How to Solve General Chemistry Problems

seventh edition

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Preface

In this seventh edition, we have tried to preserve the self-teaching nature of previous editions, while responding to suggestions from users of the sixth edition. We have rewritten many sections for improved clarity, introduced many new problems and redone many of the existing ones, and modernized certain aspects of chemical form and notation.

Features such as solutions for illustrative problems of almost every type, suggestions or clues for the solution of more difficult problems, and answers to all the problems have been retained.

The sequence of chapters has been changed to make it coincide more closely with the sequence in most popular general chemistry textbooks. Some long chapters have been divided and others restructured. The aim has been to make it easier for today's student to use this problem book as a supplement to any general chemistry textbook.

Among the changes that have been made are the bringing together of the stoichiometry chapters, the placing of the material on thermochemistry close to that on stoichiometry to emphasize their similarities; the placing of gas laws after stoichiometry; and the division of the material on acids and bases into two chapters.

The beginning chemistry student of today will be required to master a broad range of problem solving skills. We believe that many students will need a self-teaching text from which they can obtain a mastery of these skills in an active way. Our goal in preparing the seventh edition has been to make available the examples, problems, and explanations the student will need for success in the study of beginning chemistry. The student only need use the text with diligence.

We particularly wish to thank Prof. Donald F. Gaines (University of Wisconsin at Madison) and Prof. Carl Trindle (University of Virginia) for their useful suggestions, and Prof. Ramesh Agarwal and our other colleagues at Rutgers for their valuable suggestions and criticisms.

Robert S. Boikess *C. Harvey Sorum*

Contents

1

How to Solve a Problem

Every problem you encounter, whether in chemistry or elsewhere, is solved in essentially the same fashion. *First,* you size up the situation or read the problem carefully, and decide *what* you are supposed to do and what you have to do it with. *Second,* having determined what you are supposed to do and what you have to do it with, you figure out *how* to do it. *Third,* you go ahead and *solve it* according to plan. Finally, you look at what you have done. The first two steps represent the *analysis* of the problem. The third step represents the *arithmetical calculations.* Some problems are knottier than others, but they are all solved by these three fundamental steps. The last step is equally important. It will help you to avoid many errors.

To be more specific, when you go about solving any problem in this or any other book or in any test or examination:

1. Read the problem carefully. Note exactly what is given and what is sought. Recognize that most chemistry problems contain more information than is explicitly given. For example, if a problem specifies a given mass of water, it is also specifying a given number of moles, molecules, and atoms. Note any and all special conditions. Be sure you understand the meaning of all terms and units and that you are familiar with all chemical principles relevant to the problem. Every problem in this book is designed to illustrate some principle, some relationship, some law, some definition, or some fact. If you understand the principle, relationship, law, definition, or fact, you should have no difficulty solving the problem. The one big reason, almost the only reason in fact, why students have difficulty with chemistry problems is failure to understand, exactly and well, the relevant chemical principles and the meaning and value of all terms and units that are used in the problem.

2. Plan, in detail, just how the problem is to be solved. Get into the habit of visualizing the entire solution before you execute a single step. Insist on knowing what you are going to do and why you are going to do it. Aim to learn to solve every problem in the most efficient manner; this generally means doing it the shortest way, with the fewest steps.

3. Specify definitely what each number represents and the units in which it is expressed when you actually carry out the mathematical operation of solving the problem. Don't just write

$$\frac{192}{32} = 6$$

Write $$\frac{192 \text{ g sulfur}}{32 \text{ g sulfur per mole of sulfur}} = 6 \text{ mol sulfur}$$

or whatever the case may be. Always divide and multiply the *units* as well as the *numbers*. This procedure is one way to give exactness to your thought process and is a very good way to help avoid errors. You should jot down the unit or units in which your answer is to be expressed as the first step in the actual solution. For instance, if you are solving for the mass of oxygen in 200 g of silver oxide, you should jot down the fact that the answer will be "= . . . . . . g of oxygen." Many problems are worked backward, since you will often first focus your attention on the units in which the answer is to be expressed and then plan the solution with these units in mind.

