CURRICULUM GUIDELINES

Douglas College

A:	Division:	**Academic**	Date:	**23 May 2000**
B:	Department/ Program Area:	**Science and Technology**	New Course [ ]	Revision [X]
			If Revision, Section(s) Revised:	**J, P**
			Date Last Revised:	**13 September 1994**

C: **CHEM 104**	D: **Preparation for General Chemistry**	E: **4**
Subject & Course No.	Descriptive Title	Semester Credits

F: Calendar Description: This course is a basic introduction and is intended for students with little or no background in chemistry. Topics will include: exponential notations, significant figures, dimensional analysis, metric system, density, symbols, formulas, equations, atomic mass, mole, stoichiometry, solutions, molarity, classification of matter, gases and the periodic table, simple organic compounds, acids and bases, oxidation-reduction.

G: Allocation of Contact Hours to Types of Instruction/Learning Settings

Primary Methods of Instructional Delivery and/or Learning Settings:
Lecture / Laboratory

Number of Contact Hours: (per week / semester for each descriptor)

Lecture	**4 hours**
Laboratory	**2 hours**

Number of Weeks per Semester: 14

H: Course Prerequisites:

Douglas College MATH 101 or equivalent

I. Course Corequisites:

J. Course for which this Course is a Prerequisite:

CHEM 108

K. Maximum Class Size: 36

L: PLEASE INDICATE:

- [ ] Non-Credit
- [X] College Credit Non-Transfer
- [ ] College Credit Transfer: Requested [ ] Granted [ ]

SEE BC TRANSFER GUIDE FOR TRANSFER DETAILS (www.bccat.bc.ca)

M: Course Objectives/Learning Outcomes

The student will be able to:

1. explain the basis for the units of time and give one reason for the invention of more precise instruments for measuring time;
2. establish factors to convert one unit to another unit eg. hours to seconds;
3. list the basic metric units of mass, length, volume and also, the units used to record time;
4. distinguish between dependent and independent variables;
5. explain the term: function;
6. convert measurements in English units to metric units and vice-versa;
7. list the decimal equivalents of the metric prefixes kilo-, milli-, centi and deci-;
8. convert measurements given in one unit of a system (mass, volume, or linear to other units of the system);
9. express any number in exponential notation;
10. convert numbers written in exponential notation to their regular form;
11. perform addition, subtraction, multiplication and division of numbers written exponentially;
12. given certain data, prepare a bar graph using a uniform scale of measurement;
13. explain the difference between mass and weight;
14. distinguish between precision and accuracy;
15. express data to the correct number of significant figures;
16. give the number of significant figures in any number;
17. round off numbers correctly;
18. express the results of arithmetic operations to the proper number of significant figures;
19. interpret circle graphs by relating the sizes of sections with corresponding fractions and percentages;
20. determine the volume of solids and liquids;
21. calculate the volume of a regular object by using area and height;
22. list four types of volumetric glassware;
23. use a graduate cylinder, pipette and burette in the prescribed manner;
24. explain the TC and TD markings which occur on certain glassware;
25. use a triple beam and analytical balance;
26. explain the meaning of the precision range markings on simple apparatus;
27. list the general properties of gases;
28. list the more common gases by name and formula;
29. describe, in general terms, the effects of changes in temperature and pressure on gases;
30. state Boyle's and Charles' Laws;
31. given a temperature in degrees Celsius, convert it to degrees Kelvin and vice-versa;
32. list the measurable properties of gases;
33. explain the meaning of 'one atmosphere' of pressure and identify its equivalent in millimeters or centimeters of mercury;
34. apply Boyle's Law to find the new pressure or volume when the other is changed;
35. apply Charles' Law to find the new temperature or volume when the other is changed;
36. apply the combined gas laws to find the volume of a gas when both the volume and temperature are changed;
37. state Dalton's Law of Partial Pressures;
38. define density and specific gravity;
39. explain the procedure for determining the density of solids and liquids;
40. calculate the density, volume or mass of a substance from a given set of data;
41. draw broken-line graphs and explain their use;
42. distinguish between potential energy, kinetic energy, heat and temperature;
43. define the terms: calorie, specific heat and calorimetry;
44. calculate the amount of heat required to cause a given rise in temperature for a known mass of water;
45. given three of the following, calculate the fourth: (i) mass, (ii) specific heat, (iii) temperature change and (iv) amount of heat absorbed or released;
46. explain the construction and principle of operation of the mercury thermometer;
47. construct simple graphs and determine the slope of the line;
48. calculate proportionality constants from data arranged in tables or graphically represented;

