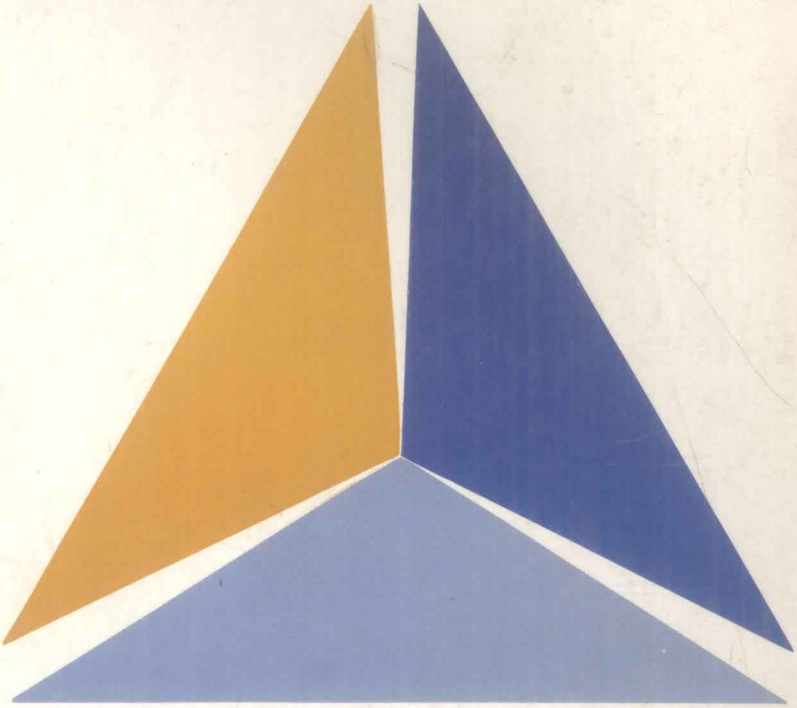
JAMES C. HILL



STUDENT'S GUIDE TO BROWN AND LeMAY
chemistry
the central science

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STUDENT'S GUIDE TO BROWN AND LEMAY

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introduction

This Student's Guide was written to help you understand the introductory chemistry material presented by Theodore L. Brown and H. Eugene LeMay in Chemistry: The Central Science, hereafter referred to as the text. The order of chapters in the Student's Guide parallels those in the text.

In the Student's Guide each chapter is divided into four sections: OVERVIEW OF THE CHAPTER, TOPIC SUMMARIES AND EXERCISES, KEY TERMS, and SELF-STUDY QUIZ. A brief description of each section and suggestions of how to use it follow.

OVERVIEW OF THE CHAPTER: The subject matter in each chapter of the text can be more easily learned by studying small segments of interrelated material. To help you identify appropriate segments, the major topics in a chapter of the text are listed in OVERVIEW OF THE CHAPTER. As a further help, included with each topic are section numbers of related material in the chapter, a list of objectives that will guide you in your study, and often references to review matter in previous chapters or in the Appendix. The objectives are called learning goals in the text.

Before you read a chapter in the text, you should read OVERVIEW OF THE CHAPTER so that you will have familiarity with the major topics, with the objectives, and with material you may want to review. Use this section in the Student's Guide as a roadmap of the chapter.

TOPIC SUMMARIES AND EXERCISES: Each topic in OVERVIEW OF THE CHAPTER is summarized in this section. Further explanations of key material are also included. After each topic summary are exercises with detailed solutions, similar in style to the sample exercises in the text. The solutions to the exercises often include further explanations of important or difficult material.

After having read the entire chapter in the text, you should study the sections that relate to the first topic. Review the suggested material listed in OVERVIEW OF THE CHAPTER if you have difficulty with the topic. Use the objectives to help you identify the more important areas for studying. After you are satisfied that you know the material associated with the first topic, continue by studying each topic sequentially.

KEY TERMS: Chemistry has a language with which you must become proficient. Listed at the end of each chapter in the text are key terms with their definitions. Without an understanding of these terms, you will have difficulty in mastering chemistry. In this section of the Student's Guide, statements with missing key terms are provided as exercises. After reviewing the key terms, complete the statements by supplying the appropriate terms.

SELF-STUDY QUIZ: Before an exam or after studying a chapter, you can check your understanding of the material by answering the questions provided in this section. The solutions to the questions are provided at the back of the Student's Guide. Although not as detailed as the solutions to the exercises, the solutions to the questions are elaborated sufficiently so that you can see how the problems are solved.