4. Having solved the problem, examine the answer to see if it is reasonable and sensible. A student who reports that 200 g of silver oxide contains 1380 g of oxygen should know that such an answer is not sensible. Get into the habit of checking the answer to see if it makes sense.

5. If you do not understand how to solve a problem have it explained to you at the very earliest possible time. To be able to solve the later problems you must understand the earlier ones. After a problem has been explained to you, fix the explanation in your mind by working other similar problems at once, or at least within a few hours, while the explanation is still fresh in your mind. Test yourself to be sure that you can apply your understanding of the solution.

2

Units of Measurement

We will assume in this book that every student is familiar, through laboratory experience, with the common units of measure in the metric system and has a fair idea of the volume represented by 1 L, 100 cm^3, and 1 cm^3, the mass represented by 10 g, 100 g, or 1 kg, and the length represented by 760 mm, 10 cm, and 1 m, etc. Also, we will assume that you are familiar with the Celsius (centigrade) thermometer scale.

You should recall that the metric system employs decimal notations in which the prefix *micro-* means one-millionth, *milli-* means one-thousandth, *centi-* means one-hundredth, and *deci-* means one-tenth, while *kilo-* means one thousand times and *mega-* means one million times.

Conversions of metric units (grams, liters, cubic centimeters, centimeters, etc.) to other units (pounds, quarts, inches, feet, etc.) are not often required. The following table will serve where such conversions are called for.

CONVERSION UNITS

1 meter (m) = 10 decimeters (dm) = 100 centimeters (cm)
= 1000 millimeters (mm)
= 1 000 000 micrometers (μm)
= 39.37 inches (in.) = 1.09 yards (yd)
1 kilogram (kg) = 1000 grams (g) = 1 000 000 milligrams (mg)
= 2.2046 pounds (lb)
1 pound (lb) = 453.6 grams (g)

1 liter (L) = 1000 cubic centimeters (cm^3)
= 0.264 U.S. gallons (gal) = 1.06 U.S. quarts (qt)
1 cubic centimeter is the volume of about 20 drops of water
A new U.S. 5-cent piece has a mass of 5 g

THE SI

Over the years many improvements have been introduced into the metric system. The result is a new and improved system of units that is slowly being adopted by the international scientific community. It is called the International System of Units or the SI from the initials of its name in French. Building on the metric system, the SI starts with seven base units, which correspond to a set of independent physical quantities. These base units and their symbols are listed in Table 2-1.

The units for all other physical quantities are called derived units and are combinations of the appropriate base units without any numerical factors. For example the unit for area is the square meter (m^2), derived from the relationship between length and area. The unit for volume is the cubic meter (m^3), from the relationship between length and volume. Such a system of units is called a *coherent system* and has an obvious advantage. We do not need to learn conversion factors. Some other important derived units are given in Table 2-2.

The SI also includes a set of prefixes that are used to form decimal fractions and decimal multiples of the base and derived SI units. Prefixes corresponding to all powers of 10 are not provided. Those smaller than 10^{-2} or larger than 10^2 must all have exponents that are divisible by 3. Some of these prefixes are listed in Table 2-3. Not all features of the SI have been adopted by chemists. In some cases the older systems are much more convenient. In this book we will try to use the SI to the extent consistent with chemical convenience.

CONVERSION OF UNITS

One very common type of scientific calculation is the conversion of the value of a physical quantity given in one unit into a value measured in another unit. The

Table 2-1

Physical Quantity	*Unit*	*Symbol*
Length	meter	m
Mass	kilogram	kg
Time	second	s
Amount	mole	mol
Temperature	kelvin	K
Electric current	ampere	A
Luminous intensity	candela	cd

Table 2-2

Physical Quantity	*Unit*	*Symbol*	*Definition*
Energy	joule	J	kg-m^2/s^2
Force	newton	N	kg-m/s^2 (or J/m)
Pressure	pascal	Pa	kg/m-s^2 (or N/m^2)
Electric charge	coulomb	C	A-s

