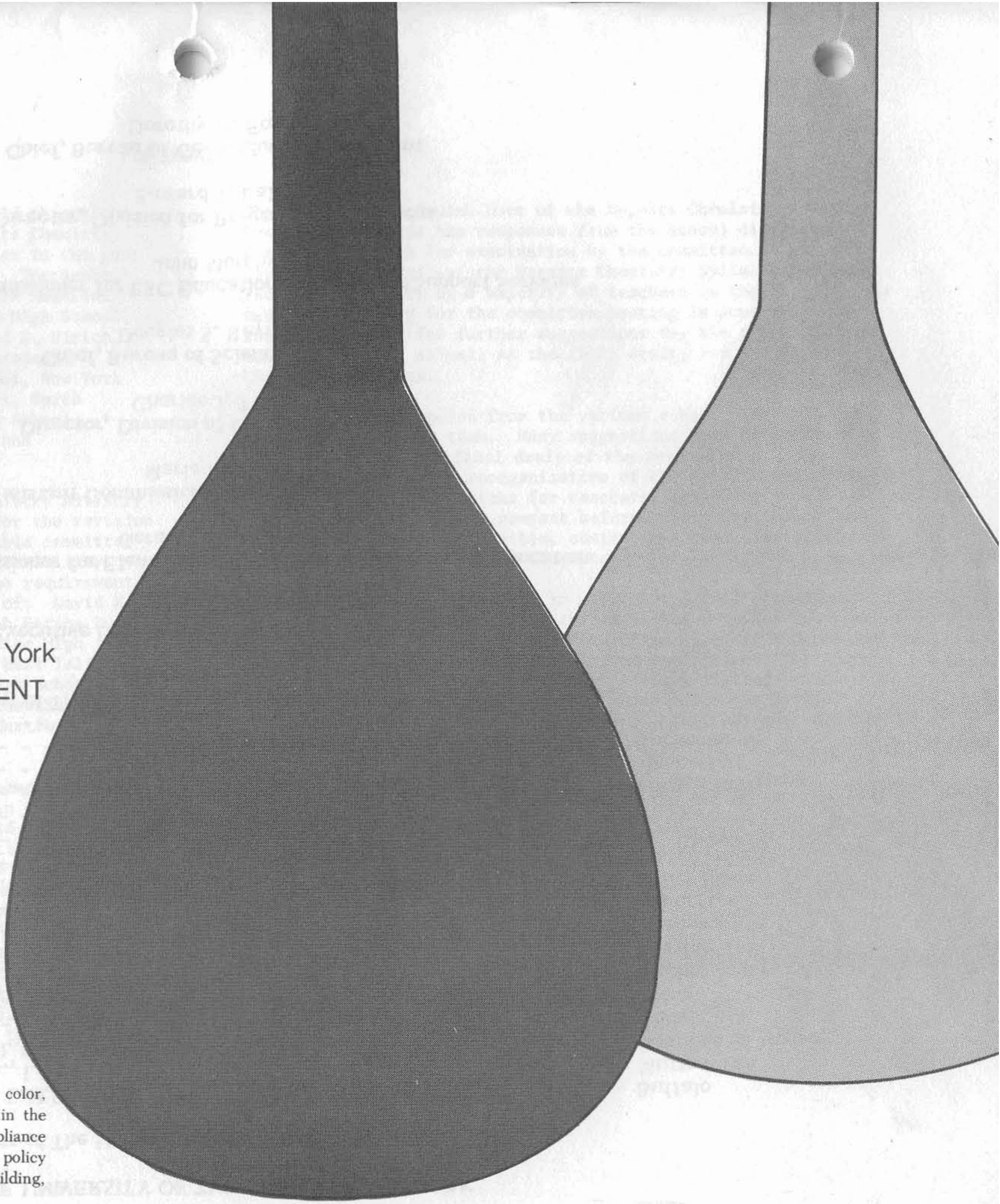
REGENTS CHEMISTRY SYLLABUS



The University of the State of New York
THE STATE EDUCATION DEPARTMENT
Bureau of Curriculum Development
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FOREWORD

During the spring of 1977, 15 area meetings were held throughout New York State to discuss the contents of the Regents Chemistry Syllabus. A committee of teachers was convened later in the year to collate the information gathered from the field. The members of this committee were: Frank Eickler, New Hyde Park Memorial High School; Beulah Patton Durr, William Nottingham High School, Syracuse; Richard Quest, Oneonta High School; Ronald N. Ulrich, Avon Central School; Martin L. Franklin, Franklin Academy, Malone; I. Fred Socolof, Grover Cleveland High School, New York City; Anthony A. Galitsis, Sleepy Hollow High School, North Tarrytown; Harry A. Kranepool, Bishop Loughlin High School, Brooklyn; Peter E. Demmin, Amherst Central School; and Lys Keneally, John Adams High School, New York City.

