Chemistry in the World

First Edition

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cognella
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For my students, from whom I have learned more than I could ever teach.
The pictures on the cover of this text are some examples of the chemistry that surrounds us in the real world. At top left, poppies bloom in the Sonoran desert of Arizona in the spring as plants increase their growth rate. Not only does this depend on the chemical process of photosynthesis, it also increases the rate at which the greenhouse gas carbon dioxide is removed from the atmosphere. At top right and bottom left, we see two examples of water in nature—in liquid form as it cascades over Upper Mesa Falls in Idaho, and in solid form as snow in Bryce Canyon in Utah. There’s also water vapor in the atmosphere, where it serves the important role of helping to keep the Earth warm and habitable. Water has many interesting and unique chemical properties, and fresh surface water, as in this waterfall, is rare and valuable. At bottom right, aspen leaves in the San Francisco Peaks near Flagstaff, Arizona, turn vibrant shades of yellow in the fall. As seasons change, changing intensity of light prompts trees to adjust concentrations of light-harvesting chemicals in the leaves. This leads to the yellows, oranges, and reds that decorate the trees as the air turns crisp.

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Authoring this book has been a work in progress for the last ten or more years, though the actual writing took (thankfully!) somewhat less than that. There are many without whom this text would never have transitioned from dream to reality—you know who you are, and I thank you. Mostly, to my husband, Scott, for being the catalyst, and for providing the gorgeous photographs that bring the ideas to life.
To the Student and the Instructor:
How to Use this Book

This text was written to provide a perspective on socially important topics in the field of chemistry, which as the student will soon ascertain, is a broad field indeed! Of course, even an applications-based approach (as this book is) must by necessity introduce sufficient principle and theory to allow for meaningful discussion. It was a primary goal in writing this text to include only those principles absolutely necessary to understanding observable chemical phenomena and relevant chemical issues. At every step of the way, I strove to make readily apparent the answer to the question, *When am I ever going to use this again?* For this reason, I recommend reading this book more like a literature than a science text—it tells a story. Granted, the story is occasionally punctuated by equations and symbols, but it’s a story nevertheless. Rather than trying to internalize the applications in terms of the chemical principles, use the applications to make sense of the principles—the big picture will help you understand the details.

Each chapter introduces a topic or group of related topics in which we see chemistry on a day-to-day basis, in our lives or on the news. Principles are addressed where needed in the course of discussing the topic. Where relevant, practice problems are provided in boxes labeled *Try This.* It’s worth taking a moment to try a problem on your own before moving on—there are many additional practice problems at the end of the chapter. Conceptual questions are occasionally presented in boxes labeled *Concept Check.* It should be fairly easy to answer the question posed in the box after reading a particular section—if not, that’s an indication that you may have missed a main point from the previous paragraphs. Occasionally, *Concept Check* boxes will refer to older concepts that are being brought in to the current discussion, and in such cases, the answers will refer you to the relevant chapter.

Problems at the end of the chapter are divided into two sections. *Questions and Topics for Discussion* are intended to stimulate thought—in many cases, there are no right or wrong answers. These questions are meant to be suitable for essay or paragraph assignments, stimulation of thinking, or in-class discussions. *Problems* are mathematical or factual in nature—in particular, math problems occur in sets of two (grouped together) for ease of studying and assigning work.

The last section in each chapter is called *The Last Word,* and is an offshoot of something I’ve done in my classes that has been very popular with my students. In class, I call it “Five Minutes of Really Cool Chemistry,” and I’ll talk about something nearly relevant and always neat that is based upon the principles we’ve been discussing. *The Last Word* is my adaptation of that into written form—these sections are meant to be interesting vignettes that relate to the concepts addressed in the chapter. They’re not chapter summaries (which do, incidentally, appear separately at the end of each chapter), but are hopefully a few moments’ worth of relaxation and wonder—a few minutes of really cool chemistry.

