

## CHAPTER 2

# Atoms, Ions, and Molecules: Matter Starts Here

### OVERVIEW

Chapter 2 begins with a relatively traditional discussion of the pioneering work of J. J. Thomson, Ernest Rutherford, and Robert A. Millikan but emphasizes how experimental evidence led to the currently accepted model of the atom. The concept of the scientific method (Chapter 1) shows how the interpretation of data from critical experiments was essential to each scientist's work.

Although we are not unique in mentioning the contributions of F. W. Aston to identifying isotopes, we feel that citing Aston's work by name is also important. Why? Surely we are not suggesting that rote memorization of names is crucial to understanding isotopic composition—it isn't. Rather, we seek to convey how Aston built on Thomson's work and how it is related to Millikan's and Rutherford's contributions.

At this point we examine some of the simple compounds in Earth's crust and atmosphere. Discussing composition and chemical properties of matter creates the need to cover the conventions used to name compounds. We have chosen compounds found in the crust or present in the early atmosphere to illustrate the rules of nomenclature. You could substitute your "favorite" compounds during classroom discussions. You might take a consumer chemistry approach, as we have done with some of our illustrations. (Chapter 4 features a few of these products.) You can find further information on the chemistry of consumer products in *1001 Chemicals in Everyday Products*, 2nd ed., by Grace Ross Lewis (Wiley, 1999). With a model of the atom, the shape of the periodic table begins to make sense, as we show students how elements in a group form ions with the same charge. (We do not introduce electronic structure until Chapter 7.)

We end the chapter by addressing the process by which elements are made in stars: nucleosynthesis. This section closes the circle by explaining the statements in the final lines of Chapter 1, namely, how elements were created. Although we urge you to teach the section on nucleosynthesis, you may use Section 2.7 at your discretion. If the academic calendar dictates that it is time to move on to Chapter 3, failing to cover the story of nucleosynthesis will not harm students' ability to comprehend the concepts in later chapters. You can even return to Section 2.7, if time allows, when teaching nuclear chemistry (Chapter 21).

Sometimes instructors follow Chapter 2 with Chapter 21, "Nuclear Chemistry." This approach presents students with a topic not covered in great detail in their previous chemistry classes. Therefore, students will need to develop study habits before the first exam and not just slide through the first few chapters, relying on their experiences. Developing these study habits early in the semester enables the long-term success of more students.

### TEACHING THE CONTEXT

In our years of teaching general (introductory) college chemistry, we have observed that students are generally well versed in the overall structure of the atom. As early as middle school, they learn that atoms consist of a nucleus, which contains protons (atomic number) and neutrons and which is surrounded by electrons. The periodic table is based on atoms having different atomic numbers. Secondary school chemistry classes discuss variability in the number of neutrons (forming isotopes). For an average student this foundation is quite solid. The names for compounds and ions, however, have probably faded from their memory. These

observations are relevant because these concepts constitute the bulk of Chapter 2. Why should an average student pay attention to lectures on atomic structure when he or she already knows it? This is the challenge to all instructors: how to make the material in this chapter new and relevant. Many approaches exist. We propose one solution: use the historical development of atomic structure to illustrate the application of the scientific method. By engaging students in how we know about atomic structure, as well as in what we know, we challenge them on a different level from their prior experiences with chemistry.

How and when did scientists first discern the commonly accepted picture of an atom? The notion of an atom as “plum pudding” (Figure 2.4) is just over 100 years old. Within 5 years of the publication of his model, J. J. Thomson’s concept of atoms as a sea of **electrons** dotted with **protons** had been replaced by Rutherford’s model of a nuclear atom surrounded by electrons (Figure 2.7). Our understanding of atomic structure developed in slightly more time than the average student in your class will take to graduate from college. We believe that the notion of a 5-year window of discovery will be more meaningful to a student than the knowledge that these discoveries were made between 1904 and 1909. One hundred years is too long ago for all but a smattering of history buffs in your audience to appreciate.

