

CHAPTER 1

Matter and Energy: The Origin of the Universe

OVERVIEW

The objective of the opening chapter is to make a case for the relevance of chemistry in our lives and to prospective careers. After this introduction, the chapter defines matter and energy, discusses mixtures and their separation, and outlines the scientific method. The rest of the chapter focuses on measurement and unit conversions. Our approach introduces many of these topics, using substances familiar to most students such as gold and water.

For many students, the material in Chapter 1 is neither new nor particularly exciting. Some students are well versed in the topics in the chapter and may find the repetition of this information boring. Others, however, are not as confident and need the foundation that Chapter 1 presents. The challenge for an instructor is to accommodate the needs of both groups. Our solution is to start by asking how the universe may have begun. By examining selected observations that support the Big Bang theory, we hope to catch the attention of students at all levels. For example, the residual thermal signal from the Big Bang leads to a review of temperature scales. We present the Big Bang's microwave afterglow, which was predicted in the 1940s and discovered by accident in the 1960s, as further evidence. We believe this contextual approach powerfully reinforces the central role of chemistry among the sciences. It also engages students in examining the workings of the scientific method by using a topic with which many are familiar.

A more detailed discussion of the Big Bang theory requires an introduction to the pioneering work of Edwin Hubble in the 1920s. Hubble's discovery that galaxies outside the Milky Way were receding supplied evidence to support the Big Bang hypothesis. To address Hubble's evidence for an expanding universe, we need to discuss redshifts and the Doppler effect. Although Chapter 1 does

not cover these topics, you may wish to supplement the textbook on these topics. A discussion of redshifts requires an introduction to the wave properties of radiation and to equations describing their behaviour—for example, $c = \lambda\nu$ and $E = h\nu$. Section 7.1 of Chapter 7 covers the fundamentals of the wave nature of light. To challenge your audience, you may wish to move that material forward to Chapter 1. For a more adventurous instructor, the simple relation between temperature and wavelength given off by an object heated to a very high temperature, $\lambda_{\text{max}} = 2.897 \times 10^6 \text{ nm} \cdot \text{K}/T$, is key to understanding the microwave afterglow of the Big Bang and may prove stimulating to students who have studied measurements and unit conversions in middle and high school. We believe these discussions and those of the discovery of heterogeneities in the microwave background in the 1990s add an intriguing perspective on the dynamics of scientific discovery.

The Big Bang leads us to constructing atoms from subatomic particles and work toward more complex structures. Thus, we favor the bottom-up approach, starting with a microscopic view and building toward a macroscopic picture of matter. We hope to help students understand atomic structure before turning to the chemical properties of the elements and the formation of compounds.

We recognize that you, the instructor, may have your own favorite contexts within which to frame your classroom discussion of the traditional topics in Chapter 1. We hope you will use them. Additional contexts will challenge your students to see the similarities between those in the book and those presented in class. For example, one member of the author team uses a can of carbonated beverage to introduce mass, volume, and density, as well as the differences among the states of matter and among pure elements and heterogeneous and homogeneous mixtures. This context is one that every student can identify and understand.

Chapter 1 and those that follow it include three features that we hope prove useful. Each Sample Exercise is solved and paired with an unsolved Practice Exercise. Concept Tests are intended to compel the reader to pause and provide a nonnumerical answer. As an instructor, you might use these questions as the basis for in-class discussion. Each Sample Exercise is solved using a set of steps, known collectively as COAST: collect and organize, analyze, solve, and think about it. COAST is by no means the only effective problem-solving approach. You may have a time-tested method of your own. However, for the novice student, COAST offers one approach to problem solving. COAST is consistently applied in the solutions to all the Sample Exercises in this textbook and to the support package.