Although every attempt was made to produce an error-free book, some mistakes may have been overlooked. If you notice any errors, I would appreciate hearing from you. Also, suggestions for improving the Student's Guide or criticisms of it will be welcome.

I wish to thank Professor Theodore L. Brown and H. Eugene LeMay for their detailed reviews of the Student's Guide, and Professor Paul Noble for his helpful comments. I am indebted to Sharon Hoover and Sharon Melia for their typing of the manuscript from which this book was printed. Most of all, I appreciate the understanding, love, and encouragement given by my wife, Jan, and my children, Jason and Jeanina, during my writing of this book.

James C. Hill
Sacramento, California

contents

introduction: some basic concepts

1

OVERVIEW OF THE CHAPTER

Metric System (1.3)
Review: Concept of fraction; exponential notation (see Appendix)
Objective: You should be able to use the metric system, to list the basic metric units and common prefixes.

SI (International System) Units (1.3)
Objective: You should be able to use the metric system, to list the basic metric units and common prefixes. The SI units are an improved form of the metric unit system.

Significant Figures (1.4)
Review: Exponential notation (see Appendix)
Objective: You should be able to determine the number of significant figures in a derived quantity.

Dimensional Analysis (1.5)
Review: Concepts of fraction and ratio
Objective: You should be able to interconvert metric and English system measurements using dimensional analysis.

Temperature (1.3)
Objective: You should be able to interconvert degrees Fahrenheit, Celsius, and Kelvin.

Density (1.3)

Review: Concept of ratio

Objective: You should be able to perform calculations involving density.

TOPIC SUMMARIES AND EXERCISES

Metric System

The metric system was developed during the French Revolution. It is the measuring system used throughout the world today. An extended and improved version of the metric system, the Systeme Internationale d'Unites (SI) or the International System, has been used since 1960 and will be discussed in the next section.

EXERCISE 1

Complete the following table which contains some of the metric units commonly used by chemists.

Physical Quantity	Unit Name	Unit Symbol
Length		
	Liter	
	Cubic centimeter	
		g
Time	Second	

Solution:

Physical Quantity	Unit Name	Unit Symbol
Length	Meter	m
Volume	Liter	l.
Volume	Cubic centimeter	cm^3
Mass	Gram	g
Time	Second	sec

The unit liter is used when measuring the volume of a solution and the unit cubic centimeter is used when measuring the volume of a solid.

EXERCISE 2

Smaller or larger units of measurement than a basic unit are indicated by a prefix before the name of the basic unit. For example, 1/100 meter can be expressed as one centimeter. The prefix centi- always means a hundreth part of something. The commonly used prefixes in chemistry are presented in the table at the top of page 3.

Prefix	Fraction or Multiple of Basic Unit	Abbreviation
Centi	10^{-2} (1/100)	c
Milli	10^{-3} (1/1,000)	m
Micro	10^{-6} (1/1,000,000)	μ
Nano	10^{-9} (1/1,000,000,000)	n
Pico	10^{-12} (1/1,000,000,000,000)	p
Kilo	10^{3} (1,000)	k
Mega	10^{6} (1,000,000)	M
Giga	10^{12} (1,000,000,000,000)	G

Which quantity of each pair is the largest: (a) 1 nm or 1 micrometer; (b) 1 picogram or 1 cg; (c) 1 megagram or 1 milligram?

Solution:

(a) 1 nm = 10^{-9} m = 10^{-9} meter
1 micrometer = 1 μm = 10^{-6} meter

A micrometer is larger than a nanometer because 10^{-6} is larger than 10^{-9}.

(b) 1 picogram = 1 pg = 10^{-12} gram
1 cg = 10^{-2}g = 10^{-2} gram

A centigram is larger than a picogram because 10^{-2} is larger than 10^{-12}.

(c) 1 megagram = 1 Mg = 10^{6} gram
1 mg = 10^{-3}g = 10^{-3} gram

A megagram is larger than a milligram because 10^{6} is larger than 10^{-3}.

EXERCISE 3

Two alternate units for the unit of length, which are not encountered in Chapter 1 but are extensively used in chemistry, are the millimicron and the Angstrom.

Unit	Symbol	Quantity in Basic Unit
Millimicron	mμ	10^{-9} m
Angstrom	Å	10^{-10} m

A millimicron is equivalent to 10^{-9} m. What is the metric name which is associated with 10^{-9} m?

