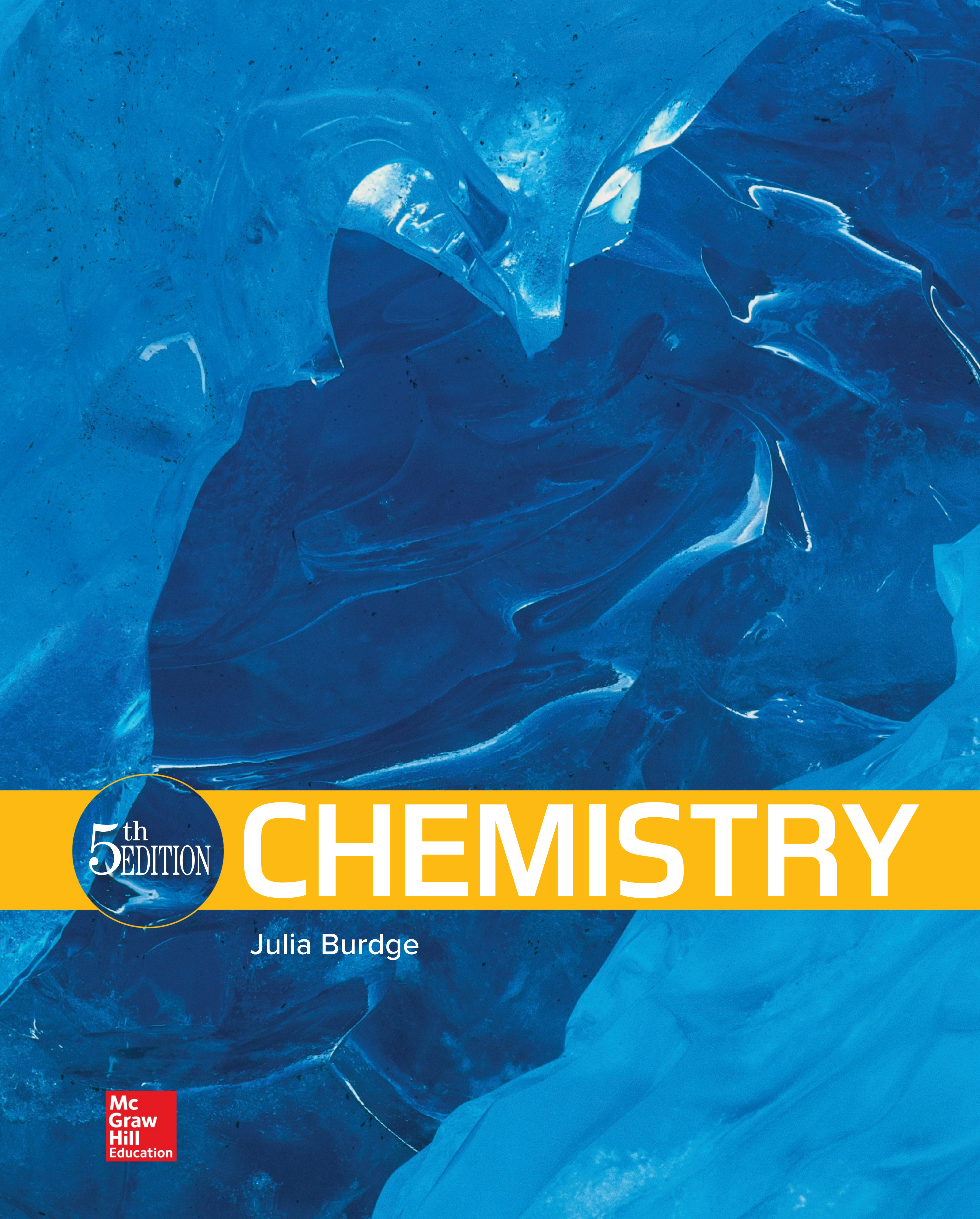

5th EDITION

CHEMISTRY

Julia Burdge

Mc Graw Hill Education

Fundamental Constants

Avogadro's number (N_A)	6.0221418×10^{23}
Electron charge (e)	1.6022×10^{-19} C
Electron mass	9.109387×10^{-28} g
Faraday constant (F)	96,485.3 C/mol e^-
Gas constant (R)	0.08206 L · atm/K · mol
	8.314 J/K · mol
	62.36 L · torr/K · mol
	1.987 cal/K · mol
Planck's constant (h)	6.6256×10^{-34} J · s
Proton mass	1.672623×10^{-24} g
Neutron mass	1.674928×10^{-24} g
Speed of light in a vacuum	2.99792458×10^{8} m/s

Some Prefixes Used with SI Units

tera (T)	10^{12}	centi (c)	10^{-2}
giga (G)	10^{9}	milli (m)	10^{-3}
mega (M)	10^{6}	micro (μ)	10^{-6}
kilo (k)	10^{3}	nano (n)	10^{-9}
deci (d)	10^{-1}	pico (p)	10^{-12}

Useful Conversion Factors and Relationships

1 lb = 453.6 g

1 in = 2.54 cm (exactly)

1 mi = 1.609 km

1 km = 0.6215 mi

1 pm = 1×10^{-12} m = 1×10^{-10} cm

1 atm = 760 mmHg = 760 torr = 101,325 N/m^2 = 101,325 Pa

1 cal = 4.184 J (exactly)

1 L · atm = 101.325 J

1 J = 1 C × 1 V

$$?^\circ C = (^\circ F - 32^\circ F) \times \frac{5^\circ C}{9^\circ F}$$

$$?^\circ F = \frac{9^\circ F}{5^\circ C} \times (^\circ C) + 32^\circ F$$

$$?K = (^\circ C + 273.15^\circ C)\left(\frac{1K}{1^\circ C}\right)$$

Periodic Table of the Elements

Key

6 C Carbon 12.01

Atomic number — 6; Symbol — C; Name — Carbon; Average atomic mass — 12.01

An element

Main group | Transition metals | Main group

Period number	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	1 H Hydrogen 1.008																	2 He Helium 4.003	1
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18	2
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95	3
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80	4
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3	5
6	55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)	6
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (271)	107 Bh Bohrium (272)	108 Hs Hassium (270)	109 Mt Meitnerium (276)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (280)	112 Cn Copernicium (285)	113 Nh Nihonium (286)	114 Fl Flerovium (289)	115 Mc Moscovium (289)	116 Lv Livermorium (293)	117 Ts Tennessine (293)	118 Og Oganesson (294)	7

Lanthanides	6	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0	6
Actinides	7	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)	7

Metals

Nonmetals

Metalloids

List of the Elements with Their Symbols and Atomic Masses*

Element	Symbol	Atomic Number	Atomic Mass†	Element	Symbol	Atomic Number	Atomic Mass†
Actinium	Ac	89	(227)	Mendelevium	Md	101	(258)
Aluminum	Al	13	26.9815386	Mercury	Hg	80	200.59
Americium	Am	95	(243)	Molybdenum	Mo	42	95.94
Antimony	Sb	51	121.760	Moscovium	Mc	115	(289)
Argon	Ar	18	39.948	Neodymium	Nd	60	144.242
Arsenic	As	33	74.92160	Neon	Ne	10	20.1797
Astatine	At	85	(210)	Neptunium	Np	93	(237)
Barium	Ba	56	137.327	Nickel	Ni	28	58.6934
Berkelium	Bk	97	(247)	Nihonium	Nh	113	(286)
Beryllium	Be	4	9.012182	Niobium	Nb	41	92.90638
Bismuth	Bi	83	208.98040	Nitrogen	N	7	14.0067
Bohrium	Bh	107	(272)	Nobelium	No	102	(259)
Boron	B	5	10.811	Oganesson	Og	118	(294)
Bromine	Br	35	79.904	Osmium	Os	76	190.23
Cadmium	Cd	48	112.411	Oxygen	O	8	15.9994
Calcium	Ca	20	40.078	Palladium	Pd	46	106.42
Californium	Cf	98	(251)	Phosphorus	P	15	30.973762
Carbon	C	6	12.0107	Platinum	Pt	78	195.084
Cerium	Ce	58	140.116	Plutonium	Pu	94	(244)
Cesium	Cs	55	132.9054519	Polonium	Po	84	(209)
Chlorine	Cl	17	35.453	Potassium	K	19	39.0983
Chromium	Cr	24	51.9961	Praseodymium	Pr	59	140.90765
Cobalt	Co	27	58.933195	Promethium	Pm	61	(145)
Copernicium	Cn	112	(285)	Protactinium	Pa	91	231.03588
Copper	Cu	29	63.546	Radium	Ra	88	(226)
Curium	Cm	96	(247)	Radon	Rn	86	(222)
Darmstadtium	Ds	110	(281)	Rhenium	Re	75	186.207
Dubnium	Db	105	(268)	Rhodium	Rh	45	102.90550
Dysprosium	Dy	66	162.500	Roentgenium	Rg	111	(280)
Einsteinium	Es	99	(252)	Rubidium	Rb	37	85.4678
Erbium	Er	68	167.259	Ruthenium	Ru	44	101.07
Europium	Eu	63	151.964	Rutherfordium	Rf	104	(267)
Fermium	Fm	100	(257)	Samarium	Sm	62	150.36
Flerovium	Fl	114	(289)	Scandium	Sc	21	44.955912
Fluorine	F	9	18.9984032	Seaborgium	Sg	106	(271)
Francium	Fr	87	(223)	Selenium	Se	34	78.96
Gadolinium	Gd	64	157.25	Silicon	Si	14	28.0855
Gallium	Ga	31	69.723	Silver	Ag	47	107.8682
Germanium	Ge	32	72.64	Sodium	Na	11	22.98976928
Gold	Au	79	196.966569	Strontium	Sr	38	87.62
Hafnium	Hf	72	178.49	Sulfur	S	16	32.065
Hassium	Hs	108	(270)	Tantalum	Ta	73	180.94788
Helium	He	2	4.002602	Technetium	Tc	43	(98)
Holmium	Ho	67	164.93032	Tellurium	Te	52	127.60
Hydrogen	H	1	1.00794	Tennessine	Ts	117	(293)
Indium	In	49	114.818	Terbium	Tb	65	158.92535
Iodine	I	53	126.90447	Thallium	Tl	81	204.3833
Iridium	Ir	77	192.217	Thorium	Th	90	232.03806
Iron	Fe	26	55.845	Thulium	Tm	69	168.93421
Krypton	Kr	36	83.798	Tin	Sn	50	118.710
Lanthanum	La	57	138.90547	Titanium	Ti	22	47.867
Lawrencium	Lr	103	(262)	Tungsten	W	74	183.84
Lead	Pb	82	207.2	Uranium	U	92	238.02891
Lithium	Li	3	6.941	Vanadium	V	23	50.9415
Livermorium	Lv	116	(293)	Xenon	Xe	54	131.293
Lutetium	Lu	71	174.967	Ytterbium	Yb	70	173.04
Magnesium	Mg	12	24.3050	Yttrium	Y	39	88.90585
Manganese	Mn	25	54.938045	Zinc	Zn	30	65.409
Meitnerium	Mt	109	(276)	Zirconium	Zr	40	91.224

*These atomic masses show as many significant figures as are known for each element. The atomic masses in the periodic table are shown to four significant figures, which is sufficient for solving the problems in this book.

†Approximate values of atomic masses for radioactive elements are given in parentheses.

Chemistry

Julia Burdge
COLLEGE OF WESTERN IDAHO

McGraw Hill Education

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Dedication

In loving memory of an extraordinary coauthor, mentor, and friend: Raymond Chang.

About the Author

Courtesy of Julia Burdge

Julia Burdge received her Ph.D. (1994) from the University of Idaho in Moscow, Idaho. Her research and dissertation focused on instrument development for analysis of trace sulfur compounds in air and the statistical evaluation of data near the detection limit.

In 1994, she accepted a position at The University of Akron in Akron, Ohio, as an assistant professor and director of the Introductory Chemistry program. In the year 2000, she was tenured and promoted to associate professor at The University of Akron on the merits of her teaching, service, and research in chemistry education. In addition to directing the general chemistry program and supervising the teaching activities of graduate students, she helped establish a future-faculty development program and served as a mentor for graduate students and post-doctoral associates. In 2008, Julia relocated back to the northwest to be near family. She lives in Boise, Idaho, and holds an adjunct faculty position at the College of Western Idaho in Nampa.

In her free time, Julia enjoys the company of her children and Erik Nelson, her husband and best friend.

