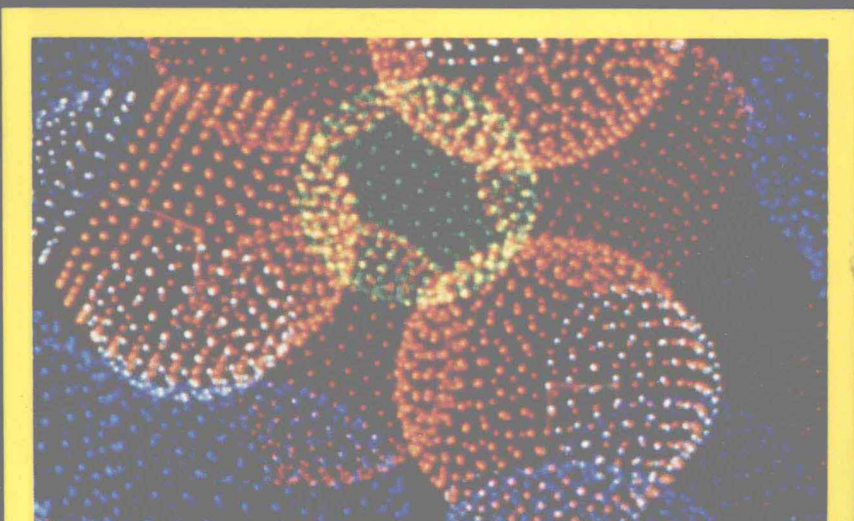
PROBLEM EXERCISES FOR GENERAL CHEMISTRY

THIRD EDITION



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PROBLEM EXERCISES FOR GENERAL CHEMISTRY

Preface

In writing this *third edition*, we have adhered to our objectives of providing a large number and variety of exercises that are pertinent to almost every general chemistry course, in the style used for examinations in large general chemistry programs and on professional school aptitude examinations—for example MCAT and VCAT. The exercises emphasize the importance of problem-solving, equation writing, and the mastery of notation and symbolism that are the "tools of the chemical trade" and fundamental to the understanding of chemistry.

In this edition we have added 'Drill Sheets' throughout the book. These may be assigned as homework to be turned in and graded. Most of the sheets have one or two entries that the professor can furnish data so as to make each homework assignment distinctive from semester to semester. The back sides of these sheets are blank and the professor may ask to show here how a specific problem was calculated, the 'Set-Up'. In addition, we have expanded upon the material in Chapters 7 and 8, Chemical Reactions.

The following order is generally followed within a chapter: Introduction-to-Subsection, Illustrative Examples on the Subsection, Exercises-on-Subsection. This order is repeated with a chapter for any new topic. Each chapter closes with a large collection of exercises covering the topics of all of the subsections. Answers to all of the exercises are given at the end of the chapter.

An acid-base table, along with acid dissociation constants, and a table of reduction potentials have been included in appropriate chapters; a concise section on significant figures and thermodynamic tables have been included as appendices. The thermodynamic tables have been given not only for the convenience in the use of this book, but also because values do vary somewhat from one book to another. Answers to the exercises are consistent with this tabulated data.

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Introductory Concepts

Chemistry is the study of *matter,* the transformations that matter undergoes, and the *energy changes* that accompany these transformations. Such a study requires a wide variety of laboratory measurements. Many of these are similar to ones that you use daily: items purchased at the grocery store may be counted in dozens (measure of *number*) or weighed in pounds or ounces (measure of *mass*); fuel is frequently purchased by the gallon (a measure of *volume*); *linear* measurements are used to express distances between towns (miles) or to indicate the size of a sheet of paper, *e.g.* 8-1/2 × 11 inches. To express a measured quantity, it is necessary to indicate **both** the **numerical value** and an **appropriate unit**. If someone asks a price of 15 (unit unspecified) for an object, it certainly makes a big difference whether the price is 15 dollars or 15 cents. Also, it would be ridiculous to ask the attendant to put 10 acres of gasoline in the tank of an automobile (an inappropriate unit). Indeed, the unit is just as important as the numerical value!

