Manahan, Stanley E. "INTRODUCTION TO CHEMISTRY"
Fundamentals of Environmental Chemistry
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1 INTRODUCTION TO CHEMISTRY

1.1 CHEMISTRY AND ENVIRONMENTAL CHEMISTRY

Chemistry is defined as the science of matter. Therefore, it deals with the air we breathe, the water we drink, the soil that grows our food, and vital life substances and processes. Our own bodies contain a vast variety of chemical substances and are tremendously sophisticated chemical factories that carry out an incredible number of complex chemical processes.

There is a tremendous concern today about the uses—and particularly the misuses—of chemistry as it relates to the environment. Ongoing events serve as constant reminders of threats to the environment ranging from individual exposures to toxicants to phenomena on a global scale that may cause massive, perhaps catastrophic, alterations in climate. These include, as examples, evidence of a perceptible warming of climate; record weather events—particularly floods—in the United States in the 1990s; and air quality in Mexico City so bad that it threatens human health. Furthermore, large numbers of employees must deal with hazardous substances and wastes in laboratories and the workplace. All such matters involve environmental chemistry for understanding of the problems and for arriving at solutions to them.

Environmental chemistry is that branch of chemistry that deals with the origins, transport, reactions, effects, and fates of chemical species in the water, air, earth, and living environments and the influence of human activities thereon.[1] A related discipline, **toxicological chemistry**, is the chemistry of toxic substances with emphasis upon their interaction with biologic tissue and living systems.[2] Besides its being an essential, vital discipline in its own right, environmental chemistry provides an excellent framework for the study of chemistry, dealing with "general chemistry," organic chemistry, chemical analysis, physical chemistry, photochemistry, geochemistry, and biological chemistry. By necessity it breaks down the barriers that tend to compartmentalize chemistry as it is conventionally addressed. Therefore, this book is written with two major goals—to provide an overview of chemical science within an environmental chemistry framework and to provide the basics of environmental chemistry for those who need to know about this essential topic for their professions or for their overall education.

1.2 A MINI-COURSE IN CHEMISTRY

It is much easier to learn chemistry if one already knows some chemistry! That is, in order to go into any detail on any chemical topic, it is extremely helpful to have some very rudimentary knowledge of chemistry as a whole. For example, a crucial part of chemistry is an understanding of the nature of chemical compounds, the chemical formulas used to describe them, and the chemical bonds that hold them together; these are topics addressed in Chapter 3 of this book. However, to understand these concepts, it is very helpful to know some things about the chemical reactions by which chemical compounds are formed, as addressed in Chapter 4. To work around this problem, Chapter 1 provides a highly condensed, simplified, but meaningful overview of chemistry to give the reader the essential concepts and terms required to understand more-advanced chemical material.

1.3 THE BUILDING BLOCKS OF MATTER

All matter is composed of only about a hundred fundamental kinds of matter called **elements**. Each element is made up of very small entities called **atoms**; all atoms of the same element behave identically chemically. The study of chemistry, therefore, can logically begin with elements and the atoms of which they are composed.

Subatomic Particles and Atoms

Figure 1.1 represents an atom of deuterium, a form of the element hydrogen. It is seen that such an atom is made up of even smaller **subatomic particles**—positively charged **protons**, negatively charged **electrons**, and uncharged (neutral) **neutrons.** Protons and neutrons have relatively high masses compared with electrons and are contained in the positively charged **nucleus** of the atom. The nucleus has essentially all the mass, but occupies virtually none of the volume, of

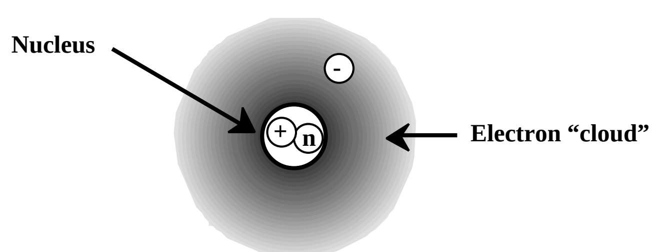
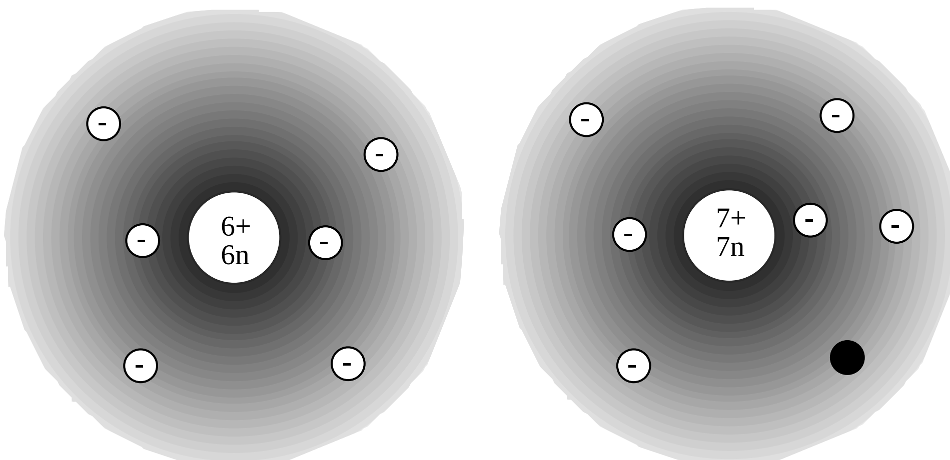


Figure 1.1 Representation of a deuterium atom. The nucleus contains one proton (+) and one neutron (n). The electron (-) is in constant, rapid motion around the nucleus, forming a cloud of negative electrical charge, the density of which drops off with increasing distance from the nucleus.

the atom. An uncharged atom has the same number of electrons as protons. The electrons in an atom are contained in a cloud of negative charge around the nucleus that occupies most of the volume of the atom.

Atoms and Elements

All of the literally millions of different substances are composed of only around 100 elements. Each atom of a particular element is chemically identical to every other atom and contains the same number of protons in its nucleus. This number of protons in the nucleus of each atom of an element is the **atomic number** of the element. Atomic numbers are integers ranging from 1 to more than 100, each of which denotes a particular element. In addition to atomic numbers, each element has a name and a **chemical symbol**, such as carbon, C; potassium, K (for its Latin name kalium); or cadmium, Cd. In addition to atomic number, name, and chemical symbol, each element has an **atomic mass** (atomic weight). The atomic mass of each element is the average mass of all atoms of the element, including the various isotopes of which it consists. The **atomic mass unit**, **u** (also called the **dalton**), is used to express masses of individual atoms and molecules (aggregates of atoms). These terms are summarized in Figure 1.2.



An atom of carbon, symbol C. Each C atom has 6 protons (+) in its nucleus, so the atomic number of C is 6. The atomic mass of C is 12.

An atom of nitrogen, symbol N. Each N atom has 7 protons (+) in its nucleus, so the atomic number of N is 7. The atomic mass of N is 14.

Figure 1.2 Atoms of carbon and nitrogen

Although atoms of the same element are *chemically* identical, atoms of most elements consist of two or more **isotopes** that have different numbers of neutrons in their nuclei. Some isotopes are **radioactive isotopes** or **radionuclides**, which have unstable nuclei that give off charged particles and gamma rays in the form of **radioactivity**. This process of **radioactive decay** changes atoms of a particular element to atoms of another element.

Throughout this book reference is made to various elements. A list of the known elements is given on page 120 at the end of Chapter 3. Fortunately, most of the chemistry covered in this book requires familiarity with only about 25 or 30 elements. An abbreviated list of a few of the most important elements that the reader should learn at this point is given in Table 1.1.

Table 1.1 List of Some of the More Important Common Elements

Element	Symbol	Atomic Number	Atomic Mass (relative to carbon-12)
Argon	Ar	18	39.948
Bromine	Br	35	79.904
Calcium	Ca	20	40.08
Carbon	C	6	12.01115
Chlorine	Cl	17	35.453
Copper	Cu	29	63.546
Fluorine	F	9	18.998403
Helium	He	2	4.00260
Hydrogen	H	1	1.0080
Iron	Fe	26	55.847
Magnesium	Mg	12	24.305
Mercury	Hg	80	200.59
Neon	Ne	10	20.179
Nitrogen	N	7	14.0067
Oxygen	O	8	15.9994
Potassium	K	19	39.0983
Silicon	Si	14	28.0855
Sodium	Na	11	22.9898
Sulfur	S	16	32.06

The Periodic Table

When elements are considered in order of increasing atomic number, it is observed that their properties are repeated in a periodic manner. For example, elements with atomic numbers 2, 10, and 18 are gases that do not undergo chemical reactions and consist of individual molecules, whereas those with atomic numbers larger by one—3, 11, and 19—are unstable, highly reactive metals. An arrangement of the elements in a manner that reflects this recurring behavior is known as the **periodic table** (Figure 1.3). The periodic table is extremely useful in understanding chemistry and predicting chemical behavior. The entry for each element in the periodic table gives the element's atomic number, name, symbol, and atomic mass. More-detailed versions of the table include other information as well.

Transition Elements

Period	IA 1	IIA 2	IIIB 3	IVB 4	VB 5	VIB 6	VIIB 7	VIII 8	VIII 9	VIII 10	IB 11	IIB 12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	Noble gases 18
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.3											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.9	36 **Kr** 83.8
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** 98.91	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9 *	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (210)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227) **	104 **Rf** (261)	105 **Ha** (262)	106 **Sg** (263)	107 **Ns** (262)	108 **Ha** (265)	109 **Mt** (266)									

Inner Transition Elements

Lanthanide series * 6	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** 144.9	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.2	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
Actinide series ** 7	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** 237.0	94 **Pu** 239.1	95 **Am** 243.1	96 **Cm** 247.1	97 **Bk** 247.1	98 **Cf** 252.1	99 **Es** 252.1	100 **Fm** 257.1	101 **Md** 256.1	102 **No** 259.1	103 **Lr** 260.1

Figure 1.3 The periodic table of the elements.

Features of the Periodic Table

The periodic table gets its name from the fact that the properties of elements are repeated periodically in going from left to right across a horizontal row of elements. The table is arranged such that an element has properties similar to those of other elements above or below it in the table. Elements with similar chemical properties are called **groups** of elements and are contained in vertical columns in the periodic table.