Enjoy the journey!
Unit 1: All Around Us and Inside Us
The word *chemistry*, especially in the title of a class or book, conjures up images of goggled scientists at their lab benches, surrounded by odd bubbling solutions in beakers and flasks. Depending upon the power of your imagination, this scene might even be punctuated in your mind by the occasional explosion. In any case, as you imagine the white-coated chemist going about his business, mixing and stirring and writing down equations rife with symbols and letters of the Greek alphabet, I practically guarantee the two things you are thinking to yourself are *What in the world does this have to do with real life?* and *When am I ever going to use this information!?* The answers to those (very real, very important) questions are EVERYTHING, and EVERY DAY!

In order to really see the applicability of chemistry to real life, we have to stop imagining it as something that takes place in a laboratory and start looking for it in the real world. Chemists work in laboratories because they are attempting to study what goes on everywhere ELSE, not because the lab is the only place chemistry happens! In this book, we’ll look for chemistry everywhere EXCEPT the lab. We’ll learn that there’s chemistry all around us—in leaves changing colors in the fall, in the atmosphere (which keeps our planet warm and habitable the way a greenhouse protects roses from winter frost), in nuclear power plants, and in the food we eat every day.

**Image 1.1.** (From top to bottom) 1) Leaves change color in the fall as the light-absorbing chemical chlorophyll (which gives them their green color) is no longer produced, leaving behind other light-absorbing chemicals such as carotenes (yellow) and anthocyanin (red). 2) Just as a greenhouse made of glass traps heat to keep plants warm during cooler months, so the gases of the atmosphere trap heat from the Sun, keeping the planet warm and habitable. 3) Nuclear reactions have the potential to release massive amounts of energy that can be used to generate power. 4) The food we eat is made up of chemicals that provide us with energy and raw material for building muscle and other tissue.
This isn't just a chemistry textbook. It also addresses ways in which chemistry, chemicals, and chemists affect society. However, this is more complex than it sounds. It's very important to think about the relationship between chemistry—any science, really—and society as a reciprocal one. To be sure, chemistry affects society: pollution, for instance, has a clear and unequivocal negative impact upon those who have to breathe it. However, society also affects chemistry, in that our actions produce and release chemicals into the environment, and we formulate legislation that affects what can and can't be done chemically. In this text we will be exploring the nature of this relationship and some of the ways in which chemistry and society affect each other.

Image 1.2. The complex interplay between chemistry and society is apparent in this picture. We affect the chemistry of the oceans by adding environmental pollutants and toxins that disrupt the ecosystem. As a result, the altered chemistry of the ocean affects us. For example, mercury builds up in high-level predator fish that we use as food (tuna and shark are among these), rendering them toxic to us if consumed more than occasionally.

Image 1.3. Chemistry on the individual level includes such topics as pharmaceuticals (above) and food (bottom right). Cities need to consider the chemistry of local air pollution (top right) and water safety (purification plant, top left). As a global community, we are all concerned with and affected by climate change. The countries shown in green (bottom left) have signed the Kyoto Protocol for reducing carbon emissions. (Why hasn’t the United States signed? See Chapter 13.)
It’s important to note that society is a multifaceted concept. Individuals, local communities, and the entire global population are different levels of society. Any level can affect any other (and be affected by any other) through chemistry. For instance, making appropriate choices regarding nutrition and exercise positively impacts personal health, but also reduces the need for medical care, which decreases insurance rates for all policy holders. A community that installs a light-rail system benefits from a reduction in local pollution, but there are also global benefits through reduced carbon emissions.

In order to address the reciprocal relationship between chemistry and each level of society, this text is divided into four units. In the first, we’ll be discussing some introductory chemical principles in the context of the chemistry we encounter every day. The second unit will address chemical concerns within local communities—things like air pollution and water safety. In Unit 3, we’ll shift our focus to the chemistry that takes place within our own bodies. Finally, in Unit 4, we’ll expand our thinking outward to encompass issues of global concern. At every step along the way, however, we will be focusing not simply on chemical concepts, but on how we affect the chemistry and how the chemistry affects us.

1.1 CLASSIFICATIONS OF MATTER

Stated very simply, chemistry is the study of matter. Matter is defined as anything that occupies space and has mass, so it is what makes up the physical universe. Everything we can touch is made of matter, but because there are several different types of matter, we need to define some general categories.

