KELLER PLAN for SELF-PACED STUDY USING MASTERTON AND SLOWINSKI'S CHEMICAL PRINCIPLES

JOSEPH L. CLOUSER

W. B. SAUNDERS COMPANY

KELLER PLAN

FOR

SELF-PACED STUDY

USING

MASTERTON AND SLOWINSKI'S

CHEMICAL PRINCIPLES

JOSEPH L. CLOUSER

Associate Professor of Chemistry
William Rainey Harper College

SAUNDERS GOLDEN SERIES

1974

W. B. SAUNDERS COMPANY • Philadelphia • London • Toronto

W. B. Saunders Company: West Washington Square
Philadelphia, PA 19105

12 Dyott Street
London, WC1A 1DB

833 Oxford Street
Toronto, Ontario M8Z 5T9, Canada

Keller Plan for Self-Paced Study Using
Masterton and Slowinski's *Chemical Principles* ISBN 0-7216-2610-6

Last digit is the print number: 9 8 7 6 5 4 3 2 1

To The Student

Introduction The format of this course will vary somewhat from the lecture-discussion approach. The system that will be used is designed to allow you to proceed, within limits, at your own pace, mastering principles in one unit before proceeding to the next unit. The key to success in this course is to become actively involved with the materials of the course. You must become a participant instead of passively receiving the information that is conveyed in a lecture. This is not a "spectator" course but one that requires your active involvement.

The System The course is divided into units that correspond to chapters in your text, *Chemical Principles,* third edition, by Masterton and Slowinski. This guide lists the performance objectives of each unit and activities that will help you to reach these objectives. The activities include readings in the text, suggested problems and questions with detailed solutions, and assignments in the *Student's Guide to Chemical Principles* by R. Boyington.

How To Use The System This guide will detail the strategy you should pursue to master chemical principles and concepts. The following procedures are suggested for each unit.

1. Read the goals and objectives for the unit as listed in this guide.
2. Read "Questions to Guide Your Study" in Boyington.
3. With the objectives and study questions in mind and at hand for ready reference, read the assigned text material.
4. Take the self-test in Boyington.
5. Now work the assigned problems and questions.
6. After making your best effort on the problems, consult the detailed solutions. You need not solve a problem exactly as shown in the solutions, but you should feel that you thoroughly understand your method of solution. Be sure that you use the proper units throughout your solution.
7. After reviewing the detailed solutions to the assigned problems, see the instructor for help with any of the problems that you do not understand. For the problems that gave you trouble, work the

corresponding problem for which the answer is provided in the text. Notice that the problems at the end of each chapter are "paired." For example, on p. 22, Problem 1.20, which is answered in Appendix 5, corresponds to the assigned problem, 1.7, located directly across from it. Both problems illustrate the same principle.

8. Look over the list of references in Boyington and select those that seem most valuable to you. Consult these at this time.
9. Review the goals and objectives listed in this guide. If you do not feel confident of your comprehension of the unit, consult the instructor or tutor for additional help to clear up difficult items or concepts.
10. Take the unit quiz. Be sure to discuss any errors with the proctor.

A Few Final Comments Set work schedules and deadlines for yourself. You will find that some units will not require as much time and effort as others. Since you have the freedom to set your pace, allow additional time for the units that you find more difficult and move at a more rapid rate for units that are, for you, easier. Your instructor will announce a grading policy, probably set some guidelines for completion of certain units, indicate the degree of mastery that is required on unit quizzes before you can proceed to the next unit, and set requirements for lectures if, in fact, any lectures are scheduled. **Responsibility for success in this course rests with you.** You are graded on your individual performance and not on a class curve. If you develop self-discipline, don't procrastinate, and budget your time, you can employ this system for successful mastery of the principles covered by this course.

Acknowledgments I would like to acknowledge the assistance provided by members of the Harper Chemistry Department. I am especially indebted to Professor William L. Masterton of the University of Connecticut for his thoughtful suggestions, review, and much of the rewriting of the manuscript. I appreciate the encouragement and support of the W. B. Saunders Company, particularly that of Mrs. Joan Garbutt, Editorial Assistant. Finally, I would like to thank Cathy for her work in typing the manuscript.

CONTENTS

UNIT 1

Basic Concepts

The goal of this unit is to consider some of the basic properties of chemical systems and learn how they are measured.

