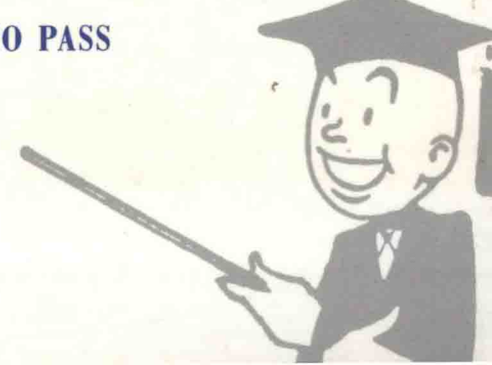
More than 30 million Schaum's Outlines sold!

College Chemistry

SCHAUM'S easy OUTlines

CRASH COURSE

- INCLUDES FULLY SOLVED PROBLEMS FOR EVERY TOPIC
- EXPERT TIPS FOR MASTERING COLLEGE CHEMISTRY
- ALL YOU NEED TO KNOW TO PASS THE COURSE



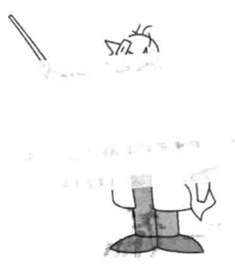
JEROME L. ROSENBERG
LAWRENCE M. EPSTEIN

SCHAUM'S *Easy* OUTLINES

COLLEGE CHEMISTRY

BASED ON SCHAUM'S
Outline of College Chemistry
BY JEROME L. ROSENBERG AND
LAWRENCE M. EPSTEIN

ABRIDGEMENT EDITOR:
PHILIP H. RIEGER



SCHAUM'S OUTLINE SERIES
McGRAW-HILL, INC.

New York San Francisco Washington, D.C. Auckland Bogotá Caracas Lisbon London Madrid Mexico City Milan Montreal New Delhi San Juan Singapore Sydney Tokyo Toronto

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. He is Professor Emeritus of Biological Sciences at the University of Pittsburgh.

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 was Associate Professor and supervisor of the General Chemistry program at the University of Pittsburgh until he retired in 1986.

PHILIP H. RIEGER is a graduate of Reed College and earned a Ph.D. from Columbia University. He is Professor of Chemistry at Brown University, where he has taught since 1961.

1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 DOC DOC 9 0 9 8 7 6 5 4 3 2 1 0 9

ISBN 0-07-052714-8

Sponsoring Editor: Barbara Gilson
Production Supervisor: Tina Cameron
Editing Supervisor: Maureen B. Walker

McGraw-Hill

A Division of The McGraw-Hill Companies

Contents

Chapter 1

QUANTITIES AND UNITS

IN THIS CHAPTER:

- ✔ *Significant Figures*
- ✔ *Propagation of Errors*
- ✔ *The International System of Units*
- ✔ *Dimensional Analysis*
- ✔ *Estimation of Numerical Answers*

Introduction

Most of the measurements and calculations in chemistry are concerned with quantities such as pressure, volume, mass, and energy. Every quantity 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. The *number* indicates how many of the reference units are contained in the quantity being measured. If we say that the mass of a sample is 20 grams, we mean that the mass is 20 times the mass of 1 gram, the unit of mass chosen for comparison. Although *20 grams* has the dimension of mass, *20* is a pure dimensionless number, being the ratio of two masses, that of the sample and that of the reference, 1 gram.

Significant Figures

The numerical value of every observed measurement is an approximation, since no physical measurement—of temperature, mass, volume, etc.—is ever exact. The accuracy of a measurement is always limited by the reliability of the measuring instrument.

Suppose that the recorded length of an object is 15.7 cm. By convention, this means that the length was measured to the *nearest* 0.1 cm and that its exact value lies between 15.65 and 15.75 cm. If this measurement were exact to the nearest 0.01 cm, it would have been recorded as 15.70 cm. We say that the first measurement is accurate to 3 significant figures and the second to 4.

A recorded volume of 2.8 L represents two significant figures. If this same volume were written 0.028 m^3, it would still contain only two significant figures. Zeroes appearing as the first digits of a number are not significant, since they merely locate the decimal point.

