CHAPTER 1. CHEMISTRY AND PROBLEM SOLVING

Chemists as Problem Solvers

Chemistry is a science for those who want to explore the mysteries of matter. It is also a science for problem-solvers. Each day chemists use their skills to answer questions that may be of critical importance. Sometimes the answers must be found quickly: Is the level of medication in the patient's blood within safe levels? Is the city's water supply safe to drink? Other problems require long-term work: How do chemical substances cause the warming of the atmosphere? Can drugs be discovered that will cure cancer or AIDS? How do substances behave at very low temperatures?

What Are the Problems?

Today our society faces some serious problems. Do we have enough oil to meet the needs of our economy? If not, can we find acceptable substitutes? How can we solve the enormous problems created by the waste we generate: toxic waste, nuclear waste, and the tons of solid waste that have exhausted the capacity of our landfills? Are the food we eat and the consumer products we use really safe? These are only a few of the contemporary issues that are directly related to chemistry. Chemists deal every day with problems such as these, but chemists are not alone in facing them. Policy makers in industry and in government at the national, state, and local level must make decisions on these issues. Voters must make decisions at the polls, where, increasingly, initiatives based on environmental issues are becoming as numerous as the candidates on the ballot. Each visit to the grocery store or pharmacy involves consumer choices between chemical products. Do food products contain harmful chemicals? Are cosmetics safe to use? Does it make a difference which headache medicine we choose?

Technology has made life so complex that every day we need some chemical knowledge to make intelligent choices. Mastering chemical concepts and chemical skills can be very satisfying, for chemistry is truly an everyday science with immediate applications for solving problems in everyone's life.

Problem-Solving Tools: Concepts, Facts and Names, Skills

Chemistry is unlike many other fields of study in that it involves several different types of learning: not just learning ideas as in a philosophy course, or a vocabulary as in a language course, or skills as in a math course (or as in a sport), but all of these. A successful chemistry student must have successful learning strategies for different kinds of learning, and must know when to apply them.

This text is designed to give the non-scientist the problem-solving tools of chemistry to deal with contemporary problems. These tools are of three types: **concepts, facts and names, and skills**. When learning any subject, it is important to distinguish among concepts, facts, and skills. A concept offers a new way of approaching a problem or viewing an issue. Once understood, a concept is seldom forgotten. Facts and names can be memorized by regular contact with them, as with learning a language, or looked up in appropriate sources. Problem-solving skills, like any skills, usually involve practice over time for best results. What would happen if a basketball player prepared for a big game by staying up all night the night before the game shooting baskets? The result would be disastrous. Yet many students, not understanding that solving a certain type of problem involves acquiring a skill, make a similar mistake in their study patterns.

Tiger Woods
2008 Buick Invitational
Torrey Pines, California
1/27/2008
Paul Gallegos / PR Photos

Figure 1-1: ***Even the best players know that a sport requires practice, because it involves skill.***

Scientists are only beginning to learn why the brain functions the way it does in processing information and acquiring skills. Successful students know how to master these three basic types of learning by using the appropriate study technique for the type of problem. To learn a concept, read about it, listen to lectures, but also involve yourself by using the idea. Discuss the concept and apply it in different situations. To learn a name or a fact, try memory aids like flash cards and lists. To learn a skill, repetition of the skill over time is best. Problem-solving skills should become faster and easier with each effort, if they are repeated several times in the course of a week.

At the end of each chapter are listed concepts to understand, facts and names to learn, and skills to practice from that chapter. Sometimes a chapter will feature concepts with very few names to learn or problem-solving skills to master. Others will emphasize problem-solving or names of chemicals to recognize. To learn all the material in the chapter, make sure you use each check list, with appropriate study techniques for each type of learning. Read and discuss concepts until you understand them; spend time going over names facts until you are familiar with them; and practice skills until you are able to use them with confidence.

Chemical Concepts

Learning chemical concepts is the most important objective of a chemistry course. Chemical concepts provide insights that apply to many types of situations. Perhaps the greatest value of a liberal arts education lies in such insights, which involve not just facts, but a new way of looking at the world.

For instance, the idea that all matter is made up of building blocks from a very limited list of substances called the elements is a powerful concept. It enables us to look at the vast number of substances that make up the universe in an organized and understandable way. The concept of the elements has applications far beyond the science of chemistry. For instance, if making new substances is simply a matter of rearranging these few indestructible elements, it follows also that we are limited to the amounts of each element we have to start with. If we are limited to the set of building blocks found on this planet, then the distribution of the elements over the earth's surface has enormous economic and political consequences. South Africa's geologic resources of rare elements, for instance, are a source of wealth and political leverage for that country. Uranium, used for both nuclear power and nuclear weapons, is another example of an element that is distributed very unevenly over the earth's crust. When the elements are presented in a chapter of the text, merely to memorize the definition of an element would be to miss entirely the point that not just a fact, but an important concept is involved.

Risk assessment is currently an important area of controversy. Individuals, corporations, and government agencies all must struggle with choices about the acceptable degree of risk. We need facts, from research and other sources, to determine the effects on people and the environment of the many substances we are putting into our food, air, and water. But, even more importantly, we must master basic concepts to interpret the overwhelming mass of factual information.

In this book, understanding chemical concepts and being able to apply them to today's problems are the most important learning objectives. Each chapter begins with a summary of the important chemical concepts that will be presented, along with some of the contemporary problems to which they apply.

