Carl H. Snyder

THE EXTRAORDINARY CHEMISTRY OF ORDINARY THINGS

SECOND EDITION

Carl H. Snyder
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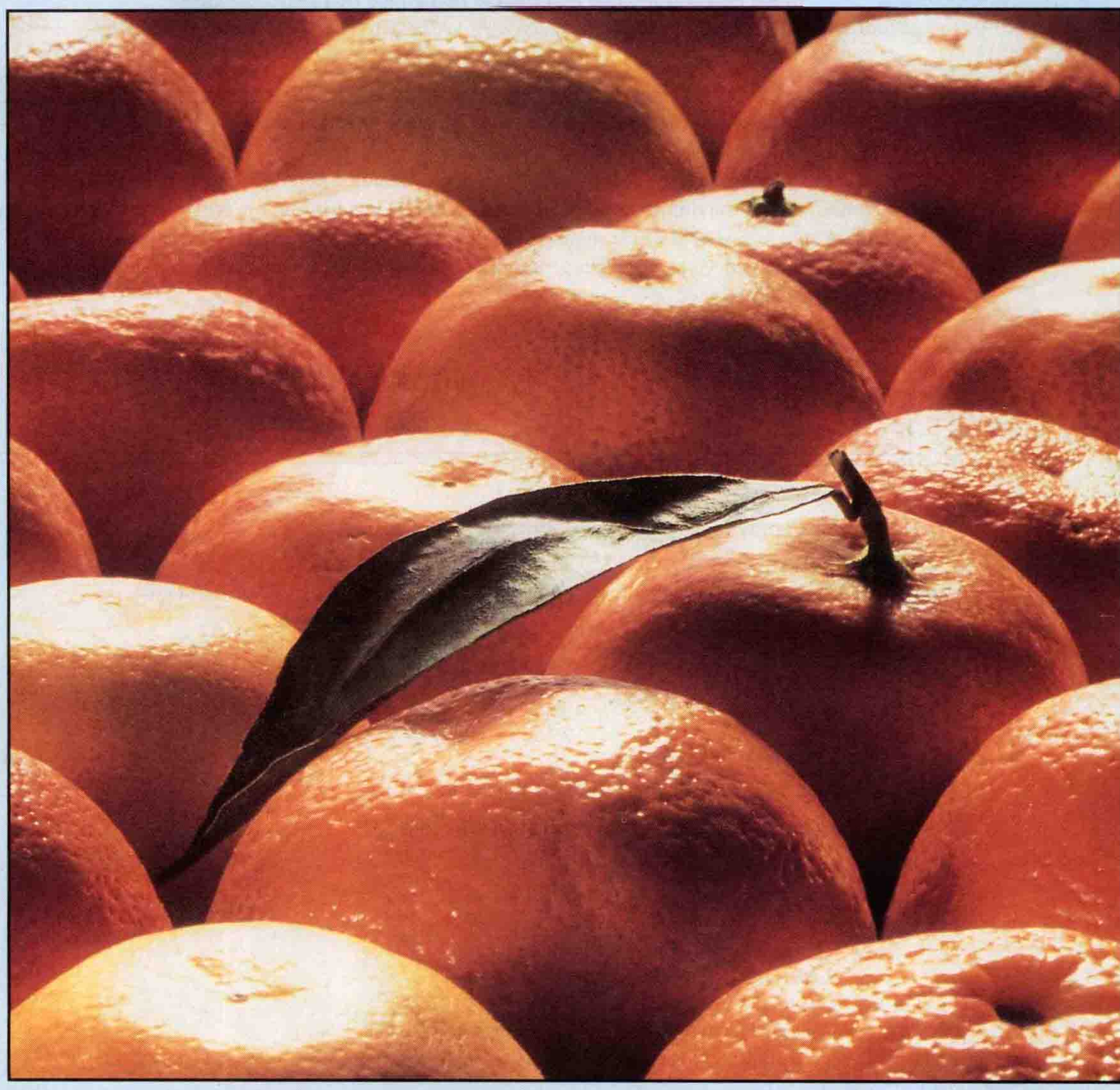
The Extraordinary Chemistry of Ordinary Things

Second Edition



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For Jean

Preface

This second edition of *The Extraordinary Chemistry of Ordinary Things* continues and expands on the central thesis of the first edition: Since we live our daily lives immersed in chemicals, the most effective way to teach and to learn chemistry is by examining the goods and substances that we use in our daily lives and that affect us and our environment.

The first edition grew from a course originally titled "Consumer Chemistry." In the early 1970s, at a time of severe inflation, it occurred to me that such a course combining chemistry with consumerism and directed at non-science students was badly needed. I developed this as a one-credit, one-semester course in which I taught chemistry through its applications to consumer products, by using consumer products to illustrate chemical principles. Each area—chemistry and consumerism—reinforced the other in examinations of gasoline and petroleum, detergents, foods and food additives, plastics, and the like.

With time, the course expanded beyond consumerism and the more common of our consumer products, but without losing sight of either. Although radioactive substances, for example, aren't commonly classified as consumer products, we do encounter them as consumers of medical care. Although ozone isn't itself a consumer product, our use of the gasoline engine and of chlorofluorocarbons affects both the undesirable *generation* of ozone (in the air we breathe) and the undesirable *destruction* of ozone (in the stratospheric ozone layer).

