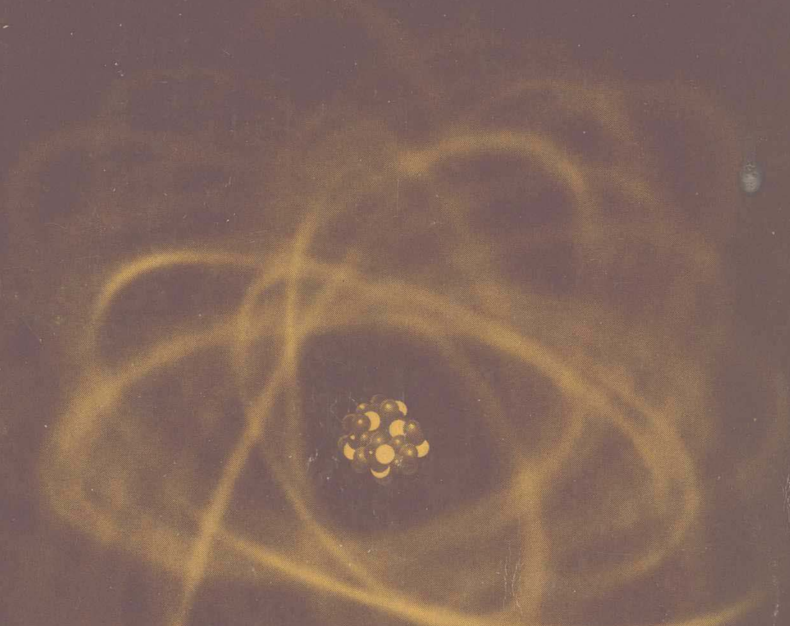
Understanding Problems For



Chemical Principles

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Understanding Problems For Chemical Principles

W. B. SAUNDERS COMPANY
Philadelphia London Toronto

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

1 St. Anne's Road
Eastbourne, East Sussex BN 21 3UN, England

1 Goldthorne Avenue
Toronto, Ontario M8Z 5T9, Canada

Library of Congress Cataloging in Publication Data

Musgrave, Ted R.

Understanding problems for chemical principles.

1. Chemistry, Physical and theoretical—Problems, exercises, etc. I. Title.

QD456.M87 1978 540'.76 77–80749

ISBN 0–7216–6626–4

Cover photo courtesy of Jay M. Pasachoff, as illustrated in his book, *Contemporary Astronomy*, published by W. B. Saunders Company, 1977.

Understanding Problems For Chemical Principles ISBN 0–7216–6626–4

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

To Pat

PREFACE

"The stumbling way in which even the ablest of scientists in every generation have had to fight through thickets of erroneous observations, misleading generalizations, inadequate formulations and unconscious prejudice is rarely appreciated by those who obtain their scientific knowledge from textbooks."

. . . *James Bryant Conant*

At the beginning level of understanding the science of Chemistry, textbooks are the principal source of acquiring some knowledge of chemical principles. Regardless of the quality of a text, the competence of an instructor or the interest and enthusiasm of students, there always seems to be a need (by students) for "more help" with problems or "another source" to guide them toward an understanding of the complexities of chemistry problems. This little book was written to provide a student with some additional "tools" by which chemistry problems can become more understandable. The emphasis is on understanding, not just the *mechanics* of problem solving.

I have tried to put into writing some of the approaches to problems that I have found to be successful in the classroom at Colorado State University and the University of Hawaii. Throughout the book I have occasionally asked unanswered questions, made comments and generally tried to be informal.

Chapters are organized by categories. Sometimes this format of organization suggests artificial boundaries that do not, in fact, exist in chemistry. The format is used primarily to *aid* the student in sorting out the seemingly vast and often confusing *types* of problems encountered in a beginning chemistry course. The book contains 445 problems. The emphasis lies not on sheer numbers of problems but on explanations and *brief* discussion of principles underlying the problems. The sequence of material and general tone of the book are designed to coincide with the 4th Edition of *Chemical Principles* by Masterton and Slowinski.

I extend my appreciation to Dr. W. L. Masterton of the University of Connecticut for his thorough review of the manuscript and his many helpful suggestions. Even Dr. Masterton's most critical comments were a pleasure to read! My thanks to Jackie Swinehart for her invaluable help in preparing the manuscript and to Marilyn Bain-Ackerman for checking the problem solutions.