1.4. CHEMICAL BONDS AND COMPOUNDS

Only a few elements, particularly the noble gases, exist as individual atoms; most atoms are joined by chemical bonds to other atoms. This can be illustrated very simply by elemental hydrogen, which exists as **molecules**, each consisting of 2 H atoms linked by a **chemical bond** as shown in Figure 1.4. Because hydrogen molecules contain 2 H atoms, they are said to be diatomic and are denoted by the **chemical formula** H_2. The H atoms in the H_2 molecule are held together by a **covalent bond** made up of 2 electrons, each contributed by one of the H atoms, and shared between the atoms.

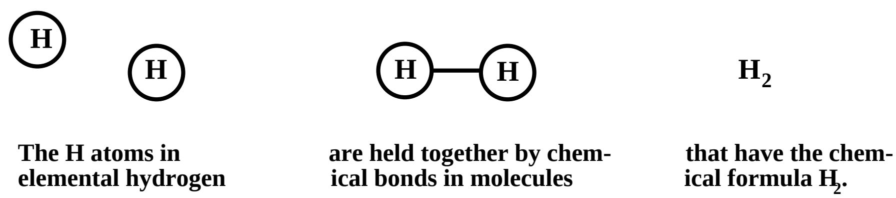


Figure 1.4 Molecule of H_2.

Chemical Compounds

Most substances consist of two or more elements joined by chemical bonds. As an example, consider the chemical combination of the elements hydrogen and oxygen shown in Figure 1.5. Oxygen, chemical symbol O, has an atomic number of 8 and an atomic mass of 16.00 and exists in the elemental form as diatomic molecules of O_2. Hydrogen atoms combine with oxygen atoms to form molecules in which 2 H atoms are bonded to 1 O atom in a substance with a chemical formula of H_2O (water). A substance such as H_2O that consists of a chemically bonded com-

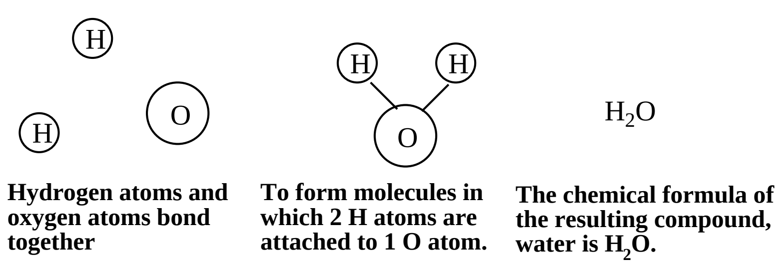


Figure 1.5 A molecule of water, H_2O, formed from 2 H atoms and 1 O atom held together by chemical bonds.

bination of two or more elements is called a **chemical compound**. (A chemical compound is a substance that consists of atoms of two or more different elements bonded together.) In the chemical formula for water the letters H and O are the chemical symbols of the two elements in the compound and the subscript 2 indicates that there are 2 H atoms per O atom. (The absence of a subscript after the O denotes the presence of just 1 O atom in the molecule.) Each of the chemical bonds holding a hydrogen atom to the oxygen atom in the water molecule is composed of two electrons shared between the hydrogen and oxygen atoms.

Ionic Bonds

As shown in Figure 1.6, the transfer of electrons from one atom to another produces charged species called **ions**. Positively charged ions are called **cations** and negatively charged ions are called **anions**. Ions that make up a solid compound are held together by **ionic bonds** in a **crystalline lattice** consisting of an ordered arrangement of the ions in which each cation is largely surrounded by anions and each anion by cations. The attracting forces of the oppositely charged ions in the crystalline lattice constitute the ionic bonds in the compound.

The formation of the ionic compound magnesium oxide is shown in Figure 1.6. In naming this compound, the cation is simply given the name of the element from which it was formed, magnesium. However, the ending of the name of the anion, ox*ide*, is different from that of the element from which it was formed, ox*ygen.*

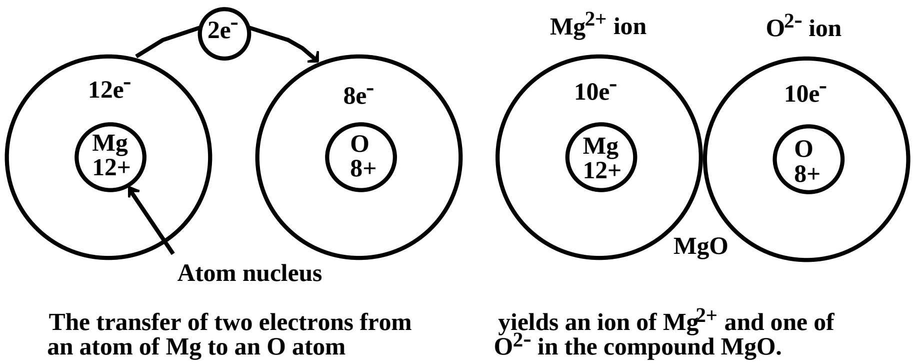
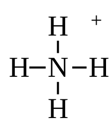


Figure 1.6 Ionic bonds are formed by the transfer of electrons and the mutual attraction of oppositely charged ions in a crystalline lattice.

Rather than individual atoms that have lost or gained electrons, many ions are groups of atoms bonded together covalently and having a net charge. A common example of such an ion is the ammonium ion, NH_4^+,



Ammonium ion, NH_4^+

consisting of 4 hydrogen atoms covalently bonded to a single nitrogen (N) atom and having a net electrical charge of +1 for the whole cation.

Summary of Chemical Compounds and the Ionic Bond

The preceding several pages have just covered some material on chemical compounds and bonds that are essential to understand chemistry. To summarize, these are the following:

- Atoms of two or more different elements can form *chemical bonds* with each other to yield a product that is entirely different from the elements.
- Such a substance is called a *chemical compound.*
- The *formula* of a chemical compound gives the symbols of the elements and uses subscripts to show the relative numbers of atoms of each element in the compound.
- *Molecules* of some compounds are held together by *covalent bonds* consisting of shared electrons.
- Another kind of compound consists of *ions* composed of electrically charged atoms or groups of atoms held together by *ionic bonds* that exist because of the mutual attraction of oppositely charged ions.

Molecular Mass

The average mass of all molecules of a compound is its **molecular mass** (formerly called molecular weight). The molecular mass of a compound is calculated by multiplying the atomic mass of each element by the relative number of atoms of the element, then adding all the values obtained for each element in the compound. For example, the molecular mass of NH_3 is 14.0 + 3 x 1.0 = 17.0. As another example consider the following calculation of the molecular mass of ethylene, C_2H_4.

1. The chemical formula of the compound is C_2H_4.
2. Each molecule of C_2H_4 consists of 2 C atoms and 4 H atoms.
3. From the periodic table or Table 1.1, the atomic mass of C is 12.0 and that of H is 1.0.
4. Therefore, the molecular mass of C_2H_4 is

$$\underbrace{12.0 + 12.0}_{\text{From 2 C atoms}} + \underbrace{1.0 + 1.0 + 1.0 + 1.0}_{\text{From 4 H atoms}} = 28.0$$

1.5. CHEMICAL REACTIONS AND EQUATIONS

Chemical reactions occur when substances are changed to other substances through the breaking and formation of chemical bonds. For example, water is produced by the chemical reaction of hydrogen and oxygen:

Hydrogen plus oxygen yields water

Chemical reactions are written as **chemical equations**. The chemical reaction between hydrogen and water is written as the **balanced chemical equation**

$$2H_2 + O_2 \rightarrow 2H_2O \tag{1.5.1}$$

in which the arrow is read as "yields" and separates the hydrogen and oxygen **reactants** from the water **product**. Note that because elemental hydrogen and elemental oxygen occur as *diatomic molecules* of H_2 and O_2, respectively, it is necessary to write the equation in a way that reflects these correct chemical formulas of the elemental form. All correctly written chemical equations are **balanced**, in that *they must show the same number of each kind of atom on both sides of the equation*. The equation above is balanced because of the following:

On the left

- There are 2 H_2 *molecules*, each containing 2 H *atoms* for a total of 4 H atoms on the left.
- There is 1 O_2 *molecule*, containing 2 O *atoms* for a total of 2 O atoms on the left.

On the right

- There are 2 H_2O *molecules* each containing 2 H *atoms* and 1 O atom for a total of 4 H atoms and 2 O atoms on the right.

The process of balancing chemical equations is relatively straightforward for simple equations. It is discussed in Chapter 4.

1.6. NUMBERS IN CHEMISTRY: EXPONENTIAL NOTATION

An essential skill in chemistry is the ability to handle numbers, including very large and very small numbers. An example of the former is Avogadro's number, which is discussed in detail in Chapters 2 and 3. Avogadro's number is a way of expressing quantities of entities such as atoms or molecules and is equal to 602,000,000,000,000,000,000,000. A number so large written in this decimal form is very cumbersome to express and very difficult to handle in calculations. It can be expressed much more conveniently in exponential notation. Avogadro's number in exponential notation is 6.02×10^{23}. It is put into decimal form by moving the decimal in 6.02 to the right by 23 places. Exponential notation works equally well to express very small numbers, such as 0.000,000,000,000,000,087. In exponential notation this number is 8.7×10^{-17}. To convert this number back to decimal form, the decimal point in 8.7 is simply moved 17 places to the left.

A number in exponential notation consists of a *digital number* equal to or greater than exactly 1 and less than exactly 10 (examples are 1.00000, 4.3, 6.913, 8.005, 9.99999) multiplied by a *power of 10* (10^{-17}, 10^{13}, 10^{-5}, 10^{3}, 10^{23}). Some examples of numbers expressed in exponential notation are given in Table 1.2. As seen in the second column of the table, a positive power of 10 shows the number of times that the digital number is multiplied by 10 and a negative power of 10 shows the number of times that the digital number is divided by 10.

