COURSE OUTLINE

Honors Chemistry

Rutherford High School
Rutherford, New Jersey

I. BASIC PHILOSOPHY

Recognizing that a student's attitudes or feelings about chemistry are just as important in the long run as his or her acquisition of specific chemical concepts, it is our goal to instill in our students the belief that chemistry is an exciting, relevant, human activity that can be enjoyable to study. To this end, the extensive use of laboratory experimentation, demonstrations and other hands-on activities are an integral part of the course.

Chemistry I Honors is taken by recommended students who have successfully completed Biology I Honors. While this is an introductory Chemtry course, emphasis is placed on logical reasoning and deductive thinking. More detail is placed on various topics and the material is covered at a faster pace.

II. METHODS EMPLOYED

direct teacher instruction
demonstrations
laboratory experiments
mini-activities (e.g. simulations) and laboratories
computer-assisted instruction
cooperative learning - problem solving
filmstrips and videos
library research
problem and question & answer sessions
homework

III. TEXT

Modern Chemistry , Tzimopoulos, Metcalfe, Williams and Castka. Holt, Rinehart and Winston, 1990

IV. BEHAVIORAL OBJECTIVES

At the completion of this course students should be able to:

I-1
1. define chemistry.
2. identify some applications of chemistry in everyday life.
3. list the major areas of chemistry.
4. describe the scientific method.

I-2
5. explain the difference between mass and weight.
6. define energy and list some types of energy.
7. state the laws of conservation of mass and energy.

8. explain the gaseous, liquid and solid states in terms of particles.
9. distinguish between physical and chemical properties of matter.
10. classify changes in matter as physical and chemical and give reasons for their choices.
11. distinguish between exothermic and endothermic reactions.

I-3 12. distinguish between homogeneous and heterogeneous matter.

13. distinguish between a pure substance and a mixture.
14. list the types of homogeneous matter.
15. state the law of definite composition and explain its meaning for the analysis of chemical compounds.

I-4 16. name the common elements, given their symbols.
17. write the symbols of the common elements, given their names.
18. describe the arrangement of the Periodic Table.
19. discuss the differences between metals, nonmetals, and metalloids.

II-1 20. name SI and other common units of length, mass and time.
21. perform unit conversions using the factor-label method.
22. name SI and other common units of volume and density.
23. perform density calculations.

II-2 24. define temperature and state the units in which it is measured.
25. convert temperatures between the Celsius scale and the Kelvin scale.
26. define heat and state its units.
27. perform specific heat calculations.

II-3 28. distinguish between accuracy and precision.
29. determine the number of significant figures in measurement.
30. perform mathematical operations involving significant figures.
31. use scientific notation to write numbers and carry out arithmetic operations.

II-4 32. use a given method to solve quantitative problems.
33. define and state equations for directly and indirectly proportional relationships.

34. plot graphs showing direct and indirect proportionality between two variables.

III-1 35. summarize the five essential points of Dalton's atomic theory.

36. explain the relationships between Dalton's atomic theory and the laws of conservation of matter and definite composition.

37. explain the law of multiple proportions.

III-2 38. summarize the observed properties of cathode rays that led to the discovery of the electron.

39. summarize the experiment conducted by Rutherford that led to the discovery of the nucleus.

40. describe the properties of protons, neutrons and electrons.

41. define "atom" and "isotope".

42. describe the atomic structures of the isotopes of hydrogen.

III-3 43. define "atomic number" and "mass number" and describe how they apply to isotopes and nuclides.

44. determine the number of protons, neutrons and electrons in a nuclide given the identity of the nuclide.

IV-1 45. discuss the dual wave-particle nature of light

46. explain the mathematical relationship among the velocity, wavelength and frequency of electromagnetic radiation.

47. discuss the significance of the line emission spectrum of hydrogen to the model of atomic structure.

48. describe the Bohr model of the hydrogen atom.

49. distinguish between an orbit and an orbital.

IV-2 50. list the four quantum numbers and describe their significance.

51. explain the number of sublevels corresponding to each of the main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

52. discuss the significance of the spin quantum number.

53. list the total number of electrons needed to fully occupy each main-energy level.

IV-3 54. state the Aufbau principle, Hund's rule, and the Pauli exclusion principle.

55. describe the arrangement of electrons around the atoms of any element using orbital notation, electron-dot notation or electron-configuration notation.

56. build up the electron configuration for atoms of any element, given the atomic number or identity of the element.

57. describe the noble-gas configuration and write it for any noble gas.

V-1 58. explain the role of Mendeleev in the development of the Periodic Table.

59. explain how the Periodic Law can be used to predict the physical and chemical properties of elements.

60. describe the modern Periodic Table.

61. describe how the elements belonging to a group of the Periodic Table are related by atomic number.

V-2 62. describe the relationship between electrons in sublevels and the length of each period.

63. locate and name the four blocks of the Periodic Table. Explain the reasons for these names.