Chemical Facts and Chemical Names

Chemical facts can be fascinating. Few people know that arsenic, a favorite poison in murder mysteries, is actually an essential nutrient of the body --in extremely small amounts!-- or that some expensive cleaning products can be duplicated for pennies a bottle. Lead products have been eliminated from gasoline, and lead paint is removed from the home environment, sometimes at great expense. Yet few people realize that the "hair color for men" in the home medicine cabinet contains a lead compound as its active ingredient.

Some chemical names are so useful that they are worth memorizing. For instance, if ascorbic acid is listed on a label, what is it? Is it a helpful or a harmful substance? Some people avoid sugar in their foods. How should they react if dextrose and fructose are listed on a label? Simply knowing some chemical names is useful information, especially when some basic chemical concepts about the substances are known as well. Do nuclear power plants employ nuclear fission or nuclear fusion? Are there dangers to the area near a nuclear plant, and if so, what are they? Knowing the facts is important when your safety and peace of mind are involved.

It is impossible for anyone to learn all the chemical facts and names that will be of use in solving all of the chemistry-related problems that might arise. Learning where to find the necessary information is an important skill. In this text, for instance, the periodic table of the elements is given, and it will be an increasingly useful source of information as we learn its organization. Important chemical names and facts are included throughout the text. Sometimes science reference books or consumer guides are appropriate information sources. Government agencies or poison information hotlines can be useful sources of chemical facts. The science sections of newspapers and magazines often provide up-to-date information on technical and environmental issues.

Learning to use the chemical facts in this text can be a first step in gaining the confidence to find and use the chemical facts that you need in everyday life. Additional information sources are listed at the ends of chapters.

Decoding Skills: Using Metric Units

Every laboratory in the world uses the same system of measurement: the metric system, or International system (abbreviated SI from Systeme Internationale). Though the United States has not officially adopted metric units, soft drink bottles are labeled in liters and races are run in distances of meters and kilometers. The reason for the increasing use of the metric system in the United States is simple: the United States is the last country in the industrialized world that has not officially adopted this system. It would be inefficient for international corporations to manufacture two sets of products, one for this country and one for all others. Athletes want to perform in internationally accepted events. Consequently, metric units once unfamiliar are seen more and more often in the U.S. The result has been a rather confusing system in which the supermarket sells soft drinks by the liter and milk by the quart or gallon. Unlike the international soft drink companies, dairy companies are usually local, and have had no incentive to change container size. (Of course, in all other countries, milk, like other liquids, is sold by the liter.) Another example of international trade is the illegal drug trade; hence, drug traffickers are likely to use metric units.

Coke is sold by the liter, but milk is sold by the quart.

Scientists have used metric units routinely for a long time for similar reasons. The scientific community is international, and experimental discoveries have worldwide impact. In this course you will use the metric system in calculations. You may appreciate the advantages of this skill when the United States finally changes to the "new" system.

A good way to begin the study of the metric system is to look at the three basic units of measure: **the meter, the gram, and the liter,** used to measure length, mass, and volume, respectively. Most laboratories are equipped with a meter stick for measuring length; it looks like the yard stick still common in the United States. While the yard stick is 36 inches long, the meter stick measures 39.37 inches. The gram is a rather small unit of weight; a nickel weighs about 5 grams. Often a laboratory will have weights of one, ten, or one hundred grams used for calibrating balances. The liter is perhaps the most familiar of these three units, since soft drinks are sold in one-liter and two-liter sizes.

If you are not familiar with the meter, the liter, and the gram, it is worth taking the time actually to look over an example of each of them. Like many of the topics we deal with in chemistry, these units of measure involve not just abstract ideas, but actual materials you can become familiar with by looking them over. Notice carefully examples you see in the lecture or laboratory. What does one gram of salt or one hundred grams of salt look like? Being familiar with the approximate sizes of sample associated with metric weights can be helpful in estimating the answers to problems. **Having a concrete, common-sense knowledge base and using it to check the reasonableness of your calculations is an important problem-solving skill**.

Using the meter, the liter, and the gram, we can easily scale up or down to larger or smaller units in the metric system, because in this system all units are related by powers of ten. Table 1-1 shows, for example, that one meter equals 1000 millimeters or 100 centimeters. Looking at a meter stick with its thousand tiny markings denoting millimeters, it is easy to associate the prefix "milli-" with a very small unit, and natural to expect a larger number when converting meters to millimeters. Just as there are 1000 millimeters in a meter, there are 1000 milliliters in a liter and 1000 milligrams in a gram. Scaling up to larger numbers, 1000 meters equal a kilometer, and 1000 grams a kilogram. The metric system is easy to use because there are very few new names to memorize or complicated calculations to perform; once you know the common prefixes and the powers of ten they represent, you know the basics of a powerful measurement system capable of representing very large or very small numbers. Table 1-2 lists some prefixes that are used to form units in the metric system. Some of them are used more frequently than others. The centimeter is frequently used as a unit of measure; the centiliter and centigram are equally respectable units, but less popular in use. Dekameters and hectometers are even rarer in use, and for the sake of simplicity such prefixes as giga- (meaning "billion") and pico- (meaning 10^{-12}, or "one trillionth") have been omitted from the table, though such terms are found useful by those who must routinely deal with very large or very small numbers.

Table 1-1: Some Basic Units of the Metric System

Length	*Mass*	*Volume*
1 meter = 1000 millimeters	1 gram = 1000 milligrams	1 liter = 1000 milliliters
1 meter = 100 centimeters		
1000 meters = 1 kilometer	1000 grams = 1 kilogram	

or, to look at it another way,

1 millimeter = 0.001 meter	1 milligram = 0.001 gram	1 milliliter = 0.001 liter
1 centimeter = 0.01 meter		

1 meter = 0.001 kilometer 1 gram = 0.001 kilogram

Fig. 1-3. 1 cm^3 is the same volume as 1 mL. 1 cm^3 (or 1 mL) of water weighs about 1 g.