Table 1.2 Numbers in Exponential and Decimal Form

Exponential form of number	Places decimal moved for decimal form	Decimal form
$1.37 \times 10^5 = 1.37 \times 10 \times 10 \times 10 \times 10 \times 10$	→ 5 places	137,000
$7.19 \times 10^7 = 7.19 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10$	→ 7 places	71,900,000
$3.25 \times 10^{-2} = 3.25/(10 \times 10)$	← 2 places	0.0325
$2.6 \times 10^{-6} = 2.6/(10 \times 10 \times 10 \times 10 \times 10 \times 10)$	← 6 places	0.000 0026
$5.39 \times 10^{-5} = 5.39/(10 \times 10 \times 10 \times 10 \times 10)$	← 5 places	0.000 0539

Addition and Subtraction of Exponential Numbers

An electronic calculator keeps track of exponents automatically and with total accuracy. For example, getting the sum $7.13 \times 10^3 + 3.26 \times 10^4$ on a calculator simply involves the following sequence:

$\boxed{7.13\text{ EE3}}$ $\boxed{+}$ $\boxed{3.26\text{ EE4}}$ $\boxed{=}$ $\boxed{3.97\text{ EE4}}$

where 3.97 EE4 stands for 3.97×10^4. To do such a sum manually, the largest number in the sum should be set up in the standard exponential notation form and each of the other numbers should be taken to the same power of 10 as that of the largest number as shown, below for the calculation of $3.07 \times 10^{-2} - 6.22 \times 10^{-3} + 4.14 \times 10^{-4}$:

3.07×10^{-2} (largest number, digital portion between 1 and 10)
$- 0.622 \times 10^{-2}$ (same as 6.22×10^{-3})
$+ 0.041 \times 10^{-2}$ (same as 4.1×10^{-4})

Answer: 2.49×10^{-2}

Multiplication and Division of Exponential Numbers

As with addition and subtraction, multiplication and division of exponential numbers on a calculator or computer is simply a matter of (correctly) pushing buttons. For example, to solve

$$\frac{1.39 \times 10^{-2} \times 9.05 \times 10^8}{3.11 \times 10^4}$$

on a calculator, the sequence below is followed:

$\boxed{1.39\text{ EE-2}}$ $\boxed{\times}$ $\boxed{9.05\text{ EE8}}$ $\boxed{\div}$ $\boxed{3.11\text{ EE4}}$ $\boxed{=}$ $\boxed{4.04\text{ EE2}}$ (same as 4.04×10^2)

In multiplication and division of exponential numbers, the digital portions of the numbers are handled conventionally. For the powers of 10, in multiplication exponents are added algebraically, whereas in division the exponents are subtracted algebraically. Therefore, in the preceding example,

$$\frac{1.39 \times 10^{-2} \times 9.05 \times 10^{8}}{3.11 \times 10^{4}}$$

the digital portion is

$$\frac{1.39 \times 9.05}{3.11} = 4.04$$

and the exponential portion is,

$$\frac{10^{-2} \times 10^{8}}{10^{4}} = 10^{2} \text{ (The exponent is -2 + 8 - 4)}$$

So the answer is 4.04 $\times 10^2$.

Example: Solve

$$\frac{7.39 \times 10^{-2} \times 4.09 \times 10^{5}}{2.22 \times 10^{4} \times 1.03 \times 10^{-3}}$$

without using exponential notation on the calculator.

Answer: Exponent of answer = $\underbrace{-2 + 5}_{\text{Algebraic addition of exponents in the numerator}} - \underbrace{(4 - 3)}_{\text{Algebraic subtraction of exponents in the denominator}} = 2$

$$\frac{7.39 \times 4.09}{2.22 \times 1.03} = 13.2 \quad \text{The answer is } 13.2 \times 10^{2} = 1.32 \times 10^{3}$$

Example: Solve

$$\frac{3.49 \times 10^{3}}{3.26 \times 10^{18} \times 7.47 \times 10^{-5} \times 6.18 \times 10^{-8}}$$

Answer: 2.32×10^{-4}

1.7 SIGNIFICANT FIGURES AND UNCERTAINTIES IN NUMBERS

The preceding section illustrated how to handle very large and very small numbers with *exponential notation*. This section considers **uncertainties** in numbers, taking into account the fact that numbers are known only to a certain degree of **accuracy**. The accuracy of a number is shown by how many **significant figures** or **significant digits** it contains. This can be illustrated by considering the atomic masses of elemental boron and sodium. The atomic mass of boron is given as 10.81. Written in this way, the number expressing the atomic mass of boron contains four significant digits—the 1, the 0, the 8, and the l. It is understood to have an uncertainty of + or - 1 in the last digit, meaning that it is really 10.81±0.01. The atomic mass of sodium is given as 22.98977, a number with seven significant digits understood to mean 22.98977±0.00001. Therefore, the atomic mass of sodium is known with more *certainty* than that of boron. The atomic masses in Table 1.1 reflect the fact that they are known with much more certainty for some elements (for example fluorine, 18.998403) than for others (for example, calcium listed with an atomic mass of 40.08).

The rules for expressing significant digits are summarized in Table 1.3. It is important to express numbers to the correct number of significant digits in chemical calculations and in the laboratory. The use of too many digits implies an accuracy in the number that does not exist and is misleading. The use of too few significant digits does not express the number to the degree of accuracy to which it is known.

Table 1.3 Rules for Use of Significant Digits

Example number	Number of significant digits	Rule
11.397	5	1. Non-zero digits in a number are always significant. The 1,1,3,9, and 7 in this number are each significant.
140.039	6	2. Zeros between non-zero digits are significant. The 1, 4, 0, 0, 3, and 9 in this number are each significant.
0.00329	3	3. Zeros on the left of the first non-zero digit are not significantbecause they are used only to locate the decimal point. Only 3, 2, and 9 in this number are significant.
70.00	4	4. Zeros to the right of a decimal point that are preceded by a significant figure are significant. All three 0s, as well as the 7, are significant.
32 000	Uncertain	5. The number of significant digits in a number with zeros to the left, but not to the right of a decimal point (1700, 110 000) may be uncertain. Such numbers should be written in exponential notation.
3.20×10^3	3	6. The number of significant digits in a number written in exponential notation is equal to the number of significant digits in the decimal portion.
Exactly 50	Unlimited	7. Some numbers, such as the amount of money that one expects to receive when cashing a check or the number of children claimed for income tax exemptions, are defined as **exact numbers** without any uncertainty.

Exercise: Referring to Table 1.3, give the number of significant digits and the rule(s) upon which they are based for each of the following numbers:

(a) 17.000	(b) 9.5378	(c) 7.001
(d) $50	(e) 0.00300	(f) 7400
(g) 6.207×10^{-7}	(h) 13.5269184	(i) 0.05029

Answers: (a) 5, Rule 4; (b) 5, Rule 1; (c) 4, Rule 2; (d) exact number; (e) 3, Rules 3 and 4; (f) uncertain, Rule 5; (g) 4, Rule 6; (h) 9, Rule 1; (i) 4 Rules 2 and 3

Significant Figures in Calculations

After numbers are obtained by a laboratory measurement, they are normally subjected to mathematical operations to get the desired final result. It is important that the answer have the correct number of significant figures. It should not have so few that accuracy is sacrificed or so many that an unjustified degree of accuracy is implied. The two major rules that apply, one for addition/subtraction, the other for multiplication/division, are the following:

1. In addition and subtraction, the number of digits retained to the right of the decimal point should be the same as that in the number in the calculation with the fewest such digits.

   Example: 273.591 + 1.00327 + 229.13 = 503.72427 is rounded to 503.72 because 229.13 has only two significant digits beyond the decimal.

   Example: 313.4 + 11.0785 + 229.13 = 553.6085 is rounded to 553.6 because 313.4 has only one significant digit beyond the decimal.

2. The number of significant figures in the result of multiplication/division should be the same as that in the number in the calculation having the fewest significant figures.

   Example: $\dfrac{3.7218 \text{ x } 4.019 \text{ x } 10^{-3}}{1.48} = 1.0106699 \times 10^{-2}$ is rounded to

   $1.01 \text{ x} 10^{-2}$ (3 significant figures because 1.48 has only 3 significant figures)

   Example: $\dfrac{5.27821 \times 10^{7} \times 7.245 \times 10^{-5}}{1.00732} = 3.7962744 \times 10^{3}$ is rounded

   to 3.796×10^{3} (4 significant figures because 7.245 has only 4 significant figures)

It should be noted that an exact number is treated in calculations as though it has an unlimited number of significant figures.

Exercise: Express each of the following to the correct number of significant figures:

(a) $13.1 + 394.0000 + 8.1937$ (b) $1.57 \times 10^{-4} \times 7.198 \times 10^{-2}$

(c) $189.2003 - 13.47 - 2.563$ (d) $221.9 \times 54.2 \times 123.008$

(e) $\dfrac{603.9 \times 21.7 \times 0.039217}{87}$ (f) $\dfrac{3.1789 \times 10^{-3} \times 7.000032 \times 10^{4}}{27.130921}$

(g) $100 \times 0.7428 \times 6.82197$ (where 100 is an exact number)

Answers: (a) 415.3, (b) 1.13×10^{-5}, (c) 173.17, (d) 1.48×10^{6}, (e) 5.9, (f) 8.2019, (g) 506.7

Rounding Numbers

With an electronic calculator it is easy to obtain a long string of digits that must be rounded to the correct number of significant figures. The rules for doing this are the following:

1. If the digit to be dropped is 0, 1, 2, 3, or 4, leave the last digit unchanged

   Example: Round 4.17821 to 4 significant digits
   Answer: 4.178
   Last retained digit Digit to be dropped

2. If the digit to be dropped is 5,6,7,8 or 9, increase the last retained digit by 1

   Example: Round 4.17821 to 3 significant digits
   Answer: 4.18
   Last retained digit Digit to be dropped

Use of Three Significant Digits

It is possible to become thoroughly confused about how many significant figures to retain in an answer. In such a case it is often permissible to use 3 significant figures. Generally, this gives sufficient accuracy without doing grievous harm to the concept of significant figures.

1.8 MEASUREMENTS AND SYSTEMS OF MEASUREMENT

The development of chemistry has depended strongly upon careful measurements. Historically, measurements of the quantities of substances reacting and produced in chemical reactions have allowed the explanation of the fundamental nature of chemistry. Exact measurements continue to be of the utmost importance in chemistry and are facilitated by increasing sophisticated instrumentation. For example, atmospheric chemists can determine a small degree of stratospheric ozone depletion by measuring minute amounts of ultraviolet radiation absorbed by ozone with satellite-mounted instruments. Determinations of a part per trillion or less of a toxic substance in water may serve to trace the source of a hazardous pollutant. This section discusses the basic measurements commonly made in chemistry and environmental chemistry.

