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SCHAUM'S OUTLINE OF

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OF

COLLEGE CHEMISTRY

SEVENTH EDITION

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JEROME L. ROSENBERG, Ph.D.
Professor of Biological Sciences and Chemistry
University of Pittsburgh

and

LAWRENCE M. EPSTEIN, Ph.D.
Associate Professor of Chemistry, Emeritus
University of Pittsburgh

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JEROME L. ROSENBERG did his graduate work at Columbia University in physical chemistry, receiving his M.A. in 1944 and his Ph.D. in 1948. His research activities at several institutions resulted in many papers in photosynthesis, general photochemistry, and molecular electronic structure. Formerly Dean of the Faculty of Arts and Sciences and Vice Provost at the University of Pittsburgh he is currently Chairman of the Department of Biological Sciences in addition to being professor of biological sciences and chemistry there. Dr. Rosenberg has been associated with the preparation of *Schaum's Outline of College Chemistry* since the third edition, published in 1949.

LAWRENCE M. EPSTEIN started his career as a chemical engineer, then earned his M.S. in 1952 and Ph.D. in 1955 from Polytechnic University in the field of physical chemistry. He did research in the fields of radiation chemistry and in Mössbauer effect spectroscopy at Westinghouse Research Laboratories, and then served as Associate Professor and supervisor of the General Chemistry program at the University of Pittsburgh until he retired in 1986.

Schaum's Outline of Theory and Problems of
COLLEGE CHEMISTRY

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Preface

This book is designed to help the student of college chemistry by summarizing the chemical principles of each topic and relating the solution of quantitative problems to those fundamentals. Although the book is not intended to replace a textbook, its solved problems, with complete and detailed solutions, do cover most of the subject matter of a first course in college chemistry. Both the solved and the supplementary problems are arranged to allow a progression in difficulty within each topic.

Several important features had been introduced into the sixth edition, notably the kinetic theory of gases, a more formal treatment of thermochemistry, a modern treatment of atomic properties and chemical bonding, and a chapter on chemical kinetics.

In this seventh edition the early chapters were revised to conform more closely to the methods used in current textbooks to introduce calculational skills to the beginning student. Some changes in notation were made, and there is now a more consistent usage of SI units. Every problem in the book was checked for accuracy, and the latest values of various physical and chemical constants were used. Some problems were removed and new ones written to replace them. An attempt was made to increase the variety of stoichiometry problems, especially in the chapters on gases and solutions, while eliminating some of the very complex problems that arise in gaseous and aqueous equilibria.

In the treatment of chemical bonding the subject of molecular orbitals has been deemphasized in favor of VSEPR theory. An elementary treatment of bonding in metals has been added, and more attention is given to intermolecular weak forces in liquids and solids.

A new chapter on Organic and Biochemistry has been added, conforming to the trend in current texts.

The use of SI units is still not universal; liter, atmosphere, and calorie are still used where appropriate, but the joule is favored, and the reader is made aware of the conversion to the bar as the standard-state pressure.

The authors acknowledge with thanks the careful word processing of Masha Epstein and the perceptive and helpful editing of David Beckwith.

Jerome L. Rosenberg
Lawrence M. Epstein

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Chapter 1

Quantities and Units

INTRODUCTION

Most of the measurements and calculations in chemistry and physics are concerned with different kinds of *quantities*, e.g., length, velocity, volume, mass, energy. Every measurement includes both a number and a unit. The *unit* simultaneously identifies the kind of dimension and the magnitude of the reference quantity used as a basis for comparison. Many units are commonly used for the dimension of length, e.g., inch, yard, mile, centimeter, meter, kilometer. The *number* obviously indicates how many of the reference units are contained in the quantity being measured. Thus the statement that the length of a room is 20 feet means that the length of the room is 20 times the length of the foot, which in this case is the unit of length chosen for comparison. Although *20 feet* has the dimension of length, *20* is a pure number and is dimensionless, being the ratio of two lengths, that of the room and that of the reference foot.

It will be assumed that readers are familiar with the use of exponents, particularly powers-of-ten notation, and with the rules for significant figures. If not, Appendices A and B should be studied in conjunction with Chapter 1.

SYSTEMS OF MEASUREMENT

Dimensional calculations are greatly simplified if the unit for each kind of measure is expressed in terms of the units of selected reference dimensions. The three independent reference dimensions for mechanics are *length*, *mass*, and *time*. As examples of relating other quantities to the reference dimensions, the unit of speed is defined as unit length per unit time, the unit of volume is the cube of the unit of length, etc. Other reference dimensions, such as those used to express electrical and thermal phenomena, will be introduced later. There are several systems of units still in use in the English-speaking nations, so that one must occasionally make calculations to convert values from one system to another, e.g., inches to centimeters, or pounds to kilograms.