50. determine extrapolated and interpolated points on a straight line;
51. list the three physical states of matter and explain their differences in terms of energy;
52. list the physical properties of matter;
53. explain the following terms: substance, matter, element, compound, mixture, homogeneous and heterogeneous;
54. classify common materials as compounds, elements or mixtures;
55. distinguish between atoms, molecules and ions;
56. list the postulates of Dalton's atomic theory;
57. state the Law of Conservation of Mass;
58. state the Law of Definite Proportions;
59. distinguish between chemical and physical properties;
60. classify changes as chemical or physical;
61. identify the three major subatomic particles by charge, location and relative mass;
62. explain the terms;:atomic mass, atomic number, isotope and atomic mass unit;
63. use the periodic table to assign atomic numbers and masses to the elements;
64. distinguish between metals and non-metals and locate their position in the periodic table;
65. write the symbols of the common elements;
66. name the element when given its symbol;
67. interpret chemical formulas in terms of number of atoms of each element present;
68. write formulas in terms of number of atoms of each element present;
69. explain the terms: formula, combining number, binary compound;
70. given the name, write the formula for some common inorganic compounds and vice-versa;
71. write the name and formula of any binary compound formed by the combination of any metal that you have learned from Groups lA, llA or lllA with any non-metal that you have learned from Groups VA, VlA and VllA;
72. list the elements which occur as diatomic molecules;
73. list the formula and name of the common polyatomic ions;
74. distinguish between an empirical and molecular formula;
75. balance simple chemical equations when all the formulas are given;
76. distinguish between reactants and products;
77. distinguish between chemical, word and balanced equation;
78. state the Law of Conservation of Energy;
79. from the formula of a substance, calculate its formula mass in atomic mass units;
80. distinguish between atomic and molecular mass;
81. define the terms: mole, molar mass and Avogadro's number;
82. given the chemical formula of any species, calculate the gram-formula mass and the percentage composition by mass;
83. given the number of grams of a chemical species of known mass, calculate the number of moles and (and vice-versa);
84. given the mass of a pure substance of known or calculable molar mass find the number of atoms, molecules or formula units, or, given the number of atoms, molecules or formula units, find the mass;
85. given the data from which the ratio of relative masses of the elements in a compound may be determined, calculate the empirical formula;
86. given the molar mass, calculate the molecular formula from the empirical formula;
87. identify and explain the symbols commonly used in chemical equations;
88. given a word description of a chemical reaction in which all reactants and products are identified, write the balanced chemical equation for the reaction;
89. given a chemical equation interpret it in terms of (a) atoms, molecules and/or formula units and (b) moles;
90. classify reactions as combination, decomposition, single replacement, double replacement or combustion;
91. explain the terms: exothermic, endothermic and heat of reaction;
92. list the four steps in the general method of mole ratio calculations;
93. determine the mole ratios involving any two specified substances when given the chemical equation;
94. given a chemical equation, or a reaction for which the equation can be written, and the number of moles of one species in the reaction, calculate the number of moles of any other species involved;
95. given a chemical equation, or reaction for which the equation can be written and the number of grams on one

M. 96. given two of the following, or information from which two of the following may be found, calculate the third: theoretical, actual and percentage yield;

97. given a thermochemical equation, or the information from which it may be written, calculate the amount of heat evolved or absorbed for a given amount of reactant of product; alternately, calculate how many grams of a reactant are required to produce a given amount of heat;
98. write and balance sample ionic equation;
99. define the terms: solution, solute, solvent, suspension and colloid;
100. construct a procedure that will classify a given substance as a pure material, a mixture or a solution;
101. list the general properties of solutions;
102. identify and explain the factors which determine the time required to dissolve a given solute;
103. explain the following terms;: dilute, concentrated, saturated, unsaturated and super-saturated;
104. given the necessary data, calculate percentage by weight and molarity of a solution;
105. calculate the molarity of a solution by diluting a stock solution by a known amount;
106. explain the term: dilution factor;
107. compare a solute to a pure solvent with respect to (a) boiling point and (b) freezing point;
108. distinguish between an acid and a base in terms of the classical properties and the ions associated with those properties;
109. explain the terms: neutralized, neutral, pH, indicator, titration, salt and end-point;
110. given the necessary titration data, calculate the molarity of an acid or base of unknown concentration;
111. name and use common laboratory equipment;
112. separate, by various means (decantation, filtration, extraction and distillation), mixtures into their pure components;
113. determine whether or not a chemical reaction occurs when two solutions are combined and state the criteria for this decision;
114. gain familiarity with a few substances and be able to identify them by name, formula and properties;
115. form a compound from its elements under controlled conditions;
116. find the percentage composition of this compound experimentally;
117. determine experimentally the empirical formula;
118. collect and weight a precipitate formed in a chemical reaction;
119. test for the presence of ions in solution;
120. list and organize data in the prescribed fashion;
121. operate and explain the laboratory techniques of titration;
122. define oxidation, reduction, oxidizing and reducing agents;
123. determine the oxidation number of some elements by using the periodic table;
124. given the formula for a compound, determine the oxidation number of each element in the compound;
125. given the oxidation numbers of component parts of a compound, write the formula for a compound.