The course evolved into an examination of the chemistry of the substances of our everyday world, from the banal to the contentious, from table salt to perception-altering drugs, from drinking water to nuclear power. It evolved to include questions of safety; of the impact of the use and disposal of consumer goods on the quality of the environment; of the meaning and measurement of pollution; and particularly of the ambiguity of terms like "good" and "bad" as they are applied to chemicals. It evolved to bring students themselves into the realm of chemistry, not only to demonstrate that we ourselves are constructed physically of chemicals but also to show that we can and must have the power of choice in how we use the chemicals of our universe. Appropriate choices require wisdom, and wisdom is founded on knowledge.

The course changed in other ways as well. It grew from a one-credit, one-semester offering into a two-semester sequence of two three-credit courses, acceptable toward the science requirement for graduation. With these changes I have tried always to remain true to my original goal: teaching chemistry through illustrations taken from the common substances, objects, and processes of the world around us.

Objectives

Throughout this evolution, the objectives of this approach have remained constant. They are to teach chemistry:

- In the context of chemistry as an experimental science.
- In the context of the ordinary things of our everyday lives, and some that aren't quite so ordinary but that nonetheless can and do affect our lives.
- In the context of the larger realm of science, drawing on chemical principles and examples to illustrate the workings of science as a whole and the scientific method.
- In the context of the need for science literacy to enable all, scientist and nonscientist alike, to make reasoned judgments on societal issues that are founded on the processes and fruits of science in general and chemistry in particular.

Chemistry, an Experimental Science

Many of us teach chemistry as we know it to be, with an understanding shaped by many years, even decades, of study. We see chemistry as a coherent, rational whole, and we transmit this model of the chemical universe to our students. Yet I have found, and I suspect many others would agree, that merely transmitting this model is insufficient and unsatisfying to both the teacher and the student. It's important to teach not only the coherent model of the universe that the science of chemistry presents to us, but to demonstrate why we are forced to accept this particular model as preferable, more useful, and more intellectually satisfying than any other.

I use the word "forced" because the model of the world that chemistry presents to us is one that we are absolutely and unconditionally required to accept. We are forced to this particular model by our contact with physical reality, by our tests of physical reality, by the questions we frame as we test this real universe experimentally, and by the answers we receive from our experimental tests. Chemistry is, above and beyond all else, an experimental science.

We are forced to mold the universe into one particular intellectual construct because of our commitment to the scientific method and its experimental approach to knowledge. To teach chemistry, I am convinced, requires teaching the broad outlines of the scientific method, explicitly or implicitly. We have no choice, for example, but to acknowledge that atoms, subatomic particles, and chemical bonds do actually exist. But *why* are we forced to this view of the world? This is what students must come to understand if they are to learn chemistry in its richest context: why we are forced to see the physical universe as we do.

We accept the reality of atoms and all the other structures and concepts of chemistry because we have no other rational choice. Our experimental tests of our universe, through the scientific method, lead us to them and only to them. Let us then give our students hard, physical, real, demonstrable evidence that what we are about to tell them is, indeed, true chemistry, real chemistry. Let us show them in lectures and in textbooks that what we tell them is true not because we say it is, but because they see it is.

The Magic of Chemistry

Some of the chapters start with what appear to be demonstrations of magic. Household bleach, for example, mysteriously makes colors appear rather than disappear, exhaled breath mysteriously causes colors to change, we mysteriously "squeeze" air out of a glass bottle, and so on. As each of these chapters unfolds, the "magic" is explained as the operation of a chemical principle, and the "magic" is seen to be no more than the rational operation of the laws of the universe. The "magic" is transformed into "chemistry" as the student comes to understand how the chemical universe about us works. In this way we can illustrate to students one of the most important contributions that science in general and chemistry in particular have made to the development of our civilization: the conversion of superstition into understanding, of fearsome magic into useful science, all through the acquisition of knowledge. After all, *the difference between "magic" and "science" is knowledge.*

Demonstrations

To emphasize the experimental basis of chemistry, all but two chapters begin with a demonstration or an action of some sort that the students themselves can perform or may already have performed with simple equipment and common substances. The first chapter, for example, begins with an illustration of the electrical conductivity of salt water and the nonconductivity of sugar water that employs table sugar, table salt, and a simple flashlight. The materials of the demonstration are about as common and ordinary as any we can find. Yet we see, at the first moment of contact with this realm of chemistry, that there's something demonstrably different about salt and sugar, other than mere taste, and that this difference *forces* us to the concept of ions. Ions are real not because we say they are, but because students *see* that they are.

Each of these initial demonstrations leads us to observations and conclusions about the chemistry of (mostly) ordinary things that we will soon run across again, as textbook chemistry, somewhere within the chapter. These can be used as lecture demonstrations, but they are more than that. All the demonstrations can be repeated by students, using common household goods. (Some chapters, like the two on nuclear chemistry, are better left without descriptions of hands-on experiments.) To integrate this experimental approach with the principles covered in the chapter, a newly introduced section near the end of each chapter reviews the initial demonstration in light of chemical principles covered within the chapter.

Sequence of Chapters

The sequence of chapters allows the text to be used for either a one- or a two-semester course. Of the 21 chapters, the first 11 cover most topics considered to be fundamental to the science of chemistry. The first 3 are introductory, dealing with atoms, ions, molecules, elements, compounds, and the periodic table. The next 2 deal with the nucleus. Chapter 4 covers nuclear chemistry in a roughly chronological narrative spanning the half-century from the discovery of radioactivity to the development and use of the atomic bomb. Chapter 5 takes us to the peaceful uses of nuclear energy. The remaining chapters of this first set cover the arithmetic of chemistry (with emphasis on the mathematics of pollution), organic chemistry (featuring hydrocarbons and their use as fuels), acids and bases, electrochemistry, and the states of matter, especially gases.