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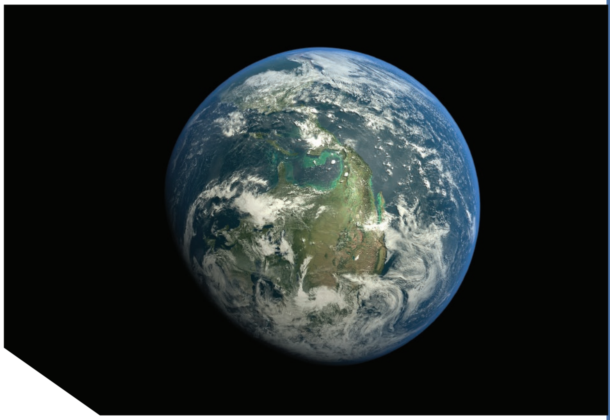

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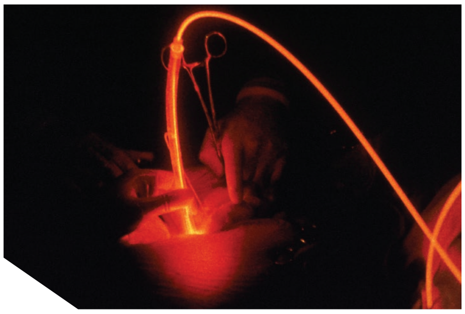

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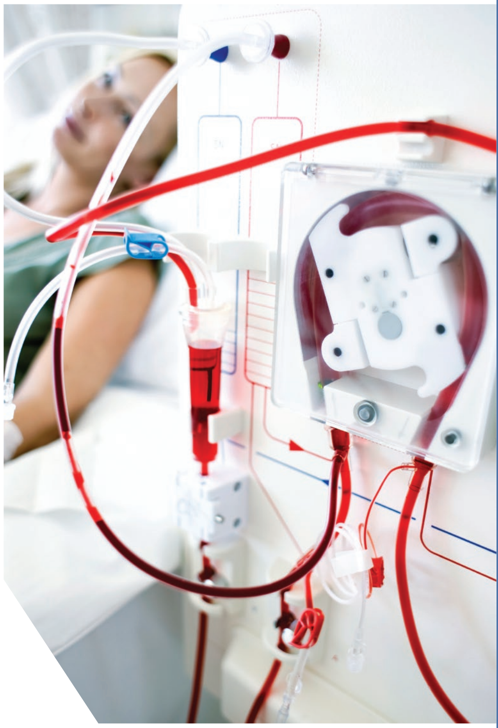

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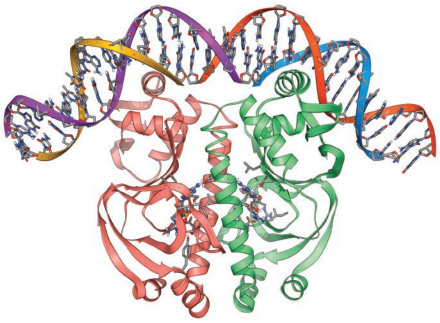

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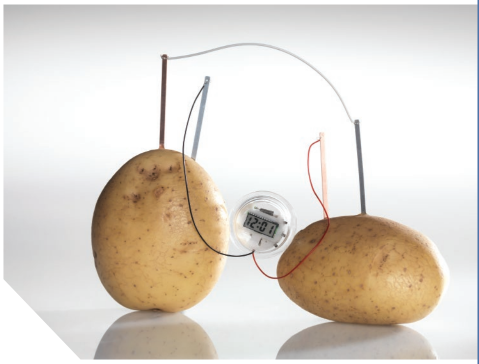

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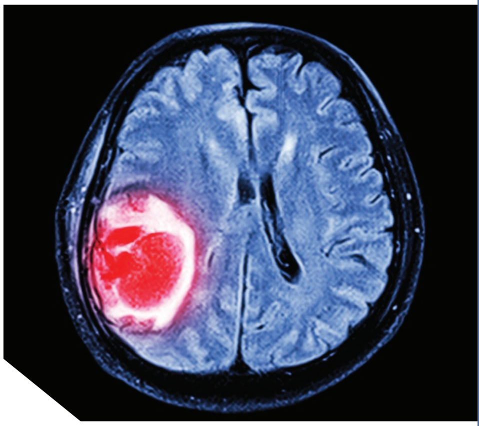

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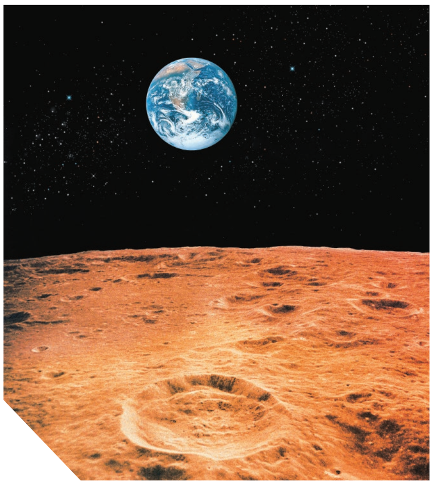

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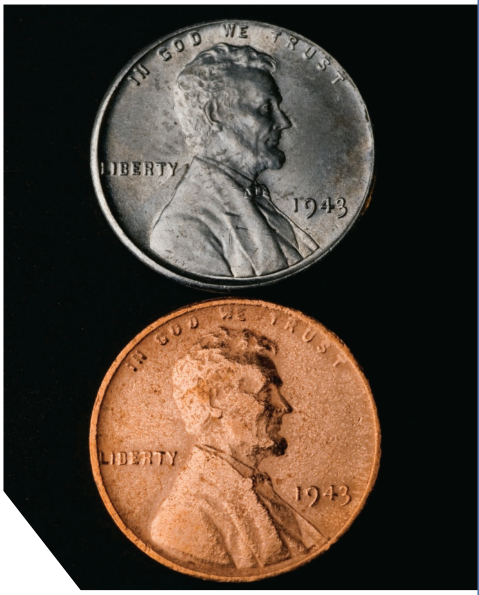

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Preface

Welcome to the exciting and dynamic world of Chemistry! My desire to create a general chemistry textbook grew out of my concern for the interests of students and faculty alike. Having taught general chemistry for many years, and having helped new teachers and future faculty develop the skills necessary to teach general chemistry, I believe I have developed a distinct perspective on the common problems and misunderstandings that students encounter while learning the fundamental concepts of chemistry—and that professors encounter while teaching them. I believe that it is possible for a textbook to address many of these issues while conveying the wonder and possibilities that chemistry offers. With this in mind, I have tried to write a text that balances the necessary fundamental concepts with engaging real-life examples and applications, while utilizing a consistent, step-by-step problem-solving approach and an innovative art and media program.

Key Features

Problem-Solving Methodology

Sample Problems are worked examples that guide the student step-by-step through the process of solving problems. Each Sample Problem follows the same four-step method: Strategy, Setup, Solution, and Think About It (check).

SAMPLE PROBLEM 4.8

For an aqueous solution of glucose ($C_6H_{12}O_6$), determine (a) the molarity of 2.00 L of a solution that contains 50.0 g of glucose, (b) the volume of this solution that would contain 0.250 mol of glucose, and (c) the number of moles of glucose in 0.500 L of this solution.

Strategy Convert the mass of glucose given to moles, and use the equations for interconversions of *M*, liters, and moles to calculate the answers.

Setup The molar mass of glucose is 180.2 g.

$$\text{moles of glucose} = \frac{50.0\text{ g}}{180.2\text{ g/mol}} = 0.277\text{ mol}$$

Solution (a) $\text{molarity} = \dfrac{0.227\text{ mol } C_6H_{12}O_6}{2.00\text{ L solution}} = 0.139\ M$

A common way to state the concentration of this solution is to say, "This solution is 0.139 *M* in glucose."

(b) $\text{volume} = \dfrac{0.250\text{ mol } C_6H_{12}O_6}{0.139\ M} = 1.80\text{ L}$

(c) moles of $C_6H_{12}O_6$ in 0.500 L = 0.500 L × 0.139 *M* = 0.0695 mol

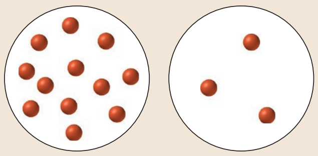
THINK ABOUT IT

Check to see that the magnitudes of your answers are logical. For example, the mass given in the problem corresponds to 0.277 mol of solute. If you are asked, as in part (b), for the volume that contains a number of moles smaller than 0.277, make sure your answer is smaller than the original volume.

Practice Problem ATTEMPT For an aqueous solution of sucrose ($C_{12}H_{22}O_{11}$), determine (a) the molarity of 5.00 L of a solution that contains 235 g of sucrose, (b) the volume of this solution that would contain 1.26 mol of sucrose, and (c) the number of moles of sucrose in 1.89 L of this solution.

Practice Problem BUILD For an aqueous solution of sodium chloride (NaCl), determine (a) the molarity of 3.75 L of a solution that contains 155 g of sodium chloride, (b) the volume of this solution that would contain 4.58 mol of sodium chloride, and (c) the number of moles of sodium chloride in 22.75 L of this solution.

Practice Problem CONCEPTUALIZE The diagrams represent solutions of two different concentrations. What volume of solution 2 contains the same amount of solute as 5.00 mL of solution 1? What volume of solution 1 contains the same amount of solute as 30.0 mL of solution 2?



solution 1 solution 2

Strategy: plan is laid out for solving the problem.

Setup: necessary information is gathered and organized.

Solution: problem is worked out.

Think About It:
– Assess the result.
– Provides information that shows the relevance of the result or the technique.
– Sometimes shows an alternate route to the same answer.

Each Sample Problem is followed by my ABC approach of three Practice Problems: **A**ttempt, **B**uild, and **C**onceptualize.

ATTEMPT

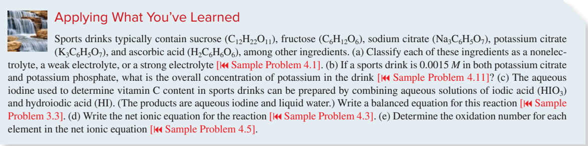
Practice Problem **A** (or "**A**ttempt") asks the student to apply the same Strategy to solve a problem very similar to the Sample Problem. In general, the same Setup and series of steps in the Solution can be used to solve Practice Problem A.

BUILD

Practice Problem **B** (or "**B**uild") assesses mastery of the same skills as those required for the Sample Problem and Practice Problem A, but everywhere possible; Practice Problem B cannot be solved using the same Strategy used for the Sample Problem and for Practice Problem A. This provides the student an opportunity to develop a strategy independently, and combats the tendency that some students have to want to apply a "template" approach to solving chemistry problems. Practice Problems "Attempt" and "Build" have been incorporated into the problems available in Connect (R) and can be used in online homework and/or quizzing.

CONCEPTUALIZE

Practice Problem **C** (or "**C**onceptualize") provides an exercise that probes the student's conceptual understanding of the material. Practice Problems C often include concept and molecular art.


Applying What You've Learned

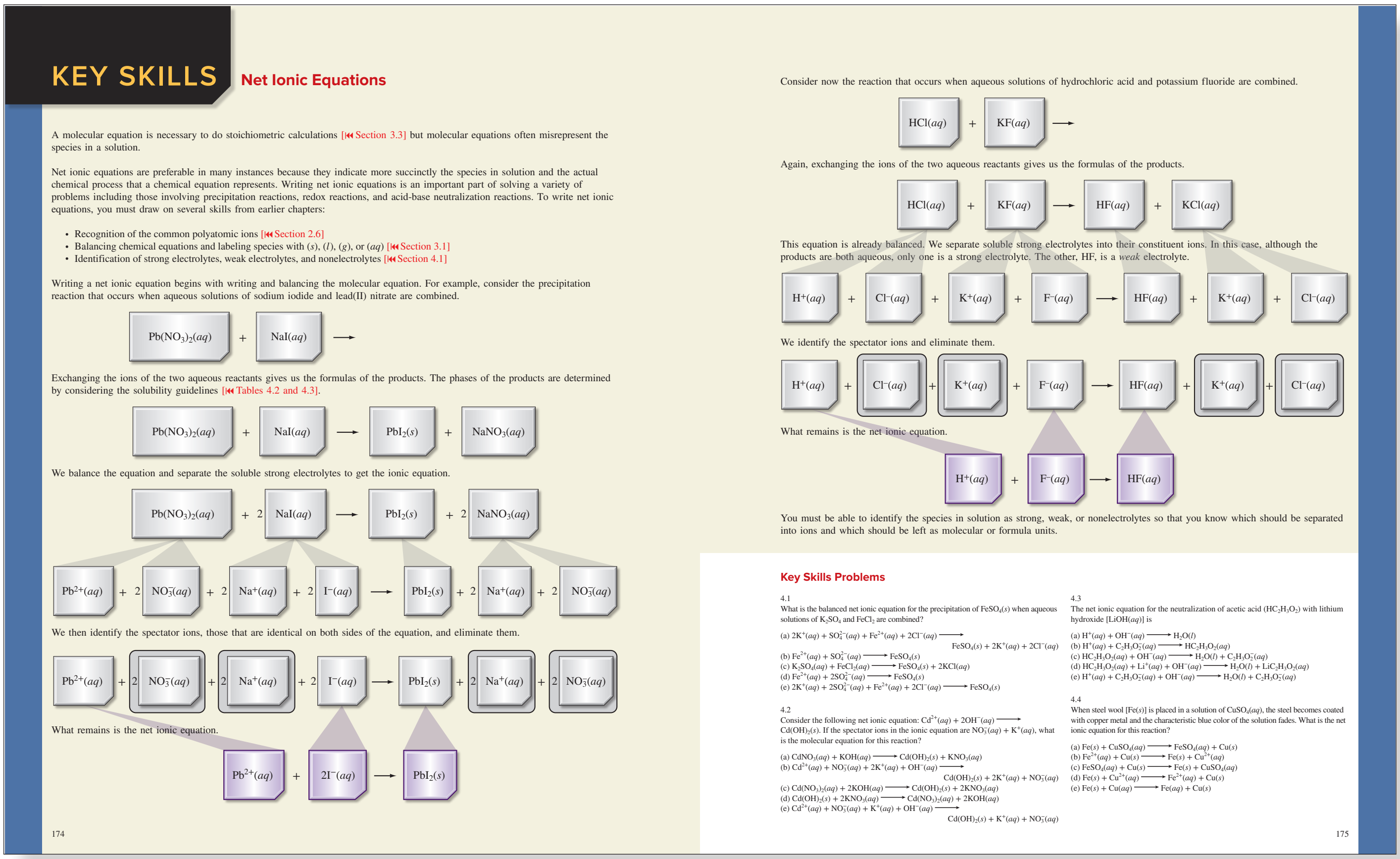
Sports drinks typically contain sucrose ($C_{12}H_{22}O_{11}$), fructose ($C_6H_{12}O_6$), sodium citrate ($Na_3C_6H_5O_7$), potassium citrate ($K_3C_6H_5O_7$), and ascorbic acid ($H_2C_6H_6O_6$), among other ingredients. (a) Classify each of these ingredients as a nonelectrolyte, a weak electrolyte, or a strong electrolyte [⏮ Sample Problem 4.1]. (b) If a sports drink is 0.0015 *M* in both potassium citrate and potassium phosphate, what is the overall concentration of potassium in the drink [⏮ Sample Problem 4.11]? (c) The aqueous iodine used to determine vitamin C content in sports drinks can be prepared by combining aqueous solutions of iodic acid (HIO_3) and hydroiodic acid (HI). (The products are aqueous iodine and liquid water.) Write a balanced equation for this reaction [⏮ Sample Problem 3.3]. (d) Write the net ionic equation for the reaction [⏮ Sample Problem 4.3]. (e) Determine the oxidation number for each element in the net ionic equation [⏮ Sample Problem 4.5].

Each chapter's end-of-chapter questions and problems begin with an **Integrative Problem,** titled *Applying What You've Learned.* These integrative problems incorporate multiple concepts from the chapter, with each step of the problem providing a specific reference to the appropriate Sample Problem in case the student needs direction.

