JAMES E.
BRADY / GERARD E.
HUMISTON

INSTRUCTOR'S MANUAL TO ACCOMPANY

GENERAL CHEMISTRY

PRINCIPLES AND STRUCTURE

THIRD EDITION

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INTRODUCTION

The text, General Chemistry: Principles and Structure, was written for a full year General Chemistry course for science students. These would include students majoring in agriculture, biology, chemistry, engineering, geology, pharmacy and the allied health professions, physics, and those preparing for medical, dental, or veterinary school.

The scope and emphasis of general chemistry courses at different colleges and universities vary greatly. Textbooks naturally reflect this by trying to incorporate sufficient material so that they can be adapted to a variety of educational situations. This text covers all of the topics any ordinary science student might be expected to learn in general chemistry. This does not mean, of course, that everything in the book must necessarily be presented. After reviewing the detailed table of contents you may wish to de-emphasize certain topics, depending on the specific educational needs of your students, their prior chemical training, and their overall intellectual potential.

The sequence of topics in the text has changed somewhat from one edition to another and reflects both the authors' bias with respect to the order of presentation, as well as input from users of the previous editions. Surely there are other equally valid orders of presentation, and in general you will find that topics are treated so that their order can be easily modified to suit your own preferences. For example, in this edition we have moved the second bonding chapter forward so that it now occurs immediately following the first one, because that is where the majority of users want it. However, this chapter, and those that follow, are written so that instructors who wish to treat the more advanced bonding topics during the second semester can easily do so.

Descriptive Chemistry

An area of particular concern to teachers of chemistry in recent times has been the subject of descriptive chemistry. In this third edition we have completely overhauled our treatment of this material to provide both greater depth of coverage, as well as greater flexibility for the instructor in deciding how much descriptive chemistry he or she wishes to present.

In addition to the descriptive chemistry woven into discussions, worked examples, and end-of-chapter questions and problems, there are now three principal locations where descriptive chemistry is concentrated. The first of these is in Chapter 6 (Chemical Reactions in Aqueous Solutions), where ionic reactions and redox reactions are discussed. We consider this to be especially important for all students, and in our classes we cover this chapter thoroughly.

The second place where descriptive chemistry is concentrated is in Chapter 9 (The Periodic Table Revisited). Its location - following chapters on bonding and the properties of the states of matter - permits a meaningful discussion of some important trends in the physical and chemical properties of the elements. Although the location of the chapter places it toward the end of the first semester, the chapter is sufficiently portable that it could easily be presented later in the course.

In courses that place a heavy emphasis on "Principles," instructors may wish to limit their discussion of descriptive chemistry to Chapters 6 and 9. However, if a more detailed treatment of the chemistry of the elements is desired, additional topics can be chosen from Chapters 18 through 21. In these chapters the goal has been to describe the principal compounds and reactions of the most important elements. As you will see, there have been frequent references to familiar substances and applications of chemistry.

Scheduling

Since the division of the academic year into semesters, trimesters, quarters, etc., varies so much from school to school, we feel it would be best to simply provide you with our estimate of the number of 50-minute periods needed to cover the subject matter in the various chapters. A little arithmetic will reveal that the total number of hours probably exceeds the number that you have available. Therefore, as mentioned earlier, you will have to be somewhat selective in choosing which chapters you will cover completely.

Chapter	Periods	Chapter	Periods
1	3	13	3
2	4	14	3
3	4	15	5
4	4	16	2
5	3	17	4
6	5	18	2
7	4	19	2
8	4	20	3
9	3	21	4
10	4	22	3
11	4	23	3
12	4	24	2

Examinations

For a two-semester course we feel it is desirable to give at least three (and preferably four) full period exams per semester, plus a final exam at the end of each semester which covers the entire term's work. This permits adequate depth of coverage of the topics dealt with on each exam. You may also find it worthwhile to give a number of shorter quizzes throughout the term as well. (These would be given in recitation sections, if your course has them.)