64. discuss the relationship between group configurations and group numbers.

65. describe the location in the Periodic Table and the general properties of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.

V-3 66. define atomic and ionic radii, ionization energy, electron affinity and electronegativity.

67. compare the periodic variations of atomic radii, ionization energy, and electronegativity and state the reasons for these variations.

68. define valence electrons and state how many are present in atoms of each s and p block element.

69. describe electron configurations of common ions of elements in the s and p blocks.

70. compare the atomic radii, ionization energies, and electronegativities of the d block elements with those of

the main group elements.

VI-1 71. define chemical bond.
72. describe ionic and covalent bonding.
73. classify bonds according to electronegativity differences.
74. explain why most chemical bonds are neither purely ionic nor purely covalent.

VI-2 75. explain the relationships among potential energy, distance between approaching atoms, bond length, and bond energy.
76. state the octet rule.
77. list the six basic steps used in writing Lewis structures.
78. explain how to determine Lewis structures for polyatomic ions or molecules containing multiple bonds.
79. explain how to determine Lewis structures for polyatomic ions or molecules containing multiple bonds.
80. write the Lewis structure for a molecule or polyatomic ion, given the identity of the atoms combined and other specific information.

VI-3 81. compare the meaning of a chemical formula for a molecular compound and one for an ionic compound.
82. discuss the arrangements of ions in a crystal.
83. list and compare the distinctive properties of ionic and molecular compounds.

VI-4 84. explain the VSEPR theory.
85. predict the shapes of molecules or ions using VSEPR.
86. describe dipole-dipole forces, hydrogen bonding, and London forces.

VII-1 87. explain the significance of a chemical formula.
88. determine the formula of an ionic compound between any two given ions.
89. explain the two systems for distinguishing different ionic compounds of the same two elements.
90. name an ionic compound, given its formula.
91. using prefixes, name a binary molecular compound from its formula.
92. write the formula of a binary molecular compound, given its name.
93. list the names and formulas of the common laboratory

acids.

VII-2 94. list the rules for assigning oxidation numbers.

95. give the oxidation number for each element in the formula of a chemical compound.

96. name binary compounds using oxidation numbers and the Stock system.

VII-3 97. calculate the formula mass or molar mass of any given compound.

98. use molar mass to convert between mass in grams and amount in moles of a chemical compound.

99. give the number of molecules, formula units, or ions in a given molar amount of a chemical compound.

100. calculate the percent composition of a given chemical compound.

VII-4 101. define simplest formula, and explain how the term applies to ionic and molecular compounds.

102. find a simple formula from either percent or mass composition.

103. explain the relationship between the simplest formula and the molecular formula of a given compound.

104. find a molecular formula from a simple formula.

VIII-1 105. list three requirements for a correctly written chemical equation.

106. translate chemical equations into sentences.

107. write a word equation and a formula equation, given a description of a chemical reaction.

108. list three things you can determine about chemical reactants and products from a chemical equation.

109. balance a formula equation by inspection.

VIII-2 110. define and give general equations for synthesis, decomposition, single-replacement, and double-replacement reactions.

111. classify a reaction as synthesis, decomposition, single-replacement, double-replacement, or combustion.

112. predict the products of simple reactions, given the reactants.

VIII-3 113. explain the significance of an activity series.

114. list the generalizations based on the activity series that apply to single-replacement reactions.

115. list the generalizations based on the activity series that apply to synthesis reactions.
116. use the activity series to predict whether or not a given reaction will take place and what the products will be.

IX-1 117. define stoichiometry and distinguish between composition and reaction stoichiometry.
118. define mole ratio and describe its role in stoichiometric calculations.
119. name the four types of reaction-stoichiometric calculations.

IX-2 120. calculate the amount in moles of a reactant or product, given the amount in moles of a different reactant or product.
121. calculate the mass of a reactant or product, given the amount in moles of a different reactant or product.
122. calculate the amount in moles of a reactant or product, given the mass of a different reactant or product.
123. calculate the mass of a reactant or product, given the mass of a different reactant or product.

IX-3 124. define limiting reactant.
125. describe the method for determining which of the two reactants is a limiting reactant.
126. define theoretical yield, actual yield and percent yield
127. calculate percent yield, given actual yield and quantity of reactant.

X-1 128. state the kinetic theory of matter and describe how it explains certain properties of matter.
129. list the five assumptions of the kinetic theory of gases Define the terms ideal gas and real gas.
130. describe each of the characteristic properties of gases: expansion, low density, fluidity, compressibility, and diffusion.
131. describe the conditions under which a real gas deviates from ideal behavior.

X-2 132. explain why the measurable quantities volume, pressure, temperature, and number of molecules of gas are needed to describe properly the state or condition of a gas.
133. name the variables one must hold constant in order to study the (1) pressure-volume, (2) temperature-volume,

(3) pressure-temperature, (4) pressure-number of molecules, and (5) volume-number of molecules relationship of gases.

