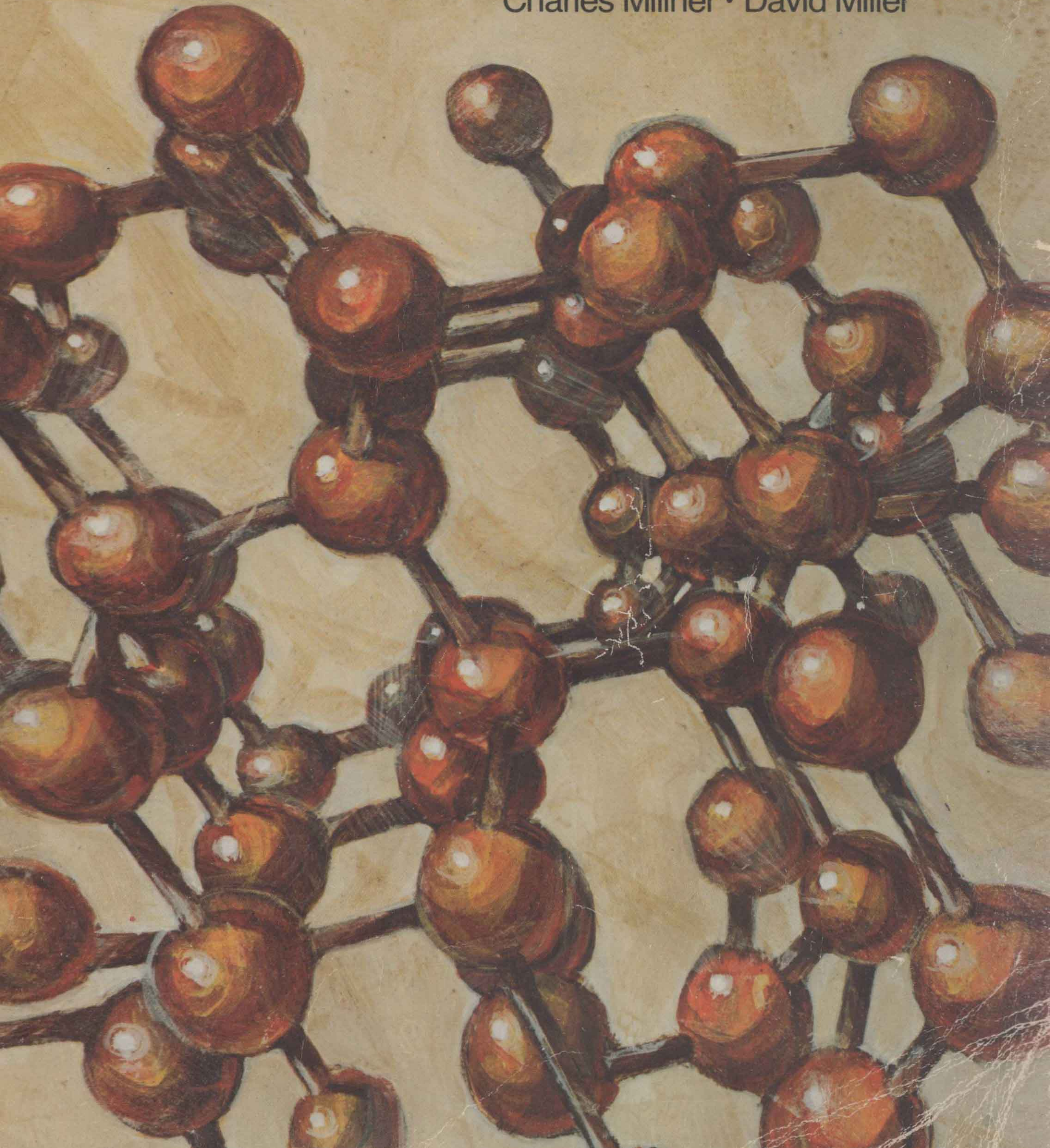
Study Guide to accompany

Chemistry

Abrash · Hardcastle

Prepared by
Charles Millner · David Miller



STUDY GUIDE

TO ACCOMPANY

CHEMISTRY

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GLENCOE PUBLISHING CO., INC.
Encino, California
Collier Macmillan Publishers
London

Glencoe Publishing Co., Inc.
17337 Ventura Boulevard
Encino, California 91316
Collier Macmillan Canada, Ltd.

