

Stop Faking It!

Finally Understanding Science So You Can Teach It

MORE CHEMISTRY BASICS

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By William C. Robertson, PhD
In consultation with Michael S. Kralik, PhD,
and Ann Cutler, PhD
Illustrated by Brian Diskin

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Preface

Prior to the publication of this book, teachers often asked me whether the book would be more advanced than the first book, *Chemistry Basics*. The answer to that question is yes and no. First, this book builds on concepts discussed in the first book. Although the first chapter of this book is a review of the concepts covered in the first book, the review will not suffice if you are not somewhat familiar with those concepts. So, you need to know a few things about basic chemistry before tackling this book. Second, there are a number of topics in this book that might be considered advanced chemistry, such as the uncertainty principle in quantum mechanics, the formation of different kinds of bonds, and oxidation-reduction reactions. Many of the reviewers of the book said things such as, “The teachers in your audience will never teach some of these concepts to their students. The students won’t see some of these concepts until high school or even college.” That might be true, but I believe that your understanding of the so-called higher-level concepts will enhance your teaching. When little Susie says, “These orbits of electrons aren’t correct; my mom says we can’t really know what electrons are doing,” what will be your response? It could be, “We don’t discuss that in sixth grade—you’ll learn that in high school.” But if you actually know what Susie is talking about, you have a choice. You can discuss it with her individually, or you might even decide that the subject will be great for a class discussion. If you don’t know what Susie is talking about, you have no choice and no control over the situation. You need to ask yourself whether you want the content of your curriculum to be in charge of what happens, or if you want your knowledge of the subject matter to be in charge of what happens. I prefer the latter. Also, it is empowering to know more than you will ever need to teach the students. It gives you a comfort level and just might affect how you address certain topics. Teaching is more fun when you are secure in the depth of your understanding.

One more thing on this topic. There is a lot of tradition in the teaching of science. You simply don’t address certain concepts until you get to a particular grade level. This implies that some ideas are much more difficult to understand than others. I strongly believe, however, that anyone can understand, say, the basics of quantum mechanics. Yes, a complete mathematical treatment of the

Preface

subject requires a sophistication with math, but the basic concepts are not that difficult to understand. You can avoid the math complexities (as I do, for the most part, in this book) and still grasp the major concepts. As someone (I forget who) said, you can teach anyone of any age just about any concept, as long as you do so in an intellectually honest manner. This means that anyone can learn any concept, but you must take care not to alter the concept in the process to the point that you are teaching lies, which unfortunately happens too often when we try to “dumb down” certain material. Suffice to say that I try to remain intellectually honest in everything I write.

Content aside, this book is arranged the same as all the other books in the *Stop Faking It!* series. I ask you to do a few simple activities prior to the explanation of any concepts. We understand concepts better when we can anchor them to concrete experiences, and that’s the purpose of these sections. It might be tempting to just read through the Things To Do sections, but I strongly encourage you to do the activities. They could make the difference between memorizing concepts and truly understanding them. Memorization fades, but understanding lasts.

Finally, I want to be clear that this is not a comprehensive chemistry book, even when combined with the first chemistry book in the series. For example, I do not deal with equilibrium in chemical equations or solubility of solids immersed in solutions, both traditional chemistry concepts. I also do not do justice to the vast field of organic chemistry. This book is not a textbook, and none of the books in the *SFI* series are intended to be comprehensive textbooks. My goal is to help you obtain a deep understanding of a number of basic chemistry concepts. I hope that with this understanding, you will know what those short descriptions of content in activity books are all about and will have the basis for knowing what’s going on in a textbook. Also, although there are activities in this book, the activities are there to help you understand concepts and are not necessarily intended as classroom activities. That’s not to say that you can’t use some of these activities in the classroom, but rather that you should consider safety and appropriateness before using the activities with your students.

Safety Note

Though the activities in this volume don't require anything more volatile than household vinegar, safety should always be in the forefront of the mind of every teacher. (This is not intended as a book of classroom activities, by the way. Rather, the activities are designed to enhance your understanding of the subject before you get into the classroom.) Your individual school, or possibly the school system of which your school is a part, likely has rules and procedures for classroom and laboratory safety.