134. explain the relationship between the variables in each of the five pairs listed above in terms of the kinetic theory.

X-3 135. define pressure, explain how it is measured, and state the standard conditions of temperature and pressure.

136. state Boyle's Law and use it to calculate volume-pressure changes at fixed temperatures.

137. discuss the significance of absolute zero temperature. Use the Kelvin scale in calculations.

XI-1 138. describe the motion of liquid particles according to the kinetic theory.

139. discuss the properties of liquids in terms of the particle model.

XI-2 140. describe the motion of solid particles according to the kinetic theory.

141. discuss the properties of solids in terms of the particle model.

142. distinguish between the two types of solids.

XI-3 143. explain the relationship between equilibrium and changes of state.

144. predict changes in equilibrium using LeChatelier's principle.

145. explain what is meant by equilibrium vapor pressure.

146. describe the processes of boiling, freezing, melting and sublimation.

XII-1 147. distinguish between homogeneous and heterogeneous mixtures.

148. distinguish between electrolytes and nonelectrolytes.

150. compare the properties of suspensions, colloids, and solutions.

XII-2 151. list and explain three factors that influence the rate of dissolving of a solid in a liquid.

152. explain solution equilibrium and distinguish among saturated, unsaturated and supersaturated solutions.

153. explain the meaning of "like dissolves like" in terms of polar and nonpolar solvents.

154. compare the effects of temperature and pressure on solubility.

XII-3 155. define concentration using molarity, molality and percent by mass.

156. given the concentration of a solution, find the amount of solute in a given amount of solution.

157. given the concentration of a solution, find the amount of solution that contains a given amount of solute.

XIII-1 158. list five general properties of aqueous acids.

159. define and give an example of a traditional acid, a Bronsted acid and a Lewis acid.

160. name five acids commonly found in the laboratory.

XIII-2 161. list five general properties of aqueous bases.

162. define and given an example of a traditional base, a Bronsted base and a Lewis base.

163. name five bases commonly found in the laboratory.

XIII-3 164. define conjugate acid, conjugate base and conjugate acid-base pair.

165. write the formula for the conjugate acid of a base and for the conjugate base of an acid.

166. explain why the conjugate base of a strong acid is a base and why the conjugate acid of a strong base is a weak acid.

XIII-4 167. define acid anhydride and basic anhydride and given an example of each.

168. write the equations for the reactions of acid anhydrides and basic anhydrides with water.

XIII-5 169. describe the self-ionization of water.

170. define pH and give the pH of a neutral solution.

171. explain and use the pH scale.

172. given $[H_3O^+]$ or $[OH^-]$, find pH.

173. given pH, find $[H_3O^+]$ or $[OH^-]$.

XIV-1 174. explain heat of reaction, heat of formation and enthalpy

175. solve problems involving heats of reaction, heats of formation and heats of combustion.

176. explain the concept of bond energy and its relationship to heats of reaction.

XIV-2 177. explain the relationship between enthalpy change and the

tendency of a reaction to occur.
178. explain the relationship between entropy change and the tendency of a reaction to occur.
179. discuss the concept of free energy and explain how the value of this quantity is calculated and interpreted.

XIV-3 180. define kinetics and explain the two conditions necessary for a chemical reaction to occur.
181. discuss the five factors that influence reaction rates.
182. state what a catalyst is, and explain its effect on reaction rates.

XV-1 183. define chemical equilibrium.
184. explain the nature of the equilibrium constant.
185. write chemical equilibrium expressions and carry out calculations involving them.

XV-2 186. discuss the factors that disturb equilibrium.
187. discuss conditions under which reactions go to completion.

XVI-1 188. distinguish between chemical and nuclear reactions.
189. define mass defect and nuclear binding energy.
190. discuss the factors affecting the stability of atomic nuclei.
191. identify four types of nuclear reactions.

XVI-2 192. identify the unusual properties of radioactive nuclides.
193. define half-life and give an example.
194. name three types of radioactive emissions and write their nuclear symbols.
195. define the term transmutation and distinguish between naturally occurring and induced radioactive nuclides.
196. identify transuranium elements as the result of the bombardment of nuclei by neutrons.

XVI-3 197. describe the use of half-life to determine the age of an object.
198. give examples of the use of radioisotopes in medicine.
199. explain how radioisotopes can be used as tracers.

XVI-4 200. define nuclear fission, chain reaction, and nuclear fusion and distinguish among them.
201. describe the design of a nuclear reactor and explain the function of the control rods, nuclear fuel, and coolant.