1 2 3 4 5 6 7 8 9 10 83 82 81

ISBN 0-02-471120-9

TO THE STUDENT

This study guide supplements *Chemistry* by Henry I. Abrash and Kenneth I. Hardcastle. For each chapter in *Chemistry*, there is a corresponding chapter in this guide. The material in this study guide neither re-states nor simply amplifies what is in the text, however. Instead, it is designed to help you focus your study practices on the important concepts covered in the text and to provide you with an additional means of measuring your understanding of these concepts. This is accomplished by providing you with additional example problems and solutions, plus a variety of questions (including answers) to test your understanding.

Each study guide chapter includes the following sections.

Learning objectives:	A list of the important goals of the chapter focuses your attention on key points.
Summary:	The relevance of the material presented in the chapter is explained, and example problems are introduced.
Examples:	Typical problems involving the chapter material are posed, and complete, step-by-step solutions are discussed.
Key terms:	The important vocabulary of the chapter is summarized. You should be prepared not only to define these terms but to explain the significance or application of each as well.
Thought exercises:	A number of true/false and multiple-choice questions are asked to test your understanding of the key concepts.
Self-test:	Additional problems are provided to help you further check your grasp of the subject matter.

The proper way to use this study guide is first to study the text, then to work the problems at the end of each chapter. When you encounter a difficult topic or wish to test further your understanding of the text material, consult the study guide for additional examples and practice problems. When working text or study guide problems, treat them as examination questions; that is, first attempt to answer each without referring to the text or similar, solved examples. "Knowing" how someone else solves problems is important, but your ability to solve problems unaided is a much more valid measure of your understanding of the concepts and procedures.

Perhaps the best advice to follow in the study of chemistry is: Budget your study time wisely and diligently complete the homework assignments. Avoid falling behind the class schedule. Chemistry is not a subject that lends itself to last-minute cramming. Hopefully, this study guide will help you organize your study of chemistry and will add to your enjoyment and success in this field.

CONTENTS

LEARNING OBJECTIVES

After studying this chapter, you should be able to:

- Explain the scientific method.
- Describe the relationships among atoms, molecules, isotopes, compounds, elements, mixtures, physical and chemical changes, and nuclear reactions.
- Use the SI system and perform unit conversions.
- Explain the laws of conservation of mass, definite proportions, and conservation of energy.
- Discuss Dalton's theory and its relationship to the above laws.
- Describe the relationships and differences among heat, temperature, work, and energy.

SUMMARY

A traditional definition of chemistry is "the study of matter, and the changes it undergoes." A central feature of this study is the scientific method. The scientific method is a logical approach to problem solving. The key elements of the scientific method are:

1. *Scientific law*—an empirically determined relationship universally agreed upon. It is established by a large number of independent observers and is usually based on information gathered over a relatively long period of time.

2. *Problem statement*—a specific question concerning the law, the answer to which will be of practical benefit. A question stated in too general terms will be very difficult or impossible to answer; a question stated in too specific terms will be of very limited benefit once answered.

3. *Hypothesis*—a tentative answer to a question. A proper hypothesis must satisfy two requirements: (a) The hypothesis must yield the law as a natural consequence when methods of math, physics, biology, and so on, which are known to be valid, are applied; (b) the hypothesis must be testable.

4. *Experiment*—the actual laboratory collection of quantitative data. This must be a valid test of the hypothesis.

5. *Theory*—a hypothesis which has not been disproven by repeated experiments. Theories serve as predictive tools to help answer more complex questions and are subject to modification when more sophisticated experiments become possible.

Matter may undergo three types of changes: physical, chemical, and nuclear. When the composition of molecules is not altered, the change is physical. Examples of physical changes are the boiling of liquid water to form steam and the stretching of a rubber band. Chemical changes involve the formation of molecules from atoms or an alteration in the nature of existent molecules. Examples of chemical changes are the reaction of hydrogen molecules with oxygen molecules to form water molecules and the souring of milk. The radioactive decay of carbon-14 exemplifies a nuclear reaction. This type of change is characterized by transformations on the subatomic level, and results in the formation of different types of atoms.

