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ABSTRACT

This course guide is the fifth in a series dealing with chemistry. The course is specifically designed for students interested in taking the Advanced Placement Chemistry examination and thus is equivalent to a first year college general chemistry course. Included in the guide are performance objectives, visual aids, a course outline, a list of textbooks and laboratory manuals, and suggestions for implementation. (JC)

AUTHORIZED COURSE OF INSTRUCTION FOR THE QUINMESTER PROGRAM

TYPES OF REACTIONS

5317.64

SCIENCE

(Experimental)

DADE COUNTY PUBLIC SCHOOLS

DIVISION OF INSTRUCTION • 1971

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SCIENCE

(Experimental)

Written by John Brennan
for the
DIVISION OF INSTRUCTION
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TYPES OF REACTIONS

COURSE DESCRIPTION

This course has been designed for the student interested in taking the Advanced Placement Chemistry examination. Together with the course "The Dynamic Nature of Atoms and Molecules", which should be offered first, this course offers an approach to the quantitative or problem solving aspects of the examination. Since there is a considerable essay portion in the examination, it is hoped the student will attain the proper amount of background material and not just equations and formulas to solve problems.

Due to the time allowed to cover the material, the suggested laboratories should be used at the teacher's discretion. The labs are not intended to encourage precise quantitative work, although calculations will be involved, but to develop an understanding of the concepts included in the course.

ENROLLMENT GUIDELINES

By the time a student has completed this program he should have the background equivalent to one completing a first year college general chemistry course. Thus it is suggested the first four units of chemistry, namely "Introduction to Chemistry," "Reactions of Atoms and Molecules," "Energy of Atoms and Molecules," and "The Dynamic Nature of Atoms and Molecules" be completed before enrolling in this course. The student should be proficient in Algebra, including proportionality constants, logarithms, and the quadratic equation.

TEXTBOOKS

State Adopted Texts:

1. Brown, Theodore. General Chemistry. Columbus, Ohio: Charles E. Merrill Company, 1968.

2. Choppin, Gregory; Jaffe, Bernard; Jackson, Lynn; and Summerlin, Lee. Chemistry. Morristown, New Jersey: Silver Burdett Co., 1970.

TEXTBOOKS (Continued)

3. Mahan, Bruce. College Chemistry. Reading, Mass.: Addison-Wesley Publishing Company, 1966.

Texts Not State Adopted:

4. Mortimer, Charles. Chemistry: A Conceptual Approach. New York: Reinhold Publishing Company, 1970.

Laboratory Manuals:

5. Hutton, Wilbur. General Chemistry Laboratory Text with Qualitative Analysis. Columbus, Ohio: Charles E. Merrill Publishing Co., 1968.

6. Laubengayer, A. W. Experiments and Problems in General Chemistry. New York: Holt, Rinehart and Winston, Inc., 1960.

7. Quick, Floyd J. Laboratory Manual for Introductory College Chemistry. New York: Macmillan Company, 1969.

IMPORTANT: Advanced Placement Chemistry is a course outline that is very necessary and may be purchased for $1.00 per copy from College Entrance Examination Board, Box 592, Princeton, New Jersey 08540.

PERFORMANCE OBJECTIVES

1. Given the rates of two opposing reactions at equilibrium, the student will:

    1. Compute the equilibrium constant expression,
    2. Use this expression to calculate concentrations within the system.

2. The student will differentiate among the Arrhenius, Bronsted-Lowry, and Lewis definitions of acids and bases.

PERFORMANCE OBJECTIVES (Continued)

3. Given a sample reaction involving a metal, non-metal or an oxide with water, the student will be able to use periodic relationships to determine the relative strength of the acid and/or base formed.

4. Given the specific gravity and weight percentages of a concentrated acid (or base), the student will calculate the:

   1. Molarity,
   2. Normality, and
   3. pH of the acid (or base).

5. Given the molar concentration of a strong acid (or base), the student will calculate the pH of the solution to the nearest hundredth, (It is implied here that given the pH of a solution, the student can calculate the concentration of a strong acid or base).