INTERNATIONAL SYSTEM OF UNITS

Considerable progress is being made in the acceptance of a common international system of reference units within the world scientific community. This system, known as SI from the French name, Système International d'Unités, has been adopted by many international bodies, including the International Union of Pure and Applied Chemistry. In SI, the reference units for *length*, *mass*, and *time* are the *meter*, *kilogram*, and *second*, with the symbols m, kg, and s, respectively.

To express quantities much larger or smaller than the standard units, use may be made of multiples or submultiples of these units, defined by applying as multipliers of these units certain recommended powers of ten, listed in Table 1-1. The multiplier abbreviation is to precede the symbol of the base unit without any space or punctuation. Thus, picosecond (10^{-12} s) is ps, and kilometer (10^3 m) is km. Since for historical reasons the SI unit for mass, kilogram, already has a prefix, multiples for mass should be derived by applying the multiplier to the *gram* rather than to the *kilogram*. Thus, microgram (10^{-6}g), abbreviated μg, is 10^{-9} kg.

Compound units can be derived by applying algebraic operations to the simple units.

EXAMPLE 1 The unit for volume in SI is the cubic meter (m^3), since

$$\text{Volume} = \text{length} \times \text{length} \times \text{length} = \text{m} \times \text{m} \times \text{m} = \text{m}^3$$

Table 1-1. Multiples and Submultiples for Units

Prefix	Abbreviation	Multiplier	Prefix	Abbreviation	Multiplier
deci	d	10^{-1}	deka	da	10
centi	c	10^{-2}	hecto	h	10^2
milli	m	10^{-3}	kilo	k	10^3
micro	μ	10^{-6}	mega	M	10^6
nano	n	10^{-9}	giga	G	10^9
pico	p	10^{-12}	tera	T	10^{12}
femto	f	10^{-15}			
atto	a	10^{-18}			

The unit for speed is the unit for length (distance) divided by the unit for time, or the meter per second (m/s), since

$$\text{Speed} = \frac{\text{distance}}{\text{time}} = \frac{\text{m}}{\text{s}}$$

The unit for density is the unit for mass divided by the unit for volume, or the kilogram per cubic meter (kg/m^3), since

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{kg}}{\text{m}^3}$$

Table 1-2. Some SI and Non-SI Units

Physical Quantity	Unit Name	Unit Symbol	Definition
Length	angstrom	Å	10^{-10} m
	inch	in	2.54×10^{-2} m
Area	square meter (SI)	m^2	
Volume	cubic meter (SI)	m^3	
	liter	L	dm^3
	cubic centimeter	cm^3	
Mass	atomic mass unit (dalton)	u	$1.660\,57 \times 10^{-27}$ kg
	pound	lb	0.453 592 37 kg
Density	kilogram per cubic meter (SI)	kg/m^3	
Force	newton (SI)	N	$kg \cdot m/s^2$
Pressure	pascal (SI)	Pa	N/m^2
	bar	bar	10^5 Pa
	atmosphere	atm	101 325 Pa
	torr (millimeter of mercury)	torr (mm Hg)	atm/760 *or* 133.32 Pa

Symbols for compound units may be expressed in any of the following formats.

1. Multiplication of units. Example: *kilogram second.*
   - (a) Dot $\mathrm{kg \cdot s}$
   - (b) Spacing $\mathrm{kg\ s}$ (*not used in this book*)
2. Division of units. Example: *meter per second.*
   - (a) Division sign $\frac{\mathrm{m}}{\mathrm{s}}$ or m/s
   - (b) Negative power $\mathrm{m \cdot s^{-1}}$ (or $\mathrm{m\ s^{-1}}$)

Note that the term *per* in a word definition is equivalent to *divide by* in the mathematical notation. Note also that symbols are not followed by a period, except at the end of a sentence.

In recognition of long-standing traditions, the various international commissions have acknowledged that a few non-SI units will remain in use in certain fields of science and in certain countries during a transitional period. Table 1-2 lists some units, both SI and non-SI, which will be used in this book, involving the quantities length, mass, and time. Other units will be introduced in subsequent chapters.

TEMPERATURE

Temperature may be defined as that property of a body which determines the flow of heat. Two bodies are at the same temperature if there is no transfer of heat when they are placed together. Temperature is an independent dimension which cannot be defined in terms of mass, length, and time. The SI unit of temperature is the *kelvin*, and 1 kelvin (K) is defined as 1/273.16 times the *triple point* temperature. The *triple point* is the temperature at which water coexists in equilibrium with ice at the pressure exerted by water vapor only. The *triple point* is 0.01 K above the *normal freezing point* of water, at which water and ice coexist in equilibrium with air at standard atmospheric pressure. The SI unit of temperature is so defined that 0 K is the absolute zero of temperature; the SI or Kelvin scale is often called the *absolute temperature scale.* Although absolute zero is never actually attainable, it has been approached to within less than 10^{-4} K.