Organization of the Instructor's Manual

This manual is divided into three principal sections. In the first, learning objectives for each chapter are presented along with the answers to all the end-of-chapter review questions. (Solutions to the review problems are available in a separate supplement - see below.) The second section provides a list of suggested lecture demonstrations. In the third section you will find reprinted all the stereo illustrations that appeared in the second edition of the text.

Supplements

Supplements available to accompany the text, in addition to this Instructor's Manual, include the following:

Student Study Guide, James E. Brady. The study guide is keyed section by section to the text. Each section includes objectives, review, self-test, and a list of new terms. Answers to the self-tests are given at the ends of the chapters. The study guide also includes a complete glossary.

Card File of Test Items, David Becker. A comprehensive set of multiple choice questions.

Solutions Manual, Theodore W. Sottery. Contains complete worked-out solutions to all the numerical problems that appear at the ends of the chapters.

Laboratory Manual, Jo A. Beran and James E. Brady. This includes 46 experiments plus special introductory sections on laboratory safety and techniques, liberally illustrated with photographs. Experiments of both a qualitative and quantitative nature are included. For each experiment there is a prelab assignment and a report sheet.

Instructor's Laboratory Manual, Jo A. Beran. A complete teachers manual for the laboratory manual, listing special equipment needed, amounts of chemicals, suggested unknowns, special precautions, and answers to the prelab assignments and questions on the report sheets.

Problem Exercises for General Chemistry, 2nd Ed., G. Gilbert Long and Forrest C. Hentz. This book contains over 1300 problems and questions designed to teach students to solve problems and answer questions in a format similar to that found on examinations.

Computer Aided Instruction for General Chemistry, William A. Butler. Tutorial help for understanding chemical principles that is provided by twenty self-contained micro-computer programs. Most programs have a "menu" of 5 to 6 parts, and each part is usually divided into 10 or 12 entries consisting of problems, questions, etc. The programs will initially be available in diskette format only, for the Pet, Apple II, and TRS 80 Microcomputers.

CHAPTER 1

INTRODUCTION

Rationale

In this chapter we have provided a rather thorough introduction to basic concepts. Even though most general chemistry students have had a chemistry course in high school, the lack of uniformity of depth of coverage and of retained knowledge makes it impossible to assume any particular level of training to serve as a point of departure. Therefore, we have assumed the student mind to be a blank slate. The coverage that you give to the topics in this chapter has to be governed by your appraisal of your students' backgrounds. If students can come away from this chapter with an understanding of its contents, they should have a firm foundation upon which to build the remainder of the course.

Objectives

After completing the chapter students should be able to:

Differentiate between law and theory; qualitative and quantitative observations; precision and accuracy; mass and weight; extensive and intensive properties; physical and chemical properties; homogeneous and heterogeneous; exothermic and endothermic.

Apply the concepts of significant figures and exponential notation in carrying out mathematical computations using the factor-label method.

Perform conversions among units in the metric/SI system.

Differentiate between substances that are either elements, compounds or mixtures.

Explain the laws of definite proportions, conservation of mass and multiple proportions by applying Dalton's Atomic Theory.

Write the chemical symbols of the common elements.

Perform computations involving density and specific gravity.

Give the number of atoms specified in a chemical formula.

Determine whether or not a chemical equation is balanced.

Associate changes in potential energy with changes in attractive and repulsive forces.

Define kinetic energy.

Perform conversions among the Fahrenheit, Celsius and Kelvin temperature scales.

Perform simple calculations relating to specific heat.

Answers to Questions

1.1 A law is a generalized statement of fact that has been obtained by experiment. Theories attempt to explain how nature behaves in the operation of laws.

1.2 Significant figures indicate the reliability of measurements.

1.3 Precision refers to how closely two measurements of the same quantity come to each other while accuracy refers to how close an experimental observation lies to the actual value.

