

General Chemistry

A Mastery-Oriented Curriculum

Second Edition

John D. Mays



Austin, Texas
2016

©2014, 2016 Novare Science & Math LLC

All rights reserved. Except as noted below, no part of this book may be reproduced or transmitted in any form or by any means, electronic or mechanical, including photocopying, recording, or by information storage and retrieval systems, without the written permission of the publisher, except by a reviewer who may quote brief passages in a review.

All images attributed to others under any of the Wikimedia Commons licenses, such as CC-BY-SA-3.0 and others, may be freely reproduced and distributed under the terms of those licenses.

Scriptural quotations are from The Holy Bible, English Standard Version, copyright ©2001 by Crossway Bibles, a publishing ministry of Good News Publishers. Used by permission. All rights reserved.

Published by Novare Science & Math

novarescienceandmath.com

Printed in the United States of America

ISBN: 978-0-9972845-1-5

Novare Science & Math is an imprint of Novare Science & Math LLC.

Cover design by Nada Orlic, <http://nadaorlic.info/>



For a catalog of titles published by Novare Science & Math, visit novarescienceandmath.com.

Contents

Preface

For Teachers

1. Student Audience and Preparedness	xii
2. Our Emphasis on Mastery, Integration, and Kingdom	xiii
3. Recommendations for Teaching With This Text	xiv
4. Laboratory Work and Lab Reports	xvii

Preface

For Students

xviii

Introduction

What is Chemistry All About?

2

I.1 A Few Major Themes

3

I.1.1 Chemistry Is All About Electrons

3

I.1.2 Chemistry Is All About Electrical Forces

5

Hmm... Interesting. Why water forms beads

6

I.1.3 Chemistry Is All About Minimizing Energy

6

I.1.4 Chemistry Is All About Whole-Number Ratios of Atoms

9

I.1.5 Chemistry Is All About Modeling

10

I.2 Conclusion

12

Chapter 1

Measurements

14

1.1 Science and Measurements

16

1.1.1 No Measurements, No Science

16

1.1.2 Matter, Volume, and Mass

16

1.1.3 The US Customary System

18

1.1.4 The SI Unit System

18

1.1.5 Metric Prefixes

20

1.2 Converting Units of Measure

22

1.2.1 Basic Principles of Unit Conversion Factors

22

1.2.2 Tips for Converting Units of Measure

24

1.2.3 Converting Temperature Units

27

1.3 Accuracy and Precision

28

1.3.1 Distinguishing Between Accuracy and Precision

28

1.3.2 Significant Digits

29

1.4 Other Important Math Skills

35

1.4.1 Scientific Notation

35

1.4.2 Calculating Percent Difference

37

Chapter 2

Atoms and Substances

42

2.1 Atoms and Molecules

44

2.1.1 Atomic Facts

44

2.1.2 The History of Atomic Models

45

2.2	Types of Substances	49
2.2.1	Pure Substances: Elements and Compounds	49
2.2.2	Mixtures	53
	<i>Hmm... Interesting. Brownian motion</i>	56
2.2.3	Physical and Chemical Properties	57
2.3	Isotopes and Atomic Masses	59
2.3.1	Isotopes	59
2.3.2	The Unified Atomic Mass Unit	60
2.3.3	Atomic Masses	60
2.4	Density and Quantity of Substances	62
2.4.1	Density	62
2.4.2	The Mole and the Avogadro Constant	64
2.4.3	Molar Mass and Formula Mass	65
2.4.4	Gram Masses of Atoms and Molecules	69

Chapter 3

	Atomic Structure	74
3.1	Atomic Spectra	76
3.1.1	The Electromagnetic Spectrum	76
3.1.2	Energy in Atoms	78
	<i>Hmm... Interesting. Neon signs and phonons</i>	79
3.1.3	The Hydrogen Atom	81
3.2	The Bohr Model of the Atom	82
3.3	The Quantum Model of the Atom	83
3.3.1	Schrödinger and Pauli	83
3.3.2	Shells, Subshells, and Orbitals	83
3.3.3	The Aufbau Principle, the Madelung Rule, and Hund's Rule	88
3.4	Electron Configurations	90
3.4.1	Electron Configurations	90
3.4.2	Condensed Electron Configurations	92
3.5	Empirical Formulas	93
3.5.1	Percent Composition and Empirical Formulas	93
3.5.2	Determining a Molecular Formula from an Empirical Formula	95

Chapter 4

	The Periodic Law	100
4.1	The Periodic Table of the Elements	102
4.2	Periodic Table Nomenclature	104
4.3	Periodic Physical Properties	105
4.3.1	Atomic Radius	105
4.3.2	Ionic Radius	107
4.4	Periodic Chemical Properties	108
4.4.1	Core and Valence Electrons	108
4.4.2	Ionization Energy	109
4.4.3	Electron Affinity	113
4.4.4	Electronegativity	115

4.5	A Few Notes About Hydrogen	116
	<i>Hmm... Interesting. Hydrogen in space</i>	117

Chapter 5

	Chemical Bonding	120
5.1	Preliminaries	122
	5.1.1 Chemical Possibilities	122
	5.1.2 The Octet Rule	122
5.2	Ionic Bonding	123
	5.2.1 Ionic Bonds and Crystals	123
	5.2.2 Naming Ionic Compounds	125
	5.2.3 Energy in Ionic Bonds	127
	5.2.4 Hydrates	129
	5.2.5 Intensive and Extensive Properties	129
	5.2.6 Physical Properties of Ionically Bonded Substances	130
5.3	Covalent Bonding	131
	5.3.1 Covalent Bonds and Molecules	131
	5.3.2 Polyatomic Ions	133
	5.3.3 Ionic Compounds with Polyatomic Ions	134
	5.3.4 Polyatomic Ion Names	134
	5.3.5 Naming Acids	135
	5.3.6 Lewis Structures	136
	5.3.7 Exceptions to the Octet Rule	140
	5.3.8 Resonance Structures	141
	5.3.9 Naming Binary Covalent Compounds	142
	5.3.10 Energy in Covalent Bonds	143
	5.3.11 Physical Properties of Covalently Bonded Substances	144
5.4	Electronegativity, Polarity, and Bond Character	145
	5.4.1 Polarity and Dipoles	145
	5.4.2 The Nature of the Bond	146
	<i>Hmm... Interesting. The molecular structure of glass and quartz</i>	148

Chapter 6

	Molecular Theory and Metallic Bonding	150
6.1	Molecular Structure	152
	6.1.1 Covalent Bond Theory	152
	6.1.2 Valence Shell Electron Pair Repulsion (VSEPR) Theory	152
	6.1.3 The Effect of Nonbonding Domains on Bond Angle	156
6.2	Metallic Bonding	157
	6.2.1 Metallic Lattices	157
	6.2.2 Physical Properties of Metals	158
6.3	Intermolecular Forces	160
	6.3.1 Bonding Forces	160
	6.3.2 Intermolecular Forces	160
	<i>Hmm... Interesting. Tin pest</i>	160
	6.3.3 Hydrogen Bonding	161
	6.3.4 Van der Waals Forces	162

Chapter 7

Chemical Reactions and Stoichiometry	168
7.1 Introduction to Chemical Equations	170
7.1.1 Fascinating Chemistry	170
7.1.2 The Law of Conservation of Mass in Chemical Reactions	170
7.1.3 Reaction Notation	171
7.1.4 Balancing Chemical Equations	172
7.1.5 Oxidation States	177
<i>Hmm... Interesting. Why nitrates and nitros blow up</i>	179
7.2 General Types of Chemical Reactions	180
7.2.1 Synthesis Reactions	180
7.2.2 Decomposition Reactions	181
7.2.3 The Activity Series of Metals	181
7.2.4 Single Replacement Reactions	182
7.2.5 Double Replacement Reactions	184
<i>Hmm... Interesting. A story about aqua regia</i>	184
7.2.6 Combustion Reactions	185
7.2.7 Acid-Base Neutralization Reactions	185
7.2.8 Oxidation-Reduction Reactions	186
7.3 Stoichiometry	187
7.3.1 Stoichiometric Calculations	187
7.3.2 Limiting Reactant	192
7.3.3 Theoretical Yield and Percent Yield	194

Chapter 8

Kinetic Theory and States of Matter	200
8.1 Temperature, Kinetic-Molecular Theory, and Pressure	202
8.1.1 Temperature and Molecular Energy	202
8.1.2 Velocity Distribution of Gases	202
8.1.3 The Kinetic-Molecular Theory of Gases	203
8.1.4 Gas Pressure	204
<i>Hmm... Interesting. How barometers work</i>	206
8.2 States of Matter	208
8.2.1 The Four Basic States of Matter	208
8.2.2 Solids	208
8.2.3 Liquids	209
8.2.4 Gases	211
8.2.5 Plasmas	212
<i>Hmm... Interesting. Gas diffusion</i>	212
8.2.6 Phase Transitions and Phase Diagrams	213
8.2.7 Heat Capacity, Heat of Fusion, and Heat of Vaporization	216
8.2.8 Evaporation	220
8.2.9 Vapor Pressure	222

Chapter 9

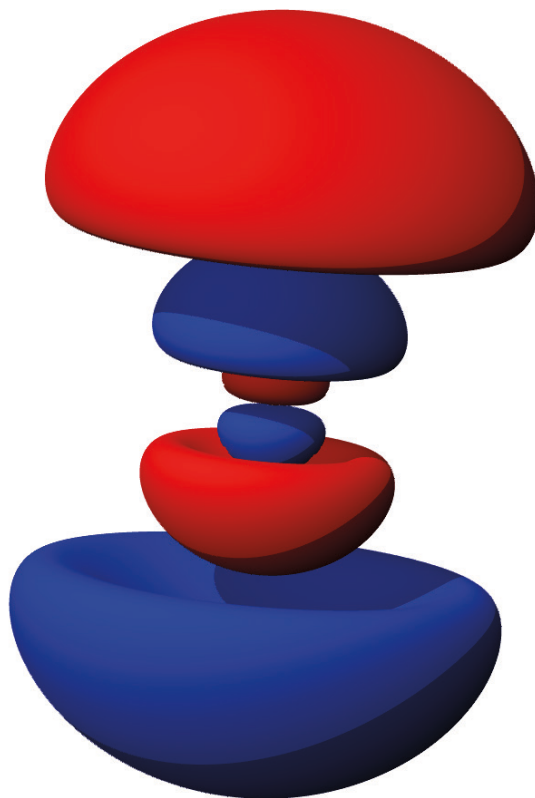
The Gas Laws	226
9.1 Early Formulations of the Gas Laws	228
9.1.1 Boyle's Law	228
9.1.2 Charles' Law	229

9.1.3 Avogadro's Law	232
9.2 The Ideal Gas Law	233
9.2.1 Standard Temperature and Pressure	233
9.2.2 The Ideal Gas Law	233
<i>Hmm... Interesting. The gas laws as models</i>	234
9.2.3 Using the Ideal Gas Law to Find Molar Mass and Density	240
9.3 The Law of Partial Pressures	243
9.3.1 Dalton's Law of Partial Pressures	243
9.3.2 Collecting a Gas Over Water	246
9.4 Stoichiometry of Gases and Effusion	248
9.4.1 Stoichiometry of Gases	248
9.4.2 Gas Diffusion and Effusion	249
<i>Hmm... Interesting. Uranium enrichment</i>	250
Chapter 10	
Solution Chemistry	256
10.1 Dissolution	258
10.1.1 The Process of Dissolving	258
10.1.2 Heat of Solution	261
10.1.3 Entropy and Free Energy	261
10.1.4 Electrolytes	263
10.2 Solubility	264
10.2.1 Ionic Solids in Water	264
10.2.2 Ionic Solids in Nonpolar Solvents	265
10.2.3 Polar Liquids	266
10.2.4 Nonpolar Liquids	267
10.2.5 Solutions of Solids	267
<i>Hmm... Interesting. How soap works</i>	268
10.2.6 Gases in Liquid Solutions	270
10.2.7 The Effect of Temperature on Solubility	271
10.3 Quantifying Solution Concentration	272
10.3.1 Molarity	272
10.3.2 Molality	274
10.4 Compounds in Aqueous Solution	275
10.4.1 Ionic Equations and Precipitates	275
10.4.2 Net Ionic Equations and Spectator Ions	277
10.5 Colligative Properties of Solutions	278
10.5.1 Vapor Pressure Lowering	278
10.5.2 Freezing Point Depression and Boiling Point Elevation	280
Chapter 11	
Acids and Bases	290
11.1 Properties and Nomenclature of Acids and Bases	292
11.1.1 Introduction	292
11.1.2 Properties of Acids and Bases	292
11.1.3 Acid Names and Formulas	294
11.2 Acid-Base Theories	295
11.2.1 Arrhenius Acids and Bases	296

11.2.2 Brønsted-Lowry Acids and Bases	298
<i>Hmm... Interesting. What is an alkali?</i>	299
11.2.3 Lewis Acids and Bases	302
11.2.4 Strength of Acids and Bases	303
11.3 Aqueous Solutions and pH	306
11.3.1 The Self-ionization of Water	306
11.3.2 Calculating $[H_3O^+]$ and $[OH^-]$	306
11.3.3 pH as a Measure of Ion Concentration and Acidity	307
11.3.4 pH Measurement, pH Indicators, and Titration	312
11.3.5 Titration Procedure	316
11.3.6 Determining $[H_3O^+]$ or $[OH^-]$ from Titration Data	318
Chapter 12	
Redox Chemistry	324
12.1 Oxidation and Reduction	326
12.1.1 Introduction to Redox Reactions	326
12.1.2 Oxidation States	326
12.1.3 Strengths of Oxidizing and Reducing Agents	330
12.2 Redox Reaction Equations	332
12.2.1 Redox Half-Reactions	332
12.2.2 Balancing Redox Equations	336
12.3 Electrochemistry	342
12.3.1 Copper and Zinc Redox	342
12.3.2 Electricity Instead of Heat	343
12.3.3 Electrochemical Cells	344
<i>Hmm... Interesting. How are salt bridges made?</i>	348
12.3.4 Electrode Potentials	350
12.3.5 Electrochemical Applications	353
Glossary	362
Answers to Selected Exercises	378
Appendix A	
Reference Data	390
Appendix B	
Scientists to Know About	394
References and Citations	396
Image Credits	398
Index	400

Introduction

What is Chemistry All About?



This computer model depicts three of the *orbitals* available in atoms for holding electrons. Shown are one of the three *p* orbitals in each of three different *subshells*. The inner pair of orbitals can hold two of the highest energy electrons for elements 5 through 10 in the Periodic Table of the Elements. The middle pair is available to hold two of the highest energy electrons for elements 13 through 18 in the periodic table, and the outer pair can hold two of the highest energy electrons belonging to elements 31 through 36 in the periodic table.

In this Introduction, we touch briefly on electron orbitals. We treat the subject in more depth in Chapter 3.

Objectives for the Introduction

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

1. Briefly explain how electrons, electrical forces, minimizing energy, whole number ratios, and modeling can each be thought of as central to understanding what chemistry is all about.
2. State and explain examples illustrating a system moving to a lower energy state and a system experiencing an increase in entropy.
3. Briefly explain hydrogen bonding and why it plays such a large role in mixtures containing water.
4. Explain the relationship between energy and atomic orbitals.

I.1 A Few Major Themes

Chemistry is the study of the elements, how they combine to form mixtures and compounds, the properties of these substances, and the processes involved. One of the astonishing things about the physical world is that as complex as the details are, we can understand a lot about how it works in terms of just a few basic principles from physics. This striking situation is a direct result of the fact that nature is governed by an orderly, mathematical set of physical laws—the laws set in place by God according to his wisdom when he created the universe.

The existence of creation and of the laws of physics are two obvious clues to God's role in creating the universe: the universe is here because God made it, and it is governed in an orderly, mathematical way because it was God's pleasure to make it so. A third clue is that *we can understand it*.

Studying chemistry involves learning a great deal of terminology, and exploring quite a few different types of processes. The amount of information involved can be daunting! But one way to help organize all this information is to be alert to a few fundamental principles that turn up time and again. In this introductory chapter, we take a brief look at a few of these principles. As you read through the chapters ahead, you will see again and again that we can understand a lot about topics such as molecular structure, solubility, and chemical reactions in terms of a few basic concepts.

I.1.1 Chemistry Is All About Electrons

You probably already know that atoms consist of a tiny nucleus containing particles called protons and neutrons, and that the nucleus is surrounded by cloud-like regions containing the atoms' electrons. The protons and electrons carry electrical charge—protons are positively charged and electrons are negatively charged. The protons in an atom stay permanently in the atom's nucleus,¹ but atoms lose or gain electrons by interacting with other atoms. An atom that gains or loses one or more electrons is called an *ion*. Ions are charged particles. Gaining an electron means gaining negative charge, resulting in a negatively charged ion. Losing electrons means losing negative charge, and ending up with more protons (positive charges) than electrons. This results in a positively charged ion.