OTHER TEMPERATURE SCALES

Other temperature scales have been employed at various times. The most useful ones are those which are linearly related to the Kelvin scale; that is, the ratio, r, of the difference between any two temperatures on such a scale to the number of kelvins separating the same two temperatures is a fixed number. One such scale, for which $r = 1$, is the *Celsius* (sometimes called the *centigrade*) scale. On the Celsius scale, the freezing point of water is 0 °C, the boiling point is 100 °C at one atmosphere pressure, and absolute zero is −273.15 °C. The *Fahrenheit scale* is one for which $r = \frac{9}{5}$. The absolute zero and the freezing point and boiling point of water on the Fahrenheit scale are −459.67 °F, 32 °F, and 212 °F, respectively. The relationships among these three scales, illustrated in Fig. 1-1, are described by the following linear equations, in which temperature on the SI scale is designated by T, and on the other scales by t.

$$\frac{t}{\mathrm{^\circ C}} = \frac{T}{\mathrm{K}} - 273.15 \qquad \text{or} \qquad t = \left(\frac{T}{\mathrm{K}} - 273.15\right)\mathrm{^\circ C}$$

$$\frac{t}{\mathrm{^\circ F}} = \frac{9}{5}\left(\frac{t}{\mathrm{^\circ C}}\right) + 32 \qquad \text{or} \qquad t = \left[\frac{9}{5}\left(\frac{t}{\mathrm{^\circ C}}\right) + 32\right]\mathrm{^\circ F}$$

$$\frac{t}{\mathrm{^\circ C}} = \frac{5}{9}\left(\frac{t}{\mathrm{^\circ F}} - 32\right) \qquad \text{or} \qquad t = \frac{5}{9}\left(\frac{t}{\mathrm{^\circ F}} - 32\right)\mathrm{^\circ C}$$

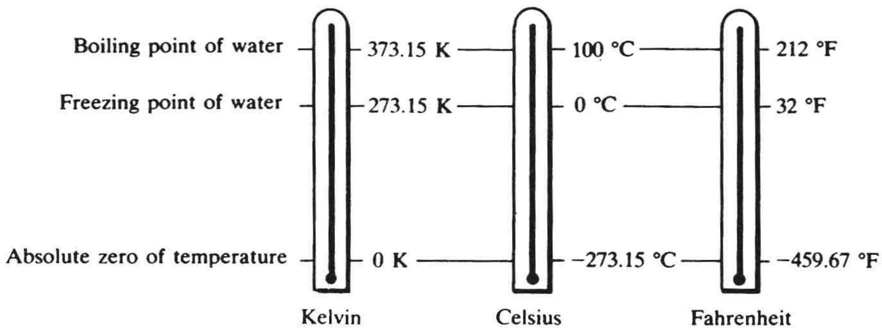

Fig. 1-1.

Here the ratios T/K, $t/°\text{C}$, $t/°\text{F}$ are the dimensionless numerical measures of the temperature on the Kelvin, Celsius, and Fahrenheit scales, respectively. Read the upper left equation as follows: The temperature in degrees *Celsius* equals the temperature in kelvins minus 273.15.

USE AND MISUSE OF UNITS

The units (e.g., cm, kg, g/mL, ft/s) must be regarded as a necessary part of the complete specification of a physical quantity. It is as foolish to separate the number of a measure from its unit as it is to separate a laboratory reagent bottle from its label. When physical quantities are subjected to mathematical operations, the units must be carried along with the numbers and must undergo the same operations as the numbers. Quantities cannot be added or subtracted directly unless they have not only the same dimensions but also the same units.

EXAMPLE 2 It is obvious that we cannot add 5 hours (time) to 20 miles/hour (speed) since *time* and *speed* have different physical significance. If we are to add 2 lb (mass) and 4 kg (mass), we must first convert lb to kg or kg to lb. Quantities of various types, however, can be combined in multiplication or division, in which *the units as well as the numbers* obey the algebraic laws of squaring, cancellation, etc. Thus:

1. $6\text{ L} + 2\text{ L} = 8\text{ L}$
2. $(5\text{ cm})(2\text{ cm}^2) = 10\text{ cm}^3$
3. $(3\text{ ft}^3)(200\text{ lb/ft}^3) = 600\text{ lb}$
4. $(2\text{ s})(3\text{ ft/s}^2) = 6\text{ ft/s}$
5. $\dfrac{15\text{ g}}{3\text{ g/cm}^3} = 5\text{ cm}^3$

ESTIMATION OF NUMERICAL ANSWERS

If one's calculator is working correctly and is accurately used, the answer will be correct. But if not, will an incorrect answer be recognized? A very important skill is to determine first, by visual inspection, an approximate answer. Especially important is the correct order of magnitude, represented by the location of the decimal point or the power of 10, which may go astray even though the digits are correct.

EXAMPLE 3 Consider the multiplication: $122\text{ g} \times 0.051\,8 = 6.32\text{ g}$. Visual inspection shows that 0.051 8 is a little more than $\frac{1}{20}$, and $\frac{1}{20}$ of 122 is a little more than 6. Hence the answer should be a little more than 6 g, which it is. If the answer were given incorrectly as 63.2 g or 0.632 g, visual inspection or mental checking of the result would indicate that the decimal point had been misplaced.