Let’s start with mixtures, because they’re actually the simplest kind of matter to understand. Mixtures are physical combinations of two or more pure substances in variable proportions. There are two key components to this definition. First, that a mixture is a physical combination means that each substance within the mixture retains its own chemical identity and properties, but is interspersed with one or more other substances in space. For instance, making a cake involves mixing sugar, baking soda, flour, spices, eggs, and so forth. Sugar is a chemical with distinct properties (it tastes sweet, it dissolves in water, and so on). Baking soda is another chemical with a separate set of distinct properties. As an ingredient like sugar is added to the batter, it mixes in and spreads out through the batter until all parts of the mixture are equally sweet. The sugar, though dispersed through the batter, is still sugar—as a chemical, it remains unchanged. That’s the nature of a mixture.

Saltwater is another example of a mixture. Salt (a common name for the chemical sodium chloride, or NaCl) has distinct properties. Water (a common name for the chemical H2O) has a different set of properties. They can be stirred together to make saltwater, but the salt is still NaCl, and the water is still H2O. In other words, physically mixing salt and water means that bits of NaCl are surrounded by bits of H2O and vice versa, but the individual
Chemicals retain their identities and properties. If we wanted to, we could “un-mix” salt water by boiling it in a pan on the stove, in which case the water would depart as steam (which is still H₂O), and the NaCl would be left behind as a crust on the bottom of the pan. At no point in the mixing or separating process does either chemical change identities or properties.

The second key part of the definition of a mixture is that the substances are combined in variable proportions. This means that it doesn’t matter how much salt we mix with a given amount of water—whether we use lots of salt or only a bit, we’ll still have a mixture of saltwater.

Chemicals that can be physically combined to produce mixtures are called pure substances, and can be divided into two sub-categories: elements and compounds. Elements are the simplest types of matter and are found listed on the periodic table (Image 1.6). They cannot be broken into simpler substances by any chemical means. When you look at the periodic table, you’ll notice that it’s made up of many elements, each with a characteristic atomic symbol, which is a letter or combination of letters that represents the element. For instance, the atomic symbol for element carbon is C, whereas the symbol for element calcium is Ca.

Elements are either metals or non-metals. Examples of metals are lithium (Li), sodium (Na), and aluminum (Al); examples of non-metals are fluorine (F), chlorine (Cl), and oxygen (O).

Compounds are made from two or more elements combined in fixed proportions. Water (H₂O) is a compound made from the elements hydrogen (H) and oxygen (O). Many compounds have characteristic names. For example, sodium chloride (NaCl) is also known as table salt. Other examples are chloroform (CHCl₃), natural gas (CH₄), and acetic acid (CH₃CO₂H).

...
Compounds are chemical combinations of two or more elements in fixed, characteristic proportions. Notice the differences between the definition of a compound and the definition of a mixture. First, where mixtures are physical combinations in which each pure substance retains its identity and properties, compounds are chemical combinations, meaning that a chemical bond forms between particles of two or more different elements. A chemical bond is a little like glue—it sticks two particles together. The compound has its own identity and properties that are different from the identities and properties of the elements that make it up. For instance, carbon dioxide (CO₂) is a compound made from a chemical combination of the elements carbon and oxygen, but CO₂ is nothing like either carbon or oxygen—it is its own separate, distinct substance. Another major difference between compounds and mixtures is that compounds must consist of fixed proportions of the constituent elements. While we can use any quantity of salt and any quantity of water to make salt water, CO₂ is only CO₂ if it’s made up of one particle of carbon and two particles of oxygen. The subscripts in the compound’s formula tell us how many particles of each element are chemically combined to form the compound, where the lack of a subscript is always taken to mean one. If one particle of carbon and one particle of oxygen combined chemically, the resulting compound would NOT be CO₂. Instead, it would be CO—carbon monoxide—which has very different properties than CO₂.
CO₂ can be classified further—in addition to being a pure substance, it’s also a molecule. A molecule is a pure substance made from the chemical combination of two or more atoms. An atom is defined as the smallest particle of an element that maintains the identity of that element. In other words, if I had a chunk of the element gold (Au on the periodic table), and I cut it in half, I’d have a smaller chunk of gold. If I cut that smaller chunk in half, and then cut the half in half, and so forth, eventually I’d get to a very, very tiny particle of gold that couldn’t be cut in half (or at least, if it were, it wouldn’t be gold anymore*). This tiny, indivisible particle of gold is an atom (the word atom comes from the Greek word atomos, meaning indivisible). All matter is made up of atoms, some of them uncombined and some of them in chemical combination with other atoms. A molecule of CO₂ is made up of one atom of carbon and two atoms of oxygen.