New Pedagogy

Key Skills

Newly located immediately before the end-of-chapter problems, Key Skills pages are modules that provide a review of specific problem-solving techniques from that particular chapter. These are techniques the author knows are vital to success in later chapters. The Key Skills pages are designed to be easy for students to find touchstones to hone specific skills from earlier chapters—in the context of later chapters. The answers to the Key Skills Problems can be found in the Answer Appendix in the back of the book.


KEY SKILLS Net Ionic Equations

A molecular equation is necessary to do stoichiometric calculations [⏮ Section 3.3] but molecular equations often misrepresent the species in a solution.

Net ionic equations are preferable in many instances because they indicate more succinctly the species in solution and the actual chemical process that a chemical equation represents. Writing net ionic equations is an important part of solving a variety of problems including those involving precipitation reactions, redox reactions, and acid-base neutralization reactions. To write net ionic equations, you must draw on several skills from earlier chapters:

- Recognition of the common polyatomic ions [⏮ Section 2.6]
- Balancing chemical equations and labeling species with (*s*), (*l*), (*g*), or (*aq*) [⏮ Section 3.1]
- Identification of strong electrolytes, weak electrolytes, and nonelectrolytes [⏮ Section 4.1]

Writing a net ionic equation begins with writing and balancing the molecular equation. For example, consider the precipitation reaction that occurs when aqueous solutions of sodium iodide and lead(II) nitrate are combined.

$$Pb(NO_3)_2(aq) + NaI(aq) \longrightarrow$$

Exchanging the ions of the two aqueous reactants gives us the formulas of the products. The phases of the products are determined by considering the solubility guidelines [⏮ Tables 4.2 and 4.3].

$$Pb(NO_3)_2(aq) + NaI(aq) \longrightarrow PbI_2(s) + NaNO_3(aq)$$

We balance the equation and separate the soluble strong electrolytes to get the ionic equation.

$$Pb(NO_3)_2(aq) + 2\,NaI(aq) \longrightarrow PbI_2(s) + 2\,NaNO_3(aq)$$

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,Na^+(aq) + 2\,I^-(aq) \longrightarrow PbI_2(s) + 2\,Na^+(aq) + 2\,NO_3^-(aq)$$

We then identify the spectator ions, those that are identical on both sides of the equation, and eliminate them.

$$Pb^{2+}(aq) + \boxed{2\,NO_3^-(aq)} + \boxed{2\,Na^+(aq)} + 2\,I^-(aq) \longrightarrow PbI_2(s) + \boxed{2\,Na^+(aq)} + \boxed{2\,NO_3^-(aq)}$$

What remains is the net ionic equation.

$$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$$

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Consider now the reaction that occurs when aqueous solutions of hydrochloric acid and potassium fluoride are combined.

$$HCl(aq) + KF(aq) \longrightarrow$$

Again, exchanging the ions of the two aqueous reactants gives us the formulas of the products.

$$HCl(aq) + KF(aq) \longrightarrow HF(aq) + KCl(aq)$$

This equation is already balanced. We separate soluble strong electrolytes into their constituent ions. In this case, although the products are both aqueous, only one is a strong electrolyte. The other, HF, is a *weak* electrolyte.

$$H^+(aq) + Cl^-(aq) + K^+(aq) + F^-(aq) \longrightarrow HF(aq) + K^+(aq) + Cl^-(aq)$$

We identify the spectator ions and eliminate them.

$$H^+(aq) + \boxed{Cl^-(aq)} + \boxed{K^+(aq)} + F^-(aq) \longrightarrow HF(aq) + \boxed{K^+(aq)} + \boxed{Cl^-(aq)}$$

What remains is the net ionic equation.

$$H^+(aq) + F^-(aq) \longrightarrow HF(aq)$$

You must be able to identify the species in solution as strong, weak, or nonelectrolytes so that you know which should be separated into ions and which should be left as molecular or formula units.

Key Skills Problems

4.1
What is the balanced net ionic equation for the precipitation of $FeSO_4(s)$ when aqueous solutions of K_2SO_4 and $FeCl_2$ are combined?

(a) $2K^+(aq) + SO_4^{2-}(aq) + Fe^{2+}(aq) + 2Cl^-(aq) \longrightarrow FeSO_4(s) + 2K^+(aq) + 2Cl^-(aq)$
(b) $Fe^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow FeSO_4(s)$
(c) $K_2SO_4(aq) + FeCl_2(aq) \longrightarrow FeSO_4(s) + 2KCl(aq)$
(d) $Fe^{2+}(aq) + 2SO_4^{2-}(aq) \longrightarrow FeSO_4(s)$
(e) $2K^+(aq) + 2SO_4^{2-}(aq) + Fe^{2+}(aq) + 2Cl^-(aq) \longrightarrow FeSO_4(s)$

4.2
Consider the following net ionic equation: $Cd^{2+}(aq) + 2OH^-(aq) \longrightarrow Cd(OH)_2(s)$. If the spectator ions in the ionic equation are $NO_3^-(aq) + K^+(aq)$, what is the molecular equation for this reaction?

(a) $CdNO_3(aq) + KOH(aq) \longrightarrow Cd(OH)_2(s) + KNO_3(aq)$
(b) $Cd^{2+}(aq) + NO_3^-(aq) + 2K^+(aq) + OH^-(aq) \longrightarrow Cd(OH)_2(s) + 2K^+(aq) + NO_3^-(aq)$
(c) $Cd(NO_3)_2(aq) + 2KOH(aq) \longrightarrow Cd(OH)_2(s) + 2KNO_3(aq)$
(d) $Cd(OH)_2(s) + 2KNO_3(aq) \longrightarrow Cd(NO_3)_2(aq) + 2KOH(aq)$
(e) $Cd^{2+}(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \longrightarrow Cd(OH)_2(s) + K^+(aq) + NO_3^-(aq)$

4.3
The net ionic equation for the neutralization of acetic acid ($HC_2H_3O_2$) with lithium hydroxide [$LiOH(aq)$] is

(a) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
(b) $H^+(aq) + C_2H_3O_2^-(aq) \longrightarrow HC_2H_3O_2(aq)$
(c) $HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$
(d) $HC_2H_3O_2(aq) + Li^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + LiC_2H_3O_2(aq)$
(e) $H^+(aq) + C_2H_3O_2^-(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$

4.4
When steel wool [$Fe(s)$] is placed in a solution of $CuSO_4(aq)$, the steel becomes coated with copper metal and the characteristic blue color of the solution fades. What is the net ionic equation for this reaction?

(a) $Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$
(b) $Fe^{2+}(aq) + Cu(s) \longrightarrow Fe(s) + Cu^{2+}(aq)$
(c) $FeSO_4(aq) + Cu(s) \longrightarrow Fe(s) + CuSO_4(aq)$
(d) $Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$
(e) $Fe(s) + Cu(aq) \longrightarrow Fe(aq) + Cu(s)$

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New to the Fifth Edition

- Use of student Heat Maps to improve presentation specifically based on student performance.
- **New chapter openers,** with emphasis on the chemistry associated with global climate change.
- **New End-of-Chapter Problems** have been added in response to user comments. These include additional conceptual problems, additional multi-concept problems, and updates of information in topical questions.
- **New Sample Problems** to improve the introduction of new concepts.
- **New figures** to help students develop conceptual understanding.
- **Continued development of truly comprehensive and consistent problem-solving.** Hundreds of worked examples (Sample Problems) help students get started learning how to approach and solve problems.
- **SmartBook™ with Learning Resources.** Our adaptive SmartBook has been supplemented with additional learning resources tied to each learning objective to provide point-in-time help to students who need it.

New and updated chapter content includes:
Incorporation of essential information from student notes into the main flow of text in each chapter. The remaining student notes are designed to help students over a variety of stumbling blocks. They include timely warnings about common errors, reminders of important information from previous chapters, and general information that helps place the material in an easily understood context.

Chapter 1—New chapter opener with environmental focus and earlier placement of the FAQ box "How Can I Enhance My Chances of Success in Chemistry Class?"

Chapter 2—New end-of-chapter problems

Chapter 5—New end-of-chapter problems, including conceptual and multi-concepts problems

Chapter 6—New conceptual illustration of the photoelectric effect

Chapter 7—New chapter opener with environmental focus

Chapter 8—New conceptual end-of-chapter problems

Chapter 10—New conceptual end-of-chapter problems

Chapter 11—New conceptual end-of-chapter problems

Chapter 17—New chapter opener with environmental focus and new conceptual end-of-chapter problems

Chapter 18—New conceptual Checkpoint and end-of-chapter problems

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All students will have access to **chemistry animations** for the animated Visualizing Chemistry figures as well as other chemistry animations in Connect. Within the text, the animations are mapped to the appropriate content.

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Additionally, students can purchase a Student Solution Manual that contains detailed solutions and explanations for the odd-numbered problems in the main text.

For me, this text will always remain a work in progress. I encourage you to contact me with any comments or questions.

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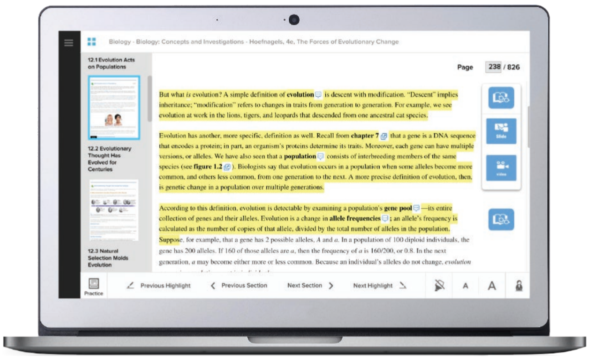

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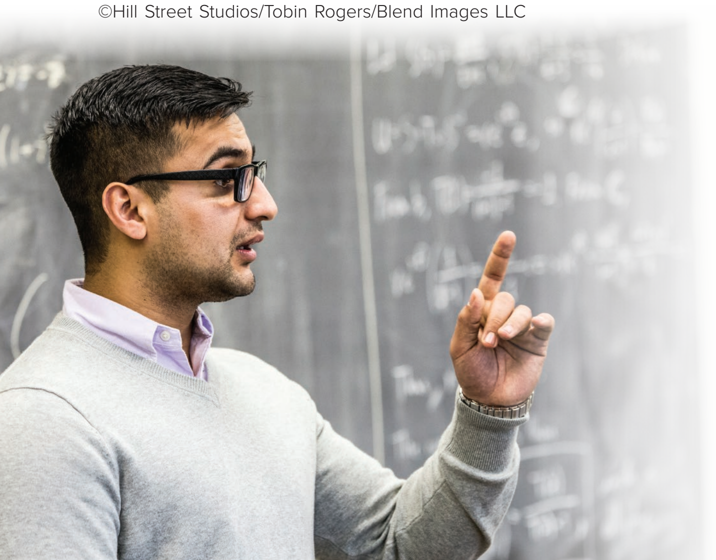

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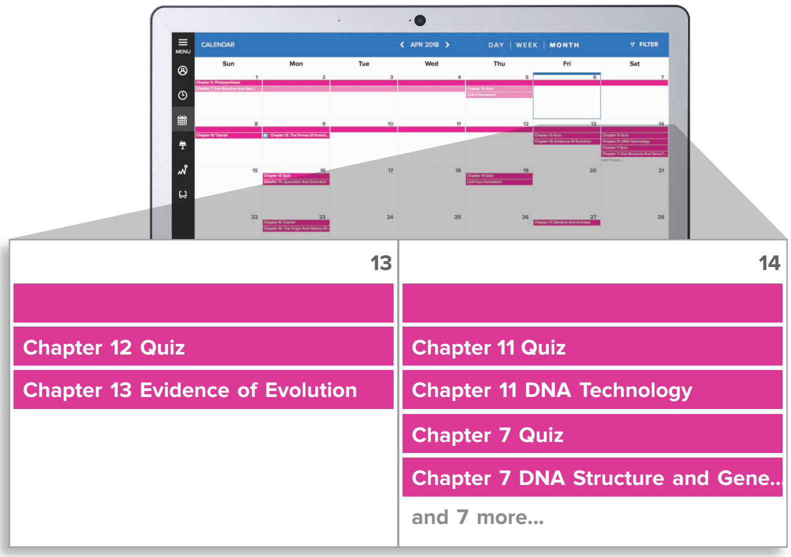

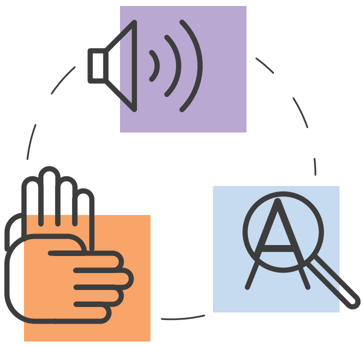

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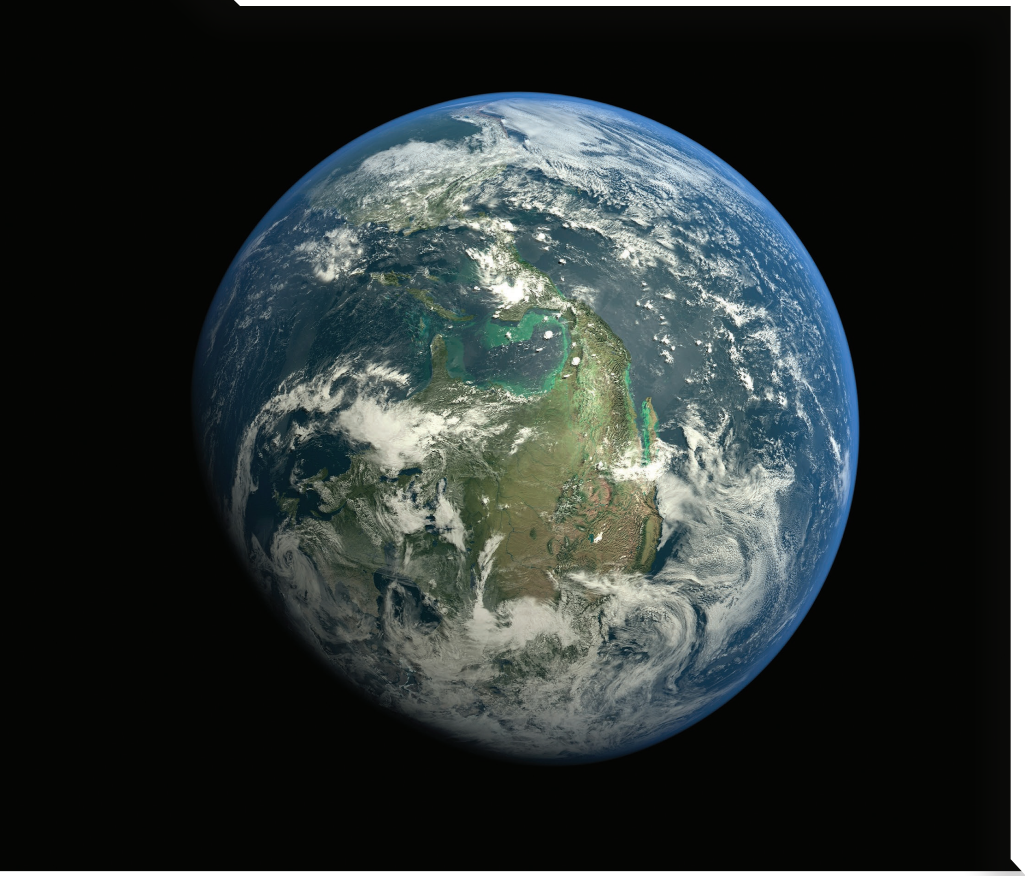
Chemistry

CHAPTER 1

Chemistry: The Central Science



Earth photographed from space.