N: Course Content

1. **Introduction and Orientation**
   Evaluation system, models, observation, classification, experimentation and the scientific method.

2. **Matter**
   Properties, heat, fusion and vaporization, melting and boiling point, density, purity, chemical and physical changes, homogeneous and heterogeneous, elements, mixtures and compounds.

3. **Measurement and Math Skills**
   Exponential notation, algebraic manipulation, units, SI system, conversions.

4. **Graphing**

5. **Treatment of Measurements**
   Uncertainty, significant figures, accuracy and precision.

6. **Atoms and Molecules**
   Atoms, molecules, compounds, physical and chemical change.

7. **Elements and Compounds**
   Symbols, periodic table, metals and nonmetals, formulas, nomenclature, oxidation number. Naming simple organic compounds if time permits.

8. **Atomic Mass and Moles**
   Mass, mole mass, molecular mass, empirical formula, molecular formula, percent composition.

9. **Electrical Nature of Matter**
   Charge, electrons, nucleus, protons, neutrons, ions, electrolysis, conductivity, ionic compounds, polyatomic ions, isotopes.

10. **Chemical Equations**
    Conservation of mass, balancing equations, special symbols, combustion, types of reactions.

11. **Stoichiometry**
    Reactions, mole relationships, mole-mole and gram-gram conversions, limiting reagent, theoretical yield, activity series.

12. **Solutions**
    Terms, concentration, solubility, solution stoichiometry, dissociation, precipitation, ionic and net ionic equations.

13. **Acids and Bases**
    Classification, properties of acids and bases, strong and weak acids, hydronium ion, electrolytes, titration.

14. **Gases**
    Direct proportionality, graphing review, inverse proportionality, variables - pressure, volume, temperature, moles - Gas Laws, Ideal Gas Equation, molar volume, molecular mass.

15. **Oxidation-Reduction Review**

P. Laboratory Course Content

Experiments will be selected from the following list:

1. Introduction and analytical balance
2. Chemistry in the kitchen.
3. Physical and chemical changes
4. Mass measurements.
5. Volume measurements.
6. Graphing
7. Conservation of mass.
8. Types of chemical reactions.
9. Periodic table.
10. The mole.
11. Gas Laws.
12. Stoichiometry.
13. Factors Influencing The Rates of Chemical Reactions.

O: Methods of Instruction

This course will consist of reading assignments, lectures, class discussions and problem sessions. Films, film strips, overhead projections, cassettes as well as any other audio-visual aids or programmed material will be used where appropriate. Close coordination will be maintained between laboratory and classroom work whenever possible. This will be accomplished by discussing laboratory work in class, using lab periods for problem solving and choosing experiments which will illustrate the practical aspects of the course material.

P: Textbooks and Materials to be Purchased by Students

Steven S. Zumdahl. INTRODUCTORY CHEMISTRY (4th Edition) D.C. Heath & Company. Toronto. 1999

Douglas College CHEMISTRY LABORATORY MANUAL

Douglas College 104 REPORT SHEET BOOKLET

Q: Means of Assessment

The student’s performance in the course will be evaluated in the following fashion:

1. Laboratory work (24%)

   Laboratory reports will be written in the laboratory note book and graded either pass or fail.

2. Examinations (76%)

   a) Course material will be assessed through tests and/or quizzes and/or problem sets. Throughout the semester a minimum of three in class evaluations will be given (46%)

   b) A final comprehensive examination during the exam period (30%)

R:	Prior Learning Assessment and Recognition: specify whether course is open for PLAR

Course Designer(s)

Education Council/Curriculum Committee Representative

Dean/Director

Registrar