All chemical changes have some fundamentally important features in common. Adherence to the law of mass conservation is one of these features. The law of mass conservation is a generalization of the observation that during the course of ordinary chemical reactions, mass is neither destroyed nor created. In other words, the total mass of the products is equal to the total mass of the reactants.

A second feature common to all chemical reactions is adherence to the law of definite proportions. According to this law, a pure compound, regardless of how it is prepared, always contains the same elements in the same definite proportions by mass. At the beginning of the nineteenth century, John Dalton unified these two laws to obtain an atomic theory, the success of which firmly established the concept of the quantification of mass in terms of atoms. Even though there have been modifications to Dalton's theory (for example, by the discovery of isotopes), the fundamental importance of his theory remains. Dalton's theory, of course, originated as a hypothesis which was subsequently shown by experiment to be valid. As practice in the application of the scientific method you may find it interesting to convince yourself that Dalton's theory actually yields the laws of definite proportions and mass conservation as a natural consequence. To do this, apply logical reasoning to the essential principles of Dalton's theory as stated in the text, section 1.4.

The law of multiple proportions provides a good example of the predictive value of a good theory such as Dalton's. Dalton was able to predict that when several different compounds are composed of the same elements but have different percentage compositions, then there should be a simple relationship between the relative masses of the elements in the different compounds. When two elements, *A* and *B*, can react to form different compounds, the ratio of the number of grams of *A* per gram of *B* in one compound to the number of grams of *A* per gram of *B* in a second compound is a ratio of simple whole numbers. As an example, onsider the compounds phosphorus trichloride and phosphorus pentachloride. Each compound is composed solely of phosphorus and chlorine. They are represented by the formulas PCl_3 and PCl_5, respectively. In phosphorus trichloride, there are 3.4338 grams of chlorine per gram of phosphorus. There are 5.7231 grams of chlorine per gram of phosphorus in phosphorus pentachloride. The relationship between these two numbers is clearly apparent:

$$\frac{5.7231 \frac{Cl}{P}}{3.4338 \frac{Cl}{P}} = 1.6667 = \frac{5}{3}$$

As suggested in the calculation above, it will be useful in the future to be able to recognize decimal equivalents of fractions. Some common fractions and their decimal equivalents are listed in Table 1-1.

Table 1-1. Decimal Equivalents of Fractions

Decimal number	Fractional equivalent
0.5000	1/2
0.3333	1/3
0.2500	1/4
0.2000	1/5
0.1667	1/6
0.1429	1/7
0.1250	1/8
0.1111	1/9
0.1000	1/10

Energy changes are frequently associated with chemical reactions. Typically, these energy changes take the form of a flow of heat (thermal energy). We find that molecular motion is closely connected to thermal energy or heat. This makes necessary a means of measuring heat and thereby thermal energy. The measurement of temperature and temperature changes is the means used to quantify thermal energy transfers. The two common temperature scales are the Fahrenheit and the Celsius scales. These two temperature scales are related according to equation (1-1), where F is the Fahrenheit temperature and C is the Celsius temperature.

$$°C = \frac{5}{9}(°F - 32) \tag{1-1}$$

When performing calculations involving heat and temperature, an additional conservation law comes into play. The law of conservation of energy states that the total energy of the universe is constant. This law will be discussed in considerable detail later.

KEY TERMS: DEFINE AND EXPLAIN

scientific law	isotope
theory	atomic mass unit (amu)
mixture	heat
reactant	temperature
product	force
element	momentum
compound	work
mass	energy
weight	kinetic energy
quantum	potential energy
atom	Celsius scale
molecule	heat capacity

THOUGHT EXERCISES

Label each of the following statements as True or False.

1. _______ Chemical reactions result in the transformation of one element into another.

2. _______ A mixture may be separated into its constituent pure substances by physical methods.

3. _______ A force performs work on an object by moving it through a distance.

4. _______ A molecule is the smallest division of matter possible.

5. _______ The burning of a 10-pound log to produce 3 pounds of ash represents a violation of the law of conservation of mass.