Table 1-2: Prefixes and Abbreviations Used in the Metric System

nano-	= 0.000000001	= 10^{-9}	nanometer	=	nm	=	10^{-9} m
micro-	= 0.000001	= 10^{-6}	micrometer	=	μm	=	10^{-6} m
milli-	= 0.001	= 10^{-3}	millimeter	=	mm	=	10^{-3} m
			milligram	=	mg	=	10^{-3} g
			milliliter	=	mL	=	10^{-3} L
centi-	= 0.01	= 10^{-2}	centimeter	=	cm	=	10^{-2} m
			centigram	=	cg	=	10^{-2} g
deci-	= 0.1	= 10^{-1}	decimeter	=	dm	=	10^{-1} m
			decigram	=	dg	=	10^{-1} g
deca-	= 10	= 10^{1}	decameter	=	dm	=	10^{1} m
kilo-	= 1000	= 10^{3}	kilometer	=	km	=	10^{3} m
			kilogram	=	kg	=	10^{3} g
mega-	= 1,000,000	= 10^{6}					

DECODING SKILLS: METRIC UNITS

Problem example 1-1. Which radio station operates at a higher frequency, an AM radio station that appears on the dial at a frequency of 103 x 10 kilohertz or an FM radio station at 98.5 megahertz?

From Table 1.2, we see that *mega-* means 1,000,000 or 10^6, while *kilo-* means 1000 or 10^3. So, even though the AM number is ten times higher (depending on how the radio dial is labelled, this might also appear as 1030), the FM station has the higher frequency since the megahertz unit is 10^3 or on thousand times greater than the kilohertz unit.

One cubic centimeter (1 cm^3) measures exactly the same volume as 1 mL, so these are interchangeable units of volume. One mL (or 1 cm^3) of water weighs about 1 g. This is an approximate relationship, since temperature slightly affects the volume of 1 g of water; only at 3.98 degrees Celsius is the weight of 1 ml exactly 1 g. Nevertheless, since water is such a common substance, this relationship is often convenient to use in calculations.

One subject that must be mentioned in a scientific discussion of measurement is the relationship between weight and mass. Even in the most sophisticated scientific laboratories, people often talk about "weighing" a sample or an object, when most likely they are more interested in its measuring its mass. **Weight** refers to the gravitational force exerted on an object; **mass** refers to the amount of the substance that is present. For example, when traveling on the space shuttle, passengers become weightless because of the absence of gravitational pull. Regardless of their "weight loss", however, they wear the same size clothing, because their mass has not changed. Using the gram balance in a scientific laboratory, then, might be referred to as taking the mass of a substance rather than its weight, since the amount of the substance is what we are really interested in. On the other hand, we are using the earth's gravitational pull to perform our measurement, so we are measuring weight as well. And the term "weight" is more familiar to most people than "mass." References to weight, then, will not be banished from this text, though there will be situations in which a concept can be described correctly only by referring to mass.

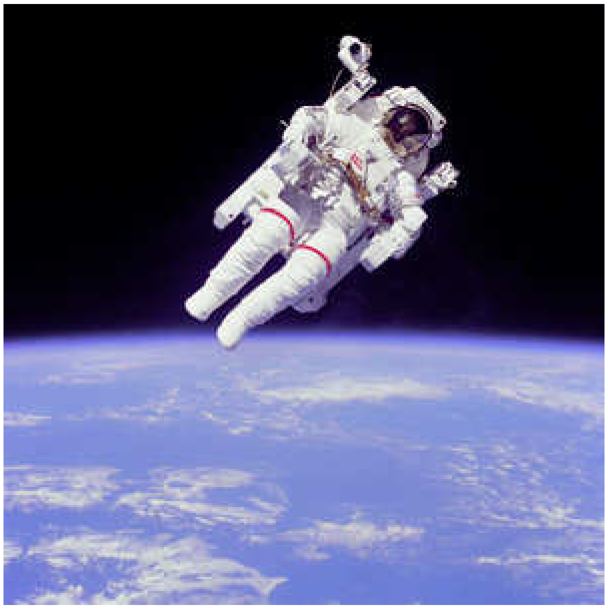

Fig.1-4. In space, matter becomes weightless because of the lack of gravity, but has the same mass as on earth.

Density is a very useful quantitative concept, not to be confused with mass or weight. The distinction is perhaps best made in the very old riddle, "Which is heavier, a pound of feathers or a pound of lead?" The answer, of course, is that they weigh exactly the same: one pound. They differ greatly, however, in the volume they occupy. A pound of lead can easily be held in one hand, while one could fill a small room full of feathers (a daunting prospect!) and still have less than a pound of feathers. In comparing two substances, one must compare the mass of a unit of volume. This is the definition of **density**: mass per unit volume, or the mass of a substance divided by its volume.

$$density = \frac{mass}{volume} \tag{1-1}$$

Any appropriate units of mass and volume will do in calculating density. Normally, densities for solids are expressed as g/cm^3, liquids as g/mL, and gases as g/L. If the densities of two substances are being compared, the units must be the same.

Fig. 1-5. A pound of feathers weighs just the same as a pound of lead. We say that lead is "heavier" because it has a higher density. The same mass of feathers occupies a much greater volume than the same amount of lead.