Solved Problems

UNITS BASED ON MASS OR LENGTH

1.1. The following examples illustrate conversions among various units of length, volume, or mass.

1 inch = 2.54 cm = 0.025 4 m = 25.4 mm = 2.54×10^7 nm
1 foot = 12 in = 12×2.54 cm = 30.48 cm = 0.304 8 m = 304.8 mm
1 liter = 1 dm^3 = 10^{-3} m^3
1 mile = 5 280 ft = 1.609×10^5 cm = 1.609×10^3 m = 1.609 km = 1.609×10^6 mm
1 pound = 0.453 6 kg = 453.6 g = 4.536×10^5 mg
1 metric ton = 1 000 kg = 10^6 g

1.2. Convert 5.00 inches to (*a*) centimeters, (*b*) millimeters, (*c*) meters.

(*a*)
$$5.00\text{ in} = (5.00\text{ in})(2.54\text{ cm/in}) = 12.7\,\frac{\text{in}\cdot\text{cm}}{\text{in}} = 12.7\text{ cm}$$

The procedure can be understood easily in terms of the word definition of the conversion factor: 2.54 is the number of cm *per* in; that is, the number of cm in 1 in. Thus the number of cm in 5 in is 5×2.54.

More formally, the conversion factor may be considered to be a statement of equality between 2.54 cm and 1 in. Since 2.54 cm = 1 in,

$$2.54\text{ cm per in} = 2.54\text{ cm/in} = \frac{2.54\text{ cm}}{1\text{ in}} = \frac{2.54\text{ cm}}{2.54\text{ cm}} = 1$$

Thus the conversion factor, 2.54 cm/in, is mathematically equal to 1, so that any quantity may be multiplied or divided by the conversion factor without changing the essential value of the quantity.

The inverse of the above conversion factor, i.e., 1 in/2.54 cm, must likewise be equal to 1 and is also a conversion factor.

In this problem the multiplication of 5.00 in by 2.54 cm/in leads to the useful result 12.7 cm, a statement of the same quantity of length in different units. The multiplication of 5.00 in by 1 in/2.54 cm, although mathematically allowed, does not lead to a useful result because the result, 1.97 in^2/cm, is not expressed in a simple unit which has a physical interpretation. In general, any proper conversion factor is equal to 1 and either it or its inverse may be used as a multiplier of any physical quantity. The choice depends on the particular problem.

(*b*)
$$12.7\text{ cm} = (12.7\text{ cm})(10\text{ mm/cm}) = 127\,\frac{\text{cm}\cdot\text{mm}}{\text{cm}} = 127\text{ mm}$$

(*c*)
$$12.7\text{ cm} = (12.7\text{ cm})(1\text{ m/100 cm}) = 0.127\,\frac{\text{cm}\cdot\text{m}}{\text{cm}} = 0.127\text{ m}$$

Obviously multiplication by a conversion factor is equivalent to division by its inverse. Thus:

$$12.7\text{ cm} = \frac{12.7\text{ cm}}{100\text{ cm/m}} = 0.127\,\frac{\text{cm}\cdot\text{m}}{\text{cm}} = 0.127\text{ m}$$

Here it was appropriate to divide by the conversion factor 100 cm/m. If by mistake we had multiplied 12.7 cm by 100 cm/m, the answer would have been expressed in cm^2/m, since

$$\text{cm} \times \frac{\text{cm}}{\text{m}} = \frac{\text{cm}^2}{\text{m}}$$

We would immediately realize our error, seeing that the answer is not expressed in meters.

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1.3. Convert (*a*) 14.0 cm and (*b*) 7.00 m to inches.

(*a*)
$$14.0\text{ cm} = (14.0\text{ cm})\left(\frac{1\text{ in}}{2.54\text{ cm}}\right) = 5.51\text{ in} \qquad \text{or} \qquad 14.0\text{ cm} = \frac{14.0\text{ cm}}{2.54\text{ cm/in}} = 5.51\text{ in}$$

Factors such as $\frac{1\text{ in}}{2.54\text{ cm}}$ will frequently appear in the form 1 in/2.54 cm, especially in computer displays, because it allows the entire calculation to appear on one line of type.

(*b*)
$$7.00\text{ m} = (7.00\text{ m})\left(\frac{100\text{ cm}}{1\text{ m}}\right)\left(\frac{1\text{ in}}{2.54\text{ cm}}\right) = 276\text{ in}$$

Note that the use of two successive conversion factors was necessary. Both the units, m and cm, canceled out leaving the desired unit, in.

1.4. How many square inches in one square meter?

$$1\text{ m} = (1\text{ m})(100\text{ cm/1 m})(1\text{ in/2.54 cm}) = 39.37\text{ in}$$
$$1\text{ m}^2 = (1\text{ m})^2 = (39.37\text{ in})^2 = 1\,550\text{ in}^2$$

alternatively,

$$1\text{ m}^2 = (1\text{ m})^2(100\text{ cm/1 m})^2(1\text{ in/2.54 cm})^2$$
$$= [(100)^2/(2.54)^2]\text{ in}^2 = 1\,550\text{ in}^2$$

Note that since a conversion factor is equal to 1, it may be squared (or raised to any power) without changing its value.