With applications intimately tied to concepts throughout, there is no sacrifice of applications if the book is used in a one-semester course. The later chapters continue the integration of principles and applications with examinations of the chemistry of soaps and detergents; chemicals as environmental pollutants; the chemicals of food; chemical hazards and the question of safety; polymers and plastics; personal care products; and medicines and drugs, especially the effects of chemicals on our perceptions of the world we live in. Any of these, in whole or in part, can be included in a one-semester course with little or no modification.

New to the Second Edition

Several revisions have reduced the number of chapters to 21, yet have expanded the scope of the subject matter presented in the first edition. The environmental chapter that was available as a shrink-wrapped supplement to the first edition has become a new Chapter 13 in this edition. A discussion of DNA has been added to the chapter on proteins (Chapter 16), in which the genetic code, viewed as the information of heredity written in a chemical script, directs the strategy of protein synthesis.

Even with these additions I have maintained a reasonable length by eliminating peripheral discussions of foods and nutrition and by condensing four of the chapters of the first edition into two. • The two chapters of the first edition that covered nutritional aspects of energy and the chemistry of food triglycerides have been combined into Chapter 14, which presents triglycerides as both our highest density providers of calories (our dietary fats and oils) and the material that stores our principal bodily reserves of energy (body fat). • Two other chapters of the first edition, on micronutrients and food additives, have been combined into Chapter 17, which examines both micronutrients (dietary minerals and vitamins) and food additives as important and sometimes controversial, yet minor components of our foods when compared with the macronutrients.

Beyond additions and condensations, a reorganization of topics has produced a more cohesive and more effective presentation. • A new sequence of chapters allows earlier examinations of the mathematics of chemistry and pollution, and defers electrochemistry until after introductions to organic chemistry and to acids and bases. • A reorganization of topics integrates both acid rain and the threat to the ozone layer into the chapter on chemical pollutants (Chapter 13). • The transfer of the discussion of the states of matter into the chapter on gases and the gas laws (Chapter 11) allows the discussion of soaps and detergents (Chapter 12) to focus more clearly on surfactant chemistry.

• A newly included list of additional readings appears at the end of each chapter. Coordinated to icons appearing at selected topics within the chapter, the entries of this list lead to more extensive or more detailed discussions of the topics covered. • New sections in each chapter review the opening demonstration as an illustration of the chemical principles covered.

• Many new and revised worked examples, and both end-of-section and end-of-chapter questions and problems, provide a large assortment of study aids for students.

Pedagogical Structure

Virtually every section is followed by a question designed to induce the student to reflect on or review the material just covered. Exercises at the end of each chapter are divided into three categories: (1) review, written for a straightforward reexamination of the factual material of the chapter; (2) mathematical, for those who wish to emphasize the mathematical aspects of chemistry; and (3) thought-provoking. Exercises in this last category sometimes have no "right" answer but are intended to stimulate thought about the interconnection of chemistry, society, and individual values. Answers are provided for virtually all of the even-numbered problems of the first two categories, and several of the third category.

Many of the chapters, especially the earlier ones, contain worked examples to ease the student's way through the more difficult concepts. Other characteristics of the presentation include the introduction of definitions, concepts, symbols, and the like, largely on a need-to-know basis. It seems to me to make more sense to explain and describe the world about us as we encounter it, rather than to start by defining and categorizing ideas well before we need to use them. It's also clear that I like etymologies. I've found that students learn technical terms more easily if they know where they came from. I have other preferences that I'm unaware of, and I'm sure they show up in the book here and there, beneficially I hope.

Supplements

An innovative package of supplements to accompany *The Extraordinary Chemistry of Ordinary Things 2/E* is available to assist both the instructor and the student.

1. **Study Guide,** by Ann Ratcliffe of Oklahoma State University and David Dever of Macon College. This Guide is an invaluable tool for the student, containing unusual, illustrative scenarios as well as the more traditional study guide features such as chapter overviews and solutions to the in-text questions. Also included are worked-out solutions to the problems in the text along with additional exercises of the same nature and level of difficulty.
2. **Laboratory Manual,** by Bruce Richardson of Highline Community College and Thomas Chasteen of Sam Houston State University. Twenty-five laboratory exercises are included in this manual, all written in a clear, concise, and unintimidating fashion. The themes emphasized in the Laboratory Manual closely parallel those of the text, incorporating experiments with both consumer and environmental applications. The Laboratory Manual *Chemistry: The Experience,* by Ann Ratcliffe is also appropriate for use with this text.
3. **Instructor's Manual,** by John Thompson of Texas A&M University—Kingsville. In addition to chapter overviews, learning objectives, chapter

outlines, discussion and critical thinking questions, postscripts/chapter lead-ins, key terms, and additional class demonstrations for each chapter in the text, the Manual also contains background information and suggestions for using the *The Extraordinary Chemistry of Ordinary Things* videotape.

4. **Test Bank.** Written by the text author, the Test Bank contains over 1000 multiple-choice questions.
5. **Computerized Test Bank.** IBM and Macintosh versions of the entire Test Bank are available with full editing features to help you customize your tests.
6. **Full-Color Overhead Transparencies.** Over 100 full-color illustrations are provided in a form suitable for projection in the classroom.
7. **Videotape.** Over 15 experiments are demonstrated by the author on this videotape. A few selected chapter-opening experiments are brought to life; other demonstrations illustrate other pertinent chapter material.