Problem-Solving Strategies: The Importance of Labelling Units

An important strategy in solving any type of problem is to be sure not to neglect any information that is given. If a given value in a problem is "5.0 grams," for example, it would be a mistake to calculate an answer by writing that value simply as "5.0," for the unit label can be an important tool in solving the problem. In solving for density as in problem example 1-1, using the unit labels to help you in identifying the values of mass and volume, then using the definition of density as mass per unit volume, or mass divided by volume, makes finding density a simple matter.

Notice that whenever the word *per* is used as in miles *per* gallon or grams *per* milliliter, that the first term is *divided by* the second.

Problem example 1-2: Diamond and coal are both made of carbon. If a lump of coal with a volume of 10 cubic centimeters weighs 12 grams, and a diamond with a volume of 0.10 cubic centimeters weighs 0.30 grams, what are the densities of the two substances? Which substance is more dense?

Using Eqn. 1-1, we see that density is defined as mass divided by volume:

$$density = \frac{mass}{volume} \tag{1-1}$$

In this problem we are given values for both the mass and the volume of each substance. For coal, the mass is given as 12 grams, and volume is given as 10 cubic centimeters. Dividing the value given for mass by the value given for volume, For coal,

$$\frac{12g}{10\,cm^3} = 1.2\frac{g}{cm^3}$$

gives us the density, in units of grams per cubic centimeter.

. For diamond,

$$\frac{0.30g}{0.10\,cm^3} = 3.0\frac{g}{cm^3}$$

very different sizes and masses, calculating their densities allows us to compare the two, finding that the diamond has a much higher density than the coal. Moviegoers may recall the film in which Superman changed a lump of coal into a diamond by compressing it in his hands, thus producing the higher density form of the substance. Most viewers probably did not realize he was demonstrating his knowledge of chemistry. He was not alone in using this technique of making diamonds by reproducing the conditions of extremely high temperature and pressure which form diamonds deep within the earth. GE has been producing small synthetic diamonds for industrial use since 1958. Other carbon-containing substances are also possible sources of diamond, as demonstrated by Robert H. Wentorf Jr. in 1955 when he made diamonds from peanut butter.

Fig. 1-6. Both diamond and coal are forms of the element carbon, but diamond has a greater density.

Problem-Solving Skills: Using Exponential Numbers

Although chemistry involves calculations, the mathematics involved in an introductory chemistry course is not usually of a high level. In this course, mathematical operations are limited to addition and subtraction, multiplication and division. The use of calculators has eliminated the more tedious aspects of calculation. Some students may be inexperienced with the handling of very large and very small numbers, which are written in exponential form, sometimes called scientific notation. In scientific notation, numbers are written as the product of a number between 1 and 10 and a whole-number power of ten. For example, rather than writing that there are 602,000,000,000,000,000,000,000 atoms of carbon in 12 grams of carbon, it is much more convenient to write the number in scientific notation, as 6.02×10^{23}. Notice that for this very large number the exponent of ten is a positive number which corresponds to the number of decimal places to the right of the 6. To express a very small number like 0.0000005 in scientific notation, move the decimal to the right until a number between 1 and 10 is obtained; the number of decimal places appears as the negative power of ten, and the number becomes 5×10^{-7}.

Problem example 1-3: Americans generate approximately 600,000,000 tons of hazardous waste and wastewater each year. Express this number in scientific notation.

$$600{,}000{,}000 \text{ tons} = 6 \times 10^{8} \text{ tons}$$

To multiply numbers written in scientific notation, exponents are added. To divide, exponents are subtracted. Multiply and divide the numbers between 1 and 10, called the coefficients, in the usual way.

Problem example 1-4:

$$\frac{(4 \times 10^{15}) \times (5 \times 10^{-1})}{2 \times 10^{3}} = 10 \times 10^{11} = 1 \times 10^{12}$$

To add or subtract numbers written in scientific notation, first make sure that the exponents of the numbers are the same. Then add or subtract as usual, retaining the exponents.

Problem example 1-5:

$2.0 \times 10^{-2} + 5.0 \times 10^{-1}$

$= 2.0 \times 10^{-2} + 50 \times 10^{-2}$

$= 52 \times 10^{-2}$

$= 5.2 \times 10^{-1}$

Problem-Solving Skills: Significant Figures

Today's computers and calculators can produce calculation results with ease. We must rely on ourselves, however, to make judgments about the meaningfulness of these results. The importance of using a commonsense knowledge base to evaluate calculation results has already been mentioned. Another issue in dealing with numbers is their exactness. If your calculator gives an answer of 7.958242, does that really mean you know the result to the millionth decimal place? There are some simple rules that serve as guidelines.

When adding or subtracting, simply aligning the decimal points shows whether some of the numbers being used are more exact then others. A weight of 10.01 g, for example, is more accurate than a weight of 7.2 g; the first has been weighed to a hundredth of a gram, the second to a tenth of a gram. When added together, they will give a result accurate to only a tenth of a gram. Similarly, a weight of 10.00 g is more accurate than a weight of 10.0 g.

Problem example 1-6: What is the total weight of 10.01 g added to 7.2 g?

10.01 g
+7.2 g
17.2 g

Problem example 1-7: What is the total weight of 10.07 g added to 7.2 g?

10.07 g
+7.2 g
17.3 g

Notice that 10.07 is closer to 10.1 than 10.0, so that the final answer rounds to 17.3 g..