1.4 The units gram and milliliter are more appropriate in the laboratory because of the size (small) of the quantities used.

1.5 An extensive property is one that depends on the size of the sample used. An intensive property is one that is independent of the size of the sample used.

Extensive Properties	Intensive Properties
Force (weight), length, number of atoms	Freezing point, color, specific heat

1.6 Density is the ratio of an object's mass to its volume. Specific gravity is the ratio of the object's density to the density of water. Units for density are g/ml; specific gravity has no units.

1.7

Homogeneous	Heterogeneous
sea water	smog
air	smoke
black coffee	club soda (with bubbles)
penny	ham sandwich

1.8 There are 4 phases: copper pan, iron nails, glass marbles, water.

1.9 The mass of an object does not vary from place to place; it is the same regardless of where it is measured.

1.10 Elements are the simplest forms of matter that can exist under ordinary chemical conditions. Compounds consist of two or more elements. Mixtures consist of two or more compounds that do not react chemically.

1.11 Atomic mass.

1.12 Atoms are the fundamental particles of all matter that cannot be further subdivided by ordinary chemical means. A molecule is a group of atoms bound tightly enough together that they behave as one.

1.13 (a) Fe (b) Na (c) K (d) P (e) Br (f) Ca (g) N (h) Ne (i) Mn (j) Mg

1.14 (a) silver (b) copper (c) sulfur (d) chlorine (e) aluminum (f) gold (g) chromium (h) tungsten (i) nickel (j) mercury

1.15 (a) 1 K, 2 S
(b) 2 Na, 1 C, 3 O
(c) 4 K, 1 Fe, 6 C, 6 N
(d) 3 N, 12 H, 1 P, 4 O
(e) 3 Na, 1 Ag, 4 S, 6 O

1.16 $CaSO_4 \cdot 2H_2O$

1.17 (a) During a chemical reaction, mass is neither created nor destroyed.

(b) In any pure chemical substance, the same elements are always combined in the same proportions, by mass.

(c) When two elements combine to form more than one compound, the masses of one of the elements that are combined with the same mass of the other in the various compounds are in ratios of small whole numbers.

1.18 The amu is one-twelfth of the mass of an atom of carbon-12.

1.19 (a) and (c)

1.20 Potential energy is the energy due to position of particles that either attract or repel each other. Kinetic energy is the energy due to motion; KE = $\frac{1}{2}mv^2$

1.21 An endothermic process absorbs heat (energy) as it proceeds. An exothermic process emits heat (energy) as it proceeds.

1.22 Repulsive

1.23 Heat is a form of energy, whereas the measure of the intensity of heat is known as temperature.

1.24 cal/g °C or J/g °C. For water, sp. ht. = 1.000 cal/g °C or 4.184 J/g °C. Water has the largest specific heat of any common substance. During the winter the water loses some of its heat and prevents the surrounding land masses from becoming too cold. During the summer, the water can absorb large amounts of heat and prevent surrounding land masses from becoming too hot.

CHAPTER 2

STOICHIOMETRY: CHEMICAL ARITHMETIC

Rationale

Stoichiometry is presented here so that you may begin quantitative laboratory experiments early in the course. You will find our introduction to the mole concept quite different from that found in most other texts. This approach has worked very well with students, particularly in getting them to think in terms of mole ratios when they look at subscripts in formulas and coefficients in equations. In this edition we have included a new section in this chapter that deals with molar concentration and explains how solutions of a desired molarity can be prepared. If you wish, at this time you can also cover additional topics on solution stoichiometry that are found in Section 6.9 (Pages 193-196). Appropriate homework problems on this topic can be assigned by referring to the Index to Questions and Problems on Page 205.

Objectives

Upon completion of this chapter, students should be able to:

Apply the factor-label method in stoichiometry calculations.

Interpret subscripts in formulas and coefficients in equations in terms of mole ratios.