How did Rutherford disprove Thomson’s model? Starting from Thomson’s hypothesis and harnessing the work of Becquerel and the Curies, in addition to his own results on radioactivity, Rutherford and his students designed the experiment shown in Figure 2.6 that showed Thomson’s model to be in error. Rutherford did not work in a vacuum. He applied the scientific method to the problem by building on the work of other scientists in synthesizing a new explanation for a structure of the atom that was more consistent with the data. Among the other scientists whose work Rutherford needed was Thomson himself. Thomson had been the first to demonstrate the existence of electrons (**cathode rays**) as **subatomic particles**, and Millikan had established the charge on an electron that allowed the calculation of its mass. Studies of atomic structure did not end with Rutherford’s experiment. A significant amount of mass in atoms was still “missing”; the mass of the protons and electrons did not add up to the known **atomic mass** of the atom. Additional experimentation by James Chadwick was needed to prove the existence of **neutrons**, whereas Francis Aston showed that atoms of the same element could exist as **isotopes** with the same **atomic number** but a different **mass number**. The **natural abundance** of different isotopes underlies the concept of **average atomic mass** for elements in the periodic table. All this work spanned only a couple of decades, a fraction of a scientist’s working life. Your students may appreciate the time frame of these discoveries more than the actual dates when the experiments were conducted. Your class may find

it interesting that Rutherford studied under Thomson, and Aston and Chadwick were both Rutherford’s students.

Carbon dioxide serves as an example of the law of definite proportions, introduced in Chapter 1. Compounds such as  $\text{SO}_2$  and  $\text{SO}_3$  that contribute to acid precipitation make excellent examples of the **law of multiple proportions**. The introduction of nomenclature and **chemical** and **molecular formulas** for **binary molecular compounds** such as  $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{SO}_3$ ,  $\text{NO}$ , and  $\text{NO}_2$  held together by **covalent bonds** fits neatly into the discussion of atmospheric molecules.

An example of **binary ionic compounds** is salt (Figure 2.16), composed of sodium **cations** and chloride **anions**. Students are all familiar with salt as a flavoring ingredient and in seawater. Consider the many possible compounds found in evaporated seawater besides sodium chloride. The **alkali** and **alkaline earth metal** halides (mostly chlorides) represent the most prevalent compounds. Lower concentrations of ionic magnesium and calcium compounds are also present in such evaporites, including those with **polyatomic ions** and **oxoanions** (Tables 2.3 and 2.4). Earth’s crust is also rich in compounds of **transition metals** such as iron and manganese, in which the metal cations often have a range of charges.

The general shape of the periodic table has been known for some 140 years. Dmitri Mendeleev’s periodic table (Figure 2.9) was based on reactivity patterns of the known elements organized by **groups** and **periods**. Differences among the physical and chemical properties of **metals**, **nonmetals**, **semimetals (metalloids)**, and **noble gases** had been documented well before anyone had a clear idea of the structure of an atom.

Where do the elements come from? Although most students have seen the periodic table of the elements before entering college, few will be aware of **nucleosynthesis**, the formation of the elements by nuclear processes in stars. Section 2.7 nicely answers the question left unanswered at the end of Chapter 1: how did the Big Bang lead to different elements? At the end of Chapter 2, students can understand the basic nuclear reactions (**fusion**, **neutron capture**, and  $\beta$  **decay**) that account for the synthesis of many elements in stars, including our sun. Most students will probably not have studied this concept in earlier courses, and we hope they will deepen their appreciation for the importance of atomic structure. After all, without the understanding of atomic structure that Rutherford and his peers developed early in the 20th century, nuclear processes and the stellar nucleosynthesis hypotheses could not have followed some 20–30 years later.