Based upon the information gathered, a Regents Chemistry Advisory Committee met in June 1978 to help set guidelines for the revision of the Regents Chemistry Syllabus. The intent of this committee was to establish a course of study having greater flexibility than the present course without lengthening the time requirement. The Regents Chemistry Advisory Committee consisted of: David W. Crane, Greece-Arcadia High School, Rochester; Beulah Patton Durr; Anthony A. Galitsis; Larry C. Grimm, Clarkstown Senior High School North; Lys Keneally; Chirakkal V. Krishnan, East Islip High School; Joseph J. Oberkrieser, Orchard Park High School; I. Fred Socolof; Edith M. Stillwaggon, Division Avenue Senior High School, Levittown; and Marlene Williams, New Hartford High School.

It was virtually impossible to get substantial agreement on what items had to be deleted from the present syllabus in order to include some additional items requested by the field. Therefore, the committee suggested a format which would essentially maintain the present content of the syllabus and at the same time incorporate the suggestions from the field. This was accomplished without altering the minimum requirements of the course.

During the summer of 1978, the writing team of Chirakkal Krishnan and Joseph Oberkrieser prepared a Development Draft of the Regents Chemistry Syllabus according to the guidelines established by the Regents Advisory Committee. Copies were mailed to all schools in October with a request for reactions to the draft and suggestions for its improvement.

Lys Keneally and Beulah Durr of the Regents Chemistry Advisory Committee collated the responses from the school districts, categorizing them for examination by the committee. The Developmental Draft of the Regents Chemistry Syllabus had been received favorably by a majority of teachers in the field. This made it necessary for the committee, meeting in June 1979, to recommend only a few further suggestions for the preparation of the Trial Draft, as well as the final draft, for the Regents Chemistry Syllabus.

Each and every response from the various school districts was read by the writing team. Many suggestions made by teachers were included in the final draft of the new syllabus. The end result has been a reorganization of the present syllabus which provides some options for teachers, resulting in a flexibility that was not present before. Some new topics have been added. Wherever possible, content has been clarified and updated according to modern concepts.

Acknowledging the variability in different school situations encountered in New York State, this syllabus provides the teacher with the opportunity to use selected units in greater depth and still have time to complete the minimum requirements for Regents examinations in chemistry. Teachers are not limited by the scope of this draft of the syllabus which lists the minimal requirements for the purpose of testing. They may add enrichment beyond the required areas.

The Developmental Draft, the Trial Draft, and the final draft of this syllabus were prepared by the Bureau of Science Education with the assistance of the Bureau of Curriculum Development. John V. Favitta, associate in Science Education, supervised the development of the revised syllabus.

INTRODUCTION

OVERVIEW

This course of study presents a modern view of chemistry suitable for students with a wide range of skills and abilities. The outline of topics provides the unifying principles of chemistry. The principles included in the outline are basic to the understanding of our environment.

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 consider important and which have local or intrinsic significance.

PREREQUISITIES

In order to be able to understand and use the principles of chemistry 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 chemistry course must have completed Ninth Year Mathematics or 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 or Course II is used directly in chemistry, the experiences in setting up and solving problems, and the analytical thinking developed in this course are most useful in chemistry.

TEACHING SEQUENCE

Teachers are at liberty to use any sequence of topics outlined on the next page and such teaching techniques as are appropriate for their students.

Teachers 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.

LABORATORY REQUIREMENT

Regents chemistry has a mandated laboratory requirement, and successful completion of this course earns for the students one unit of credit in a laboratory science. Each student must be engaged in laboratory activities for at least thirty 40-minute periods (or its equivalent) exclusive of time used in changing classes or teachers. Laboratory experiences should be related to daily classroom instruction and provide "reinforcement" for such instruction. Satisfactory written reports of these laboratory experiences must be prepared by the student. At the completion of the course, these laboratory reports must be kept in the school for six months following the date of the examination, except in instances where a senior or transferring student needs these reports for further work.