INSTRUCTIONAL OBJECTIVES

1. Define the fundamental units of length, volume, mass, and temperature and describe an instrument used to measure each of the quantities
2. Distinguish between mass and weight
3. Distinguish between heat and temperature and describe the experimental basis for their measurement
4. Express the uncertainty of a measurement by use of significant figures
5. Distinguish between intensive and extensive properties
6. Describe methods (and the principles upon which these methods are based) for the separation and identification of pure substances
7. Recognize the basic distinction between elements and compounds
8. Become familiar with the names and symbols of some of the more common elements (see Figure 1.12 and Table 1.4)
9. Distinguish between abundance and availability of elements

SKILLS TO BE DEVELOPED

1. Express numbers in exponential notation (Appendix 4, Prob. 1, p. 670)
2. Perform multiplication, division, addition, and subtraction operations using exponential notation (Appendix 4, Prob. 1.2)
3. Convert units within the metric system and between the metric and English system through the use of the *conversion factor* method (Probs. 1.9, 1.28)

4. Apply the rules of significant figures in calculations (Prob. 1.8)
5. Employ mass, volume, and density relationships in calculations (Prob. 1.7)

SUGGESTED ASSIGNMENT I

Text: Chapter 1, pages 1-21, Appendices 1 and 4

Boyington: Preface and pages 1-7

Problems: Text pages 22-24: Numbers 1.2a & b, 1.3, 1.4, 1.5a, 1.7, 1.8, 1.9, 1.12, 1.13, 1.28. Text pages 670-671: Numbers 1, 2

If you have difficulty with the above problems, consult the Study Aids listed on page V and VI of the Preface in Boyington.

Solutions to Assigned Problems Set I

1.2

a) Mass is a measure of the amount of matter in a substance and is invariable. Weight is a measure of the attractive force of gravity and varies at different locations.

b) The principle of the common mercury-in-glass thermometer is based on the expansion of mercury. When the mercury is heated the increase in volume of the mercury contained in a fine capillary is detected. The scale used (Celsius) is based on the freezing and normal boiling points of water, assigned values of 0° and 100° respectively. The distance between these two points is divided into 100 equal parts, and designated as a degree Celsius. A reading of 50°C corresponds to a mercury level 50 per cent of the way from the 0° mark to the 100° mark.

1.3

a) Use an analytical balance to measure the mass of a sample of mercury delivered from a pipet. Divide the mass of the sample by the volume delivered.

b) Fill the flask with water delivered from a buret.

c) Purify the methanol by fractional distillation, then measure the boiling point of the purified methanol.

d) Purify the DDT by fractional crystallization, then measure the melting point of the purified DDT.

1.4

a) distillation b) fractional crystallization c) fractional distillation d) VPC or distillation after liquefaction.

1.5

a) An intensive property is independent of the amount of sample while an extensive property is directly proportional to the amount of sample.

1.7

The value 85.279 g is discarded since it is not consistent with the first three measurements. Mass of water = (mass of beaker + water) − mass of beaker.

Trial 1: 84.136 g − 64.232 g = 19.904 g
2: 84.151 g − 64.232 g = 19.919 g
3: 84.141 g − 64.232 g = 19.909 g

$$\text{Average mass} = \frac{19.904\text{ g} + 19.919\text{ g} + 19.909\text{ g}}{3} = 19.911\text{ g}$$

$$\text{Volume} = \frac{19.911\text{ g}}{0.9970\text{ g/ml}} = 19.97\text{ ml}$$

(Note: Only four significant figures in the answer. Refer to the text page 7 if you have more or less than four significant figures)

1.8

a) First convert all masses to units of grams, then apply the rule of absolute uncertainty to the addition as indicated in the text pages 7 and 8

$$500\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} = 0.500\text{ g}$$

$$20\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} = 0.020\text{ g}$$

since each measurement is accurate to 0.001 g, addition is 10.000 g + 1.000 g + 0.500 g + 0.020 g = 11.520 g

b) Solve as in 1.8 (a) by first converting to kg and applying the rule of absolute uncertainty to the addition 20.00 kg + 0.012 kg + .209 kg = 20.22 kg