Ted R. Musgrave

CONTENTS

1 THE FACTOR-LABEL METHOD

The factor-label method, variously called dimensionanalysis, unity-factor, or unit conversion, is a convenient problem solving technique that can be applied to a variety of chemistry problems and, indeed, to many "everyday" practical arithmetic problems. The method involves two steps: (1) stating the problem as a mathematical equation, and (2) multiplying the right side of the equation by conversion (unity) factors until the units on the right side of the equation match the units on the left.

With a current national interest in conversion to the metric system, a good way to introduce the factor-label method is to apply it to some typical metric conversions.

Example 1.

How many centimeters are there in one foot?

Solution.

First, we must state the problem as an equation. When you first attempt to use the factor-label method, this is often the most difficult step. We must set something "equal to" or "equivalent to" something else, and to do this usually requires some mental translation of the way the problem is stated. Another way of stating the problem here would be to ask, "How many centimeters are equivalent to one foot?" The unknown, "how many," we can abbreviate by calling it "X," and we have our equation

$$X \text{ cm} = 1.00 \text{ ft}$$

Next, we must multiply the right side of our equation by suitable conversion factors until "ft" becomes "cm." We must know some numerical conversion for length between the English and metric systems. If we knew that 1 foot = 30.48 cm, that would be fine, and in fact would be the answer to our problem, but let us suppose we did not remember that particular conversion, but we did remember that 1 inch = 2.54 cm. This will serve nicely as a conversion factor, or let's call it a "unity" factor. Unity simply means the number one; any conversion factor can be expressed as the number one. Thus, if we take the expression 1 inch = 2.54 cm and divide both sides by 1 inch, we have

$$\frac{1 \text{ inch}}{1 \text{ inch}} = \frac{2.54 \text{ cm}}{1 \text{ inch}}$$

or

$$1 = \frac{2.54 \text{ cm}}{1 \text{ inch}}$$

and if we divide both sides by 2.54 cm,

$$\frac{1 \text{ inch}}{2.54 \text{ cm}} = \frac{2.54 \text{ cm}}{2.54 \text{ cm}} = 1$$

Similarly, we know that 12 inches = 1 foot, so

$$\frac{12 \text{ inches}}{1 \text{ foot}} = 1, \text{ or } \frac{1 \text{ foot}}{12 \text{ inches}} = 1$$

In the factor-label method, it makes sense to think of conversion factors as unity factors. The reason is that we start with an equation, such as

$$X \text{ cm} = 1 \text{ ft}$$

and then proceed to multiply one side of the equation by conversion factors. Mathematically, in order for the equation to remain true, those conversions must be equal to one. We can multiply one side of an equation by 1 an infinite number of times without changing the original equation! For our problem, we can convert feet to inches by using the unity factor 12 inches/1 foot = 1. Thus,

$$X \text{ cm} = 1.00 \cancel{\text{ft}} \times \frac{12 \text{ in}}{1 \cancel{\text{ft}}}$$

The units of ft cancel out. We can then convert to cm using the unity factor 2.54 cm/1 in = 1.

$$X \text{ cm} = 1.00 \cancel{\text{ft}} \times \frac{12 \cancel{\text{in}}}{1 \cancel{\text{ft}}} \times \frac{2.54 \text{ cm}}{1 \cancel{\text{in}}}$$

The units on the right side now match the units on the left and the problem is solved, except for the arithmetic. Multiplying 12 X 2.54, the answer is 30.5 cm. Once the solution "set up" is correct, the multiplication and division may be accomplished quickly using a pocket calculator or slide rule. Notice that a little inspection (and practice) will show us which way to use a unity factor. Thus, 2.54 cm/1 in = 1, and 1 in/2.54 cm = 1, and in this particular problem we use it so that inches are in the denominator and will cancel to leave the desired units, cm, in the numerator.

Example 2.

How many grams are in one ton?

Solution.