Convert grams of a substance to moles, and vice versa.

Calculate molecular weight (formula weight) from a chemical formula.

Calculate percentage composition.

Calculate an empirical formula from percentage composition or data from a chemical analysis.

Determine the molecular formula from an empirical formula and the molecular weight of the substance.

Balance a chemical equation by inspection, and interpret the equation on a mole basis.

Use a balanced chemical equation to perform chemical calculations.

Determine the limiting reactant given the quantities of two or more reactants.

Calculate theoretical yield and percentage yield if given the actual yield.

Calculate the molarity of a solution given the amount of solute and the volume of the solution.

Explain how a solution having a desired molarity should be prepared.

Use molarity as a conversion factor in calculating (a) the amount of solute in a known volume of a solution of a given molarity, (b) the volume of a solution of known molarity that is required to contain a specified amount of solute, and (c) the amount of solute needed to prepare a given volume of solution having a specified molarity.

Answers to Questions

2.1 All three represent a set number of objects; the mole is 6.02×10^{23} things, the dozen is 12 things and the gross means 144 things.

2.2 Formula weights are preferred in compounds where formula units are used to describe a neutral aggregate of ions, e.g. NaCl.

2.3 Structural formulas represent the way the atoms in a molecule are linked together. Molecular formulas specify the actual numbers of each kind of atom found in a molecule. Empirical formulas give the relative numbers of atoms of each element present in a molecule.

2.4 NH_4SO_4, Fe_2O_3, $AlCl_3$, CH, $C_3H_8O_3$, CH_2O, Hg_2SO_4

2.5 molecular formula is $C_2H_6O_2$; empirical formula is CH_3O.

2.6 A simplest formula is calculated from experimentally observed or measured data obtained by a chemical analysis of the compound. Webster's defines *empirical* as "Pertaining to, or founded upon, experiment or experience."

2.7 The law of conservation of mass.

2.8 Coefficients are: (a) 1,2,1,1 (b) 6,2,2,3 (c) 8,3,4,9 (d) 1,1,2 (e) 2,1,2,1

2.9 Coefficients are: (a) 2,3,1,6 (b) 2,1,1,2,2 (c) 3,2,1,6 (d) 1,3,1,3,3 (e) 1,1,1,1,2

2.10 Coefficients are:
(a) 2, 13, 8, 10 (1, 13/2, 4, 5)
(b) 2, 17, 14, 6 (1, 17/2, 7, 3)
(c) 1, 6, 4
(d) 4, 11, 2, 8 (2, 11/2, 1, 4)
(e) 4, 5, 4, 6 (2, 5/2, 2, 3)

2.11 That reactant that is completely consumed before the remaining reactants are used up. First calculate the number of moles of each reactant present. Then compare their ratios with that for the balanced chemical equation. From this deduce which reactant will be depleted first.

2.12 The theoretical yield is the maximum amount of product that could be produced from a given quantity of reactant regardless of any other products. The percent yield is a comparison of the theoretical yield to the yield that is actually obtained; it is the efficiency of the reaction. The actual yield is the amount of product that you actually obtain in a given experiment when the reaction is carried out.

2.13 $$\text{molar concentration} = \frac{\text{number of moles of solute}}{\text{total volume of the solution in liters}}$$

2.14 Place 180 g of $C_6H_{12}O_6$ in a 1.00-liter volumetric flask. Dissolve the sugar in some water; then dilute to a total volume of 1.00 liter.