Table 2-3

Fraction	*Prefix*	*Symbol*	*Multiple*	*Prefix*	*Symbol*
10^{-1}	deci	d	10	deka	da
10^{-2}	centi	c	10^2	hecto	h
10^{-3}	milli	m	10^3	kilo	k
10^{-6}	micro	μ	10^6	mega	M
10^{-9}	nano	n	10^9	giga	G

first step in such a calculation is to find a relationship between the two units. The relationship may merely be a power of 10 that is indicated by the prefix. For example

$$1 \text{ m} = 100 \text{ cm}$$

$$1 \text{ m}^3 = (100 \text{ cm})^3 = 10^6 \text{ cm}^3$$

Or the relationship may require a knowledge of a numerical factor. For example

$$1 \text{ kg} = 2.205 \text{ lb}$$

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ Torr}$$

Conversion factors that are used to perform unit conversions can be obtained by rewriting these types of relationships between units as fractions. For example

$$\frac{1 \text{ m}}{100 \text{ cm}} = 1 \quad \text{or} \quad \frac{100 \text{ cm}}{1 \text{ m}} = 1$$

These two conversion factors are used to convert between meters and centimeters. The choice of conversion factor can be based on the logic of the conversion or on the canceling of units. Thus to convert a length of 3.2 m to centimeters

$$3.2 \cancel{\text{m}} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} = 320 \text{ cm}$$

while to convert a length of 53.7 cm to meters

$$53.7 \cancel{\text{cm}} \times \frac{1 \text{ m}}{100 \cancel{\text{cm}}} = 0.537 \text{ m}$$

The conversion factor used in each case is the one which has the unit of the original measurement in the denominator and the unit of the result in the numerator. As you can see, the original units cancel, leaving the units of the result when the correct conversion factor is chosen.

You should always check such conversions to see if the answer makes sense. Since a centimeter is a smaller length than a meter, there are more centimeters in a given length than meters. Thus the conversion from meters to centimeters should result in a numerically larger answer. As we see, 3.2 m is 320 cm. Also, 53.7 cm is 0.537 m, a numerically smaller answer for the reverse conversion. If you had used the wrong conversion factor, for example,

$$3.2 \text{ m} \times \frac{1 \text{ m}}{100 \text{ cm}} = 0.032 \frac{\text{m}^2}{\text{cm}}$$

not only would the answer be illogical, indicating a smaller number of centimeters than meters in a given length, but the units would not cancel and the result would have units that make no sense in the context of the problem.

Problems

2.1 The mass of a paperback text is 410 g. Express this mass in kilograms and in pounds.

Solution: We start by expressing the relationship between the units of the conversion.

$$1 \text{ kg} = 1000 \text{ g}$$

We use this relationship to derive the necessary conversion factor. The factor is the one that has the unit of the answer in the numerator. In this case it is

$$\frac{1 \text{ kg}}{1000 \text{ g}}$$

The conversion is then carried out

$$410 \cancel{\text{g}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} = 0.410 \text{ kg}$$

Note that when the units are canceled only the desired unit remains. The conversion to units of pounds is carried out in the same way.

$$1 \text{ lb} = 454 \text{ g}$$

$$410 \cancel{\text{g}} \times \frac{1 \text{ lb}}{454 \cancel{\text{g}}} = 0.903 \text{ lb}$$

2.2 Express the height of a 71-in. person in units of feet, meters, and centimeters.

2.3 There are 32 fluid ounces in a U.S. quart. Express the volume of a 750-cm^3 bottle of consumable liquid in units of ounces, quarts, and liters.

2.4 Express a volume of 1 cubic meter in units of liters.

2.5 There are 5280 ft in 1 mile. The highway speed limit is 55 mph. Express this speed in units of kilometers per hour and meters per second.

2.6 A rectangular building lot is 42 m × 35 m. Express its area in units of square feet.

2.7 Find the number of cubic nanometers in 1 km^3.

Solution: Both these units can be related to the meter, which is the base unit, and to the meter cubed, which is the derived unit for volume.

$$1 \text{ m} = 10^9 \text{ nm}, \qquad (1 \text{ m})^3 = (10^9 \text{ nm})^3$$

$$1 \text{ m} = 10^{-3} \text{ km}, \qquad (1 \text{ m})^3 = (10^{-3} \text{ km})^3$$