SI Units of Measurement

Several systems of measurement are used in chemistry and environmental chemistry. The most systematic of these is the **International System of Units**, abbreviated **SI**, a self-consistent set of units based upon the metric system recommended in 1960 by the General Conference of Weights and Measures to simplify and make more logical the many units used in the scientific and engineering community. Table 1.4 gives the seven base SI units from which all others are derived.

Multiples of Units

Quantities expressed in science often range over many orders of magnitude (many factors of 10). For example, a mole of molecular diatomic nitrogen contains 6.02×10^{23} N_2 molecules and very small particles in the atmosphere may be only about 1×10^{-6} meters in diameter. It is convenient to express very large or very small multiples by means of **prefixes** that give the number of times that the basic unit is multiplied. Each prefix has a name and an abbreviation. The ones that are used in this book, or that are most commonly encountered, are given in Table 1.5.

Metric and English Systems of Measurement

The **metric** system has long been the standard system for scientific measurement and is the one most commonly used in this book. It was the first to use multiples of 10 to designate units that differ by orders of magnitude from a basic unit. The **English** system is still employed for many measurements encountered in normal everyday activities in the United States, including some environmental engineering measurements. Bathroom scales are still calibrated in pounds, well depths may be given in feet, and quantities of liquid wastes are frequently expressed as gallons or barrels. Furthermore, English units of pounds, tons, and gallons are still commonly used in commerce, even in the chemical industry. Therefore, it is still necessary to have some familiarity with this system; conversion factors between it and metric units are given in this book.

1.9 UNITS OF MASS

Mass expresses the degree to which an object resists a change in its state of rest or motion and is proportional to the amount of matter in the object. **Weight** is the gravitational force acting upon an object and is proportional to mass. An object weighs much less in the gravitational force on the Moon’s surface than on Earth, but the object’s mass is the same in both places (Figure 1.7). Although mass and weight are not usually distinguished from each other in everyday activities, it is important for the science student to be aware of the differences between them.

Table 1.4 Units of the International System of Units, SI

Physical quantity Measured	Unit name	Unit symbol	Definition
Base units			
Length	metre	m	Distance traveled by light in a vacuum in $\frac{1}{299\,792\,458}$ second
Mass	kilogram	kg	Mass of a platinum-iridium block located at the International Bureau of Weights and Measures at Sevres, France
Time	second	s	9 192 631 770 periods of a specified line in the microwave spectrum of the cesium-133 isotope
Temperature	kelvin	K	1/273.16 the temperature interval between absolute zero and the triple point of water at 273.16 K (0.01˚C)
Amount of substance	mole	mol	Amount of substance containing as many entities (atoms, molecules) as there are atoms in exactly 0.012 kilograms of the carbon-12 isotope
Electric current	ampere	A	—
Luminous intensity	candela	cd	—
Examples of derived units			
Force	newton	N	Force required to impart an acceleration of 1 m/s^2 to a mass of 1 kg
Energy (heat)	joule	J	Work performed by 1 newton acting over a distance of 1 meter
Pressure	pascal	Pa	Force of 1 newton acting on an area of 1 square meter

The **gram** (g) with a mass equal to 1/1000 that of the SI kilogram (see Table 1.4) is the fundamental unit of mass in the metric system. Although the gram is a convenient unit for many laboratory-scale operations, other units that are multiples of the gram are often more useful for expressing mass. The names of these are obtained by affixing the appropriate prefixes from Table 1.5 to “gram.” Global burdens of atmospheric pollutants may be given in units of teragrams, each equal to 1×10^{12} grams. Significant quantities of toxic water pollutants may be measured in micrograms (1×10^{-6} grams). Large-scale industrial chemicals are marketed in units of megagrams (Mg). This quantity is also known as a metric ton, or tonne, and is somewhat larger

(2205 lb) than the 2000-lb short ton still used in commerce in the United States. Table 1.6 summarizes some of the more commonly used metric units of mass and their relationship to some English units.

Table 1.5 Prefixes Commonly Used to Designate Multiples of Units

Prefix	Basic unit is multliplied by	Abbreviation
Mega	1 000 000 (10^{6})	M
Kilo	1 000 (10^{3})	k
Hecto	100 (10^{2})	h
Deka	10 (10)	da
Deci	0.1 (10^{-1})	d
Centi	0.01 (10^{-2})	c
Milli	0.001 (10^{-3})	m
Micro	0.000 001 (10^{-6})	µ
Nano	0.000 000 001 (10^{-9})	n
Pico	0.000 000 000 001 (10^{-12})	p

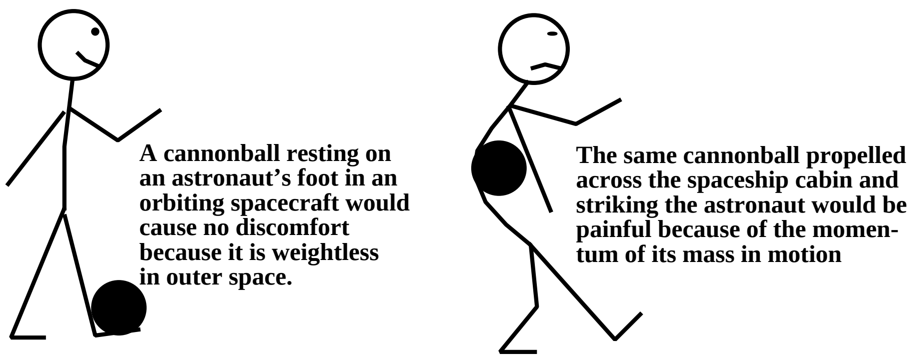


Figure 1.7 An object maintains its mass even in the weightless surroundings of outer space.

1.10 UNITS OF LENGTH

Length in the metric system is expressed in units based upon the **meter**, m (SI spelling *metre*, Table 1.4). A meter is 39.37 inches long, slightly longer than a yard. A kilometer (km) is equal to 1000 m and, like the mile, is used to measure relatively great distances. A centimeter (cm), equal to 0.01 m, is often convenient to designate lengths such as the dimensions of laboratory instruments. There are 2.540 cm per inch, and the cm is employed to express lengths that would be given in inches in the English system. The micrometer (µm) is about as long as a typical bacterial cell. The µm is also used to express wavelengths of infrared radiation by which Earth reradiates solar energy back to outer space. The nanometer (nm), equal to 10^{-9} m, is a convenient unit for the wavelength of visible light, which ranges from 400 to 800 nm.

Atoms are even smaller than 1 nm; their dimensions are commonly given in picometers (pm, 10^{-12} m). Table 1.7 lists common metric units of length, some examples of their use, and some related English units.

Table 1.6 Metric Units of Mass

Unit of mass	Abbreviation	Number of grams	Definition
Megagram or metric ton	Mg	10^6	Quantities of industrial chemicals (1 Mg = 1.102 short tons)
Kilogram	kg	10^6	Body weight and other quantities for which the pound has been commonly used (1 kg = 2.2046 lb)
Gram	g	1	Mass of laboratory chemicals (1 ounce = 28.35 g and 1 lb = 453.6 g)
Milligram	mg	10^{-3}	Small quantities of chemicals
Microgram	µg	10^{-6}	Quantities of toxic pollutants

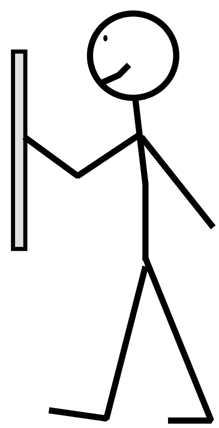


Figure 1.8 The meter stick is a common tool for measuring length.

1.11 UNITS OF VOLUME

The basic metric unit of **volume** is the **liter**, which is defined in terms of metric units of length. As shown in Figure 1.9, a liter is the volume of a decimeter cubed, that is, 1 L = 1 dm^3 (a dm is 0.1 meter, about 4 inches). A milliliter (mL) is the same volume as a centimeter cubed (cm^3 or cc), and a liter is 1000 cm^3. A kiloliter, usually designated as a cubic meter (m^3), is a common unit of measurement for the volume of air. For example, standards for human exposure to toxic substances in the workplace are frequently given in units of µg/m^3. Table 1.8 gives some common metric units of volume. The measurement of volume is one of the more frequently performed routine laboratory measurements; Figure 1.10 shows some of the more common tools for laboratory volume measurement of liquids.

Table 1.7 Metric Units of Length

Unit of length	Abbreviation	Number of meters	Definition
Kilometer	km	10^3	Distance (1 mile = 1.609 km)
Meter	m	1	Standard metric unit of length (1 m = 1.094 yards)
Centimeter	cm	10^{-2}	Used in place of inches (1 inch = 2.54 cm)
Millimeter	mm	10^{-3}	Same order of magnitude as sizes of letters on this page
Micrometer	µm	10^{-6}	Size of typical bacteria
Nanometer	nm	10^{-9}	Measurement of light wavelength

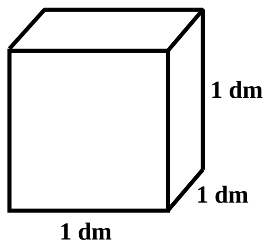


Figure 1.9 A cube that is 1 decimeter to the side has a volume of 1 liter.

Table 1.8 Metric Units of Volume

Unit of volume	Abbreviation	Number of liters	Example of use for measurement
Kiloliter or cubic meter	kL	10^3	Volumes of air in air pollution studies
Liter	L	1	Basic metric unit of volume (1 liter = 1 dm^3 = 1.057 quarts; 1 cubic foot = 28.32 L)
Milliliter	mL	10^{-3}	Equal to 1 cm^3. Convenient unit for laboratory volume measurements
Microliter	µL	10^{-6}	Used to measure very small volumes for chemical analysis

1.12 TEMPERATURE, HEAT, AND ENERGY

Temperature Scales

In chemistry, temperatures are usually expressed in metric units of **Celsius degrees**, ˚C, in which water freezes at 0˚C and boils at 100˚C. The **Fahrenheit** scale, still used for some non-scientific temperature measurements in the U.S., defines the freezing temperature of water at 32 degrees Fahrenheit (˚F) and boiling at 212˚F, a range of 180˚F. Therefore, each span of 100 Celsius degrees is equivalent to one of 180 Fahrenheit degrees and each ˚C is equivalent to 1.8˚F.