As it turns out, a lot of chemistry can be understood in terms of the atoms' electrons—where they are, how many there are, whether an atom has ionized by gaining or losing electrons, whether an atom is sharing electrons with another atom, and so on.

¹ Except in the case of radioactive elements.

The cloud-like regions containing an atom's electrons are called *orbitals*, and electrons reside in different orbitals according to how much energy they have. The arrangement of the orbitals is the same for all atoms, although the specific energies associated with each orbital vary from atom to atom, depending on the size of the nucleus and how many electrons an atom has. The orbitals in atoms are grouped into different energy groupings called *shells*. There are seven main shells containing the orbitals with the electrons of all the elements discovered so far. There are additional shells above these that high-energy electrons can move into when they absorb more energy.

In each shell, there is a specific number of orbitals, and each orbital and set of orbitals holds a specific number of electrons. One of the essential facts about atomic behavior is that atoms seek to gain, lose, or share electrons until they have just the right number of electrons so that they

have only full shells, without any extra electrons and without electrons missing from any orbitals in the full shells. If only the first shell is full, an atom has two electrons. If the first two shells are full, 10 electrons. If the first three are full, 28 electrons, and so on. Significantly, these numbers relate to the numbers of elements in the rows of the Periodic Table of the Elements. With this one fact, we can understand a great deal about how atoms of one element bond with atoms of other elements to form compounds.

The position of the electrons within an atom also has a lot to do with how an atom behaves. One aspect of atoms that affects the position of electrons is the shapes of the different orbitals. Some orbitals are spherically shaped, some are shaped in pairs of protruding lobes often described as “dumbbells,” and some are shaped as rings. There are other more complex shapes as well. Since all electrons repel each other due to their negative electrical charge, electrons located in the lobes of dumbbell-shaped orbitals push away from each other, resulting in molecules with very particular shapes. Examples are the water, ammonia, and methane molecules illustrated in Figure I.1.

Electron position is also affected by the fact that within molecules some atoms attract electrons more strongly than others, an effect denoted by a value called the *electronegativity* of the atom. We discuss this in more detail later, but I will mention an important example here to illustrate this point. The electronegativity values for oxygen and hydrogen are 3.44 and 2.20, respectively. This means the oxygen atoms in water molecules attract electrons more strongly than the hydrogen atoms do. As a result, the four bonding electrons in the molecule crowd over toward the oxygen atom, making the oxygen region of the water molecule more electrically negative and the hydrogen regions more electrically positive. These differences make the water molecule electrically imbalanced—or *polar*, as we say—negative on one side and positive at the ends on the other side, as illustrated in Figure I.2. In this diagram, the arrows point from the positive regions of the molecule toward the negative region of the molecule.

The shapes and polarizations affect atomic behavior because of electrical attractions and repulsions, the basic theme we discuss in the next section.

The shapes and polarizations affect atomic behavior because of electrical attractions and repulsions, the basic theme we discuss in the next section.

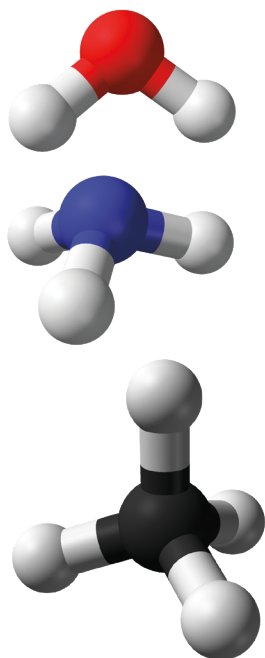


Figure I.1. Representations of the H_2O water molecule (top), the NH_3 ammonia molecule (middle), and the CH_4 methane molecule (bottom).

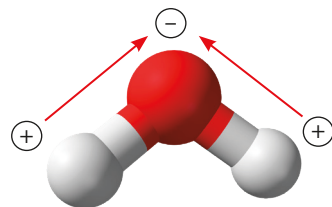


Figure I.2. The higher electronegativity of oxygen atoms compared to hydrogen atoms results in the polar water molecule.

1.1.2 Chemistry Is All About Electrical Forces

Consider again the two types of electrical charge, the positive protons and the negative electrons. You know that like charges repel each other (such as two positive charges) and opposite charges attract each other (positive and negative).

These electrical attractions and repulsions are highly important for chemistry because atoms and molecules are as prickly as porcupines with charges that repel or attract other charges. Some of these attractions and repulsions are stable and long-lasting, like the attractions between ions that hold together the atoms in a crystal of table salt. Other electrical interactions are sort of semi-stable, you might say, given the fact that molecules are moving around all the time. The world-class example of this is *hydrogen bonding*, which we examine in detail later. Hydrogen bonding takes us back to the polar water molecule described in the previous section. Since water molecules are polar, the positive regions of one water molecule are attracted toward the negative regions of other water molecules, as illustrated in Figure I.3.

The importance of hydrogen bonding cannot be overstated. Water is everywhere, and thus so is hydrogen bonding. Hydrogen bonding explains why so many things dissolve in water. It explains why water travels upwards against the force of gravity when soaking into the fibers of a towel. And it explains why water gets less dense right before it freezes (which in turn explains why ice floats). Figure I.4 is a model of how the water molecules are arranged in ice. The dashed lines in the figure indicate the hydrogen bonds between water molecules. The result of these bonds is the three-dimensional, hexagonal structure of ice.

There are several other ways electrical forces between atoms and molecules are made manifest. In general, these different attractions and repulsions are called *intermolecular forces*. There is an electron cloud around every atom (except in the case of a hydrogen atom that has lost its only electron due to ionization). There is also an electron cloud around and between the atoms of every molecule. As the electrons swarm around in these clouds, there are moments when some regions in the molecule are more negatively charged because of electrons crowding together. There are other moments when regions are more positively charged because electrons have temporarily moved away and the positive charge on the protons in atomic nuclei are dominant in the area. These electron movements and crowding go on all the time and at extremely high speeds, giving rise to ever-changing patterns of intermolecular forces. Intermolecular forces are all caused by electrical attractions.

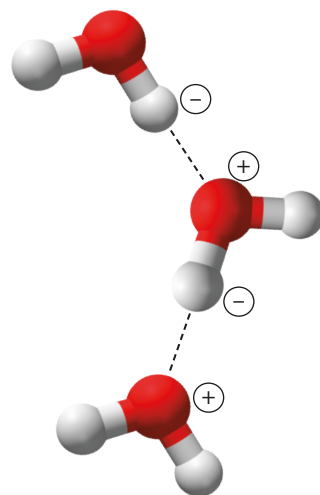


Figure I.3. Hydrogen bonding in water molecules. Dashed lines indicate hydrogen bonds between water molecules.

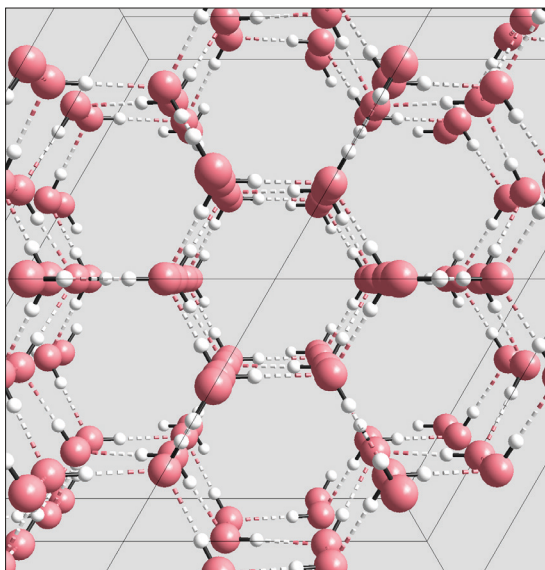
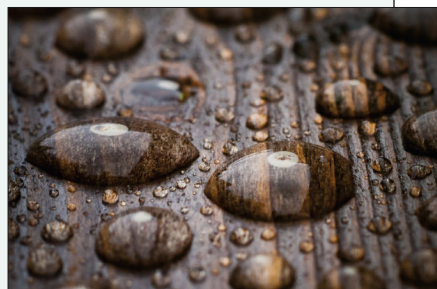
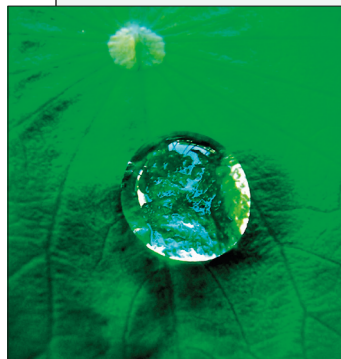


Figure I.4. The crystal structure in ordinary water ice. In this model, oxygen atoms are red and hydrogen atoms are white. Hydrogen bonds are shown as dashed lines. Thick black lines indicate the bonding of hydrogen to oxygen inside individual water molecules.

Hmm... Interesting.**Why water forms beads**

As described in this chapter, the polarity of water molecules makes them cling to one another. When nonpolar molecules are in contact with water, the water molecules are attracted to each other but not to the nonpolar molecules. The molecules in waxy leaves and oil-based wood finishing products are nonpolar. When water molecules rest on a surface of nonpolar molecules, they cling to each other but not to the surface, and the result is the formation of water drops. Small drops are nearly spherical because this shape minimizes the energy between the molecules. Larger drops flatten out due to their greater weight.



Nonpolar molecules do not dissolve in water. The attractions between the polar water molecules squeeze out the nonpolar molecules, causing the two substances to separate. This is why oil and vinegar separate—oil molecules are nonpolar and vinegar is mostly water.

1.1.3 Chemistry Is All About Minimizing Energy

One of the primary drivers causing atoms to do what they do is the natural tendency of all things to minimize the energy associated with the state they are in. Minimizing energy is a concept that explains a great deal of chemical behavior. Here, we look at several examples of objects in different energy states. Then we apply the concept of minimizing energy to phenomena we see occurring in chemistry.

To begin, in a previous science course you may have studied different forms of potential energy. For example, gravitational potential energy is the energy an object has after being lifted up in a gravitational field. Figure I.5 shows a ball up on the side of the hill. The ball is trapped in a small valley or depression. The ball is located up above the ground, so it has gravitational potential energy. The ball always acts to reduce its potential energy if given a chance. If a tunnel opens up to a lower energy state, the ball goes there, releasing potential energy into some other form of energy (such as kinetic energy) as it goes. Another way for the ball to release potential energy and move to a lower state is for someone to hit it or kick it so that it has enough kinetic energy to get over the small hill where it is trapped. The point is that given the chance, the ball releases potential energy and moves to a lower energy state.

Another example of this idea is shown in Figure I.6. A cone held on its point has potential energy that is released if the cone is released and allowed to fall. In this case, the cone doesn't even need any kind of push or kick; it spontaneously moves to the lower energy state (laying down on its side) if released.

As a third example, consider the act of stretching a rubber band. To stretch out a rubber band, you have to supply energy.

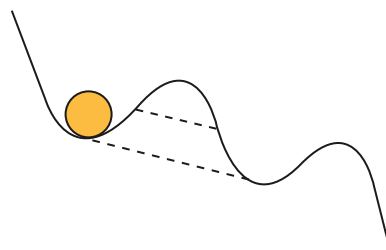
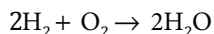


Figure I.5. The ball is trapped in the lowest-energy region in its vicinity. However, if a tunnel to a lower-energy region opens up, the ball goes there.

That is, you have to do mechanical work on the rubber band. If you release the stretched rubber band, it spontaneously contracts back to its lower energy (unstretched) state.

Let's now apply the idea of minimizing energy to a chemical reaction you may already be familiar with: the combustion of hydrogen to produce water. This reaction is represented by the following chemical equation:



The left side of this equation indicates that the *reactants*—the substances taking part in the reaction—are molecules of hydrogen (H_2) and oxygen (O_2). Each of these gases exists as *diatomic* molecules, meaning that each hydrogen molecule is a pair of hydrogen atoms bonded together, and each oxygen molecule is a pair of oxygen atoms bonded together, as illustrated in Figure I.7. At room temperature, these gas molecules zoom around inside their container, colliding with one another several billion times per second, but otherwise nothing else happens.

In terms of the energy of these molecules, they are in a situation similar to the ball in Figure I.5: there are lower energy states the molecules can go to, releasing energy in the process, but they can't get there without a boost of energy to get the process started. Now, if a spark or flame is introduced to this gas mixture, the heat from the spark or flame excites the nearby molecules, causing them to move much faster and slam into each other with enough energy to break the bonds holding the molecules together. The result—which only lasts for a tiny fraction of a second—is a soup of unbonded gas atoms.

At this point, we have a situation similar to the stretched rubber band the instant after being released, before it has had a chance to shrink. Electrical attractions between the protons and electrons in the isolated atoms of oxygen and hydrogen draw the atoms toward each other at an extremely high rate. Consider the collapse of the rubber band after it is released. It collapses to its unstretched state—a lower energy state—releasing energy in the process. The energy released might result in a snap (kinetic energy) that stings your hand and a sound wave (kinetic energy in moving air molecules) producing a snapping sound. Just as the relaxed rubber band is at a lower energy state and releases energy to get there, the hydrogen and oxygen atoms collapse together to the lowest energy state they can find, which is to form water molecules (H_2O). As they do so, they release a lot of energy in the form of light and heat and all this happens in an instant.

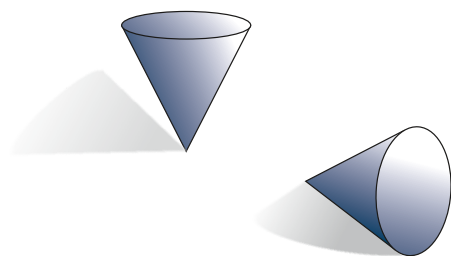


Figure I.6. The cone on the left is in a higher energy state. If released, it falls to the lower energy state shown on the right.

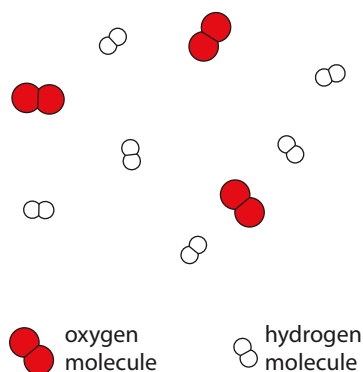


Figure I.7. Diatomic oxygen and hydrogen gas molecules.



Figure I.8. Energy released as heat and light as hydrogen and oxygen atoms combine to form water molecules.

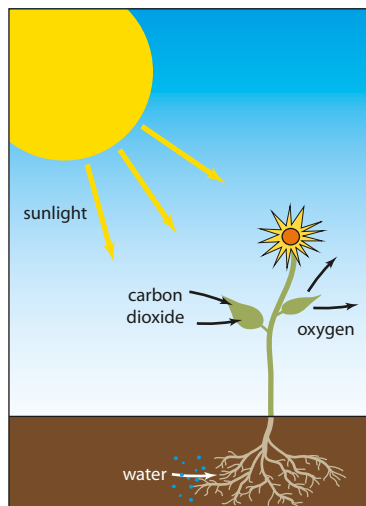
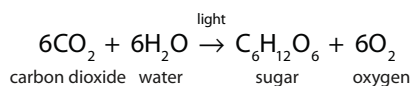


Figure I.9. The photosynthesis reaction is endothermic, as light from the sun is used by plants to convert water and carbon dioxide into sugar and oxygen.

This is the explosion of hydrogen, captured in the photograph of Figure I.8.

This release of heat indicates the reaction is *exothermic*—the reaction releases energy. When considering the way energy relates to various chemical processes, I have found it very helpful to remember the rubber band and to compare it in my mind to the way positive and negative ions are attracted to each other. If separated positive and negative ions are released and allowed to fly together, energy is released—the light and heat of the exothermic reaction—as the ions move to a lower energy state. To separate them, one has to pull them apart by putting in energy (doing work on them) and thus moving them to a higher energy state—just like stretching the rubber band. This is an *endothermic* process, where energy is being absorbed by the ions. The most well-known example of an endothermic chemical reaction is the photosynthesis reaction that occurs in plants, depicted in Figure I.9.

And I can't help pointing out in passing the exquisite elegance of the photosynthesis reaction, a process both simple and incredibly complex that happens automatically and continuously all over the world. Consider the care with which God placed oxygen-breathing creatures like ourselves on a planet covered with oxygen-producing vegetation.

Of course, every school kid learns about photosynthesis, but do we also learn that the delicate balance displayed everywhere in the environment around us in creation is a tremendous gift? I encourage you, as a young student made in God's image, to give thanks and worship to our loving Creator for this most wonderful gift!