Pursuant to Section 207 of the Education Law, Section 8.2(c) of the Rules of the Board of Regents states, "Only those persons who have satisfactorily met the laboratory requirements as stated in the State syllabus for a science shall be admitted to the Regents examination in such science." This Rule of the Board of Regents applies to all students whether regularly enrolled in a Regents science class or studying independently. For students with severe physical handicaps, admission to science Regents examinations will be considered on an individual basis. Questions pertaining to the admission of students with severe physical handicaps to science Regents examinations should be directed to the Chief of the Bureau of Science Education, New York State Education Department, Albany, New York, 12234.

NOTE: The Unit 12* on Laboratory Activities presented in this syllabus is not intended to be comprehensive. Rather, the topics outlined in this unit represent the minimum requirements for testing in the Regents examination if this unit is elected by the students.

This course of study has been developed for a minimum of six 40-minute periods per week of instruction or its equivalent. These time allotments should include at least one double period each week for laboratory work. Because of the strong emphasis on the development of laboratory skills, seven periods per week of instruction is recommended.

STATE DIPLOMA CREDIT

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

SYLLABUS SCOPE AND CONTENT

The material in this syllabus is organized under three headings: Topics, Understandings and Fundamental Concepts, and Supplementary Information. Materials presented under all the headings will be included in the Regents examination. All the material organized under the headings, Activities, Minimum Requirements, and Supplementary Information in Unit 12*, Laboratory Activities, will also be subject to the Regents examination.

Units 1-9 are required of all students. In addition, a minimum of two Units from Units 5*, 6*, 8*, 9*, 10*, 11*, and 12* are required of all students.

TOPICAL OUTLINE

Units 1-9 are required of all students. A minimum of two units with an asterisk (*) are required of all students. Numbers in parenthesis refer to page location.

This material need not be taught as a unit, but may be incorporated in various places in the syllabus at the discretion of the teacher.

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Teachers who elect this Unit 5* may wish to teach materials outlined here in the appropriate places in the syllabus.

TOPICAL OUTLINE

Teachers who elect this Unit 6* may wish to teach these outlined materials in the appropriate places in Unit 6.

Teachers who elect this Unit 8* may wish to teach these outlined materials in the appropriate places in Unit 8.

Teachers who elect this Unit 9* may wish to teach these outlined materials in the appropriate places in Unit 9.

Teachers who elect this unit 10* may wish to teach materials outlined here in the appropriate places in the syllabus.

TOPICAL OUTLINE

UNIT 11* - NUCLEAR CHEMISTRY (77)

Teachers who elect this Unit 11* may wish to teach these outlined materials in the appropriate places in Unit 2.

I. Artificial Radioactivity (77)

  A. Artificial transmutation (77)
    1. Accelerators

II. Nuclear Energy (77)

  A. Fission reaction (77)
    1. Fuels
    2. Moderators
    3. Control rods
    4. Coolants
    5. Shielding
  B. Fusion reaction (78)
    1. Fuels
    2. High energy requirement
  C. Radioactive wastes (79)
  D. Uses of radioisotopes (79)
    1. Based on chemical reactivity
    2. Based on radioactivity
    3. Based on half-life

UNIT 12* - LABORATORY ACTIVITIES (81)

This unit is not intended to include all of the Regents Chemistry laboratory requirements. See introduction to Syllabus, p. iv.