*Note

This statement seems paradoxical at first—how can we say that elements are the simplest type of matter, and then say that if we were to divide an atom of an element (gold, for instance), we’d get something that was NOT gold? We’ll discuss this more in Chapter 2, but as it turns out, we can’t really “cut” an atom and get pieces that hang around at all. If an atom were (theoretically) divided, the resulting pieces would be what are called sub-atomic particles, and would not be stable on their own. So, stated simply, atoms are the smallest particles of matter that are independently persistent.
found as uncombined atoms: examples include helium (He) and neon (Ne). Other elements occur most frequently as diatomic molecules (molecules consisting of only two atoms) such as O₂ and N₂. Still others occur naturally as polyatomic molecules, or molecules composed of many atoms. Sulfur, for instance, is most commonly found in nature as S₈. Atoms, molecules, elements, compounds, and mixtures are visually summarized in Image 1.11.

**CONCEPT CHECK:**

What's the difference between a molecule and a compound?

**Answer:** A molecule consists of two or more atoms, chemically combined. A compound consists of two or more elements, chemically combined.

**TRY THIS:**

Which of the following are pure substances, and which are mixtures? Of the pure substances, which are elements and which are compounds? Which are atoms and which are molecules?

Fe, FeCl₂, sugar water, silver (Ag on the periodic table), maple syrup, Cl₂.

Answers: Fe is a pure substance, an element, and an atom. FeCl₂ is a pure substance, a compound, and a molecule. Sugar water is a mixture. Ag is a pure substance, an element, and an atom. Maple syrup is a mixture. Cl₂ is a pure substance, an element, and a molecule.

**Image 1.10.** The polyatomic molecule S₈ is a crystalline, yellow solid that smells faintly of rotten eggs.

**Image 1.11.** An atom of oxygen (top left) is a pure substance, an element, and, of course, an atom. As a diatom, oxygen (O₂, top right) is a pure substance, an element, and a molecule. A molecule of carbon dioxide (CO₂, bottom right) is a pure substance, a compound, and a molecule. A mixture of ethanol (C₂H₅O) and water contains molecules of both compounds, physically stirred together, but chemically unchanged.
One question students often have regarding elements is how to know whether they are found in nature as single atoms, diatomic molecules, or polyatomic molecules. The answer is that frankly, it’s a bit complicated! There are a few rules of thumb, though, that are helpful to know. The elements that naturally form diatomic molecules are H, N, O, F, Cl, Br, and I, and it’s quite useful to know this about them, because several of these are elements we’ll be talking about frequently in this text.

The elements in the far-right column of the periodic table tend to be found as free atoms, and we’ll learn why in Chapter 3. One other element whose form in nature is worth mentioning is carbon. Most of the carbon atoms we’ll see in this class will be chemically combined with other elements, because carbon is a very common component of compounds in nature. However, elemental carbon is also common—probably more so than you think—and it can be found in several different forms, with VERY different properties. Two of the most frequently encountered forms of elemental carbon are graphite and diamond. These differ only in the way in which carbon atoms are arranged and chemically bonded to one another.* Looking at the arrangement of atoms in these two forms (Image 1.13) helps us to understand some of the differences in their physical properties and appearances. Graphite’s structure is comprised of carbon atoms bonded into flat plates. These plates lie on top of one another, but are independent and can slide past each other easily. This sliding capability within the structure of graphite makes it soft and easily spread onto a surface (hence its utility as pencil “lead”). Also, it makes graphite a good lubricant in mechanical applications. Diamond, on the other hand, has carbon atoms

*NOTE
If graphite and diamond are both made of carbon, why is diamond so much more valuable than graphite? The quick answer is that it’s much harder and more durable because of the arrangement of atoms. It’s also considered aesthetically pleasing. These two factors combine to make it a good candidate for jewelry (among other things), and because it’s rare, it comes at a high price. From a chemical perspective, diamond is actually LESS STABLE than graphite, meaning that very slowly, diamond turns into graphite! Don’t worry—it doesn’t happen in a human lifetime (or even multiple human lifetimes), so your diamonds will still be around for your descendants. Still, it’s interesting to think about!
arranged in a crystal lattice in which all atoms are bonded to all nearby atoms. This structure is incredibly rigid, giving diamond its unique hardness.