6. _______ *Heat* and *temperature* have the same meaning.

7. _______ There are two types of mechanical energy: kinetic energy and potential energy.

8. _______ A mixture is always heterogeneous.

9. _______ One atom of carbon-12 weighs 1 amu.

10. _______ An increase in the momentum of an object results in a decrease in the object's kinetic energy.

11. _______ New theories become a part of the accepted body of scientific knowledge by receiving a majority vote of approval of the board of governors of the International Union of Pure and Applied Scientists, which is headquartered in Paris, France.

12. _______ Atoms of the same element with different masses are called isotopes.

13. _______ A joule is the fundamental unit of measurement of work.

14. _______ Chemical reactions bring about changes in molecular structure.

15. _______ Pure substances cannot be separated into their component parts by physical methods.

Choose the best response to complete each of the following statements.

1. Which of the following is the smallest division of matter which can exhibit the properties of a particular element?

   a. Molecule
   b. Atom
   c. Milligram
   d. amu

2. Temperature is

   a. a measure of the intensity of heat energy.
   b. independent of the size of the object.
   c. measured in degrees Celsius.
   d. all of the above.

3. An elastic collision is one in which

   a. kinetic energy is converted to heat.
   b. kinetic energy is converted to potential energy.
   c. kinetic energy is simply conserved.
   d. there is no change in the velocity of the colliding objects.

4. The most important feature of a good experiment is

   a. that it be cheap to perform.
   b. that it be a valid test of the hypothesis.
   c. that it not require a great deal of time to perform.
   d. that the results are repeatable.

5. The weight of an object

   a. depends only on the amount of matter in the object.
   b. is the same as the mass of the object.
   c. depends on its environment.
   d. is measured in terms of liters.

6. The best description of a quantum is

   a. an indivisible unit of definite size.
   b. a particle involved in all nuclear reactions.
   c. the internationally accepted unit of astronomical time measurement.
   d. the distinguishing feature among different isotopes.

7. Energy can be best described as

   a. a measure of heat content.
   b. the ability to do work.
   c. being proportional to an object's temperature.
   d. being directly related to an object's velocity.

8. 4.6 meters equals

   a. 4.6×10^{3} mm
   b. $4.6 \times 10^{+6}$ μm
   c. 4.6×10^{3} km
   d. all of the above.

9. The heat capacity of an object is

   a. characteristic of its elemental composition.
   b. the same as the object's specific heat.
   c. measured in degrees Fahrenheit.
   d. none of the above.

10. An object whose velocity is changing is

   a. being acted on by a force.
   b. undergoing an acceleration.
   c. experiencing a change in its momentum.
   d. all of the above.

SELF-TEST

1. One inch equals 2.54 cm. Calculate how many kilometers are in one mile.

2. The density of tin is 7.28 g/cm^3. Express this density in kg/m^3.

3. The compounds nitrous oxide and dinitrogen tetroxide are composed solely of nitrogen and oxygen. Nitrous oxide is 63.64% by weight nitrogen, and dinitrogen tetroxide is 30.43% by weight nitrogen. Show that these data are consistent with the law of multiple proportions.

4. What type of change does each of the following exemplify?

   a. The melting of ice

   b. Baking a cake

   c. Hydrogen and lithium reacting to produce helium

   d. The dissolving of sugar in water

   e. The production of hydrogen and oxygen gas by passing an electric current through water

5. One Fahrenheit thermometer and one Celsius thermometer are simultaneously enclosed in the same temperature controlled environment. Both thermometers indicate the same numerical value for the temperature. What is this temperature?

6. What is the kinetic energy of a 907 kg car traveling at 88.6 km/hr?

7. If the energy of the car in problem 6 were converted to heat and used to warm a swimming pool containing 9.45×10^7 g of water (about 25,000 gallons), what change in the temperature of the water would occur?

LEARNING OBJECTIVES

After studying this chapter you should be able to:

- Determine both empirical and molecular formulas.
- Describe the mole concept and its application.
- Balance chemical equations.
- Explain how to use a balanced equation to do stoichiometric calculations.
- Apply the limiting reagent concept.