1.9

a) $1.00 \text{ g/ml} \times \dfrac{1000 \text{ ml}}{1 \text{ liter}} = 1.00 \times 10^3 \text{ g/liter}$

b) $1.00 \text{ g/ml} \times \dfrac{1 \text{ lb}}{453.6 \text{ g}} \times \dfrac{28.32 \times 10^3 \text{ ml}}{1 \text{ ft}^3} = 62.4 \text{ lb/ft}^3$

c) $1.00 \text{ pint} \times \dfrac{1 \text{ liter}}{2.114 \text{ pt}} \times \dfrac{1.00 \times 10^3 \text{ g}}{1 \text{ liter}} \times \dfrac{16 \text{ oz}}{453.6 \text{ g}} = 16.7 \text{ oz}$

Why are there three significant figures in each of the above answers?

1.12

a) Employ equation 1.2 in the text in the form $°C = \dfrac{°F - 32}{1.8}$ to calculate the body temperature in °C: $\dfrac{98.6 - 32}{1.8} = 37.0°C$. Recognize the linear relationship $°F = a°C + b$ applies and evaluate the constants $b = 32$, $a = \dfrac{100 - 32}{37 - 0} = 1.84$ then $°F = 1.84°C + 32$

b) Employ the realtionship derived in (a), $°F = 1.84\ (100) + 32 = 216°$

1.13

a) $65 \text{ g tartaric acid} \times \dfrac{100 \text{ g water}}{98 \text{ g tartaric acid}} = 66$ g water dissolves all of the sample (up to $71 \text{ g} \times 0.66 = 47$ g of succinic acid would dissolve in 66 g of water, far more than is present in the sample).

b) grams tartaric acid remaining $= 66 \text{ g water} \times \dfrac{18 \text{ g T.A.}}{100 \text{ g water}} = 12$ g T.A.

$65 \text{ g} - 12 \text{ g} = 53$ g T.A. crystallized

grams succinic acid remaining $= 66 \text{ g water} \times \dfrac{7 \text{ g S.A.}}{100 \text{ g water}} = 5$ g S.A.

$12 \text{ g} - 5 \text{ g} = 7$ g S.A. crystallized

(Note that the % of T.A. in the solid has increased)

c) Stop cooling at a higher temperature (50°C would be O.K.) or use more water so that all the succinic acid remains in solution at 20°C

1.28*

$$\text{Vol} = 100 \text{ barrels} \times 31.5 \text{ gal/barrel} \times \frac{4 \text{ qt}}{\text{gal}} \times \frac{57.75 \text{ in}^3}{1 \text{ qt}} \times \frac{1 \text{ ft}^3}{1728 \text{ in}^3} \times \frac{1 \text{ mile}^3}{(5280 \text{ ft})^3} = 2.87 \times 10^{-9} \text{ mile}^3$$

$$h = 1.000 \times 10^3 \text{ A} \times \frac{1 \times 10^{-8} \text{ cm}}{1\text{A}} \times \frac{1 \text{ in}}{2.54 \text{ cm}} \times \frac{1 \text{ ft}}{12 \text{ in}} \times \frac{1 \text{ mile}}{5280 \text{ ft}} = 6.21 \times 10^{-11} \text{ mile}$$

$$\text{area} = \frac{2.87 \times 10^{-9} \text{ mile}^3}{6.211 \times 10^{-11} \text{ mile}} = 46.2 \text{ mile}^2$$

Answers are provided to the exercises that are assigned in Appendix 4

*Note that problems marked with an asterisk tend to be rather difficult.

UNIT 2

Atoms, Molecules, and Ions

The goal of this unit is to examine the fundamental units of matter and relate these units to the properties of substances containing them.

INSTRUCTIONAL OBJECTIVES

1. State the postulates of Dalton's atomic theory
2. Relate the postulates of Dalton's theory to the Laws of Conservation of Mass, Constant Composition, and Multiple Proportions
3. Summarize the experimental work of Thomson, Milliken, and Rutherford and explain how they contributed to a model of the atom
4. List the properties (charge, mass, diameter) of the electron, proton, and neutron
5. Illustrate the relationship of atomic number and mass number to the number of protons, neutrons, and electrons (Prob. 2.2)
6. Recognize the existence of isotopes and the nuclear differences between isotopes of an element
7. Distinguish between atom, molecule, and ion
8. Explain the basis of the atomic weight scale
9. Define atomic weight, gram atomic weight, molecular weight and gram molecular weight
10. Define Avogadro's number and explain what it means
11. Describe the basic principles of the mass spectrometer and how it can be used to determine atomic weights (Prob. 2.8)