Stated as an equation, we can "ask":

$$X \text{ g} = 1.0 \text{ ton}$$

We need an English to metric conversion for mass, such as 1 pound = 454 grams. In fact, all English to metric conversions can be carried out by the factor-label method if one remembers three conversion factors. One for length, e.g., 1 in = 2.54 cm; one for mass, 1 lb = 454 g; and one for volume, such as, 1 liter = 1.06 quarts. This problem can be solved by using two unity factors:

$$X \text{ g} = 1.0 \cancel{\text{ton}} \times \frac{2000 \cancel{\text{lbs}}}{1.0 \cancel{\text{ton}}} \times \frac{454 \text{ g}}{1 \cancel{\text{lb}}}$$

Units of tons cancel to give lbs, then units of lbs cancel to give the desired units of grams. The numerical answer is left to you.

Example 3.

How many milliliters of gasoline are there in a metric and in a U. S. gallon of gasoline?

Solution.

A metric gallon is 4.0 liters; hence there are 4000 ml in a metric gallon. For a U. S. gallon, we can use factor-label:

$$\text{X ml} = 1\ \cancel{\text{gal}} \times \frac{4\ \cancel{\text{qt}}}{1.0\ \cancel{\text{gal}}} \times \frac{1.0\ \cancel{\text{liter}}}{1.06\ \cancel{\text{qt}}} \times \frac{1000\ \text{ml}}{1.0\ \cancel{\text{liter}}} = 3774\ \text{ml}$$

Example 4.

In 1976, the world's record in the 100 meter dash was 9.9 seconds. To accomplish this feat, what would be the athlete's average velocity in miles per hour?

Solution.

The question may be translated to "How many miles per hour are equivalent to 100 meters per 9.9 seconds?" Since "per" just means "divided by," the problem may be stated mathematically as

$$X\ \frac{\text{miles}}{\text{hour}} = \frac{100\ \text{meters}}{9.9\ \text{seconds}}$$

In this type of problem, there are units in both numerator and denominator that must be converted. It is most convenient to convert them one at a time. In our problem, let's operate on the numerator first and convert meters to miles, then convert seconds to hours in the denominator.

If we knew that 1 kilometer (1000 meters) = 0.62 miles, the former conversion could be accomplished in one step. But let's assume we know only that 1 in = 2.54 cm. Then,

$$X\ \frac{\text{mi}}{\text{h}} = \frac{100\ \text{m}}{9.9\ \text{s}} \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{1\ \text{in}}{2.54\ \text{cm}} \times \frac{1\ \text{ft}}{12\ \text{in}} \times \frac{1\ \text{mi}}{5280\ \text{ft}}$$

Successively, units of m, cm, in, and ft cancel out until we are left with units of miles in the numerator. If we stopped at this point, the "answer" would be in miles per second, so we must next convert the denominator to hours:

$$X\ \frac{\text{mi}}{\text{h}} = \frac{100\ \text{m}}{9.9\ \text{s}} \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{1\ \text{in}}{2.54\ \text{cm}} \times \frac{1\ \text{ft}}{12\ \text{in}} \times$$

$$\frac{1\ \text{mi}}{5280\ \text{ft}} \times \frac{60\ \text{s}}{1\ \text{min}} \times \frac{60\ \text{min}}{1\ \text{h}}$$

We now have the units we want, $\frac{\text{mi}}{\text{h}}$. The numerical answer is 22.6 mph. The world record in the 100 yard dash is 9.0 seconds. How many miles per hour is that? The record in the 1000 meters is 1 minute, 43.7 seconds. What would that average in mph? (ans. 21.5 mph)

Example 5.

A graphite pencil signature weighs 1.0 milligram. How many carbon atoms are there in the signature?

Solution.

The "question" equation is

$$X \text{ C atoms} = 1.0 \text{ mg C}$$

For the appropriate conversion factors, we need a basic concept of chemistry, one which you may not have covered at this point. That is, that there are 6.02×10^{23} atoms in 12.0 g of carbon: $12.0 \text{ g C} = 6.02 \times 10^{23}$ C atoms. The problem may now be solved as follows:

$$X \text{ C atoms} = 1.0 \text{ mg C} \times \frac{1 \text{ g C}}{1000 \text{ mg C}} \times \frac{6.02 \times 10^{23} \text{ C atoms}}{12.0 \text{ g C}} = 5.0 \times 10^{19} \text{ atoms}$$

Example 6.

The velocity of light is 3.00×10^{10} cm per second. How many miles does light travel in one year (i.e., one light-year)?