One of the themes that we will see repeatedly in this class is the pervasiveness of chemistry. Even when we don’t think we’re talking about chemistry, much of the time we are. For instance, let’s ponder jewelry for a moment. This is a topic that seems very far removed from anything we might traditionally associate with chemistry. However, because it occupies space and has mass, jewelry must be made of matter, and chemistry is the study of matter. It becomes a little more complicated when we try to ask ourselves of what KIND of matter jewelry is made. Take, for example, a gold ring. If you look inside a gold ring, you’ll see a stamp that says 14K, 18K, or some other number followed by a “K” to indicate the material of the ring.*

Pure gold is an element (Au on the periodic table), and while aesthetically pleasing and rare (which combine to make it valuable), it’s far too soft to be used in jewelry designed for daily wear. As such, elemental gold is mixed with other elements like nickel (Ni) and palladium (Pd). A “gold” ring, therefore, is actually made of a mixture rather than a pure substance.

*NOTE

If you’re curious, the “K” is for karat, which is just a way of indicating what fraction of the metal mixture is elemental gold. By definition, 24K is elemental gold, or 24/24 parts of the metal are gold. 18K gold has 18/24 parts elemental gold, and 6/24 parts other metals. Metallurgists blend in various metals to affect the color and properties of the resulting alloy (mixture) in addition to strengthening the gold.
There's one liquid nonmetal—bromine (Br₂).
halfway between those of metals and nonmetals, and which are often referred to as semiconductors. They are useful in making photovoltaic cells for solar power panels (as we’ll see in Chapter 11), as well as in computer applications. The element hydrogen (H) shows up in two places in many periodic tables. This is because hydrogen has some properties in common with metals, and some properties in common with nonmetals. Its technical classification, however, is as a nonmetal.

**TRY THIS:**

Classify each element as a metal, metalloid, or nonmetal.

Fe, B, I, Kr, Ca, U, Si.

Answers: The metals are Fe (iron), Ca (calcium), and U (uranium). The metalloids are B (boron) and Si (silicon). The nonmetals are I (iodine) and Kr (krypton).

**Image 1.16. An “extended” periodic table of the elements.**

**Image 1.17. Copper (top left) has all the properties of metal: it’s shiny, ductile, conductive, and malleable. Sodium (top right) is a metal (shown here in liquid paraffin to keep it from reacting with air), but we normally don’t think of it as such, because we find it combined with other elements in compounds like table salt. Silicon (bottom right) is a metalloid (semiconductor) commonly used in computer chips. The nonmetals include gases such as chlorine (a toxic yellowish gas, bottom center), and dull, brittle, non-conductive solids such as iodine (bottom left).**
As you look at the different elements of the periodic table, you’ll notice a number of names among the metals that belong to substances you may never have thought of as metallic, like sodium (Na) and calcium (Ca). Both sodium and calcium are, in their elemental form, shiny metals. They have all the properties of common metals like aluminum and silver, but interestingly, they are soft. Sodium can actually be cut with a butter knife! The reason we don’t find these elements in their metallic form in nature is that they are very reactive and readily combine with other elements to form compounds. For instance, you probably encounter the element sodium most commonly as part of the compound sodium chloride (NaCl), which is table salt.

1.3 THINKING ABOUT CHEMICALS, CHEMICAL REACTIONS, AND CHEMICAL EQUATIONS

It’s fundamentally a bit difficult to be a student of chemistry, because unlike some of the other sciences, chemistry doesn’t really afford us the opportunity to see and manipulate the object of our examination. In biology, we can touch the fetal pig we’re dissecting, or turn over the leaf we’re sketching. Even bacteria, which are invisible to the naked eye, can be viewed with the help of a microscope. Chemistry is a little different. Even though we can go to the lab, mix a few chemicals and note a color change or bubbling or some other visible evidence of a reaction, we have to take our instructor or lab manual at face value when they tell us what the chemicals are and what reaction is occurring. We can’t see the molecules, and we can’t watch bonds break and reform. As a result, the study of chemistry involves employing a variety of visualization aids. Let’s take the example of water. We can think about water on a macroscopic scale, in terms of what we actually see—it’s a clear liquid. We can also think about water on a microscopic scale, in terms of how the atoms are arranged—an atom of oxygen is bonded to two separate atoms of hydrogen.