6. Given the K_a (or K_b) of a weak acid (or base) and the concentration, the student will calculate the pH of the solution. (It is implied here that given the pH and the K_a (or K_b) of a solution the student will be able to compute the concentration of the solution).

7. Given the ionization constant for each ionizing step for a specific polyprotic acid, the student will:

   1. Write the balanced ionic equation for each step.
   2. Calculate the concentration for any given species from the balanced equation.

8. Given an acid and a base of known concentration, the student will:

   1. Calculate the pH at any given point during the titration and,
   2. Plot a titration curve of volume vs. pH.

9. Given the K_a (or K_b) of a weak acid (or base) and the total volume required, the student will calculate the amount in grams of the soluble salt of the given acid (or base) needed to produce a buffer solution of desired pH.

10. The student will predict the spontaniety of a chemical reaction (of an isothermal system at constant pressure) from a change in enthropy (ΔH) and/or a change in enthalopy ($T\Delta S$) of the system using the Gibbs Free Energy equation.

PERFORMANCE OBJECTIVES (Continued)

11. Given the reactants (at unit activity) in an oxidation-reduction reaction and a standard electrode potential table, the student will:

    1. Predict the products of the reaction
    2. Balance the complete equation
    3. Compute the electrical potential (voltage) of the cell.

12. Given a sample solution (capable of undergoing oxidation or reduction) of unknown concentration and a standard electrode potential table, the student will select a proper oxidizing (or reducing) agent to titrate with the sample solution and, as a result of this titration calculate the concentration of the sample solution.

13. The student will use the Nernst equation to calculate the electrical potential (voltage) of an electro-chemical cell whose initial concentrations vary from unit activity by a given specific amount.

14. Given an instability constant for a complex ion system at equilibrium, the student will:

    1. Give the equation for the system
    2. Calculate the concentration of each species.

15. Given an element in any of the 7 Group A families or any of the 8 Group B families of the periodic table the student will predict the product (s) when the given element reacts with:

    1. An acid
    2. A halogen
    3. Oxygen
    4. Sulfur
    5. Nitrogen.

COURSE OUTLINE

I. Concentration of Equilibrium Systems

   A. Law of Mass Action

      1. Homogeneous systems

      2. Heterogeneous systems

COURSE OUTLINE (Continued)

B. Rate constant expression derived from law of mass action

C. Equilibrium constant expression derived from rates of opposing reactions at equilibrium

1. All systems at equilibrium have a Keq

2. Different subscripts (i.e. Keq, Kw, Ksp) are used only to differentiate specific equilibrium systems.)

II. Definitions of Acids and Bases

A. Arrhenius

1. Hydrogen ion concentration

2. Limited definition of an acid (or base)

B. Bronsted-Lowry

1. Conjugate pair

2. Strong vs. weak acid (or base)

3. Incorporates Arrhenius expanation of an acid

C. Lewis

1. Proton donor and proton acceptor

2. Includes ionic and non-ionic reactions

III. Periodic Properties of Acids and Bases

A. The chemical family or group

B. Acidic and basic anhydrides

C. Amphoterism

IV. Measuring the Strength of Acids (or Bases)

A. Ion concentration of strong acids (or bases)

B. K_A, K_B of weak acids (or bases)

1. Ion concentration from the K_A, or K_B

2. Percent ionization from K_A, or K_B

C. Computations involved when diluting reagents

V. Polyprotic Acids and Related Salts

A. Stepwise ionization and equations

1. Decreasing strength of the acid

2. Calculation of species concentration

B. Formation of salts of polyprotic acids and the dependency of the formula on the stochiometric ratio of the reactants

VI. Titrations

A. Measuring concentrations

1. Normality

2. Molarity

B. Selection of proper indicator for desired endpoint

C. Plotting and interpreting titration curves

1. Strong acids (or bases)

2. Weak acids (or bases)

VII. The Common Ion Effect

A. The effect of a common ion on equilibrium concentrations

1. Soluble salts of weak acids (or bases)

2. Le Chatelier's principle

B. Significance of pH values

C. Preparation of buffer solutions

COURSE OUTLINE (Continued)