Acknowledgments

As in the first edition, I want to acknowledge the contribution of my former departmental chairman, Harry P. Schultz. I wrote down what appears on the following pages mostly because of Harry. After my initial suggestion, many years ago, that we introduce a course for nonscientists, he gave me unreserved support, encouragement, and recognition. He also asked, repeatedly, "Why don't you put all this down on paper?" He asked once too often, so I did. With his enthusiastic support for the course, and, I must add, for our students as well, and his repeated urgings that I put it all on paper, this book owes its existence more to Harry Schultz than to any other person. Without Harry neither the course nor the book would exist.

From a more personal point of view, I thank my wife Jean once again for her patience and unfailing good cheer, both of which have eased the effort of a work like this. Especially with this second edition, she has graciously accepted this textbook as an additional member of our family, one that requires its own, unique kind of care, feeding, and particular attention.

I'm particularly grateful to all those at John Wiley & Sons who brought this book into being, and especially to Rachel Nelson and Nedah Rose, whose combined vision of a second edition has inspired and informed this revision. Both this edition and I have benefited from the assistance and encouragement generously given by others at John Wiley & Sons, especially Katharine Rubin and Suzanne Ingrao, Production; Ed Starr, Illustration; Joan Kalkut, Supplements; Pete Noa, Design; Stella Kupferberg and Mary Ann Price, Photos; Kaye Pace, Editorial. I would also like to thank Barbara Burke of California State Polytechnic University, Pomona, for contributing the additional readings at the end of each chapter.

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Carl H. Snyder

Featured in This Book

Demonstrations...

To emphasize the experimental basis of chemistry, every chapter but the two on nuclear chemistry begins with a demonstration or activity of some sort that students can perform with simple equipment and common household goods. In the spirit of the experimental approach, results of the demonstration are explained in the context of the principles of chemistry developed within the chapter.

Demonstration Demonstra

Atoms and Paper Clips

To gain some insight into the modern view of an atom, try this demonstration. It draws an analogy between a pile of paper clips and a small amount of one of the 109 known elements, a bar of gold for example. Although it's only a simple analogy, one that can't be stretched very far, it does illustrate an important point about atoms.

Place a pile of 15 to 20 paper clips on a level surface and imagine a bar of pure gold sitting next to them (Fig. 2.1). Now divide the pile of paper clips roughly in half. Subdivide one of the new, smaller piles in half and repeat the process again and again until you are down to a "pile" that consists of a single paper clip. In our analogy, that single paper clip represents an "atom" of paper clips, the smallest part of the original pile of paper clips that you can still identify as a paper clip. In dividing and subdividing the original pile of paper clips into ever smaller piles you finally came to the smallest part of the original pile that you can still identify as a paper clip. By analogy, that simple paper clip represents an "atom."

Now imagine that you perform the same operation with the bar of gold. Picture yourself dividing the (imaginary) bar in half again and again, as you did with the paper clips, until you finally reach the smallest particle of the bar that

Piles of paper clips, individual paper clips, and fragments of a paper clip. The fragments no longer represent a paper clip.

20

Perspect

2.9 Atoms and Paper Clips Revisited

In the opening demonstration we drew an analogy between atoms and paper clips. We saw that in some ways an individual paper clip is related to a handful of paper clips in the same way an individual atom is related to a few grams of an element. Each is the smallest unit that we can identify with the group it represents. We'll now extend this analogy a bit by applying it to atomic weights and isotopes.

Like elements, paper clips vary quite a bit in their characteristics. For example, there are triangular, rectangular, and the familiar oval clips. Each of these can be made of plastic or metal and can come in any of a variety of colors (Fig. 2.14). Although there aren't as many different kinds of paper clips as there are elements, the variety of paper clips that are available offers a rough parallel to the variety of the chemical elements. Moreover, in addition to variations in shape, composition, and color, paper clips come in different sizes.

For example, we can use a small, metal paper clip with rounded ends, or one of the very same kind but a bit larger (and therefore a bit heavier). In a general sort of way, the two clips of Figure 2.15 could illustrate isotopes, both representing the same kind of paper clip but one heavier than the other. (The heavier clip is larger, too, but that's not part of the analogy; isotopic atoms are about the same size.)

If we drop one of these larger and heavier clips into a pile of 6700 of the smaller clips, we see a parallel to the distribution of hydrogen atoms in the universe, with one heavier deuterium for every 6700 lighter protiums. The addition of the heavier clip raises the *average* weight of all the paper clips just a little above the weight of one of the more plentiful (smaller) clips. Similarly, the *average* atomic weight of all the hydrogen in the universe is just a bit greater than the mass of a protium atom.

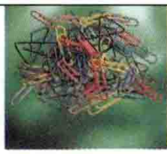


Figure 2.14 The many varieties of paper clips reflect the varieties of elements and their atoms.

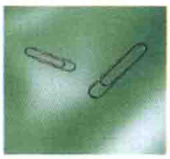


Figure 2.15 Two paper clips of the same kind, but different sizes can be used as models for atomic isotopes. Atomic isotopes represent the same element, but have different masses.

The first 10 elements, those with atomic numbers from 1 through 10, are about as diverse a group of substances as we can imagine. We'll take a quick look at some of their more interesting facets and then, in Chapter 3, examine the ways in which many of these and other elements combine to form some of the ordinary and extraordinary substances of our world.