When multiplying or dividing, evaluating the number of figures that should appear in the result is done by employing the rule that there are only as many significant figures in the final result as there are in the least significant number in the calculation. The commonsense basis of this rule is simple: the final answer is no better than the least exact number used to calculate it. It is well to keep this rule in mind whenever any figures are given in support of an argument. All figures carry with them some degree of uncertainty; failing to mention that uncertainty is a common failing.

Rounding off the large number of figures produced on a calculator will not only simplify calculations, but may also give you the beginnings of a healthy skepticism about "exact" numbers that individuals, corporations, or government agencies may use to support their arguments or positions. **Ask where the numbers come from that are used to calculate the final result, and ask how precise those numbers are.** Sometimes it is possible that using one imprecise number in a very precise calculation may give a very incorrect final result, even if all the other numbers are precise.

Problem example 1-8: Find the number of significant figures for each number below. Do not count any zeros that occur before the first digit of the number, as these are only place markers to show the position of the decimal point. For the same reason, the numbers in exponents do not count as significant figures.

a. 4.7	2 significant figures
b. 4.70	3 significant figures
c. 4.700	4 significant figures
d. 0.47	2 significant figures
e. 0.047	2 significant figures
f. 4.7×10^{-8}	2 significant figures

Problem example 1-9: Find the solution to the following numerical problem, using the correct number of significant figures to express the answer.

GETTING COMFORTABLE WITH YOUR CALCULATOR

Calculators can differ considerably in the ways they present numbers; it is best to consult the instructions that come with the calculator you are using. The calculator used in this course need not be an expensive or complex one. To be comfortable in operating the calculator, it will be necessary to know:
Can the calculator show exponents? If so, how do they appear?
What operations are used to add, subtract, multiply, and divide? These are the only mathematical operations you will need to perform in this course.
Calculators almost never know about significant figures! If the calculator gives an answer of 7.77777777, it will be up to you to determine how many of those numbers are "real."

$$\frac{(2.1 \times 10^{-5}) \times (3.48 \times 10^{3})}{1.635} = 4.5 \times 10^{-2}$$

Problem-Solving Skills: Analyzing the Problem

Some problem-solving procedures are designed specifically for solving chemistry problems. But many problem-solving techniques are applicable to a wide variety of problems, including those encountered in daily life. A good general problem-solving technique is to start by analyzing the problem. First, identify clearly what information has been given in the problem. No matter how simple that information might be, it should be written down in an organized manner, making sure to include all parts of the information, such as the unit or identifying label of each number given. Then the objective of the problem should be written down, again indicating any units or labels of an expected number. Once the starting point of the problem and the desired final objective have been clearly defined, it becomes much easier to solve the problem. Solving the problem becomes a matter of finding a way from "where you are" to "where you want to be." This strategy is not only for chemistry problems; it has been recommended by experts in fields ranging from financial planning to career counselling. No matter what your goal, it's a good starting point actually to write down exactly what you want to have and what you have to start with.

ANALYZING THE PROBLEM

1. What do you have to start with? Write it down!

2. What do you want to end up with? Write it down!

3. How do you get from what you have to what you want? Form a strategy!

WHAT YOU HAVE ->->->->->->->->->-> WHAT YOU WANT

Forming a strategy to solve a problem can be relatively obvious and straightforward, or it may involve several steps. If you have carefully written down all the units for the numbers in a numerical problem, sometimes the strategy becomes obvious after an examination of the units of the starting point and the desired answer. For example, recall that if you know that the units of density are grams per cubic centimeter, if you know that "per" means "divided by," and you have a mass value in grams and a volume in cubic centimeters, it becomes a straightforward strategy to divide grams by cubic centimeters to get the answer. Whenever possible it helps to look at problems that have already been solved by someone else to see how the problem-solving skills have been applied. This text includes solved sample problems as part of the chapter discussion; they will be useful guides in solving problems found at the end of the chapters. Simple examples of analyzing a problem can be found in the dimensional analysis problems which follow.

Problem-Solving Skills: Using Conversion Factors

Converting numbers from one type of unit to another is a type of problem that often appears in chemistry; frequently numbers must be converted to appropriate units as a first step to solving a larger problem. This relatively simple type of problem is a good example of a situation in which analysis of the problem results in a straightforward solution. In this case **what we are given** is the starting number and its unit label. **What we want** is a number with the desired units. The route to get from what we are given to what we want involves setting up a **conversion factor** using these units. This method is often called **dimensional analysis**.

Problem example 1-10: How many millimeters in 0.40 meters?

First analyze the starting point and the objective of the problem:

WE HAVE	*WE WANT*
0.40 meters	*millimeters*

Clearly, the conversion factor will involve both meters and millimeters. We know the relationship between them: one meter is equal to 1000 millimeters. Because we want our answer to be in millimeters, millimeters should appear in the <u>top</u> of the conversion factor. Meters will appear in the <u>bottom</u> of the conversion factor, where they will cancel out the units in our given information, leaving us with final units of millimeters.

WE ARE GIVEN		*CONVERSION FACTOR*		*WE WANT*

$$0.400m \quad x \quad \frac{1000mm}{m} \quad = \quad 400mm$$

From your experience with the relative size of these units, does this seem like a reasonable answer?

Problem example 1-11: How many centimeters in 400 millimeters?

WE ARE GIVEN	*WE WANT*
400 millimeters	*centimeters*

Table 1 does not give us conversion units from millimeters to centimeters. It does give us information about both millimeters and centimeters: 1 meter equals 1000 millimeters, and 1 meter equals 100 centimeters. Construct two conversion factors using these relationships, making sure that the desired unit, centimeters, is on <u>top</u> of its factor and the given unit, millimeters, is on the <u>bottom</u> of its factor where it will be canceled out. Notice how meters

$$400mm \; x \; \frac{1m}{1000mm} \; x \; \frac{100cm}{1m} = 40cm$$

cancel out when the conversion factors are set up properly.