You can also find specific guidelines for the safe storage, use, and disposal of thousands of types of chemical products in the Material Safety Data Sheets (MSDS). Start with <http://www.ilpi.com/msds/#Internet>. This site links to dozens of free searchable databases, including those of top American and European universities.

NSTA has also published several award-winning titles covering the safety theme at all school levels. For the elementary level, there's *Exploring Safely: A Guide for Elementary Teachers* and the *Safety in the Elementary Science Classroom* flipchart. Middle school-level offerings are *Inquiring Safely: A Guide for Middle School Teachers* and the *Safety in the Middle School Science Classroom* flipchart. Finally, there's *Investigating Safely: A Guide for High School Teachers*.

How to Get Special Materials

You can visit the following websites to get supplies for chemistry activities.

Arbor Scientific. www.arborsci.com/SearchResult.aspx?CategoryID=4

Carolina. www.carolina.com

Edmund Scientific. <http://scientificsonline.com>

Educational Innovations. www.teachersource.com/Chemistry.aspx

Hi-Valley Chemical. www.hvchemical.com

(*Note:* Although other suppliers require that you are affiliated with a school or other institution, Hi-Valley will sell supplies to private individuals—which is important for home schoolers or others who are not affiliated with an institution.)

Sargent-Welch. <http://sargentwelch.com/chemicals/c/4749>

Science Kit. <http://sciencekit.com>

Ward's Natural Science. http://wardsci.com/category.asp_Q_c_E_1251

About the Author

As the author of NSTA Press's *Stop Faking It!* series, Bill Robertson believes science can be both accessible and fun—if it's presented so that people can readily understand it. Robertson is a science education writer, reviews and edits science materials, and frequently conducts inservice teacher workshops as well as seminars at NSTA conferences. Bill has published research in cognitive science that addresses the difference between memorizing and understanding and how that affects learning. He has also taught college-level physics and math and developed K–12 science curricula, teacher materials, and award-winning science kits. He earned a master's degree in physics from the University of Illinois and a PhD in science education from the University of Colorado.

About the Illustrator

The recently-out-of-debt, soon-to-be-famous, humorous illustrator Brian Diskin grew up outside of Chicago. He graduated from Northern Illinois University with a degree in commercial illustration, after which he taught himself cartooning. His art has appeared in many books, including *The Beerbellie Diet* and *How a Real Locomotive Works*. You can also find his art in newspapers, on greeting cards, on T-shirts, and on refrigerators. At any given time he can be found teaching watercolors and cartooning, and hopefully working on his ever-expanding series of *Stop Faking It!* books. You can view his work at www.briandiskin.com.

About the Consultants

Michael Kralik received his PhD in chemistry from the University of Utah with postdoctorate studies in chemistry, pharmacology, and toxicology. He has been faculty at the university and has conducted many faculty and staff development seminars. Kralik has established product development and manufacturing operations domestically and internationally, and has directed the development of hundreds of products for major corporations in chemical, medical, pharmaceutical, and electronics industries. He has developed K–12 science curricula, teacher inservice workshops, and many award-winning educational toys, games, and science kits.

Ann Cutler is the field editor for the *Journal of College Science Teaching*, a position she has held since 2006. She is an associate professor of chemistry at the University of Indianapolis in Indiana and has a doctorate in inorganic chemistry from Purdue University.

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This book was reviewed by Dennis Huffman (Lanier Middle School, Texas), Sue Vogel (Thomas Jefferson High School, Iowa), and Peggy Carlisle (Pecan Park Elementary School, Mississippi).

Dedication

I dedicate this book to two people. The first is Michael Kralik, a good friend and valuable colleague for many years who has served as a consultant on both chemistry books. There's a saying in science education that all activities are recycled—there really isn't anything new. People who say that have never met Michael, who consistently comes up with inventive ways to introduce concepts. The second person is Ann Cutler, who was a reviewer for both chemistry books. I can honestly say that Ann has fundamentally changed the content of both books and, with this second book, saved me from embarrassing myself with hastily written, wrong explanations.