1.5. (*a*) How many cubic centimeters in one cubic meter? (*b*) How many liters in one cubic meter? (*c*) How many cubic centimeters in one liter?

(*a*)
$$1\text{ m}^3 = (1\text{ m})^3\left(\frac{100\text{ cm}}{1\text{ m}}\right)^3 = (10^2\text{ cm})^3 = 10^6\text{ cm}^3$$

(*b*)
$$1\text{ m}^3 = (1\text{ m})^3\left(\frac{10\text{ dm}}{1\text{ m}}\right)^3\left(\frac{1\text{ L}}{1\text{ dm}^3}\right) = 10^3\text{ L}$$

(*c*)
$$1\text{ L} = 1\text{ dm}^3 = (1\text{ dm})^3\left(\frac{10\text{ cm}}{1\text{ dm}}\right)^3 = 10^3\text{ cm}^3$$

1.6. Find the capacity in liters of a tank 0.6 m long, 10 cm wide, and 50 mm deep.

Convert to decimeters, since 1 L = 1 dm^3.

$$\text{Volume} = (0.6\text{ m})(10\text{ cm})(50\text{ mm})$$
$$= (0.6\text{ m})\left(\frac{10\text{ dm}}{1\text{ m}}\right) \times (10\text{ cm})\left(\frac{1\text{ dm}}{10\text{ cm}}\right) \times (50\text{ mm})\left(\frac{1\text{ dm}}{100\text{ mm}}\right)$$
$$= (6\text{ dm})(1\text{ dm})(0.5\text{ dm}) = 3\text{ dm}^3 = 3\text{ L}$$

1.7. Determine the mass of 66 lb of sulfur in (*a*) kilograms and (*b*) grams. (*c*) Find the mass of 3.4 kg of copper in pounds.

(*a*)
$$66\text{ lb} = (66\text{ lb})(0.453\,6\text{ kg/lb}) = 30\text{ kg} \qquad \text{or} \qquad 66\text{ lb} = (66\text{ lb})(1\text{ kg/2.2 lb}) = 30\text{ kg}$$

(*b*) $$66 \text{ lb} = (66 \text{ lb})(453.6 \text{ g/lb}) = 30\,000 \text{ g, or } 3.0 \times 10^4 \text{ g}$$

(*c*) $$3.4 \text{ kg} = (3.4 \text{ kg})(2.2 \text{ lb/kg}) = 7.5 \text{ lb}$$

COMPOUND UNITS

1.8. Fatty acids spread spontaneously on water to form a monomolecular film. A benzene solution containing 0.1 mm^3 of stearic acid is dropped into a tray full of water. The acid is insoluble in water but spreads on the surface to form a continuous film area of 400 cm^2 after all of the benzene has evaporated. What is the average film thickness in angstroms?

$$1 \text{ mm}^3 = (10^{-3} \text{ m})^3 = 10^{-9} \text{ m}^3 \qquad 1 \text{ cm}^2 = (10^{-2} \text{ m})^2 = 10^{-4} \text{ m}^2$$

$$\text{Film thickness} = \frac{\text{volume}}{\text{area}} = \frac{(0.1 \text{ mm}^3)(10^{-9} \text{ m}^3/\text{mm}^3)}{(400 \text{ cm}^2)(10^{-4} \text{ m}^2/\text{cm}^2)} = 2.5 \times 10^{-9} \text{ m}$$

$$= 2.5 \times 10^{-9} \text{ m} \times 10^{10} \text{ Å/m} = 25 \text{ Å}$$

1.9. A pressure of one atmosphere is equal to 101.3 kPa. Express this pressure in pounds force (lbf) per square inch.

The pound force (lbf) is equal to 4.448 N.

$$1 \text{ atm} = 101.3 \text{ kPa} = (101.3 \times 10^3 \text{ N/m}^2)\left(\frac{1 \text{ lbf}}{4.448 \text{ N}}\right)\left(\frac{2.54 \times 10^{-2} \text{ m}}{1 \text{ in}}\right)^2 = 14.69 \text{ lbf/in}^2$$

Notice that the conversion factor between m and in is squared to give the conversion factor between m^2 and in^2.

1.10. A table is to be prepared showing the dependence of the braking distance of an automobile upon speed of travel. The braking distance is to be expressed in feet and the speed in miles per hour. If the entries in the table are to be pure (dimensionless) numbers, how should the columns be headed for the two physical quantities?