SUMMARY

Chemical calculations that deal with the relative amounts of materials involved in chemical reactions are described as stoichiometric calculations. Formulas specify the number and type of atoms that make up molecules. It is possible to classify formulas into two broad categories: empirical (or simplest) and molecular formulas. Empirical formulas specify the types of atoms in molecules and their percent composition by mass. If, in addition to the percent composition, the molecular mass of a compound is known, it is possible to determine the molecular formula.

Because atoms and molecules and similar chemical species have very small masses and physical dimensions, chemists rarely deal with very small numbers of such particles. Instead, large groups of atoms or molecules are dealt with, and the mole is the basic unit used by chemists for counting such particles. One mole describes a collection of particles numbering 6.023×10^{23}.

In writing chemical equations, molecular formulas are used when available; simple, whole-number coefficients are used to indicate the relative amounts of reactants and products. Frequently, the proper coefficients may be determined by trial and error. When the trial and error method becomes unwieldy, algebraic methods can be used.

A balanced chemical equation is very useful in calculating the quantitative relationships that exist among substances in a chemical reaction. Often it becomes apparent that the amount of products in a chemical reaction is dependent upon the amount of some limiting reactant available. Examples of these stoichiometric calculations are provided in this chapter.

Examples 2.1 and 2.2 illustrate the procedures used for determining empirical formulas, and example 2.3 illustrates those used for determining molecular formulas. Examples 2.4 through 2.7 illustrate the utility of the mole concept. Examples 2.8 and 2.9 illustrate the procedures of stoichiometric calculations, and examples 2.10 and 2.11 demonstrate the principles involved in limiting reactants.

Example 2.1

A compound composed of only nitrogen and oxygen is found to be 26.00 percent nitrogen by weight. What is the empirical formula of this compound?

The empirical formula takes the general form N_xO_y, where the subscripts x and y are the smallest whole numbers that indicate the correct ratio of nitrogen to oxygen atoms in the compound. To determine x and y, calculate the number of nitrogen and oxygen atoms present in some arbitrary amount of the compound. One hundred amu of the compound is a convenient amount to use because the percent composition data given can be applied easily to give the masses of each element present in this amount of compound. Because the compound is 26.00 percent nitrogen by weight, the remaining 74.00 percent must be oxygen, and in 100 amu of the compound there are 26.00 amu of nitrogen and 74.00 amu oxygen. These masses can be converted into number of atoms of each element by using the respective atomic masses.

$$x = 26.00\ \cancel{\text{amu N}} \times \frac{1\ \text{atom N}}{14.01\ \cancel{\text{amu N}}} = 1.856\ \text{atoms N}$$

$$y = 74.00\ \cancel{\text{amu O}} \times \frac{1\ \text{atom O}}{16.00\ \cancel{\text{amu O}}} = 4.625\ \text{atoms O}$$

The formula becomes $N_{1.865}O_{4.625}$. To convert the subscripts to whole numbers, first divide each subscript by the smaller one, that is

$$N_{\frac{1.856}{1.856}}O_{\frac{4.625}{1.856}} = N_{1.000}O_{2.492}$$

The subscript of oxygen seems to present a problem because it is not yet a whole number, and it is not clear whether it should be rounded off to 2 or to 3. A good rule of thumb is "never round off subscripts, either up or down, by more than a tenth of a unit (0.100)." Rather than rounding off by more than 0.1, some integer multiplier should be used that will convert all subscripts to whole numbers. In this example, the obvious choice of a multiplier is the integer 2. Multiplying both subscripts by 2 gives $N_2O_{4.984}$, which is rounded off correctly to N_2O_5.

Example 2.2

Ethylenediamine is a compound composed solely of carbon, hydrogen, and nitrogen. On a mass basis, it contains 40.00 percent carbon and 46.60 percent nitrogen. What is the empirical formula of ethylenediamine?