V. COURSE OUTLINE

I. Matter, Energy and Change
   1. What is Chemistry?
      a. Chemistry is a Physical Science
      b. Chemistry in Modern Society
      c. What is a Chemical?
      d. Branches of Chemistry
      e. The Scientific Method
   2. Matter and Energy
      a. Definition of Matter
      b. Definition of Energy
      c. Law of Conservation of Energy
      d. States of Matter
      e. Properties and Changes of Matter
   3. Classification of Matter
      a. Mixtures
      b. Pure Substances
   4. The Chemical Elements
      a. Introduction to the Periodic Table
      b. Symbols for the Elements
      c. Types of Elements
      d. Metals

II. Measurements and Solving Problems
   1. Units of Measurement
      a. The SI Measurement System
      b. Fundamental SI Units
      c. Units of Measurement in Calculations
      d. Derived SI Units
   2. Heat and Temperature
      a. Definitions of Heat and Temperature
      b. Units of Heat
      c. Heat Capacity and Specific Heat
   3. Using Scientific Measurements
      a. Accuracy and Precision
      b. Significant Figures
      c. Scientific Notation
   4. Solving Quantitative Problems
      a. Steps in Solving Quantitative Problems
      b. Direct and Inverse Proportional Relationships

III. Atoms: The Building Blocks of Matter

1. The Atom: From Philosophical Idea to Scientific Theory
    a. Dalton's Atomic Theory
    b. Law of Multiple Proportions
2. The Structure of the Atom
    a. Discovery of the Electron
    b. Discovery of the Atomic Nucleus
    c. Composition of the Atomic Nucleus
    d. The Isotopes of Hydrogen
    e. The Sizes of Atoms
3. Weighing and Counting Atoms
    a. Atomic Numbers and Mass Numbers
    b. Relative Atomic Masses
    c. Average Atomic Masses of Elements
    d. The Mole, Avogadro's Number, and Molar Mass

IV. Arrangement of Electrons in Atoms

1. Refinements of the Atomic Model
    a. Wave-Particle Nature of Light
    b. Bohr Model of the Hydrogen Atom
    c. Spectroscopy
    d. Quantum Model of the Atom
2. Quantum Numbers and Atomic Orbitals
    a. Quantum Numbers
    b. Electrons in Each Main Energy Level
3. Electron Configurations
    a. Rules Governing Electron Configurations
    b. Representing Electron Configurations
    c. Elements of the Second and Third Periods
    d. Elements of the Fourth and Fifth Periods
    e. Elements of the Sixth and Seventh Periods

V. The Periodic Law

1. History of the Periodic Table
    a. Mendeleev and Chemical Periodicity
    b. Moseley and the Periodic Law
    c. The Modern Periodic Table
2. Electron Configuration and the Periodic Table
    a. Groups, Periods, and Blocks of the Periodic Table
    b. The s-Block Elements: Groups 1 and 2
    c. The p-Block Elements: Groups 3-8

d. The d-Block Elements: Transition Metals
e. The f-Block Elements: Lanthanides and Actinides

3. Electron Configuration and Periodic Properties
   a. Atomic Radii
   b. Ionization Energy
   c. Electron Affinity
   d. Ionic Radii
   e. Valence Electrons
   f. Electronegativity

VI. Chemical Bonding

1. Introduction to Chemical Bonding
   a. Types of Chemical Bonds
   b. Why Chemical Bonding Occurs
2. Covalent Bonding and Molecular Compounds
   a. Formation of Covalent Compounds
   b. The Octet Rule
   c. Lewis Structures
   d. Multiple Covalent Bonds
   e. Polyatomic Ions
3. Ionic Bonding and Ionic Compounds
   a. Formation of Ionic Bonds
   b. A Comparison of Ionic and Molecular Compounds
4. The Properties of Molecular Compounds
   a. VSEPR Theory
   b. Intermolecular Forces

VII. Chemical Formulas and Chemical Compounds

1. Chemical Names and Formulas
   a. Significance of a Chemical Formula
   b. Monatomic Ions
   c. Binary Ionic Compounds
   d. Compounds Containing Polyatomic Ions
   e. Binary Molecular Compounds
   f. Acids and Salts
2. Oxidation Numbers
   a. Assigning Oxidation Numbers
   b. Using Oxidation Numbers for Formulas and Names
3. Using Chemical Formulas
   a. Formula Masses and Molar Masses
   b. Molar Mass as a Conversion Factor

c. Percent Composition

4. Determining Chemical Formulas
    a. Calculation of Simplest Formula
    b. Calculation of Molecular Formula

VIII. Chemical Equations and Reactions

1. Chemical Equations
    a. Reading and Writing Chemical Equations
    b. Significance of a Chemical Equation
    c. Balancing Chemical Equations
2. Types of Chemical Reactions
    a. Synthesis Reactions
    b. Decomposition Reactions
    c. Single-Replacement Reactions
    d. Double-Replacement Reactions
    e. Combustion Reactions
3. Activity Series of the Elements
    a. Principles of the Activity Series
    b. Useful Generalizations Based on the Activity Series