There are two more important concepts about the role of energy in chemistry to note here. The first involves a quantity called *entropy*. Entropy is a term that originated in the field of thermodynamics. Entropy is a measure of the *disorder* present in a system, and the second law of thermodynamics states that left to themselves, physical processes go in a direction that increases the entropy (disorder) in the system. As an example, consider a glass of water you may be holding in your hand versus a broken glass on the floor with water splashed everywhere. While the glass is intact with the water contained in it, the system of glass and water is in an orderly state. When you release the glass, disorder increases—the glass breaks and the water goes everywhere on the floor. If you leave the mess like this, the disorder continues to increase: the water evaporates and the water molecules are not even together any more at all. Instead, they are randomly distributed around in the atmosphere. And with time, the chunks of glass get trampled and broken more and more until the remnants of the glass are completely gone. You will never see this process occur in reverse!

For some chemical processes, the minimizing of energy and the increase of entropy both pull in the same direction. In other cases, they try to pull the system in opposite directions. This sets up a sort of tug of war, and the process goes in the most favorable direction. We discuss this in more detail later.

Finally, some detail is in order regarding the boundaries surrounding energy minimization. Try this little thought experiment: imagine a hydrogen ion, which is simply a proton with its positive charge. Nearby is a negatively charged electron, as illustrated in Figure I.10. Since these particles have opposite charges, they are strongly attracted to each other, and since the

proton's mass is 1,836 times greater than the electron's mass, the electron dashes toward the proton while the proton essentially stays put, waiting for the electron to arrive. You might expect that the electron would crash right into the proton, bringing the potential energy between them right down to zero. But this is not what happens.

In 1905, Albert Einstein theorized that energy is *quantized*—it comes in discrete chunks or packets. Since 1905, a host of scientists have explored the quantization of energy, confirming Einstein's proposal over and over and giving birth to the now well-developed theory of quantum mechanics. What quantum mechanics suggests for our proton-electron scenario is that an electron in an orbital of an atom cannot possess just any old amount of energy; it can only possess particular values of energy. In the context of dropping into one of the orbitals surrounding the proton, the electron can only possess an amount of energy corresponding to the one of the energies of the proton's orbitals. The bottom line is that instead of crashing into the proton and sticking to it like cat hair sticking to your pants, the electron instead pops into the lowest energy orbital available around the proton and stays there, captive, buzzing around furiously like a bee in a bottle. (But though this analogy may be suggestive, it is strictly metaphorical. Electrons are not at all like bees. For one thing, they don't have wings. And they don't make honey, either.)



Figure I.10. Oppositely charged particles strongly attracted to each other.

I.1.4 Chemistry Is All About Whole-Number Ratios of Atoms

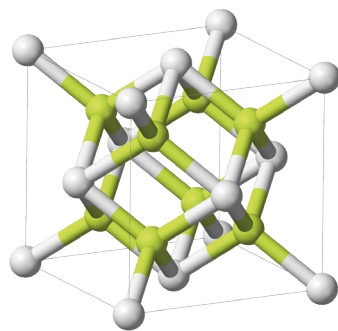
It is strange to think that even as recently as the beginning of the 20th century there was no consensus among scientists as to whether atoms even existed. In 1803, English scientist John Dalton put forward the first detailed, scientific atomic theory. Dalton proposed that all material substances are composed of atoms, and that the way different compounds are formed is by atoms combining together. Since various substances are composed of discrete, individual particles and not just a continuum of matter, there is always a whole number of each type of atom in the substance.

Although many scientists throughout the 19th century refused to accept the existence of atoms, we now agree that Dalton was correct. Compounds do form with whole-number ratios of the atoms involved. (Back then, those who accepted the existence of atoms were called “atomists.” Today, everyone is an atomist, so we don't need a name for this view any more.) As an example, sulfuric acid, H_2SO_4 , has two hydrogen atoms, one sulfur atom, and four oxygen atoms in every molecule, so the ratio of oxygen atoms to hydrogen atoms in the molecule is 2 to 1. The

Figure I.11. The ratio of fluorine atoms (greenish-yellow) to calcium atoms (white) in fluorite is 2 to 1. To see this, note that there are 8



fluorine atoms completely within this crystalline cell. Each of the 8 calcium atoms at the corners is shared by 8 cells—the one shown and 7 other surrounding cells. Eight calcium atoms each shared 8 ways contributes a net of 1 calcium atom to the cell. Then there are 6 calcium atoms on the faces of the cell, each shared by the cell shown and the adjacent cell. Six atoms each shared 2 ways contributes a net of 3 atoms to the cell. In total then, this cell claims 8 fluorine atoms and 4 calcium atoms, a ratio of 2 to 1.



ratio of oxygen atoms to sulfur atoms is 4 to 1. Of course, there are strange exceptions to every rule, including this one. Nevertheless, it is correct to say that just about every compound, regardless of how the atoms are structured, consists of atoms of different elements joined together in predictable whole-number ratios. Figure I.11 contains a photo of the mineral fluorite along with a computer model of the crystal structure of fluorite, or calcium fluoride, which has the formula CaF_2 . In this crystal structure, the ratio of fluorine atoms to calcium atoms is 2 to 1. The caption in the figure explains this, using the computer model of the crystal structure as an aid.

The fact that atoms combine in whole-number ratios is a powerful computational tool. When we get into the math behind chemical reactions (*stoichiometry*, as it is called), we will appeal often to the whole-number ratios of atoms involved in order to compute how much of one compound reacts with a given quantity of another compound.

I.1.5 Chemistry Is All About Modeling

Chemical reactions are happening around us all the time. Just pour a can of soft drink into a glass and watch the carbonic acid (H_2CO_3) in the can convert into the carbon dioxide bubbles and water (CO_2 and H_2O). Light up the gas grill and watch propane (C_3H_8) reacting with the oxygen (O_2) in the air to produce carbon dioxide (CO_2) and water (H_2O). Heat up a pan of cake batter in an oven and the rather complicated molecules in the batter react and change into different complicated molecules in a cake.

In these examples, even though we see bulk materials going into a chemical reaction (the *reactants*) and resulting from the chemical reaction (the *products*), we are not able to see the actual atoms and molecules as they zoom around, combining with and separating from each other. Understanding the behavior of things we cannot see is tricky business.

As mentioned in the previous section, even in recent scientific history the existence of atoms was debated for a hundred years. The issue was finally resolved with experiments in the early 20th century that gave more and more support to the theory that material substances were composed of atoms. We certainly know a lot more today about atoms and their internal structure than we did just a few decades ago. We can even put this knowledge to use in designing amazing new engineering materials, specialty drugs, and even chemical delivery systems to get the drugs into our bodies. But there remains much we do not understand about atoms.

It is helpful to think of science as the process of building “mental models” of the natural world. These mental models are called *theories*. The information we use to build our mental models—scientific facts—comes from experiments, observations, and inferences from these.

Since chemistry deals so much with atoms and molecules, which we can't see, we are almost completely dependent on inferences to develop atomic models describing how the atomic world works. Knowing that the gunpowder in a firecracker explodes when ignited doesn't require a model. It is obvious to all of us that gunpowder is explosive. But *why* is it explosive? What are the rules governing how the atoms in those compounds behave? Understanding why gunpowder explodes does require a model. And the models we work with in chemistry come at us from two different directions.

First, there is the information we gather from experiments. Chemical experimentation has been going on for hundreds of years. In the early days of the scientific revolution, scientists were amazed to discover quantitative laws such as Dalton's whole number ratios and the inverse relationship between the pressure and volume of a gas, a relationship known as Boyle's Law. Second, there is the theoretical modeling that occurs when scientists attempt to apply physical principles from quantum mechanics, thermodynamics, and statistical mechanics to the solution of chemical problems. The shapes and sizes of the atomic orbitals, which we address in Chapter 3, are an example of this type of theoretical modeling.

The theoretical models developed by scientists are the basis for our entire understanding of how the natural world functions. Successful theories are those that account for the facts we know and lead to new hypotheses (predictions) that can be put to the test. It is helpful to think about the relationship between facts, theories, hypotheses, and experiments as illustrated in Figure I.12. This diagram illustrates what I call the *Cycle of Scientific Enterprise*. It is important for every student to develop a correct understanding of the kind of knowledge scientific study provides for us. The *goal* of science is to uncover the truth about how nature works, but scientific theories are always works in progress. Even our best theories are provisional and subject to change. For this reason, science is not in the business of making truth claims about scientific knowledge. Science is in the business of modeling how nature works with theories based on research.

As our theories develop over time, our hope is that they get closer and closer to the truth—the amazing and profound truth about mysteries such as what protons and electrons are, why they have the properties they have, and how the two most successful theories of the 20th century—quantum mechanics and general relativity—can be reconciled with each other. But the truth about nature is always out in front of us somewhere, always outside our grasp. To know the truth about nature, we would have to understand nature as God understands it. We are nowhere close to that.

Here are some definitions to keep in mind as you consider the models we discuss in future chapters.

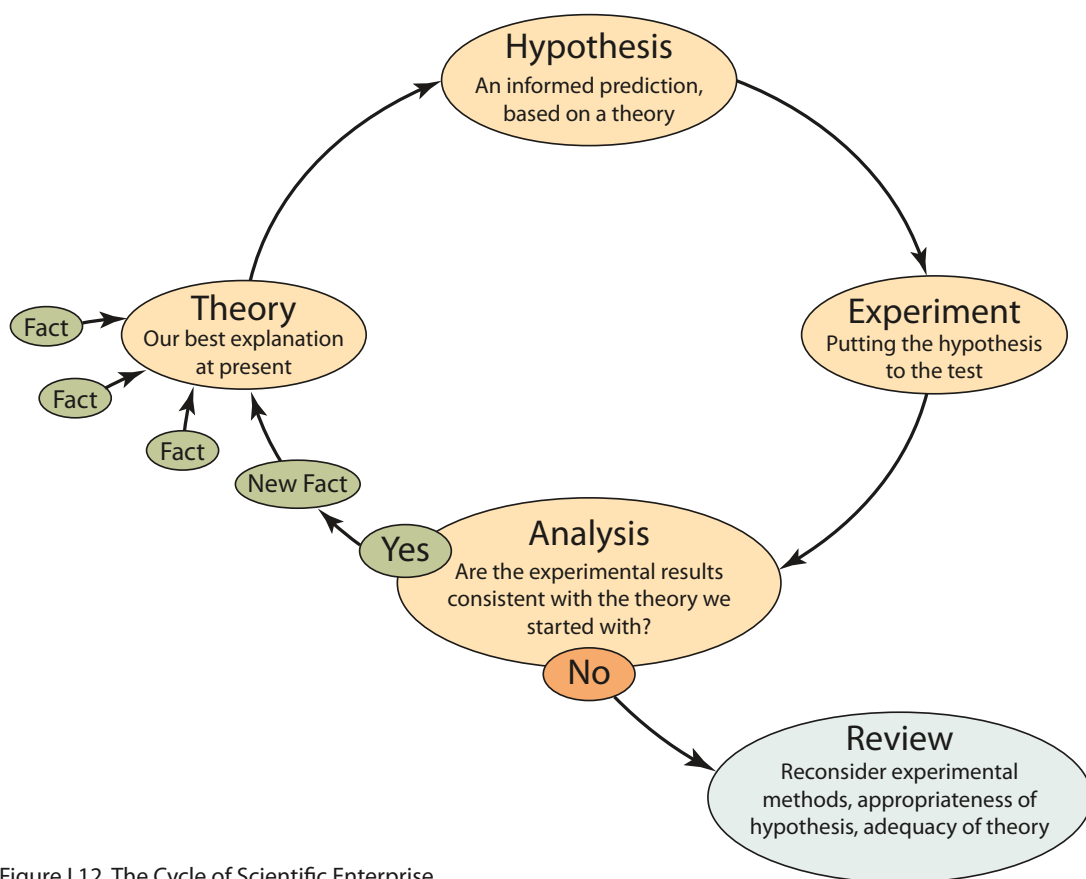


Figure I.12. The Cycle of Scientific Enterprise.

<i>Fact</i>	A proposition based on a large amount of scientific data that is correct so far as we know. Facts are discovered by experiment, observation, and inferences from experiments and observations. Facts can and do change as new scientific knowledge—new data—is acquired. Since facts are always subject to change, we generally avoid terms like <i>true</i> or <i>proven</i> . Instead, we say a fact is correct so far as we know.
<i>Theory</i>	A mental model that accounts for the data (facts) in a certain field of research, and attempts to relate them together, interpret them, and explain them. Scientific theories are successful if they repeatedly allow scientists to form new hypotheses that can be put to experimental test. <i>Successful theories are the glory and goal of science.</i> Nevertheless, theories, like facts, are provisional and subject to change. Indeed, theories are almost constantly evolving as research continues. And as with facts, when referring to theories we avoid terms like <i>true</i> or <i>proven</i> . Instead, we speak in terms of how successful theories have been in generating hypotheses that are confirmed by experiments. A widely accepted scientific theory should be understood as our best explanation at present—our best <i>model</i> of how nature works. ²
<i>Hypothesis</i>	An informed prediction about what will happen in certain circumstances. Every hypothesis is based on a particular theory. It is hypotheses that are tested in scientific experiments.
<i>Experiment</i>	A test designed to confirm or disconfirm a particular hypothesis. If a hypothesis is confirmed through experiment, and if other scientists are able to validate the confirmation by replicating the experiment, then the new facts gained from the experimental results become additional support for the theory the hypothesis came from.

Chemistry is a subject loaded with facts and heavily based on theories—*models*—that we know are incomplete descriptions of nature. That is why the research continues, as our models (hopefully) get nearer and nearer to the truth.

I.2 Conclusion

The goal of this introductory chapter is simply to alert you to some of the key concepts undergirding our understanding of chemical processes. Over and over in the coming chapters, you will find that thinking about the content in terms of one or more of these central ideas will help you develop a better grasp of the material.

Introduction Study Questions

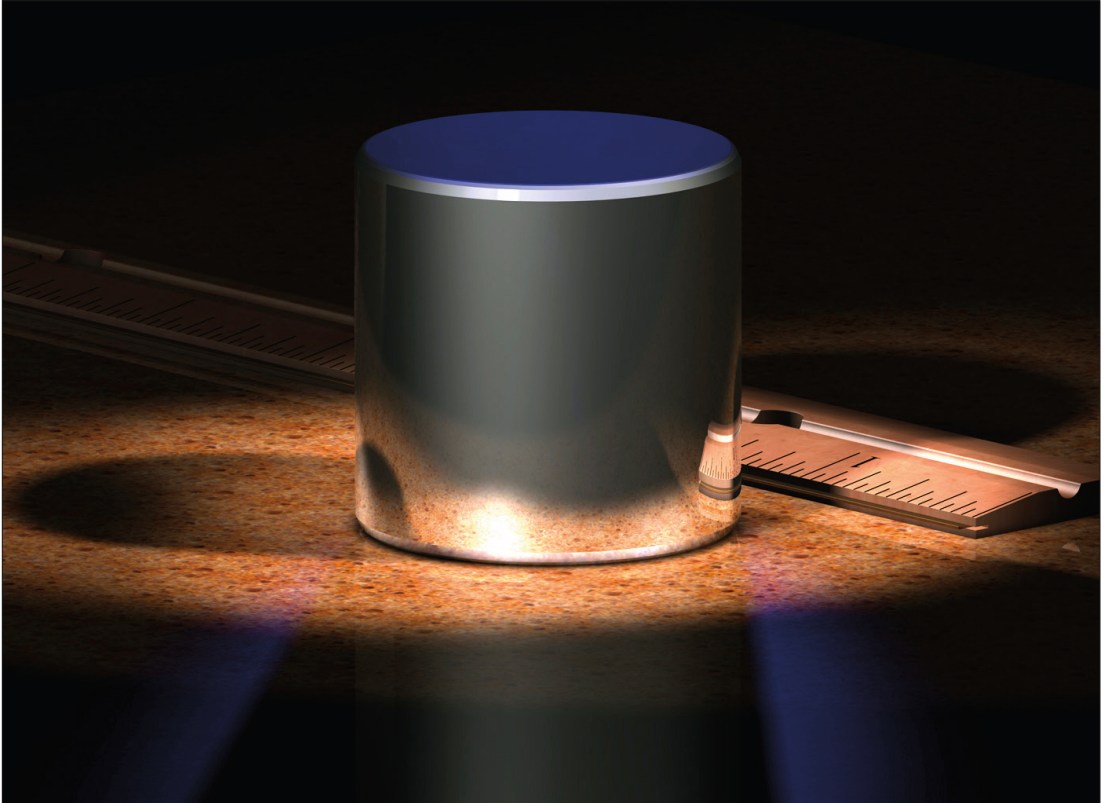
1. Write five brief paragraphs summarizing the main ideas behind the titles of Sections I.1.1 through I.1.5.
2. Describe two examples, other than those in this Introduction, of a system of some kind spontaneously (without help) moving from a higher energy state to a lower energy state. (Hint: If energy is being released, it means the entities involved in the process are moving to a lower energy state.)

² Note that the term *law* is simply an obsolete term for what we call a theory. For historical reasons, the term is still in use.