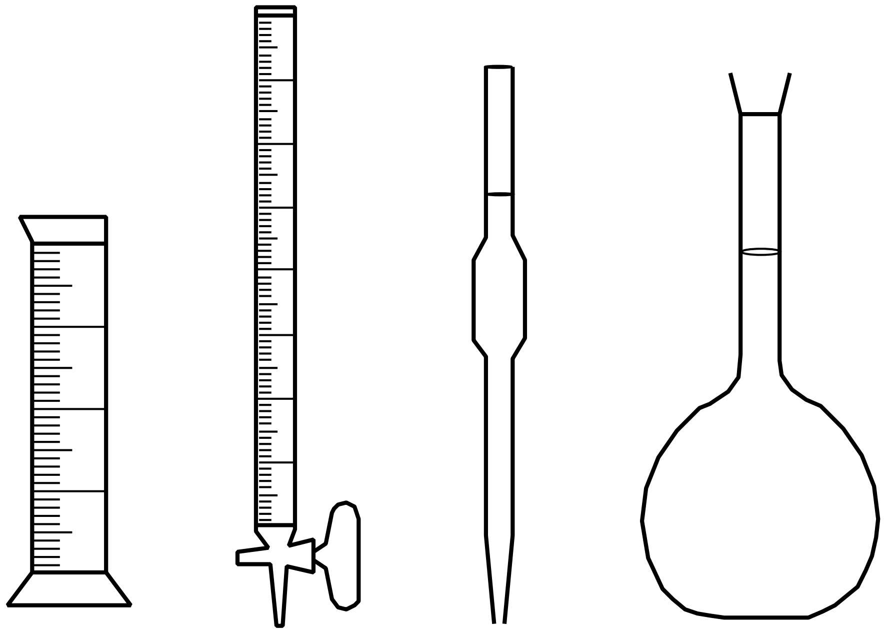


Figure 1.10 Glassware for volume measurement in the laboratory.

The most fundamental temperature scale is the **Kelvin** or **absolute** scale, for which zero is the lowest attainable temperature. A unit of temperature on this scale is equal to a Celsius degree, but it is called a **kelvin**, abbreviated K, not a degree. Kelvin temperatures are designated as K, not ˚K. The value of absolute zero on the Kelvin scale is -273.15˚C, so that the Kelvin temperature is always a number 273.15 (usually rounded to 273) higher than the Celsius temperataure. Thus water boils at 373 K and freezes at 273 K. The relationships among Kelvin, Celsius, and Fahrenheit temperatures are illustrated in Figure 1.11.

Converting from Fahrenheit to Celsius

With Figure 1.11 in mind, it is easy to convert from one temperature scale to another. Examples of how this is done are given below:

Example: What is the Celsius temperature equivalent to room temperature of 70˚F?

Answer: Step 1. Subtract 32 Fahrenheit degrees from 70 Fahrenheit degrees to get the number of Fahrenheit degrees above freezing. This is done because 0 on the Celsius scale is at the freezing point of water.

Step 2. Multiply the number of Fahrenheit degrees above the freezing point of water obtained above by the number of Celsius degrees per Fahrenheit degree.

$$^\circ C = \underbrace{\frac{1.00^\circ C}{1.80^\circ F}}_{\substack{\uparrow \\ \text{Factor for conversion} \\ \text{from } ^\circ F \text{ to } ^\circ C}} \times \underbrace{(70^\circ F - 32^\circ F)}_{\substack{\uparrow \\ \text{Number of } ^\circ F \\ \text{above freezing}}} = \frac{1.00^\circ C}{1.80^\circ F} \times 38^\circ F = 21.1^\circ C \qquad (1.12.1)$$

In working the above example it is first noted (as is obvious from Figure 1.11) that the freezing temperature of water, zero on the Celsius scale, corresponds to 32˚F on the Fahrenheit scale. So 32˚F is subtracted from 70˚F to give the number of Fahrenheit degrees by which the temperature is above the freezing point of water. The number of Fahrenheit degrees above freezing is converted to Celsius degrees above the freezing point of water by multiplying by the factor 1.00˚C/1.80˚F. The

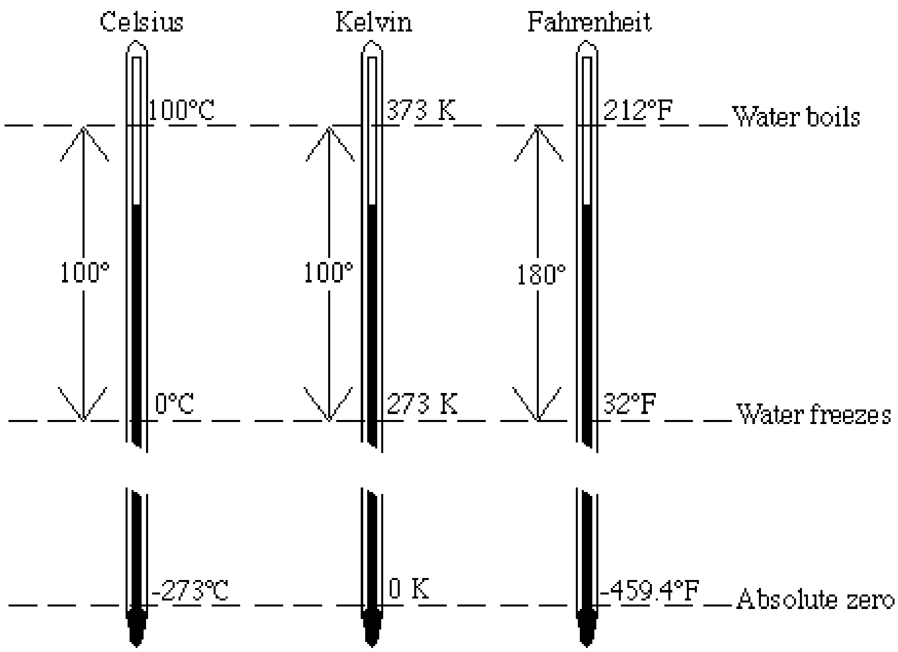


Figure 1.11 Comparison of temperature scales.

origin of this factor is readily seen by referring to Figure 1.11 and observing that there are 100˚C between the freezing and boiling temperatures of water and 180˚F over the same range. Mathematically, the equation for converting from ˚F to ˚C is simply the following:

$$^\circ C = \frac{1.00^\circ C}{1.80^\circ F} \times (^\circ F - 32) \qquad (1.12.2)$$

Example: What is the Celsius temperature corresponding to normal body temperature of 98.6˚F?

Answer: From Equation 1.12.2

$$˚C = \frac{1.00˚C}{1.80˚F} \times (98.6˚F - 32˚F) = 37.0˚C \qquad (1.12.3)$$

Example: What is the Celsius temperature corresponding to -5˚F?

Answer: From Equation 1.12.2

$$˚C = \frac{1.00˚C}{1.80˚F} \times (-5˚F - 32˚F) = ˚C = -20.6˚C \qquad (1.12.4)$$

Converting from Celsius to Fahrenheit

To convert from Celsius to Fahrenheit first requires multiplying the Celsius temperature by 1.80˚F/1.00˚C to get the number of Fahrenheit degrees above the freezing temperature of 32˚F, then adding 32˚F.

Example: What is the Fahrenheit temperature equivalent to 10˚C?

Answer: Step 1. Multiply 10˚C by 1.80˚F/1.00˚C to get the number of Fahrenheit degrees above the freezing point of water.

Step 2. Since the freezing point of water is 32˚F, add 32˚F to the result of Step 1.

$$˚F = \frac{1.80˚F}{1.00˚C} \times ˚C + 32˚F = \frac{1.80˚F}{1.00˚C} \times 10˚C + 32˚F = 50˚F \qquad (1.12.5)$$

The formula for converting ˚C to ˚F is

$$˚F = \frac{1.80˚F}{1.00˚C} \times ˚C + 32˚F \qquad (1.12.6)$$

To convert from ˚C to K, add 273 to the Celsius temperature. To convert from K to ˚C, subtract 273 from K. All of the conversions discussed here can be deduced without memorizing any equations by remembering that the freezing point of water is 0˚C, 273 K, and 32˚F, whereas the boiling point is 100˚C, 373 K, and 212˚F.

Melting Point and Boiling Point

In the preceding discussion, the melting and boiling points of water were both used in defining temperature scales. These are important thermal properties of any substance. For the present, **melting temperature** may be defined as the temperature at which a substance changes from a solid to a liquid. **Boiling temperature** is defined as the temperature at which a substance changes from a liquid to a gas. More-exacting definitions of these terms, particularly boiling temperature, are given later in the book.

Heat and Energy

As illustrated in Figure 1.12, when two objects at different temperatures are placed in contact with each other, the warmer object becomes cooler and the cooler one warmer until they reach the same temperature. This occurs because of a flow of energy between the objects. Such a flow is called **heat**.

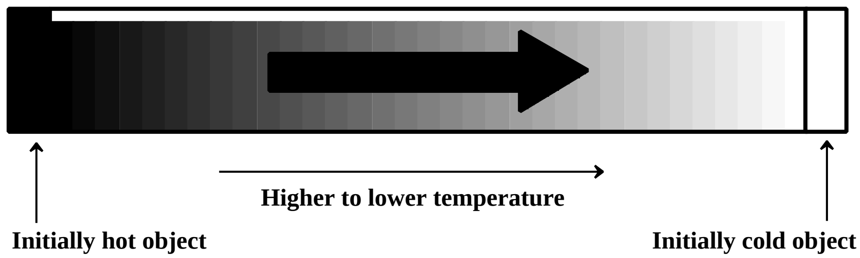


Figure 1.12 Heat energy flow from a hot to a colder object.