IX. Stoichiometry

1. Introduction to Stoichiometry
    a. Applications of Stoichiometry
    b. Reaction-Stoichiometry Problems
2. Ideal Stoichiometric Calculations
    a. Mole-Mole Calculations
    b. Mole-Mass Calculations
    c. Mass-Mole Calculations
    d. Mass-Mass Calculations
3. Limiting Reactants and Percent Yield
    a. Limiting Reactant
    b. Percent Yield

X. Physical Characteristics of Gases

1. The Kinetic Theory of Matter
    a. The Kinetic-Molecular Theory of Gases
    b. The Kinetic Theory and the Nature of Gases
    c. Deviations of Real Gases from Ideal Behavior
2. Qualitative Description of Gases
    a. Pressure and Volume at Constant Temperature
    b. Temperature and Volume at Constant Pressure
    c. Pressure and Temperature

d. Relationship Between Pressure and Moles
e. Relationship Between Volume and Moles

3. Quantitative Description of Gases
   a. Pressure
   b. Boyle's Law: Pressure-Volume Relationship
   c. Charles' Law: Temperature-Volume Relationship
   d. Gay-Lussac's Law
   e. The Combined Gas Laws
   f. Dalton's Law of Partial Pressures

XI. Liquids and Solids

1. Liquids
   a. Kinetic Theory Description of the Liquid State
   b. Properties of Liquids and the Particle Model
2. Solids
   a. Kinetic Theory Description of the Solid State
   b. Properties of Solids and the Particle Model
3. Changes of State
   a. Equilibrium
   b. Equilibrium Vapor Pressure of a Liquid
   c. Boiling
   d. Freezing and Melting

XII. Solutions

1. Types of Mixtures
   a. Solutions
   b. Suspensions
   c. Colloids
2. The Solution Process
   a. Factors Affecting the Rate of Dissolving
   b. Solubility
   c. Factors Affecting Solubility
3. Concentrations of Solutions
   a. Molarity
   b. Molality

XIII. Acids, Bases and pH

1. Acids
   a. General Properties of Aqueous Acids
   b. Definitions of Acids
   c. Some Common Acids
2. Bases and Acid-Base Reactions

a. General Properties of Aqueous Bases
b. Definitions of Bases and Acid-Base Reactions
c. Types of Bases

3. Relative Strengths of Acids and Bases
a. Bronsted Acids and Bronsted Bases
b. Bronsted Acid-Base Pairs

4. Oxides, Hydroxides and Acids
a. Basic Oxides
b. Acidic Oxides

5. Aqueous Solutions and the Concept of pH
a. Self-Ionization of Water
b. The pH Scale
c. Calculations Involving pH

XIV. Reaction Energy and Reaction Kinetics

1. Thermochemistry
a. Heat of Reaction
b. Heat of Formation
c. Stability and Heat of Formation
d. Bond Energy and Reaction Heat

2. Driving Force of Reactions
a. Enthalpy and Reaction Tendency
b. Entropy and Reaction Tendency
c. Free Energy

3. Reaction Rate
a. Rate-Influencing Factors

XV. Chemical Equilibrium

1. The Nature of Chemical Equilibrium
a. Reversible Reactions
b. Equilibrium, A Dynamic State
c. The Equilibrium Constant

2. Shifting Equilibrium
a. Reactions That Run to Completion
b. Le Chatelier's Principle

XVI. Nuclear Chemistry

1. The Composition and Structure of the Nucleus
a. Mass Defect and Nuclear Binding Energy
b. Relationship Between Nuclear Stability and Neutron/Proton Ratio
c. Types of Nuclear Reactions

2. The Phenomenon of Radioactivity
    a. Naturally Occurring Radioactive Nuclides
    b. Artificially Induced Radioactive Nuclides
3. Applications of Radioactivity
    a. Radioactive Dating
    b. Radioisotopes Used in Medicine as Tracers
4. Energy from the Nucleus
    a. Nuclear Fission
    b. Nuclear Fusion

VI. SUGGESTED LABORATORY EXPERIMENTS

1. Observing Candles
2. Energy and Entropy phase changes
3. Density and Measurement
4. Precision of Measurements
5. Measuring Mass
6. Physical and Chemical Changes
7. Recovering the Parts
8. Identifying Elements, Compounds, and Mixtures
9. Quantitative Study of a Chemical Reaction
10. Mass Relationships Accompanying Chemical Changes
11. Volume and Weight of Liquids
12. Paper Chromatography
13. Flame Tests
14. Glass Tubing
15. Introduction to Qualitative Analysis
16. Molecular Size
17. The Periodic law
18. Halide Ions
19. Activity of Groups I, II, and III
20. Atomic Structure - Electron Arrangement
21. Chemical Bonds
22. Shapes of Covalent Molecules and Polarity
23. Water of Crystallization
24. Determining an Empirical Formula
25. Determining the Chemical Formula of MgO and CuS
26. Synthesis of Zinc and Iodine
27. Structure of Compounds
28. Formulas and Oxidation Numbers