TEACHING THE CONTEXT

Writing the opening chapter of a general chemistry textbook challenges a textbook author. Typically, the content is light on chemistry. Certain topics such as measurement and unit conversions are unlikely to evoke an enthusiastic response from students. One is often tempted to skip this material altogether in favor of “real chemistry” in later chapters. Nevertheless, with students’ diverse backgrounds, instructors must include and discuss these topics. In Chapter 1 we have tried to combine standard introductory material with a more interesting story.

Consistently throughout this book we start with observations of the world around us. We see familiar objects around us, all of which are examples of **matter** (Section 1.1). **Chemistry** is the science of matter. As an instructor, you might focus on some simple examples of matter in the classroom—some copper wire, a bottle of water a student may have brought to class, or a piece of gold jewelry. What are these examples of matter composed of? Matter has **mass** (an **extensive property**) and occupies space. All matter is composed of **atoms** (Figure 1.5). Copper wire is an example of a **pure substance** as well as of an **element**, copper, which has certain **physical properties** such as **density**, an **intensive property**. Copper is malleable, is ductile, and has high electrical conductivity. The tarnish on the wire is a **chemical property** of copper. One might even drop a little copper wire into nitric acid to further demonstrate the chemical properties of copper, a **chemical reaction** that you can describe with a **chemical equation**. The text uses gold as an example of density determination (Figure 1.20). Is the gold jewelry you wear a pure substance? The answer is yes only if the gold is 24 karat. (One karat is equal to 1/24 [4.17%] of the weight of the metal in a mixture.) Drop a gold ring into nitric acid—nothing happens, showing the different chemical properties of copper and gold. You may need to cheat here and use a piece of cheap 14-karat gold jewelry rather than find some more expensive 22- or 24-karat gold.

At this point, revealing the periodic table and locating copper, gold, and the other 110 elements might be useful.

Which element is water? This may be a silly question to nearly everyone in the audience but is one worth asking on the first day of class. A mixture of ice and water is a useful visual aid here. Water, of course, isn’t an element at all but a pure substance that is a **compound** with a **chemical formula**: H_2O . Water is a **molecule** composed of the elements hydrogen and oxygen. Showing that water is composed of two gaseous elements is easy if you have a simple electrolysis apparatus available. Connecting a large dry cell to two pencils immersed in a beaker of water is sufficient to illustrate the generation of hydrogen and oxygen. Water exists in three physical states: **gas**, **liquid**, and **solid**. Boiling the water on a hot plate illustrates the concept of water vapor. Choosing water, pre-1980 pennies, and 24-karat jewelry as examples has allowed you to cover most of Figure 1.2, a standard illustration of the organization of matter.

Why is restricting our discussion to copper wire and 24-karat gold jewelry important? Contemporary “copper” pennies are examples of **heterogeneous mixtures**: a zinc core surrounded by a copper cladding. A 14-karat earring is a **homogeneous** mixture of 58.3% Au and 41.7% Cu by mass. How would you separate the copper from the gold in 14-karat gold? Here is an opportunity to combine several observations. We have seen that nitric acid dissolves copper but not gold. Both metals and the homogeneous mixture of 14-karat gold (alloys are defined in Chapter 12) will dissolve in a mixture of HCl and HNO_3 (aqua regia, 3:1). In principle, electrolysis (of the aqueous solution) will deposit gold at a lower potential than that required for copper. The same procedure applies to separating copper from zinc in a penny. Both metals dissolve in nitric acid yet are deposited electrochemically at different potentials. (Demonstrating the etching of the copper from the outside of a penny with acid to show the zinc core is tempting.) **Distillation** (Figure 1.12) of the homogeneous mixture obtained by dissolving copper in nitric acid separates the water from the ionic compound copper nitrate on the basis of **volatility**. If you restrict yourself to pennies and gold jewelry, then illustrating **filtration** becomes a bit trickier. One choice is to evaporate the copper nitrate solution until crystals (solid) form, followed by filtration. You can do this during the lecture as part of the distillation demonstration.