From your experience with the relative size of these units, does this seem like a reasonable answer? The relationships in this problem are simple enough that the final answer may be intuitively obvious to you if you are familiar with the use of a meter stick.

PARTS PER MILLION

Often very low concentrations of substances are of interest to us; even a small amount of lead in a child’s bloodstream can be a cause for concern, for example. The sophisticated instruments available to chemists today make possible the measurement of concentrations as low as one part in a million, or considerably lower, and environmental standards are often written in terms of parts per million. Parts per million and milligrams per liter are frequently used interchangeably as units. Here’s why: A milligram is a thousandth of a gram, and a liter of water is equivalent to 1000 milliliters, or about 1000 grams. Therefore, a milligram per liter is equivalent to 10^{-3} g/10^{3}g or $1/10^{6}$; that is, one part in a million.

DECODING SKILLS: METRIC UNITS

The Sodium Content of Chicken Noodle Soup

Figure 1-7 .It's important to check the sodium content of canned products.

Problem example 1-12: The label of a can of vegetable soup indicates that 890 mg of sodium is present in 1/2 cup of the condensed soup. How many grams of sodium are present in one cup of the condensed soup? (After dilution this would make a 2-cup serving of soup).

One cup of the condensed soup contains twice as much as the 1/2-cup serving, or 1840 mg. To convert to grams,

WE HAVE	***WE WANT***
1780 mg	g

Costructing a conversion factor with the desired unit on top,

$$1780 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} = \text{ g}$$

Solving the numerical problem,

$$1780 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 1.780 \text{ g}$$

In order to decode the meaning of the soup can label we will need to learn some chemistry as well as having some familiarity with the metric system. Actually, the sodium is not present in the soup in the form of the pure element sodium. (This is fortunate for the consumer, since elemental sodium explodes on contact with water!) We learn from the label that most of the sodium comes from "salt," or sodium chloride, a compound of the elements sodium and chlorine. We will learn in Chapter 2 about elements and compounds, and how to use chemical names and chemical concepts to decode product labels. For now, we will simply point out that to determine the amount of salt that must have been added, we must take into account the presence of both sodium and chloride in the table salt. If we do so, we find that the amount of table salt in the soup serving is actually 4.53 grams! The picture below shows 4.68 grams of sodium chloride next to the two-cup serving of soup; as a relative measure of size, the soup spoon near the salt holds about one tablespoon. *(Fig. 1-7a. Soup, spoon, 4.68 grams sodium chloride)*

Problem-Solving Skills: Visualizing Molecules

Mathematical problem-solving skills are only one type of skill that may be required to solve problems in chemistry. Few chemists spend all their time performing mathematical calculations. Many chemists, of course, spend a significant amount of time in the laboratory, using a variety of laboratory skills. Another, very different, skill is the ability to visualize the shapes and structures of molecules, the infinitesimally small units of which matter is made. After all, molecules are much too small for us to be able to see them with our eyes. Only very recently and with the most sophisticated technology have we been able to "see" the first fuzzy pictures of atoms and molecules. Chemists have had to use the information they had about the properties of molecules and then guess what the molecular structures and shapes must be like.

Some of the most fascinating problems in chemistry have involved figuring out the structures of molecules. In Chapter 10, for example, is the famous story of how the German chemist Kekule first visualized the structure of the benzene molecule in 1865. More recently, finding the structure of the DNA molecule was a problem that engaged the minds of brilliant chemists; among the skills needed to solve this problem was the ability to visualize a three-dimensional form for this highly complex molecule.

Chemists use three-dimensional aids called **molecular models** to help them to understand the shapes of molecules. Molecular models can vary from the simplest to the most complex. Elaborate representations of the shapes of molecules can now be generated with the aid of computers; but very simple, home-made models have played an important role in scientific discovery. Linus Pauling, the Nobel-prize-winning chemist, explains that he made his landmark discovery of an important protein structure while in bed with a cold. Writing the structure of the long-chain molecule on a sheet of paper, he found that the bonding patterns in the molecule made sense if he twisted the paper into a spiral, forming the now-famous helical structure.

To solve the problem of finding the structure of DNA, the important molecule which determines heredity, Francis Crick and James Watson built a large scale model, using the X-ray data of Maurice Wilkins to determine the distances between atoms and the angles at which they were connected. They received the Nobel Prize in 1962 for solving this structural problem, which they described as a "three-dimensional jigsaw puzzle." Crick and Watson's model was constructed out of metal with the help of a machine shop. At one point, though, James Watson was so impatient with the wait for metal components that he made his own out of cardboard. In the popular book "The Double Helix" Watson describes the efforts of his research group to solve the structural puzzle of DNA, and the intense rivalry with Pauling's group to be the first to do so.

James Watson with Francis Crick in 1953, the year of their discovery of the DNA molecule. (UPI/Bettman)

Richard Smalley, professor of chemistry and physics at Rice University, while trying in 1985 to solve the puzzle of finding the shape of a symmetrical molecule with 60 carbons, used paper, scissors, and tape to make an interlocking structure made of 12 pentagons and 20 hexagons. The structure, which fitted perfectly with the known information about the molecule and was very stable, turned out to be identical to that of a soccer ball!