If d is to be the symbol for braking distance, the column showing braking distances should be headed, d/ft. An entry under that heading, such as 200, would then be a pure number, the ratio of the braking distance to the standard foot. If $d/\text{ft} = 200$, then $d = 200$ ft. This symbolic heading is more acceptable than the headings

$$d(\text{ft}) \qquad d, \text{ ft}$$

Similarly, if speed is to be designated as s, the column should be headed, $s/\text{mi} \cdot \text{h}^{-1}$ or $s/(\text{mi/h})$. An entry in this column, such as 62, would be the ratio between the car speed and the standard mile per hour. (The heading $s \cdot \text{h/mi}$ might also be used, but the ratio form is clearer.)

1.11. New York City's 7.9 million people in 1978 had a daily per capita consumption of 173 gallons of water. How many tons of sodium fluoride (45% fluorine by weight) would be required per year to give this water a tooth-strengthening dose of 1 part (by weight) fluorine per million parts water? One U.S. gallon of water at normal room temperature weighs 8.34 lbf (i.e., has a mass of 8.34 lb). One ton is 2 000 lb.

Mass of water, in tons, required per year

$$= (7.9 \times 10^6 \text{ persons})\left(\frac{173 \text{ gal water}}{\text{person} \cdot \text{day}}\right)\left(\frac{365 \text{ days}}{\text{yr}}\right)\left(\frac{8.34 \text{ lb water}}{1 \text{ gal water}}\right)\left(\frac{1 \text{ ton}}{2\,000 \text{ lb}}\right)$$

$$= 2.08 \times 10^9 \frac{\text{tons water}}{\text{yr}}$$

Note that all units cancel out except tons water/yr, which appears in the result.

Mass of sodium fluoride, in tons, required per year

$$= \left(2.08 \times 10^9 \frac{\text{tons water}}{\text{yr}}\right)\left(\frac{1 \text{ ton fluorine}}{10^6 \text{ tons water}}\right)\left(\frac{1 \text{ ton sodium fluoride}}{0.45 \text{ ton fluorine}}\right)$$

$$= 4.6 \times 10^3 \frac{\text{tons sodium fluoride}}{\text{yr}}$$

1.12. A tennis ball was observed to travel at a speed of 95 miles per hour. Express this in meters per second.

$$95 \text{ mi/h} = \left(95 \frac{\text{mi}}{\text{h}}\right)\left(1.609 \times 10^3 \frac{\text{m}}{\text{mi}}\right)\left(\frac{1 \text{ h}}{3.6 \times 10^3 \text{ s}}\right) = 42.5 \frac{\text{m}}{\text{s}}$$

The conversion factor relating the meter to the mile was taken from Problem 1.1.

1.13. Calculate the density, in g/cm³, of a body that weighs 420 g (i.e., has a mass of 420 g) and has a volume of 52 cm³.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{420 \text{ g}}{52 \text{ cm}^3} = 8.1 \text{ g/cm}^3$$

1.14. Express the density of the above body in the standard SI unit, kg/m³.

$$(8.1 \text{ g/cm}^3)\left(\frac{1 \text{ kg}}{1\,000 \text{ g}}\right)\left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 = 8.1 \times 10^3 \text{ kg/m}^3$$

1.15. What volume will 300 g of mercury occupy? Density of mercury is 13.6 g/cm³.

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{300 \text{ g}}{13.6 \text{ g/cm}^3} = 22.1 \text{ cm}^3$$

1.16. The density of cast iron is 7 200 kg/m³. Calculate its density in pounds per cubic foot.

$$\text{Density} = \left(7\,200 \frac{\text{kg}}{\text{m}^3}\right)\left(\frac{1 \text{ lb}}{0.453\,6 \text{ kg}}\right)\left(\frac{0.304\,8 \text{ m}}{1 \text{ ft}}\right)^3 = 449 \text{ lb/ft}^3$$

The two conversion factors were taken from Problem 1.1.

1.17. A casting of an alloy in the form of a disk weighed 50.0 g. The disk was 0.250 inch thick and had a circular cross section of diameter 1.380 in. What is the density of the alloy, in g/cm³?

$$\text{Volume} = \left(\frac{\pi d^2}{4}\right)h = \left(\frac{\pi(1.380 \text{ in})^2(0.250 \text{ in})}{4}\right)\left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^3 = 6.13 \text{ cm}^3$$

$$\text{Density of the alloy} = \frac{\text{mass}}{\text{volume}} = \frac{50.0 \text{ g}}{6.13 \text{ cm}^3} = 8.15 \text{ g/cm}^3$$

1.18. The density of zinc is 455 lb/ft³. Find the mass in grams of 9.00 cm³ of zinc.

First express the density in g/cm³.

$$\left(455 \frac{\text{lb}}{\text{ft}^3}\right)\left(\frac{1 \text{ ft}}{30.48 \text{ cm}}\right)^3\left(\frac{453.6 \text{ g}}{1 \text{ lb}}\right) = 7.29 \frac{\text{g}}{\text{cm}^3}$$

$$(9.00 \text{ cm}^3)(7.29 \text{ g/cm}^3) = 65.6 \text{ g}$$

1.19. Battery acid has a density of 1.285 g/cm^3 and contains 38.0% by weight H_2SO_4. How many grams of pure H_2SO_4 are contained in a liter of battery acid?