VIII. Thermochemical Reactions

A. Enthalpy - Hess's Law

B. Entropy

1. As a measure of order

2. At constant temperature

C. Free energy (Gibbs free energy)

1. Relationship between enthalpy, entropy and the free energy equation

2. As a measure of available work from an isolated chemical system

IX. Oxidation - Reduction Reactions

A. Balancing oxidation - reduction reactions

1. Half-reaction method

2. Oxidation number method

B. Standard electrode potential table

1. As an aid in predicting spontaneous reactions

2. As a means of computing cell potential when reactants are at unit activity

3. As an aid in selecting a suitable oxidizing (or reducing) agent for a redox titration

C. Reactants other than at unit activity

1. Faradays laws of electricity

2. Relationship between electrode potential and free energy of a redox system

3. Use of Nernst equation to calculate electrode potential

COURSE OUTLINE (Continued)

X. Complex Ions (Coordination Compounds)

  A. Structure and nomenclature

  B. Instability constants for equilibrium systems

  C. Ion exchange resins (organic compounds)

XI. Chemical Reactions and Periodic Relationships

  A. Properties of chemical families

    1. Group I A through VIII A

    2. Transition elements

  B. Variation of chemical properties within a given family as a function of atomic weight

  C. Typical reactions with acids and common elements

    1. Acids

    2. Halogens

    3. Oxygen

    4. Nitrogen

    5. Sulfur

EXPERIMENTS

Quick, Floyd J. Laboratory Manual for Introductory College Chemistry. New York: Macmillan Company, 1969.

1. Oxidation-Reduction (pp. 151-157)
2. Sulfur and its compounds (pp. 193-196)
3. Carbon and its compounds

EXPERIMENTS (Continued)

Laubengayer, A. W. Experiments and Problems in General Chemistry. New York: Holt, Rinehart and Winston Inc., 1960.

4. Equilibrium and complex ions (p. 117)
5. Equilibrium involving substances that are only slightly soluble (p. 118)
6. Ionic equilibrium and pH (pp. 121-127)
7. Determination of equilibrium constants (pp. 127-129)
8. The common ion effect (pp. 129-130)
9. Changes in pH in acid-base titrations (pp. 130-132)
10. Buffer solutions (p. 133)
11. Oxides of metals (pp. 211-221)
12. Oxidation-reduction titrations (pp. 233-236)

The following experiments are more quantitative in nature and require more exacting work. These experiments would not be used to illustrate certain principles, but to show good lab technique and subject knowledge.

Hutton, Wilbur. General Chemistry Laboratory Text with Qualitative Analysis. Columbus, Ohio: Charles E. Merrill Publishing Co., 1968.

13. Kinetics of the reaction between hydrogen peroxide and the iodide ion (pp. 107-112)
14. The determination of the ionization constant for Iodic Acid (pp.139-144)
15. Acid-base indicators - A study of the dissociation of a weak acid (pp. 157-166)
16. The neutralization of a polyprotic acid with a strong base (pp. 173-182)
17. Single electrode potentials (pp. 199-206)
18. Potentiometric measurements of equilibria (pp. 207-214)

RELATED PROBLEMS

The following chart lists suggested problems for specific topics according to specific texts.

Under each author the chapter is given first, when applicable, and then the problem number, i.e. 8-15 is chapter 8, problem 15.

The problems are only suggested and are not listed in any degree of difficulty.