Nobel laureate Richard Smalley with a nanotube model
Photographs by Mark Thiessen

He was weakened by cancer, but Nobel laureate Richard Smalley was still eager to pull on rubber boots, wade into water with a light tube, and pose before a model of a molecule he discovered—all to make his point: Nanotechnology can help us. In 1985 Smalley and his team discovered a remarkably strong molecule made of 60 carbon atoms and resembling a geodesic dome; they dubbed it a buckyball (after Buckminster Fuller, who invented the dome).

His collaborator, Harold Kroto, of the University of Sussex, England, was less successful in finding the correct structure, using toothpicks and Juicy Fruit gumballs. The molecule, which they named buckminsterfullerene in honor of the famous architect of geodesic domes, is now the subject of intense research because of its unique properties.

How can we visualize the shapes of molecules that we cannot see with the naked eye? Just as with mathematical ability, this ability to visualize a three-dimensional object seems to come more naturally to some people than others. Indeed, one of the interesting aspects of a chemistry course is the opportunity both to test your problem-solving abilities of different types and to develop them further. Even the most brilliant chemists, as we have seen, need molecular models to help them "see" complex molcular structures. If you have access to molecular models, you will find them an invaluable aid. Some important concepts in chemistry cannot be fully understood unless the shapes of molecules are clearly visualized.

Problem-Solving Skills: Applying the Scientific Method

Science is more than performing calculations and reasoning through the answer to a problem. The heart of science lies in the use of experiments to test ideas. First, a **hypothesis**, or a logical guess about the answer to a question, is formulated. Then an experiment is devised to test the hypothesis. If the results of all experiments performed support the hypothesis, then it becomes a **theory**, or accepted explanation for observed phenomena. Formulating hypotheses and devising experiments present a creative intellectual challenge. Laboratory work requires measurement skills, observational skills, organization, and the honesty to accept disappointing or surprising data. The discovery of the atomic nucleus is only one example of a major scientific advance that occurred through unexpected experimental results. From the earliest days of chemistry to the present high-technology laboratories, chemists have found satisfaction in learning about the substances that form our universe and the changes that they undergo through applying the scientific method.

DECODING SKILLS: LEARNING THE LANGUAGE OF CHEMISTRY

One of the most powerful skills to learn from a chemistry course is how to interpret common chemical names and symbols. Intimidated by unfamiliar terms, consumers often neglect to read labels on foods, vitamins, cosmetics, and cleaning products, only to find later they have chosen unwisely. Manufacturers can mislead consumers with reassuring statements like "contains only naturally-occurring elements" or "contains body-building branched-chain amino acids" - both perfectly true, but not, as we will learn, good reasons to buy a product. In the chapters to come we will learn how to decode chemical names and formulas, often using product examples provided by students in this course.

CONCEPTS TO UNDERSTAND FROM CHAPTER 1

You should be able to explain in your own words the following:

How does weight differ from mass?
How does weight differ from density?
What is the scientific method?
What is chemistry?
What is a molecular model?

FACTS TO LEARN FROM CHAPTER 1

Common metric system units (Table 1)
Metric system prefixes (Table 2)
1 mL equals 1cm^3; 1 mL of water weighs approximately 1 g

SKILLS TO PRACTICE FROM CHAPTER 1

This chapter introduces several kinds of numerical problems; each is represented by an example in the text. In solving problems from the problem set, first identify the type of problem and use the example as a model. When you have mastered a type of problem, you should be able to obtain a solution without consulting the text. After finishing this chapter,
you should be able to:

Convert from one type of metric unit to another using
dimensional analysis
Convert numbers to scientific notation
Perform calculations using scientific notation
Recognize the number of significant figures in a number
Give the correct number of significant figures for a
calculation
Calculate density, given mass and volume

Several general problem-solving skills are introduced in this chapter. When solving problems, remember to:

- Analyze the problem. Define what is given and what is wanted.

- Use dimensional analysis. Label all units. Try to construct a conversion factor that will be a bridge between what is given and what is wanted.

- Check your answer against your knowledge base and your common sense.

- Check significant figures. How exact are the numbers used in the calculation? What effect do they have on the final result?

PROBLEMS TO SOLVE
USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 1

1-1. Identify each of these phrases or statements as referring to a concept, a fact, or a skill.

a. Bowling

b. World peace

c. Your height

d. The boiling point of water under standard conditions is 100°C.

e. When water boils, molecules gain enough energy to escape from the liquid state to the gas state.

f. Using the relationships between the temperature scales, convert Fahrenheit temperature to Celsius temperature.

1-2. Match each of the following with the appropriate unit from the opposite column.

1. Mass	A. Millimeters
2. Length	B. Milliliters
3. Volume	C. Grams/milliliter
4. Density	D. Grams

1-3. Which is larger?

a. 1 km or 1 mm?

b. 1 mm or 1 μm?

c. 1 mL or 1 L?

d. 1 mL or 1 μL?

e. 1 g or 1 mg?
f. 1 mm or 1 nm?

1-4. Give the number of significant figures in each of the following quantities.

a. 2 g

b. 2.01 g

c. 2.00 g

d. 0.2 g

e. 0.20 g

f. 2.0×10^3 g

g. "Americans generate 600,000,000 tons of waste every year."
How many significant figures are implied in this figure?
How many do you think there really are?
What would be the "plus or minus" error figure implied in the number of significant figures that are actually in this number?

1-5. Find the densities of the following substances:

a. A piece of chalk that weighs 10.0 g and has a volume of 4.00 cm^3.

b. A piece of cardboard that weighs 13.8 g and has a volume of 20.0 cm^3.

c. A cork that weighs 10.0 g and has a volume of 40.0 cm^3.

d. A sample of gasoline that weighs 68 g and has a volume of 100 ml.