29. Changes and Equations
30. Types of Chemical Reactions
31. Predicting Chemical Reactions
32. Relating Moles to Coefficients in a Chemical Reaction
33. Decomposition of $KClO_3$
34. Mass-Mass Relationships
35. Mole-Mass Relatioships
36. Mole Relatioships in a Chemical Reaction
37. Periodicity and Bonding - Oxygen Gas
38. Behavior of Gases
39. Effect of Temperature on Volume
40. Molar Mass of Butane
41. Mass-Volume Relationship in a Reaction
42. Molar Volume of a Gas - Hydrogen
43. Stoichiometry: Mass-Volume
44. Density of Solids
45. Boiling Point Elevation
46. Boiling Point of Water
47. Changes in State
48. Anti-Freeze is Anti-Boil
49. Popcorn Lab
50. Effect of Temperature on Solubility
51. Ionic Reactions
52. Solubility Curve for KNO_3
53. Reactions Between Ions
54. LeChatelier's Principle
55. Conductivity and Bonding
56. Electroplating
57. Hydronium Ion With Indicators
58. Hydrolysis - Water Reaction With Normal Salts
59. Effects of Acidic, Basic, and Salt Solutions on Nonmetals
60. Effects of Acidic, Basic, and Salt Solutions on Metals
61. Acid-Base Titrations
62. Heat of a Chemical Reaction
63. Energy Relationships of Metallic Ions
64. Heat of Combustion
65. Corrosion - An Electrochemical Problem
66. Quantitative Titration Involving Redox

VII. SUGGESTED DEMONSTRATIONS

I	1.	Ira Remsen Reaction
	2.	Cyalume Light Stick
	3.	Dancing Raisins
II	4.	Heat vs. Temperature
	5.	Density of Pennies
III	6.	Compounds Exemplifying The Law of Multiple Proportions
	7.	Black Box Experiment
	8.	Counting Out Atoms
IV	9.	Spectroscopes and the Excited Electrons
	10.	Shapes of Orbitals
V	11.	How Mendeleev Discovered the Periodic Table
	12.	Reactivity of the Alkaline Earth Metals
VI	13.	Reaction of Magnesium Metal
	14.	Size of a Molecule
	15.	Crystal Formation
VII	16.	Models of Chemical Compounds
	17.	Oxidation States of Manganese
VIII	18.	A Single Replacement Reaction of Zinc and Copper (II) Sulfate Solution
	19.	A Double Replacement Reaction of Lead (II) Nitrate and Potassium Iodide
IX	20.	Single Replacement of Zinc and Hydrogen Chloride
	21.	Bottle Caps and Bottles
	22.	Limiting Reactant of Hydrogen Chloride and Magnesium
X	23.	Preparation of Oxygen
	24.	Identifying Hydrogen
	25.	Properties of Carbon Dioxide
XI	26.	Pressure and Volume
	27.	Temperature and Volume
	28.	Soda Can and Air Pressure
XII	29.	Diffusion of Gases
XIII	30.	Phase Changes of Water Vapor, Water, and Ice
XIV	31.	1+1 Does Not Always Add Up To 2
	32.	Solution Equilibrium
	33.	Molar vs. Molal Solution
XV	34.	Conductivity of Solutions and Pure Substances
	35.	Dissociation vs. Ionization

XVI 36. Identifying Common Acids and Bases
37. Neutralization
38. Acids and Bases in The Kitchen
39. Some Effects of Acid Precipitation
XVII 40. Conductivity of Water
41. Behavior of Indicators
XVIII 42. Spontaneous Endothermic Reactions
43. Concentration vs. Reaction Rate
44. Temperature vs. Reaction Rate
XVIX 45. Equilibrium of Soda
XX 46. Bio-Clock
47. Electrochemical Cells

VIII. SUGGESTED FILMSTRIPS, SLIDES, VIDEOS AND COMPUTER SOFTWARE

Slides

1. Bonding Between Atoms of Different Elements: Metals
2. Periodic Table of the Elements: The Group Relations
3. Balancing Chemical Equations by Inspection
4. Collecting and Plotting Linear Data

Carousels

5. Nuclear Energy: Peril or Promise? Part I
6. Nuclear Energy: Peril or Promise? Part II
7. Chemistry of Life: Hormones and Endocrine System Part I
8. Chemistry of Life: etc. Part II
9. Chemistry of Life: etc. Part III

Filmstrips

10. Basic Chemistry for the Biologist: Part I
11. Basic Chemistry for the Biologist: Part II
12. Science and Society: Energy, Part I
13. Science and Society: Energy, Part II
14. Electron Transfer in Chemical Change
15. Proton Transfer in Chemical Change
16. Energy and Enthalpy in Chemical Change
17. Entropy and Free Energy: Why Chemical Reactions Occur
18. Subatomic Structure of the Atom: Isotopes
19. Subatomic Structure of the Atom: The Periodic Chart