In This Chapter, You Will Learn

Some of what chemistry is and how it is studied using the scientific method. You will learn about the system of units used by scientists and about expressing and dealing with the numbers that result from scientific measurements.

Before You Begin, Review These Skills

- Basic algebra
- Scientific notation [⏭ Appendix 1]

Global Climate Change and the Scientific Method

To advance understanding of science, researchers use a set of guidelines known as the *scientific method*. The guidelines involve careful observations, educated reasoning, and the development and experimental testing of hypotheses and theories. One field of study in which the scientific method has informed our understanding of the world is that of *global climate change*.

Late in the nineteenth century, Swedish chemist Svante Arrhenius used the principles of chemistry to describe the "greenhouse effect," the process by which certain components of the atmosphere absorb some of the energy radiating from Earth's surface and prevent it from escaping into space—thereby warming the planet. The greenhouse effect is a natural phenomenon, responsible in part for Earth's average global temperature being hospitable to humans and other forms of life. But Arrhenius also predicted what he perceived to be an inevitable, eventual consequence of the burning of coal and other fossil fuels, which increased significantly during the industrial revolution. He believed that, unchecked, the dramatic increase in atmospheric CO_2 caused by human activities would cause a potentially dangerous increase in global temperature via the "enhanced greenhouse effect."

Several groups of climate scientists, including those at the National Aeronautics and Space Administration's Goddard Institute for Space Studies (NASA/GISS) at Columbia University, study global temperature trends by analyzing observations from many thousands of data sets gathered using a variety of different measurement techniques over the course of more than a century. Their findings have consistently validated Arrhenius's prediction. There is no doubt that the temperature of our planet is increasing. Moreover, the connection between global temperature change and human activities—most importantly the burning of fossil fuels—is undeniable.

The issue of global climate change is one that appears frequently in the popular press. Unfortunately, it has become something of a political issue, with some people dismissing its importance or denying its existence outright. As a student of science, you will want to develop an informed perspective. To do this, you must understand how observations, hypotheses, theories, and experimentation contribute to a self-correcting scientific narrative; and how they have given rise to the current scientific consensus regarding climate change and humankind's role in it.

At the end of this chapter, you will be able to answer several questions related to the study of global climate change **[⏭ Applying What You've Learned, page 30]****.**

1.1 The Study of Chemistry

Chemistry often is called the *central science* because knowledge of the principles of chemistry can facilitate understanding of other sciences, including physics, biology, geology, astronomy, oceanography, engineering, and medicine. ***Chemistry*** is the study of *matter* and the *changes* that matter undergoes. Matter is what makes up our bodies, our belongings, our physical environment, and in fact our universe. ***Matter*** is anything that has mass and occupies space.

Although it can take many different forms, all matter consists of various combinations of atoms of only a relatively small number of simple substances called *elements*. The properties of matter depend on which of these elements it contains and on how the atoms of those elements are arranged.

Chemistry You May Already Know

You may already be familiar with some of the terms used in chemistry. Even if this is your first chemistry course, you may have heard of *molecules* and know them to be tiny pieces of a substance—much too tiny to see. Further, you may know that molecules are made up of *atoms,* even smaller pieces of matter. And even if you don't know what a chemical formula is, you probably know that H_2O is water and CO_2 is carbon dioxide. You may have used, or at least heard, the term *chemical reaction;* and you are undoubtedly familiar with a variety of chemical reactions, such as those shown in Figure 1.1.

Familiar chemical reactions, such as those shown in Figure 1.1, are all things that you can observe at the *macroscopic level*. In other words, these processes and their results are visible to the human eye. In studying chemistry, you will learn to understand and visualize many of these processes at the *molecular level*.

Because atoms and molecules are far too small to observe directly, we need a way to visualize them. One way is through the use of molecular models. Throughout this book, we will represent matter at the molecular level using *molecular art,* the two-dimensional equivalent of molecular models. In these pictures, atoms are represented as spheres, and atoms of particular

(a)

(b)

(c)

(d)

(e)

Figure 1.1 Many familiar processes are chemical reactions: (a) The flame of a creme brulee torch is the combustion of butane. (b) The bubbles produced when Alka-Seltzer dissolves in water are carbon dioxide, produced by a chemical reaction between two ingredients in the tablets. (c) The formation of rust is a chemical reaction that occurs when iron, water, and oxygen are all present. (d) Many baked goods "rise" as the result of a chemical reaction that produces carbon dioxide. (e) The glow produced when luminol is used to detect traces of blood in crime-scene investigations is the result of a chemical reaction.

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FAQ How Can I Enhance My Chances of Success in Chemistry Class?

Success in a chemistry class depends largely on problem-solving ability. The Sample Problems throughout this text are designed to help you develop problem-solving skills. Each is divided into four steps: Strategy, Setup, Solution, and Think About It.

Strategy: Read the problem carefully and determine what is being asked and what information is provided. The Strategy step is where you should think about what skills are required and lay out a plan for solving the problem. Give some thought to what you expect the result to be. If you are asked to determine the number of atoms in a sample of matter, for example, you should expect the answer to be a whole number. Determine what, if any, units should be associated with the result. When possible, make a ballpark estimate of the magnitude of the correct result, and make a note of your estimate.

Setup: Next, gather the information necessary to solve the problem. Some of the information will have been given in the problem itself. Other information, such as equations, constants, and tabulated data (including atomic masses), should also be brought together in this step. Write down and label clearly all of the information you will use to solve the problem. Be sure to write appropriate units with each piece of information.

Solution: Using the necessary equations, constants, and other information, calculate the answer to the problem. Pay particular attention to the units associated with each number, tracking and canceling units throughout the calculation. In the event that multiple calculations are required, carefully label any intermediate results.

Think About It: Consider your calculated result and ask yourself whether or not it makes sense. Compare the units and the magnitude of your result with your ballpark estimate from the Strategy step. If your result does not have the appropriate units, or if its magnitude or sign is not reasonable, check your solution for possible errors. A very important part of problem solving is being able to judge whether the answer is reasonable. It is relatively easy to spot a wrong sign or incorrect units, but you should also develop a sense of magnitude and be able to tell when an answer is either way too big or way too small. For example, if a problem asks how many molecules are in a sample and you calculate a number that is less than 1, you should know that it cannot be correct.

For additional practice, each Sample Problem is followed by three Practice Problems: A, B, and C. Practice Problem A, "Attempt," typically is very similar to the Sample Problem and can be solved using the same strategy. Practice Problem B, "Build," generally tests the same skills as Practice Problem A, but usually requires a slightly different approach. Practice Problem B lets you practice devising your own problem-solving strategy—an indispensable skill in any science curriculum. Practice Problem C, "Conceptualize," specifically probes your understanding of the underlying chemical concepts associated with the Sample Problem.

Regular use of the Sample Problems and Practice Problems A, B, and C in this text can help you develop an effective set of problem-solving skills. They can also help you assess whether you are ready to move on to the next new concepts. If you struggle with the Practice Problems, then you probably need to review the corresponding Sample Problem and the concepts that led up to it.

elements are represented using specific colors. Table 1.1 lists some of the elements that you will encounter most often and the colors used to represent them in this book.

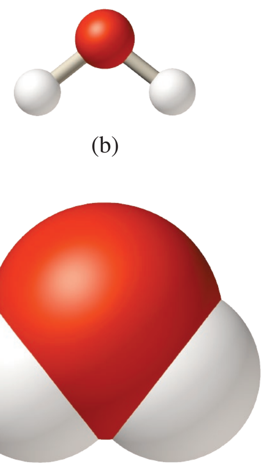
Molecular art can be of *ball-and-stick* models, in which the bonds connecting atoms appear as sticks [Figure 1.2(b)], or of *space-filling* models, in which the atoms appear to overlap one another [Figure 1.2(c)]. Ball-and-stick and space-filling models illustrate the specific, three-dimensional arrangement of the atoms. The ball-and-stick model does a good job of illustrating the arrangement of atoms, but exaggerates the distances between atoms, relative to their sizes. The space-filling model gives a more accurate picture of these *interatomic* distances but can obscure the details of the three-dimensional arrangement.

TABLE 1.1 Colors of Elements Commonly Used in Molecular Art

Hydrogen	Sodium
Boron	Phosphorus
Carbon	Sulfur
Nitrogen	Chlorine
Oxygen	Bromine
Fluorine	Iodine

H_2O

(a)



(b)

(c)

Figure 1.2 Water represented with a (a) molecular formula, (b) ball-and-stick model, and (c) space-filling model.

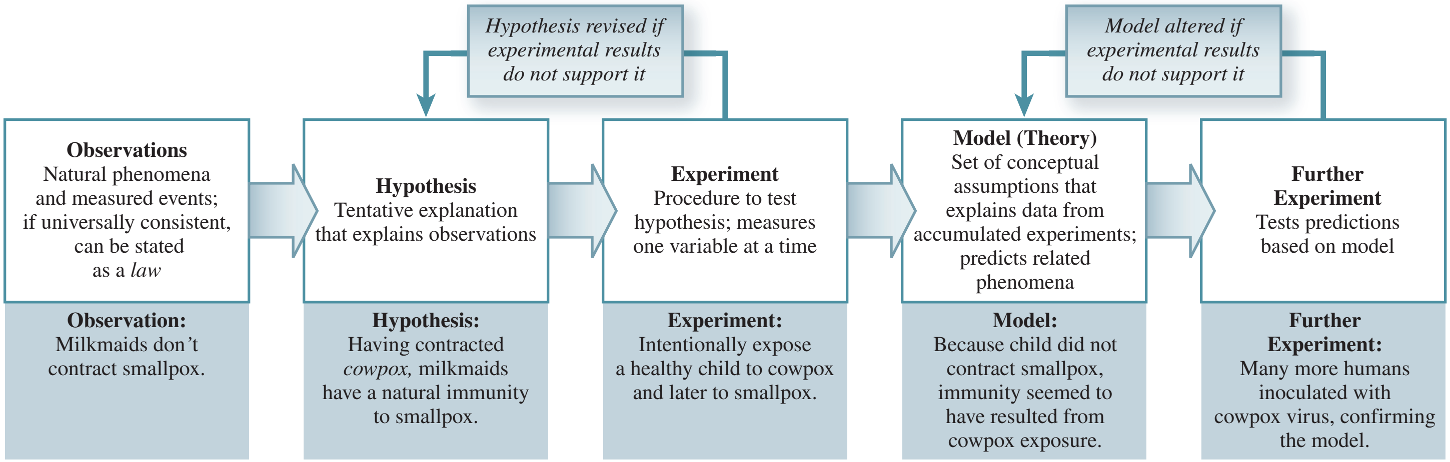


Figure 1.3 Flowchart of the scientific method.

The Scientific Method

Experiments are the key to advancing our understanding of chemistry—or any science. Although not all scientists will necessarily take the same approach to experimentation, they all follow a set of guidelines known as the ***scientific method*** to add their results to the larger body of knowledge within a given field. The flowchart in Figure 1.3 illustrates this basic process. The method begins with the gathering of data via observations and experiments. Scientists study these data and try to identify *patterns* or *trends*. When they find a pattern or trend, they may summarize their findings with a ***law,*** a concise verbal or mathematical statement of a reliable relationship between phenomena. Scientists may then formulate a ***hypothesis,*** a tentative explanation for their observations. Further experiments are designed to test the hypothesis. If experiments indicate that the hypothesis is incorrect, the scientists go back to the drawing board, try to come up with a different interpretation of their data, and formulate a new hypothesis. The new hypothesis will then be tested by experiment. When a hypothesis stands the test of extensive experimentation, it may evolve into a theory. A ***theory*** is a unifying principle that explains a body of experimental observations and the laws that are based on them. Theories can also be used to predict related phenomena, so theories are constantly being tested. If a theory is disproved by experiment, then it must be discarded or modified so that it becomes consistent with experimental observations.

A fascinating example of the use of the scientific method is the story of how smallpox was eradicated. Late in the eighteenth century, an English doctor named Edward Jenner observed that even during outbreaks of smallpox in Europe, milkmaids seldom contracted the disease. He reasoned that when people who had frequent contact with cows contracted *cowpox,* a similar but far less harmful disease, they developed a natural immunity to smallpox. He predicted that intentional exposure to the cowpox virus would produce the same immunity. In 1796, Jenner exposed an 8-year-old boy to the cowpox virus using pus from the cowpox lesions of an infected milkmaid. Six weeks later, he exposed the boy to the *smallpox* virus and, as Jenner had predicted, the boy did *not* contract the disease. Subsequent experiments using the same technique (later dubbed *vaccination* from the Latin *vacca* meaning *cow*) confirmed that immunity to smallpox could be induced.