UNITS, CONVERSIONS BETWEEN UNITS, AND SIGNIFICANT FIGURES

Although we readily recognize units such as gallons and feet, they are not commonly used in most of the rest of the world. The system more widely used throughout the world is the *metric system,* a modified form of which is used in the sciences and is called the *Système International d'Unités* (abbreviated, *SI*). In the SI system seven fundamental units (see Table 1) are defined and all measurements are made by using either these units or units derived directly from them.

Numerical quantities are most conveniently expressed when one or two digits is followed by a decimal point and then an appropriate number of digits beyond the decimal point; thus, 62,500 meters is more conve-

TABLE 1. FUNDAMENTAL SI UNITS

Quantity	SI Unit	Symbol	SI/English Conversion Factor
Mass	kilogram*	kg	$\frac{0.4536\ \text{kg}}{1\ \text{pound}}$ or $\frac{453.6\ \text{g}}{1\ \text{pound}}$
Length	meter	m	$\frac{0.02540\ \text{m}}{1\ \text{inch}}$ or $\frac{2.540\ \text{cm}}{1\ \text{inch}}$
Amount	mole	mol	$\frac{\text{number of items}}{6.023 \times 10^{23}\ \text{number of items/mole}}$
Time	second	s	The same in both systems
Temperature	kelvin**	K	See p. 133
Electric Current	ampere	A	The same in both systems
Luminous Intensity	candela	cd	The same in both systems

*For experiments in the laboratory the kilogram usually represents an inconveniently large mass. Laboratory quantities of chemicals are measured in gram quantities; the gram is one-thousandth of a kilogram. **The centigrade (or Celsius) temperature, °C, is also accepted as an alternate to the Kelvin temperature.

TABLE 2. PREFIXES USED AS DECIMAL MULTIPLERS FOR UNITS

Prefix	Abbreviation	Multiplication Factor
exa	E	10^{18}
peta	P	10^{15}
tera	T	10^{12}
giga	G	10^{9}
mega†	M	10^{6}
kilo†	k	10^{3}
deci	d	10^{-1}
centi†	c	10^{-2}
milli†	m	10^{-3}
micro†	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}
atto	a	10^{-18}

†Prefixes that are most commonly used in chemistry and should be memorized by the student.

niently expressed by its equivalent 62.5 km. The SI system uses prefixes to indicate large or small quantities. Hence, one thousandth of a gram, 0.001 g, would be more simply expressed as 1 mg. For purposes of comparison with other data, it may be desirable to express this mass as 1,000 μg. All of these refer to exactly the same mass.

DERIVED UNITS

Upon inspection of Table 1, one notes that no units are specified for the very important quantities *area* and *volume.* This is because suitable units for these can readily be derived by multiplication of fundamental units. Area can be expressed by squaring a linear unit, the meter or the meter with an appropriate prefix. Likewise, volume can be represented by cubing a linear unit; the volume of a cube having 2-cm edges is 8 cm^3 [a cm^3 is also referred to as a *cubic centimeter* (*cc*)]. The *liter* (L) has been used as a measure of volume in the metric system for many years and is exactly 1 dm^3. Due to its wide usage, the liter is an accepted part of the SI. Thus, 1 milliliter (1 mL) is equal to 1 cm^3 (or 1 cc); carry out the arithmetic to check this yourself. With respect to the English system, the liter is just a little larger than a quart; the relationship between these two units is 0.9461 L = 1 quart.