Solution: The nanometer because 1 nm = 10^{-9} m.

EXERCISE 4

The English system of basic units is not as simple as the metric unit system. A few of the English-metric unit equivalences are

Physical Quantity	English Unit Symbol	Metric Unit Symbol	Equivalents
Mass	lb	g	1 g = 0.002205 lb
Length	ft	m	1 m = 3.272 ft
Volume	qt	L	1 L = 1.06 qt

What is the advantage of the metric system compared to the English system?

Solution: All larger and smaller quantities of the basic unit in the metric system are in fractions of 10's (e.g. 1/10, 1/100, 1/1000) or multiple of 10's (e.g. 10, or 100) of the basic unit. This is not true in the English system. Smaller or larger quantities of the basic unit in the English system are new defined units. For example 4,000 quarts equals 1,000 gallons, not 4 kiloquarts. Many more conversion factors are required in the English unit system than in the metric unit system.

SI Units

The International System (SI) of units was adopted by the General Conference of Weights and Measures in 1960. Some of the basic and derived SI units are presented in the following table. Some of these units will not be encountered until later chapters.

Physical Quantity	Name of Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	sec
Energy	Joule	J
Volume	Cubic meter	m^3
Pressure	Newton per square meter	N/m^2

The equivalence relations between the above SI units and the metric units are presented in the following table:

Physical Quantity	Metric Unit	SI Unit	Equivalence
Length	Meter	Meter	Same
Mass	Gram	Kilogram	10^3 g = 1 kg
Time	Second	Second	Same
Energy	Calorie	Joule	1 cal = 4.184 J
Volume	Liter	Cubic meter	10^3 l = 1 m^3
Pressure	Torr (mm Hg)	Newton per square meter	1 torr = 133.3222 N/m^3

EXERCISE 5

One milligram is one ____________________ of the basic SI mass unit, not one thousandth as is indicated by the prefix milli.

Solution: Millionth. We do not use the quantity 1 microkilogram, 1 μkg, for the quantity one milligram because compound (more than one) prefixes should not be used. This is a flaw in thè SI system. The name for the basic unit of mass involves a prefix, kilo.

Significant Figures

There are two types of numbers used in chemistry. Measured values are numbers determined from a measurement or the result of a calculation using numbers determined from measurements. Defined values are set exactly by definition.

EXERCISE 6

Except for numbers determined by counting observable objects, measured values are never exact. Measured values are reported so that the last digit of the number is uncertain to ±1 unless otherwise stated. What is the uncertain digit in each of the following numbers: (a) 10.03 kg; (b) 5 apples; (c) 5.02± 0.02 m?

Solution: (a) The 3 in 10.03 kg is uncertain to ±1. (b) This is an exact measured value. There is no uncertain digit. (c) The 2 in 5.02 m is uncertain to ±2.

EXERCISE 7

The precision of a measurement is indicated by the number of significant figures. The number of significant figures is determined by counting the total number of certain digits plus the first uncertain digit. Zeroes used to indicate a decimal point are not significant. How many significant figures do each of these numbers have: (a) 225; (b) 10004; (c) 0.0025; (d) 1.0025; (e) 0.002500; (f) 14100; (g) 14100. ; (h) 14100.0?

Solution: (a) Three. (b) Five. The middle zeroes do not fix the decimal point. (c) Two. The zeroes immediately to the right of a decimal, which has only zeroes to its left, fix the decimal point. (d) Five. The zeroes to the right of the decimal do not fix the decimal. The 1. fixes the decimal. (e) Four. The last two zeroes do not fix the decimal and are significant. (f) Three, four, or five. The zeroes may or may not be significant since the decimal is not indicated. Without more information, we usually assume the zeroes are not significant. (g) Five. The zeroes do not fix the decimal because the decimal is explicitly stated. (h) Six. The placement of the zero after 14100. indicates it is significant.

EXERCISE 8

The ambiguity in determining the number of significant figures in a number such as 14100 is avoided by using exponential notation. The decimal is placed after the first digit in a number; all the significant figures are written and multiplied by 10 raised to a number which fixes the decimal point. For example 1420 is written as 1.42×10^3 if it has three significant figures, or as 1.420×10^3 if the last zero in 1420 is significant. Write the numbers in Exercise 7 using exponential notation.

Solution:

(a) $225 = 2.25 \times 10^2 = 2.25 \times 10^2$. The decimal is moved two digits to the left; thus the power of 10 is 2.

