GENERAL CHEMISTRY

JOHN D. MAYS

Second Edition

General Chemistry

A Mastery-Oriented Curriculum

Second Edition

NOVARE
SCIENCE & MATH

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Preface

For Teachers

Thank you for taking the time to begin here at the beginning. This chemistry text is quite different from other high school texts covering similar material. So to ensure that your experience with this text is successful, it is important to be aware of its unique features and the logic behind its structure. In this Preface for Teachers, I address the following points:

1. Student Audience and Preparedness

2. Our Emphasis on Mastery, Integration, and Kingdom

3. Recommendations for Teaching With This Text

4. Laboratory Work and Lab Reports

Several important resources are described in the pages that follow. For your convenience, they are all listed here. Please visit our website (novarescienceandmath.com) or contact us for more information.

- *Introductory Physics* (John D. Mays, 2013, 2016)

- *Accelerated Studies in Physics and Chemistry* (John D. Mays, 2012)

- *Chemistry for Accelerated Students* (John D. Mays, 2014)

- *Teaching Science so that Students Learn Science: A Paradigm for Christian Schools* (John D. Mays, 2009)

- *Chemistry Experiments for High School* (Christina Swan and John D. Mays, 2014)

- *Chemistry Experiments for High School at Home* (Christina Swan and John D. Mays, 2014)

- *Science for Every Teacher* (John D. Mays, 2013)

- *The Student Lab Report Handbook* (John D. Mays, 2009, 2014)

1. Student Audience and Preparedness

This text is designed for grade-level students. This means the typical student using this text is a junior in high school concurrently taking Algebra 2. In my view, students need to be at or above Algebra 2 in mathematics in order to study chemistry. Computations with pH involve logarithms and power functions, and computations with the gas laws involve a lot of algebraic manipulation. For these reasons, chemistry is usually undertaken when students are in Algebra 2 or later.

I promote a physics-first approach to secondary science programming. According to this program, students take an introductory physics course in 9th grade. The obvious choice for 10th grade science is then biology, followed by chemistry in 11th grade. My text *Introductory Physics* is specifically designed for this purpose. The benefits of using *Introductory Physics* in 9th grade followed two years later by *General Chemistry* in 11th grade are *significant*. Much of the content covered in the first two chapters of the present text is also addressed in *Introductory Physics*. Students who have had this background should find the first two chapters of *General Chemistry* to be a breeze—it will mostly be review for them. One of the results of this is that the study of

chemistry, which can be heart-breakingly difficult for some students, becomes much more accessible.

Contrast this preparedness with the more typical scenario of students arriving in chemistry with very weak unit conversion skills and knowing virtually nothing about significant digits. In this situation, the challenge of chemistry is much greater simply because of students' lack of appropriate background in basic math skills. Like the present text, *Introductory Physics* is a mastery-oriented text. By using it, students get repeated, sustained practice at unit conversions throughout the year. With such extensive practice in 9th grade, brushing up those skills two years later in chemistry comes easily. *Introductory Physics* also gives students practice in the use of significant digits, an added bonus. (However, the addition rule for computations does not appear in *Introductory Physics*. That rule is introduced in Chapter 1 of the present text.)

Another important benefit of the physics-first approach is that students have a solid conceptual understanding of mass, potential energy, kinetic energy, and electric charge, all of which are hugely important in the study of chemistry. And we could add the topic of DC circuits as well, since the application section in the redox chapter is all about electrochemical cells and batteries.

It is important to mention here that we have a different set of texts for use by accelerated students. Accelerated students complete Algebra 1 in 8th grade. After Geometry in 9th grade they take Algebra 2 in 10th grade. For these students, we recommend *Accelerated Studies in Physics and Chemistry* for the 9th-grade course and *Chemistry for Accelerated Students* for the 10th-grade chemistry course. For details on these texts please visit our website at novarescience-andmath.com

The benefits of the physics-first program and the logic behind the separation of science and math into separate pathways for grade-level and accelerated students are discussed in much more detail in past articles of the Novare Newsletter. These are available on our website under the Newsletter tab. In particular, you may wish to read "Sequencing the Upper School Science Curriculum" in Volume 4 Number 3, and "Stratifying Math Students" in Volume 2 Number 1.