We often use *scientific notation* to express very large or very small numbers, indicating the number of significant figures by the number multiplied by 10^x. Thus, for example,

$$22400 = 2.24 \times 10^4 \qquad 0.00306 = 3.06 \times 10^{-3}$$

When two exponentials are multiplied (or divided), the exponents are added (or subtracted). For example,

$$(1.5 \times 10^5) \times (2.0 \times 10^{-3}) = 3.0 \times 10^2$$

$$(4.0 \times 10^7)/(2.0 \times 10^2) = 2.0 \times 10^5$$

When an exponential is raised to a power, the exponents are multiplied; for example,

$$(10^4)^{-2} = 10^{-8} \qquad (10^6)^{1/2} = 10^3$$

Some numbers are exact. These include π (3.14159...), numbers arising from counting (e.g., the number of experimental determinations of an observed measurement), and numbers which involve a definition (the mass of one atom of ^{12}C is exactly 12 u and the conversion of cm to m involves exactly 10^{-2} m/cm).

A number is rounded off to the desired number of significant figures by dropping one or more digits from the right. When the first digit dropped is less than 5, the last digit retained should remain unchanged; when it is greater than 5, the last digit is rounded up. When the digit dropped is exactly 5, the number retained is rounded up or down to get an even number. When more than one digit is dropped, rounding off should be done in a block, not one digit at a time.

Illustration:

The following numbers are accurate to two significant figures: 2.3, 0.023, 2.3×10^4.

The following numbers are rounded to two significant figures: $64500 \rightarrow 6.4 \times 10^4$, $5.75 \rightarrow 5.8$, $1.653 \rightarrow 1.7$, $1.527 \rightarrow 1.5$

Propagation of Errors

When we perform a calculation using numbers of limited accuracy, the result should be written with the appropriate number of significant figures.

When we add or subtract numbers, the number of significant figures in the answer is limited by the number with the smallest number of significant figures to the right of the decimal, e.g.,

$$4.20 + 1.6523 + 0.015 = 5.8673 \rightarrow 5.87$$

This rule is an approximation to a more exact statement that the error in a sum or difference is the square root of the sum of the squares of the errors in the numbers being added or subtracted. Thus in the above example, the error in the result is

$$\sqrt{(0.01)^2 + (0.0001)^2 + (0.001)^2} = 0.010$$

When multiplying or dividing two numbers, the result should contain only as many significant figures as the least accurate factor without regard for the position of the decimal point, e.g.,

$$7.485 \times 8.61 = 64.4 \qquad 0.1642/1.52 = 0.108$$

This rule is an approximation to a more exact statement that the fractional error of a product or quotient is the square root of the sum of the squares of the fractional errors in the numbers being multiplied or divided.

Thus in the above examples, the fractional and absolute errors in the results are:

$$\sqrt{\left(\frac{0.001}{7.485}\right)^2+\left(\frac{0.01}{8.61}\right)^2}=0.0012 \qquad 0.0012\times 64.4=0.08\approx 0.1$$

$$\sqrt{\left(\frac{0.0001}{0.1642}\right)^2+\left(\frac{0.01}{1.52}\right)^2}=0.0066 \qquad 0.0066\times 0.108=0.0007\approx 0.001$$

The approximate and more exact approaches sometimes lead to different results when numbers beginning with 1 or 2 are involved. For example, 9.84/8.9 = 1.106. The approximate method suggests writing the result to two significant figures, 1.1, but the more exact method leads to relative error of 0.011 and an absolute error of 0.012, so that writing the result with three significant figures, 1.11, is appropriate.

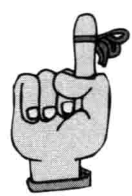
When propagating errors through more complex expressions, two or more steps of error estimation may be needed. For example,

$$29.7 \times (7.250 + 3.6554) = 29.7 \times 10.905 = 324$$

Remember:

Approximate rules for propagation of errors:

- When adding or subtracting numbers, the number of significant figures to the right of the decimal determines the accuracy of the result.
- When multiplying or dividing numbers, the total number of significant figures determines the accuracy of the result.



The International System of Units

Dimensional calculations are greatly simplified if a consistent set of units is employed. The three major reference dimensions for mechanics are *length, mass,* and *time*, but length can be measured in units of inches, feet, centimeters, meters, etc. Which should be used? The scientific community has made considerable progress toward a common system of reference units. This system is known as SI from the French name *Systéme International d'Unités.* In SI, the reference units for *length, mass,* and *time* are the *meter, kilogram,* and *second*, with symbols m, kg, and s, respectively.