Various other derived units have been developed to describe particular properties of matter. For instance, for comparing the "heaviness" of different substances, *density* (*d*) is defined as mass per unit volume, which for liquids and solids (in the SI and metric systems) is usually expressed in g/cm^3. Thus, a cube of wood, 2.0 cm on an edge, weighing 6.0131 g, would have a density of:

$$d = \frac{6.0131\ g}{(2.0\ cm)^3} = 0.75\ g/cm^3$$

If the above arithmetic had been carried out on a calculator, one would have obtained additional decimal places, *i.e.* 0.751638. It is necessary to *round off* this number to two *significant figures,* 0.75 (or equivalently, using *exponential* or *scientific notation,* to 7.5×10^{-1}), so that the calculated answer shows an *accuracy* that is in agreement with the measured numbers. See pp. 349–351 in the Appendix.

A wide variety of important numerical conversions can be made on the basis of a knowledge of the SI and English systems. The successful chemistry student must learn to handle such conversions easily, since the problem-solving techniques are exactly the same as those used in solving many chemical problems. Carefully read the section, "Reporting Numerical Results", in the Appendix (pp. 446–448) and then study the following example problems, paying particular attention to the units of each factor and how they cancel, as well as the number of significant figures reported in the final answer.

Example Problems Involving Units, Conversions Between Units, and Significant Figures

1-A. Conversion of Metric Volume to English Volume: The "Factor-Label" or "Unit Conversion" Method

A soft drink bottler distributes its product in 2.00-L bottles. What is the volume in quarts?

(A) 1.89 qt (B) 2.00 qt (C) 2.11 qt (D) 8.00 qt (E) 9.39 qt

Solution 2.00 L → ? qt

This is a metric-to-English volume conversion. In the factor-label (or dimensional-analysis) method of problem solving, the original quantity, 2.00 L, is multiplied by one or more unity factors, cancelling units until the desired unit, quarts, is obtained. A unity factor is a ratio in which the numerator and denominator are equivalent.

In the section, "Derived Units", the connection, 0.9461 L = 1 qt, was given. By dividing both sides of this relationship by 1 qt, we have the ratio,

$$\frac{0.9461\ \text{L}}{1\ \text{qt}} = 0.9461\ \text{L/qt} = 1;$$

similarly, by dividing both sides by 0.9461 L we obtain the inverse ratio,

$$\frac{1\ \text{qt}}{0.9461\ \text{L}} = 1.057\ \text{qt/L} = 1,$$

as available unity factors. We see that the second of these will permit cancellation of the "L" from 2.00 L, and simultaneously interject "qt" into the numerator.

Answer Thus, (2.00 L) (Unity Factor) = ? qt

or

$$(2.00\,\cancel{\text{L}})\left(\frac{1\ \text{qt}}{0.9461\,\cancel{\text{L}}}\right) = 2.11\ \text{qt}$$ [C]

1-B. Mileage-Volume of Gasoline Required: The "Factor-Label" Method

A car is rated with a highway mileage of 41 miles per gallon of gasoline. Given that 1 mile = 1.6093 kilometers, how many liters of gasoline will be needed for a highway trip of 500. kilometers?

(A) 74 (B) 3.0×10^3 (C) 2.0 (D) 5.2 (E) 29

Solution

Again, one first writes down the numerical value and the unit(s) of the given condition, here 500. km. Then one writes down the unit(s) required by the answer. Compare these two sets of units, and from other available data, write down the sequence of steps that will be needed. In this case these are:

$$\text{km} \rightarrow \text{mi} \rightarrow \text{gal} \rightarrow \text{qt} \rightarrow \text{L}$$

The steps used will depend upon the unity factors available since each step requires the use of a unity factor. Thus, for our problem, the set-up takes the form:

$$(500.\ \text{km})\,(\text{Factor A})\,(\text{Factor B})\,(\text{Factor C})\,(\text{Factor D}) = ?\ \text{L}$$

It is worth noting that the word "per" used in the problem indicates a division line. The required unity factors are:

$$\left(\frac{1\ \text{mi}}{1.6093\ \text{km}}\right), \left(\frac{41\ \text{mi}}{1\ \text{gal}}\right), \left(\frac{1\ \text{gal}}{4\ \text{qt}}\right), \text{ and } \left(\frac{1\ \text{qt}}{0.9461\ \text{L}}\right).$$

