
d in the least accurate number;

Divide

5.1 by 2.13

$$\begin{array}{r l} 2.39 & = 2.4 \\ 2.13 \overline{)5.1000} & \\ \underline{4.26} & \\ 840 & \\ \underline{639} & \\ 2010 & \end{array}$$

racting, multiplying, or dividing,
off to one more than the number
to be carried in the answer
n is carried out; for example,
3.2 = 8.8.

Standard Notation (scientific or exponential notation)

Standard notation should be used to indicate the number of significant figures and to facilitate mathematical operations with large and small numbers.

In standard notation numbers are expressed in the form $A \times 10^n$, where A is any number with one digit to the left of the decimal point and n is an integer. All the digits in A are significant. To change a number into standard notation move the decimal point so that only one non-zero digit is to the left of it. Round off nonsignificant figures. The value of n is determined by counting the number of places the decimal point was moved. If the decimal point was moved to the left, n is positive. If it was moved to the right, n is negative. For example, 186,000 becomes 1.86×10^5, and 0.0000520 becomes 5.20×10^{-5}. In standard notation it is possible to indicate any desired number of significant figures. For example, if the figure 186,000 were known to four significant figures it would be written 1.860×10^5.

Multiplication and division in standard notation: To multiply or divide numbers in standard notation, multiply or divide the significant figure factors to obtain the new value of A, retaining the correct number of significant figures (opposite col.), and add or subtract the powers of 10 to obtain the new value of n. Adjust the decimal point if the new A has more or less than one non-zero digit to the left of the decimal point. Examples:

$$2.2 \times 10^4 \times 3.01 \times 10^2 = 6.6 \times 10^6$$

$$2.2 \times 10^{-4} \times 3.01 \times 10^2 = 6.6 \times 10^{-2}$$

$$6.0 \times 10^3 \times 3.01 \times 10^4 = 18 \times 10^7 = 1.8 \times 10^8$$

$$6.0 \times 10^5 \div 3.0 \times 10^2 = 2.0 \times 10^3$$

$$6.0 \times 10^5 \div 3.0 \times 10^{-2} = 2.0 \times 10^7$$

$$3.0 \times 10^2 \div 6.0 \times 10^5 = 0.50 \times 10^{-3} = 5.0 \times 10^{-4}$$

Addition and subtraction in standard notation: Numbers expressed in standard notation can be added or subtracted only if the powers of 10 are the same; for example, $5 \times 10^3 + 2 \times 10^3 = (5 + 2) \times 10^3 = 7 \times 10^3$. If the numbers to be added or subtracted have different powers of 10, then the powers must be equalized.

For example,

$2 \times 10^2 + 3 \times 10^3 = 2 \times 10^2 + 30 \times 10^2 = 32 \times 10^2 = 3.2 \times 10^3$.

Manipulation of

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their uncertaint

subtraction in standard notation:
in standard notation can be added or
f the powers of 10 are the same; for
$+ 2 \times 10^3 = (5 + 2) \times 10^3 = 7 \times 10^3$.
be added or subtracted have different
n the powers must be equalized.

$= 2 \times 10^2 + 30 \times 10^2 = 32 \times 10^2 =$

Manipulation of Units

In mathematical manipulations, units behave like algebraic quantities. In any equation the units on each side must be equivalent.

Graphs

Graphs should be used to illustrate mathematical relationships. A line representing the relationship should be smooth and probably will not pass through all measured points. Points should be circled to indicate their uncertainty.