To express quantities much larger or smaller than the standard units, multiples or submultiples of these units are used, as shown in the Table 1-1. Thus, 10^{-12} s is a picosecond (ps), and 10^3 m is a kilometer (km). Since for historical reasons the SI reference unit for mass, the kilogram, already has a prefix, multiples for mass should be derived by applying the multiplier to the unit *gram* rather than to the *kilogram.* Thus 10^{-9} kg is a microgram (10^{-6} g), abbreviated μg.

Table 1-1 Multiples and Submultiples for Units

Prefix	Abbr.	Multiplier	Prefix	Abbr.	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}

Many non-SI units remain in common use; some of these are given in the Table 1-2.

Compound units can be derived by applying algebraic operations to the simple units. For example, the SI units of volume and density are m^3 and kg/m^3, since

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

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

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

Quantity	Unit Name	Unit Symbol	Definition
length	angstrom inch	Å in	10^{-10} m 2.54×10^{-2} m
volume	cubic meter liter cubic centimeter	m^3 L cm^3, mL	(SI unit) 10^{-3} m^3 10^{-6} m^3
mass	atomic mass unit pound	u lb	1.6605×10^{-27} kg 0.45359 kg
density	gram per milliliter	g/mL or g/cm^3	10^3 kg/m^3
force	newton	N	$kg{\cdot}m/s^2$ (SI unit)

Note that symbols for multiplied units may be separated by a dot or a space, e.g., kg•s or kg s. Symbols for divided units may be written with a solidus or an exponent, e.g., m/s or $m{\bullet}s^{-1}$ or m s^{-1}.

Temperature is an independent dimension which cannot be defined in terms of mass, length, and time. The SI unit of temperature is the *kelvin* (K), defined as 1/273.16 times the *triple point* temperature of water (the temperature at which ice, liquid water, and water vapor coexist at equilibrium). 0 K is the absolute zero of temperature.

On the *Celsius* (or *centigrade*) scale, a temperature difference of 1°C is 1 K (exactly). The normal boiling point of water is 100°C, the normal freezing point 0°C, and absolute zero –273.15°C. On the *Fahrenheit* scale, a temperature difference of 1°F is 5/9 K (exactly). The boiling point and freezing point of water, and absolute zero are 212°F, 32°F and –459.67°F, respectively. Conversions from one temperature scale to another make use of the following equations:

$$t/°\text{C} = T/\text{K} - 273.15$$

$$t/°\text{F} = (9/5)(t/°\text{C}) + 32$$

In these equations, we have used a particularly convenient notation: t/°C, T/K, and t/°F refer to the numerical values of temperatures on the Celsius, Kelvin, and Fahrenheit scales. By dividing the temperature in degrees Celsius by 1°C, we obtain a pure number.

There Is Good News and Bad News!

The good news is that the SI system of units is consistent so that substitution of quantities with SI units into an equation will give a result in SI units.

The bad news is that many chemical calculations involve non-SI units.

Dimensional Analysis

Units are a necessary part of the specification of a physical quantity. 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, for example:

$$6\ \text{L} + 2\ \text{L} = 8\ \text{L} \qquad (5\ \text{cm})(2\ \text{cm}^2) = 10\ \text{cm}^3$$

In solving problems, one often can be guided by the units to the proper way of combining the given values. Some textbooks refer to this method as the *factor-label* or *unit-factor* method; we will call it *dimensional analysis.* In essence, one goes from a given unit to the desired unit by multiplying or dividing such that unwanted units cancel. For example, consider converting 5.00 inches to centimeters, given the conversion factor 2.54 cm/in. We might try two approaches:

$$5.00\ \text{in} \times \frac{\text{in}}{2.54\ \text{cm}} = 1.97\ \text{in}^2/\text{cm} \qquad 5.00\ \text{in} \times \frac{2.54\ \text{cm}}{\text{in}} = 12.7\ \text{cm}$$

The first try gives nonsensical units, signaling a misuse of the conversion factor. On the second try, the inch units cancel, leaving the desired centimeter units.

Estimation of Numerical Answers

If you input the numbers correctly into your calculator, the answer will be correct. But will you recognize an incorrect answer? You will if you obtain an approximate answer by visual inspection. Especially important is the 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. For example, consider a calculation of the power required to raise a 639 kg mass 20.74 m in 2.120 minutes:

$$\frac{639\ \text{kg} \times 20.74\ \text{m} \times 9.81\ \text{m s}^{\pm 2}}{2.120\ \text{min} \times 60\ \text{s/min}} = 1022\ \text{J/s} = 1022\ \text{watts}$$

This example involves concepts and units which may be unfamiliar to you, so that you can't easily judge whether the result "makes sense," so we'll check the answer by estimation. Write each term in exponential notation, using just one significant figure. Then mentally combine the powers of 10 and the multipliers separately to estimate the result:

$$\frac{6 \times 10^2 \times 2 \times 10^1 \times 1 \times 10^1}{2 \times 6 \times 10^1} = \frac{12 \times 10^4}{12 \times 10^1} = 1000$$

The estimate agrees with the calculation to one significant figure, showing that the calculation is very likely correct.