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Special Reactions

I covered the basics of chemical reactions in the first chemistry book, including how to write and balance chemical equations that represent those reactions. There is also a quick review of chemical reactions in Chapter 1 of this book. We're going to delve a bit deeper into certain kinds of reactions—ones that occur often enough that it's useful for them to have their own special categories. As you go through this chapter, try to avoid seeing these separate categories as conceptually different from other reactions. The same basic principles govern all chemical reactions, regardless of the category.



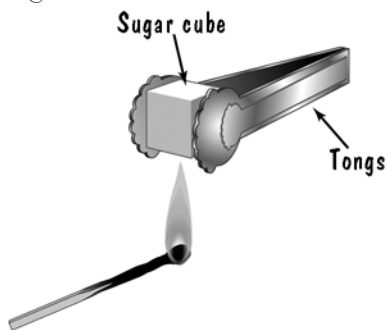
4 Chapter

Things to do before you read the science stuff

Here's a neat trick to amaze friends and family members. Challenge someone to make a sugar cube catch on fire. For safety's sake, hold the sugar cube with tongs and place a lit match or a candle under it. See Figure 4.1.

Although your chosen person might get the sugar cube to melt or caramelize, he or she will be out of luck when it comes to the cube going up in flames. You come to the rescue with a new sugar cube. Before you place the cube in the tongs, dust it on all sides with cigarette ashes.¹ When you place your dusted cube over a flame, it should catch fire easily. Neat.

Figure 4.1



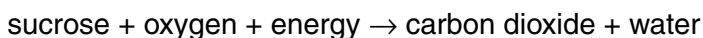
The science stuff

What you just did is an example of a **catalyzed reaction**. A **catalyst** is something that helps a reaction occur or increases the rate of the reaction. Although catalysts affect a reaction, they themselves are not affected by the reaction as a whole. The particular kind of reaction you catalyzed when getting the sugar cube to burn is known as **combustion**, which is just a fancy name for something burning. When a substance burns, oxygen combines with carbon and hydrogen atoms to form carbon dioxide and water. The products—carbon dioxide and water—are always the same in combustion. Therefore, we have a classification known as combustion reactions. All of them are of the form



The term **organic molecule** refers to molecules that contain primarily carbon and hydrogen.

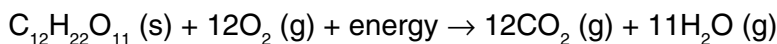
The above reaction doesn't include the fact that you often have to add energy (via the flame) to get the reaction to go forward. For our sugar cube (composed of sucrose), we can write this as



or using chemical formulas (and balancing the equation):²

¹ For the record, we here at the Stop Faking It! Institute of Higher Learning do not promote smoking. I have to admit, though, that I love the smell of a good cigar.

² Those letters in parentheses—(s), (g), and (aq)—refer to the state of the chemical involved in the reaction. These symbols refer to solid, gas, and aqueous, respectively.



Whenever you have to add energy to get a reaction to occur, that energy is referred to as the **activation energy**. We can represent the activation energy with a graph of the energy involved in the reaction versus time, as shown in Figure 4.2.³

You can think of the activation energy shown in Figure 4.2 as an “energy hump” you have to get over for the reaction to proceed. Imagine the graph is a track on which you can place a bowling ball. For the bowling ball to go from the energy level on the left side to the lower energy on the right side (you eventually get energy *out* of this process because the ball is now rolling faster on the right side), you have to push the ball (add energy) up to the top of the hump. See Figure 4.3.