Hydrogen, the first in this series, was originally identified as a distinct substance in 1766 by Henry Cavendish, a British scientist. In examining this newly identified gas, Cavendish was struck by its ability to form water when it burns in air. This observation led Antoine Lavoisier—a French chemist who was a member of the French aristocracy and a contemporary of Cavendish—to suggest the name "hydrogen," which comes from Greek words meaning "produces water." (In 1794, shortly after he made this suggestion, Lavoisier was executed at the guillotine by the leaders of the French Revolution.)

The isotope deuterium was first identified in 1931 by the American chemist Harold C. Urey, who received the 1934 Nobel Prize in Chemistry for this discovery. In a real sense hydrogen is one of the most important of all the elements: Our own bodies contain more atoms of hydrogen than of all the other elements combined.

Perspective

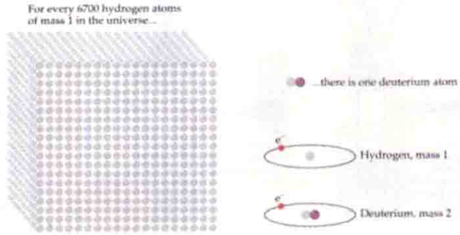
The First 10 Elements

...Revisited

The results of each chapter-opening demonstration are revisited near the end of each chapter, providing a wrap-up of the concepts involved and the conclusions drawn in the experiment. The student is now able to see how the results were obtained given the information developed throughout the chapter.

30 Chapter Two • Atoms and Elements

Figure 2.11 Relative abundance of deuterium.



in nature (Fig. 2.11). Because the ratio of the two isotopes overwhelmingly favors the atom of mass number 1, the word *hydrogen* commonly refers either to the naturally occurring mixture of the two or simply to the isotope of mass number 1. Where confusion can occur, the term *protium* is used for the isotope of mass number 1. Furthermore, the symbol *D* represents specifically an atom of deuterium.

Naturally occurring hydrogen consists almost entirely of only the two isotopes, protium and deuterium. But it's possible to manufacture a third isotope, *tritium*, by adding a second neutron to the nucleus. Tritium, with a nucleus containing one proton and two neutrons, has a mass number of 3 and an atomic number of 1. Tritium is used along with deuterium to produce the explosive force of the hydrogen bomb. The hydrogen of the universe consists of about 99.985% protium, 0.015% deuterium, and just a trace of tritium.

Question

As we've just seen, an atom with a nucleus consisting of *one proton* and *two neutrons* is tritium, an isotope of hydrogen. Is an atom with a nucleus consisting of *one neutron* and *two protons* still another isotope of hydrogen? Explain your answer. ______________________________

2.7 Building up the Elements: Hydrogen through Neon

Although adding a neutron to an atomic nucleus increases its mass number and thereby generates a different isotope of the same element, adding a proton produces an entirely different element. (In actual practice it isn't nearly as easy to add a proton to a nucleus as it is to add a neutron. In any case, what concerns us here isn't the specific procedure we might use for the addition, but rather the consequence of adding a proton to an atomic nucleus.)

Adding protons, as we have seen, increases atomic numbers as well as mass numbers. Adding one proton to a hydrogen nucleus, for example, produces an atom of the element *helium*. Virtually all the helium atoms in the universe have two neutrons in their nuclei as well as two protons, so a helium atom's mass number is 4 and its atomic number is 2. With two positively charged protons

Questions

Virtually every section is followed by a question designed to induce the student to reflect on or review the material just covered. The questions serve as quick checks to ensure that the student comprehends what he or she has just read before moving on in the chapter.

Worked Examples

Worked examples are provided throughout the text to help ease the student through some of the more difficult concepts and work through the more quantitative aspects of chemistry. Frequently broken down into step-by-step stages, these examples serve as models for some of the end-of-chapter exercises.

...uction to Chemistry

Reacting in Ratios

Suppose we allow 10.0 g of sodium to react with an equal weight of chlorine. What is the composition of the product?

We know that a ratio of 23.0 g of sodium to 35.5 g of chlorine produces pure sodium chloride and that if either sodium or chlorine is present in excess, the product is a mixture of sodium chloride and the element that is present in excess. Since 35.5 g of chlorine reacts with a *smaller* weight of sodium (23.0 g) to produce sodium chloride, it's clear that with equal weights of the two there's an excess of sodium. Our problem, then, is to calculate just how much excess sodium is present. To find this value we multiply the 10.0 g of chlorine we're given by the ratio 23.0 g sodium/35.5 g chlorine.

$$10.0 \text{ g chlorine} \times \frac{23.0 \text{ g sodium}}{35.5 \text{ g chlorine}} = 6.5 \text{ g sodium}$$

This means that to maintain the ratio of 23.0 g sodium to 35.5 g chlorine, 6.5 g of sodium must react with the 10.0 g of chlorine provided in this illustration. The reaction consumes all of the chlorine present (10.0 g) and 6.5 g of the original 10.0 g of sodium to produce 16.5 g of sodium chloride (from the 10.0 g of chlorine and 6.5 g of sodium), with

$$\begin{array}{r} 10.0 \text{ g sodium} \\ \underline{-6.5 \text{ g sodium}} \\ 3.5 \text{ g sodium left over} \end{array}$$

The product, then, is composed of 16.5 g sodium chloride and 3.5 g of sodium. (You might want to consult the appendices at the end of the book for help with this example and others throughout the text.)