3. Describe two examples, other than those in this Introduction, of a system that will move to a lower energy state if allowed to, but which needs an initial boost of energy to get started (like the ball in Figure I.5 being kicked and then having enough energy to get out of the valley).
4. Describe two examples of processes in which entropy *decreases*. In each case, describe what source of energy and/or intelligence must be present for the decrease in entropy to occur. Here is an example to assist your thinking: an oxygen tank contains pressurized oxygen gas. The oxygen in this tank is more ordered than the oxygen in air because it has been separated from the air; there is a boundary (the tank) between the oxygen and the air. And if the valve on the tank is opened, the oxygen flows out into the air to increase the entropy (disorder). What we will never see: opening the tank valve and oxygen atoms from the atmosphere spontaneously flow into the tank. But the oxygen is put into the tank somehow, and the process that put it there decreases the entropy of that oxygen.
5. What is the ratio of nitrogen atoms to hydrogen atoms in ammonia molecules? What is the ratio of hydrogen atoms to carbon atoms in propane molecules?
6. Why are water molecules polar and what is the significance of this fact?
7. If oppositely charged objects attract, why can't a free electron and a free proton collide into one another and stick together because of their opposite charges?
8. A hydronium ion is a water molecule that has gained an extra proton. (A proton is identical to a hydrogen ion.) Hydronium ions form spontaneously in water, and are formed in greater quantities any time an acid is poured into water. What is the ratio of hydrogen atoms to oxygen atoms in hydronium ions?
9. What is hydrogen bonding?
10. Distinguish between endothermic and exothermic processes.
11. In a previous course, you may have learned about the "gold foil experiment" conducted by Ernest Rutherford in 1909. (I describe this experiment in Chapter 2.) This experiment led Rutherford to propose that the positive charge in atoms is concentrated in a tiny nucleus in the center of the atom. Think about this experiment and explain why Rutherford had to depend on inference as he interpreted his experimental data.
12. Why doesn't oil dissolve in water?
13. Distinguish between facts, theories, and hypotheses.
14. Explain why it is scientifically inappropriate to say, "no theory is true until it is proven."

Chapter 1

Measurements



The image above shows a small cylinder made of 90% platinum (element 78). This image is computer-generated, but it is an accurate representation of the official one-kilogram mass maintained in a vault in Sèvres, France by the International Bureau of Weights and Measures. The mass of the platinum cylinder in that vault is the official definition of the kilogram.

The kilogram is the only one of the seven base units in the metric system still defined by an artifact (a man-made physical object). The others are now defined in terms of various constants found in nature. Officials are looking to change the definition of the kilogram so that it, too, is defined in terms of a constant instead of an artifact. In 2014, research commenced to determine the best definition.

The kilogram is also the only base unit in the metric system that includes a metric prefix in its name.

Objectives for Chapter 1

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

SECTION 1.1

1. Define *matter* and *mass*.
2. Describe the advantages the SI system has over the USCS system for scientific work.
3. State the SI base units for length, mass, and time.
4. For the SI system of units, define the terms *base unit* and *derived unit*.
5. State several examples of base units and derived units in the SI system of units.
6. Use the metric prefixes listed in Table 1.4 from memory with various units of measure to perform unit conversions and solve problems.

SECTION 1.2

7. Given appropriate unit conversion factors, convert the units of measure for given quantities to different units of measure.
8. Given appropriate conversion factors, convert USCS units to SI units and vice versa.
9. Convert units of temperature measurements between °F, °C, and K.

SECTION 1.3

10. Define *accuracy* and *precision*.
11. Name several possibilities for sources of error in experimental measurements.
12. Explain how the significant digits in a measurement relate to the precision of the measurement.
13. Use significant digits correctly to record measurements from digital and analog measurement instruments.
14. Use significant digits correctly to perform computations, including multiplication, division, addition, subtraction, and combinations of these operations.

SECTION 1.4

15. Convert numerical values from standard notation into scientific notation and vice versa.
16. Use scientific notation when recording measurements and performing computations.
17. Calculate the percent difference between an experimental value and a theoretical value or accepted value.

1.1 Science and Measurements

1.1.1 No Measurements, No Science

One of the things that distinguishes scientific research from other fields of study is the central role played in science by *measurement*. In every branch of science, researchers study the natural world, and they do it by making measurements. The measurements we make in science are the data we use to quantify the facts we have and to test new hypotheses. These data—our measurements—answer questions such as What is its volume?, How fast is it moving?, What is its mass?, How much time does it take?, What is its diameter?, What is its frequency and wavelength?, When will it occur again?, and many others. Without measurements, modern science would not exist.

Units of measure are crucial in science. Since science is so deeply involved in making measurements, you will work with measurements a lot in this course. The value of a measurement is always accompanied by the units of measure—a measurement without its units of measure is a meaningless number. For this reason, your answers to computations in scientific calculations must always show the units. In this chapter, we discuss units of measure at some length. The material in the chapter is very important—for the rest of this book, we will be engaged with calculations involving units of measure.

However, before we launch into our study of units, there is a topic of great importance we need to nail down—*mass*. Students often do not fully understand what mass is, so we will start by reviewing the topic.

1.1.2 Matter, Volume, and Mass

The best way to understand mass is to begin with *matter* and its properties. The term *matter* refers to anything composed of atoms or parts of atoms. Note that there are many things that are not material, that is, they are not matter, as illustrated in Figure 1.1. Your thoughts, your soul, and your favorite song are not matter. You can write down your thoughts in ink, which is matter, and your song can be recorded onto a CD, which is matter. But ideas and souls are not material and are not made of what we call matter. Another part of this world that is not matter

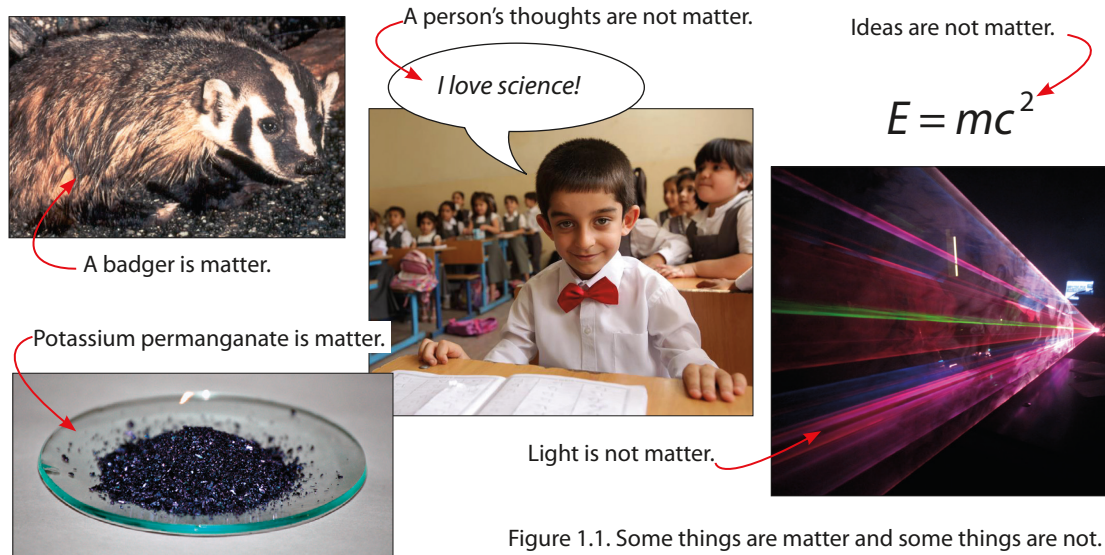


Figure 1.1. Some things are matter and some things are not.

is *electromagnetic radiation*—light, radio waves, X-rays, and all other forms of electromagnetic radiation. Light is pure energy; it is not matter and it has no mass.

We get into electromagnetic radiation a bit in Chapter 3. For now, we are going to focus on matter. A lot of what we discuss in this text is about different properties of matter. There are many different properties to discuss, but here we focus on just two properties that all matter possesses: all matter takes up space and all matter has inertia. Describing and comparing these two properties helps make clear what we mean by the term *mass*.

All matter takes up space. Even individual atoms and protons inside of atoms take up space. Now, how do we *quantify* how much space an object takes up? That is, how do we put a numerical measurement to it? The answer is, of course, by specifying its *volume*. Volume is the name of the variable we use to quantify how much space an object takes up. There are many different units of measure we use to specify an object's volume. Examples are gallons, liters, cubic meters, and pints. When we say that the volume of an object is 338 cubic centimeters, what we mean is that if we could hollow the object out and fill it up with little cubes, each with a volume of one cubic centimeter, 338 of them are needed to fill up the hollowed object.

All matter has inertia. The effect of this property is that objects resist being accelerated. The more inertia an object has, the more difficult it is to accelerate the object. For example, if the inertia of an object is small, as with, say, a golf ball, the object is easy to accelerate. Golf balls are easy to throw, and if you hit one with a golf club it accelerates at a high rate to a very high speed. But if the amount of inertia an object has is large, as with say, a grand piano, the object is difficult to accelerate. Just try throwing a grand piano or hitting one with a golf club and you will see that it doesn't accelerate at all. This is because the piano has a great deal more inertia than a golf ball.

As with the property of taking up space, we need to quantify the property of inertia. The way we do this is with the variable we call *mass*. The mass of an object is a numerical measurement specifying the amount of inertia the object has. Since inertia is a property of matter, and since all matter is composed of atoms, it should be pretty obvious that the more atoms there are packed into an object, the more mass it has. And since the different types of atoms themselves have different masses, an object made of more massive atoms has more mass than an object made of an equal number of less massive atoms.

The main unit of measure we use to specify an object's mass is the *kilogram*. There are other units such as the gram and the microgram. The kilogram (kg) is one of the base units in the metric system, our topic in Section 1.1.4. On the earth, an object weighing 2.2 pounds (lb) has a mass of one kilogram. To give you an idea of what a kilogram mass feels like in your hand on the earth, the lantern battery pictured in Figure 1.2 weighs 2.2 lb, and thus has a mass of 1 kg.

We have established that the mass of an object is a measure of its inertia, which in turn depends on how many atoms it is composed of and how massive those atoms are. The implication of this is that an object's mass does not depend on where it is. A golf ball on the earth has the same mass as a golf ball at the bottom of the ocean, on the moon, or in outer space. Even where there is no gravity, the mass of the golf ball is the same. This is what distinguishes the *mass* of an object from its *weight*.

Weight is caused by the force of gravity acting on an object composed of matter (which we often simply refer to as *a mass*). The weight of an object depends on where it is. An object—or mass—on the moon weighs only about 1/6 its weight on earth, and in outer space, where there is no gravity, a mass has no weight at all. But the mass of an object



Figure 1.2. The mass of this battery is about one kilogram.

does not depend on where it is. This is because an object’s mass is based on the matter the object is made of. The lantern battery in the figure has a certain weight on the earth (2.2 lb). In outer space, it weighs nothing and floats right in front of you. But if you try to throw the battery, the force you feel on your hand is the same on the earth or in space. That’s because the force you feel depends on the object’s mass.

Here is a summary using slightly different terminology that may help even more. Inertia is a *quality* of all matter; mass is the *quantity* of a specific portion of matter. Inertia is a quality or property all matter possesses. Mass is a quantitative variable, and it specifies an amount of matter, a quantity of matter.

1.1.3 The US Customary System

The two major systems of units students should know about are the *International System of Units*, known as the *SI system* or the metric system, and the *U.S. Customary System*, or USCS. You have probably studied these systems before and should be already familiar with some of the SI units and prefixes. In this course, we do not make much use of the measurement system you are most familiar with—the USCS. For scientific work, the entire international scientific community uses the SI system. But here I address the USCS briefly before moving on.

Americans are generally comfortable with measurements in feet, miles, gallons, inches, and degrees Fahrenheit because they grow up using this system and are very familiar with it. But in fact, the USCS is rather cumbersome. One problem is that there are many different units of

Unit	Symbol	Quantity
meter	m	length
kilogram	kg	mass
second	s	time
ampere	A	electric current
kelvin	K	temperature
candela	Cd	luminous intensity
mole	mol	amount of substance

Table 1.1. The seven base units in the SI unit system.

measure for every kind of physical quantity. Just for measuring length or distance, for example, we have the inch, foot, yard, and mile. The USCS is also full of random numbers such as 3, 12, and 5,280. A third problem is that there is no inherent connection between units for different types of quantities. Gallons have nothing whatsoever to do with feet, and quarts have nothing to do with miles.

The USCS may be familiar ground, and it may even feel patriotic to prefer it, but it is not the system of measurement scientists use. Scientists everywhere use the SI, and it is to that system we now turn.

1.1.4 The SI Unit System

In contrast to the USCS, the SI system is simple and has many advantages. There is usually only one basic unit for each kind of quantity, such as the meter for measuring length. Instead of having many unrelated units of measure for measuring quantities of different sizes, prefixes based on powers of ten are used on all the units to accommodate various sizes of quantities. And

Unit	Symbol	Quantity
joule	J	energy
newton	N	force
cubic meter	m ³	volume
watt	W	power
pascal	Pa	pressure

Table 1.2. Some SI System derived units.

units for different types of quantities relate to one another in some way. Unlike the gallon and the foot, which have nothing to do with each other, the cubic meter is 1,000,000 cubic centimeters. For all these reasons, the USCS is not used much at all in scientific work. The SI system is the international standard.

There are seven *base units* in the SI System, listed in Table 1.1. All other SI units of measure,

such as the joule (J) for measuring quantities of energy and the newton (N) for measuring amounts of force, are based on these seven base units. Units based on combinations of the seven base units are called *derived units*. A few common derived units are listed in Table 1.2.

You are already familiar with the SI unit for time: the second. You may or may not be familiar with some of the other base units, so here are some facts and photos to help familiarize you with these. A meter is just a few inches longer than a yard (3 feet). Figure 1.3 shows a wooden measuring rule one meter long, commonly called a meter stick, along with a metal yardstick for comparison.

On earth, a mass of one kilogram weighs about 2.2 pounds. The six-volt lantern battery shown in Figure 1.2 weighs just under 2.2 pounds, so the mass of the battery is just about one kilogram, as I mentioned before.

I can't show you a picture of one ampere of electric current, but it may be helpful to know that a standard electrical receptacle

(or "outlet") such as the one shown in Figure 1.4 is rated to carry 15 amperes of current. (However, the largest continuous current that the receptacle is allowed to supply is 80% of its rating, or 12 amperes. This is why vacuum cleaners are often advertised as having 12-amp motors. That's the upper limit of the current available to run them.)

Regarding temperature units, the Celsius scale is generally used for making scientific temperature measurements, but the Kelvin scale must be used for nearly all scientific calculations involving temperature. You must become familiar with both scales.

On the Celsius scale, water freezes at 0°C and boils at 100°C . Since Celsius temperature measurements can take on negative values, the Celsius scale (like the Fahrenheit scale) is not an absolute temperature scale. The Kelvin scale is an absolute scale, with 0 K (0 kelvins) being equal to absolute zero, theoretically the lower limit of possible temperatures. Note that the term "degrees" is not used when stating or writing values in kelvins.

A temperature change of one kelvin is the same as a temperature change of one degree Celsius, and both are almost double the change that a change of one degree Fahrenheit is. For reference, room temperature on the three scales is 72°F , 22.2°C , and 295.4 K .

The kilogram is the only base unit in the SI system defined by an artifact (a man-made physical object). Formerly, the meter was also defined that way; one meter used to be defined as the length of a metal bar, such as the one depicted in Figure 1.5. The very bar shown in the figure was the standard in the U.S. for one meter from 1893 to 1960. But this method of definition is not at all convenient.



Figure 1.3. A meter stick (left), with a yardstick for comparison.



Figure 1.4. A standard receptacle in American homes is rated for a current of 15 amperes.



Figure 1.5. Standard meter bar number 27, owned by the U.S. and used as the standard meter from 1893 to 1960.

Multiples	Prefix	deca-	hecto-	kilo-	mega-	giga-	tera-	peta-	exa-	zetta-	yotta-
	Symbol	da	h	k	M	G	T	P	E	Z	Y
	Factor	10	10 ²	10 ³	10 ⁶	10 ⁹	10 ¹²	10 ¹⁵	10 ¹⁸	10 ²¹	10 ²⁴
Fractions	Prefix	deci-	centi-	milli-	micro-	nano-	pico-	femto-	atto-	zetto-	yocto-
	Symbol	d	c	m	μ	n	p	f	a	z	y
	Factor	1/10	1/10 ²	1/10 ³	1/10 ⁶	1/10 ⁹	1/10 ¹²	1/10 ¹⁵	1/10 ¹⁸	1/10 ²¹	1/10 ²⁴

Table 1.3. The SI System prefixes.