2. Our Emphasis on Mastery, Integration, and Kingdom

Mastery The norm for classes in contemporary schools is what I call the *Cram–Pass–Forget cycle*. Students cram for tests, pass them, and then forget most of what they crammed in just a few weeks. Teachers across the nation know well what this looks like because they see it day after day. This destructive cycle is a waste of time for teachers and students. Instead of cramming and forgetting, students should *learn*, *master*, and *retain* what they have learned. Realizing this in the classroom requires both the teacher and the students to make significant changes in the ways they approach the tasks of teaching, testing, practicing, and studying. For textbooks, a mastery orientation requires that the curriculum cover fewer topics, and cover them more deeply. When students engage with content, their efforts should be directed at comprehension, practice, and retention of important concepts, not mere exposure to masses of material they will not remember. Methods that promote the *Cram–Pass–Forget cycle* must be eliminated through pedagogy based on more effective methods and a more sensible classroom dynamic.

In a mastery-oriented learning environment, students notice that the workload is demanding, and initially some of them resist. But all the students learn far more than they ever have in a course. They remember the material for years, and most of them—even those who aren't much interested in school—come to recognize the value of such a learning experience. Some recognize it immediately, some only after they go off to college and find themselves ahead of their peers because of things they remember from years before.

Integration For a variety of complex reasons, effective science instruction requires a number of content areas to be deeply embedded in the science curriculum—content areas often con-

sidered tangential and not represented adequately in curriculum materials. These content areas include the scientific process of modeling, basic epistemology of science (e.g., understanding the difference between scientific claims and truth claims), mathematics, scientific history, and English language usage. Typically, science classes do not place the necessary emphasis on these areas, and as a result science courses fail to present students with a coherent and effective learning experience. Science teachers need to think deeply about how their courses address these and other key areas of integration and make adjustments to curriculum, teaching methods, assessments, and coordination between science courses.

Kingdom Science and mathematics provide us with unique ways of seeing God's creative presence in the world. Bringing biblical faithfulness to science classes is not accomplished by simply folding in a few Bible verses or prayers. In fact, much more is involved. Science and math teachers need to think very broadly about how we fulfill Christ's mandate to love God with all our mind, how we teach our students effectively to engage issues, and how we perceive God's fingerprints in creation.

We also need to engage thoughtfully with the scientific claims of our day. It is not a scientific claim to say that the universe got here by itself; that is a metaphysical claim based on an atheistic worldview that Christians reject. But it is a scientific claim to say that the universe began with the Big Bang and is now 13.77 billion years old. The scientific evidence behind this claim is vast, and I believe an appropriate science text is one that teaches students how to engage such claims productively. I do not believe it is appropriate to teach students to be dismissive of claims such as this one simply because they do not line up with certain ways of interpreting Genesis. I say this as one who fully believes the Bible, who loves reading Genesis and the rest of Scripture, and who accepts the strong evidence for an old universe. (Note that I use the age of the earth issue to make a general point about science instruction, but the topic is not addressed in this text.)

My ideas about all three parts of this core philosophy are described in detail in my book *Teaching Science so that Students Learn Science*. This book, along with a more detailed description of our textbook philosophy, may be found on our website.

3. Recommendations for Teaching With This Text

In this section, I make a few remarks pertaining to some specific chapters in this text, followed by some comments about teaching for Mastery and Integration.

Regarding the chapter content, there are some sections you may wish to skip, depending on the circumstances at your school. Details are as follows:

- Chapter 1   If your students were taught from my text *Introductory Physics* in a prior course, you may find you can move through this chapter fairly quickly. Feel free to speed through it if your students have mastered the content and are ready to move on. However, mastery of unit conversions and significant digits is so important that you should cover the chapter in depth if needed.

- Chapter 2   My comments about Chapter 1 apply here as well; most of this content is covered in *Introductory Physics*. Again, move through the chapter quickly if you find that your students remember this material well. However, I suggest that at a minimum you spend some time in Section 2.2 on the types of substances. Students often forget (or are confused about) this important information.

- Chapters 3–7   These chapters contain standard material and should be covered in full.

- Chapter 8   As before, students who have studied *Introductory Physics* can probably move quickly through this material, particularly the discussion of states of matter.

Topics in this chapter that are not in *Introductory Physics* include calculations with molar heat capacity, molar heat of fusion/vaporization, and phase diagrams.

- **Chapters 9–10** These chapters contain standard material and should be covered in full.

- **Chapter 11** In this chapter (acids and bases), some teachers may prefer to stick to the Arrhenius model; others will consider Brønsted-Lowry theory indispensable. My advice is to think about your students and consider the merits of racing through material that students will not be able to master. If there is not sufficient time in your schedule to bring students to mastery, it would be better to skip the Brønsted-Lowry material and move on to pH and titrations.