(b) $10004 = 1.0004 \times 10^4 = 1.0004 \times 10^4$. The power of 10 is 4 because the decimal is moved four places to the left.

(c) $0.0025 = 0002.5 \times 10 = 2.5 \times 10^{-3}$. The power of 10 is -3 because the decimal is moved three places to the right.

(d) 1.0025. We do not write 1.0025×10^0.

(e) $0.002500 = 0002.500 \times 10^{-3} = 2.500 \times 10^{-3}$. The zeroes after the 25 are written because they do not define the decimal.

(f) $14100 = 1.41 \times 10^4$ if the zeroes in 14100 are not significant.

(g) $14100. = 1.4100 \times 10^4 = 1.4100 \times 10^4$. The zeroes are significant.

(h) $14100.0 = 1.41000 \times 10^4 = 1.41000 \times 10^4$.

EXERCISE 9

The final answer in a calculation using the modern calculator usually has more significant figures than any of the numbers used in the calculation. The final answer obtained when adding or subtracting must be rounded off so that the final answer has the same number of significant figures to the right of the decimal as in the number in the calculation which has the fewest digits to the right of the decimal. Round off the final answer in each calculation.

(a) $12.25 + 1.32 + 1.2 = 14.770$

(b) $13.7325 - 14.21 = -0.4775$

Solution: (a) The number with the fewest digits to the right of the decimal is 1.2; thus the final answer can have only one digit to the right of the decimal: 14.770 is rounded off to 14.8. (b) The number with the fewest digits to the right of the decimal is 14.21; thus, the final answer can have only two digits to the right of the decimal: -0.4775 is rounded off to -0.48.

EXERCISE 10

The rule for rounding off the final calculated answer when multiplying or dividing numbers is to round off the final calculated answer so that it has the same number of significant figures as the least precise number (fewest significant figures) in the calculation. Round off the final answer in each calculation.

(a) $(1.256)(2.42) = 3.03952$

(b) $\dfrac{16.231}{2.20750} = 7.352661$

(c) $\dfrac{(1.1)(2.62)(13.5278)}{2.650} = 14.712119$

Solution: (a) The least precise number in the calculation is 2.42 with three significant figures. The final answer must be rounded off to three significant figures: 3.04. (b) The least precise number in the calculation is 16.231 with five significant figures. The final answer must be rounded off to five significant figures: 7.3527. (c) The least precise number in the calculation is 1.1 with two significant figures. The final answer must be rounded off to two significant figures: 14.

Dimensional Analysis

Most calculation problems require interconverting units. One approach to solving unit conversion problems is called dimensional analysis, or the unit factoring method, or the factor-label method. This method requires that the units be included with all numbers in the calculation. The conversion between units in the problem is done by a conversion factor that equals 1. That is, the conversion factor is a ratio of units equal to 1. This method is explored in detail in the next exercise.

EXERCISE 11

Make the conversion of 10 m to centimeters.

Solution: The required operations for converting 10 m to centimeters are as follows:

1. State the general relation required to convert units.

   ? cm = (10 m)(unit conversion factor that converts m to cm).

2. Change all numbers to exponential notation.

   ? cm = (1×10^{1} m)(unit conversion factor).

3. Determine the unit conversion factor that converts meter to centimeter. First state the equivalence between units: 1 cm = 10^{-2} m. Using this equivalence relation, determine the ratio of units that converts meter to centimeter. To convert meter to centimeter, we multiply m by cm/m so that the unit meter cancels. If we divide the equivalence relation 1 cm = 10^{-2} m on both sides of the equal sign by 10^{-2} m we obtain the required ratio of units.

   $$\frac{1 \text{ cm}}{10^{-2} \text{ m}} = \frac{10^{-2} \text{ m}}{10^{-2} \text{ m}} = 1$$

   The ratio 1 cm/10^{-2} m is a unit conversion factor since it equals one.
4. Substitute 1 cm/10^{-2} m for the unit conversion factor in the equation in step 2.

   $$? \text{ cm} = (1 \times 10^{1} \text{ m}) \left(\frac{1 \text{ cm}}{10^{-2} \text{ m}}\right)$$

5. Check that the units properly cancel to yield the desired unit.

Note: If we had multiplied 1×10^{1} m by $\frac{10^{-2} \text{ m}}{1 \text{ cm}}$ instead of $\frac{1 \text{ cm}}{10^{-2} \text{ m}}$ the result would have been

$$? \text{ cm} = (1 \times 10^1 \text{ m}) \left(\frac{10^{-2} \text{ m}}{\text{cm}}\right) = (1 \times 10^1) \left(10^{-2} \frac{\text{m}^2}{\text{cm}}\right)$$

The unit cm does not equal the unit m^2/cm.