In 1960, the definition of one meter was changed to be equal to a certain number of wavelengths of a certain color of light emitted by a certain isotope of the element krypton. In 1983, the definition of the meter was changed again. Today, the meter is defined as equal to the distance light travels in a vacuum in exactly 1/299,792,458 seconds.

The SI system includes not only the base and derived units, but all the other units that can be formed by adding metric prefixes to these units. We address the prefixes in the next section. But first I will mention a particular subset of the SI system known as the MKS system. MKS stands for meter-kilogram-second. The MKS system uses only the base and derived units without the prefixes (except for the kilogram, the only base unit with a prefix). The nice thing about the MKS system is that any calculation performed with MKS units produces a result in MKS units. For this reason, the MKS system is used almost exclusively in physics. However, in chemistry, it is common to use SI units that are not MKS units. Some commonly used non-MKS units are the gram (g), the centimeter (cm), the cubic centimeter (cm³), the liter (L), and the milliliter (mL). The liter is not actually an official SI unit, but it is used all the time in chemistry anyway.

1.1.5 Metric Prefixes

In the system of units commonly used in the U.S., different units are used for different sizes of objects. For example, for short distances we might use the inch or the foot, whereas for longer distances we switch to the mile. For the small volumes used in cooking, we use the fluid ounce (or pint, quart, teaspoon, tablespoon, etc.), but for larger volumes like the gasoline in the gas tank of a car, we switch to the gallon. (That's six different volume units I just listed!)

The SI System is much simpler. Each type of quantity—such as length or volume—has one main unit of measure. Instead of using several different units for different sizes of quantities, the SI System uses multipliers on the units to multiply them for large quantities, or to scale them down for smaller quantities. We call these multipliers the *metric prefixes*. The complete list of the 20 metric prefixes is in Table 1.3. You do not need to memorize all these; some are rarely used. But you do need to memorize some of them. I recommend that all science students commit to memory the prefixes listed in Table 1.4.

Fractions			Multiples		
Prefix	Symbol	Factor	Prefix	Symbol	Factor
centi-	c	1/10 ²	kilo-	k	10 ³
milli-	m	1/10 ³	mega-	M	10 ⁶
micro-	μ	1/10 ⁶	giga-	G	10 ⁹
nano-	n	1/10 ⁹	tera-	T	10 ¹²
pico-	p	1/10 ¹²			

Table 1.4. SI System prefixes to commit to memory.

Table 1.5 shows a few representative examples of how to use the prefixes to represent multiples (quantities larger than the SI base unit) and fractions (quantities smaller than the SI base unit). As illustrations, let's look at a couple of these more closely.

	Prefix	Symbol	Meaning	Examples of usage
Multiples	kilo–	k	1,000	One kilojoule is 1,000 joules. There are 1,000 joules in one kilojoule, so $1,000\text{ J} = 1\text{ kJ}$.
	mega–	M	1,000,000	One megawatt is 1,000,000 watts. There are 1,000,000 watts in one megawatt, so $1,000,000\text{ W} = 1\text{ MW}$.
Fractions	centi–	c	1/100	One centimeter is 1/100 of a meter. There are 100 centimeters in one meter, so $100\text{ cm} = 1\text{ m}$.
	milli–	m	1/1,000	One milligram is 1/1,000 of a gram. There are 1,000 milligrams in one gram, so $1,000\text{ mg} = 1\text{ g}$.
	micro–	μ	1/1,000,000	One microliter is 1/1,000,000 of a liter. There are 1,000,000 microliters in one liter, so $1,000,000\text{ }\mu\text{L} = 1\text{ L}$.

Table 1.5. Examples of correct usage of metric prefixes.

The prefix *kilo–* is a *multiple*, and it means 1,000. One kilogram is 1,000 grams, one kilometer is 1,000 meters, and so on. Figure 1.6 illustrates with my favorite dairy product—chocolate chip cookie dough ice cream. One gram is only a fraction of a taste (and contains only two chocolate chips and no cookie dough). A kilogram of ice cream is 1,000 grams of ice cream, equivalent to two large bowls of ice cream.

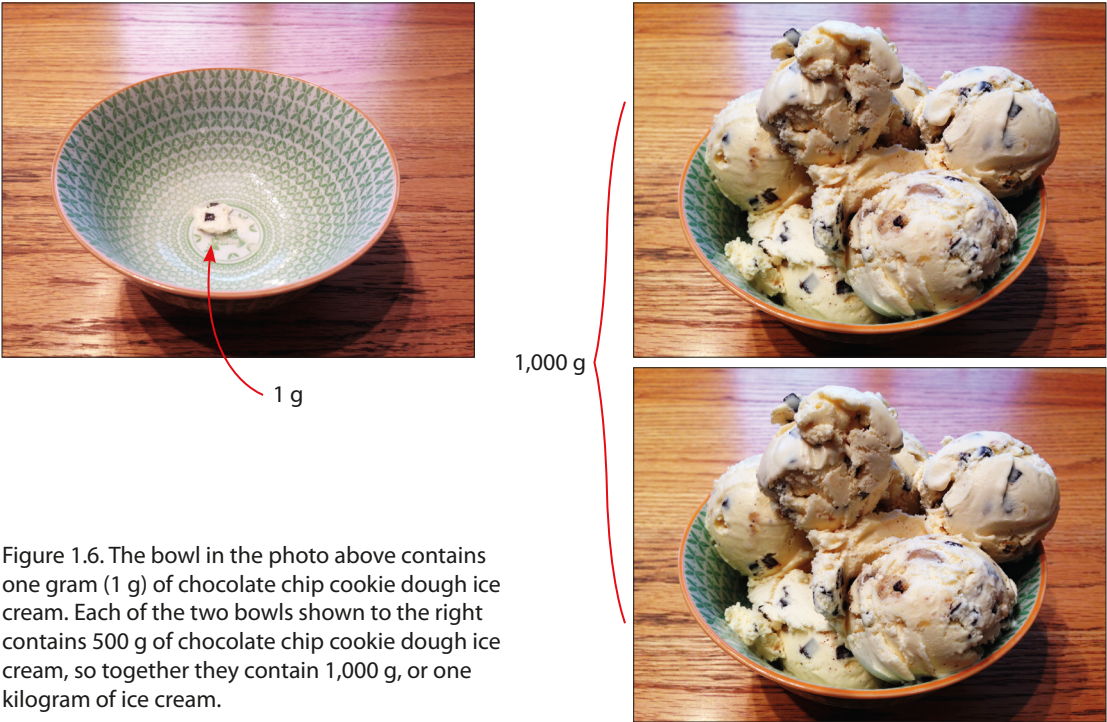


Figure 1.6. The bowl in the photo above contains one gram (1 g) of chocolate chip cookie dough ice cream. Each of the two bowls shown to the right contains 500 g of chocolate chip cookie dough ice cream, so together they contain 1,000 g, or one kilogram of ice cream.

The prefix *milli-* is a *fraction*, and it means one thousandth. One millimeter is one thousandth of a meter, and so on. The wooden rule in Figure 1.3 is one meter in length. A millimeter is one thousandth of this length, equal to the width of the line in Figure 1.7.

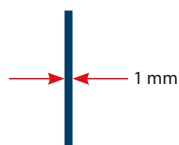


Figure 1.7. One millimeter, which is one thousandth of a meter.

We conclude this introduction to metric prefixes with a few brief notes. First, when using the prefixes for quantities of mass, prefixes are never added to the kilogram. Prefixes are only added to the gram, even though the kilogram is the base unit in the SI system, not the gram. Second, note that when writing the symbols for metric prefixes, the case of the letter matters: *kilo-* always takes a lower-case k, *mega-* always takes an upper-case M, and so on. Third, one of the prefix symbols is not an English letter. The prefix μ for *micro-* is the lower-case Greek letter *mu*, the m in the Greek alphabet. Finally, pay close attention to the difference between multiplier prefixes and fraction prefixes. Learning to use the fraction prefixes properly is the most challenging part of mastering the SI System of units, and using them incorrectly in unit conversion factors (our next topic) is a common student error.

1.2 Converting Units of Measure

1.2.1 Basic Principles of Unit Conversion Factors

For scientists and engineers, one of the most commonly used skills is re-expressing quantities into equivalent quantities with different units of measure. These calculations are called *unit conversions*. Mastery of this skill is essential for all students studying science. In this section, I describe what unit conversion factors are and how they are used.

Let's begin with the basic principles of how unit conversions work. First, we all know that multiplying any value by *unity* (one) leaves its value unchanged. Second, we also know that in any fraction, if the numerator and denominator are equivalent, the value of the fraction is unity (one). For example, the expression "12 bricks over 12 bricks" is equal to one:

$$\frac{12 \text{ bricks}}{12 \text{ bricks}} = 1$$

This is because the numerator and denominator are equivalent, and any time this is the case, the value of the fraction is one, or unity.

A *unit conversion factor* is simply a fractional expression in which the numerator and denominator are equivalent ways of writing the same physical quantity with different units of measure. This means a conversion factor is just a special way of writing unity (one).

The third basic principle is that when multiplying fractions, factors that appear in both the numerator and denominator may be "cancelled out." So when performing ordinary unit conversions, what we are doing is repeatedly multiplying a given quantity by unity so that cancellations alter the units of measure until they are expressed the way we wish. Since all we are doing is multiplying by one, the value of our original quantity is unchanged; it simply looks different because it is expressed with different units of measure.

There are many different units of measure and there are many different conversion factors used for performing unit conversions. Table A.3 in Appendix A lists a number of important ones.

Let me elaborate a bit more on the idea of unity I mentioned above, using one common conversion factor as an example. American school kids all learn that there are 5,280 feet in one mile, which means $5,280 \text{ ft} = 1 \text{ mi}$. One mile and 5,280 feet are equivalent ways of writing the same length. If we place these two expressions into a fraction, the numerator and denominator

are equivalent, so the value of the fraction is unity, regardless of the way we write it. The equation $5,280 \text{ ft} = 1 \text{ mi}$ can be written as a conversion factor two different ways, and the fraction equals unity either way:

$$\frac{5280 \text{ ft}}{1 \text{ mi}} = \frac{1 \text{ mi}}{5280 \text{ ft}} = 1 \quad (1.1)$$

Now, I need to make an important clarification about the use of the equal sign in the expressions I just wrote. In mathematics, the equal sign means *identity*. An expression such as $3 \text{ miles} = 3 \text{ miles}$ is a mathematical *identity*; the first expression, “3 miles” is *identical* to the second expression, also “3 miles.” But when we are dealing with converting units of measure from one set of units to a different set of units, we don’t use the equal sign to mean identity. Obviously, the expression

$$\frac{5280 \text{ ft}}{1 \text{ mi}}$$

is not *identical* to the expression

$$\frac{1 \text{ mi}}{5280 \text{ ft}}$$

One of these has units of ft/mi and the other has units of mi/ft. But even though these expressions are not identical, they are *equivalent*. When we are dealing with converting units of measure, this is the sense in which we interpret the equal sign. We are using it to mean *equivalent*. This is why I can write Equation (1.1) using equal signs. The three terms in Equation (1.1) are not identical, but they are equivalent.

Suppose you have a measurement such as 43,000 feet that you wish to re-express in miles. To convert the units from feet to miles, first write down the quantity you are given, with its units of measure:

$$43,000 \text{ ft}$$

Next, select a unit conversion factor containing the units you presently have and the ones you want to convert to. (This is not always possible. Sometimes more than one conversion factor is required, as Example 1.1 below illustrates.) In this case, those units are feet and miles, and the conversion factors containing these units are the two written in Equation (1.1). As I explain below, the one we need for converting 43,000 ft into miles is the second one. So to perform the conversion, you multiply your given quantity by the conversion factor. Then you cancel any units that appear both in the numerator and the denominator, as follows:

$$43,000 \cancel{\text{ ft}} \cdot \frac{1 \text{ mi}}{5280 \cancel{\text{ ft}}} = 8.1 \text{ mi}$$

There are two important comments to make here. First, since any conversion factor can be written two ways (depending on which quantity is placed in the numerator), how do we know which way to write the conversion factor? Well, we know from algebra that when we have quantities in the numerator of a fraction that are multiplied (factors), and quantities in the denominator of the fraction that are multiplied (factors), any quantities that appear in both the numerator and denominator can be cancelled out. In the example above, we want to cancel out the “feet” in the given quantity (which is in the numerator), so the conversion factor needs to be written with feet in the denominator and miles in the numerator.

Second, if you perform the calculation above ($43,000 \div 5,280$), the result that appears on your calculator screen is 8.143939394. So why didn't I write down all those digits in my result? Why did I round my answer off to simply 8.1 miles? The answer to that question has to do with the significant digits in the value 43,000 ft that we started with. We address the issue of significant digits later in this chapter, but in the examples that follow I always write the results with the correct number of significant digits for the values involved in the problem.

The following example illustrates the use of conversion factors based on metric prefixes. This example also illustrates how to perform a conversion when more than one conversion factor is required.

▼ Example 1.1

Convert the value 2,953,000 μg into kilograms.

Referring to Tables 1.4 and 1.5, you see that the symbol μ means *micro*-, which means one millionth. Thus, μg means millionths of a gram. We use this information to make conversion factors. Since it takes 1,000,000 millionths of a gram to make one gram, $1,000,000 \mu\text{g} = 1 \text{ g}$, and thus

$$\frac{1,000,000 \mu\text{g}}{1 \text{ g}} = \frac{1 \text{ g}}{1,000,000 \mu\text{g}} = 1$$

Note that converting from μg to g only gets us part of the way toward the solution. We need another conversion factor to get from g to kg. Looking again at Tables 1.4 and 1.5, we see that there are 1,000 grams in one kilogram, or $1,000 \text{ g} = 1 \text{ kg}$. From this we can make additional conversion factors:

$$\frac{1000 \text{ g}}{1 \text{ kg}} = \frac{1 \text{ kg}}{1000 \text{ g}} = 1$$

Now to perform the conversion, first convert micrograms into grams. Then convert grams into kilograms. You can do this two-step conversion at the same time by simply multiplying both conversion factors at the same time as follows:

$$2,953,000 \mu\text{g} \cdot \frac{1 \text{ g}}{1,000,000 \mu\text{g}} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 0.002953 \text{ kg}$$

The μg in the given quantity cancels with the μg in the denominator of the first conversion factor. The g in the first conversion factor cancels with the g in the denominator of the second conversion factor. The units we are left with are the kg in the numerator of the second conversion factor. The kg did not cancel out with anything, so these are the units of our result.



1.2.2 Tips for Converting Units of Measure

There are several important points you must remember in order to perform unit conversions correctly. I illustrate them below with examples. You should rework each of the examples on your own paper as practice to make sure you can do them correctly. The conversion factors used in the examples below are all listed in Table A.3 in Appendix A.

Point 1 *Never use slant bars in your unit fractions. Use only horizontal bars.*

In printed materials, we often see values written with a slant fraction bar in the units, as in the value 35 m/s. Although writing the units this way is fine for a printed document, you should not write values this way when you are performing unit conversions. This is because it is easy to get confused and not notice that one of the units is in the denominator in such an expression (s, or seconds, in my example), and the conversion factors used must take this into account.

▼ **Example 1.2**

Convert 57.66 mi/hr into m/s.

Writing the given quantity with a horizontal bar makes it clear that the “hours” is in the denominator. This helps you write the hours-to-seconds factor correctly.

To perform this conversion, we must convert the miles in the given quantity into meters, and we must convert the hours into seconds. From Table A.3, we select the two conversion factors we need. Then we multiply the given quantity by them to convert the mi/hr into m/s. When doing the multiplying, we write all the unit fractions with horizontal bars.

$$57.66 \frac{\text{mi}}{\text{hr}} \cdot \frac{1609 \text{ m}}{1 \text{ mi}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 25.77 \frac{\text{m}}{\text{s}}$$

As you see, the miles cancel and the hours cancel, leaving meters in the numerator and seconds in the denominator. Now that you have your result, you may write it as 25.77 m/s if you wish, but do not use slant fraction bars in the units when you are working out the unit conversion.



Point 2 *The term “per,” abbreviated p, implies a fraction.*

Some units of measure are commonly written with a “p” for “per,” such as mph for miles per hour, or gps for gallons per second. Change these expressions to fractions with horizontal bars when you work out the unit conversion.

▼ **Example 1.3**

Convert 472.2 gps to L/hr.

When you write down the given quantity, change the gps to gal/s and write these units with a horizontal bar:

$$472.2 \frac{\text{gal}}{\text{s}} \cdot \frac{3.785 \text{ L}}{1 \text{ gal}} \cdot \frac{3600 \text{ s}}{\text{hr}} = 6,436,000 \frac{\text{L}}{\text{hr}}$$



Point 3 *Use the $\boxed{\times}$ and $\boxed{\div}$ keys correctly when entering values into your calculator.*

When dealing with several numerator terms and several denominator terms, multiply all the numerator terms together first, hitting the $\boxed{\times}$ key between each, then hit the $\boxed{\div}$ key and enter

all the denominator terms, hitting the \div key between each. This way you do not need to write down intermediate results, and you do not need to use any parentheses.