2.15 0.20 M Na_3PO_4 means

$$\frac{0.20\ \text{mol}\ Na_3PO_4}{1000\ \text{ml solution}} \quad \text{and} \quad \frac{1000\ \text{ml solution}}{0.20\ \text{mol}\ Na_3PO_4}$$

CHAPTER 3

ATOMIC STRUCTURE AND THE PERIODIC TABLE

Rationale

This chapter introduces students to the detailed structure of matter for the first time. In this chapter one of our goals has been to show students how theory must relate to experimentally observable phenomena. For example, Sections 3.1 through 3.7 examine, historically, the development of our current picture of the atom. The structure of the periodic table and the periodic law are discussed because the theory relating to electronic structure must explain the structure of the periodic table. The Bohr theory is examined because it was successful at accounting for the line spectrum of hydrogen. We also see here how theories must be abandoned when they fail. Another reason for discussing the Bohr theory is to show how a theory relating to atomic structure is checked against "reality," i.e. by deriving an equation from the theory that matches with one derived empirically (Page 80).

The remainder of the chapter deals with the modern view of atomic structure and electron configuration, as well as the variation of some properties with variations in electronic structure. A more qualitative treatment of this chapter might include a condensation of Sections 3.1 to 3.7, with the elimination of numerical calculations. You may also wish to limit the discussion of the Bohr theory to simply an examination of the Bohr model of the atom and its reason for failure.

Objectives

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

Relate the contributions of the following scientists toward our picture of the atom: Faraday, Thomson, Millikan, Becquerel, Rutherford, Moseley, Chadwick, Bohr, Planck, Einstein, DeBroglie and Schrödinger.

Define isotopes and compute the average atomic weight of an element given the relative abundances and actual masses of its isotopes.

Perform calculations relating wavelength and frequency for electromagnetic radiation.

Relate the contributions of Mendeleev and Meyer to the periodic classification of the elements.

For the periodic table, identify: groups and periods; representative, transition and inner transition elements; rare earth elements; alkali metals; alkaline earth metals; halogens; noble gases; metals; nonmetals; metalloids.

Explain the difference between a continuous spectrum and a line spectrum.

Explain how an atomic spectrum is obtained experimentally.

Give the relationship between energy and frequency for a photon.

Describe the Bohr model of the atom.

Explain how a diffraction pattern is formed.

Identify the four quantum numbers and specify their allowed values.

State the Pauli exclusion principle and Hund's rule.

Use the periodic table to predict the electron configurations of the elements.

Use the location of a representative element in the periodic table to write the electron configuration of its outer shell.

Describe the shapes of s and p atomic orbitals.

Describe and explain the periodic trends in the following properties: atomic size, ionic size, ionization energy, and electron affinity.

Answers to Questions

3.1 Cathode rays travel in straight lines, cast shadows, turn pinwheels, heat metal foil, and can be bent by an electric or magnetic field.

3.2 A coulomb is the amount of electric charge moving past a given point in a wire when an electric current of one ampere flows for one second.

3.3 See Figure 3.1 on Page 62.

3.4 Electrons (cathode rays) passing through the gas knock electrons off neutral molecules, which leaves positively charged particles behind.

3.5 Canal rays are rays of positive particles that pass through a perforated cathode.

3.6 Hydrogen is the lightest of all elements and has the largest e/m ratio for any positive ion.

3.7 Alpha, α, He^{2+} particles; beta, β, composed of electrons; gamma, γ, rays, high energy light waves.

3.8 The path of the particles is guided by a magnetic field which deflects them into a circular path. The degree of curvature of these paths is determined by the charge-to-mass ratio of the ions.

3.9 Because some of the alpha particles were strongly deflected by the thin foil.

3.10 See Table 3.1, Page 68.

3.11

	protons	neutrons	electrons
$^{132}_{55}Cs$	55	77	55
$^{115}_{48}Cd^{2+}$	48	67	46
$^{194}_{81}Tl$	81	113	81
$^{105}_{47}Ag^{1+}$	47	58	46
$^{78}_{34}Se^{2-}$	34	44	36

3.12

	protons	neutrons	electrons
$^{131}_{56}Ba$	56	75	56
$^{109}_{48}Cd^{2+}$	48	61	46
$^{36}_{17}Cl^{-}$	17	19	18