A superbly coordinated international effort on the part of healthcare workers was successful in eliminating smallpox worldwide. In 1980, the World Health Organization declared smallpox officially eradicated in nature. This historic triumph over a dreadful disease, one of the greatest medical advances of the twentieth century, began with Jenner's astute observations, inductive reasoning, and careful experimentation—the essential elements of the *scientific method.*

1.2 Classification of Matter

Chemists classify matter as either a *substance* or a *mixture* of substances. A ***substance*** is a form of matter that has a specific composition and distinct properties. Examples are salt (sodium chloride), iron, water, mercury, carbon dioxide, and oxygen. Substances can be further classified as either *elements* (such as iron, mercury, and oxygen) or *compounds* (such as salt, water, and carbon dioxide). Different substances differ from one another in composition and properties, and each can be identified by its appearance, taste, smell, or other properties.

Student Note: Some books refer to substances as *pure substances.* These two terms generally mean the same thing although the adjective *pure* is unnecessary in this context because a substance is, by definition, pure.

States of Matter

Every substance can, in principle, exist as a solid, a liquid, and a gas, the three physical states depicted in Figure 1.4. Solids and liquids sometimes are referred to collectively as the *condensed phases*. Liquids and gases sometimes are referred to collectively as *fluids*. In a solid, particles are held close together in an orderly fashion with little freedom of motion. As a result, a solid does not conform to the shape of its container. Particles in a liquid are close together but are not held rigidly in position; they are free to move past one another. Thus, a liquid conforms to the shape of the part of the container it fills. In a gas, the particles are separated by distances that are very large compared to the size of the particles. A sample of gas assumes both the shape and the volume of its container.

The three states of matter can be interconverted without changing the chemical composition of the substance. Upon heating, a solid (e.g., ice) will melt to form a liquid (water). Further heating will vaporize the liquid, converting it to a gas (water vapor). Conversely, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid form. Figure 1.5 shows the three physical states of water.

Animation
Matter—three states of matter.

Elements

An ***element*** is a substance that cannot be separated into simpler substances by chemical means. Iron, mercury, oxygen, and hydrogen are just 4 of the 118 elements that have been identified. Most of the known elements occur naturally on Earth. The others have been produced by scientists via nuclear processes, which are discussed in Chapter 20. As shown in Figure 1.6(a) and (b), an element may consist of atoms or molecules.

For convenience, chemists use symbols of one or two letters to represent the elements. Only the first letter of an element's chemical symbol is capitalized. A list of the elements and their symbols appears at the beginning of this book. The symbols of some elements are derived from their Latin names—for example, Ag from *argentum* (silver), Pb from *plumbum* (lead), and Na from *natrium* (sodium)—while most of them come from their English names—for example, H for hydrogen, Co for cobalt, and Br for bromine.

Compounds

Most elements can combine with other elements to form compounds. Hydrogen gas, for example, burns in the presence of oxygen gas to form water, which has properties that are distinctly different from those of either hydrogen or oxygen. Thus, water is a ***compound,*** a substance composed of atoms of two or more elements chemically united in fixed proportions [Figure 1.6(c)]. The elements that make up a compound are called the compound's *constituent elements*. For example,

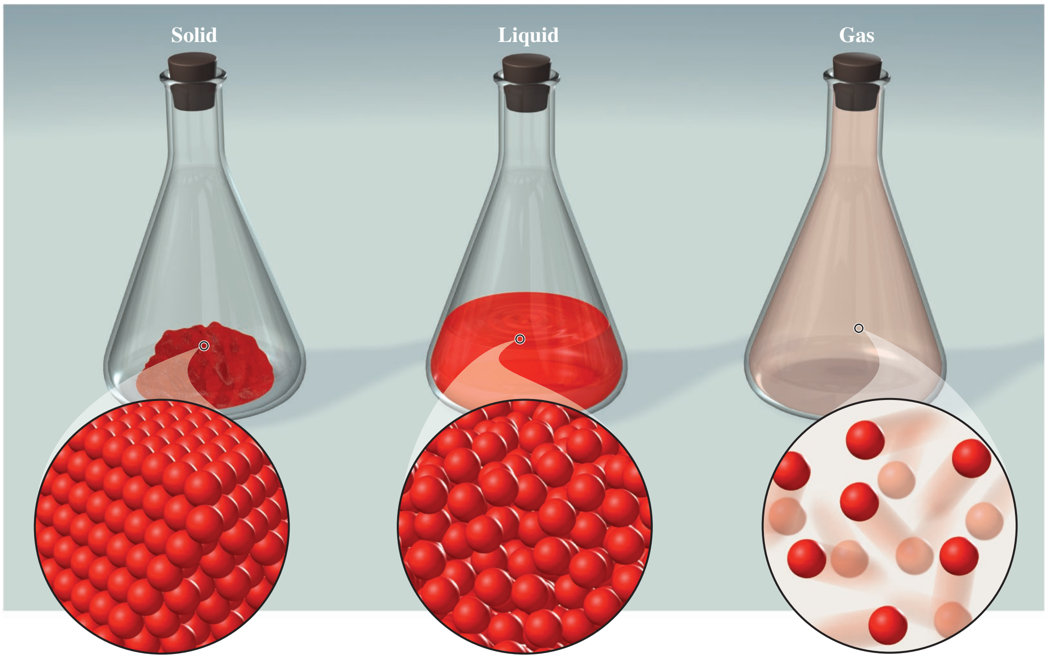


Figure 1.4 Molecular-level illustrations of a solid, liquid, and gas.

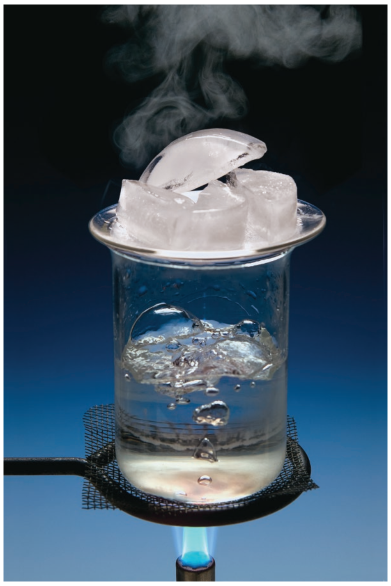


Figure 1.5 Water as a solid (ice), liquid, and gas. (We can't actually see water vapor, any more than we can see the nitrogen and oxygen that make up most of the air we breathe. When we see steam or clouds, what we are actually seeing is water vapor that has condensed upon encountering cold air.)

the constituent elements of water are hydrogen and oxygen; and water always contains twice as many hydrogen atoms as oxygen atoms (fixed proportions).

A compound cannot be separated into simpler substances by any physical process. (A physical process [⏭ Section 1.4] is one that does not change the identity of the matter. Examples of physical processes include boiling, freezing, and filtering.) Instead, the separation of a compound into its constituent elements requires a *chemical reaction*.

Mixtures

A ***mixture*** is a combination of two or more substances [Figure 1.6(d)] in which the substances retain their distinct identities. Like pure substances, mixtures can be solids, liquids, or gases. Some familiar examples are mixed nuts, 14-carat gold, apple juice, salt water, and air. Unlike compounds, mixtures do not have a universal constant composition. Therefore, samples of air collected in different locations will differ in composition because of differences in altitude, pollution, and other factors. The ratio of salt to water in different samples of salt water will vary depending on how they were prepared.

Mixtures are either *homogeneous,* having uniform composition throughout; or *heterogeneous,* having variable composition. When we dissolve a teaspoon of sugar in a glass of water, we get a ***homogeneous mixture.*** However, if we mix sand with iron filings, we get a a ***heterogeneous mixture*** in which the two substances remain distinct and discernible from each other (Figure 1.7).

Mixtures, whether homogeneous or heterogeneous, can be separated into pure components by physical means—without changing the identities of the components. Thus, sugar can be recovered from a water solution by evaporating the solution to dryness. Condensing the vapor will give us back the water component. To separate the sand–iron mixture, we can use a magnet to remove the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.7(b)]. After separation, the components of the mixture will have the same composition and properties as they did prior to being combined. The relationships among substances, elements, compounds, and mixtures are summarized in Figure 1.8.

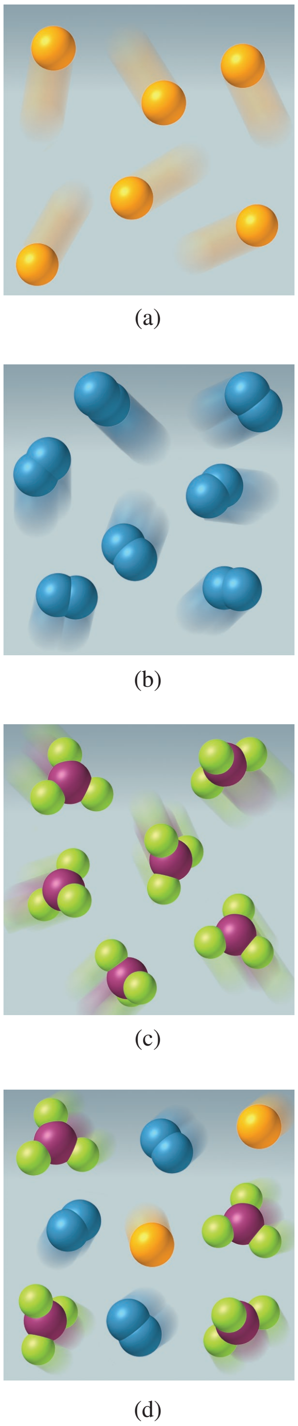


Figure 1.6 (a) Isolated atoms of an element. (b) Molecules of an element. (c) Molecules of a compound, consisting of more than one element. (d) A mixture of atoms of an element and molecules of an element and a compound.

Student Note: A compound may consist of *molecules* or *ions,* which we discuss in Chapter 2.

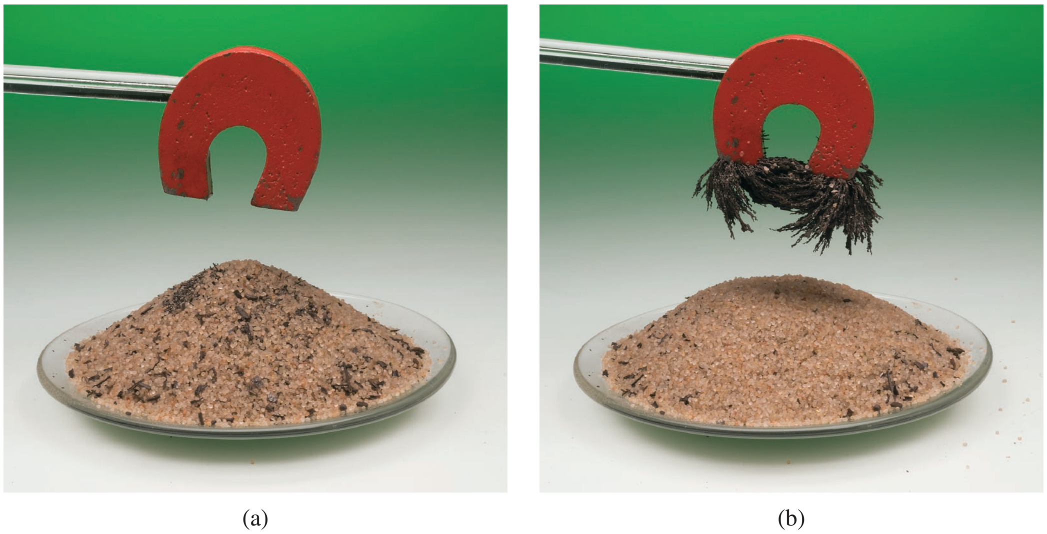


Figure 1.7 (a) A heterogeneous mixture contains iron filings and sand. (b) A magnet is used to separate the iron filings from the mixture.

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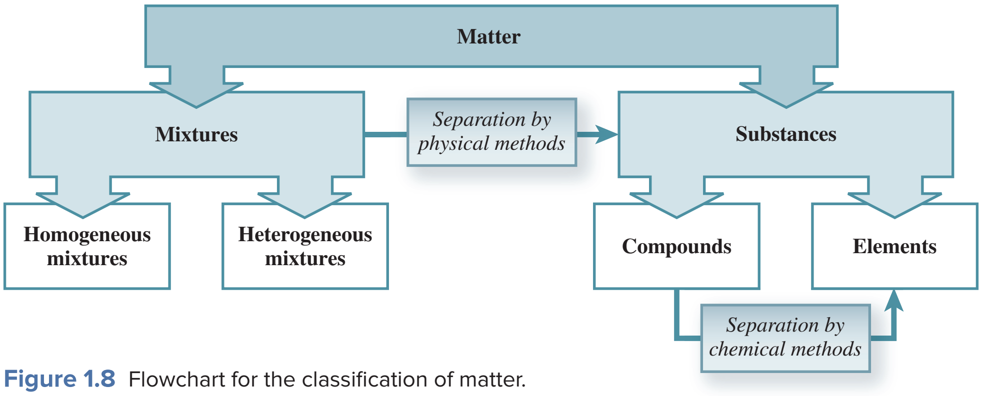


Figure 1.8 Flowchart for the classification of matter.

1.3 Scientific Measurement

Student Note: According to the U.S. Metric Association (USMA), the United States is "the only significant holdout" with regard to adoption of the metric system. The other countries that continue to use traditional units are Myanmar (formerly Burma) and Liberia.

Scientists use a variety of devices to measure the properties of matter. A meterstick is used to measure length; a burette, pipette, graduated cylinder, and volumetric flask are used to measure volume (Figure 1.9); a balance is used to measure mass; and a thermometer is used to measure temperature. Properties that can be measured are called *quantitative* properties because they are expressed using numbers. When we express a measured quantity with a number, though, we must always include the appropriate unit; otherwise, the measurement is meaningless. For example, to say that the depth of a swimming pool is 3 is insufficient to distinguish between one that is 3 *feet* (0.9 meter) and one that is 3 *meters* (9.8 feet) deep. Units are essential to reporting measurements correctly.

The two systems of units with which you are probably most familiar are the *English system* (foot, gallon, pound, etc.) and the *metric system* (meter, liter, kilogram, etc.). Although there has been an increase in the use of metric units in the United States in recent years, English units still are used commonly. For many years, scientists recorded measurements in metric units, but in 1960, the General Conference on Weights and Measures, the international authority on units, proposed a revised metric system for universal use by scientists. We use both metric and revised metric (SI) units in this book.