For your students, accepting that the origins of all matter can be traced to the Big Bang should seem like a giant leap of faith. What evidence supports this hypothesis? Chapter 1 presented some of the evidence. Whether a student truly accepts the Big Bang hypothesis does not affect the observation that Earth is constructed from compounds of some 82 naturally occurring elements.

## ALTERNATIVE CONTEXTS (Key Terms Bolded)

### A Chemist's View of a Soda Can Revisited

The readily identifiable can of soda you may have used as an alternative context in Chapter 1 works equally well in Chapter 2. The can is composed of aluminum, a **metal** or perhaps a **metalloid**. Carbon dioxide ( $\text{CO}_2$ ) is a **binary molecular compound** (Section 2.5) containing **nonmetals** held together by **covalent bonds**. Carbon dioxide remains a convenient starting point to introduce the law of definite proportions, the **law of multiple proportions** (Section 2.5), **chemical formulas**, and the nomenclature for binary molecular compounds (Section 2.6). The presence of 50 mg of sodium **cations** in many sodas introduces nomenclature rules for **alkali** (and, by extension, **alkaline earth**) **metals**. An **anion** must be present to balance the charge of the sodium ion. Two **acids**—phosphoric acid ( $\text{H}_3\text{PO}_4$ ) and carbonic acid ( $\text{H}_2\text{CO}_3$ )—are present in sufficient concentration to justify their inclusion on the list of ingredients. Both acids contain **oxoanions**, themselves examples of **polyatomic ions** (Section 2.6). All the elements found in many sodas can be located on the periodic table.

As an element, aluminum can serve as a model for studying atomic structure. Aluminum atoms have nuclei containing **protons** (atomic number = 13) and **neutrons**, surrounded by 13 electrons. We can infer the structure of aluminum atoms from Rutherford's experiments. Although aluminum is 100% natural abundance of  $^{27}\text{Al}$  (**average mass** = 26.98 amu), other **isotopes** exist. For example,  $^{26}\text{Al}$  has a half-life of about 740,000 years and is seen in stars, where **nucleosynthesis** occurs. In fall 2007, scientists detected a new, very heavy isotope of aluminum ( $^{42}\text{Al}$ ).

## ACTIVE AND COLLABORATIVE CLASSROOM EXERCISES

### Chemconnections Activity Workbook

**Activity 5** ("What Gives Fireworks Their Brilliance?") applies many topics in Chapter 2. The activity covers periodic trends and offers practice naming covalent and ionic species. Many of your students may know some of these compound names, but almost every student will need to learn some if not all of the naming practices. You could lift it up as a challenge that you will know all your students' names by the first exam if they will be able to name the compounds by then. Learning all their names can be challenging, especially with classes with more than 70 students, but if you take pictures of the students in small groups it is possible.

## 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 [wwnorton.com/instructors](http://wwnorton.com/instructors).

**Question 1** is a nice review if you assigned or want to encourage the ChemTours.

**Question 2** is great to assess how well students can do simple math and use their calculators. It can also serve as an opportunity to review significant figures.

**Questions 3–7** are all applicable to assess their learning of atomic structure and understanding of isotopes. Question 7 is especially pertinent to assess whether students can combine multiple concepts within one problem.

**Questions 8 and 9** drive home the concept of average atomic masses as well as the ability to estimate, which we want students to develop.

**Questions 10–13** are great ways to cover the structure of the periodic table. Some caution is needed here with combining this with your expectations of knowing the names of elements. We recommend that you be clear with your expectations of how many elements you expect the students to know by the time of the exam. They may not know these names yet. Thus, are you assessing their understanding of the table or their knowledge of the element names? By the first exam, one member of our author group has the students be able to name the first 36 elements, along with a few extra elements such as silver, gold, tin, lead, all the alkali and alkaline earth metals, halogens, and noble gases.

**Questions 15–23** present a variety of ways to assess students' knowledge of naming compounds. These vary in difficulty as well as the concept being covered. We will let you decide which areas you want to emphasize but have found such questions helpful.