▼ Example 1.4

Convert 43.2 mm/hr into km/yr.

The setup with all the conversion factors is as follows:

$$43.2 \frac{\text{mm}}{\text{hr}} \cdot \frac{1 \text{ m}}{1000 \text{ mm}} \cdot \frac{1 \text{ km}}{1000 \text{ m}} \cdot \frac{24 \text{ hr}}{1 \text{ dy}} \cdot \frac{365 \text{ dy}}{1 \text{ yr}} = 0.378 \frac{\text{km}}{\text{yr}}$$

To execute this calculation in your calculator, you enter the values and operations in this sequence:

$$43.2 \times 24 \times 365 \div 1000 \div 1000 =$$

If you do so, you get 0.37843200. (Again, significant digits rules require us to round to 0.378.)



Point 4

When converting units for area and volume such as cm^2 or m^3 , you must use the appropriate length conversion factor twice for areas and three times for volumes.

The units “ cm^2 ” for an area mean the same thing as “ $\text{cm} \times \text{cm}$.” Likewise, “ m^3 ” means “ $\text{m} \times \text{m} \times \text{m}$.” So when you use a length conversion factor such as $100 \text{ cm} = 1 \text{ m}$ or $1 \text{ in} = 2.54 \text{ cm}$, you must use it twice to get squared units (areas) or three times to get cubed units (volumes).

▼ Example 1.5

Convert $3,550 \text{ cm}^3$ to m^3 .

$$3550 \text{ cm}^3 \cdot \frac{1 \text{ m}}{100 \text{ cm}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 0.00355 \text{ m}^3$$

Notice in this example that the unit cm occurs three times in the denominator, giving us cm^3 when they are all multiplied together. This cm^3 term in the denominator cancels with the cm^3 term in the numerator. And since the m unit occurs three times in the numerator, they multiply together to give us m^3 for the units in our result.



The issue of needing to repeat conversion factors only arises when you are using a unit raised to a power, such as a when a length unit is used to represent an area or a volume. When using a conversion factor such as $3.785 \text{ L} = 1 \text{ gal}$, the units of measure are written using units that are strictly volumetric (liters and gallons), and are not obtained from lengths the way they are with in^2 , ft^2 , cm^3 , and m^3 . Another common unit that uses an exponent is acceleration, which has units of m/s^2 in the MKS unit system.

▼ Example 1.6

Convert 5.85 mi/hr^2 into MKS units.

The MKS unit for length is meters (m), so we must convert miles to meters. The MKS units for time is seconds (s), so we must convert hr^2 into s^2 .

$$5.85 \frac{\text{mi}}{\text{hr}^2} \cdot \frac{1609 \text{ m}}{1 \text{ mi}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 0.000726 \frac{\text{m}}{\text{s}^2}$$

With this example, you see that since the “hours” unit is squared in the given quantity, the conversion factor converting the hours to seconds must appear twice in the conversion calculation. The “miles” unit in the given quantity has no exponent, so the conversion factor used to convert miles to meters only appears once in the calculation.



1.2.3 Converting Temperature Units

Converting temperature values from one scale to another requires the use of equations rather than conversion factors. This is due to the fact that the Fahrenheit and Celsius scales are not absolute temperature scales. If all temperature scales were absolute scales like the Kelvin scale is, temperature conversions could be performed with conversion factors just as other conversions are. To convert a temperature in degrees Fahrenheit (T_F) into degrees Celsius (T_C), use this equation:

$$T_C = \frac{5}{9}(T_F - 32^\circ)$$

Using a bit of algebra, we can work this around to give us an equation that can be used to convert Celsius temperatures to Fahrenheit values:

$$T_F = \frac{9}{5}T_C + 32^\circ$$

To convert a temperature in degrees Celsius into kelvins (T_K), use this equation:

$$T_K = T_C + 273.15$$

Again, some algebra gives us the equation the other way around.

$$T_C = T_K - 273.15$$

All four of the temperature conversion equations above are exact. This is important to know later when we discuss significant digits. These equations are listed in Table A.3 in Appendix A.

▼ Example 1.7

The normal temperature of the human body is 98.6°F . Express this value in degrees Celsius and kelvins.

Since the given value is in degrees Fahrenheit, write down the equation that converts values from $^\circ\text{F}$ to $^\circ\text{C}$.

$$T_C = \frac{5}{9}(T_F - 32^\circ)$$

Now insert the Fahrenheit value and calculate the Celsius value.

$$T_C = \frac{5}{9}(98.6^\circ - 32^\circ) = 37.0^\circ\text{C}$$

Now we are able to use the Celsius value to compute the Kelvin value.

$$T_K = T_C + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

The reason the answer is 310.2 K instead of 310.15 K is due to the significant digits rule for addition. Again, the topic of significant digits is coming up next.



▼ Example 1.8

The melting point of aluminum is 933.5 K. Express this temperature in degrees Celsius and degrees Fahrenheit.

Write down the equation that converts Kelvin values to Celsius values.

$$T_C = T_K - 273.15$$

From this we calculate the Celsius value as

$$T_C = T_K - 273.15 = 933.5 - 273.15 = 660.3^\circ\text{C}$$

Next, write down the equation for converting a Celsius value to a Fahrenheit value.

$$T_F = \frac{9}{5}T_C + 32^\circ$$

Insert the Celsius value into this equation and calculate the Fahrenheit value.

$$T_F = \frac{9}{5}T_C + 32^\circ = \frac{9}{5} \cdot 660.3^\circ\text{C} + 32^\circ = 1220.5^\circ\text{F}$$



1.3 Accuracy and Precision

1.3.1 Distinguishing Between Accuracy and Precision

The terms *accuracy* and *precision* refer to the practical limitations inherent in making measurements. Science is all about investigating nature, and to do that we must make measurements.

Accuracy relates to error—that is, to the lack of it. Error is the difference between a measured value and the true value. The lower the error is in a measurement, the better the accuracy. Error arises from many different sources, including human mistakes, malfunctioning equipment, incorrectly calibrated instruments, vibrations, changes in temperature or humidity, or unknown causes that are influencing a measurement without the knowledge of the experimenter. All measurements contain error, because (alas!) perfection is simply not a thing we have access to in this world.

Precision refers to the resolution or degree of “fine-ness” in a measurement. The limit to the precision that can be obtained in a measurement is ultimately dependent on the instrument being used to make the measurement. If you want greater precision, you must use a more precise instrument. The degree of precision in every measurement is signified by the measurement value itself because the precision is a built-in part of the measurement. *The precision of a measurement is indicated by the number of significant digits (or significant figures) included in the measurement value when the measurement is written down* (see below).

Here is an example that illustrates the idea of precision and also helps distinguish between precision and accuracy. The photograph in Figure 1.8 shows a machinist’s rule and an architect’s scale set one above the other. Since the marks on the two scales line up consistently, *these two scales are equally accurate*. But the machinist’s rule (on top) is more precise. The architect’s scale is marked in 1/16-inch increments, but the machinist’s rule is marked in 1/64-inch increments. Thus, *the machinist’s rule is more precise*.

It is important that you are able to distinguish between accuracy and precision. Here is another example to help illustrate the difference. Let’s say Shana and Marius each buy digital thermometers for their homes. The thermometer Shana buys costs \$10 and measures to the nearest 1°F. Marius pays \$40 and gets one that reads to the nearest 0.1°F. Shana reads the directions and properly installs the sensor for her new thermometer in the shade. Marius doesn’t read the directions and mounts his sensor in the direct sunlight, which causes a significant error in the thermometer reading when the sun is shining on it; thus Marius’ measurements are not accurate. The result is that Shana has lower-precision, higher-accuracy measurements!

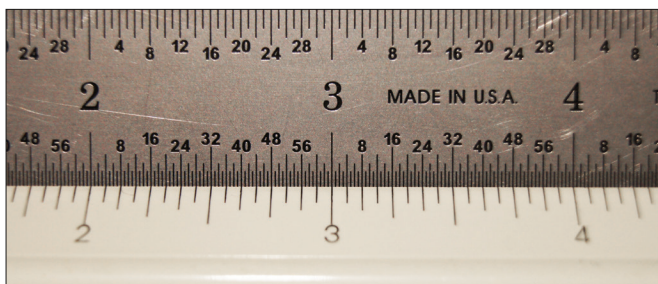


Figure 1.8. The accuracy of these two scales is the same, but the machinist’s rule (above) is more precise than the architect’s scale (below).

1.3.2 Significant Digits

The precision in any measurement is indicated by the number of *significant digits* it contains. Thus, the number of digits we write in any measurement we deal with in science is very important. The number of digits is meaningful because it shows the precision inherent in the instrument used to make the measurement.

Let’s say you are working a computational exercise in a science book. The problem tells you that a person drives a distance of 110 miles at an average speed of 55 miles per hour and wants you to calculate how long the trip takes. The correct answer to this problem *is different* from the correct answer to a similar problem with given values of 110.0 miles and 55.0 miles per hour. And if the given values are 110.0 miles and 55.00 miles per hour, the correct answer is different yet again. Mathematically, of course, all three answers are the same. If you drive 110 miles at 55 miles per hour, the trip takes two hours. But scientifically, the correct answers to these three problems are different: 2.0 hours, 2.00 hours, and 2.000 hours, respectively. The difference between these cases is in the precision indicated by the given data, which are *measurements*. (Even though this is just a made-up problem in a book and not an actual measurement someone made in an experiment, the given data are still measurements. There is no way to talk about distances or speeds without talking about measurements, even if the measurements are only imaginary or hypothetical.)

So when you perform a calculation with physical quantities (measurements), you can't simply write down all the digits shown by your calculator. The precision inherent in the measurements used in a computation governs the precision in any result you might calculate from those measurements. And since the precision in a measurement is indicated by the number of significant digits, data and calculations must be written with the correct numbers of significant digits. To do this, you need to know how to count significant digits, and you must use the correct number of significant digits in all your calculations and experimental data.

Correctly counting significant digits involves four different cases:

1. Rules for determining how many significant digits there are in a given measurement.
2. Rules for writing down the correct number of significant digits in a measurement you are making and recording.
3. Rules for computations you perform with measurements—multiplication and division.
4. Rules for computations you perform with measurements—addition and subtraction.

We address each of these cases below, in order.

Case 1

We begin with the rule for determining how many significant digits there are in a given measurement value. The rule is as follows:

- *The number of significant digits in a number is found by counting all the digits from left to right beginning with the first nonzero digit on the left. When no decimal is present, trailing zeros are not considered significant.*

Let's apply this rule to the following values to see how it works.

15,679 This value has five significant digits.

21.0005 This value has six significant digits.

37,000 This value has only two significant digits because when there is no decimal, trailing zeros are not significant. Notice that the word *significant* here is a reference to the precision of the measurement, which in this case is rounded to the nearest thousand. The zeros in this value are certainly *important*, but they are not *significant* in the context of precision.

0.0105 This value has three significant digits because we start counting digits with the first nonzero digit on the left.

0.001350 This value has four significant digits. Trailing zeros count when there is a decimal.

The significant digit rules enable us to tell the difference between two measurements such as 13.05 m and 13.0500 m. Again, these values are obviously equivalent *mathematically*. But they are different in what they tell us about the process of how the measurements were made—and science deals in measurements. The first measurement has four significant digits. The second measurement is more precise—it has six significant digits and was made with a more precise instrument.

Now, just in case you are bothered by the zeros at the end of 37,000 that are not significant, here is one more way to think about significant digits that may help. The precision in a measurement depends on the instrument used to make the measurement. If we express the measure-

ment in different units, this cannot change the precision of the value. A measurement of 37,000 grams is equivalent to 37 kilograms, as shown in the following calculation:

$$37,000 \text{ g} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 37 \text{ kg}$$

Whether we express this value in grams or kilograms, it still has two significant digits.

Case 2

The second case addresses the rules that apply when you are recording a measurement yourself, rather than reading a measurement someone else has made. When you make measurements yourself, as when conducting the laboratory experiments in this course, you must know the rules for which digits are significant in the reading you are making on the measurement instrument. The rule for making measurements depends on whether the instrument you are using is a digital instrument or an analog instrument. Here are the rules for these two possibilities:

- **Rule 1 for digital instruments** *For the digital instruments commonly found in introductory science labs, assume all the digits in the reading are significant except leading zeros.*
- **Rule 2 for analog instruments** *The significant digits in a measurement include all the digits known with certainty, plus one digit at the end that is estimated between the finest marks on the scale of your instrument.*

The first of these rules is illustrated in Figure 1.9. The reading on the left has leading zeros, which do not count as significant. Thus, the first reading has three significant digits. The second reading also has three significant digits. The third reading has five significant digits.

The fourth reading also has five significant digits because with a digital display the only zeros that don't count are the leading zeros. Trailing zeros are significant with a digital instrument. However, when you write this measurement down, you must write it in a way that shows those zeros to be significant. The way to do this is by using scientific notation. When a value is written in scientific notation, *the digits that are written down in front of the power of 10* (the stem, also called the mantissa) *are the significant digits*. Thus, the right-hand value in Figure 1.9 must be written as 4.2000×10^4 . We address scientific notation in more detail in the next section.

Dealing with digital instruments is actually more involved than the simple rule above implies, but the issues involved go way beyond what we can deal with in introductory science classes. So, simply make your readings and assume that all the digits in the reading except leading zeros are significant.

Now let's look at some examples illustrating the rule for analog instruments. Figure 1.10 shows a machinist's rule being used to measure the length in millimeters (mm) of a brass block. We know the first two digits of the length with certainty; the block is clearly between 31 mm and 32 mm long. We have to estimate the third significant digit. The scale on the rule is marked in increments of 0.5 mm. Comparing the edge of the block with these marks, I estimate the next digit to be a 6, giving a measurement of 31.6 mm. Others might estimate the last digit to be a 5 or a 7; these small differences in the last digit are unavoidable because the last digit is estimated. Whatever you estimate the last digit to be, two digits of this measurement are known with certainty, the third digit is estimated, and the measurement has three significant digits.



Figure 1.9. With digital instruments, all digits are significant except leading zeros. Thus, the numbers of significant digits in these readings are, from left to right, three, three, five, and five.

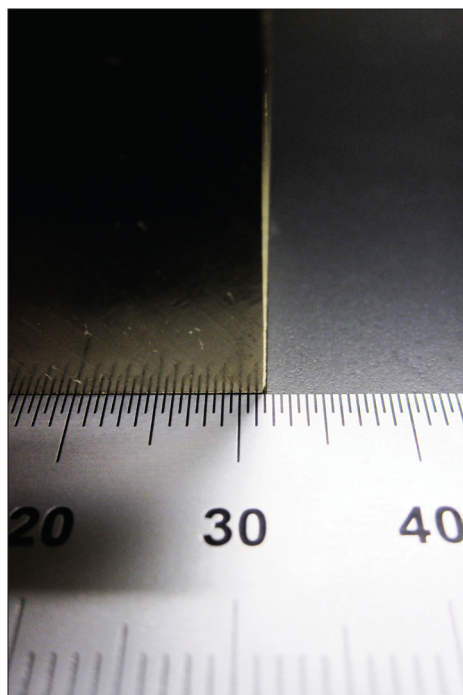


Figure 1.10. Reading the significant digits with a machinist's rule.

The photograph in Figure 1.11 shows a liquid volume measurement in milliliters (mL) being made with an article of apparatus called a *buret*. Notice in this figure that when measuring liquid volume the surface of the liquid curls up at the edge of the cylinder, forming a bowl-shaped surface on the liquid. This curved surface is called a *meniscus*. For most liquids, liquid measurement readings are taken at the bottom of the meniscus. Liquid mercury is the major exception, because the meniscus in liquid mercury is inverted—liquid mercury curves down at the edges. In that case, the measurement is read at the top of the meniscus. But that is an unusual case. For most liquids, the reading is made at the bottom of the meniscus.

For the buret in the figure, you can see that the scale is marked in increments of 0.1 milliliters (mL). This means we are to estimate to the nearest 0.01 mL. To one person, it may look like the bottom of the meniscus (where the black curve touches the bottom of the silver bowl) is just above 2.2 mL, so that person would call this measurement 2.19 mL. To someone else, it may seem that the bottom of the meniscus is right on 2.2, in which case that person would call the reading 2.20 mL. Either way, the reading has three significant digits and the last digit is estimated to be either 9 or 0.