I. Measurement (81)

II. Laboratory Skills (82)

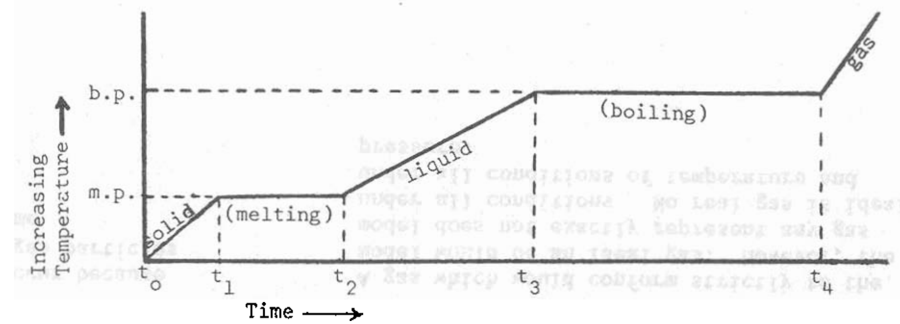
III. Laboratory Activities (83)

IV. Laboratory Reports (84)

	Understandings and Fundamental Concepts	Supplementary Information
I. 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.	
II. Matter		
A. Substances	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.	*For the purpose of Regents examinations, the slight differences in composition of a substance due to differences in isotopic masses will not be considered.*
1. Elements	An element is a substance which cannot be decomposed by chemical change.	All samples of an element are composed of atoms with the same atomic number.
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 chemically combined in a definite ratio.	All samples of a compound have identical composition. Binary compounds are compounds consisting of only two elements.
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 homogeneous (example, solutions or mixtures of gases) or heterogeneous (example, a mixture of iron and sulfur).
III. 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 include reactions involving a variety of forms of energy.
B. Energy and Chemical Change	Energy is either absorbed or given off in chemical changes.	Energy is absorbed by molecules when chemical bonds are broken. Energy is liberated when stronger bonds are formed. This concept is developed in Units 3 and 6.
		Although it is sometimes convenient to distinguish between physical changes and chemical changes, the distinction is not clear-cut and to most scientists is of little importance.

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Exothermic reaction	Energy is released in an exothermic reaction.	The role of energy in reactions is developed in Unit 6 and Unit 10*.
2. Endothermic reaction	Energy is absorbed in an endothermic 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.	
1. Calorie	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.	*For the purpose of Regents examinations this definition of calorie will be used.*
2. 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 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.
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.	
	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 equlibrium temperature occurs at 373 K.
IV. Phases of Matter	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 both heating and cooling curves.*

Topics	Understandings and Fundamental Concepts	Supplementary Information



This curve, if read from right to left, would illustrate a cooling curve.