Solution.

Examine the solution step by step. It is lengthy but not difficult.

$$X \text{ mi} = 1 \text{ yr} \times \frac{365 \text{ d}}{1 \text{ y}} \times \frac{24 \text{ h}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{3.00 \times 10^{10} \text{ cm}}{1 \text{ s}} \times \frac{1 \text{ in}}{2.54 \text{ cm}} \times \frac{1 \text{ ft}}{12 \text{ in}} \times \frac{1 \text{ mi}}{5280 \text{ ft}} = 5.88 \times 10^{12} \text{ miles}$$

or 5,880,000,000,000 miles!

PROBLEMS

1. How many yards are there in 100 meters?
2. How many kilograms are in a 5 lb bag of sugar?
3. The density of mercury is 13.6 g per ml. What is the weight (in lbs) of one quart of mercury?
4. The distance to the sun is 93 million miles and light travels at 3.0×10^{10} cm per second. How long does it take for light from the sun to reach the earth?
5. How long would it take an automobile averaging 55 mph to travel 200 kilometers?
6. The artificial element californium 252 sells for \$100 per tenth of a microgram (10^{-7} g). If one pound of ^{252}Cf were for sale, what would it cost?
7. A metric wrench has a span of 20 millimeters. How many inches is that?
8. If a carbon dioxide molecule were 5.0 angstroms in length, how many CO_2 molecules laid end to end would it take to reach 1 mile? (one angstrom = 10^{-8} cm)
9. Convert 55 miles per hour to meters per second.
10. What would be the weight (in lbs) of 100 billion atoms of lead? (at. wt. 207)

2 MOLES, FORMULAS, AND STOICHIOMETRY

The term stoichiometry refers to relationships between quantities of reactants and products in chemical systems. Formulas tell us the relative numbers of different atoms making up a chemical compound. One of the most important quantitative concepts in chemistry is the *mole.* Since individual atoms or molecules are so small as to be virtually impossible to work with, the idea of a large collection of atoms or molecules, namely the "mole," was developed. The mole is defined as the number of carbon atoms in 12.000 grams of the isotope $^{12}_{6}C$ – ordinary carbon. That number, known as *Avogadro's Number*, is 6.023×10^{23}. Strictly speaking, one mole of anything – atoms, molecules, ions, stars, grains of sand, etc. – is defined as 6.023×10^{23} of that thing. This is a huge number! One mole of marbles, for example, would completely cover the surface of 250 planets the size of the earth to a depth of five feet.

The *atomic weight* listed on periodic tables gives the mass of an atom of an element as it occurs in nature relative to that of a ^{12}C atom, which is taken to be exactly 12. Atomic weights are average values. Often they are not whole numbers because most natural elements consist of more than one isotope. Carbon, for example, is mostly $^{12}_{6}C$ but contains small amounts of $^{13}_{6}C$ and $^{14}_{6}C$, so the average atomic weight of carbon is 12.01115. The atomic weight of hydrogen is 1.00797; of iron, 55.847, and so on. While these numbers are often known to six or seven significant figures, the chemist usually rounds off to three or four figures. The *gram atomic weight* is the mass of an element, in grams, which is numerically equal to its atomic weight.

Molecular weight may be determined by summing atomic weights when the formula of a molecular substance is known. Thus, if we know that the formula of water is H_2O, this tells us that each water molecule contains two hydrogen atoms and one oxygen atom. One mole of water molecules would therefore contain two moles of hydrogen atoms and one mole of oxygen atoms. The molecular weight of H_2O is $(2 \times 1.01) + (1 \times 16.00) = 18.02$. The *gram molecular weight* of water is 18.02 grams.

Formula weight is often used instead of molecular weight. Strictly speaking, the term "molecular weight" should be used only with a molecular substance. Ionic compounds, such as NaCl, do not consist of molecules. Therefore, it is incorrect to refer to the "molecular weight" of sodium chloride, but it is perfectly accurate to say that the formula weight of NaCl is 23.0 + 35.5 = 58.5. One mole of sodium chloride weighs 58.5 grams.

Problem Categories

Problems in this chapter are presented in three categories:

I Conversions between grams, atoms, moles, and molecules.