20. Subatomic Structure of the Atom: Electrolytes, Acids, Bases
22. Subatomic Structure of the Atom: Double Replacement Reactions
23. Chemical Symbols, Formulas and Equations: Part I
24. Chemical Symbols, Formulas and Equations: Part II
25. Chemical Symbols, Formulas and Equations: Part III
26. Chemical Symbols, Formulas and Equations: Part IV Balancing Equations
27. Chemistry - Dissecting the Atom: Part I
28. Chemistry - Dissecting the Atom: Part II
29. Chemistry - Dissecting the Atom: Part III
30. Chemistry - Dissecting the Atom: Part IV
31. Chemistry - Dissecting the Atom: Part V
32. Chemistry - Dissecting the Atom: Part VI
33. Chemistry - Dissecting the Atom: Part VII
34. Energy and Entropy in Chemical Reactions: Entropy
35. Introduction to Energy and Radiant Energy
36. Chemical Energy
37. Electrical Energy and Mechanical Energy
38. Nuclear Energy and Energy: What Lies Ahead
39. Bonding Between Atoms of Different Elements
40. How to Use Thin Layer Chromatography
41. Periodic Table of the Elements: The Group Relationships
42. Titration Curves and Indicators
43. Balancing Chemical Equations by Inspection
44. Solar Energy
45. Nuclear Energy
46. Atom: From Atomic Hypothesis to Atomic Fact, Parts I-III
47. Models of the Atom: Parts I-V
48. Basic Biochemistry
49. Alternate Energy
50. Basics of Laboratory Safety
51. Nuclear War
52. An Introduction to Chemistry: Part I
53. An Introduction to Chemistry: Part II
54. An Introduction to Chemistry: Part III
55. All About Acids and Bases: Part I
56. All About Acids and Bases: Part II

57. All About Acids and Bases: Part III
58. All About Acids and Bases: Part IV
59. Radioactivity

Videotapes

60. Right to Know Educational Series: Corrosives

Computer Software

1. General Chemistry Disk #1 - The Elements (a)
2. General Chemistry Disk #2 - Nomenclature (a)
3. General Chemistry Disk #3 - Balancing Equations (a)
   (with Chemical Formulas and Chemical Equations)
4. General Chemistry Disk #4 - Atomic Weights (a)
5. General Chemistry Disk #5 - Percent Composition (a)
6. General Chemistry Disk #6 - CHEMAZE
7. General Chemistry Disk #7 - Ideal Gases
8. General Chemistry Disk #8 - pH: Acids & Bases in Water (a)
9. General Chemistry Disk #9 - The Metric System (a)
10. General Chemistry Disk #10 - Solutions (a)
11. Arrakis Advantage - Volume I, Disk 1 (b)
    Introduction to Atoms
    Scientific Models
    Dalton's Atomic Theory
    Law of Conservation of Matter
    Law of Definite Proportions
    Dalton's postulates
    Dalton's model
    discharge tube
    cathode rays
12. Arrakis Advantage - Volume I, Disk 2 (b)
    canal rays
    Thomson's e/m ratio
    Thomson's model
    Millikan's experiment
    Rutherford's experiment
13. Arrakis Advantage - Volume I, Disk 3 (b)
    Bohr's model
    Modern atomic model
    subatomic particles
    electrons, protons, neutrons, nucleus

atomic notation
atomic number
mass number
a.m.u.

14. Arrakis Advantage - Volume I, Disk 4 (b)
electron configuration
energy levels
sublevels
orbitals
valence shells and electrons
Final Exam

15. Arrakis Advantage - Volume II, Disk 1 (b)
Introduction
History of the Periodic Table
The Periodic Law

16. Arrakis Advantage - Volume II, Disk 2 (b)
Periods 1 through 3
Periods 4 and 5

17. Arrakis Advantage - Volume II, Disk 3 (b)
Periods 6 and 7
Periodic Properties
Family 0

18. Arrakis Advantage - Volume II, Disk 4 (b)
Family 7A
Family 1A
Metals
Nonmetals
Final Exam

19. Albion Chemical Element Game
20. Bonding in Molecules
22. Bonding Between Molecules
24. Moving Molecules
25. Reaction Time: Balancing Chemical Equations
26. Experiments in Chemistry

IX. EVALUATION

Student evaluation for this course is in accordance with the Board-approved policy on grading, including the mandatory **Homework** requirement and **Final Exams.**

Grades are also assigned based upon:

1. tests (teacher-made and/or standardized)
2. quizzes
3. class participation
4. labwork

The following evaluation criteria may also be used:

1. notebooks
2. research reports
3. special projects

Special projects to include:

1. label collection from common products showing the chemicals within those products.
2. article summaries for assigned and student-choice selections. Article summaries may include the following:
   a. children at risk in article
   b. CO_2 warming
   c. popcorn
   d. Bell lab scientists confirm new form of matter
   e. lead poisoning
   f. laser chemistry
3. career project using newspaper advertisements for chemistry-related positions.
4. poster to depict some aspect of the interaction of chemistry with the environment.