These are now used directly or inverted so that cancellations leave only liters (L) as the final unit. This leads to the set-up:

$$(500.\ \cancel{\text{km}})\left(\frac{1\ \cancel{\text{mi}}}{1.6093\ \cancel{\text{km}}}\right)\left(\frac{1\ \cancel{\text{gal}}}{41\ \cancel{\text{mi}}}\right)\left(\frac{4\ \cancel{\text{qt}}}{1\ \cancel{\text{gal}}}\right)\left(\frac{0.9461\ \text{L}}{1\ \cancel{\text{qt}}}\right) = 29\ \text{L} \qquad [\text{E}]$$

Answer

If you used a calculator to carry out the indicated arithmetic, the calculator read-out would have been 28.67782235. This answer was rounded off to 2 significant figures so that it would agree in accuracy with the datum (41 mi/gal) used in the calculation. See the section in the Appendix (pp. 349–351) entitled "Reporting Numerical Results" for a discussion of significant figures and rounding off of numbers. Here, the factor, 4 qt/1 gal, is exact to any desired accuracy, *i.e.*, a "defined number", as opposed to a "measured number".

1-C. Conversion of Area in English Units to Metric Units: The "Factor-Label" Method—Unit Factors Raised to a Power

A typical freshman chemistry exam consists of pages measuring 8-1/2 × 11 inches (or a rather imposing 93-1/2 square inches, *i.e.* 93-1/2 in^2). What is the approximate area (reported to one significant figure) of one side of such a page in square meters, *i.e.* m^2?

(A) 0.001 m^2 (B) 2 m^2 (C) 0.01 m^2 (D) 0.06 m^2 (E) 0.0006 m^2

Solution

If the unity factor relating square meters and square inches were known, all that would be required would be to multiply the given area in square inches by this unity factor. However, there are no unity factors involving square measure in our table. Some related linear factors that do appear are:

$$\frac{2.540\text{ cm}}{1\text{ inch}}, \text{ and } \frac{10^{-2}\text{ m}}{1\text{ cm}} \text{ or } \frac{100\text{ cm}}{1\text{ m}}.$$

Linear unity factors cannot be directly used, but raising any of these to a power generates a new unity factor. Since the problem deals with area, these factors will need to be squared to give unity factors involving area, *i.e.*

$$\left(\frac{2.540\text{ cm}}{1\text{ inch}}\right)^2 \text{ and } \left(\frac{10^{-2}\text{ m}}{1\text{ cm}}\right)^2 \text{ or } \left(\frac{100\text{ cm}}{1\text{ m}}\right)^2.$$

Using these factors appropriately, the following steps are needed:

$$\text{in}^2 \rightarrow \text{cm}^2 \rightarrow \text{m}^2$$

$$(93\text{-}1/2\ \cancel{\text{in}^2})\left(\frac{2.540\ \cancel{\text{cm}}}{1\ \cancel{\text{in}}}\right)^2\left(\frac{1\text{ m}}{100\ \cancel{\text{cm}}}\right)^2 = 0.060322\text{ m}^2$$

Alternate Solution. The result would have been exactly the same if the linear English measurements had first been converted to meters and then multiplied together to get the area in m^2. Thus,

$$\left(8.5\text{ in} \times \frac{2.540\text{ cm}}{1\text{ in}} \times \frac{1\text{ m}}{100\text{ cm}}\right)\left(11\text{ in} \times \frac{2.540\text{ cm}}{1\text{ in}} \times \frac{10^{-2}\text{ m}}{1\text{ cm}}\right) = 0.060322\text{ m}^2$$

Answer

The first two zeroes are not significant figures, but merely hold the decimal point. Rounding to one significant figure gives 0.06 m^2 or equivalently 6 $\times$ 10^{-2} m^2. (Each page of the chemistry test now looks much less imposing, doesn't it?) **[D]**

You should note that $\left(\frac{2.540\text{ cm}}{\text{in}}\right)^2 = \frac{(2.540)^2\text{ cm}^2}{1^2\text{ in}^2} = \frac{6.452\text{ cm}^2}{1\text{ in}^2}$.