- **Chapter 12** This is the redox chapter. At a minimum, students should learn to balance equations using the half-reaction method, and understand what the electrode potentials represent—how they relate to the activity series of metals and redox reactions in general. However, if time is short at the time you are covering this final chapter, you could focus on a qualitative encounter with the electrochemistry.

Now for some remarks about teaching for Mastery and Integration, beginning with Mastery.

First, students need continually to be working with previously learned material to keep old skills fresh. Included in the Chapter Exercises at the end of each chapter is a set of General Review Exercises covering material from previous chapters. Students should always work through these exercises.

Second, the teacher's assessment regimen should support the goal of students retaining previously learned material. For students at this level, I recommend the following four-part assessment regimen: (a) Do not award credit in students' grades for homework. (Do, however, require students to complete their assignments and hold them accountable for doing so.) The logic behind this principle is explained in detail in *Teaching Science so that Students Learn Science* and in my article "Putting Homework in its Place" in the March 11, 2013 issue of our newsletter, available for free at the Newsletter tab of our website. (b) Between each chapter test, administer one or two quizzes covering material presented in class during the previous week. This motivates students to stay current with their studies instead cramming the night before the test. (It also gives them direct information about how prepared they are for that part of the chapter test.) Count the average of all the quizzes for a single semester together as equivalent to one chapter test in the semester grade. (c) Use the "Standard Problems List" technique described in *Teaching Science*. Inform students at the beginning of the year that each chapter test includes problems from the Standard Problems List on material from previously covered chapters. (d) On each chapter test, allocate about 20% of the exam to material on the Standard Problems List from previously covered chapters. I have placed a Standard Problems List for Chapters 1–11 in the Preface for Students. Individual teachers may wish to modify that list to suit what the teacher feels are the most important topics to emphasize.

Additionally, as described above, you should feel free to cut some of the material from any chapter if the students in a particular class simply can't take it in fast enough. Classes are different year to year, and some groups of students can handle more than others. Teaching according to mastery principles is quite different from conventional *Cram–Pass–Forget* methods, and it is more demanding for teacher and students alike. However, the rewards are huge, which is why I constantly promote mastery methods. But teachers need to administer course content with wisdom.

Now I will make a few points about how the principle of Integration should work with the course.

First, in addition to computations, I always promote questions on tests and quizzes requiring responses written in complete sentences. In their responses, students should be required to demonstrate competence with standard English. I discourage so-called objective items requiring true/false, matching, and multiple-choice responses.

Second, the incorporation of language skills into your course will be enhanced even further if you require students to write their lab reports from scratch. I address this more in the next section.

Third, a healthy epistemology of science should pervade all science courses. The *Cycle of Scientific Enterprise* model briefly described in the Introduction is the place to start. This topic is developed at length in *Teaching Science* and in the first chapter of *Introductory Physics*. I strongly encourage all teachers to draw upon these resources and become fluent in the concepts and terminology about the nature of science and scientific knowledge. Then bring the subject up as often as possible in class. Science teachers are very busy, and there are constant pressures threatening to cause courses to lag behind in the curriculum. It is tempting to stick to the technical content and neglect teaching our students what science *is* and how science *works*. From neglect of this topic, only a small percentage of our adult population understands what the statement "science is modeling" means, or has any idea of the distinction between scientific claims and truth claims from Scripture. One of the results of this massive lack of understanding is fuel thrown on the fires of the "science-faith debate," as if one has to choose between godly faith and robust science. I argue that one does not have to make such a choice. A critical component of the science teacher's role is helping students learn to participate in a healthy, faithful, responsible dialog concerning what we know about the world from Scripture and what we know about the world from scientific inquiry.

Finally, as I state in Appendix B, Wolfgang Goethe once wrote that "the history of science *is* science." As I mention in the previous paragraph, science is about modeling nature. Our models—theories—are never perfect and never complete, and as a result they change over time. One of the best ways to help students grasp this is to get into the history of the subject—the scientists with their theories and discoveries. In my texts for younger students, I place learning objectives pertaining to the history prominently alongside those pertaining to the technical content. In this text, I do not do so. The reason is that there are quite a few historical references in this text, and I do not think it reasonable to require students to memorize a paragraph of information pertaining to every one of them. So instead of stipulating historical learning objectives, I leave it up to the individual teacher to decide how much history to include and which scientists to focus on. However, it is very important that some history is included in student learning objectives and assessments, and that the historical circumstances are related to the nature of scientific knowledge and the *Cycle of Scientific Enterprise* wherever possible.