Copper wire and gold jewelry (if you did not sacrifice them to the acid treatment) provide the context for discussing measurement, **conversion factors**, the uncertainty inherent in all measurements, and the need to consider **significant figures** in both measurements and calculations involving experimental data (Section 1.8). One can calculate the density of either object from measurements of the mass and volume (Figures 1.19 and 1.20). For making each measurement, the uncertainty depends in part on the **precision** and **accuracy** of each measurement (Figure 1.22). Calculating density

forces you to take into account the number of significant figures dictated by your measurements.

Where did matter come from? Unknowingly, perhaps, it is a question we have all probably asked on several occasions in our lives. The answer to this question may place you, the instructor, in a delicate position of addressing the differences among **hypotheses**, **theories**, beliefs, and the role of the **scientific method** (Sections 1.7 and 1.10 and Figure 1.18). We present the Big Bang as an example of a scientific hypothesis and evaluate some of the experimental evidence supporting this hypothesis to illustrate the scientific method. For some students, coming to terms with the Big Bang hypothesis may prove problematic, so we have mentioned Hoyle's skepticism in the same sentence as LeMaître's hypothesis. Observations of receding galaxies are viewed as support for the Big Bang hypothesis but do not have to be presented with this emphasis.

The residual thermal signal from the Big Bang offers an opportunity to review the three commonly used temperature scales: **Celsius**, **Fahrenheit**, and **Kelvin**.

ALTERNATIVE CONTEXTS (Key Terms Bolded)

A Chemist's View of a Soda Can

We suspect that some instructors struggle with the question of how to make the study of chemistry seem relevant to a class of science majors from a variety of disciplines, of which chemistry majors are likely to be a minority or even absent. One answer is to give students examples from the world around them. This approach is the basis for many of the science pages in magazines and newspapers as well as popular books written for the nonscientist.

Although we have chosen the Big Bang and origin of the universe as a medium to introduce the fundamentals of chemistry, one could choose from many alternative contexts that may be less controversial. One example is to focus on a can of soda, such as Coke or Pepsi. Advertisements for these products abound, and a student in your class has probably brought a can of soda to the lecture. You might want to bring one yourself. A can of soda contains examples of all three phases of matter: solid, liquid, and gas. The aluminum can (excluding the paint on the outside) is a **pure substance** and an example of an **element**. Aluminum is a shiny (**lustrous**), **malleable**, **ductile**, and **electrically conductive metal** (or **metalloid**), all of which are examples of **physical properties**. How can we use a can of Coca-Cola to illustrate **chemical properties**? Ask students to reflect on the fate of steel versus aluminum containers left out in the rain by the side of the road. Iron rusts, but aluminum does not (ignore for the moment that aluminum will slowly oxidize to aluminum oxide).

The contents of the sealed can represent a **homogeneous mixture** containing the ingredients listed on the label. Once the can is opened, the unequal distribution of the gas bubbles makes it a **heterogeneous mixture**. The gas is carbon dioxide, a **molecular compound** containing two nonmetallic elements, carbon and oxygen, with the **molecular formula** CO_2 . The sodium present in the soda is, of course, not in the elemental form but is found as the sodium **cation** with some unspecified **anion** to balance the charge.

How could we separate this mixture? **Distillation** would allow us to remove the highly volatile CO_2 as well as to separate the water from the other ingredients on the basis of **volatility**, leaving behind a mixture of molecular and ionic compounds. One could argue that as the volume of the solution decreases, we could remove the precipitated solids by **filtration**.

A can of soda typically is labeled with its volume, 12 fl oz or 355 mL. This label represents a fine introduction to the **units** used in measurement and serves as an example of unit conversions. Is 12 fl oz really 355 mL? The two values contain different numbers of **significant figures**. How does this concept affect a calculation that converts 12 fl oz to milliliters? The mass of the soda is not given; how could we determine the mass of the solution (assuming that the mass of the can itself is minimal)? Weighing the can forces us to ask about the **accuracy** and **precision** of our measurement. Furthermore, with the mass of the soda in hand, we can calculate the **density** of the solution. Finally, most of us would prefer to drink our soda cold rather than warm. What we mean by *cold* depends on which temperature scale we use: **Fahrenheit**, **Celsius**, or **Kelvin**. Although 0°C might be a nice temperature for a "cold" soda, 0°F would probably leave us with a slush, and 0 K represents an unattainable value, **absolute zero**. Throughout this *Instructor's Resource Manual* we will return to the soda can as an example of an alternative context.