SKILLS TO BE DEVELOPED

1. Calculate the average atomic weight of an element given the percentage and mass of each component isotope. Conversely, given the

average atomic weight, calculate the percentage abundance of each of two isotopes in the element. (Probs. 2.9, 2.10)

2. Calculate the mass in grams of a single atom given the atomic weight of the element; the mass in grams of a single molecule given the GMW of the substance

3. Make conversions such as:
number of GAW's $\longrightarrow$ number of atoms (Prob. 2.11a)
number of grams $\longrightarrow$ number of GAW's (Prob. 2.11c)
number of grams $\longrightarrow$ number of GMW's (Prob. 2.12b)
number of grams $\longrightarrow$ number of molecules (Prob. 2.12c)

4. Illustrate the Law of Multiple Proportions given appropriate data (Prob. 2.5)

SUGGESTED ASSIGNMENT II

Text: Chapter 2, pages 25 - 44

Boyington: Pages 9 - 16

Problems: Text pages 44 - 46: Numbers 2.2 - 2.6, 2.8 - 2.12, 2.14, 2.36

Solutions to Assigned Problems Set II

2.2

		$^{23}_{11}Na$	$^{30}_{14}Si$	
a)	no. of protons	11	14	(atomic no. = no. of protons)
b)	no. of neutrons	12	16	(no. of neutrons = mass no. − no. of protons
c)	no. of electrons	11	14	(no. of electrons = no. of protons in neutral atom)
d)	no. of protons	11	14	
	no. of neutrons	12	16	
	no. of electrons in +1 ion	10	13	(see text page 32)

2.3

a) There are no molecules of NaCl

b) They differ in mass due to the existence of isotopes

c) The $^{1}_{1}H$ nucleus contains no neutrons

d) The mass of *Avogadro's number* of C-12 atoms is 12 g

2.4

a) 12 (no. of electrons = atomic number in a neutral atom)

b) $$\frac{6 \text{ electrons}}{\text{C atom}} \times \frac{6.02 \times 10^{23} \text{ C atoms}}{\text{GAW}} = 3.61 \times 10^{24} \frac{\text{electrons}}{\text{GAW}}$$

c) $1.00 \text{ cc} \times \frac{1.00 \text{ g}}{1 \text{ cc } H_2O} \times \frac{1 \text{ GMW}}{18 \text{ g}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{\text{GMW}} \times \frac{10 \text{ electrons}}{1 \text{ molecule}} = 3.34 \times 10^{23} \text{ electrons}$

d) $\frac{1 \text{ electron}}{1.60 \times 10^{-19} \text{ coulomb}} = 6.25 \times 10^{18} \frac{\text{electrons}}{\text{coulomb}}$

2.5

Assume 100 g compound, therefore:

Compound I: 44.0 g Fe, 56.0 g Cl; $\frac{44.0 \text{ g Fe}}{56.0 \text{ g Cl}} = 0.786 \frac{\text{g Fe}}{\text{g Cl}}$

Compound II: 34.4 g Fe, 65.6 g Cl; $\frac{34.4 \text{ g Fe}}{65.6 \text{ g Cl}} = 0.524 \frac{\text{g Fe}}{\text{g Cl}}$

Ratio: $\frac{0.786 \frac{\text{g Fe}}{\text{g Cl}}}{0.524 \frac{\text{g Fe}}{\text{g Cl}}} = 1.5:1 \text{ or } 3:2$

The mass of Fe that combines with one gram of Cl in the two different compounds is in the ratio of the small whole numbers 3:2

2.6

$\frac{2.02 \text{ g H}}{2.02 \text{ g H} + 12.0 \text{ g C}} \times 100 = 14.4\% \text{ H}; \frac{12.0 \text{ g C}}{2.0 \text{ g H} + 12.0 \text{ g C}} = 85.6\%$

2.8

$12.0 \times \frac{704 \text{ v}}{273 \text{ v}} = 30.9$

2.9

$10.02\,(0.1883) + 11.01\,(0.8117) = 10.82$