To assist both teacher and student in managing historical content, all the scientists discussed in the text are listed in Appendix B. I know that some teachers and home school families using this text may wish for more prescriptive requirements when it comes to the history. Accordingly, I offer this proposal: have the students write a paper on the history of chemistry that summarizes the contributions to chemistry from at least six or eight scientists. Require that the contributions discussed be tied together in a unified narrative, rather than simply quoted one after the other from material in the text. Alternatively, have the students write papers in which they describe how several of the experiments or other contributions described in this text relate to the *Cycle of Scientific Enterprise* model discussed in the Introduction.

However you choose to go about it, the history of science is important. The historical material should be an important component of your course, not something students simply skip over.

4. Laboratory Work and Lab Reports

A laboratory component is essential for every high school chemistry class. Not only does a lab *practicum* give students direct knowledge and experience that are virtually impossible to obtain from a text, but the report writing component of the lab work provides a rich enhancement to the overall learning objectives for the course. As I state above, development of English language skills should be deeply integrated into our science courses. Requiring students to write lab reports from scratch—rather than by filling out blank spots in a workbook—provides a premier opportunity to do this.

High school science teachers should require that students write full-length lab reports from scratch, on a computer, five or six times per year. In *Teaching Science so that Students Learn Science*, I devote a chapter to laboratory work and lab reports. That chapter outlines assessment guidelines and learning objectives for lab reports at different grade levels throughout high school.

There is a lot that goes into writing a quality lab report, and teaching students how to do it without a guide to help is a difficult and time-consuming task. Thus, I commend to you and your students my manual, *The Student Lab Report Handbook*. I recommend that schools distribute copies at the beginning of the school year to each high school freshman. Let each student keep the book for use at home for the next four years. Students should begin learning how to write lab reports in their freshman science class and continue writing reports in all science classes throughout their high school years. Students trained this way astonish their lab instructors when they get to college, and are prepared for college in a way few students are.

Students in high school chemistry courses typically conduct at least 10 or 12 lab exercises during the year. But requiring students to write 10 or 12 lab reports from scratch in a year would be unduly burdensome. This is why I recommend that you require students to write reports from scratch five or six times during the year—three times in the fall and three times in the spring. For other experiments, use a short-form lab report in which students present data and interact with a few key questions.

To accompany this text in schools, Dr. Christina Swan and I produced *Chemistry Experiments for High School*. The experiments in the book are designed for use in a fully-equipped chemistry lab. (An alternative resource for those without a lab is discussed below.) At the end of each experiment in *Chemistry Experiments*, there are a few pages that may be used as a short-form lab report. The book is printed on perforated sheets that can be written on, removed from the book, and submitted. This provision is *not* intended to suggest that all the experiments should be documented this way. But for those that are, the book makes short-form reports easy.

For those that do not have access to a laboratory—and this includes most home-school students—we offer *Chemistry Experiments for High School at Home*. The experiments in this manual are adapted from *Chemistry Experiments for High School*, substituting alcohol burners for Bunsen burners, gravity filtration for vacuum filtration, etc. Some of the experiments do require special chemicals or apparatus—for students to have a legitimate chemistry laboratory experience there is just no way around it. But we have sought to keep these special items to a minimum.

Preface

For Students

You probably don't normally read the Preface to Teachers in your textbooks. (You may not normally read the one to students either, but I am glad you are reading this one.) However, the books I write are quite different from other textbooks you may have used in the past. So in this case, I recommend that you read the Preface to Teachers so you learn about how this book is structured and what I recommend to teachers about how to use it. In particular, it is my view that students should master and *remember* what they are taught, instead of cramming for tests and forgetting everything a few weeks later.

For this to happen, your teacher needs to teach and test in ways you may not be accustomed to. My recommendation to your teacher is that all your chapter exams include problems and questions from prior chapters. Now, that doesn't mean you need to remember every detail from every chapter. But it does mean that there are certain questions and types of problems that are considered very foundational in any chemistry course. I call these Standard Problems. My advice to your teacher is that you use a list of standard problems in your course, and that your exams always include problems from the standard problems list. At the end of this Preface, you will find the Standard Problems List I recommend to your teacher for this chemistry course.

Naturally, your teacher is the person who decides whether to use the Standard Problems List. But if you do use it, then you need to study in a way that enables you to stay current with the material on the list. Simply doing your homework each night and cramming for tests won't cut it. You must have specific study strategies that help you remember definitions and concepts from previous chapters and how to solve previously learned types problems.

Here are my recommendations for how to do that.