Setting equal things equal to each other gives

$$(10^9 \text{ nm})^3 = (10^{-3} \text{ km})^3$$

$$10^{27} \text{ nm}^3 = 10^{-9} \text{ km}^3$$

$$10^{36} \text{ nm}^3 = 1 \text{ km}^3$$

INTERCONVERSION OF CELSIUS (CENTIGRADE) AND FAHRENHEIT TEMPERATURE READINGS

The thermometers used in the laboratory are graduated in Celsius degrees, designated by the letter C. You should note that the correct term is *Celsius degrees* or *degrees Celsius* (rather than centigrade degrees or degrees centigrade) in honor of the Swedish scientist Anders Celsius, who devised the scale. The Celsius degree is related to the SI unit of temperature, the kelvin, by the equation

$$T = t + 273$$

where T is the temperature in kelvins (K) and t is the temperature in degrees Celsius (°C). In the SI, the unit of temperature is the kelvin; the term degree is not used. To convert between Celsius and SI we merely add or subtract 273.

Most household thermometers in the United States are graduated in Fahrenheit degrees, designated by the letter F. The fixed points on both the Celsius and Fahrenheit temperature scales are the boiling point and freezing point of water. On the Celsius scale the freezing point of water is 0°C and the boiling point is 100°C; the space between the fixed points is divided into 100 units and the space above 100°C and below 0°C is divided into the same size units. On the Fahrenheit scale the freezing point of water is 32°F and the boiling point is 212°F; the space between the fixed points is divided into 180 units and the space above 212°F and below 32°F is divided into the same size units. Since the space between the freezing point and boiling point of water is divided into 100° on the Celsius scale and 180° on the Fahrenheit scale, it follows that 100 Celsius degrees must represent the same temperature change as 180 Fahrenheit degrees. Thus 1 Celsius degree is equal to 1.8 Fahrenheit degrees; or expressing it in fractional form, 1 Celsius degree is equal to $\frac{9}{5}$ Fahrenheit degrees, and 1 Fahrenheit degree is equal to $\frac{5}{9}$ Celsius degree.

With these facts in mind we see that if we wish to find the Fahrenheit value, F, of a certain number of Celsius degrees, C, we first multiply the Celsius reading by $\frac{9}{5}$; this gives us $\frac{9}{5}$ C. Since the reference temperature (the freezing point of water) on the F scale is 32° above zero we must add 32 to $\frac{9}{5}$ C in order to get the actual reading on the Fahrenheit scale.

$$\text{Fahrenheit temperature} = \tfrac{9}{5}\,\text{Celsius temperature} + 32$$

or

$$F = \tfrac{9}{5}C + 32 \qquad (2\text{-}1)$$

Equation (2-1) can be transposed to the form

$$C = \tfrac{5}{9}(F - 32) \qquad (2\text{-}2)$$

Equation (2-2) tells us that to find the value, in degrees Celsius, of a Fahrenheit temperature, we first subtract 32 from the Fahrenheit temperature (because the Fahrenheit freezing point reference is 32° above zero) and then take $\frac{5}{9}$ of that answer.

The uses of Eqs. (2-1) and (2-2) are illustrated by the following problems.

(a) Convert 144°F to a Celsius reading.

In thinking our way through this problem we note that 144°F is (144 − 32) or 112°F above the freezing point of water. Since 1 Fahrenheit degree is equal to $\frac{5}{9}$ Celsius degree, 112 Fahrenheit degrees must be equal to 112 × $\frac{5}{9}$ or 62.2 Celsius degrees. That means that 144°F is 62.2 Celsius degrees above the freezing point of water. Since the freezing point of water is 0°C, 62.2 Celsius degrees above the freezing point of water is 62.2°C.

(b) Convert 80°C to a Fahrenheit reading.

In thinking our way through this problem we note that 80°C is 80 Celsius degrees above the freezing point of water. Since 1°C equals $\frac{9}{5}$°F, 80°C is equal to $\frac{9}{5}$ × 80 or 144 Fahrenheit degrees above the freezing point of water. But the freezing point of water on the Fahrenheit scale is 32°. Therefore, we must add 32 to our 144 to get the actual Fahrenheit temperature, 176°F.

(c) Convert 100°C to SI.

To convert from Celsius to SI we add 273. Thus 100°C is 373 K.