II Relationship of chemical formulas to elementary composition.
III Mass relationships in chemical equations.

CATEGORY I CONVERSIONS BETWEEN GRAMS, ATOMS, MOLES, AND MOLECULES

Example 1.

How many moles of sulfur atoms are there in 8.23 grams of sulfur, S?

Solution.

One may convert from grams to moles by using the factor-label method. Here, we make use of the unity factor 1 mole S atoms = 32.06 grams S. Thus,

$$\text{X moles S} = 8.23 \text{ grams S} \times \frac{1 \text{ mole S (atoms)}}{32.06 \text{ grams}} = 0.257 \text{ mole}$$

Example 2.

How many iron atoms are there in 1.00 gram of iron, Fe?

Solution.

Solving the problem requires two steps. First, convert grams to moles, then use Avogadro's number as a conversion factor since 1 mole of Fe atoms is equivalent to 6.02×10^{23} Fe atoms. Thus,

$$\text{X Fe atoms} = 1.00 \text{ g Fe} \times \frac{1 \text{ mole Fe (atoms)}}{55.85 \text{ g Fe}} \times \frac{6.02 \times 10^{23} \text{ Fe atoms}}{1 \text{ mole Fe (atoms)}}$$

$$= 1.08 \times 10^{22} \text{ Fe atoms}$$

Example 3.

How many moles of carbon dioxide, CO_2, are there in 20.0 grams of CO_2?

Solution.

The molecular weight of CO_2 is 12.01 + 16.00 + 16.00 = 44.01 grams; thus,

$$1 \text{ mole } CO_2 = 44.01 \text{ g } CO_2$$

By the factor-label method:

$$\text{X moles } CO_2 = 20.0 \text{ g } CO_2 \times \frac{1 \text{ mole } CO_2}{44.01 \text{ g } CO_2} = 0.454 \text{ mole}$$

Example 4.

How many sugar molecules are there in 5.00 grams of sugar, $C_{12}H_{22}O_{11}$?

Solution.

The solution is similar to Example 2, except that we now have molecules instead of atoms. The molecular weight of $C_{12}H_{22}O_{11}$ is $(12 \times 12.01) + (22 \times 1.01) + (11 \times 16.00) = 342.3$ grams.

$$X\ C_{12}H_{22}O_{11} \text{ molecules} = 5.00 \text{ g } C_{12}H_{22}O_{11} \times \frac{1 \text{ mole } C_{12}H_{22}O_{11}}{342.3 \text{ g } C_{12}H_{22}O_{11}}$$

$$\times \frac{6.02 \times 10^{23}\ C_{12}H_{22}O_{11} \text{ molecules}}{1 \text{ mole } C_{12}H_{22}O_{11}} = 8.79 \times 10^{21} \text{ molecules}$$

Example 5.

If a box contained 9,000,000,000,000 oxygen molecules, O_2, how many moles of oxygen would that be?

Solution.

$$X \text{ moles } O_2 = 9 \times 10^{12} \text{ molecules } O_2 \times \frac{1 \text{ mole } O_2}{6.02 \times 10^{23} \text{ molecules } O_2}$$

$$= 1.5 \times 10^{-11} \text{ mole}$$

RELATIONSHIP OF CHEMICAL FORMULAS TO ELEMENTARY COMPOSITION

CATEGORY II

Example 1.

How many grams of carbon and of chlorine are there in 0.50 mole of carbon tetrachloride, CCl_4?

Solution.

From the formula CCl_4, we know that each mole of carbon tetrachloride contains one mole of carbon atoms and four moles of chlorine atoms. Furthermore, we know that 1 mole C (atoms) = 12.01 g C (atoms), so by factor-label

$$X \text{ g C} = 0.50 \text{ mole } CCl_4 \times \frac{1 \text{ mole C (atoms)}}{1 \text{ mole } CCl_4 \text{ (molecules)}} \times \frac{12.01 \text{ g C (atoms)}}{1 \text{ mole C (atoms)}} = 6.01 \text{ grams}$$

Since there are four moles of chlorine atoms per mole of CCl_4, and the atomic weight of chlorine is 35.45, by factor-label

$$X \text{ g Cl} = 0.50 \text{ mole } CCl_4 \times \frac{4 \text{ moles Cl}}{1 \text{ mole } CCl_4} \times \frac{35.45 \text{ g Cl}}{1 \text{ mole Cl}} = 70.90 \text{ grams}$$

Example 2.