X. CORRELATION WITH STATE CORE COURSE PROFICIENCIES

PROFICIENCIES

1. Identify the components of the atom, i.e. location, charge, mass, name.
2. Utilize models (physical or mental) of molecules to write formulas for compounds.
3. Use appropriate basic chemical terminology.
4. Describe and predict the nature of elements and chemical reactions with the assistance of the Periodic Table.
5. Determine how energy and matter are related in many ways through their transportation, transformation and conservation.
6. Apply their knowledge of atomic structure to show its relationship to the chemical behavior of the elements.

7. Explain how the behavior of matter under various common circumstances is dependent on its physical state, i.e. solid, liquid, plasma, or gas.
8. Apply the mole concept to explain the behavior of matter and calculate quantitative relationships.
9. Compare and contrast physical, chemical, and nuclear changes.
10. Denote the conditions that establish an equilibrium (balance of forces) system and recognize the existence of equilibrium (balance of forces) systems in the real world.
11. Explain how matter undergoes chemical reactions whose nature, occurrence, and rates are dependent upon the intrinsic features of atoms and molecules and upon the surrounding environment.
12. Compare and contrast the changes of properties between reactants and products in a chemical transformation.
13. Illustrate how chemical systems control the natural and man-made world.
14. Cite examples of how technologies have been influenced by changes in our understanding of atomic theory from the early Greeks through Dalton to the modern models.
15. Logically gather, order, and interpret data through appropriate use of measurement and tools.

Minimum Proficiencies - Honors Chemistry

The course proficiencies listed below represent the minimum requirements to receive a passing grade for this course.

1. Students will acquire knowledge and understanding in each of the following areas:
    a. What is chemistry?
    b. The nature of properties and chemical change.
    c. How to analyze data and to write scientific research reports which emphasize good technical English.
    d. The meaning and use of chemical symbols and the writing of chemical equations.
2. Students will gain knowledge and understanding of atomic structure, including each of the following areas:
    a. Comparing and contrasting Dalton's and the Modern Atomic Theory.
    b. The contributions of J. J. Thompson, Milliken and Rutherford leading to the idea of a nuclear atom.
    c. Familiarity with Quantum concepts and the contributions made by early physicists, chemists and mathematicians.
3. Students will gain knowledge and understanding of the electronic structure of atoms and be able to:
    a. Write electron dot formulas.
    b. Write orbital notation.
    c. Become familiar with the d-orbitals.
4. Students will gain an understanding of the periodic table and the relationship between periodic properties of the elements and their electronic structure.
5. Students will learn to make predictions of how elements react based on their position on the periodic table.
6. Students will acquire knowledge and understanding of the basic concepts of chemical bonding and molecular orbital geometry, including understanding the differences between the groung and hybridized states.
7. Students will be able to make predictions concerning bond character.

8. Students will gain the knowledge and necessary skills enabling them to solve problems in chemistry which involve the Avogadro number and the mole concept, i.e. stoichiometric relationships.
9. Students will gain the knowledge and understanding of the properties and structures of matter in the liquid and solid states.
10. Students will acquire an understanding of the Kinetic Molecular theory as it applies to the three states of matter-solids, liquids and gases.
11. Students will acquire knowledge and understanding of the Gaseous State and be able to solve a variety of problems utilizing the following:

    a. Avogadro's Principle
    b. Charles and Boyle's Laws
    c. Gay Lussac's Principle
    d. The equation of State PV+nRT
    e. Molar Volume of a Gas

12. Students will gain knowledge and understanding of the different type of solutions and of the colligative properties of solutions.

13. Students will gain knowledge and understanding of the Bronsted-Lowry and Lewis theories of Acids and Bases and be able to do calculations involving:

    a. Neutralization Reactions
    b. Preparation of Molar Solutions

14. Students will acquire knowledge and understanding of oxidation reduction reactions, including half cell reactions and electrochemistry.

15. Students will gain an understanding of the Principles of Chemical Equilibrium and of the factors which can dis a system in equilibrium.

16. Students will gain an understanding of why reactions take

place by considering such factors as enthalpy and entropy as they relate to free energy.

17. Students will acquire a knowledge of introductory organic chemistry through a study of the homologous series of the Alkane, Alkene and Alkyne groups of hydrocarbons.

18. Students should be familiar with the use of a wide variety of lab equipment, including calculators, buret, balance, eudiometer, pH meter and micro computers.

19. Students should develop an awareness of possible hazards in the lab situation and know the safety procedures.

20. Students must meet the standards for a passing grade as outlined in the policy on grading in the Student/Parent Handbook.

21. Students must meet the requirements for school attendance as outlined in the Board of Education policy on attendance.