Problems

1.1 The color of light depends on its wavelength. Red light has a wavelength on the order of 7.8×10^{-7} m. Express this length in micrometers, in nanometers, and in angstroms.

Ans. 0.78 μm, 780 nm, 7800 Å

1.2 The blue iridescence of butterfly wings is due to striations that are 0.15 μm apart, as measured by an electron microscope. What is this distance in centimeters? How does this spacing compare with the wavelength of blue light, about 4500 Å?

Ans. 1.5×10^{-5} cm, 1/3 the wavelength

1.3 In a crystal of Pt, individual atoms are 2.8 Å apart along the direction of closest packing. How many atoms would lie on a 1.00-cm length of a line in this direction? *Ans.* 3.5×10^7

1.4 The bromine content of average ocean water is 65 parts by weight per million. Assuming 100% recovery, how many cubic meters of ocean water must be processed to produce 0.61 kg of bromine? Assume that the density of sea water is 1.0×10^3 kg/m^3.

Ans. 9.4 m^3

1.5 Find the volume in liters of 40 kg of carbon tetrachloride, whose density is 1.60 g/cm^3. *Ans.* 25 L

1.6 A sample of concentrated sulfuric acid is 95.7% H_2SO_4 by weight and its density is 1.84 g/cm^3. (a) How many grams of pure H_2SO_4 are contained in one liter of the acid? (b) How many cubic centimeters of acid contain 100 g of pure H_2SO_4?

Ans. (a) 1.76×10^3 g; (b) 56.8 cm^3

1.7 A quick method of determining density utilizes Archimedes' principle, which states that the buoyant force on an immersed object is equal to the weight of the liquid displaced. A bar of magnesium metal attached to a balance by a fine thread weighed 31.13 g in air and 19.35 g when completely immersed in hexane (density 0.659 g/cm^3). Calculate the density of this sample of magnesium in SI units. *Ans.* 1741 kg/m^3

1.8 A piece of gold leaf (density 19.3 g/cm^3) weighing 1.93 mg can be beaten into a transparent film covering an area of 14.5 cm^3. What is the volume of 1.93 mg of gold? What is the thickness of the transparent film in angstroms?

Ans. 1.00×10^{-4} cm^3, 690 Å

1.9 Sodium metal has a very wide liquid range, melting at 98°C and boiling at 892°C. Express the liquid range in degrees Celsius, kelvins, and degrees Fahrenheit. *Ans.* 794°C, 794 K, 1429°F

Chapter 2
MOLES AND EMPIRICAL FORMULAS

IN THIS CHAPTER:

- ✔ *Atoms and Isotopes*
- ✔ *Relative Atomic Masses*
- ✔ *Moles and Molar Masses*
- ✔ *Empirical Formula from Composition*
- ✔ *Chemical Formulas from Mass Spectrometry*

Atoms and Isotopes

In the atomic theory proposed by John Dalton in 1805, all atoms of a given element were assumed to be identical. Eventually it was realized that atoms of a given element are not necessarily identical; an element can exist in several *isotopic* forms that differ in atomic mass.

Every atom has a positively charged nucleus and one or more electrons that form a charge cloud surrounding the nucleus. The nucleus contains over 99.9% of the total mass of the atom. Every nucleus may be described as being made up of two different kinds of particles, *protons* and *neutrons*, collectively called *nucleons*. Protons and neutrons have nearly the same mass, but only the proton is charged, so that the total charge of a nucleus is equal to the number of protons times the charge of one proton. The magnitude of the proton charge is equal to that of the electron so that a neutral atom has an equal number of protons and electrons.

The atoms of all isotopes of an element have the same number of protons, the *atomic number*, *Z*. The nuclei of different isotopes differ, however, in the number of neutrons and therefore in the total number of nucleons per nucleus. The total number of nucleons is *A*, the *mass number*. Atoms of different isotopic forms of an element, *nuclides*, are distinguished by using the mass number as a left superscript on the symbol of the element, e.g., ^{15}N refers to the isotope of N with mass number 15.