SI Base Units

The revised metric system is called the ***International System of Units*** (abbreviated SI, from the French *Système Internationale d'Unités*). Table 1.2 lists the seven SI base units. All other units of measurement can be derived from these base units. The ***SI unit*** for *volume,* for instance, is derived by cubing the SI base unit for *length*. The prefixes listed in Table 1.3 are used to denote decimal fractions and multiples of SI units. This enables scientists to tailor the magnitude of a unit to a particular application. For example, the meter (m) is appropriate for describing the dimensions of a classroom, but the kilometer (km), 1000 m, is more appropriate for describing the distance between two cities. Units that you will encounter frequently in the study of chemistry include those for mass, temperature, volume, and density.

Mass

Although the terms *mass* and *weight* often are used interchangeably, they do not mean the same thing. Strictly speaking, weight is the force exerted by an object or sample due to gravity. ***Mass***

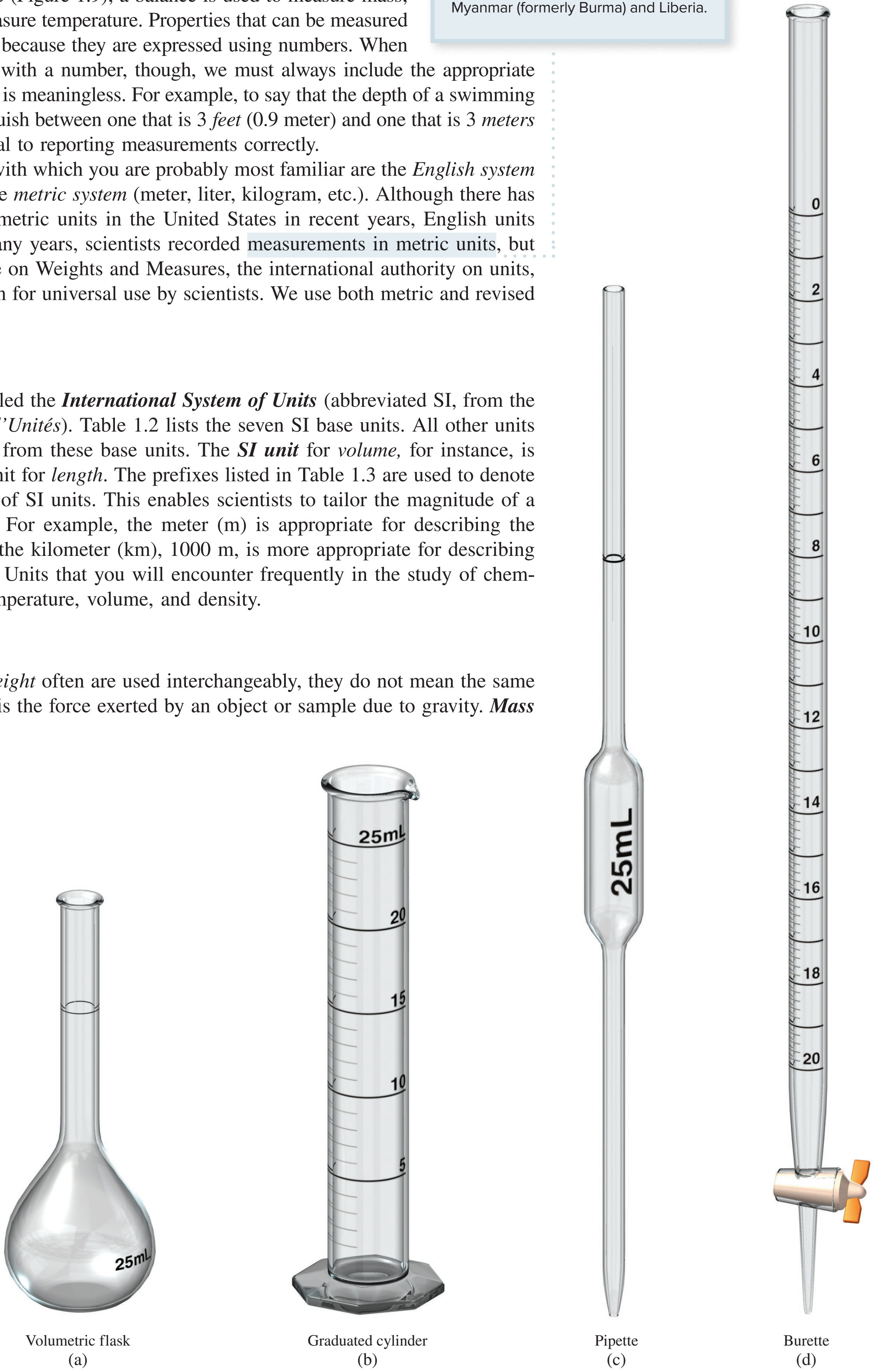


Figure 1.9 (a) A volumetric flask is used to prepare a precise volume of a solution for use in the laboratory. (b) A graduated cylinder is used to measure a volume of liquid. It is less precise than the volumetric flask. (c) A volumetric pipette is used to deliver a precise amount of liquid. (d) A burette is used to measure the volume of a liquid that has been added to a container. A reading is taken before and after the liquid is delivered, and the volume delivered is determined by subtracting the first reading from the second.

TABLE 1.2 Base SI Units

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

Student Note: Only one of the seven SI base units, the kilogram, itself contains a prefix.

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9} W
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6} Hz
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	μ	1×10^{-6} (0.000001)	1 microliter (μL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g

is a measure of the amount of matter in an object or sample. Because gravity varies from location to location (gravity on the moon is only about one-sixth that on Earth), the weight of an object varies depending on where it is measured. The mass of an object remains the same regardless of where it is measured. The SI base unit of mass is the kilogram (kg), but in chemistry the smaller gram (g) often is more convenient and is more commonly used:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

Temperature

There are two temperature scales used in chemistry. Their units are degrees Celsius (°C) and kelvin (K). The Celsius scale was originally defined using the freezing point (0°C) and the boiling point (100°C) of pure water at sea level. As Table 1.2 shows, the SI base unit of temperature is the ***kelvin.*** Kelvin is known as the *absolute* temperature scale, meaning that the lowest temperature possible is 0 K, a temperature referred to as "absolute zero." No *degree* sign (°) is used to represent a temperature on the Kelvin scale. The theoretical basis of the Kelvin scale has to do with the behavior of gases and is discussed in Chapter 10.

Units of the Celsius and Kelvin scales are equal in magnitude, so *a degree Celsius* is equivalent to *a kelvin*. Thus, if the temperature of an object increases by 5°C, it also increases by 5 K. Absolute zero on the Kelvin scale is equivalent to –273.15°C on the Celsius scale. We use the following equation to convert a temperature from units of degrees Celsius to kelvin:

Equation 1.1 $$\text{K} = {}^\circ\text{C} + 273.15$$

Depending on the precision required, the conversion from degrees Celsius to kelvin often is done simply by adding 273, rather than 273.15. Sample Problem 1.1 illustrates conversions between these two temperature scales.

SAMPLE PROBLEM 1.1

Normal human body temperature can range over the course of the day from about 36°C in the early morning to about 37°C in the afternoon. Express these two temperatures and the range that they span using the Kelvin scale.

Strategy Use Equation 1.1 to convert temperatures from the Celsius scale to the Kelvin scale. Then convert the range of temperatures from degrees Celsius to kelvin, keeping in mind that 1°C is equivalent to 1 K.

Setup Equation 1.1 is already set up to convert the two temperatures from degrees Celsius to kelvin. No further manipulation of the equation is needed. The range in kelvin will be the same as the range in degrees Celsius.

Solution 36°C + 273 = 309 K, 37°C + 273 = 310 K, and the range of 1°C is equal to a range of 1 K.

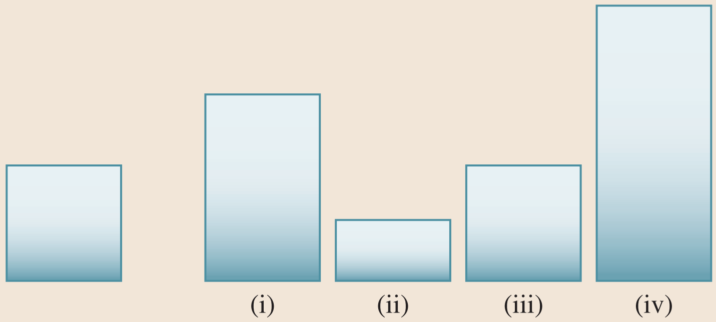
THINK ABOUT IT

Check your math and remember that converting a temperature from degrees Celsius to kelvin is different from converting a *difference* in temperature from degrees Celsius to kelvin.

Practice Problem ATTEMPT Express the freezing point of water (0°C), the boiling point of water (100°C), and the range spanned by the two temperatures using the Kelvin scale.

Practice Problem BUILD According to the website of the National Aeronautics and Space Administration (NASA), the average temperature of the universe is 2.7 K. Convert this temperature to degrees Celsius.

Practice Problem CONCEPTUALIZE If a single degree on the Celsius scale is represented by the rectangle on the left, which of the rectangles on the right best represents a single kelvin?



Bringing Chemistry to Life

Fahrenheit Temperature Scale

Outside of scientific circles, the Fahrenheit temperature scale is the one most used in the United States. Before the work of Daniel Gabriel Fahrenheit (German physicist, 1686–1736), there were numerous different, arbitrarily defined temperature scales, none of which gave consistent measurements. Accounts of exactly how Fahrenheit devised his temperature scale vary from source to source. In one account, in 1724, Fahrenheit labeled as 0° the lowest artificially attainable temperature at the time (the temperature of a mixture of ice, water, and ammonium chloride). Using a traditional scale consisting of 12 degrees, he labeled the temperature of a healthy human body as the twelfth degree. On this scale, the freezing point of water occurred at the fourth degree. For better resolution, each degree was further divided into eight smaller degrees. This convention makes the freezing point of water 32° and normal body temperature 96°. Today we consider normal body temperature to be somewhat higher than 96°F.

The boiling point of water on the Fahrenheit scale is 212°, meaning that there are 180 degrees (212° − 32°) between the freezing and boiling points. This separation is considerably more than the 100 degrees between the freezing point and boiling point of water on the Celsius scale [named after Swedish physicist Ander Celsius (1701–1744)]. Thus, the size of a degree on the Fahrenheit scale is only 100/180 or five-ninths of a degree on the Celsius scale. Equations 1.2 and 1.3 give the relationship between Fahrenheit and Celsius temperatures.

$$\text{temp in °C} = (\text{temp in °F} - 32\text{°F}) \times \frac{5\text{°C}}{9\text{°F}} \qquad \textbf{Equation 1.2}$$

and

$$\text{temp in °F} = \frac{9\text{°F}}{5\text{°C}} \times (\text{temp in °C}) + 32\text{°F} \qquad \textbf{Equation 1.3}$$

Sample Problem 1.2 illustrates the conversion between Celsius and Fahrenheit scales.

SAMPLE PROBLEM 1.2

A body temperature below 35.0°C constitutes hypothermia, whereas one above 39.0°C constitutes a high fever. Convert each of these temperatures to the Fahrenheit scale.

Strategy We are given temperatures in Celsius and are asked to convert them to Fahrenheit.

Setup We use Equation 1.3:

$$\text{temp in °F} = \frac{9\text{°F}}{5\text{°C}} \times (\text{temp in °C}) + 32\text{°F} \qquad \textbf{Equation 1.3}$$

Solution

$$\text{temp in °F} = \frac{9\text{°F}}{5\text{°C}} \times 35.0\text{°C} + 32\text{°F} = 95.0\text{°F}$$

$$\text{temp in °F} = \frac{9\text{°F}}{5\text{°C}} \times 39.0\text{°C} + 32\text{°F} = 102.2\text{°F}$$

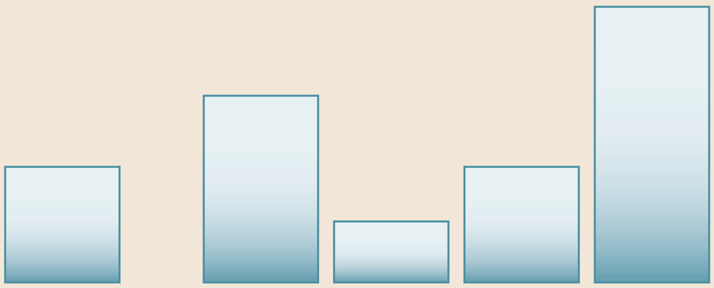
THINK ABOUT IT

"Normal" body temperature on the Fahrenheit scale is generally considered to be 98.6°F. The temperatures of hypothermia and high fever should be *below* and *above* that number, respectively. Therefore, 95.0°F and 102.2°F seem like reasonable results.

Practice Problem ATTEMPT Convert the temperatures 45.0°C and 90.0°C, and the difference between them, to degrees Fahrenheit.

Practice Problem BUILD In Ray Bradbury's 1953 novel *Fahrenheit 451,* 451°F is said to be the temperature at which books, which have been banned in the story, ignite. Convert 451°F to the Celsius scale.

Practice Problem CONCEPTUALIZE If a single degree on the Fahrenheit scale is represented by the rectangle on the left, which of the rectangles on the right best represents a single degree on the Celsius scale? Which best represents a single kelvin?



(i) (ii) (iii) (iv)

Oil floating on water is a familiar demonstration of density differences.

©David A. Tietz/Editorial Image, LLC

Derived Units: Volume and Density

There are many quantities, such as volume and density, that require units not included in the base SI units. In these cases, we must combine base units to *derive* appropriate units for the quantity.

The derived SI unit for volume, the meter cubed (m^3), is a larger volume than is practical in most laboratory settings. The more commonly used metric unit, the *liter* (L), is derived by cubing the *decimeter* (one-tenth of a meter) and is therefore also referred to as the cubic decimeter (dm^3). Another commonly used metric unit of volume is the *milliliter* (mL), which is derived by cubing the centimeter (1/100 of a meter). The milliliter is also referred to as the cubic centimeter (cm^3). Figure 1.10 illustrates the relationship between the liter (or dm^3) and the milliliter (or cm^3).