ACTIVE AND COLLABORATIVE CLASSROOM EXERCISES

ChemConnections Activity Workbook

Activity 1 ("What Am I Eating?") in the *ChemConnections Activity Workbook* is a wonderful way to relate unit conversion to students' previous knowledge of units in their lives. It starts with simple one-step conversions and builds up to multiple unit conversions within a problem. The problems are simple enough that anyone can do them, but everyone will have to refer to the conversion table at the start for some of the rarer and yet common conversions.

Activity 4 ("How Can We Separate Plastics for Recycling?") is set up as a lab activity. It brings together the **density** of

plastic versus water at different temperatures as well as when salt is dissolved in water to create a **homogeneous mixture**. The activity brings to light the common experience of recycling plastic, but few students will have actually thought about how the process actually happens. Other ways of sorting plastic exist, such as Near-IR, but this does not apply to every situation.

Lecture Suggestions with *Clickers in Action*

The following clicker questions have been selected from Margaret R. Asirvatham's *Clickers in Action* (W. W. Norton) and are available at www.norton.com/instructors.

In general, clicker questions requiring calculations are more time consuming; however, students should be able to take an educated guess at the answer to a question. By limiting the time they have to answer the question, you can force them to practice estimation, or the “Analyze” step of the COAST method for solving problems.

One way to proceed is to start with **Questions 1 and 2**. These examples drive home the idea that you will expect students to know chemistry concepts and terms. You can start the semester by explaining that students can think of chemistry as a foreign language since it is filled with terms not used in everyday discussions. What is a molecule? What is sublimation? Asking these questions allows you to build an understanding with the students and lets them develop an appreciation of your lecture style. To have a positive classroom experience, you need to define your expectations of the students early on as well as what you will be giving them. Will you define all terms in lecture? Will you give the students a list of terms you expect them to know? Will you expect them to go through the textbook on their own and define each boldface term?

Most questions in this chapter are reviewing terms defined in the book. However, **Question 12**, the last question, brings to light the difference between accuracy and precision—a hard concept for anyone to grasp, which is evident from the histogram for this question.

One great example to bring together several *topics* from this chapter would be to do the activity “Bowling for Density” (K. Holley, D. Mason, and K. Hunter, *J. Chem. Educ.*, 2004, vol. 81, p. 1312A). Most bowling balls weighing less than 12 lbs will float on water, and those greater than 12 lbs will sink. Few students will have known this before class; maybe you did not either. You can either bring a balance with you to class or give the mass for the bowling balls. The only other materials needed to calculate **density** are a few bowling balls of different masses, meter sticks, and string to measure the circumference by using **dimensional analysis**. Having students discuss how to account for the volume of the finger holes and the **significant figures** in their density value is more beneficial than just presenting this information. The student values for the density of the 12 lb ball bring this

activity to a climax, as this value is closest to the density of water. Will it float or sink? How valid were students' assumptions? With this example, you can also introduce students to the **scientific method** and the COAST approach to problem solving.

Another great visual is listed under chemical demonstrations, which involves comparing the density of regular and diet soda cans. Have you ever wondered when attending family reunions in the summertime why the diet sodas are always on top? With less sugar dissolved, diet cans float and regular soda cans sink.

ChemTours

The following ChemTours are available online at www.npag.es/chemtours.

Big Bang

Section 1.7

This animation explores the early formation of matter and radioactive decay rates in the Big Bang.