The SI unit of heat is the **joule** (J, see Table 1.4). The kilojoule (1 kJ = 1000 J) is a convenient unit to use to express energy values in laboratory studies. The metric unit of energy is the **calorie** (cal), equal to 4.184 J. Throughout the liquid range of water, essentially 1 calorie of heat energy is required to raise the temperature of 1 g of water by 1˚C. The "calories" most people hear about are those used to express energy values of foods and are actually kilocalories (1 kcal = 4.184 kJ).

1.13 PRESSURE

Pressure is force per unit area. The SI unit of pressure is the pascal (Pa), defined in Table 1.4. The kilopascal (1 kPa = 1000 Pa) is often a more convenient unit of pressure to use than is the pascal.

Like many other quantities, pressure has been plagued with a large number of different kinds of units. One of the more meaningful and intuitive of these is the **atmosphere** (atm), and the average pressure exerted by air at sea level is 1 atmosphere. One atmosphere is equal to 101.3 kPa or 14.7 lb/in^2. The latter means that an evacuated cube, 1 inch to the side, has a force of 14.70 lb exerted on each side due to atmospheric pressure. It is also the pressure that will hold up a column of liquid mercury metal 760 mm long, as shown in Figure 1.13. Such a device used to measure atmospheric pressure is called a **barometer**, and the mercury barometer was the first instrument used to measure pressures with a high degree of accuracy. Consequently, the practice developed of expressing pressure in units of **millimeters of mercury** (mm Hg), where 1 mm of mercury is a unit called the **torr**.

Pressure is an especially important variable with gases because the volume of a quantity of gas at a fixed temperature is inversely proportional to pressure. The temperature/pressure/volume relationships of gases (Boyle's law, Charles' law, general gas law) are discussed in Chapter 2.

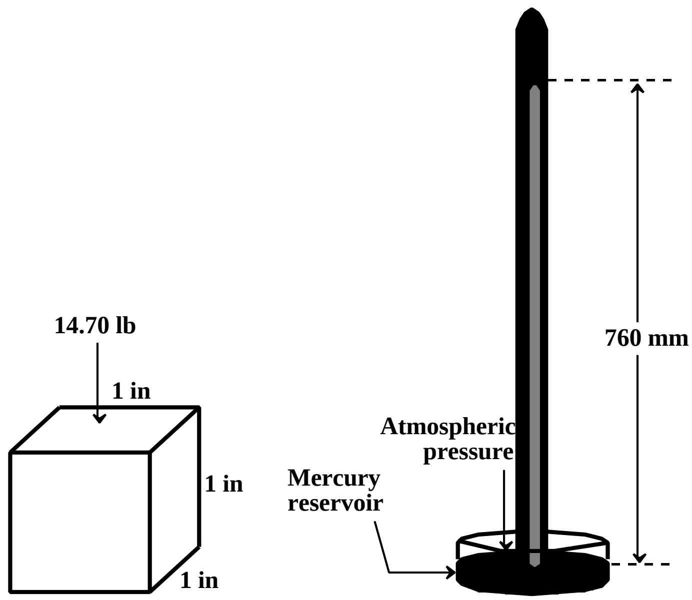


Figure 1.13 Average atmospheric pressure at sea level exerts a force of 14.7 pounds on an inch-square surface. This corresponds to a pressure sufficient to hold up a 760 mm column of mercury.

1.14 UNITS AND THEIR USE IN CALCULATIONS

Most numbers used in chemistry are accompanied by a **unit** that tells the type of quantity that the number expresses and the smallest whole portion of that quantity. For example, “36 liters” denotes that a volume is expressed and the smallest whole unit of the volume is 1 liter. The same quantity could be expressed as 360 deciliters, where the number is multiplied by 10 because the unit is only 1/10 as large.

Except in cases where the numbers express relative quantitities, such as atomic masses relative to the mass of carbon-12 or specific gravity, it is essential to include units with numbers. In addition to correctly identifying the type and magnitude of the quantity expressed, the units are carried through mathematical operations. The wrong unit in the answer shows that something has been done wrong in the calculation and it must be checked.

Unit Conversion Factors

Most chemical calculations involve calculating one type of quantity, given another, or converting from one unit of measurement to another. For example, in the chemical reaction

$$2H_2 + O_2 \rightarrow H_2O$$

someone might want to calculate the number of grams of H_2O produced when 3 g of H_2 react, or they might want to convert the number of grams of H_2 to ounces. These kinds of calculations are carried out with **unit conversion factors**. Suppose, for example, that the mass of a 160-lb person is to be expressed in kilograms; the person doing the calculation does not know the factor to convert from lb to kg, but does know that a 551-lb motorcycle has a mass of 250 kg. From this information the needed unit conversion factor can be derived and the calculation completed as follows:

$$\text{Mass of person in kg} = 160 \text{ lb} \times \text{unit conversion factor} \quad \text{(problem to be solved)} \tag{1.14.1}$$

$$250 \text{ kg} = 551 \text{ lb} \quad \text{(known relationship between lb and kg)} \tag{1.14.2}$$

$$\frac{250 \text{ kg}}{551 \text{ lb}} = \frac{551 \text{ lb}}{551 \text{ lb}} = 1 \quad \text{(The unit of kg is left on top because it is the unit needed; division is by 551 lb.)} \tag{1.14.3}$$

$$\frac{0.454 \text{ kg}}{1.00 \text{ lb}} = 1 \quad \text{(The unit conversion factor in the form } 250 \text{ kg}/551 \text{ lb could have been used, but dividing 250 by 551 gives the unit conversion factor in a more concise form.)} \tag{1.14.4}$$

$$\text{Mass of person} = 160 \text{ lb} \times \frac{0.454 \text{ kg}}{1.00 \text{ lb}} = 72.6 \text{ kg} \tag{1.14.5}$$

It is permissible to multiply 160 lb by 0.454 kg/1.00 lb because, as shown by Equation 1.14.4, this unit conversion factor has a value of exactly 1. Any quantity can be multiplied by 1 without changing the quantity itself; the only change is in the units in which it is expressed.

As another example of the use of a unit conversion factor, calculate the number of liters of gasoline required to fill a 12-gallon fuel tank, given that there are 4 gallons in a quart and that a volume of 1 liter is equal to that of 1.057 quarts. This problem can be worked by first converting gallons to quarts, then quarts to liters. For the first step, the unit conversion factor is

$$1 \text{ gal} = 4 \text{ qt} \tag{1.14.6}$$

$$\frac{1 \text{ gal}}{1 \text{ gal}} = \frac{4 \text{ qt}}{1 \text{ gal}} = 1 \quad \text{(Conversion from gallons to quarts)} \tag{1.14.7}$$

$$1.057 \text{ qt} = 1 \text{ L} \tag{1.14.8}$$

$$\frac{1.057 \text{ qt}}{1.057 \text{ qt}} = \frac{1 \text{ L}}{1.057 \text{ qt}} = 1 \quad \text{(Conversion from quarts to liters)} \tag{1.14.9}$$

Both unit conversion factors are used to calculate the capacity of the tank in liters:

$$\text{Tank capacity} = 12 \text{ gal} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{1 \text{ L}}{1.057 \text{ qt}} = 45.4 \text{ L} \tag{1.14.10}$$

Cancellation of Units

The preceding examples show that units are canceled in mathematical operations, just as numbers may be. When the same unit appears both above and below the line in a mathematical operation, the units cancel. An example of such an operation is shown for lb in the following:

$$160 \text{ lb} \times \frac{0.454 \text{ kg}}{1.00 \text{ lb}}$$

The unit of lb simply cancels, leaving kg as the unit remaining.

Calculation of Some Unit Conversion Factors

Several values of units are given that enable conversion between metric and English units in Table 1.6 (mass), Table 1.7 (length), and Table 1.8 (volume). For example, Table 1.6 states that a megagram (Mg, metric ton) is equal to 1.102 short tons (T). By using this equality to give the correct unit conversion factors, it is easy to calculate the number of metric tons in a given number of short tons of material or vice versa. To do this, first write the known equality given that a megagram is equal to 1.102 short tons:

$$1 \text{ Mg} = 1.102 \text{ T} \tag{1.14.11}$$

If the number of Mg is to be calculated given a mass in T, the unit conversion factor needed is

$$\frac{1 \text{ Mg}}{1.102 \text{ T}} = \frac{1.102 \text{ T}}{1.102 \text{ T}} = 1 \tag{1.14.12}$$

leaving Mg on top. Suppose, for example, that the problem is to calculate the mass in Mg of a 3521 T shipment of industrial soda ash. The calculation involves simply multiplying the known mass in T times the unit conversion factor required to convert to Mg:

$$3521 \text{ T} \times \frac{1 \text{ Mg}}{1.102 \text{ T}} = 3195 \text{ Mg} \tag{1.14.12}$$

If the problem had been to calculate the number of T in 789 Mg of copper ore, the following steps would be followed:

$$1.102 \text{ T} = 1 \text{ Mg}, \qquad \frac{1.102 \text{ T}}{1 \text{ Mg}} = \frac{1 \text{ Mg}}{1 \text{ Mg}} = 1, \tag{1.14.13}$$

$$789 \text{ Mg} \times \frac{1.102 \text{ T}}{1 \text{ Mg}} = 869 \text{ T copper ore} \tag{1.14.14}$$

Table 1.9 gives some unit conversion factors calculated from the information given in Tables 1.6–1.8 and in preceding parts of this chapter. Note that in each case, two unit conversion factors are calculated; the one that is used depends upon the units that are required for the answer.