Now let’s apply this idea to the sugar cube. For the cube to burn, you have to get over the energy hump to the right side of the graph. A match simply can’t provide enough energy to do this. You can solve the problem by

Figure 4.2

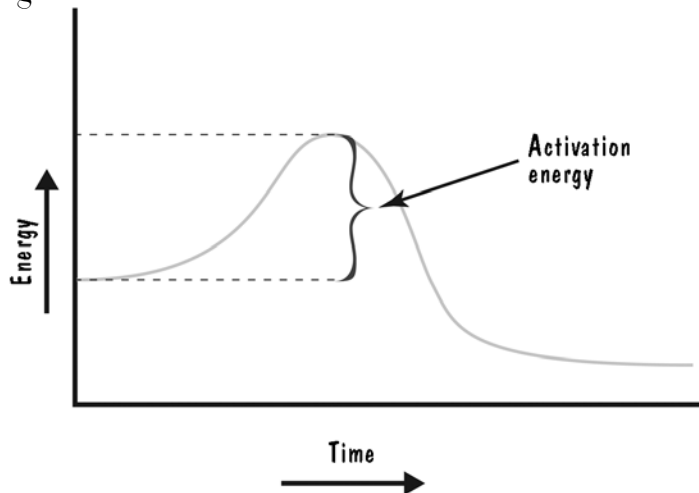
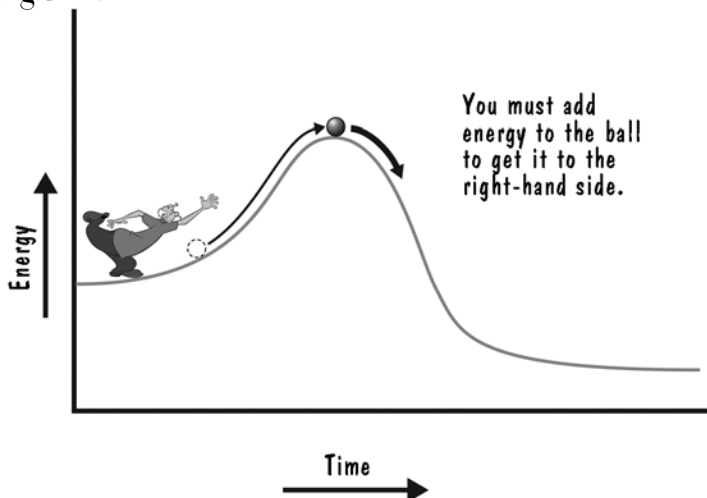


Figure 4.3

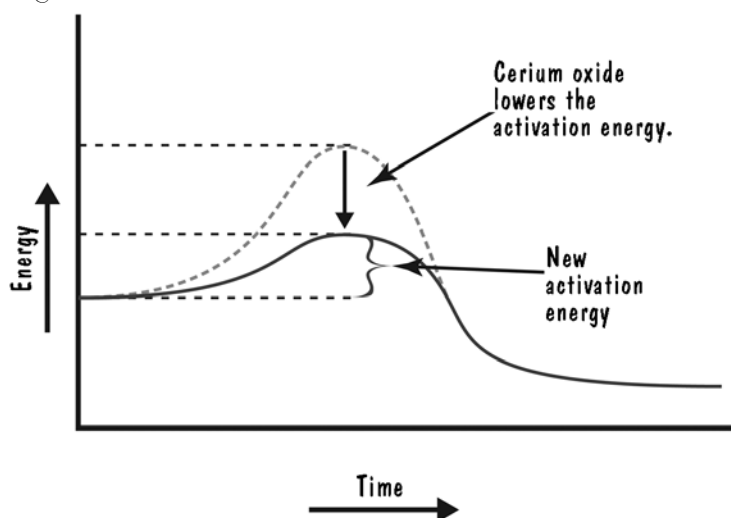


³ There’s potential for confusion in this graph because it looks similar to energy level diagrams for individual atoms, such as Figure 1.9 (p. 9) in Chapter 1. The vertical axis is energy in both cases, but the horizontal axis in Figure 4.2 is time, while the horizontal axis in Figure 1.9 has no meaning. In Figure 4.2, we are tracking the energy of the molecules involved in a chemical reaction as the reaction proceeds, and in Figure 1.9 we are illustrating the energy levels available to electrons in an atom.

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adding more energy than a match can provide.⁴ Alternatively, you can *lower* the activation energy, which is where the cigarette ash comes in. The ash contains a substance called cerium oxide. The cerium oxide acts as a go-between for the oxygen molecules in the air to combine with the sucrose molecules. With cerium oxide as the mediary, the energy required for the combustion of sucrose is lower than without the cerium oxide, low enough that the energy provided by a lit match is enough for the reaction to proceed. See Figure 4.4.