Significant Figures

Section 1.8

This ChemTour reviews the rules for assigning significant figures and walks students through sample calculations. It concludes with a series of interactive Practice Exercises that require students to express answers to addition, subtraction, multiplication, and division problems in significant figures.

Scientific Notation

Section 1.8

This ChemTour explains how to use scientific notation to express very large and very small numbers, and how to easily convert back and forth between decimal numbers and scientific notation. It includes Practice Exercises.

Dimensional Analysis

Section 1.9

Students learn to keep track of the units associated with numerical values. This ChemTour includes worked examples and interactive Practice Exercises.

Temperature Conversion

Section 1.10

Students practice converting among the Fahrenheit, Celsius, and Kelvin temperature scales. This ChemTour includes Practice Exercises.

Calculations in Chemistry

You will probably have students with a variety of experiences and thus a variety of mathematical abilities. Chapter 1 of any general chemistry textbook introduces the basics of many different types of mathematical calculations and definitions

used in the rest of the textbook. The authors of *Calculations in Chemistry* realize this and have provided its first four chapters to cover in great detail every math ability needed. You will need to assess and determine what math abilities a student may lack or need further development of. Then you can review the section devoted to each ability in these four chapters.

REFERENCES

Classroom Demonstrations

- “Science Demonstrations, Experiments, and Resources,” D. A. Katz, *J. Chem. Educ.*, 1991, 68, 235.
- “Ira Remsen’s Investigation of Nitric Acid,” in *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 2, L. R. Summerlin, C. L. Borgford, and J. B. Ealy, American Chemical Society, Washington, DC, 1988, p. 4.
- “The Mysterious Sunken Ice Cube,” in *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 2, L. R. Summerlin, C. L. Borgford, and J. B. Ealy, American Chemical Society, Washington, DC, 1988, p. 15.
- “Sugar in a Can of Soft Drink: A Density Exercise,” in *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 2, L. R. Summerlin, C. L. Borgford, and J. B. Ealy, American Chemical Society, Washington, DC, 1988, p. 126.
- “Densities and Miscibilities of Liquids and Liquid Mixtures,” D. A. Franz, *J. Chem. Educ.*, 1991, 68, 594.
- “Colorful Mixture Separation,” in *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 2, L. R. Summerlin, C. L. Borgford, and J. B. Ealy, American Chemical Society, Washington, DC, 1988, p. 17.
- “Separating Liquids: Fractional Distillation,” in *Chemical Demonstrations: A Handbook for Teachers of Chemistry*,

Vol. 3, B. Z. Shakhshiri, University of Wisconsin Press, Madison, 1989, p. 258.

“Meter Sticks in the Demonstration of Error Measurements,” R. Suder, *J. Chem. Educ.*, 1989, 66, 437.

“Bowling for Density,” K. Holley, D. Mason, and K. Hunter, *J. Chem. Educ.*, 2004, 81, 1312A.

Laboratory Exercises

The following laboratory exercises address some of the topics covered in this chapter.

The *Journal of Chemical Education* has published few lab exercises dealing with general chemistry chapter 1 material in recent years, and not many lab activities in this area have been published recently elsewhere. The following are a few topics that you may find useful in limiting your search.

- Observing some physical and chemical changes
- Studying density
- Classifying matter

“Endothermic Chemical and Physical Changes: An Introductory Chemistry Experiment” Margaret J. Steffel. The Ohio State University at Marion, Marion, OH 43302 *J. Chem. Educ.*, 2006, 83, 1185 **DOI:** 10.1021/ed083p1185. Publication Date (Web): August 1, 2006

“Floating Plastics: An Initial Chemistry Laboratory Experience” Enrique A. Hughes, Helena M. Ceretti, and Anita Zalts Instituto de Ciencias, Área Química, Universidad Nacional de General Sarmiento, Roca 850, 1663 San Miguel, Prov. Bs. As., Argentina. *J. Chem. Educ.*, 2001, 78, 522 **DOI:** 10.1021/ed078p522. Publication Date (Web): April 1, 2001