Topics	Understandings and Fundamental Concepts	Supplementary Information
A. Gases	Gases take the shape and volume of the container.	*See also Unit 5*, V, p. 35.*
1. Boyle's law	At constant temperature, the volume of a given mass of gas varies inversely with the pressure exerted on it.	It is useful to introduce the equation 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.*
2. Charles' law	At constant pressure, the volume of a given mass of 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. 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 liquefy 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.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
3. Standard temperature and pressure (STP)	Standard temperature and pressure (STP) of a gas are defined as 0°C (273 K) and 760 mm of mercury (760 torr), or 1 atmosphere pressure.	Except for gases, standard temperature is usually 25°C (298 K). Since the volume of a given mass of gas varies with change in temperature and pressure, gas volumes are usually calculated to an arbitrary standard, STP.
4. Partial pressures	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.
5. Kinetic theory	Study of gas behavior has led to a model based on the following assumptions: • 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.	*In explaining and interpreting observed behavior, it is often convenient to use a model, which may be a "picture," a mathematical expression, or some 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 under new conditions.* 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.
6. Deviations from the gas laws	Deviations from the gas laws occur because the model is not perfect; the gas particles do have volume and do exert some attraction for each other.	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.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	These factors become significant when the space between gas particles is reduced, as in conditions of relatively high pressure and low temperature.	*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, van der Waals forces, p. 17.)*
7. Avogadro's hypothesis	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 amount of matter that contains 6.02×10^{23} (Avogadro's number) particles is called a mole of matter. One mole of any substance contains as many particles (molecules, atoms, ions, etc.) as there are atoms of ^{12}C in 12.000g of ^{12}C isotope.	Since it is inconvenient to work with individual particles (atoms, molecules, ions, electrons, etc.), chemists have chosen a unit containing many particles for comparing amounts of different materials. The mole is a unit which contains 6.02×10^{23} particles.
	A mole of particles of any gas occupies a volume of 22.4 liters at STP.	
B. Liquids	Liquids have definite volume but take the shape of the container.	Particles of a liquid have no regular arrangement and are in constant motion.
1. Vapor pressure	When a liquid substance changes to a gas the process is called evaporation. Evaporation tends to take place at the surface of a liquid and 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. The vapor pressure of water at various temperatures is given in the Reference Tables for Chemistry.*
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.	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	Supplementary Information
3. Heat of vaporization	The energy required to vaporize a unit mass of a liquid at constant temperature is called its heat of vaporization.	The energy involved in the change of phase is required to overcome the forces of attraction between particles and does not increase their average kinetic energy. Thus there is no increase in temperature during this phase change. *For the purpose of the Regents examinations, heat of vaporization at the normal boiling temperature will be used. See also Unit 5*, V, A. p. 35.*
C. Solids	Solids have definite shape and volume.	
1. Crystals	All true solids have crystalline structure. Crystals contain particles arranged in a regular geometric pattern.	Certain materials often considered solids are really super-cooled liquids; for example, glass, some plastics.
	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.
2. Melting point	The normal melting point is the temperature at which a solid substance will change to a liquid at 1 atmosphere pressure.	Melting points can be determined from cooling curves which are obtained experimentally. A melting point may also be defined as the temperature at which the solid and liquid phases can exist in 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.	*See also Unit 5*, V, B, p. 35.*
4. Sublimation	Sublimation is a change from the solid phase directly to the gas phase without passing through an apparent liquid phase.	*For the purpose of Regents examinations this definition of Sublimation will be used.* Solids that sublime have high vapor pressures and low intermolecular attractions. Examples of solids that sublime at room temperature are solid carbon dioxide (dry ice), solid iodine, and naphthalene.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Atoms		
A. Introduction to atomic structure	Our 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.*
B. Important subatomic particles	The atom is a complex unit of various particles	
1. Electrons	An electron has a mass of approximately 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 mass unit is defined as exactly 1/12 the mass of the ^{12}C atom.
b. Neutrons	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.
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.	
1. "Empty space" concept	Most of the atom consists of empty space.	Rutherford's gold foil experiments indicated the atom to be mostly empty space with the size of the nucleus very small compared to the size of the atom.
2. Nucleus	The mass of the atom is concentrated almost 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.
a. Atomic number	The atomic number indicates the number of protons in the nucleus.	*Teachers 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.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
b. Isotopes	Isotopes are atoms with the same atomic number but a different number of neutrons.	The atomic number identifies the element The difference in the number of neutrons affects the mass of the atom.
	For a given element the number of protons in the nucleus remains constant, but the number of neutrons may vary.	
c. Mass number	The mass number indicates the total 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.
d. Atomic mass (weight)	The mass of a neutral atom, atomic mass, is measured in atomic mass units based on ^{12}C equal to 12.000 atomic mass units.	
	The atomic mass of an element given in the Reference Tables for Chemistry 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.	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.
		The atomic mass (weight) of a single <u>atom</u> (isotope) such as neon-20 is 19.992 amu while the atomic mass (weight) of the <u>element</u> neon (which is the weighted average of all its natural isotopes) is 20.183 amu.
	The gram atomic mass (the mass of one mole of atoms) 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 is numerically equal to the atomic mass. See also Unit 5, Section II, A, p. 28.
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. Atomic structure models	The model for atomic structure of the elements has passed through many stages of development.	*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 structures of atoms and ions and to indicate electronic charges using the model.*
	In the Bohr model, electrons were considered to revolve around the nucleus in one of several concentric circular orbits.	

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Principal energy levels	The principal energy levels can be denoted by the letters K, L, M, N, O, P, Q, or by the numbers 1, 2, 3, 4, 5, 6, 7.	The principal energy level approximates how far the electron is from the nucleus.
	Electrons in orbits near the nucleus are at lower energy levels than those in 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."
	When atoms absorb energy, electrons may 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."
	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.	
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 lines has provided much of the evidence regarding energy levels within the atom.
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 accounted for the lines of the hydrogen spectrum, it did not account for the spectra of heavier and more complicated atoms.
	The term orbital refers to the average region of most probable electron location. Electrons occupy orbitals that may differ in size, shape, or orientation in space.	
1. Energy levels	The energy levels of electrons within an atom are represented by quantum numbers.	
a. Principal quantum numbers	The principal quantum number (n) represents the principal energy level.	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 in the periodic table.
b. Sublevels	The principal energy levels may be divided 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.