Density is the ratio of mass to volume. Oil floats on water, for example, because, in addition to not mixing with water, oil has a lower density than water. That is, given *equal volumes* of the two liquids, the oil will have a *smaller mass* than the water. Density is calculated using the following equation:

Equation 1.4 $$d = \frac{m}{V}$$

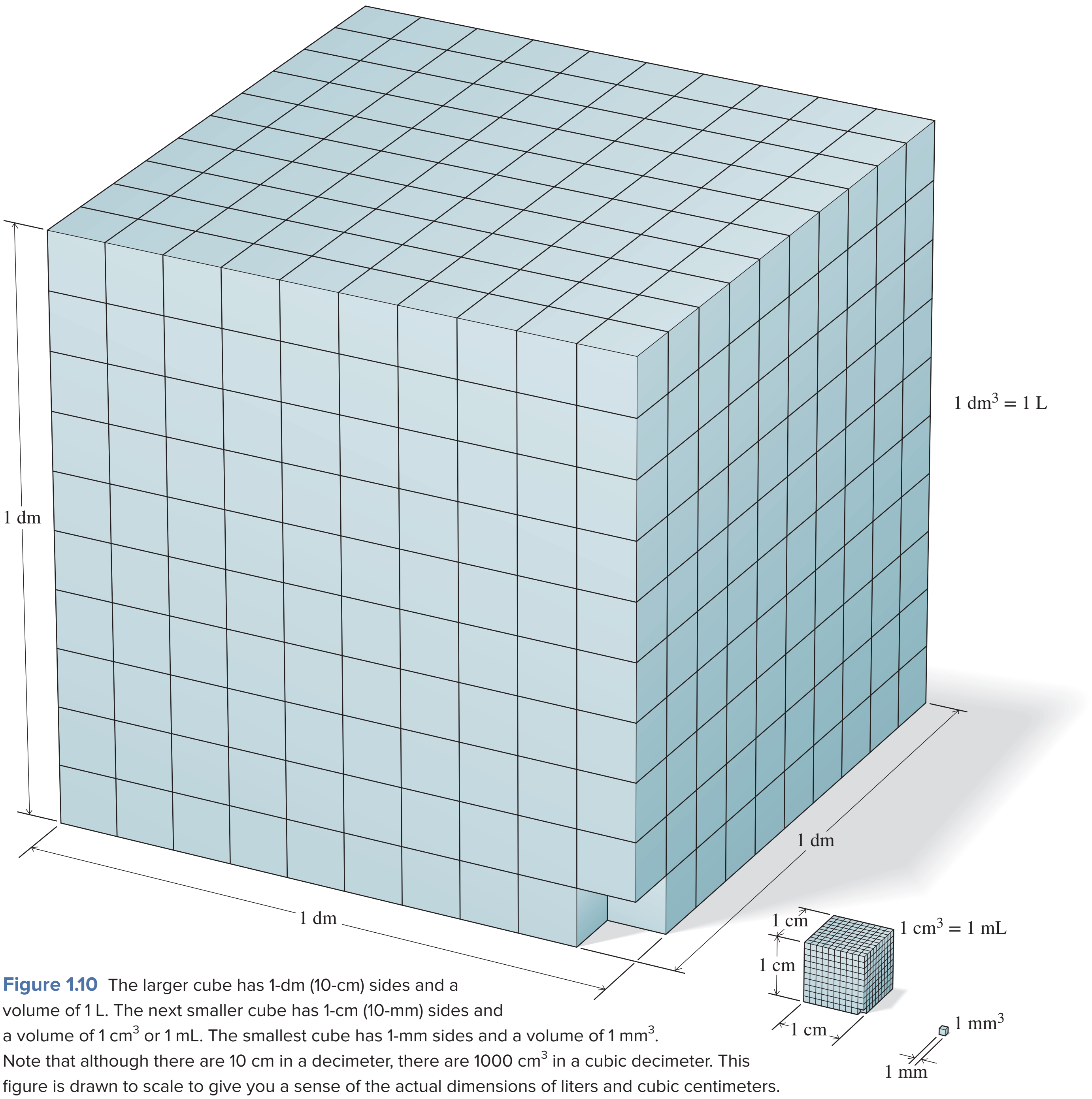


Figure 1.10 The larger cube has 1-dm (10-cm) sides and a volume of 1 L. The next smaller cube has 1-cm (10-mm) sides and a volume of 1 cm^3 or 1 mL. The smallest cube has 1-mm sides and a volume of 1 mm^3. Note that although there are 10 cm in a decimeter, there are 1000 cm^3 in a cubic decimeter. This figure is drawn to scale to give you a sense of the actual dimensions of liters and cubic centimeters.

where *d, m,* and *V* denote density, mass, and volume, respectively. The SI-derived unit for density is the kilogram per cubic meter (kg/m^3). This unit is too large for most common uses, however, so grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are used to express the densities of most solids and liquids. Water, for example, has a density of 1.00 g/cm^3 at 4°C. Because gas densities generally are very low, we typically express them in units of grams per liter (g/L):

$$1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1000 \text{ kg/m}^3$$

$$1 \text{ g/L} = 0.001 \text{ g/mL}$$

Sample Problem 1.3 illustrates density calculations.

SAMPLE PROBLEM 1.3

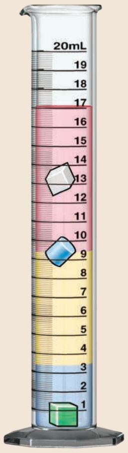
Ice cubes float in a glass of water because solid water is less dense than liquid water. (a) Calculate the density of ice given that, at 0°C, a cube that is 2.0 cm on each side has a mass of 7.36 g, and (b) determine the volume occupied by 23 g of ice at 0°C.

Strategy (a) Determine density by dividing mass by volume (Equation 1.4), and (b) use the calculated density to determine the volume occupied by the given mass.

Setup (a) We are given the mass of the ice cube, but we must calculate its volume from the dimensions given. The volume of the ice cube is $(2.0\text{ cm})^3$, or 8.0 cm^3. (b) Rearranging Equation 1.4 to solve for volume gives $V = m/d$.

Solution

(a) $d = \dfrac{7.36\text{ g}}{8.0\text{ cm}^3} = 0.92\text{ g/cm}^3$ or 0.92 g/mL (b) $V = \dfrac{23\text{ g}}{0.92\text{ g/cm}^3} = 25\text{ cm}^3$ or 25 mL



THINK ABOUT IT

For a sample with a density *less* than 1 g/cm^3, the number of cubic centimeters should be *greater* than the number of grams. In this case, $25\ (\text{cm}^3) > 23\ (\text{g})$.

Practice Problem ATTEMPT Given that 25.0 mL of mercury has a mass of 340.0 g, calculate (a) the density of mercury and (b) the volume of 155 g of mercury.

Practice Problem BUILD Calculate (a) the density of a solid substance if a cube measuring 2.33 cm on one side has a mass of 117 g and (b) the mass of a cube of the same substance measuring 7.41 cm on one side.

Practice Problem CONCEPTUALIZE Using the picture of the graduated cylinder and its contents, arrange the following in order of increasing density: blue liquid, pink liquid, yellow liquid, grey solid, blue solid, green solid.

The following box illustrates the importance of using units carefully in scientific work.

FAQ Why Are Units So Important?

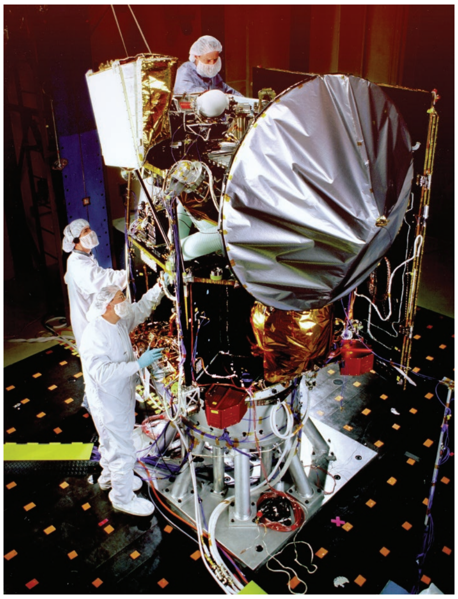
On December 11, 1998, NASA launched the 125-million-dollar Mars Climate Orbiter, which was intended to be the Red Planet's first weather satellite. After a 416-million-mile (mi) journey, the spacecraft was supposed to go into Mars's orbit on September 23, 1999. Instead, it entered Mars's atmosphere about 100 km (62 mi) lower than planned and was destroyed by heat. Mission controllers later determined that the spacecraft was lost because English measurement units were not converted to metric units in the navigation software.

Engineers at Lockheed Martin Corporation, who built the spacecraft, specified its thrust in pounds, which is an English unit of force. Scientists at NASA's Jet Propulsion Laboratory, on the other hand, who were responsible for deployment, had assumed that the thrust data they were given were expressed in *newtons,* a metric unit. To carry out the conversion between pound and newton, we would start with 1 lb = 0.4536 kg and, from Newton's second law of motion,

$$\begin{aligned}\text{force} &= (\text{mass})(\text{acceleration}) = (0.4536\text{ kg})(9.81\text{ m/s}^2)\\ &= 4.45\text{ kg}\cdot\text{m/s}^2 = 4.45\text{ N}\end{aligned}$$

because 1 newton (N) = $1\text{ kg}\cdot\text{m/s}^2$. Therefore, instead of converting 1 lb of *force* to 4.45 N, the scientists treated it as a force of 1 N. The considerably smaller engine thrust employed because of the engineers' failure to convert from English to metric units resulted in a lower orbit and the ultimate destruction of the spacecraft.

Commenting on the failure of the Mars mission, one scientist said, "This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses until the end of time."



Mars Climate Orbiter during preflight tests.

Source: *Jet Propulsion Laboratory/NASA*

CHECKPOINT – SECTION 1.3 Scientific Measurement

1.3.1 The coldest temperature ever recorded on Earth was −128.6°F (recorded at Vostok Station, Antarctica, on July 21, 1983). Express this temperature in degrees Celsius and in kelvins.

a) −89.2°C, −89.2 K
b) −289.1°C, −15.9 K
c) −89.2°C, 183.9 K
d) −173.9°C, 99.3 K
e) −7.0°C, 266.2 K

1.3.2 What is the density of an object that has a volume of 34.2 cm^3 and a mass of 19.6 g?

a) 0.573 g/cm^3
b) 1.74 g/cm^3
c) 670 g/cm^3
d) 53.8 g/cm^3
e) 14.6 g/cm^3

1.3.3 A sample of water is heated from room temperature to just below the boiling point. The overall change in temperature is 72°C. Express this temperature change in kelvins.

a) 345 K
b) 72 K
c) 0 K
d) 201 K
e) 273 K

1.3.4 Given that the density of gold is 19.3 g/cm^3, calculate the volume (in cm^3) of a gold nugget with a mass of 5.98 g.

a) 3.23 cm^3
b) 5.98 cm^3
c) 115 cm^3
d) 0.310 cm^3
e) 13.3 cm^3

1.4 The Properties of Matter

Substances are identified by their properties as well as by their composition. Properties of a substance may be ***quantitative*** (measured and expressed with a number) or ***qualitative*** (not requiring explicit measurement).

Physical Properties

Color, melting point, boiling point, and physical state are all physical properties. A ***physical property*** is one that can be observed and measured without changing the *identity* of a substance. For example, we can determine the melting point of ice by heating a block of ice and measuring the temperature at which the ice is converted to water. Liquid water differs from ice in appearance but not in composition; both liquid water and ice are H_2O. Melting is a ***physical change:*** one in which the state of matter changes, but the identity of the matter does not change. We can recover the original ice by cooling the water until it freezes. Therefore, the melting point of a substance is a *physical* property. Similarly, when we say that nitrogen dioxide gas is brown, we are referring to the physical property of color.

Chemical Properties

The statement "Hydrogen gas burns in oxygen gas to form water" describes a ***chemical property*** of hydrogen, because to observe this property we must carry out a ***chemical change***—burning in oxygen (combustion), in this case. After a chemical change, the original substance (hydrogen gas in this case) will no longer exist. What remains is a different substance (water, in this case). We *cannot* recover the hydrogen gas from the water by means of a physical process, such as boiling or freezing.

Every time we bake cookies, we bring about a chemical change. When heated, the sodium bicarbonate (baking soda) in cookie dough undergoes a chemical change that produces carbon dioxide gas. The gas forms numerous little bubbles in the dough during the baking process, causing the cookies to "rise." Once the cookies are baked, we cannot recover the sodium bicarbonate by cooling the cookies, or by *any* physical process. When we eat the cookies, we cause further chemical changes that occur during digestion and metabolism.

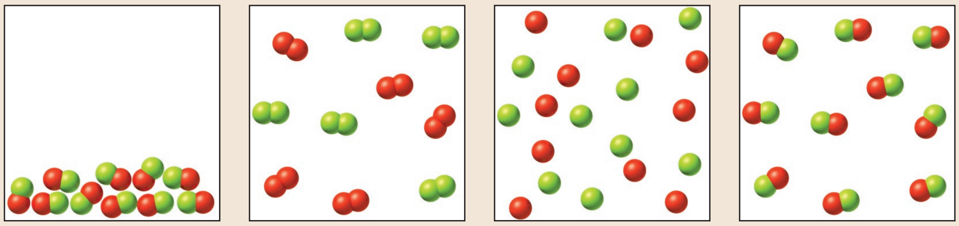
Extensive and Intensive Properties

All properties of matter are either *extensive* or *intensive*. The measured value of an ***extensive property*** depends on the amount of matter. *Mass* is an extensive property. More matter means more mass. Values of the same extensive property can be added together. For example, two gold nuggets will have a combined mass that is the sum of the masses of each nugget, and the length of two city buses is the sum of their individual lengths. The value of an extensive property depends on the amount of matter.

The value of an ***intensive property*** does *not* depend on the amount of matter. *Density* and *temperature* are intensive properties. Suppose that we have two beakers of water at the same temperature and we combine them to make a single quantity of water in a larger beaker. The density and the temperature of the water in the larger combined quantity will be the same as they were in the two separate beakers. Unlike mass and length, which are additive, temperature, density, and other intensive properties are not additive. Sample Problem 1.4 shows you how to differentiate chemical and physical processes.