1. Study the Objectives List at the beginning of each chapter carefully. Make it your goal to be able to do everything on the list (that is, for the objectives that have been covered so far in class) before your quizzes and tests occur.

2. Look over the Standard Problems List regularly. Identify any item that you cannot do or cannot remember how to do (assuming the topic has already been covered in class) and follow up on it.

3. Develop, maintain, and practice flash cards for each new chapter Objectives List and each item on the Standard Problems List.

4. Read every chapter in this text at least once, preferably twice. Ideally, every time your instructor covers new material you should read the sections in this book corresponding to that material within 24 hours.

5. Go back and read the chapters in this book again when you are a month or two down the road. You will be amazed at how much easier it is to remember things when you have reread a chapter. (Besides, reading is more fun than rehearsing flash cards.)

6. When you are working on exercises involving computations, check your answers against the answers in the back of the book. Every time you get an incorrect answer, dig in and stay with the problem until you identify your mistake and obtain the correct answer. If you can't figure out a problem after 10 or 15 minutes, raise the question in class.

7. Every time you lose significant points on a quiz or test, follow up and fill in the gaps in your learning. If you didn't understand something, raise the question with your instructor. If you

forgot something, rehearse it more thoroughly until you have it down. If you failed to commit something to memory, or didn't have it in your flash cards, then add it to the cards and commit it to memory. If you were not proficient enough at one or more of the computations, look up some similar problems from the exercises or from previous quizzes and tests and practice them thoroughly.

Finally, there is a fair bit of historical material scattered around in this text. It is my view that the history of science *matters* and that students should be held accountable for learning and remembering important historical information. In Appendix B you will find a list of all the scientists discussed in the text. This appendix is there to help you as you study. Your instructor will let you know which scientists to know about and what to know about them. I mention it here so that as you study and review you will know to use Appendix B as review a tool in addition to the Standard Problems List.

Standard Problems List

Given appropriate reference materials, students should be prepared to work the following problems on any chapter test throughout the course, assuming the material has already been covered in class.

FROM CHAPTER 1

1. Use the metric system fluently.

2. Use significant digits fluently.

FROM CHAPTER 2

3. Describe and define the various types of substances.

4. Calculate density.

5. Determine molar mass or formula mass and perform unit conversions between moles and grams.

6. Determine numbers of atoms in a given quantity of substance.

FROM CHAPTER 3

7. Describe the Bohr model and explain how it is able to explain emission spectra.

8. Write electron configurations.

9. Determine percent composition, determine an empirical formula from the percent composition, and determine the molecular formula from the empirical formula and the molar mass.

FROM CHAPTER 4

10. Describe the general structure and arrangement of the periodic table.

11. Describe general trends in the periodic table pertaining to atomic size, ionization energy, and electronegativity.

12. Identify the names of the major groups of elements and other significant regions in the periodic table.

FROM CHAPTER 5

13. Name binary ionic compounds and acids and write their formulas, including those incorporating polyatomic ions.

14. Describe ionic and covalent bonding.

15. Draw Lewis structures.

FROM CHAPTER 6

16. Apply VSEPR theory to predict molecular shapes and bond angles.

17. Describe the different intermolecular forces, when they occur, and their relative strengths.

18. Balance chemical equations.

19. Perform stoichiometric calculations.

20. Determine oxidations states of pure elements and elements in compounds.

21. Use the activity series of metals to predict whether a single-replacement reaction will occur.

22. Describe the kinetic-molecular theory of gases and the Maxwell-Boltzmann velocity distribution in gases.

23. Explain the causes of surface tension and capillary action.

24. Describe the four basic states of matter.

25. Calculate the heat that must be added to or removed from a substance to change its temperature or effect a phase transition.

26. Explain the causes of evaporation and vapor pressure.

27. Describe Boyle's law, Charles' law, and Avogadro's law.

28. Solve problems using Boyle's law, Charles' law, the ideal gas law, and the density equation.

29. Distinguish between ideal gases and real gases.

30. Use Dalton's law of partial pressures to compute partial pressures, mole fractions, and the total pressure in a gas mixture.