1 cm^3 of acid has a mass of 1.285 g. Then 1 L of acid (1 000 cm^3) has a mass of 1 285 g. Since 38.0% by weight (or by mass) of the acid is pure H_2SO_4, the number of grams of H_2SO_4 in 1 L of battery acid is

$$0.380 \times 1\,285 \text{ g} = 488 \text{ g}$$

Formally, the above solution can be written as follows:

$$\text{Mass of } H_2SO_4 = (1\,285 \text{ g acid})\left(\frac{38 \text{ g } H_2SO_4}{100 \text{ g acid}}\right) = 488 \text{ g } H_2SO_4$$

Here the conversion factor

$$\frac{38 \text{ g } H_2SO_4}{100 \text{ g acid}}$$

is taken to be equal to 1. Although the condition 38 g H_2SO_4 = 100 g acid is not a universal truth in the same sense that 1 in always equals 2.54 cm, the condition is a rigid one of association of 38 g H_2SO_4 with every 100 g acid for this particular acid preparation. Mathematically, these two quantities may be considered to be equal for this problem, since one of the quantities implies the other. Liberal use will be made in subsequent chapters of conversion factors that are valid only for particular cases, in addition to the conversion factors that are universally valid.

1.20. (*a*) Calculate the mass of pure HNO_3 per cm^3 of the concentrated acid which assays 69.8% by weight HNO_3 and has a density of 1.42 g/cm^3. (*b*) Calculate the mass of pure HNO_3 in 60.0 cm^3 of concentrated acid. (*c*) What volume of the concentrated acid contains 63.0 g of pure HNO_3?

(*a*) 1 cm^3 of acid has a mass of 1.42 g. Since 69.8% of the total mass of the acid is pure HNO_3, then the number of grams of HNO_3 in 1 cm^3 of acid is

$$0.698 \times 1.42 \text{ g} = 0.991 \text{ g}$$

(*b*) $$\text{Mass of } HNO_3 \text{ in } 60.0 \text{ cm}^3 \text{ of acid} = (60.0 \text{ cm}^3)(0.991 \text{ g/cm}^3) = 59.5 \text{ g } HNO_3$$

(*c*) 63.0g HNO_3 is contained in

$$\frac{63.0 \text{ g}}{0.991 \text{ g/cm}^3} = 63.6 \text{ cm}^3 \text{ acid}$$

TEMPERATURE

1.21. Ethyl alcohol (*a*) boils at 78.5 °C and (*b*) freezes at −117 °C, at one atmosphere pressure. Convert these temperatures to the Fahrenheit scale.

Use $$t = \left[\frac{9}{5}\left(\frac{t}{°\text{C}}\right) + 32\right] °\text{F}$$

(*a*) $$t = [\tfrac{9}{5}(78.5) + 32]\,°\text{F} = (141 + 32)\,°\text{F} = 173\,°\text{F}$$

(*b*) $$t = [\tfrac{9}{5}(-117) + 32]\,°\text{F} = (-211 + 32)\,°\text{F} = -179\,°\text{F}$$

1.22. Mercury (*a*) boils at 675 °F and (*b*) solidifies at −38.0 °F, at one atmosphere pressure. Express these temperatures in degrees Celsius.

Use $$t = \frac{5}{9}\left(\frac{t}{°\text{F}} - 32\right) °\text{C}$$

(*a*) $$t = \tfrac{5}{9}(675 - 32)\,°\text{C} = \tfrac{5}{9}(643)\,°\text{C} = 357\,°\text{C}$$

(*b*) $$t = \tfrac{5}{9}(-38.0 - 32.0)\,°\text{C} = \tfrac{5}{9}(-70.0)\,°\text{C} = -38.9\,°\text{C}$$

1.23. Change (*a*) 40 °C and (*b*) −5 °C to the Kelvin scale.

Use
$$T = \left(\frac{t}{°\mathrm{C}} + 273\right)\mathrm{K}$$

(*a*)
$$T = (40 + 273)\ \mathrm{K} = 313\ \mathrm{K}$$

(*b*)
$$T = (-5 + 273)\ \mathrm{K} = 268\ \mathrm{K}$$

1.24. Convert (*a*) 220 K and (*b*) 498 K to the Celsius scale.

Use
$$t = \left(\frac{T}{\mathrm{K}} - 273\right)°\mathrm{C}$$

(*a*)
$$t = (220 - 273)\ °\mathrm{C} = -53\ °\mathrm{C}$$

(*b*)
$$t = (498 - 273)\ °\mathrm{C} = 225\ °\mathrm{C}$$

1.25. During the course of an experiment, laboratory temperature rose 0.8 °C. Express this rise in degrees Fahrenheit.

Temperature *intervals* are converted differently than temperature *readings.* For intervals, it is seen from Fig. 1-1 that

$$100\ °\mathrm{C} = 180\ °\mathrm{F} \qquad \text{or} \qquad 5\ °\mathrm{C} = 9\ °\mathrm{F}$$

Hence
$$0.8\ °\mathrm{C} = (0.8\ °\mathrm{C})\left(\frac{9\ °\mathrm{F}}{5\ °\mathrm{C}}\right) = 1.4\ °\mathrm{F}$$

Supplementary Problems

UNITS BASED ON MASS OR LENGTH

1.26. (*a*) Express 3.69 m in kilometers, in centimeters, and in millimeters. (*b*) Express 36.24 mm in centimeters and in meters.