The third example involves a liquid volume measurement with an article of apparatus called a *graduated cylinder*. The scales on small graduated cylinders like this one are marked in increments of 1 mL. In the photo of Figure 1.12, the entire meniscus appears silvery in color with

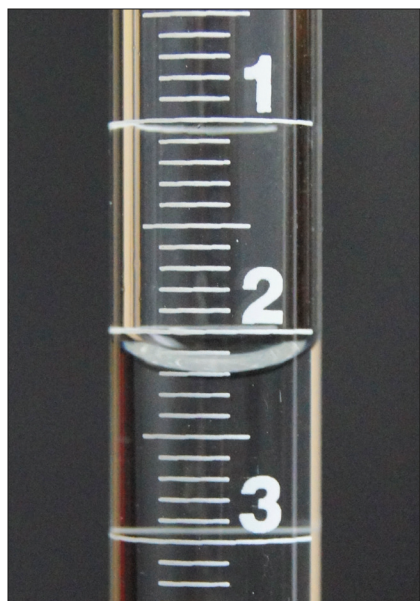


Figure 1.11. Reading the significant digits on a buret.

a black curve at the bottom. For the liquid shown in the figure, we know the first two digits of the volume measurement with certainty because the reading at the bottom of the meniscus is clearly between 82 mL and 83 mL. We have to estimate the third digit, and I estimate the edge of the meniscus to be at 60% of the distance between 82 and 83, giving a reading of 82.6 mL. Others may prefer a different value for that third digit.



Figure 1.12. Reading the significant digits on a graduated cylinder.

It is important for you to keep the significant digits rules in mind when you are making measurements and entering data for your lab reports. The data in your lab journal and the values you use in your calculations and report should correctly reflect the use of the significant digits rules as they apply to the actual instruments you use to make your measurements.

Case 3

The third and fourth cases of rules for significant digits apply to the calculations you perform with measurements. In Case 3, we deal with multiplication and division. The main idea behind the rule for multiplying and dividing is that the precision you report in your result cannot be higher than the precision that is in the measurements you start with. The precision in a measurement depends on the instrument used to make the measurement, nothing else. Multiplying and dividing things cannot increase that precision, and thus your results can be no more precise than the measurements used in the calculations. In fact, your result can be no more precise than the *least precise value* used in the calculation. The least precise value is, so to speak, the “weak link” in the chain, and a chain is no stronger than its weakest link.

Here are the two rules for using significant digits in calculations involving multiplication and division:

- **Rule 1** When multiplying or dividing, count the significant digits in each of the values you are using in the calculation, including any conversion factors involved. (However, note: Conversion factors that are exact are not considered.) Determine how many significant digits there are in the least precise of these values. The result of your calculation must have this same number of significant digits.
- **Rule 2** When performing a multi-step calculation, keep at least one extra digit during intermediate calculations, and round off to the final number of significant digits you need at the very end. This practice ensures that small round-off errors don't accumulate during a multi-step calculation. This extra digit rule also applies to unit conversions performed as part of the computation.

I illustrate the two rules above, along with some more unit conversions, in the following example problem and calculation.

▼ Example 1.9

At a chemical research lab, a stream of a reactant solution is flowing into a reaction vessel at a rate of $56.75 \mu\text{L}$ per second. A volume of 1.0 ft^3 of this solution is required in the vessel for the reaction. Determine the amount of time needed for the required volume to be collected. State your result in hours.

First note that the value of the flow rate has four significant digits, and the required volume has two significant digits. The two-digit value is the least precise of these, so our result must be rounded to two significant digits. But to avoid rounding error, we must work with values having at least three significant digits (one more than we need) until the very end.

One of the volumes in this problem is in μL and the other is in ft^3 . I begin by converting the required volume from ft^3 to μL so our volumes all have the same units. We have no conversion factor that goes directly from ft^3 to μL , so we must use a chain of conversion factors that we know or have available. Since we are dealing with relatively small volumes based on length units, the main conversion from USCS units to SI units is the inch to centimeter factor of $1 \text{ in} = 2.54 \text{ cm}$. This factor is exact and should be committed to memory. If we were starting with gallons instead of ft^3 , we might use the conversion $1 \text{ gal} = 3.785 \text{ L}$, although this factor is not exact.

We first convert ft^3 to in^3 , then from in^3 to cm^3 , then from cm^3 to L, and finally from L to μL .

$$1.0 \text{ ft}^3 \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{1 \text{ L}}{1000 \text{ cm}^3} \cdot \frac{1 \times 10^6 \mu\text{L}}{1 \text{ L}} = 28,300,000 \mu\text{L}$$

This result has three significant digits. We are keeping one extra digit during the intermediate calculations. It is not incorrect to write down all the digits your calculator shows. But it *is* pointless. A person who writes all the digits regardless of whether they are needed simply shows that he or she doesn't understand significant digits. Those extra digits are meaningless.

Notice that all the conversion factors used in the calculation above are exact; none of them are approximations. Since they are all exact, they play no role in limiting the significant digits in our result. Note that if you must use a conversion factor that is approximate, you should make sure the precision of the value in the conversion factor is at least as high as the precision in your data (two significant digits, in this case). That way your conversion factor does not limit the precision of your result. If, for some reason, you do not have a conversion factor with as many significant digits as your least precise measurement, then the precision of your result must match the precision of the conversion factor. The least precise value in the entire calculation always governs the precision in your result.

Now we compute the time required by dividing the required volume by the flow rate.

$$t = \frac{28,300,000 \mu\text{L}}{56.75 \frac{\mu\text{L}}{\text{s}}} = 499,000 \text{ s}$$

This value also has three significant digits—one more than we need. We now convert this value from seconds to hours as the problem statement requires.

$$499,000 \text{ s} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 139 \text{ hr}$$

Finally, we need to round the result to the required two significant digits. The second non-zero digit from the left (3) is in the tens place, so we round to the nearest ten.

$$t = 140 \text{ hr}$$



Case 4

The fourth case of rules for significant digits also applies to the calculations you perform with measurements. In Case 4, we deal with addition and subtraction.

The rule for addition and subtraction is completely different from the rule for multiplication and division. When performing addition, it is not the number of significant digits that governs the precision of the result. Instead, it is the *place value of the last digit that is farthest to the left in the numbers being added* that governs the precision of the result. This rule is quite wordy and is best illustrated by an example. Consider the following addition problem:

$$\begin{array}{r} 13.65 \\ 1.9017 \\ + 1,387.069 \\ \hline 1,402.62 \end{array}$$

Of the three values being added, 13.65 has digits out to the hundredths place, 1.9017 has digits out to the ten thousandths place, and 1,387.069 has digits out to the thousandths place. Looking at the final digits of these three, you can see that the final digit farthest to the left is the 5 in 13.65, which is in the hundredths place. This is the digit that governs the final digit of the result. There can be no digits to the right of the hundredths place in the result. The justification for this rule is that one of our measurements is precise only to the nearest hundredth, even though the other two are precise to the nearest thousandth or ten thousandth. We are going to add these values together, and one of them is precise only to the nearest hundredth. It makes no sense to have a result that is precise to a place more precise than that, so hundredths are the limit of the precision in the result.

Correctly performing addition problems in science (where nearly everything is a measurement) requires that you determine the place value governing the precision of your result, perform the addition, then round the result. In the above example, the sum is 1,4602.6207. Rounding this value to the hundredths place gives 1,4602.62.

Going back to Example 1.7, we saw the following equation for converting a temperature from kelvins to degrees Celsius:

$$T_K = T_C + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

The two values we are adding are 37.0, which has digits out to the tenths place, and 273.15, which has digits out to the hundredths place. The final digit in 37.0 is the one farther to the left, so it governs the final digit we can have in the sum. The final digit in 37.0 is in the tenths place, so the final digit in the sum must also be in the tenths place. Adding the values gives 310.15 K. Rounding this to the tenths place gives us 310.2 K for our result.

1.4 Other Important Math Skills

1.4.1 Scientific Notation

No doubt you have studied *scientific notation* in your math classes. However, beginning in high school, scientific notation is used all the time in scientific study. Knowing how to use scientific notation correctly—including the use of the special key found on scientific calculators for working with values in scientific notation—is very important.

Mathematical Principles

Scientific notation is a way of expressing very large or very small numbers without all the zeros, unless the zeros are *significant*. This is of enormous benefit when one is dealing with a value such as 0.0000000000001 cm (the approximate diameter of an atomic nucleus). The basic idea will be clear from a few examples.

Let's say we have the value 3,750,000. This number is the same as 3.75 million, which can be written as $3.75 \times 1,000,000$. Now, 1,000,000 itself can be written as 10^6 (which means one followed by six zeros), so our original number can be expressed equivalently as 3.75×10^6 . This expression is in scientific notation. The number in front, the stem, is always written with one digit followed by the decimal and the other digits. The multiplied 10 raised to a power has the effect of moving the decimal over as many places as necessary to recreate our original number.

As a second example, the current population of earth is about 7,200,000,000, or 7.2 billion. One billion has nine zeros, so it can be written as 10^9 . So we can express the population of earth in scientific notation as 7.2×10^9 .

When dealing with extremely small numbers such as 0.000000016, the process is the same, except the power on the 10 is negative. The easiest way to think of it is to visually count how many places the decimal in the value has to be moved over to get 1.6. To get 1.6, the decimal has to be moved to the right 8 places, so we write our original value in scientific notation as 1.6×10^{-8} .

Using Scientific Notation with a Scientific Calculator

All scientific calculators have a key for entering values in scientific notation. This key is labeled **EE** or **EXP** on most calculators, but others use a different label.¹ It is very common for those new to scientific calculators to use this key incorrectly and obtain incorrect results. So read carefully as I outline the procedure.

The whole point of using the **EE** key is to make keying in the value as quick and error free as possible. *When using the scientific notation key to enter a value, you do not press the **×** key, nor do you enter the 10.* The scientific calculator is designed to reduce all this key entry, and the potential for error, by use of the scientific notation key. You only enter the stem of the value and the power on the ten and let the calculator do the rest.

Here's how. To enter a value, simply enter the digits and decimal in the stem of the number. Then hit the **EE** key, and then enter the power on the ten. The value is then in the calculator and you may do with it whatever you need to. As an example, to multiply the value 7.2×10^9 by 25 using a standard scientific calculator, the sequence of key strokes is as follows:

7.2 **EE** 9 **×** 25 **=**

Notice that between the stem and the power, the only key pushed is the **EE** key.

When entering values in scientific notation with negative powers on the 10, the **+/-** key is used before the power to make the power negative. Thus, to divide 1.6×10^{-8} by 36.17, the sequence of key strokes is:

1.6 **EE** **+/-** 8 **÷** 36.17 **=**

Again, neither the 10 nor the **×** sign that comes before it are keyed in. The **EE** key has these built in.

Students sometimes wonder why it is *incorrect* to use the **10^x** key for scientific notation. To calculate 7.2×10^9 times 25, they are tempted to enter the following:

7.2 **×** **10^x** 9 **×** 25 **=**

The problem with this approach is that sometimes it works and sometimes it doesn't, and calculator users need to use key entries that *always* work. The scientific notation key (**EE**) keeps all the parts of a value in scientific notation together as one number. That is, when the **EE** key is used, a value such as 7.2×10^9 is not two separate numbers to the calculator; it is a single numerical value. But when the **×** key is manually inserted, the calculator treats the numbers separated by the **×** key as two separate values, and this can cause the calculator to use a different order of operations than you intend. For example, using the **10^x** key causes the calculator to render an incorrect answer for a calculation such as this:

$$\frac{3.0 \times 10^6}{1.5 \times 10^6}$$

¹ One infuriating model uses the extremely unfortunate label **×10^x** which looks a *lot* like **10^x**, a different key with a completely different function.

The denominator of this expression is exactly half the numerator, so the value of this fraction is obviously 2.0. But when using the 10^x key, the 1.5 and the 10^6 in the denominator are separated and treated as separate values. The calculator then performs the following calculation:

$$\frac{3.0 \times 10^6}{1.5} \times 10^6$$

This comes out to 2,000,000,000,000 (2×10^{12}), which is not the same as 2.0!

The bottom line is that the \boxed{EE} key, however it may be labeled, is the correct key to use for scientific notation.

Finally, when writing a result in scientific notation, it is not acceptable to write it using the EE notation your calculator uses. For example, your calculator might display a result as 3.14EE8 or 3.14E8, but you must write this as 3.14×10^8 .

1.4.2 Calculating Percent Difference

One of the conventional calculations in science experiments is the so-called “experimental error.” Experimental error is typically defined as the difference between a predicted value and an experimental value, expressed as a percentage of the predicted value, or

$$\text{experimental error} = \frac{|\text{predicted or accepted value} - \text{experimental value}|}{\text{predicted or accepted value}} \times 100\%$$

Although the term “experimental error” is widely used, it is a poor choice of words. When there is a mismatch between theory and experiment, the experiment may not be the source of the error. Often, it is the theory that is found wanting—this is how science advances.

I now prefer to use the phrase *percent difference* to describe the value computed by the above equation. When quantitative results are compared to quantitative predictions or accepted values, students should compute the percent difference as

$$\text{percent difference} = \frac{|\text{predicted or accepted value} - \text{experimental value}|}{\text{predicted or accepted value}} \times 100\%$$

Chapter 1 Exercises

For all exercises, note that physical constants and unit conversion factors are found in Tables A.2 and A.3 of Appendix A.

SECTION 1.1

1. Write a paragraph distinguishing between matter and mass.
2. Distinguish between base units and derived units in the SI system of units and give three examples of each.
3. Describe the advantages the SI system has over the USCS system for scientific work.
4. Why does the SI system use prefixes on the units of measure?

5. Re-express the quantities in the following table using only a single numerical digit followed by an SI unit symbol, with a metric prefix if necessary. Example: 5 thousand liters = 5 kL

- | | | |
|----------------------|------------------------------|----------------------------|
| a. 8 pascals | b. 5 hundredths of a meter | c. 3 million amperes |
| d. 2 thousand meters | e. 4 thousandths of a second | f. 6 thousand newtons |
| g. 8 thousand grams | h. 7 millionths of a liter | i. 1 thousandth of a joule |

6. Re-write the quantities in the following table by writing out the unit names without symbols. Example: 5 km = 5 kilometers

- | | | |
|----------------------|-------------|--------------|
| a. 14 m ³ | b. 164.1 kg | c. 250 MPa |
| d. 16.533 ms | e. 160 kA | f. 19.55 cL |
| g. 31.11 μJ | h. 2300 K | i. 13.0 mmol |

SECTION 1.2

7. Why must equations be used instead of conversion factors for most temperature unit conversions?
8. Perform the USCS unit conversions required in the following table. (Note: The answers in the back of the book are given with the correct number of significant digits.)

Convert This Quantity	Into These Units	Convert This Quantity	Into These Units
a. 12.55 ft	yd	b. 0.44556 mi	ft
c. 147.55 in	ft	d. 55.08 gal	ft ³
e. 934 ft ³	in ³	f. 739.22 ft ³ /s	gal/hr
g. 12.4 yr	hr	h. 51,083 in	mi
i. 14,560.77 gal/hr	qt/s	j. 15.90 mi/dy	in/hr

9. Perform the SI/metric unit conversions required in the following table.

Convert This Quantity	Into These Units	Convert This Quantity	Into These Units
a. 35.4 mm	m	b. 76.991 mL	μL
c. 34.44 cm ³	L	d. 6.33 g/cm ²	kg/m ²
e. 9.35 m/s ²	mm/ms ²	f. 542.2 mJ/s	J/s
g. 56.6 μs	ms	h. 44.19 mL	cm ³
i. 532 nm	μm	j. 96,963,000 mL/ms	m ³ /s
k. 295.6 cL	μL	l. 0.007873 m ³	mL
m. 8,750 mm ²	m ²	n. 87.1 cm/s ²	m/s ²
o. 15.75 kg/m ³	g/cm ³	p. 0.875 km	m
q. 16,056 MPa	kPa	r. 7,845 μA	mA

10. Reproduce the following table on your own paper and fill in the empty cells.

	°F	°C	K		°F	°C	K
a.	431.1			b.		-56.1	
c.			16.0	d.	0.0 (exact)		
e.		-77.0		f.			4,002
g.	-32.0			h.		65.25	
i.		1,958		j.			998.0

SECTION 1.3

11. Distinguish between accuracy and precision.

12. Describe the measurements you would obtain from an instrument that was very precise but not very accurate.

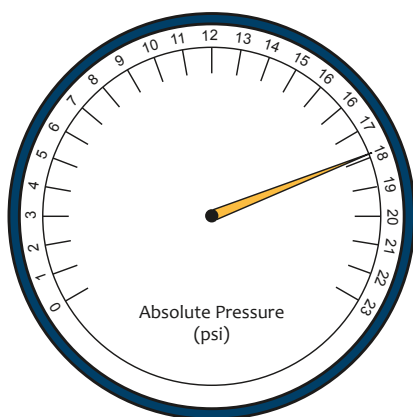
13. Which is more important on the speedometer of a car—accuracy or precision?