6. Do the math portion of the problem

$$? \text{ cm} = (1 \times 10^1 \text{ m}) \left(\frac{1 \text{ cm}}{10^{-2} \text{ m}}\right) = (1 \times 10^1)(10^2 \text{ cm}) = 1 \times 10^3 \text{ cm}$$

Therefore: 1×10^1 m = 1×10^3 cm.

Temperature

A chemist measures temperature in many types of experiments. Temperature is an example of an intensive property. In Chapter 5, we will relate temperature and the energies of motion of gas molecules. Several temperature scales are used to measure temperature: Fahrenheit, Celsius (Centigrade) and Kelvin.

EXERCISE 12

Answer the following questions: (a) What is the mathematical relation between the Celsius scale and the Fahrenheit scale? (b) What is the mathematical relation between the Celsius scale and the Kelvin scale?

Solution: (a) °F = (1.8)(°C) + 32°. A 1°C difference in temperature is equivalent to a 1.8°F difference in temperature. The 32° must be added to the 1.8°C because the freezing point of H_2O is 32° higher on the Fahrenheit scale (32°F) than on the Celsius scale (0°C). Thus the Fahrenheit scale is shifted 32° relative to the Celsius scale. (b) K = °C + 273.15°. A 1°C difference is the same as a 1 K difference.

EXERCISE 13

The temperature on a Spring day is around 22°C. What is this temperature in degrees Fahrenheit and degrees Kelvin?

Solution: Write the general equation °F = (1.8)(°C) + 32°. Substitute 22° for °C in the equation and solve for °F.

$$°F = (1.8)(22°) + 32° = 39.6° + 32° = 72° \text{ (rounded off)}$$

Repeat for K = °C + 273.15°.

$$K = 22° + 273.15° = 295° \text{ (rounded off)}$$

Density

Density is an important property of matter. It measures the mass of a definite volume of a substance. Density equals the mass of the substance divided by the volume of the substance: $d = \frac{m}{V}$. Density varies with temperature.

EXERCISE 14

Which has the greater mass, 2.0 cm^3 of iron (d = 7.9 g/cm^3) or 1.0 cm^3 of gold (d = 19.32 g/cm^3)?

Solution: The mass of a substance is related to its density by the equation $d = m/V$. Multiplying both sides of the equation by V yields $m = dV$. The mass of 2.0 cm^3 of iron is

$$m = dV = \left(7.9 \frac{g}{cm^3}\right) (2.0\ cm^3) = 16\ g$$

The mass of 1.0 cm^3 of gold is

$$m = dV = \left(19.32 \frac{g}{cm^3}\right) (1.0\ cm^3) = 19\ g$$

1.0 cm^3 of gold has a greater mass than 2.0 cm^3 of iron.

KEY TERMS

Review the definitions of the following key terms which are listed at the end of Chapter 1:

combustion	mass	scientific law
density	matter	significant figures
extensive property	paradigm	theory
intensive property		

Having reviewed the definitions of these terms, use the appropriate key term to fill in the blanks in the following statements. Correct answers are given following the statements.

1. Water, salt, baking powder, and plastic wrap are examples of ______________________________.
2. The number 6.203×10^3 has four ________________________________.
3. The burning of wood in a fireplace is an example of _______________.
4. Scientists may operate under a set of universal ideas or beliefs which is termed a __________________________.
5. A chemist may develop a hypothesis to explain a series of similar reactions. The hypothesis becomes a _____________________ after others have confirmed the hypothesis.
6. __________________________ has the unit mass/volume.
7. The ratio mass/volume does not depend on the amount of the sample and is termed an __________________________ property.
8. A liter of water contains 1.000 kg of water. Thus, the ____________ of the water is 1.000 kg.
9. The attraction between two oppositely charged particles of charge q+ and q- is E = (q+)(q-)/r where r is the distance between the particles. This relation is an example of a _____________________________.
10. The volume of a substance depends on the amount of the sample and is termed an __________________________ property.

<u>Answers</u>: (1) matter; (2) significant figures; (3) combustion; (4) paradigm; (5) theory; (6) density; (7) intensive; (8) mass; (9) scientific law; (10) extensive.