	Laubenga	Mahan	Brown	Choppin	Mortimer
Rate K Equivalent Weight	39, 43, 45	9-5, 9-8	14-2 14-4 12-12 12-15	15-1 19-6 19-7	12-1 12-3 8, 10
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Nernst Equation			16-11 16-15 16-19		9-20, 9-21 11-8, 11-9

DADE COUNTY 16 mm FILMS

1. Chlorine: A Representative Halogen
   AV# 1-10939, 15', C

2. Electric Interactions in Chemistry
   AV# 1-10839, 21', C

3. Introduction to Reaction Kinetics
   AV# 1-10859, 13', C

4. Nitric Acid Compounds and the Nitrogen Cycle
   AV# 1-11588, 19', C

5. Nitrogen and Ammonia
   AV# 1-10940, 16', C

6. Properties of Solutions
   AV# 1-30345, 28', B/W

7. Standard Solutions and Titration
   AV# 1-10926, 21', B/W

8. Sulfur And Its Compounds
   AV# 1-10937, 14', C

9. Transuranium Elements
   AV# 1-10900, 23', C

DADE COUNTY TRANSPARENCIES

1. Periodic Table of the Elements
   AV# 2-00143, C

SUGGESTIONS FOR IMPLEMENTATION

It is suggested that prior to or early in the course you poll the students as to whom will or will not take the Advanced Placement Examination. This may help you in regards to the depth in which each area should be covered.

SUGGESTIONS FOR IMPLEMENTATION (Continued)

It is also suggested, if possible, to secure an adequate number of copies of Mortimer's text (see Textbooks). This is written so the student can understand, and the depth of material covered is more than adequate for the APP examination.

When quizzing (those students interested in the APP exam), it is best to include essay as well as problem type questions, and keep in mind the APP exam does not press the students for time but comprehension of the question asked. Problem solving should include significant figures, correct units and a logical arrangement of work, as part credit is given in each case.

Most college labs are 3-4 hours in length, but the labs suggested here have convenient stopping places and even are broken into parts that can be done separately in one hour or less. It may be preferable in some cases, to do the labs as demonstrations to save time.

The laboratory apparatus and chemicals suggested for use are those that would be used in a Chem Study or equivalent course. It is strongly recommended that the students use an analytical balance, pH meter, electrolysis set-ups when available.

The chemicals suggested are common and pose no exceptional danger to the student.

REFERENCES

1. Ander, Paul and Sonessa, Anthony. Principles of Chemistry. New York: Macmillan Company, 1968.

2. Andrews, Donald and Kokes, Richard. Fundamental Chemistry. New York: John Wiley and Sons, Inc., 1965.

3. Hamm, Donald I. Fundamental Concepts of Chemistry: New York: Meredith Corporation, 1969.

4. Quick, Floyd. Introductory College Chemistry. 4th Edition. New York: Macmillan Company, 1968.

5. Runquist, Creswell and Head. Chemical Principles a Programmed Text. 3rd Edition. Minneapolis: Burgess Publishing Company, 1969.

6. West, Robert C. Handbook of Chemistry and Physics. Cleveland, Ohio: The Chemical Rubber Company, 1971.

Objectives	Texts	References	Experiments	Films
1	#1. pp. 308-316 #2. pp. 308-319,299-300 #3. pp. 151-163 #4. pp. 472-501	#1. pp. 436-440,565-583 #2. pp. 480-489 #3. pp. 159-164,197-206 #5. pp. 294-319,320-323	7,13,14	3,6
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3	#1. pp. 441-471,525-534 #3. pp. 453-470 #4. pp. 592-595	#3. pp. 240-246 #4. p. 461 #5. pp. 439-445	5,11	
4	#2. pp. 325-326,336-354 #4. pp. 282-2[illegible]4	#3. pp. 145-151,781-782 #5. pp. 66-67		
5	#1. pp. 358-359 #2. pp. 348-349 #3. pp. 191-193 #4. pp. 551-556	#3. pp. 781-782 #4. pp. 100-101 #5. pp. 68-70	6	
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15	#1. pp. 441-471,525-534 #2. pp. 384-410,462-498 #4. pp. 636-661	#1. pp. 608-631 #3. pp. 596-602,631-640 #5. pp. 339-345,124-135	2,3,11	1,4,5,8,9