Figure 4.4



All catalyzed reactions use this same principle. By adding a catalyst to a reaction, you lower the activation energy necessary for the reaction to occur. You might think of the catalyst as a facilitator. You have two molecules that have trouble getting together on issues. In comes the catalyst as mediator, helping the two sides get together. As catalysts often say, “Can’t we all just get along?”

More things to do before you read more science stuff

In this section and the next one, I’m not going to deal with reactions, but rather with chemicals that we classify as acids and bases. This little aside is necessary because following these two sections I’m going to deal with acid-base reactions. Seems silly to cover that before you know what acids and bases are! Anyway, for this section you’ll need a supply of litmus paper (both red and blue) and a supply of pH paper. You can get these from a science materials supplier or possibly at a place that sells products for swimming pools or spas. You’ll also need various liquids to test. Good ones that you can find around the house are vinegar, bleach, black coffee, ammonia, water, baking soda (dissolved in water), and lemon juice. Chemicals that are available from a science supply outfit or maybe already lying around your school include sodium hydroxide, calcium hydroxide, hydrochloric

⁴ For example, you might get a sugar cube to go up in flames if you trained a blowtorch on it.

acid, and sulfuric acid. If you have these chemicals in different concentrations (different molarities)⁵, that would be good.

Caution: Commercial chemicals such as hydrochloric acid and sodium hydroxide can come in concentrations that are downright dangerous. When using these chemicals and somewhat noxious household chemicals such as ammonia, always wear goggles and gloves and have a plan for washing out your eyes if they come in contact with the chemicals. Also, **take care not to mix any of the liquids you have**. Mixing ammonia and bleach, for example, will generate a dangerous gas. Not good (I inadvertently did this once and it's scary!). See Safety Note (p. ix).

Okay. First, dip strips of blue litmus paper in the various liquids you have. Use a new strip for each liquid. The paper will either stay blue or turn red. Then dip strips of red litmus paper in the various liquids. The paper will either stay red or turn blue. Keep track of what liquids cause what reactions in the litmus paper.

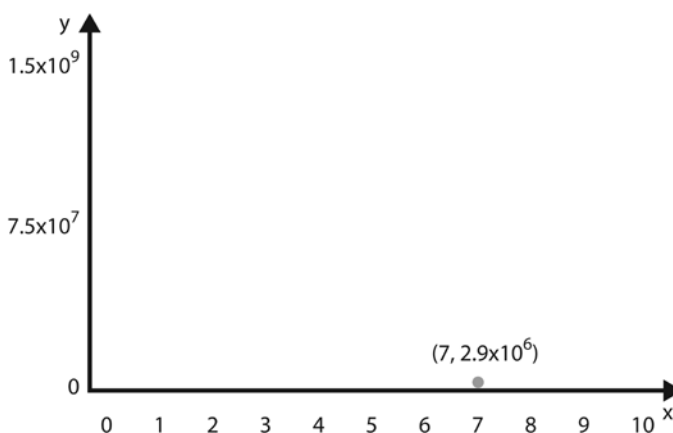
Next get your pH paper and dip those strips (a new one for each liquid) into each liquid. You can compare the color of the strip with a color code that comes with the pH paper to determine a number for the pH—something between 0 and 14.

Table 4.1

x value	y value
1.0	0.00000002
2.0	0.00006
3.0	0.004
4.0	0.8
5.0	25.0
6.0	3200
7.0	2900000
8.0	1400000000

Now for something completely different. Listed in Table 4.1 are a bunch of ordered pairs of numbers. You're going to try to graph these ordered pairs, with the x value going on the x -axis and the y value going on the y -axis. Figure 4.5 shows a sample graph with the first ordered pair plotted on the graph. Use the sample graph or draw one of your own to try to get all of the ordered pairs on one graph. Hint: It's pretty much impossible. You can give up once you know *why* it's impossible.