How many grams of sulfur are there in 10.0 grams of sulfur trioxide, SO_3?

Solution.

From the formula we see that one mole of SO_3 (molecules) contains one mole of S (atoms). Solving by the factor-label method,

$$X \text{ g S} = 10.0 \text{ g } SO_3 \times \frac{1 \text{ mole } SO_3}{80.1 \text{ g } SO_3} \times \frac{1 \text{ mole S}}{1 \text{ mole } SO_3} \times \frac{32.1 \text{ g S}}{1 \text{ mole S}} = 4.01 \text{ g}$$

The second unity factor may be eliminated by direct use of the relationship: 32.1 g S (atoms) is equivalent to 80.1 g SO_3 (molecules). Thus, we could have written

$$X \text{ g S} = 10.0 \text{ g } SO_3 \times \frac{32.1 \text{ g S}}{80.1 \text{ g } SO_3} = 4.01 \text{ g}$$

Example 3.

How many chlorine atoms are in 0.20 gram of DDT, $C_{14}H_9Cl_5$?

Solution.

There are 5 moles of Cl atoms per mole of DDT, and the molecular weight of DDT is $(14 \times 12.01) + (9 \times 1.01) + (5 \times 35.45) = 354.5$ g.

$$X \text{ Cl atoms} = 0.20 \text{ g DDT} \times \frac{1 \text{ mole DDT}}{354.5 \text{ g DDT}} \times \frac{5 \text{ moles Cl}}{1 \text{ mole DDT}} \times \frac{6.02 \times 10^{23} \text{ Cl atoms}}{1 \text{ mole Cl}} = 1.70 \times 10^{21} \text{ atoms}$$

Example 4.

A sample of a chemical compound was analyzed and found to contain 5.60 grams of nitrogen and 12.80 grams of oxygen. What is the formula of the compound?

Solution.

Formulas are written to correspond to the relative numbers of atoms in molecules, ions, etc. Formulas of chemical compounds are determined from the ratios of the number of moles of the various atoms composing the compound. We know that this compound consists of N and O and will have a formula N_zO_y. The problem is to find z and y. From analysis, we know there are 5.60 g of N and 12.80 g of O. The first step is to convert to moles of N and O.

$$X \text{ moles N} = 5.60 \text{ g N} \times \frac{1 \text{ mole N}}{14.0 \text{ g N}} = 0.40 \text{ mole N}$$

$$X \text{ moles O} = 12.80 \text{ g O} \times \frac{1 \text{ mole O}}{16.0 \text{ g O}} = 0.80 \text{ mole O}$$

Since analysis of a compound is performed on some arbitrary amount, it would be coincidental if these numbers came out exactly 1, 2, 3, etc. They are usually always non-integers. Compounds cannot be made up of fractions of atoms, so we must convert the ratio of atoms to whole numbers. This is done by dividing the larger number(s) by the smallest:

$$\frac{0.80}{0.40} = \frac{2}{1} = \frac{\text{moles O atoms}}{\text{moles N atoms}}$$

The formula of the compound is NO_2.
The formula we have found, NO_2, is called the *simplest formula*. As the name implies, it tells us the simplest whole number ratio of atoms in the compound. However, this does not necessarily mean that the compound consists of molecules of NO_2. The compound may be ionic rather than molecular. Some information other than elemental analysis would be needed to determine the chemical nature of the compound. Even if the compound "NO_2" were known to be molecular, we would not

know for sure if the molecules consisted simply of one nitrogen atom and two oxygen atoms, or two nitrogen atoms and four oxygen atoms, etc. All we know is that the ratio of N to O is 1:2. More information would be needed before we could ascertain the *molecular formula*.

Example 5.

Experiment revealed that the compound in Example 4 was molecular and had a molecular weight of 138.0. What is the molecular formula of the compound?

Solution.

If the simplest and molecular formulas were identical, i.e., NO_2, what would be the molecular weight? (14.0) + (2 X 16.0) = 46. Obviously, the molecules must consist of more than one atom of N and two of O. What if the molecules were N_2O_4? The molecular weight would be: (2 X 14.0) + (4 X 16.0) = 92. How about N_3O_6? The molecular weight would be: (3 X 14.0) + (6 X 16.0) = 138. Bingo! The molecular formula is N_3O_6.