### ChemTours

The following ChemTours are available online at [wwnpag.es/chemtours](http://wwnpag.es/chemtours).

#### Cathode Ray Tube

Section 2.1, Figure 2.2

This ChemTour explores the effects of magnetic and electric fields and cathode rays. Students can manipulate electric and magnetic field effects on the cathode ray.

**Millikan Oil-Drop Experiment**

Section 2.1, Figure 2.3

This ChemTour re-creates the experimental procedure Millikan used to determine the charge of an electron.

**Rutherford Experiment**

Section 2.1, Figure 2.6

This ChemTour re-creates Rutherford's gold-foil experiment, which led to the discovery of the atomic nucleus.

**Periodic Table**

Section 2.4, Figures 2.10 and 2.11

This ChemTour offers a guided tour of the trends that the periodic table summarizes (metallic properties, subshells, electronegativity, and atomic radius) and explains how to use this tool to predict an element's characteristics, including bonding capacity. An interactive version of the periodic table is accessible from every ChemTour (use the "Pt" button found in the upper-right corner of the screen). You should also refer to this feature when covering Chapter 8.

**NaCl Reaction**

Section 2.5

This ChemTour shows how a metal and a nonmetal combine to form a binary ionic compound, as seen in the reaction of sodium metal and chlorine gas.

**Synthesis of Elements**

Section 2.7

This ChemTour animates the neutron capture process and explains how elements are synthesized in stars.

**Calculations in Chemistry**

Even though Chapter 2 does not have many calculations, the *Calculations in Chemistry* workbook still offers a lot of help for students. For instance, the book introduces the atomic names over a series of lessons. Thus, learning all the elements can be overwhelming, but learning 10–15 of them at a time is not as daunting. Lessons 1.4, 3.4, 4.6, and 6.6 will help with this.

Chapter 6 offers a nice overview and definition of terms. Seeing terms defined slightly differently across various sources may give students the opportunity to more fully understand a

topic. This chapter also introduces students to the idea of using flash cards if they so choose. However, this chapter is a little light on calculating weighted averages of isotopes. Although the chapter covers this topic well, it offers no examples for students to work on. You will therefore need to find a different source for work on problems dealing with weighted averages.

Chapter 7 builds on students' knowledge of naming elements and now identifying formulas and naming compounds. The chapter also offers guidance for students in making their own flash cards. When you assign this chapter, you should note any ions that you will not require students to know.

**REFERENCES****Classroom Demonstrations**

"An Easy Model for Teaching Mass/Charge in Mass Spectrometry," R. V. F. Bravo and N. A. de Sousa, *J. Chem. Educ.*, 1989, 66, 1039.

"General Chemistry Demonstrations Based on Nuclear and Radiochemical Phenomena," R. H. Herber, *J. Chem. Educ.*, 1969, 46, 665.

"Nucleogenesis! A Game with Natural Rules for Teaching Nuclear Synthesis and Decay," D. J. Olbris and J. Herzfeld, *J. Chem. Educ.*, 1999, 76, 349.

"Teaching Aids for Nuclear Chemistry," C. H. Atwood, *J. Chem. Educ.*, 1994, 71, 845.

**Laboratory Exercises**

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

"Nomenclature Made Practical: Student Discovery of the Nomenclature Rules". Michael C. Wirtz, Joan Kaufmann, and Gary Hawley Concord Academy, Concord, MA 01742. *J. Chem. Educ.*, 2006, 83, 595. DOI: 10.1021/ed083p595. Publication Date (Web): April 1, 2006

"Flow Chart for Naming Inorganic Compounds" David Robson. *J. Chem. Educ.*, 1983, 60, 131. DOI: 10.1021/ed060p131. Publication Date: February 1983

"Teaching Inorganic Nomenclature. A Systematic Approach" Gerhard Lind. *J. Chem. Educ.*, 1992, 69, 613. DOI: 10.1021/ed069p613. Publication Date: August 1992