Figure 4.5



⁵ Refer to Chapter 2 of this book for a definition of molarity. We use a capital M for molarity, so a label of 0.5 M means a chemical has a molarity of 0.5 moles per liter. Also, try to keep to low molarities of these substances.

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Now use a calculator to take the logarithm of each of the y values and create a new table. I've done the first few and the last one for you below in Table 4.2.

Try graphing this second set of ordered pairs. Shouldn't be too difficult as long as you allow for negative y values.

Table 4.2

x value	Logarithm of y value
1.0	-7.7
2.0	-4.2
3.0	-2.4
4.0	?
5.0	?
6.0	?
7.0	?
8.0	9.1

More science stuff

You no doubt have used the terms *acid* and *acidic* in your everyday life. Many fruits contain citric acid, and we often refer to things we eat as being acidic or non-acidic. Movies let us know that acids can burn you (think *RoboCop*), especially the famous acids hydrochloric acid and sulfuric acid. People can have an acid tongue, but that's something different. Less common in everyday language are the words *base* and *basic*, at least as they apply to chemicals. The term *pH* is also common. We buy pH-balanced shampoos and check the pH of hot tubs and pools.

You probably knew before you even did the previous section that litmus paper helps determine whether a substance is an **acid** or a **base**. If blue litmus paper turns red when dipped in a liquid, the liquid is an acid. If red litmus paper turns blue in a liquid, the liquid is a base. If neither the red nor the blue litmus paper changes color, the liquid is neutral. pH paper takes this a step further and shows *how* acidic or basic a substance is.



Topic: Acids and Bases

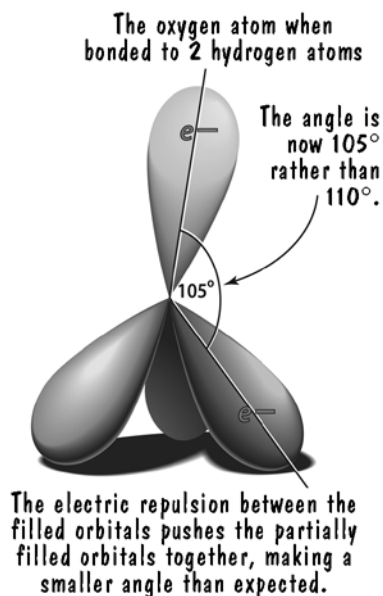
Go to: www.scilinks.org

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Great. Now if we only knew what acids and bases *are*! There are a number of ways to define acids and bases, and I'll give you two of them here. Before doing that, I need to remind you of how water interacts with other chemicals. Water molecules are **polar**, with one end being positive and the other negative. See Figure 4.6.

Figure 4.6 depicts the atoms hydrogen and oxygen. Be careful you don't confuse this drawing with the drawings in Chapter 3 that showed electron orbitals and how they interact. Of course, we can use the orbital drawings to explain exactly why oxygen is a polar molecule. Figure 3.29 is shown again on the next page. Notice that the two orbitals that do not bond with hydrogen atoms are filled with two electrons each. This makes the side of the molecule away from the hydrogen atoms more negative. Thus, a polar molecule.

Figure 3.29



plus side of the water molecule (the hydrogen side) is attracted to the negatively charged chloride ion, and the negative side of the water molecule (the oxygen side) is attracted to the positively charged sodium ion.⁶

This tendency of water molecules to break apart other molecules is part of the first definition of acids and bases, known as the **Arrhenius theory of acids and bases** and named after (surprise!) a chemist named Svante Arrhenius. According to Arrhenius, an acid is a substance that increases the concentration of H^+ ions in an aqueous (water) solution. So, hydrogen chloride (HCl) dissolved in water is an acid because the water breaks this molecule up into H^+ ions and Cl^- ions. Actually, this statement is a bit of a lie because H^+ ions (which are simply protons—take away an electron from hydrogen, which is com-

Because water molecules are polar, they interact electrically (you know, like charges repel and unlike charges attract) with other molecules. For example, when you put sodium chloride (table salt) in water, the attraction between different parts of the water molecule and the ions that make up sodium chloride are strong enough that water molecules separate the sodium chloride into positive sodium ions and negative chloride ions, as in Figure 4.7. Notice in this figure that the