The empirical formula is defined as before. Assume 100 amu of $C_xH_yN_z$, which contains 40.00 amu carbon, 46.60 amu nitrogen and, by difference (100 - 40 - 46.6), 13.40 amu hydrogen. Therefore,

$$x = 40.00 \cancel{\text{amu C}} \times \frac{1 \text{ atom C}}{12.01 \cancel{\text{amu C}}} = 3.331 \text{ atoms C}$$

$$y = 13.40 \cancel{\text{amu H}} \times \frac{1 \text{ atom H}}{1.01 \cancel{\text{amu H}}} = 13.3 \text{ atoms H}$$

$$z = 46.60 \cancel{\text{amu N}} \times \frac{1 \text{ atom N}}{14.01 \cancel{\text{amu N}}} = 13.326 \text{ atoms N}$$

and the basic formula becomes $C_{3.331}H_{13.3}N_{3.326}$. Dividing each subscript by the smallest one gives

$$C_{\frac{3.331}{3.326}}H_{\frac{13.3}{3.326}}N_{\frac{3.326}{3.326}} = C_{1.002}H_{4.00}N_{1.000}$$

Since the subscripts are sufficiently close to whole numbers to be rounded off, the empirical formula is CH_4N.

Example 2.3

The molecular mass of ethylenediamine (example 2.2) is 60.00 amu. What is its molecular formula?

The empirical formula of ethylenediamine was found to be CH_4N, and the molecular formula will be just some integer multiple of this combination of atoms [$(CH_4N)_n$]. Likewise, the molecular mass is some integer multiple of the empirical formula mass; that is

$$\text{molecular mass} = \text{empirical formula mass} \times n$$

where n = 1, 2, 3, ... For ethylenediamine,

$$\begin{aligned}\text{empirical formula mass} &= 1(12.01 \text{ amu}) + 4(1.01 \text{ amu}) + 1(14.01 \text{ amu}) \\ &= 30.06 \text{ amu}\end{aligned}$$

Thus,

$$n = \frac{60.00 \cancel{\text{amu}}}{30.06 \cancel{\text{amu}}} = 1.996$$

which may be rounded off to 2. Therefore, the molecular formula is

$$(CH_4N)_2 = C_2H_8N_2$$

Example 2.4

How many moles of carbon dioxide molecules are there in 3.80 grams of CO_2?

The mass of 1 mole of CO_2 molecules is given by the molecular mass of CO_2 expressed in grams.

$$CO_2 \text{ molecular mass} = 1(12.0 \text{ amu}) + 2(16.0 \text{ amu}) = 44.0 \text{ amu}$$

$$1 \text{ mol } CO_2 \text{ weighs } 44.0 \text{ grams} \quad \text{or} \quad \frac{1 \text{ mol } CO_2}{44.0 \text{ g } CO_2}$$

Multiplying this factor times the mass of CO_2 given above yields the desired results.

$$3.80 \cancel{\text{g } CO_2} \times \frac{1 \text{ mol } CO_2}{44.0 \cancel{\text{g } CO_2}} = 0.0864 \text{ mol } CO_2$$

Example 2.5

How many atoms of carbon are there in 5.228 grams of carbon tetrachloride, CCl_4?

First calculate the number of moles of CCl_4, as outlined above, then convert moles of molecules to number of molecules, and, finally, relate molecules of CCl_4 to atoms of carbon.

$$CCl_4 \text{ molecular mass} = 1(12.0 \text{ amu}) + 4(35.45 \text{ amu}) = 153.8 \text{ amu}$$

$$\text{moles of } CCl_4 \text{ molecules} = 5.228 \cancel{\text{g } CCl_4} \times \frac{1 \text{ mol } CCl_4}{153.8 \cancel{\text{g } CCl_4}}$$

$$= 0.03399 \text{ mol } CCl_4$$

$$\text{number of } CCl_4 \text{ molecules} = 0.03399 \cancel{\text{mol } CCl_4} \times \frac{6.023 \times 10^{23} \text{ molecules } CCl_4}{1 \cancel{\text{mol } CCl_4}}$$

$$= 2.047 \times 10^{22} \text{ molecules } CCl_4$$

$$\text{atoms of carbon} = 2.047 \times 10^{22} \cancel{\text{molecules } CCl_4} \times \frac{1 \text{ atom C}}{1 \cancel{\text{molecule } CCl_4}}$$

$$= 2.047 \times 10^{22} \text{ atoms C}$$