Finally, we can symbolically represent water in a concise way that provides information about its chemical composition—water has the formula H₂O. Depending upon what it is we’re trying to communicate or what we’re trying to imagine, we might pick any one or a combination of these strategies to visualize the compound water. For instance, if I wanted to communicate to you that the chemical methane (CH₄) burns in oxygen (O₂), making carbon dioxide (CO₂) and water (H₂O), it would be sufficient to symbolically represent chemicals with their formulas—no further information is needed. In Chapter 13, however, we’ll need to think about the shape of a water molecule in order to understand the greenhouse effect. Because of that, we’ll spend a lot of time thinking about its microscopic representation, which helps us to imagine the physical arrangement of the atoms in the molecule and its shape in space.

Image 1.18. Macroscopic (left), microscopic or molecular (top right), and symbolic (bottom right) representations of the compound water. (Havasu Falls picture © Scott Lefler, 2007.)
Let’s take a closer look at the chemical reaction mentioned in the previous paragraph. Methane (CH₄) is a member of the class of chemicals collectively called hydrocarbons, because they’re composed of carbon and hydrogen. We’ll be talking about hydrocarbons frequently in this textbook, in part because they all burn in oxygen (in chemistry, burning is called combustion) to produce carbon dioxide and water. Have you ever barbecued on a gas grill? If so, you’ve combusted the hydrocarbon propane (C₃H₈). If you’re a camper, you’ve probably used a camp stove that runs on the hydrocarbon butane (C₄H₁₀). Have you fueled your car recently? That was octane (C₈H₁₈) you put in the tank, unless you drive an alternative-fuel vehicle that runs on ethanol (C₂H₅OH).* When we describe the behavior of methane in oxygen, we are describing a chemical reaction, which is a rearrangement of atoms and molecules to form new molecules. The original chemical species are called the reactants, and the resulting chemical species are called the products. In the example above, methane and oxygen are reactants, while carbon dioxide and water are products.

In addition to describing a reaction in words, we can represent it symbolically. Let’s take the example of a very simple chemical reaction with only two reactants and one product. Sulfur (S) combines with oxygen (O₂) to produce sulfur dioxide (SO₂). Represented microscopically, the reaction looks like this:

\[ S + O_2 \rightarrow SO_2 \]

While this is a convenient way to visualize the rearrangement of atoms and molecules taking place in the reaction, it quickly becomes prohibitive to draw microscopic representations of reactants and products, especially in more complicated reactions. Chemical equations are symbolic representations of chemical reactions. The chemical equation for the reaction shown above would look like this:

\[ S + O_2 \rightarrow SO_2 \]

The arrow indicates that the reactants sulfur and oxygen* (S and O₂) are rearranging to form the product sulfur dioxide (SO₂). Aloud, this would read: sulfur and oxygen react to form sulfur dioxide.

You may be asking yourself right about now: Why would I want to know how to write out a chemical equation, since I’m not a chemist? That’s a valid

Try This:
In the reaction of carbon and oxygen to form carbon dioxide, what are the reactants and the products?
Answer: The reactants are carbon and oxygen, and the product is carbon dioxide.

*NOTE

Compounds like ethanol are called hydrocarbon derivatives—these compounds are mostly carbon and hydrogen, but contain small amounts of other elements, such as N, O, or S. Many of them, like ethanol, also combust in oxygen.