Figure 4.6

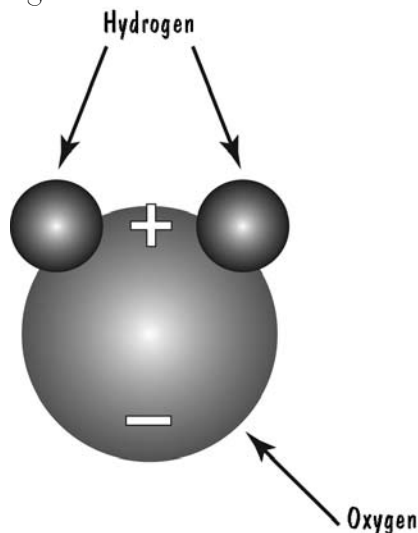
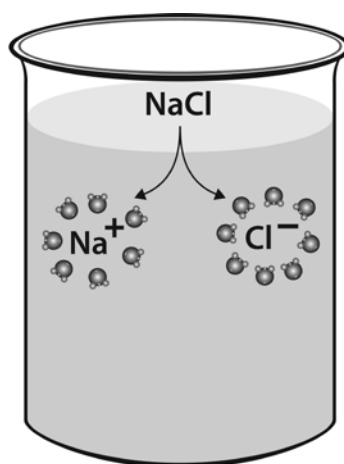


Figure 4.7



⁶ When we break apart molecules with ionic bonds, we say they have been **dissociated**. From experience, you know that the salt disappears in this process because it has dissolved into the water. While molecules with covalent bonds can dissolve in liquid (such as sugar in water), covalent bonds don't dissociate in the process. They retain their molecular identity.

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posed of one proton and one electron in its most common state, and all you're left with is a proton) don't exist by themselves in water. Instead, they latch onto water molecules to form what are known as hydronium ions, or H_3O^+ . So really, an acid is anything that increases the hydronium ion concentration in an aqueous solution. Check out Figure 4.8.

Figure 4.8

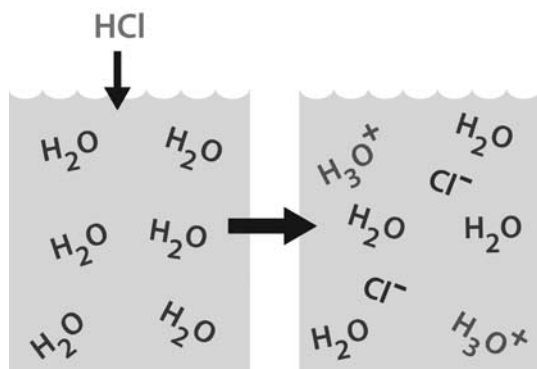
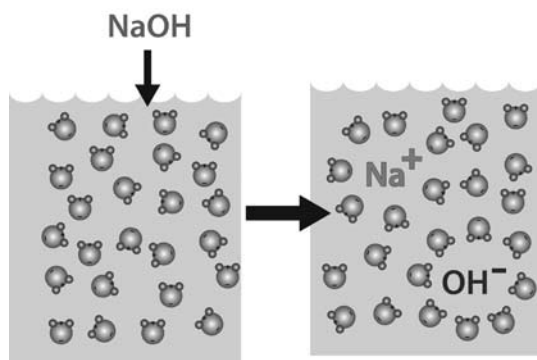


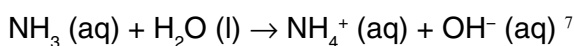
Figure 4.9



Arrhenius defines a base as anything that increases the concentration of hydroxide ions (OH^-) in an aqueous solution. For example, sodium hydroxide (NaOH) is a base because when you put it in water, the polar water molecules separate it into sodium ions (Na^+) and hydroxide ions (OH^-). Figure 4.9 shows this.