1.5 Light Bulbs, Salt, and Sugar Revisited

Let's pause here to review what we've just done and seen. We began our study of chemistry with the observation that pure water is a very poor conductor of electricity. We've seen that dissolving *table sugar* in water doesn't improve its ability to conduct an electric current but that dissolving *table salt* in water does increase this ability, quite dramatically, as demonstrated by the light bulb of Figure 1.3.

Looking back on the opening demonstration we can recognize that we used table salt, table sugar, some water, and a flashlight to answer a sequence of questions:

- Does *pure water* conduct an electric current?
- Does adding *table sugar* to water produce any change in water's ability to conduct an electric current?

Running Glossary

Key terms are boldfaced and defined both in the text and in the margin, helping students identify and recall the most important concepts.

Icons

Icons are placed in the margin of the text next to passages that pertain to specific additional readings from a mix of current and classic journal articles and books. The citations are listed at the end of each chapter.

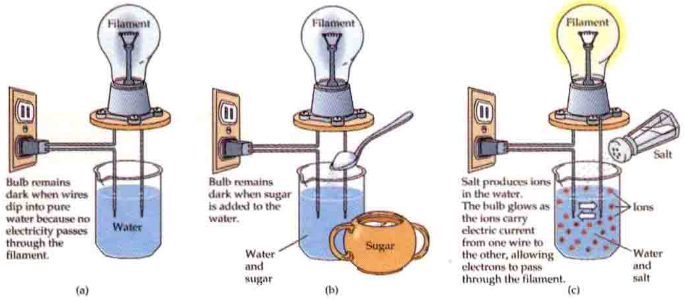


Figure 1.5

Actually, we rarely find pure water in the world around us. Most samples of water we encounter contain dissolved salts, much like the table salt of Figure 1.3, and therefore most of the water we find in our everyday lives does conduct electricity. Dissolved salts are partly responsible for the common observation that (impure) water often conducts electricity, and sometimes very well indeed.

Sucrose doesn't change water's ability to conduct electricity, but sodium chloride does. In fact, a solution of sodium chloride in water (or molten sodium chloride, at a very high temperature) is very effective at conducting an electric current. Clearly, adding the sodium chloride introduces something into the water that allows electricity to flow from one wire to the other. Adding sucrose does not.

The simplest explanation for all this and other, related observations is that sodium chloride is made up of electrically charged particles that can move about in water and can transport an electric current through water much as electrons transport current through the wires of an electrical circuit. On the other hand, we have to conclude that sucrose is *not* made up of electrically charged particles (Figure 1.5).

An **ion** is an atom or a group of atoms that carries an electrical charge.

These small particles of sodium chloride and other electrolytes, each bearing a negative or positive electric charge, are called **ions.** As we saw in Section 1.1, *an ion is an atom or a group of atoms that carries an electrical charge.* To understand more fully what ions are—and how it is that sodium chloride can carry an electric current although sucrose cannot—we have to understand what atoms are (and, later, what molecules are). We'll have more to say about atoms and ions in Chapter 2.

An **anion** is a negatively charged ion; a **cation** is a positively charged ion.

Ion is a term derived from a Greek word meaning "to go." In 1834 the English physicist Michael Faraday used the word to describe chemical particles that move to one electrical pole or another. He divided ions into two categories: **anions** are the negatively charged chemical particles that move to the positive electrical pole (the anode); **cations** are the positively charged chemical particles that move to the negative electrical pole (the cathode). We still use the

22. Suppose we could combine one electron with one proton to form a single, new subatomic particle. What mass, in amu, would the resulting subatomic particle have? What electrical charge would it carry? What known subatomic particle would the resulting particle be equivalent to?
23. Suppose someone discovered a particle that consisted of a single neutron surrounded by a shell containing a single electron. Would you classify this as an atom? Would you classify it as an ion? Would it represent a new element? Explain your answers.
24. Suppose you have two spheres, one made of lead and one made of cork. They are standing next to each other at some spot on the surface of the earth. At that location each weighs 1 kg. Compare their masses at that location. Does one have a greater mass than the other? If so, which has the greater mass? Now move both spheres to one particular loc moon and again co have a greater mas weigh more than which? Finally, leave the cork sphere on the moon and move the lead sphere back to its original location on earth. Again, compare both their masses and their weights.
25. The description of the atom that we have used in this chapter resembles in some ways our own solar system. What part of our solar system corresponds to the nucleus of an atom? What part corresponds to the electron shells surrounding the nucleus? What are some of the other similarities between the structure of the atom, as described in this chapter, and our own solar system? What are some of the more obvious differences?

Additional Reading

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Friend, J. Newton, 1961, 2nd edition. *Man and the Chemical Elements: An Authentic Account of the Successive Discovery and Utilization of the Elements, from the Earliest Times to the Nuclear Age*. New York: Scribner.

Partington, James R. 1960. 3rd edition. *A Short History of Chemistry*. New York: Harper: 357–360.

Ringnes, Vivi. 1989. Origins of the Names of Chemical Elements. *Journal of Chemical Education*. 66(9): 731–738.

Weinberg, Steven. 1983. *The Discovery of Subatomic Particles*. New York: Scientific American Library.

Additional Readings

A list of suggested readings from various periodicals and trade books is included at the end of every chapter. The readings range from classic biographical accounts to current articles on issues affecting our modern existence. The readings are referenced throughout the text with icons placed in the margins.