Ans. (*a*) 0.003 69 km, 369 cm, 3 690 mm; (*b*) 3.624 cm, 0.036 24 m

1.27. Determine the number of (*a*) millimeters in 10 in, (*b*) feet in 5 m, (*c*) centimeters in 4 ft 3 in.

Ans. (*a*) 254 mm; (*b*) 16.4 ft; (*c*) 130 cm

1.28. Convert the molar volume, 22.4 liters, to cubic centimeters, to cubic meters, and to cubic feet.

Ans. 22 400 cm^3, 0.022 4 m^3, 0.791 ft^3

1.29. Express the weight (mass) of 32 g of oxygen in milligrams, in kilograms, and in pounds.

Ans. 32 000 mg, 0.032 kg, 0.070 5 lb

1.30. How many grams in 5.00 lb of copper sulfate? How many pounds in 4.00 kg of mercury? How many milligrams in 1 lb 2 oz of sugar?

Ans. 2 270 g, 8.82 lb, 510 000 mg

1.31. Convert the weight (mass) of 500 lb of coal to (*a*) kilograms, (*b*) metric tons, (*c*) U.S. tons (1 ton = 2 000 lb).

Ans. (*a*) 227 kg; (*b*) 0.227 metric ton; (*c*) 0.250 ton

1.32. The color of light depends on its wavelength. The longest visible rays, at the red end of the visible spectrum, are 7.8×10^{-7} m in length. Express this length in micrometers, in nanometers, and in angstroms.

Ans. 0.78 μm, 780 nm, 7 800 Å

1.33. Determine the number of (*a*) cubic centimeters in a cubic inch, (*b*) cubic inches in a liter, (*c*) cubic feet in a cubic meter.

Ans. (*a*) 16.4 cm^3; (*b*) 61.0 cm^3; (*c*) 35.3 ft^3

1.34. In a crystal of platinum, centers of individual atoms are 2.8 Å apart along the direction of closest packing. How many atoms would lie on a one-centimeter length of a line in this direction?

Ans. 3.5×10^7

1.35. The blue iridescence of butterfly wings is due to striations that are 0.15 μm apart, as measured by the electron microscope. What is this distance in centimeters? How does this spacing compare with the wavelength of blue light, about 4 500 Å?

Ans. 1.5×10^{-5} cm, $\frac{1}{3}$ wavelength of blue light

1.36. The thickness of a soap bubble film at its thinnest (bimolecular) stage is about 60 Å. (*a*) What is this thickness in centimeters? (*b*) How does this thickness compare with the wavelength of yellow sodium light, which is 0.589 0 μm?

Ans. (*a*) 6.0×10^{-7} cm; (*b*) about 0.01 of wavelength of yellow light

1.37. An average man requires about 2.00 mg of riboflavin (vitamin B_2) per day. How many pounds of cheese would a man have to eat per day if this were his only source of riboflavin and if the cheese contained 5.5 μg riboflavin per gram?

Ans. 0.80 lb/day

1.38. When a sample of healthy human blood is diluted to 200 times its initial volume and microscopically examined in a layer 0.10 mm thick, an average of 30 red corpuscles are found in each 100×100 micrometer square. (*a*) How many red cells are in a cubic millimeter of blood? (*b*) The red blood cells have an average life of 1 month, and the adult blood volume is about 5 L. How many red cells are generated every second in the bone marrow of the adult?

Ans. (*a*) 6×10^6 cells/min^3; (*b*) 1×10^7 cells/s

1.39. A porous catalyst for chemical reactions has an internal surface area of 800 m^2 per cm^3 of bulk material. Fifty percent of the bulk volume consists of the pores (holes), while the other 50 percent of the volume is made up of the solid substance. Assume that the pores are all cylindrical tubules of uniform diameter d and length l, and that the measured internal surface area is the total area of the curved surfaces of the tubules. What is the diameter of each pore? (*Hint*: Find the number of tubules per bulk cm^3, n, in terms of l and d, by using the formula for the volume of a cylinder. $V = \frac{1}{4}\pi d^2 l$. Then apply the surface-area formula, $S = \pi dl$, to the cylindrical surfaces of n tubules.)

Ans. 25 Å

COMPOUND UNITS

1.40. The density of water is 1.000 g/cm^2 at 4 °C. Calculate the density of water in pounds per cubic foot at the same temperature.

Ans. 62.4 lb/ft^3

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