Table 1.9 Examples of Some Unit Conversion Factors

Equality	Conversion factors	
1 kg = 2.2046 lb	$\frac{1\ \text{kg}}{2.2046\ \text{lb}} = 1$	$\frac{2.2046\ \text{lb}}{1\ \text{kg}} = 1$
1 oz = 28.35 g	$\frac{1\ \text{oz}}{28.35\ \text{g}} = 1$	$\frac{28.35\ \text{g}}{1\ \text{oz}} = 1$
1 mi = 1.609 km	$\frac{1\ \text{mi}}{1.609\ \text{km}} = 1$	$\frac{1.609\ \text{km}}{1\ \text{mi}} = 1$
1 in = 2.54 cm	$\frac{1\ \text{in}}{2.54\ \text{cm}} = 1$	$\frac{2.54\ \text{cm}}{1\ \text{in}} = 1$
1 L = 1.057 qt	$\frac{1\ \text{L}}{1.057\ \text{qt}} = 1$	$\frac{1.057\ \text{qt}}{1\ \text{L}} = 1$
1 cal = 4.184 J	$\frac{1\ \text{cal}}{4.184\ \text{J}} = 1$	$\frac{4.184\ \text{J}}{1\ \text{cal}} = 1$
1 atm = 101.4 kPa	$\frac{1\ \text{atm}}{101.4\ \text{kPa}} = 1$	$\frac{101.4\ \text{kPa}}{1\ \text{atm}} = 1$

CHAPTER SUMMARY

The chapter summary below is presented in a programmed format to review the main points covered in this chapter. It is used most effectively by filling in the blanks, referring back to the chapter as necessary. The correct answers are given at the end of the summary.

Chemistry is defined as [1] ________________________. Environmental chemistry is [2] ____ ____. Toxicological chemistry is defined as [3] _____________________________________ ____. All matter is composed of only about a hundred fundamental kinds of matter called [4], each composed of very small entities called [5] __________ ________. The three major subatomic particles and their charges are [6] ___________ ______________________________________. Of these, the two that have relatively high masses are contained in the [7] __________________ of the atom. The subatomic particles with a relatively low mass are contained in [8] _________________________ ____ in the atom. The number of protons in the nucleus of each atom of an element is the [9] ________________________ of the element. Each element is represented by an abbreviation called a [10] _____________. In addition to atomic number, name, and chemical symbol, each element has a characteristic [11] ____________________________________. Atoms of most elements consist of two or more isotopes that have different [12] _________________________ ________________________________. An arrangement of the elements in a manner

that reflects their recurring behavior with increasing atomic number is the [13] _____
___________________ in which elements with similar chemical properties are called
[14] ________________________________ and are contained in [15] _________
___________________ in the periodic table. Instead of existing as atoms, elemental
hydrogen consists of [16] ____________________, each consisting of [17] __________
________________ linked by a [18]_____________________. Water is not an element,
but is a [19]_________________________, for which the [20] ____________
____________________ is H_2O. Species consisting of electrically charged atoms or
groups of atoms are called [21]_______________. Those with positive charges are called
[22]__________________ and those with negative charges are [23] ___________.
Compounds made of these kinds of entities are held together by [24] ______________
bonds. The average mass of all molecules of a compound is its [25]_________________,
which is calculated by [26]____
____.
[27]______________________________ occur when substances are changed to other
substances through the breaking and formation of chemical bonds and are written as
[28] _________________________. To be correct, these must be [29]____________.
In them, the arrow is read as [30]_________ and separates the [31]_________________
from the [32]_____________________. Very large or small numbers are conveniently
expressed in [33]______________________, which is the product of a [34]__________
_______________ with a value equal to or greater than [35]__________________ and
less than [36] __________ multiplied times a [37]_________________________. In such a
notation, 3,790,000 is expressed as [38]_____________ and 0.000 000 057 is expressed
as [39]____________________. The accuracy of a number is shown by how many [40]
it contains. Non-zero digits in a number are always [41]_____________. Zeros between
non-zero digits are [42]_________
_____. Zeros on the left of the first non-zero digit are [43]_______________________.
Zeros to the right of a decimal point that are preceded by a significant figure are [44]
______________________. The number of significant digits in a number written in
exponential notation is equal to [45]____
________________. Some numbers, such as the amount of money that one expects
to receive when cashing a check are defined as [46]__________________________. In
addition and subtraction, the number of digits retained to the right of the decimal
point should be [47]____
____.
The number of significant figures in the result of multiplication/division should be [48]
. In rounding numbers, if the digit to be dropped is 0, 1, 2, 3, or 4, [49]____________
_______________________________, whereas if the digit to be dropped is 5,6,7,8 or
9,[50]____. A self-
consistent set of units based upon the metric system is the [51]____________________
___________________. [52]__________ is proportional to the amount of matter in an
object, the metric unit for which is the [53]___________. Length in the metric system is
expressed in units based upon the [54]___________. The basic metric unit of volume is
the [55]_____________________. In ˚C, ˚F, and K, respectively, water freezes at [56]
and boils at [57]_____________________. Boiling temperature is defined as [58]
____.
Energy that flows from a warmer to a colder object is called [59]_________________,

commonly expressed in units of [60]__________ or [61]____________. [62]_____________ is force per unit area, some of the common units for which are [63]_______________ ____. A unit after a number tells the [64]_____________________________________ that the number expresses and the smallest [65]____________________________________. The quantity $0.454\ \text{kg}/1\ \text{lb}$ is an example of [66]____.

Answers to Chapter Summary

1. the science of matter
2. that branch of chemistry that deals with the origins, transport, reactions, effects, and fates of chemical species in the water, air, earth, and living environments and the effects of human activities thereon
3. the chemistry of toxic substances with emphasis upon their interaction with biologic tissue and living systems
4. elements
5. atoms
6. positively charged protons, negatively charged electrons, and uncharged (neutral) neutrons
7. nucleus
8. a cloud of negative charge
9. atomic number
10. chemical symbol
11. atomic mass
12. numbers of neutrons in their nuclei
13. periodic table
14. groups of elements
15. vertical columns
16. molecules
17. 2 H atoms
18. chemical bond
19. chemical compound
20. chemical formula
21. ions
22. cations
23. anions
24. ionic
25. molecular mass
26. multiplying the atomic mass of each element by the relative number of atoms of the element, then adding all the values obtained for each element in the compound
27. Chemical reactions
28. chemical equations
29. balanced
30. yields
31. reactants
32. products

33. exponential notation
34. digital number
35. exactly 1
36. exactly 10
37. power of 10
38. 3.79×10^6
39. 5.7×10^{-8}
40. significant figures or significant digits
41. significant
42. significant
43. not significant
44. significant
45. the number of significant digits in the decimal portion
46. exact numbers
47. the same as that in the number in the calculation with the fewest such digits
48. the same as that in the number in the calculation having the fewest significant figures
49. leave the last digit retained unchanged
50. increase the last retained digit by 1
51. International System of Units, SI
52. Mass
53. gram
54. meter
55. liter
56. 0˚C, 32˚F, and 273 K
57. 100˚C, 212˚F, and 373 K
58. the temperature at which a substance changes from a liquid to a gas
59. heat
60. joules
61. calories
62. Pressure
63. atmosphere, torr, mm Hg, lb/in^2
64. type of quantity
65. whole portion of that quantity
66. a unit conversion factor

LITERATURE CITED

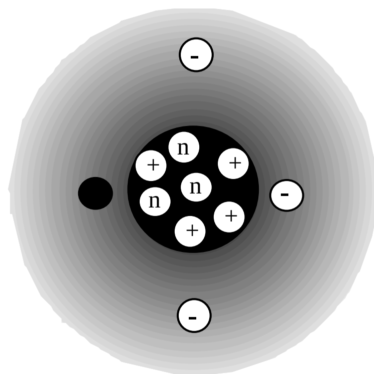
1. Manahan, Stanley E., *Environmental Chemistry*, 7th ed., CRC Press, Boca Raton, FL 2000.

2. Manahan, Stanley E., *Toxicological Chemistry,* 2nd ed., CRC Press, Boca Raton, FL 1993.

QUESTIONS AND PROBLEMS

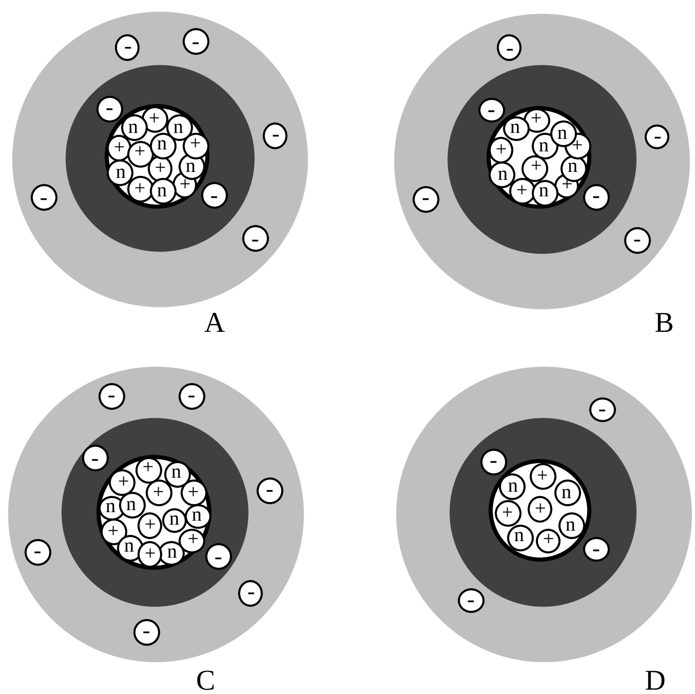
Section 1.3.   The Building Blocks of Matter

1. Consider the following atom:



   How many electrons, protons, and neutrons does it have? What is its atomic number? Give the name and chemical symbol of the element of which it is composed.

2. Give the numbers of electrons, protons, and neutrons in each of the atoms represented below. Give the atomic numbers, names, and chemical symbols of each of the elements represented.



3. What distinguishes a radioactive isotope from a “normal” stable isotope?

4. Why is the periodic table so named?

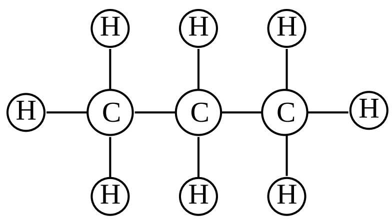
Section 1.4. Chemical Bonds and Compounds

5. Match the following:

A. O_2	1. Element consisting of individual atoms
B. NH_3	2. Element consisting of chemically bonded atoms
C. Ar	3. Ionic compound
D. NaCl	4. Covalently bound compound

6. After examining Figure 1.6, consider what might happen when an atom of sodium (Na), atomic number 11, loses an electron to an atom of fluorine, (F), atomic number 9. What kinds of particles are formed by this transfer of a negatively charged electron? Is a chemical compound formed? What is it called?