Thus, if the molecular weight of the compound is known, the molecular formula can be determined by dividing the true molecular weight by the simplest formula weight. In this case, 138 ÷ 46 = 3 shows us that the molecular formula is three times the simplest formula, NO_2, or N_3O_6.

Example 6.

Analysis of a compound gave the following composition: 1.59% H, 22.22% N, and 76.19% O. Calculate the simplest formula of the compound.

Solution.

Chemical analyses are usually reported as per cent (by weight). How do you find moles from per cent? Choose any number of grams of compound you wish, then knowing per cent by weight, you can find the number of grams of a particular element in a known weight of that compound. Thus, in our example, one might arbitrarily choose 20.0 grams of the compound. How much of that is hydrogen? 1.59% of 20 g = 0.318 g. The easy way, however, is to choose 100.0 grams. Then 1.59%, 22.22%, and 76.19% would be 1.59, 22.22, and 76.19 grams of H, N, and O respectively. Now we can convert to moles.

$$\text{X moles H} = 1.59 \text{ g H} \times \frac{1 \text{ mole H}}{1.01 \text{ g H}} = 1.57 \text{ moles H}$$

$$\text{X mole N} = 22.22 \text{ g N} \times \frac{1 \text{ mole N}}{14.0 \text{ g N}} = 1.59 \text{ moles N}$$

$$\text{X mole O} = 76.19 \text{ g O} \times \frac{1 \text{ mole O}}{16.0 \text{ g O}} = 4.76 \text{ moles O}$$

The atomic weights were rounded off to three significant figures, but the mole ratio of N/H is obviously 1/1. The ratio of O/H, or O/N, is: 4.76/1.59 = 3/1. The simplest formula, then, is: $H_1N_1O_3$, or HNO_3.

MASS RELATIONSHIPS IN CHEMICAL REACTIONS CATEGORY III

In the examples in Categories I and II we have examined various conversions between grams, atoms, molecules, and moles and related these quantities to chemical formulas.

Now, let's turn our attention to what is called stoichiometry, that is, how these quantities relate to chemical equations. A chemical equation is an abbreviated means of quantitatively describing what happens in a chemical reaction. For example, molecules of methane and oxygen react to give molecules of carbon dioxide and water. This statement can be abbreviated by using molecular formulas and writing

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

We can call this a "reaction expression," but it is not (yet) a chemical equation. In chemistry, as in mathematics, an equation must be "balanced." In the course of a chemical reaction, atoms cannot magically appear or disappear. We must end up with the same number and kind that we started with, regardless of how they become rearranged in the reaction. So, we must balance the numbers of atoms in a chemical reaction. Looking at the reaction above, the methane molecule contains 4 hydrogen atoms – what happens to them? They end up as part of water molecules. Yet, each water molecule contains only 2 hydrogen atoms, so to account for the hydrogens, it must be that two H_2O molecules are formed for each CH_4 that reacts. We should change the reaction to read

$$CH_4 + O_2 \rightarrow CO_2 + 2\ H_2O$$

So far, so good! We've now indicated that one molecule of CH_4 gives two molecules of H_2O. That takes care of balancing the hydrogen, but what about the other atoms? Inspection shows one carbon atom on each side, so the carbon is balanced, but the oxygen is not. As the expression stands, there are two O atoms on the left and four on the right. The oxygens can be balanced by using two O_2 molecules on the left:

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

We could also add subscripts to indicate whether the species were gas (g), liquid (l), or solid (s), and write, for example,

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$$

Now we have a balanced chemical equation that tells us that one CH_4 molecule combines with two O_2 molecules to produce one molecule of CO_2 and two molecules of H_2O. On a mole scale, we can say that one mole of CH_4 would react with two moles of O_2 to give one mole of CO_2 and two moles of H_2O.

Example 1.

Propane, C_3H_8, is commonly used as a fuel source. It burns (reacts with O_2) to produce CO_2 and H_2O. How many moles of O_2 are required to react with 0.50 mole of C_3H_8?

Solution.

The reaction $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ may be balanced by inspection to give the equation

$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$