*NOTE

Why do we call O₂ oxygen when an atom of O is also called oxygen? This is a slightly confusing issue, but it’s worth addressing. Even though an atom of O is called oxygen, atomic oxygen is rarely found on its own in nature—it’s almost always combined into molecules. Elemental oxygen in nature is found as the molecule O₂. It’s so common, in fact, that rather than saying molecular oxygen or diatomic oxygen, we simply call it oxygen. Generally, we can figure out what is meant by the word oxygen from context. If I say CO₂ consists of one carbon and two oxygens, you pretty well know I mean atoms of oxygen. If I say the atmosphere contains 21 percent oxygen, you should know that elemental oxygen found in nature is always O₂. If the context leaves doubt as to what is meant, we generally reserve oxygen for O₂ and say atomic oxygen or an atom of oxygen for O.
question. The answer is that in order to talk meaningfully about chemistry in everyday life, we’ll need to speak a little bit of the language of chemistry, which is written in terms of reactants, products, and reactions.

1.4 BALANCING CHEMICAL EQUATIONS

Some chemical reactions end up causing us a bit of trouble when we try to write them out as chemical equations. Take, for instance, the true statement that elemental hydrogen (found in nature as a diatomic molecule, remember?) burns in oxygen to form water. When we write out an equation, we get something that looks like this:

\[ \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \]

Looking closely, it seems that we have a problem. There are two atoms of hydrogen (the subscript tells us that one molecule of hydrogen consists of two atoms of hydrogen chemically bonded together) and two atoms of oxygen on the reactant side of the equation. However, there are two atoms of hydrogen and only ONE atom of oxygen on the products side. A microscopic representation of the equation makes this very clear:

Where did the second atom of oxygen go? The answer is nowhere—atoms can’t disappear (nor can they appear!)—that’s a fundamental law of nature. The above is therefore not an accurate representation of the chemistry, because it’s not a balanced equation. Balanced equations obey the Law of Conservation of Matter, which means that atoms cannot appear, disappear, or change into atoms of other elements through any chemical reaction. In order for our equation to be chemically correct, we must ensure that there is the same number of each type of atom on each side of the equation. In order to do this, we can specify that different AMOUNTS of each reactant or product are involved in the reaction, but under no circumstances can we change the SUBSCRIPTS associated with any element in any molecule.* To specify that there is more than one of any given chemical involved in a reaction, we use a coefficient in front of the species (the lack of a coefficient in front of a species is always taken to mean one). In this case, we need two water molecules to balance the oxygen atoms, and two hydrogen molecules to account for the hydrogen needed to form the second water molecule. Our equation, correctly balanced, becomes:

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \]

Doing a quick atomic head count, this gives us four hydrogen atoms on each side of the equation, and two oxygen atoms on each side:

*NOTE

Tempting though it might be to balance the oxygen atoms by changing the product to \( \text{H}_2\text{O}_2 \), the chemical truth is that hydrogen and oxygen combine to form water, which is \( \text{H}_2\text{O} \). \( \text{H}_2\text{O}_2 \) (while appearing to balance the equation nicely) is NOT water. It’s hydrogen peroxide, which is a very different chemical indeed!
Even though we started with three molecules and ended up with two, we’ve conserved the numbers and identities of the atoms. In chemical reactions, reactants and products can contain different numbers of molecules, but they must contain exactly the same number and type of atoms, which is why we count atoms rather than molecules when balancing equations.

**Concept Check:**

Why did we have to change the number of hydrogen molecules and water molecules in the equation, but not the number of oxygen molecules?

Answer: By changing the number of water molecules (to get equal numbers of hydrogen atoms on each side of the equation), we ended up with equal numbers of oxygen atoms on each side of the equation.

Some equations are a bit more difficult to balance, because they involve more species. Let’s return, for instance, to the reaction between methane and oxygen that produces carbon dioxide and water:

\[ \text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

A cursory glance at this equation reveals that it’s not balanced. Unlike the hydrogen and oxygen reaction, however, this one is difficult to balance simply by looking at it. A good technique for balancing reactions like this follows a series of steps:

1) Identify an element that appears in ONE compound (do not pick elements that appear alone—save these for last!) on either side of the equation. Balance that element by adjusting the coefficients for molecules containing the element.

*Oxygen appears in our reaction as a free element (O₂), so we’ll save it for last. Carbon and hydrogen are both good Step 1 candidates, as they appear in only ONE compound on either side. We can therefore start with either carbon or hydrogen.*

\[ \text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]
Carbon is already in balance; there is one atom of carbon on each side of the equation.