7. Give the chemical formula and molecular mass of the molecule represented below:



8. Calculate the molecular masses of (a) C_2H_2, (b) N_2H_4, (c) Na_2O, (d) O_3 (ozone), (e) PH_3, (f) CO_2, (g) C_3H_9O.

9. Match the following:

A. Periodic table	1. Has groups
B. Elemental hydrogen	2. Consists of ions
C. Water H_2O	3. Consists of covalently bonded (shared electrons) molecule containing two different elements
D. Magnesium oxide, MgO	4. Consists of diatomic, covalently bonded molecules

Section 1.5. Chemical Reactions and Equations

10. Is the equation, $H_2 + O_2 \rightarrow H_2O$, a balanced chemical equation? Explain. Point out the reactants and products in the equation.

Section 1.6. Numbers in Chemistry: Exponential Notation

11. Write each of the following in exponential form to three significant figures: (a) 321,000, (b) 0.000 005 29, (c) 5170, (d) 000 000 000 000 784, (e) 86,300,000,000,000.

12. Write each of the following in decimal form: (a) 7.49×10^3, (b) 9.6×10^{-5}, (c) 1.16×10^{21}, (d) 4.47×10^{-17}, (e) 2.93×10^{13}.

13. Without using a calculator, calculate the sum $4.13 \times 10^3 + 8.76 \times 10^2 + 1.22 \times 10^4$ expressed to three significant figures in the correct exponential notation.

14. Without using a calculator, calculate the sum $4.13 \times 10^{-4} + 8.76 \times 10^{-3} + 1.22 \times 10^{-2}$ expressed to three significant figures in the correct exponential notation.

15. Without using a calculator for the exponential portions, calculate $1.39 \times 10^{-2} \times 9.05 \times 10^8 \times 3.11 \text{ x } 10^4$.

16. Without using a calculator for the exponential portions, calculate

$$\frac{9.05 \times 10^{-6} \times 3.19 \times 10^3}{4.02 \times 10^5 \times 1.93 \times 10^{-7}}$$

Section 1.7. Significant Figures and Uncertainties in Numbers

17. Match the following numbers, with the significant figures for each given in parentheses with the rule for assigning significant figures that applies to each:

A. 0.00027(2)	1. Non-zero digits in a number are always significant.
B. 7.28139(6)	2. Zeros between non-zero digits are significant.
C. 7.4 x 10^3(2)	3. Zeros on the left of the first non-zero digit are not significant.
D. $50(infinite)	4. Zeros to the right of a decimal point that are preceded by a significant figure are significant.
E. 81.000(5)	5. The number of significant digits in a number written in exponential notation is equal to the number of significant digits in the decimal portion.
F. 40.007(5)	6. Some numbers are defined as exact numbers without any uncertainty.

18. Using the appropriate rules, round each of the following to the correct number of significant digits: (a) 923.527 + 3.02891 + 729.29, (b) 273.591 + 12.72489 + 0.1324, (c) 473+ 9.3827 + 349.17, (d) 693.59102 + 9.00327 + 229.461853.

19. Using the appropriate rules, round each of the following to the correct number of significant digits: (a) $3.52 \times 8.02891 \times 729$, (b) $4.52 \times 10^3 \times 8.021 \times 0.5779$,

(c) $\dfrac{7.7218 \times 10^7 \times 4.019 \times 10^{-3}}{1.48 \times 10^{-5}}$, (d) $\dfrac{7.8 \times 6.028 \times 10^{-3}}{4.183 \times 10^{-5} \times 2.19 \times 10^5}$

20. Round each of the following properly: (a) 7.32987 to 3 places, (b) 1.193528 to 4 places, (c) 7.1382×10^3 to 2 places (d) 9.04557×10^{-17} to 4 places, (e) 71235.801 to 3 places, (f) 5.8092355 to 3 places.

21. How many significant digits does each of the following numbers have?

A. 7.8231 B. 7.63×10^5

C. 0.004 D. 20.071

Section 1.8. Measurements and Systems of Measurement

22. Match the following numbers, for each of which the significant figures are given in parentheses, with the rule for assigning significant figures that applies to each:

A. Mole	1. 1/273.16 the temperature interval between absolute zero and the triple point of water at 273.16 K (0.01˚C)
B. Metre	
C. Gram	2. Metric unit of volume
D. Kelvin	3. SI unit for amount of substance
E. Liter	4. Metric unit of mass
	5. Distance traveled by light in a vacuum in $\frac{1}{299\,792\,458}$ second

23. Match the abbreviations below with the numbers they represent:

A. Centi	D. Micro	1. $\times\ 10^{-3}$	4. × 10
B. Kilo	E. Milli	2. $\times\ 10^{-6}$	5. × 0.01
C. Deka		3. × 1000	

Section 1.9. Units of Mass.

24. Denote each of the following as characteristic of mass (*m*) or characteristic of weight (*w*):

A. Varies with gravity

B. Degree to which an object resists a change in its state of rest or motion

C. Direct measure of the amount of matter in the object

D. Different on the Moon's surface than on Earth

25. Match the following units on the left, below, with the quantity they are most likely to be used to express:

A. Mg	1. Quantities of toxic pollutants
B. µg	2. Quantities of large-scale industrial chemicals
C. kg	3. Quantities of laboratory chemicals
D. g	4. Global burdens of atmospheric pollutants
E. Teragrams	5. Body mass

26. Calculate (a) the number of grams in 1.56 pounds, (b) the number of kilograms in a 2000-pound ton, (c) the number of g in 2.14 kg, (d) the number of atmospheric dust particles, each weighing an average of 2.56 µg, to make up an ounce of dust particles.

Section 1.10. Units of Length

27. Distinguish between a meter and a metre.

28. How tall is a 6-foot person in cm?

29. Estimate approximately how many bacterial cells would have to be laid end-to-end to reach an inch.

30. Match the following units on the left, below, with the quantity they are most likely to be used to measure:

A. km	1. Distance between this line and the line directly below
B. m	2. Distance run by an athlete in 5 seconds
C. Nanometer	3. Distance traveled by an automobile in 1 hour
D. Centimeter	4. Dimensions of this book
E. mm	5. Wavelength of visible light

Section 1.11. Units of Volume

31. Explain how metric units of volume can be defined in terms of length.

32. Recalling the appropriate formula from elementary geometry, what is the volume in liters of a round tank with a radius of 39.0 cm and a depth of 15.0 cm?

33. Consider gasoline at a price of $1.45 per gallon. What is its equivalent price in dollars per liter ($/L)?

34. Match the following units on the left, below, with the quantity they are most likely to be used to measure:

A. Milliliter	1. Volume of milk
B. Kiloliter	2. Volume of a laboratory chemical
C. Microliter	3. Volume of air in air pollution studies
D. Liter	4. Volume of chemical reagent in a syringe for chemical analysis

35. Give the volume in liters of cubes that are (A) 12 cm, (B) 14 cm, and (C) 17.2 cm to the side.

Section 1.12. Temperature, Heat, and Energy

36. Convert each of the following Fahrenheit temperatures to Celsius and Kelvin: (a) 237˚F, (b) 105˚F, (c) 17˚F, (d) 2˚F, (e) -32˚F, (f) -5˚F, (g) 31.2˚F.

37. Convert each of the following Celsius temperatures to Fahrenheit: (a) 237˚C, (b) 75˚C, (c) 17˚C, (d) 100˚C, (e) -32˚C, (f) -40˚C, (g) -11˚C.

38. Convert each of the following Celsius temperatures to Kelvin: (a) 237˚C, (b) 75˚C, (c) 48˚C, (d) 100˚C, (e) -0˚C, (f) -40˚C, (g) -200˚C.

39. Calculate the value of temperature in ˚F that is numerically equal to the temperature in ˚C.

40. The number 1.8 can be used in making conversions between Fahrenheit and Celsius temperatures. Explain how it is used and where it comes from in this application.

41. Calculate how many calories there are in 1 joule.

1.13. Pressure

42. Match each of the following pertaining to units of pressure:

A. pascal	1. Based on a column of liquid
B. atm	2. Takes 14.7 to equal 1 atm
C. mm Hg	3. SI unit
D. lb/in^2	4. Essentially 1 for air at sea level.

43. Try to explain why pressure is a more important variable for gases than for liquids.

44. The pressure in a typical automobile tire is supposed to be 35 lb/in^2 (above normal atmospheric pressure). Calculate the equivalent pressure in (a) pascal, (b) atm, and (c) torr.

45. Knowing that there are 12 inches per foot, calculate the normal pressure of the atmosphere in lb/ft^2.

46. Atmospheric pressure readings on weather reports in the U.S. used to be given as 29–30 "inches." Speculate on what such a reading might mean.

1.14. Units and their Use in Calculations

47. Using unit conversion factors, calculate the following:

    (a) A pressure in inches of mercury eqivalent to 1 atm pressure

    (b) The mass in metric tons of 760,000 tons of contaminated soil

    (c) The number of cubic meters in a cubic mile of atmospheric air

    (d) The cost of 100 liters of gasoline priced at $1.33/gallon.

    (e) The number of kilograms of cheese in 200 ounces of this food

    (f) The pressure in kPa equivalent to 5.00 atm pressure

48. An analyst reported some titration data as "34.52 mL." What two things are stated by this expression?

49. Why is it not wrong to give the atomic mass of aluminum as 26.98, even though a unit is not specified for the mass?

50. Explain what is meant by a unit conversion factor. How is such a factor used? Why may quantities be multiplied by a proper unit conversion factor without concern about changing the magnitude of the quantity?

51. Consider only the following information: *An object with a mass of 1 kg also has a mass of 2.2046 lb. A pound contains 16 ounces. A piece of chalk 2 inches*

long is also 5.08 cm long. There are 36 inches in 1 yard. A 1-liter volumetric flask contains 1.057 quarts. A cubic centimeter is the same volume as a milliliter. From this information show how to calculate unit conversion factors to convert the following: (a) From yards to meters, (b) from ounces to grams, (c) from quarts to deciliters, (d) from cubic inches to cubic centimeters.

52. A quantity of colored water was poured into a 50-mL graduated cylinder, raising it to the level shown on the left in the figure. Next, an object with a mass of 25.2 g was placed in the water, raising the level to that shown on the right. What was the density of the object in g/mL?