You Need to Know

Atomic nuclei contain protons and neutrons.

- An element is defined by the nuclear charge; the atomic number Z = number of protons.
- Mass number *A* = no. protons + no. neutrons.
- Isotopes of an element have the same Z, but different *A*'s.

Relative Atomic Masses

Because the mass of an atom is very small, it is convenient to define a special unit that avoids large negative exponents. This unit, called the *atomic mass unit* and designated by the symbol u (some authors use the abbreviation amu), is defined as exactly 1/12 the mass of a ^{12}C atom.

Thus the mass of a ^{12}C atom is exactly 12 u. The masses and abundances of some other nuclides are listed in Table 2-1.

Naturally occurring silicon is 92.23% ^{28}Si, 4.67% ^{29}Si, and 3.10% ^{30}Si. For chemical purposes, it is sufficient to know the *average mass* of a silicon atom in this isotopic mixture. These average masses are designated by $A_r(E)$, where E is the symbol for the particular element. For example, the average mass of silicon atoms is

$$A_r(\text{Si}) = 0.9223\times27.977 + 0.0467\times28.976 + 0.0310\times29.974 = 28.085$$

The term *atomic mass* will be understood to mean average atomic mass; *nuclidic mass* refers to one particular isotope of an element. Atomic masses are used in nearly all chemical calculations.

Table 2-1. Some Nuclidic Masses in Atomic Mass Units

^{1}H	99.985%	1.00783 u	^{16}O	99.76 %	15.99491 u
^{2}H	0.015	2.01410	^{17}O	0.04	16.99913
^{12}C	98.89	12.00000	^{18}O	0.20	17.99916
^{13}C	1.11	13.00335	^{28}Si	92.23	27.97693
^{14}N	99.64	14.00307	^{29}Si	4.67	28.97649
^{15}N	0.36	15.00011	^{30}Si	3.10	29.97377
^{32}S	95.0	31.97207	^{34}S	4.22	33.96786

Remember:

Nuclidic mass = mass of a particular nuclide, relative to mass of ^{12}C, exactly 12 u. Atomic masses, A_r, commonly found in tables, are averages of nuclidic masses, weighted to reflect the isotopic composition of the elements.

Moles and Molar Masses

Most chemical experiments involve enormous numbers of atoms or molecules. For this reason, the SI system of units defines the *mole*, abbreviated mol, as the amount of a substance that contains the same number of atoms as 12 g of ^{12}C. This number is called *Avogadro's number*, $N_A = 6.0221 \times 10^{23}$ mol^{-1}. The mole thus is a counting term analogous to "dozen." Just as a dozen eggs corresponds to 12 eggs, a mole of atoms is 6.0221×10^{23} atoms. The mole can be applied to counting atoms, molecules, ions, electrons, protons, neutrons, etc.—it always corresponds to N_A of the counted species. The mass of one mole of an element with atomic mass A_r(E) u is $N_A A_r$ u or simply A_r g/mol, e.g., the atomic mass of gold is 197.0 u or 197.0 g/mol.

One mole of atoms, molecules, ions, etc., contains one Avogadro's number ($N_A = 6.0221 \times 10^{23}$ mol^{-1}) of that species.

The chemical symbol for an element—H, C, O, etc.—is used to designate that element. Molecular substances consist of independent molecules containing two or more atoms bound together. A molecular formula specifies the identity and number of the atoms in the molecule. For instance, the formula for carbon dioxide is CO_2, one carbon atom and two oxygen atoms. The molecular mass of CO_2 is $A_r(C) + 2A_r(O) = 12.0107 + 2 \times 15.9994 = 44.0095$ u. The *molar mass* of CO_2 is the mass in grams numerically equal to the molecular mass in u, 44.0095 g/mol, i.e., 44.0095 g contains N_A CO_2 molecules.

Many common substances are ionic, e.g., sodium chloride, NaCl. A crystal of NaCl contains sodium ions, Na^+, and chloride ions, Cl^-, arranged in a regular spatial array. Although there are no NaCl molecules, the formula indicates the relative number of atoms of each element present in the crystal, and we can speak of the molar mass of NaCl, 22.98977 + 35.4527 = 58.4425 g/mol, as the mass of sodium chloride which contains N_A sodium ions and N_A chloride ions. We also speak of the molar mass of an ion, e.g., OH^-, 15.9994 + 1.00794 = 17.0073 g/mol, as the mass of N_A hydroxide ions.