SAMPLE PROBLEM 1.4

The diagram in (a) shows a compound made up of atoms of two elements (represented by the green and red spheres) in the liquid state. Which of the diagrams in (b) to (d) represent a physical change, and which diagrams represent a chemical change?



(a) (b) (c) (d)

Strategy We review the discussion of physical and chemical changes. A physical change does not change the *identity* of a substance, whereas a chemical change *does* change the identity of a substance.

Setup The diagram in (a) shows a substance that consists of molecules of a compound, each of which contains two different atoms, represented by green and red spheres. Diagram (b) contains the same number of red and green spheres, but they are not arranged the same way as in diagram (a). In (b), each molecule is made up of two identical atoms. These are molecules of *elements,* rather than molecules of a compound. Diagram (c) also contains the same numbers of red and green spheres as diagram (a). In (c), however, all the atoms are shown as isolated spheres. These are atoms of elements, rather than molecules of a compound. In diagram (d), the spheres are arranged in molecules, each containing one red and one green sphere. Although the molecules are farther apart in diagram (d), they are the same molecules as shown in diagram (a).

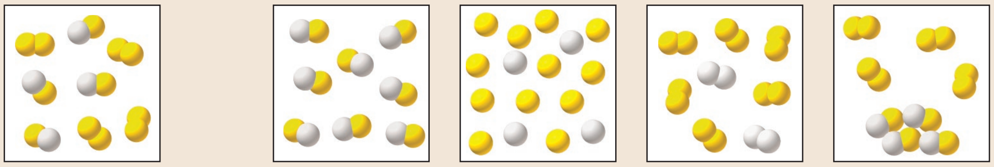
Solution Diagrams (b) and (c) represent chemical changes. Diagram (d) represents a physical change.

THINK ABOUT IT

A chemical change changes the *identity* of matter. A physical change does not.

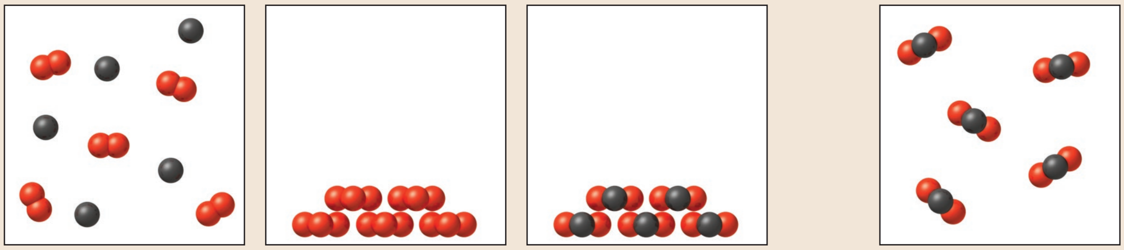
Practice Problem ATTEMPT Which of the following processes is a physical change? (a) evaporation of water; (b) combination of hydrogen and oxygen gas to produce water; (c) dissolution of sugar in water; (d) separation of sodium chloride (table salt) into its constituent elements, sodium and chlorine; (e) combustion of sugar to produce carbon dioxide and water.

Practice Problem BUILD The diagram on the left shows a system prior to a process taking place. Which of the other diagrams [(i) to (iv)] could represent the system after a *physical* process; which could represent the system after a *chemical* process; and which could not represent either?



(i) (ii) (iii) (iv)

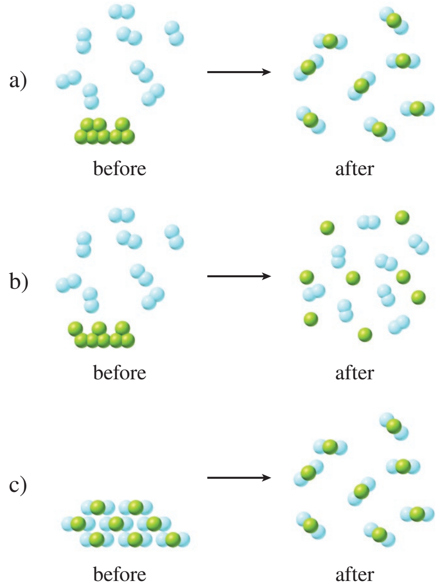
Practice Problem CONCEPTUALIZE The diagram on the right represents the result of a process. Which of the diagrams [(i) to (iii)] could represent the starting material if the process were physical, and which could represent the starting material if the change were chemical?



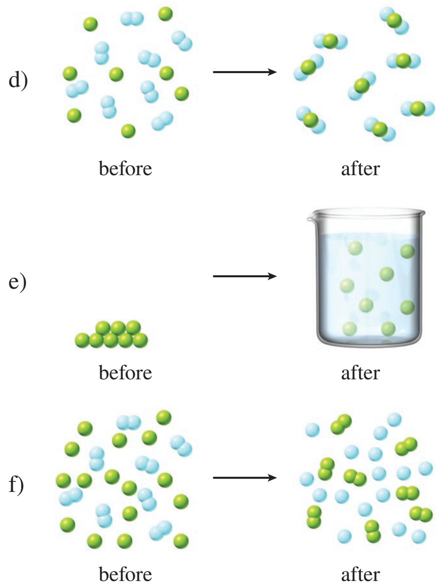
(i) (ii) (iii)

CHECKPOINT – SECTION 1.4 The Properties of Matter

1.4.1 Which of the following [(a)–(f)] represents a physical change? (Select all that apply.)



1.4.2 Which of the following [(a)–(f)] represents a chemical change? (Select all that apply.)



1.5 Uncertainty in Measurement

Chemistry makes use of two types of numbers: exact and inexact. *Exact* numbers include numbers with defined values, such as 2.54 in the definition 1 inch (in) = 2.54 cm, 1000 in the definition 1 kg = 1000 g, and 12 in the definition 1 dozen = 12 objects. (The number 1 in each of these definitions is also an exact number.) Exact numbers also include those that are obtained by counting. Numbers measured by any method other than counting are *inexact*.

Measured numbers are inexact because of the measuring devices that are used, the individuals who use them, or both. For example, a ruler that is poorly calibrated will result in measurements that are in error—no matter how carefully it is used. Another ruler may be calibrated properly but have insufficient resolution for the necessary measurement. Finally, whether or not an instrument is properly calibrated or has sufficient resolution, there are unavoidable differences in how different people see and interpret measurements.

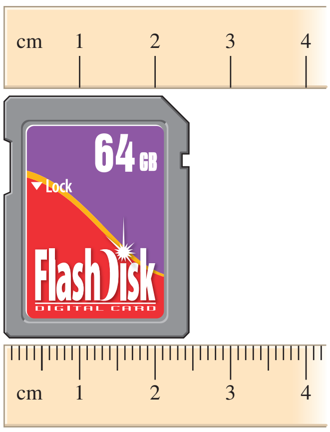


Figure 1.11 The width we report for the memory card depends on which ruler we use to measure it.

Significant Figures

An inexact number must be reported in such a way as to indicate the uncertainty in its value. This is done using significant figures. ***Significant figures*** are the *meaningful digits* in a reported number. Consider the measurement of the memory card in Figure 1.11 using the ruler above it. The card's width is slightly greater than 2 cm. We may record the width as 2.5 cm, but because there are no gradations between 2 and 3 cm on this ruler, we are *estimating* the second digit. Although we are certain about the 2 in 2.5, we are *not* certain about the 5. The last digit in a measured number is referred to as the *uncertain digit;* and the uncertainty associated with a measured number is generally considered to be ± 1 in the place of the last digit. Thus, when we report the width of the memory card to be 2.5 cm, we are implying that its width is 2.5 ± 0.1 cm. Each of the digits in a measured number, including the uncertain digit, is a significant figure. The reported width of the card, 2.5 cm, contains *two* significant figures.

A ruler with millimeter gradations would enable us to be certain about the second digit in this measurement and to estimate a third digit. Now consider the measurement of the memory card using the ruler below it. We may record the width as 2.45 cm. Again, we estimate one digit beyond those we can read. The reported width of 2.45 cm contains *three* significant figures. Reporting the width as 2.45 cm implies that the width is 2.45 ± 0.01 cm.

Student Note: It is important not to imply greater certainty in a measured number than is realistic. For example, it would be inappropriate to report the width of the memory card in Figure 1.11 as 2.4500 cm, because this would imply an uncertainty of ± 0.0001 cm.

The number of significant figures in any number can be determined using the following guidelines:

Always Significant Nonzero digits and the zeros between them:

	137.1	**209.51**	**410.05**	**10.0011**	0.0**36**	0.00**501**
significant figures	4	5	5	6	2	3

Zeros to the *right* of nonzero digits in numbers that contain decimal points:

	8.300	**161.000**	0.**50**	0.0**113**	**309.0**	0.00**52500**
significant figures	4	6	2	3	4	5

Never Significant Zeros to the *left* of leftmost nonzero digit:

	0.00**137**	0.**695**	0.0000**8**	0.0**51050**	0.00**6011**	0.000**90**
significant figures	3	3	1	5	4	2

Sometimes Significant Zeros to the *right* of the rightmost nonzero digit in a number that does *not* contain a decimal point may or may not be considered significant, depending on circumstance. For example, the number 1000 may have anywhere from one to four significant figures. Without additional information, it is not possible to know. To avoid ambiguity in such cases, it is best to express such numbers using scientific notation [⏭ Appendix 1].

	1 $\times 10^3$	**1.0** $\times 10^3$	**1.00** $\times 10^3$	**1.000** $\times 10^3$
significant figures	1	2	3	4

Student Hot Spot

Student data indicate you may struggle with significant figures. Access the eBook to view additional Learning Resources on this topic.

Sample Problem 1.5 lets you practice determining the number of significant figures in a number.

SAMPLE PROBLEM 1.5

Determine the number of significant figures in the following measurements: (a) 443 cm, (b) 15.03 g, (c) 0.0356 kg, (d) 3.000×10^{-7} L, (e) 50 mL, (f) 0.9550 m.

Strategy All nonzero digits are significant, so the goal will be to determine which of the zeros is significant.

Setup Zeros are significant if they appear between nonzero digits or if they appear after a nonzero digit in a number that contains a decimal point. Zeros may or may not be significant if they appear to the right of the last nonzero digit in a number that does not contain a decimal point.

Solution (a) 3; (b) 4; (c) 3; (d) 4; (e) 1 or 2, an ambiguous case; (f) 4.

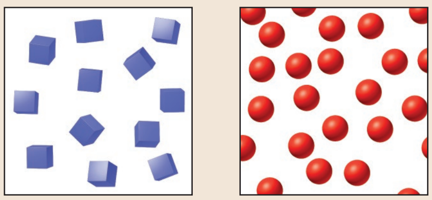
THINK ABOUT IT

Be sure that you have identified zeros correctly as either significant or not significant. They are significant in (b), (d), and (f); they are not significant in (c); and it is not possible to tell in (e).

Practice Problem ATTEMPT Determine the number of significant figures in the following measurements: (a) 1129 m, (b) 0.0003 kg, (c) 1.094 cm, (d) 3.5×10^{12} atoms, (e) 150 mL, (f) 9.550 km.

Practice Problem BUILD For each of the following numbers, determine the number of significant figures it contains, rewrite it without using scientific notation, and determine the number of significant figures in the result. (a) 3.050×10^{-4}, (b) 4.3200×10^{2}, (c) 8.001×10^{-7}, (d) 2.006080×10^{5}, (e) 1.503×10^{-5}, (f) 6.07510×10^{4}.

Practice Problem CONCEPTUALIZE Report the number of colored objects contained within each square and, in each case, indicate the number of significant figures in the number you report.



Calculations with Measured Numbers

Because we often use one or more measured numbers to calculate a desired result, a second set of guidelines specifies how to handle significant figures in calculations.

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than the original number with the smallest number of digits to the right of the decimal point. For example:

$$\begin{array}{rl} 102.50 & \leftarrow \text{two digits after the decimal point} \\ +\ 0.231 & \leftarrow \text{three digits after the decimal point} \\ \hline 102.731 & \leftarrow \text{round to } 102.73 \end{array}$$

$$\begin{array}{rl} 143.29 & \leftarrow \text{two digits after the decimal point} \\ -\ 20.1 & \leftarrow \text{one digit after the decimal point} \\ \hline 123.19 & \leftarrow \text{round to } 123.2 \end{array}$$

   The rounding procedure works as follows. Suppose we want to round 102.13 and 54.86 each to one digit to the right of the decimal point. To begin, we look at the digit(s) that will be dropped. If the leftmost digit to be dropped is less than 5, as in 102.13, we *round down* (to 102.1), meaning that we simply drop the digit(s). If the leftmost digit to be dropped is equal to or greater than 5, as in 54.86, we *round up* (to 54.9), meaning that we add 1 to the preceding digit.

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures. The following examples illustrate this rule:

   $1.4 \times 8.011 = 11.2154$ ← round to 11 (limited by 1.4 to *two* significant figures)

   $\dfrac{11.57}{305.88} = 0.037825290964$ ← round to 0.03783 (limited by 11.57 to *four* significant figures)

3. *Exact numbers* can be considered to have an infinite number of significant figures and do not limit the number of significant figures in a calculated result. For example, a penny minted after 1982 has a mass of 2.5 g. If we have three such pennies, the total mass is

$$3 \times 2.5\ \text{g} = 7.5\ \text{g}$$

   The answer should *not* be rounded to one significant figure because 3 is an *exact* number.

Student Note: Note that it is the number of pennies (3), not the mass, that is an exact number.

4. In calculations with multiple steps, rounding the result of each step can result in "rounding error." Consider the following two-step calculation:

   First step: $A \times B = C$
   Second step: $C \times D = E$

   Suppose that $A = 3.66$, $B = 8.45$, and $D = 2.11$. The value of E depends on whether we round the value of C prior to using it in the second step of the calculation (Method 1) or not (Method 2).

Method 1	Method 2
$C = 3.66 \times 8.45 = 30.9$	$C = 3.66 \times 8.45 = 30.93$
$E = 30.9 \times 2.11 = 65.2$	$E = 30.93 \times 2.11 = 65.3$

   In general, it is best to retain at least one extra digit until the end of a multistep calculation, as shown by Method 2, to minimize rounding error.