1-6. Referring to the density values calculated in 1-5,

a. Rank the four substances in order from highest to lowest density.

b. Will a piece of chalk float or sink in gasoline?

c. Will a cork float or sink in gasoline?

d. What will happen when the cardboard is put in the gasoline? Compare the

results
with what you expect to happen with the chalk and with the cork.

1-7. What is the density of a urine sample of 150 mL that has a mass of 157 g?

1-8. Perform the following calculations, expressing the results in scientific notation with the correct number of significant figures.

a. $(2.0 \times 10^{3}) \times (3.54 \times 10^{-2})$

b. $(2.1 \times 10^{3}) + (1.8 \times 10^{2})$

c. $\dfrac{(5.0 \times 10^{9}) \times (3.6 \times 10^{-5})}{4.63 \times 10^{-5}}$

1-9. Perform the following calculations, expressing the results in scientific notation with the correct number of significant figures.

a. $(6.82 \times 10^{-8}) \times (7.3 \times 10^{-2})$

b. $(7.11 \times 10^{-2}) + (6.35 \times 10^{-3})$

c. $\dfrac{(9.8 \times 10^{5}) \times (1.2 \times 10^{-3})}{5.6 \times 10^{2}}$

1-10. Use your problem-solving skills to perform the following unit conversions.

a. 4.2 cm = ? m

b. 4.2 cm = ? mm

c. 4.2 cm = ? km

d. 4.2 cm = ? μm

1-11. Use your problem-solving skills to perform the following unit conversions.

a. 12.5 mL = ? L

b. 12.5 mL = ? cm^3

c. 12.5 mg = ? g

d. 12.5 mg = ? kg

1-12. Use your problem-solving skills to perform the following unit conversions.

a. 53.6 mm = ? m

b. 4860 m = ? km

c. 2582.6 mL = ? L

d. 985 mg = ? g

1-13. *Decoding skills using metric units*: According to the nutritional labelling on a can of peas, a 1/2-cup serving of peas contains 400 mg of sodium. How many grams of sodium are found in one cup of the canned peas?

Fig. 1-10.

1-14. As we will learn in Chapter 2, the tiny building blocks of which matter is made are called atoms. These atoms combine to form molecules, which can be very complex, like DNA, or very simple, like the hydrogen molecule, which is made of two hydrogen atoms joined together. A molecule of natural gas, or methane, is made of one carbon atom attached to four hydrogen atoms. The hydrogen atoms are arranged around the carbon so that each of the hydrogen atoms is totally equivalent in position to the others, making a symmetrical, three-dimensional molecule (not a flat one).

a. Using gumdrops and toothpicks, construct a model of the methane molecule. How will you make it clear that the hydrogen atoms are different from the carbon atom? Compare your construction with the two-dimensional representation of methane in Chapter 10. Name some advantages of gumdrops and toothpicks as molecular models. Name some disadvantages.

b. The fourfold symmetrical arrangement you constructed in (a) is described as tetrahedral symmetry. This symmetry is present as well in the four-sided structure called the tetrahedron. Each of the four sides of the tetrahedron is an equilateral triangle. Construct a tetrahedron using cardboard or stiff paper and tape. To what features of the methane molecule do the points of the tetrahedron correspond?

c. Name some twentieth-century scientists who used model-building to visualize molecular structures as you have just done. What molecules were they trying to understand?

MORE PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 1

1-15. Which is larger?

a. 1 mg or 1 kg?

b. 1 cm or 1 mm?

c. 1 megaHertz or 1 kiloHertz?

d. 1 µm or 1 nm?

e. 1 centigram or 1 milligram?

f. 1 decimeter or 1 dekameter?

1-16. Give the number of significant figures in each of the following quantities.

a. 4.00 g

b. 0.40 g

c. 1.40 g

d. 1.404 g

e. $1.404 \text{ X } 10^{15}$ g

1-17. In attempting to identify fragments of glass from a burglary or a hit-and run scene, forensic scientists measure the density of the glass fragments. The volume of a fragment is determined by the volume of water it displaces in a graduated cylinder or similar device.

a. Find the density of a fragment of glass with a mass of 0.241 g and a volume of 0.90 cm^3.

b. Find the density of a glass fragment with a mass of 0.510 g and a volume of 1.70 cm^3.

c. Is it likely that the two fragments came from the same source?

1-18. Perform the following calculations, expressing the results in scientific notation with the correct number of significant figures.

a. $\dfrac{(3.042 \times 10^{-5}) \times (0.639 \times 10^{1})}{8.455 \times 10^{-3}}$

b. $6.02 \times 10^{23} \times 3.5$

c. $(6.02 \times 10^{3}) + 205$

1-19. Use your problem-solving skills to perform the following unit conversions.

a. 6.00 m = ? km

b. 6.00 m = ? mm

c. 6.00 m = ? cm

d. 6.00 m = ? nm

e. 6.00 m = ? µm

1-20. Use your problem-solving skills to perform the following unit conversions.

a. 78.32 mL = ? L

b. 6.34 mL = ? cm^3

c. 3.7 kg = ? mg

d. 4.32 g = ? mg

e. 3.6 g = ? µg

1-21. *Decoding skills using metric units*: According to the nutritional labelling on the can, a 1/2 cup serving of creamed corn contains 430 mg of sodium, and the can provides 3 servings. How many grams of sodium are in the can of creamed corn?