Perspectives

Every chapter closes with a Perspective that takes a look at the implications and consequences of the facts and concepts presented throughout the chapter. The Perspectives exemplify the basic approach of the text by presenting informed choices to the student, analyzing risks and benefits, discussing the experimental basis of science, and revealing the chemistry all around us.

2.9 Atoms and Paper Clips Revisited

In the opening demonstration we drew an analogy between atoms and paper clips. We saw that in some ways an individual paper clip is related to a handful of paper clips in the same way an individual atom is related to a few grams of an element. Each is the smallest unit that we can identify with the group it represents. We'll now extend this analogy a bit by applying it to atomic weights and isotopes.

Like elements, paper clips vary quite a bit in their characteristics. For example, there are triangular, rectangular, and the familiar oval clips. Each of these can be made of plastic or metal and can come in any of a variety of colors (Fig. 2.14). Although there aren't as many different kinds of paper clips as there are elements, the variety of paper clips that are available offers a rough parallel to the variety of the chemical elements. Moreover, in addition to variations in shape, composition, and color, paper clips come in different sizes.

For example, we can use a small, metal paper clip with rounded ends, or one of the very same kind but a bit larger (and therefore a bit heavier). In a general sort of way, the two clips of Figure 2.15 could illustrate isotopes, both representing the same kind of paper clip but one heavier than the other. (The heavier clip is larger, too, but that's not part of the analogy; isotopic atoms are about the same size.)

If we drop one of these larger and heavier clips into a pile of 6700 of the smaller clips, we see a parallel to the distribution of hydrogen atoms in the universe, with one heavier deuterium for every 6700 lighter protiums. The addition of the heavier clip raises the *average* weight of all the paper clips just a little above the weight of one of the more plentiful (smaller) clips. Similarly, the *average* atomic weight of all the hydrogen in the universe is just a bit greater than the mass of a protium atom.

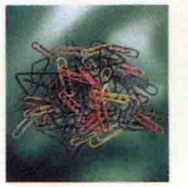

Figure 2.14 The many varieties of paper clips reflect the varieties of elements and their atoms.

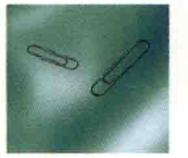

Figure 2.15 Two paper clips of the same kind, but different sizes can be used as models for atomic isotopes. Atomic isotopes represent the same element, but have different masses.

Perspective

The First 10 Elements

The first 10 elements, those with atomic numbers from 1 through 10, are about as diverse a group of substances as we can imagine. We'll take a quick look at some of their more interesting facets and then, in Chapter 3, examine the ways in which many of these and other elements combine to form some of the ordinary and extraordinary substances of our world.

Hydrogen, the first in this series, was originally identified as a distinct substance in 1766 by Henry Cavendish, a British scientist. In examining this newly identified gas, Cavendish was struck by its ability to form water when it burns in air. This observation led Antoine Lavoisier—a French chemist who was a member of the French aristocracy and a contemporary of Cavendish—to suggest the name "hydrogen," which comes from Greek words meaning "produces water." (In 1794, shortly after he made this suggestion, Lavoisier was executed at the guillotine by the leaders of the French Revolution.)

The isotope deuterium was first identified in 1931 by the American chemist Harold C. Urey, who received the 1934 Nobel Prize in Chemistry for this discovery. In a real sense hydrogen is one of the most important of all the elements: Our own bodies contain more atoms of hydrogen than of all the other elements combined.

4. Question 3 mentioned "the first 10 elements." What does this phrase refer to?
5. How many different elements are currently known to exist?
6. In terms of the number of atoms present, hydrogen is the most abundant element in the universe and also in our bodies. What is the second most abundant element in the universe, again in terms of the number of atoms present? What is the second most abundant element in our bodies?
7. Given the atomic number and the mass number of an atom, how do we determine the number of protons in the nucleus? The number of neutrons in the nucleus? The number of electrons in the surrounding shells?
8. (a) Name the three isotopes of the element whose atomic number is 1. (b) What collective name do we give to a mixture of these three isotopes when they are present in the same ratios as in the universe as a whole?
9. What is a name used for the isotope of hydrogen indicated by ${}^{3}_{1}H$?
10. What chemical symbol is used as the equivalent of ${}^{2}_{1}H$?
11. Name and give the chemical symbols for the elements with the first 10 atomic numbers.
12. Atoms of what element are used to define the atomic mass unit?
13. Can any isotope of any element have a mass number of zero? Explain your answer.
14. (a) What would remain if we removed an electron from a hydrogen atom? What charge would this particle bear? What would its mass be? (b) How do you think we could convert a sodium atom into one of the sodium cations discussed in Chapter 1?
15. Name
    a. three elements that have only a single electron in their outermost quantum shell
    b. two elements that have exactly two electrons in their outermost quantum shell
    c. three elements with filled outermost quantum shells
    d. one element that has only a single electron in its innermost quantum shell
16. (a) Which element was found on the sun before it was found on earth? (b) Which element has a name that indicates its presence in many acids or sour-tasting substances? (c) How many years passed between the identification of hydrogen as an element and the discovery of its isotope deuterium? (d) Which of gases? (e) Which is th 10 elements?