14. Explain why accuracy is important on a heart rate monitor but precision is not.

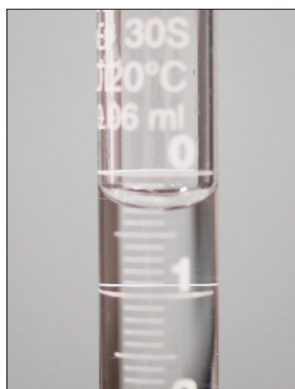
15. Sometimes we want high accuracy in a measurement, but are not too concerned about high precision. Sometimes we want both high accuracy and high precision. Explain why no one wants low accuracy and high precision.

16. On the package of a digital stopwatch I once purchased was the phrase: “1/100th second accuracy.” The stopwatch readings in seconds contained two decimal places, but the values the stopwatch actually displayed were spaced 0.03 seconds apart. Thus, it could read 12.31 s, 12.34 s, 12.37 s, etc. Comment on the accuracy and precision of this stopwatch with respect to the claim on the package.

17. Using the correct number of significant digits and the correct units of measure, record the measurements represented by the following instruments.



(a)



(b)

(c) **01320**
gallons

(d) **010**
wind speed
(mph)

SECTION 1.4

18. Using the correct number of significant digits, compute the percent difference for the experimental results in each of the following cases:

- a. A scientist measures the masses of three compounds resulting from a certain chemical reaction. Her measurements are 0.234 g, 1.678 g, and 4.446 g. Her calculations predict that the reaction results in masses of 0.239 g, 1.688 g, and 4.678 g, respectively. Determine the percent difference for each of the three compounds.
 - b. A student measures the density of aluminum and finds it to be 2.81 g/cm^3 . The accepted density value for this alloy is 2.72 g/cm^3 .
 - c. According to the Periodic Table of the Elements, the atomic mass of carbon is 12.011 g/mol. A calculation from experimental data results in a figure of 12.0117 g/mol.
 - d. The predicted yields for the products of certain chemical reaction are 23.4 kg of compound A and 2.21 kg of compound B. Careful measurements of the masses of the compounds produced indicate masses of 21.610 kg for compound A and 1.995 kg of compound B.
19. Perform each of the unit conversions indicated in the table below. Express each result using the correct number of significant digits. Where possible and appropriate, express your result in both standard notation and scientific notation. (Note: By possible, I refer to the fact that sometimes a result can only be expressed with the correct number of significant digits if it is written in scientific notation, such as a value of 100 with two or three significant digits. By appropriate, I refer to the fact that it is silly to write a value with a very large number of zeros. Such values should always be expressed in scientific notation. It is also silly to use scientific notation to express a value such as 3 or 4.1. Such values should only be expressed in standard notation.)

Convert This Quantity	Into These Units
a. 1,737 km (radius of the earth's moon)	ft
b. 2.20 g (mass of a single peanut m&m)	kg
c. 591 mL (volume of a typical water bottle)	μL
d. $7 \times 10^8 \text{ m}$ (radius of the sun)	mi
e. $1.616 \times 10^{-35} \text{ m}$ (Planck length, a fundamentally small length)	ft
f. 750 cm^3 (size of the engine in my old motorcycle)	m^3
g. $2.9979 \times 10^8 \text{ m/s}$ (speed of light in a vacuum)	mi/hr
h. 168 hr (one week)	s
i. $5,570 \text{ kg/m}^3$ (average density of the earth)	g/cm^3
j. 45 gps (gal/sec, flow rate of Mississippi River at the source)	m^3/min
k. $600,000 \text{ ft}^3/\text{s}$ (flow rate of Mississippi River at New Orleans)	L/hr
l. 5,200 mL (volume of blood in a typical man's body)	m^3
m. $5.65 \times 10^2 \text{ mm}^2$ (area of a postage stamp)	in^2
n. 32.16 ft/s^2 (acceleration of gravity, or one "g")	m/s^2
o. $10.6 \mu\text{m}$ (wavelength of light from a CO_2 laser)	in
p. 1.1056 g/mL (density of heavy water)	kg/m^3

Convert This Quantity	Into These Units
q. 13.6 g/cm ³ (density of liquid mercury metal)	mg/m ³
r. 93,000,000 mi (distance from earth to the sun)	cm
s. 65 mph (typical highway speed limit)	m/s
t. 633 nm (wavelength of light from a red laser)	in
u. 5.015% of the speed of light (see item g, or Table A.2)	mph
v. 6.01 kJ/mol (molar heat of fusion of water)	J/mol
w. 32.1 bar (pressure in saltwater at 318 m, free diving record depth)	psi
x. 0.116 nm (radius of a sodium atom)	cm
y. 6.54×10^{-24} cm ³ (volume of a sodium atom)	in ³
z. 0.385 J/(g·K) (specific heat capacity of copper)	J/(mg·K)
aa. 370 mL (volume of a soft drink can)	ft ³
ab. 268,581 mi ² (land area of Texas)	mm ²
ac. 50,200 mi ² /yr (current rate of global deforestation) ²	ft ² /s

² As a science educator who believes that we should take care of our planet, I note that items (ab) and (ac) above indicate that every 5.4 years, we lose an area of forest the size of Texas. Texas is a big place. Think about it.

Chapter 2

Atoms and Substances



Some substances pose a minimal threat to human health and safety. Others need to be labeled and classified so we know about the risks they pose. If you examine the packaging on a shipment of chemicals or look on the back of a tank truck on the highway, you may find the Fire Diamond, a symbolic representation of the various hazards associated with the substance inside. The numerical values in each of the three colored zones range from 0 to 4, with 4 representing the most extreme hazard. The blue region pertains to health, the red to flammability, the yellow to instability or reactivity, and the white to special notices. In the sample symbol above, the numerical codes specify the following:

Blue: Health 3—Short exposure could cause serious temporary or moderate residual injury. (Example: chlorine)

Red: Flammability 2—Must be moderately heated before ignition can occur (Example: diesel fuel)

Yellow: Instability/reactivity 1—Normally stable, but can become unstable at elevated temperature and pressure (Example: alcohol)

White: Special notices “W bar”—Reacts with water in an unusual or dangerous manner (Example: sodium)

Objectives for Chapter 2

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

SECTION 2.1

1. Define and describe *atom* and *molecule*.
2. State the five points of John Dalton's atomic model.
3. Write brief descriptions of J.J. Thomson's cathode ray tube experiment, Robert Millikan's oil drop experiment, and Ernest Rutherford's gold foil experiment. In your descriptions, include definitions for the terms *cathode ray* and *alpha particle*.
4. Describe the atomic models proposed by J.J. Thomson and Ernest Rutherford.

SECTION 2.2

5. Define *pure substance*, *element*, *compound*, *mixture*, *heterogeneous mixture*, and *homogeneous mixture* and give several examples of each.
6. Define *suspension* and *colloidal dispersion* and give several examples of each.
7. Explain the Tyndall effect and how it can be used to identify a mixture as colloidal.
8. Describe the two basic types of structures atoms form when bonding together.
9. Distinguish between compounds and mixtures.
10. Define and distinguish between *physical properties*, *chemical properties*, *physical changes*, and *chemical changes*. Give several examples of each.
11. Define and give examples of the terms *malleable* and *ductile*.

SECTION 2.3

12. Define *isotope*, *nuclide*, and *atomic mass*.
13. Given isotope mass and abundance data, calculate the atomic mass of an element.
14. Given the periodic table, determine the number of protons, electrons, and neutrons in the atoms of a given nuclide.
15. Define the *unified atomic mass unit*, u.

SECTION 2.4

16. Use the density equation to calculate the density, volume, or mass of a substance.
17. Define the *mole*.
18. State the Avogadro constant to four digits of precision.
19. Define *molar mass*, *formula mass*, and *molecular mass*.
20. Calculate the molar mass, formula mass, or molecular mass of a compound or molecule.
21. Calculate the mass in grams of a given mole quantity of a compound or molecule, or vice versa.
22. Calculate the number of atoms or molecules in a given quantity of substance.
23. Calculate the gram masses of an atom or molecule of a given pure substance.

2.1 Atoms and Molecules

2.1.1 Atomic Facts

We begin this chapter with a summary of the basic facts about atoms and molecules. Much of this information you probably already know.

As we saw in the previous chapter, all matter is made of atoms, the smallest basic units matter is composed of. An atom of a given element is the smallest unit of matter that possesses all the properties of that element.

Atoms are almost entirely empty space. Each atom has an incredibly tiny nucleus in the center containing all the atom's protons and neutrons. Since the protons and the neutrons are in the nucleus, they are collectively called *nucleons*. The masses of protons and neutrons are very nearly the same, although the neutron mass is slightly greater. Each proton and neutron has nearly 2,000 times the mass of an electron, so the nucleus of an atom contains practically all the atom's mass. Outside the nucleus is a weird sort of cloud surrounding the nucleus containing the atom's electrons.

We address the details about electrons in the next chapter, but here is a brief preview. The electron cloud consists of different *orbitals* where the electrons are contained. Electrons are sorted into the atomic orbitals according to the amount of energy they have. For an electron to be in a specific orbital means the electron has a certain amount of energy—no more, no less.

I wrote above that atoms are almost entirely empty space because the nucleus is incredibly small compared to the overall size of the atom with its electron cloud. It's quite easy for us to pass over that remark without pausing to consider what it means. To help visualize the meaning, consider the athletic stadium pictured in Figure 2.1. Using this stadium as an enlarged atomic model, the electrons in their orbitals would be zipping around in the region where the red seating sections are in the stadium. Each electron in this enormous atomic model is far smaller than the period at the end of this sentence. The atomic nucleus containing the protons and neutrons is located at the center of the playing field, and is the size of a pinhead. And what fills all the vast space inside the atom? Nothing, not even air, since air, of course, is also made of atoms. The inside of an atom is empty space.

Returning to our discussion of atomic facts, one of the fundamental physical properties of the subatomic particles is *electric charge*. Neutrons have no electric charge. They are electrically neutral, hence their name. Protons and electrons each contain exactly the same amount of charge, but the

charge on protons is positive and the charge on electrons is negative. If an atom or molecule has no net electric charge, it contains equal numbers of protons and electrons.

Atoms are significantly smaller than the wavelengths of light, which means light does not reflect off atoms and there is no way to see them. The same is true of *molecules*. Molecules are clusters of atoms chemically bonded together. When atoms of different elements are bonded together in a molecule they form a compound, which we discuss later in this chapter. But sometimes atoms of the same element bond together in molecules, as illustrated in Figure 2.2. Oxygen and chlorine are two of the *diatomic gases* that form molecules consisting of a pair of atoms



Figure 2.1. The head of a pin at center field in a stadium is analogous to the nucleus in the center of an atom.

chemically bound together. Hydrogen, nitrogen, and fluorine also exist naturally as diatomic gases.

2.1.2 The History of Atomic Models

The story of atomic theory starts back with the ancient Greeks. As we look at how the contemporary model of the atom developed, we also hit on some of the great milestones in the history of chemistry and physics along the way.

In the 5th century BC, the Greek philosopher Democritus proposed that everything was made of tiny, indivisible particles. Our word atom comes from the Greek word *atomos*, meaning “indivisible.” Democritus’ idea was that the properties of substances were due to characteristics of the atoms they are made from. So atoms of metals were supposedly hard and strong, atoms of water were assumed to be wet and slippery, and so on. At this same time, there were various views about what the most basic substances—that is, the elements—were. One of the most common views was that there are four elements—earth, air, water, and fire—and that everything is composed of these.

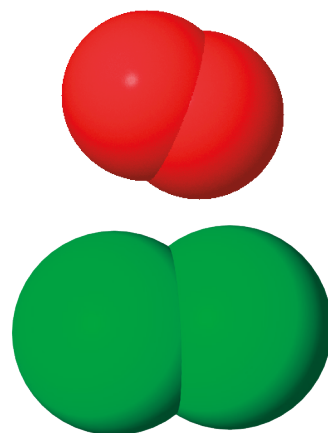


Figure 2.2. Space-filling models of the diatomic oxygen (top) and chlorine molecules.



Figure 2.3. English scientist John Dalton (1766–1844).

are made of atoms, but we now know that atoms are not indivisible. This is now obvious, since atoms themselves are composed of protons, neutrons, and electrons. The second point is correct in every respect but one. Except for the number of neutrons in the nucleus, every atom of a given element is identical. However, we now know that

Not much real chemistry went on for a very long time.

During the medieval period, of course, there were the alchemists, who sought to transform lead and other materials into gold. But this cannot be done by the methods available to them, so their efforts were not successful.

But in the 17th century, things started changing as scientists became interested in experimental research. The goal of the scientists described here was to figure out what the fundamental constituents of matter are. This meant figuring out how atoms are put together, what the basic elements are, and understanding what is going on when various chemical reactions take place. The nature of earth, air, fire, and water was under intense scrutiny over the next 200 years.

In 1803, English scientist John Dalton (Figure 2.3) produced the first scientific model of the atom. Dalton’s atomic model is based on five main points, listed in Table 2.1.

The impressive thing about Dalton’s atomic theory is that even today the last three of these points are regarded as correct, and the first two are at least partially correct. On the first point, it is still scientifically factual that all substances

1. All substances are composed of tiny, indivisible substances called atoms.
2. All atoms of the same substance are identical.
3. Atoms of different elements have different weights.
4. Atoms combine in whole-number ratios to form compounds.
5. Atoms are neither created nor destroyed in chemical reactions.

Table 2.1. The five tenets of Dalton’s 1803 atomic model.



Figure 2.4. English scientist Joseph John (J.J.) Thomson (1856–1940).

atoms of the same element can vary in the number of neutrons they have in the nucleus. These varieties of nuclei are called *isotopes*, a topic we return to soon.

After Dalton, the next breakthrough in our understanding of atomic structure came from English scientist J.J. Thomson (Figure 2.4). Thomson worked at the Cavendish Laboratory in Cambridge, England. In 1897, he conducted a series of landmark experiments that revealed the existence of electrons. Because of his work, he won the Nobel Prize in Physics in 1906. A photograph of the *cathode ray tube* Thomson used for his work is shown in Figure 2.5.

Thomson placed electrodes from a high-voltage electrical source inside a very elegantly made, sealed-glass vacuum tube. This apparatus can generate a so-called *cathode ray* from the negative electrode (1), called the *cathode*, to the positive one, called the *anode* (2). A cathode ray is simply a beam of electrons, but this was not known at the time. The anode inside Thomson's vacuum tube had a hole in it for some of the electrons to escape through, which created a beam of cathode rays heading toward the other end of the tube (5).

Thomson placed the electrodes of another voltage source in-

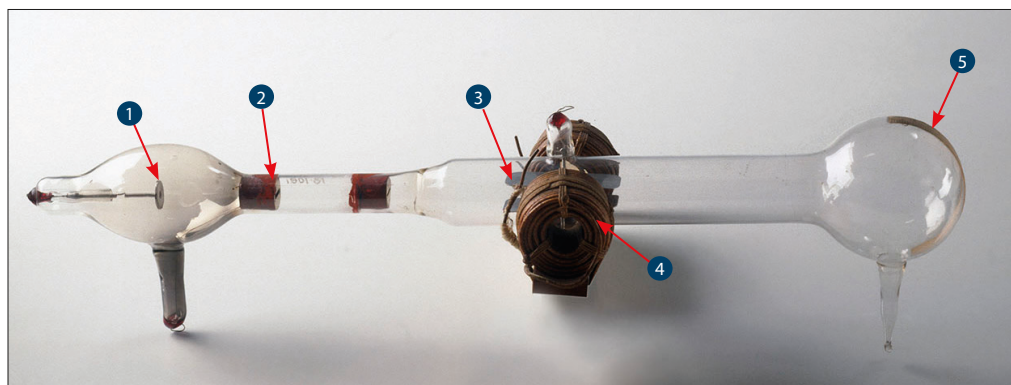


Figure 2.5. J.J. Thomson's cathode ray tube.

side the tube (3), above and below the cathode ray, and discovered that the beam of electrons deflected when this voltage was turned on. He also placed magnetic coils on the sides of the tube (4) and discovered that the electrons also deflected as they passed through the magnetic field produced by the coils. The deflection of the beam toward the positive electrode led Thomson to theorize that the beam was composed of negatively charged particles, which he called “corpuscles.” (The name *electron* was first used a few years later by a different scientist.) By trying out many different arrangements of cathode ray tubes, Thomson confirmed that the ray was negatively charged. Then using the scale

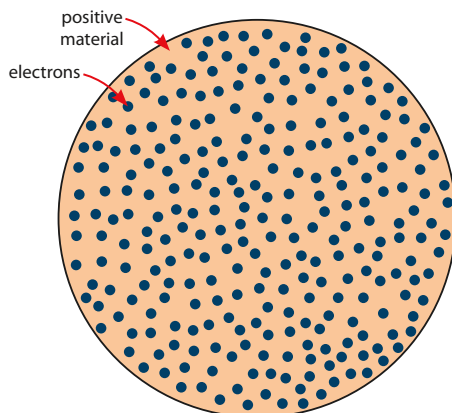


Figure 2.6. Thomson's plum pudding model.