31. Solve stoichiometric problems involving gas volumes.

32. Describe dissolution, electrolytes, and heat of solution.

33. Use solubility guidelines to classify compounds as soluble or insoluble in aqueous solution.

34. Use notions of ionization and polarity to explain the phrase "like dissolves like."

35. Calculate molarity and molality.

36. Use solubility guidelines to predict when precipitates will form.

37. Write ionic equations and net ionic equations describing precipitation reactions.

38. Describe three colligative properties of solutions.

39. Calculate boiling points and freezing points of solutions.

40. Describe two acid-base theories, including definitions for acids and bases in each.

41. Identify conjugate acid-base pairs.

42. Write ionic and net ionic equations for neutralization reactions.

43. Explain the concept of the self-ionization of water.

44. Describe the pH scale.

45. Compute pH, pOH, $[H_3O^+]$, and $[OH^-]$ from each other and from concentration data.

46. Explain the purpose of titration.

47. Determine $[H_3O^+]$ and $[OH^-]$ from titration data.

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Introduction

What is Chemistry All About?
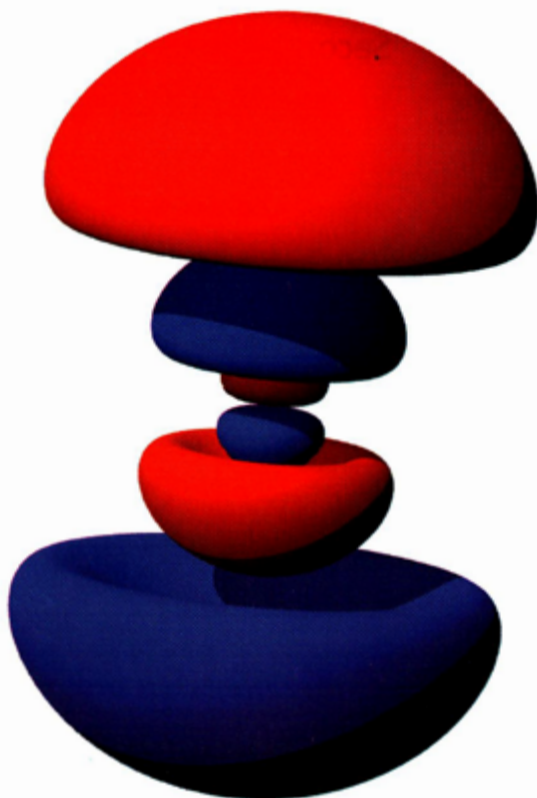
This computer model depicts three of the *orbitals* available in atoms for holding electrons. Shown are one of the three *p* orbitals in each of three different *subshells*. The inner pair of orbitals can hold two of the highest energy electrons for elements 5 through 10 in the Periodic Table of the 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.

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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.

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

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.
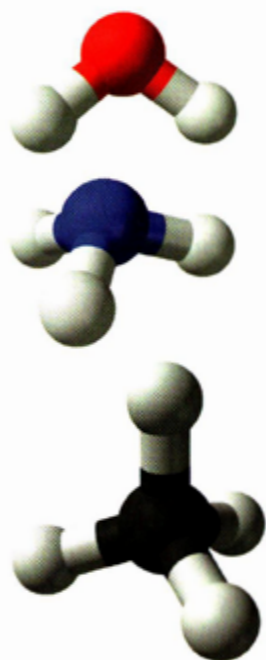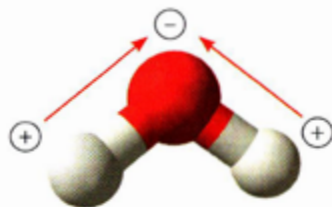

Figure I.2. The higher electronegativity of oxygen atoms compared to hydrogen atoms results in the polar water molecule.

I.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.
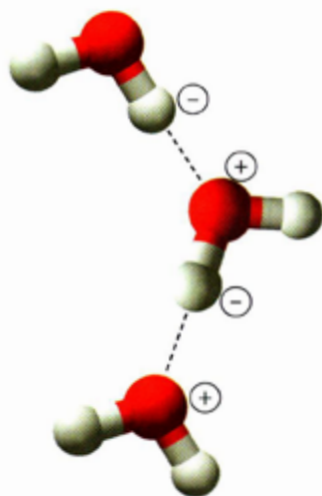

Figure I.3. Hydrogen bonding in water molecules. Dashed lines indicate hydrogen bonds between water molecules.
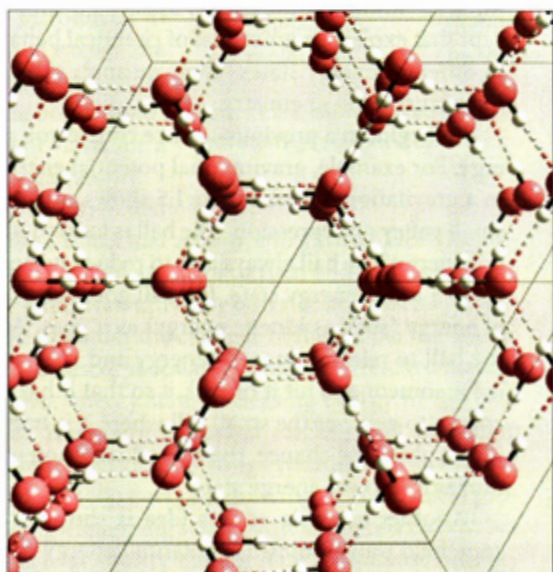