A LITTLE ARITHMETIC AND OTHER QUANTITATIVE PUZZLES

17. Give the number of protons and the number of neutrons in the nucleus of each of the following atoms and the number of electrons in the first and second quantum shells: (a) ${}^{13}_{6}C$; (b) ${}^{10}_{5}B$; (c) ${}^{18}_{8}O$; (d) ${}^{40}_{19}K$; (e) ${}^{6}_{3}Li$.
18. (a) If we could unite a boron atom with a lithium atom to form a single new atom, what element would it represent? (b) If we double the number of protons in an atom of carbon, with a mass number of 12, what element would the resulting atom represent? (c) If we double the mass number of the atom of carbon described in part (b) but do not change its atomic number, what element would the resulting atom represent?
19. Section 2.5 contains the statement that hydrogen atoms "make up two-thirds of all atoms in water, but just over 11% of the water's weight." Given that there are twice as many hydrogen atoms in water as there are oxygen atoms, and that virtually all the hydrogen atoms in water are ${}^{1}_{1}H$ and virtually all the oxygen atoms are ${}^{16}_{8}O$, how do you explain this apparent discrepancy?
20. Using 3×10^{-8} cm as the diameter of a gold atom and 3.3×10^{-22} g as its weight, calculate the weight of 1 cm^3 of gold. (You can start by finding how many gold atoms fit on a line 1 cm long; see Sec. 2.2.) How does your calculated value compare with the measured density of gold, 19.3 g/cm^3? Suggest some factors that might account for the difference.

THINK, SPECULATE, REFLECT, AND PONDER

21. Your answers to parts (a) and (b) of this exercise do not depend on whether the elements named exist as isotopes or on which isotope you choose to consider. (a) Lithium, sodium, and potassium are all metals that react with water to liberate hydrogen gas. What, if anything, do atoms of each of these metals have in common? (b) Helium is an unreactive gas and neon is a gas of extremely low reactivity. What, if anything, do their atoms have in common?

Exercises

Exercises at the end of each chapter are divided into three categories:

For Review are written for a straightforward reexamination of the factual material presented in the chapter. These include fill-in-the-blank questions that require students to choose among a number of answers and to come up with some answers that are not provided for them; matching questions; and straightforward review questions that provide students with the opportunity to build their own chapter summaries.

A Little Arithmetic and Other Quantitative Puzzles are mathematical exercises that drill the student on the more quantitative aspects of the material.

Think, Speculate, Reflect, and Ponder are thought-provoking problems that are intended to stimulate thought about the interconnections among chemistry, society, and individual values. Sometimes there are no right answers to these problems.

Demonstration Guide

The following is a complete list of the chapter-opening demonstrations and the concepts they illustrate. As you can see, many of the same household substances are used in a number of the demonstrations. The experiments can be conducted in class or at home and help set the stage for the topic of the chapter and its role in our everyday lives.

Chapter 1 Enlightenment From a Flashlight—See what the differences in the ways salt and sugar conduct (or don't conduct) electricity in water tell us about their composition.

Chapter 2 Atoms and Paper Clips—Gain insight into the modern view of an atom using a pile of paper clips.

Chapter 3 Scouring Pads and Kitchen Magnets—Demonstrate the changes in the properties of an element as it enters into a compound, using a kitchen scouring pad.

Chapter 6 The Glass Where Pollution Begins—Understand how we count chemical particles, what concentrations are, and the importance of measuring levels of pollution, using drinking glasses, a ruler, a marking pen, and some salt and sugar.

Chapter 7 A Candle Burning in a Beaker: Energy From Hydrocarbons—Observe the chemistry of a burning candle to demonstrate the power of hydrocarbons.

Chapter 8 Petroleum and Strong Tea—See how valuable products are obtained from petroleum by distilling pure water from strong tea.

Chapter 9 Breath with the Strength of Red Cabbage—Design a breath test to demonstrate the acidity of exhaled breath, using red cabbage, household ammonia, vinegar, and water.

Chapter 10 Galvanized Tacks, Drugstore Iodine, and Household Bleach—Show how galvanized tacks, iodine, and household bleach produce color changes as a result of electron transfers.

Chapter 11 Squeezing Air Out of a Bottle—Demonstrate the properties of gases through a parlor trick that involves the illusion of squeezing the air out of a bottle with your bare hands.

Chapter 12 With Nerves as Steady as a Chemical Bond—Illustrate the chemical phenomenon of surface tension by seemingly floating a thumbtack on the surface of a glass of water.

Chapter 13 How to Generate Air Pollution in Your Own Environment—Show how acid rain forms using red cabbage leaves, a pot of water, matches, a candle, and a paper towel.

Chapter 14 Warming Up with Work—Demonstrate the principle that doing work generates heat, simply by rubbing your hands together.

Chapter 15 How to Tell a Potato from an Apple, the Hard Way—Distinguish between a potato and an apple using only iodine and the sense of sight.

Chapter 16 How to Turn an Egg White White—Turn an egg white white without cooking it to demonstrate the process of denaturation.

Chapter 17 The Power of Vitamin C—Visualize Vitamin C's chemical prowess by demonstrating its ease of oxidation, using Vitamin C tablets, iodine, and household bleach.

Chapter 18 Spreading Poison All Around—Deliberately introduce a toxic chemical into the environment simply by spraying an insecticide onto the floor.

Chapter 19 Plastics, Water, and Rubbing Alcohol—Separate plastics of different densities using a mixture of rubbing alcohol and water.

Chapter 20 The Wave of Chemistry—See how a home permanent changes the shape of protein molecules in hair.

Chapter 21 Perception, Reality, and Chemicals—Present illusions that test your concept of reality.