1-D. Mass- and Volume-Unit Conversion

The density of lead, 11.2 g/cm^3, when expressed in lb/ft^3 is:

(A) 2.60 (B) 699 (C) 11.2 (D) 0.179 (E) 0.753

Solution

$$\text{g/cm}^3 \rightarrow \text{lb/cm}^3 \rightarrow \text{lb/in.}^3 \rightarrow \text{lb/ft}^3$$

$$\left(\frac{11.2\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}}\right)\left(\frac{1\text{ lb}}{453.6\ \cancel{\text{g}}}\right)\left(\frac{2.540\ \cancel{\text{cm}}}{1\ \cancel{\text{in}}}\right)^3\left(\frac{12\ \cancel{\text{in}}}{1\text{ ft}}\right)^3 = 699.1813973\text{ lb/ft}^3$$

Answer

699 lb/ft^3 (It is rounded off to 3 significant figures, since the original density was given to 3 significant figures.) **[B]**

1-E. Volume to Area and Thickness Conversion

A latex semigloss enamel is advertised as having a coverage of 450. ft^2/gal when used on sealed surfaces. What is the average thickness of the coat of paint in millimeters?

(A) 10.1 (B) 188 (C) 0.0905 (D) 0.378 (E) 1.01

Solution

$$\text{Thickness} = \frac{\text{Vol}}{\text{Area}} = \frac{\text{gal} \to \text{qt} \to \text{L} \to \text{mL} \to \text{cm}^3 \to \text{mm}^3}{\text{ft}^2 \to \text{in}^2 \to \text{cm}^2 \to \text{mm}^2}$$

$$\frac{(1\ \text{gal})\left(4\ \frac{\text{qt}}{\text{gal}}\right)\left(0.946\frac{\text{L}}{\text{qt}}\right)\left(10^3\ \frac{\text{mL}}{\text{L}}\right)\left(1\ \frac{\text{cm}^3}{\text{mL}}\right)\left(10\ \frac{\text{mm}}{\text{cm}}\right)^3}{(450\ \text{ft}^2)\left(12\frac{\text{in}}{\text{ft}}\right)^2\left(2.540\ \frac{\text{cm}}{\text{in}}\right)^2\left(10\ \frac{\text{mm}}{\text{cm}}\right)^2}$$

$$= 0.0905125267\ \text{mm}$$

Answer

0.0905 mm (3 significant figures). The accuracy implied is 1 part in 905, or about 0.1%. This is comparable in order of magnitude to the accuracy indicated in the data, *i.e.,* the area (~0.2%) and in the qt → L conversion factor (~0.1%). **[C]**

1.F. Significant Figures in Addition, Subtraction, and Division

A mixture is prepared by mixing 7.65 g of A and 6.95 g of B. The total weight of the mixture can be reported to __?__ significant figures; the ratio, wt of A/wt of B, in the mixture can be reported to __?__ significant figures; and the difference, A - B, can be reported to __?__ significant figures. The numbers required to fill the blanks are, respectively,

(A) 3, 3, 3 (B) 4, 3, 2 (C) 4, 3, 3 (D) 4, 4, 3 (E) None of these

Solution

In the summation, both weights are known to the nearest 0.01 g; therefore the sum, 14.60 g, is likewise known to the hundreths place, and one significant figure has been gained in the addition. In the ratio, the quotient, 7.65/6.95 = 1.100719424, must be rounded to 1.10, or 3 significant figures. The difference, 0.70 g, is also good to the hundreths place, but contains only two significant figures due to the loss of the units place in the subtraction.