Figure I.4. The crystal structure in ordinary water ice. In this model, oxygen atoms are red and hydrogen atoms are white. Hydrogen bonds are shown as dashed lines. Thick black lines indicate the bonding of hydrogen to oxygen inside individual water molecules.

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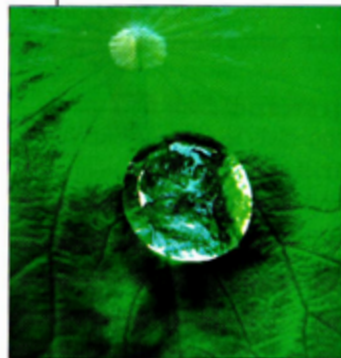
Nonpolar molecules do not dissolve in water. The attractions between the polar water molecules squeeze out the nonpolar molecules, causing the two substances to separate. This is why oil and vinegar separate—oil molecules are nonpolar and vinegar is mostly water.

I.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.
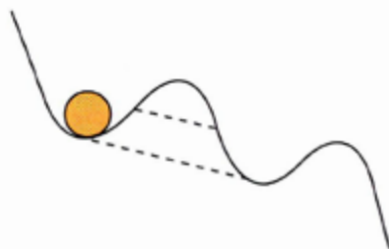

Figure I.5. The ball is trapped in the lowest-energy region in its vicinity. However, if a tunnel to a lower-energy region opens up, the ball goes there.

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:

$$2H_2 + O_2 \rightarrow 2H_2O$$
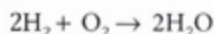
The left side of this equation indicates that the *reactants*—the substances taking part in the reaction—are molecules of hydrogen (H_2) and oxygen (O_2). Each of these gases exists as *diatomic* molecules, meaning that each hydrogen molecule is a pair of hydrogen atoms bonded together, and each oxygen 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.

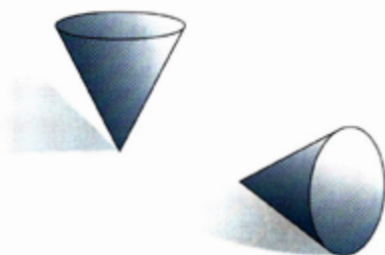

Figure I.6. The cone on the left is in a higher energy state. If released, it falls to the lower energy state shown on the right.

oxygen molecule
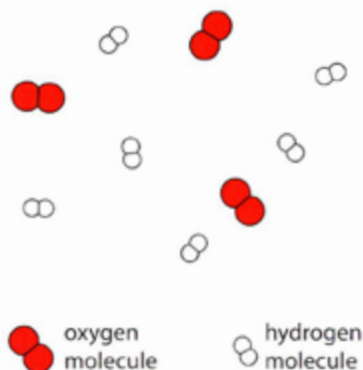
hydrogen molecule

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.

$$6CO_2 + 6H_2O \xrightarrow{\text{light}} C_6H_{12}O_6 + 6O_2$$

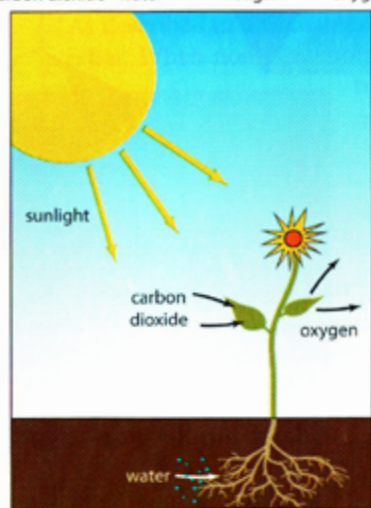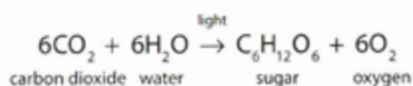carbon dioxide water   sugar   oxygen

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 an 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.

proton electron
(+) (−)

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

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.)