2) Repeat Step 1 for remaining elements OTHER than free elements.

Hydrogen needs to be balanced; we can accomplish this by adding the coefficient 2 in front of the product \( \text{H}_2\text{O} \).

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}
\]

There are now four atoms of hydrogen on each side of the equation.

3) Balance any remaining elements.

Oxygen can now be balanced. There are four total oxygen atoms on the right—two from \( \text{CO}_2 \), and one from each of the two waters—so we need four total oxygen atoms on the left. We therefore add the coefficient 2 in front of the reactant \( \text{O}_2 \).

\[
\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}
\]

This is a balanced equation.

**TRY THIS:**

Balance each of the following equations:

\[
\begin{align*}
\_ \_ \text{S}_8 + \_ \_ \text{O}_2 & \rightarrow \_ \_ \text{SO}_3 \\
\_ \_ \text{N}_2 + \_ \_ \text{O}_2 & \rightarrow \_ \_ \text{N}_2\text{O} \\
\_ \_ \text{C}_{10}\text{H}_{16} + \_ \_ \text{Cl}_2 & \rightarrow \_ \_ \text{C} + \_ \_ \text{HCl}
\end{align*}
\]

Answers:

\[
\begin{align*}
\text{S}_8 + 12 \text{O}_2 & \rightarrow 8 \text{SO}_3 \\
2 \text{N}_2 + \text{O}_2 & \rightarrow 2 \text{N}_2\text{O} \\
\text{C}_{10}\text{H}_{16} + 8 \text{Cl}_2 & \rightarrow 10 \text{C} + 16 \text{HCl}
\end{align*}
\]

**1.5 NOMENCLATURE**

Shakespeare said in *Romeo and Juliet* that a rose by any other name would smell as sweet, but as it turns out, naming is critically important in chemistry! You’ve seen a few chemical formulas coupled with names so far in this chapter, so you may be getting familiar with the way the names tend to sound. For instance, it was mentioned earlier that a compound consisting of sodium (Na) and chlorine (Cl) is called sodium chloride, and I’ll tell you right now that a compound of rubidium (Rb) and fluorine (F) is
called rubidium fluoride—there’s a pattern there. Just to throw a wrench in the works, however, a compound of carbon and oxygen is NEVER called carbon oxide. WHY? The short answer is that there are several ways in which carbon and oxygen can combine,* and using the (incorrect) name carbon oxide doesn’t help us to distinguish them from one another. Clearly, we need to outline some rules for how to name compounds. Knowing a little bit of nomenclature (which means a system for naming) helps us to understand one another when we are discussing chemistry. The easiest compounds to name are binary compounds, which are made up of two (and only two) different elements. There are lots of compounds made of more than two elements, but their nomenclature is much more complex. Of course, anything made up of only one element is not a compound at all. Binary compounds may consist of a metal and a nonmetal, or a nonmetal and a nonmetal; there are no compounds of metals with metals.

Binary Compounds—Metal and Nonmetal

To name a binary compound made of a metal and a nonmetal, we use the name of the metal followed by the name of the nonmetal, adding the suffix -ide to the nonmetal. You may need to drop some letters to do this phonetically—sulfur becomes sulfide, and oxygen becomes oxide.

Example: A compound of the metal barium (Ba) and the nonmetal oxygen (O) is called barium oxide.

Example: The compound LiF is called lithium fluoride.

Note that chemical formulas for binary compounds of metals and nonmetals are written in the same order in which the name is given—metal first.

When there are subscripts in the chemical formula of a compound, we ignore them in naming the compound.

Example: The compound AlCl3 is called aluminum chloride.

The reason we don’t reference the relative number of atoms of each element when we name the compound is that metals can only combine with a given nonmetal in ONE possible way. In other words, when lithium and fluorine combine, the ONLY possible combination is LiF. When aluminum and chlorine combine, the ONLY possible combination is AlCl3. How

Try This:

Name each of the following: K₂O, MgS, AlF₃.
Answers: Potassium oxide, magnesium sulfide, aluminum fluoride.

*NOTE

Two examples include CO (carbon monoxide) and CO₂ (carbon dioxide).