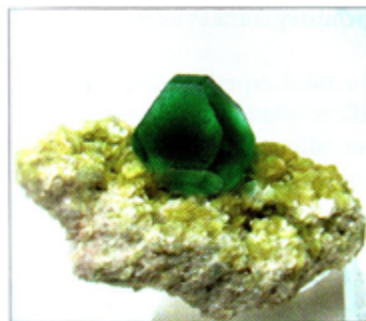### I.1.4 Chemistry Is All About Whole-Number Ratios of Atoms

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.
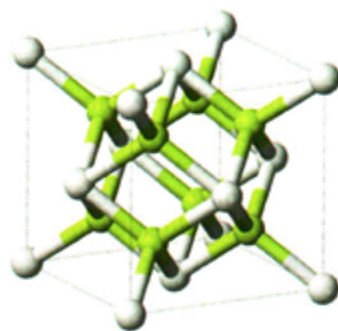
ratio of oxygen atoms to sulfur atoms is 4 to 1. Of course, there are strange exceptions to every rule, including this one. Nevertheless, it is correct to say that just about every compound, regardless of how the atoms are structured, consists of atoms of different elements joined together in predictable 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.
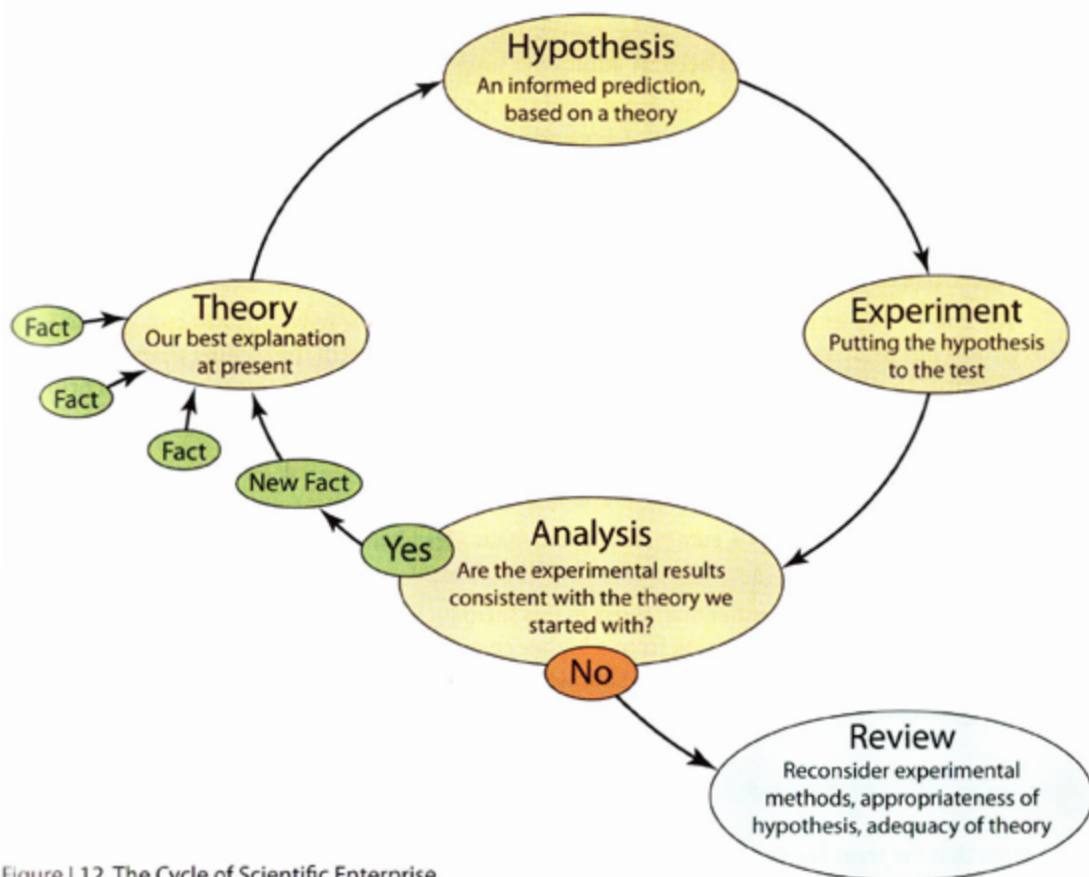

Figure I.12. The Cycle of Scientific Enterprise.

Fact	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 *true* or *proven*. Instead, we say a fact is correct so far as we know.
Theory	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. *Successful theories are the glory and goal of science.* 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 *true* or *proven*. 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 *model* of how nature works.[2]
Hypothesis	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.
Experiment	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.)

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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."

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