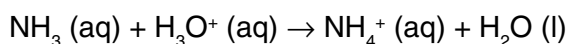
Okay, so what? What's so special about substances that form hydronium ions or hydroxide ions in water? Well, it has to do with what those substances do when they encounter other substances. If a substance produces lots of positively charged hydronium ions, then those ions have a tendency to interact with negatively charged ions. And if a substance produces lots of negatively charged hydroxide ions, then those ions tend to interact with positively charged ions. In other words, we define substances as acids and bases because that helps us figure out how they'll interact in chemical reactions, which is what most of chemistry is about.

I said I would give you two definitions of acids and bases, so I guess I'd better provide the second one. It turns out that molecules containing the hydroxide ion aren't the only ones that increase the concentration of hydroxide ions when dissolved in water. For example, here's what happens when ammonia (NH_3) dissolves in water:



⁷ In case you forgot, (g) means the substance is a gas, (s) means the substance is a solid, and (aq) means the substance is in a solution. One classification that isn't represented here is (l), which indicates a substance that is itself a liquid.

NH_3 contains no hydroxide ions, yet when you put it in water it increases the hydroxide ion concentration. To account for examples such as this, we have what is known as the **Brønsted-Lowry** (Johannes Brønsted and Thomas Lowry) **theory of acids and bases**. In this definition, an acid is any substance that donates protons in a reaction, and a base is any substance that accepts protons in a reaction. This is basically (no pun intended) the same as the Arrhenius definition of acids because if a substance readily donates protons in a reaction, it will also increase the hydronium ion concentration when placed in water. The definition of a base, though, is expanded to include any substance that readily accepts protons in a reaction, and that substance doesn't have to contain hydroxide ions. Let's continue with our above example. There, NH_3 contributed to the OH^- concentration when dissolved in water. But NH_3 is also a proton acceptor (which we have defined as a base) because it picks up a proton to become NH_4^+ , as in



Let's head back to your use of pH paper and why I had you attempt to graph those sets of numbers. You probably know that pH is a measure of how acidic something is, but you might not completely understand the pH scale, which is what I'm going to explain. First, let's talk about the two graphs I had you attempt to create. The y values in the first table (Table 4.1) have a wide range, from really small (0.00000002) to really large (1,400,000,000). If you make the graph so you can distinguish the small values, then the larger ones are on top of one another, and vice versa. Logarithms to the rescue. If you plot the logarithms of the y values instead of the original y values, then you can graph them easily.

The concentrations of hydronium ions in aqueous solutions are like the numbers in Table 4.1 in that they cover a wide range of values. For example, a strong acid might have a concentration of hydronium ions of 0.01 M, and a weak acid might have a hydronium concentration of 0.000001 M.⁸ We use logarithms to compare these numbers and graph them side by side. This leads us to the definition of pH, which is expressed in terms of hydrogen ion concentration, even though we realize that in the real world we're dealing with hydronium ions rather than hydrogen ions.⁹ Here's the definition:

pH = negative of the log of the hydrogen ion concentration

⁸ Remember that M stands for molarity, which is the number of moles of a substance per liter of solution.

⁹ Many things in science are done in the name of tradition. The definition of pH was developed before chemists realized that hydrogen ions don't exist when in water. You will sometimes see the expression p(Hydronium) instead of pH, but for the most part, the term pH persists.



Topic: pH Scale

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Code: MCB012

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Chemists use brackets to indicate concentrations, so we rewrite the hydrogen ion concentration as $[H^+]$, and we have

$$pH = -\log [H^+]$$

The reason for the negative sign is that most concentrations are less than 1.0, and as you saw previously, taking the log of these numbers gives you a negative number. So, the negative sign is just to make most pH values positive.

The nature of logarithms. How do logarithms (or logs) reduce a wide range of values to a small range of values? To understand that, you simply need to know the definition of a logarithm. To take the logarithm of a number, you ask, “10 raised to what power (exponent) will give me this number?” For example, $\log 100$ is equal to 2 because 10 raised to the second power (10 squared) equals 100. Similarly, $\log 10,000$ is equal to 4 because 10 raised to the fourth power equals 10,000. For numbers less than 1, negative exponents come into play. $\log 0.01$ is equal to -2 because 10 raised to the -2 power equals 0.01. So, just using the numbers I’ve mentioned here, the numbers 0.01, 100, and 10,000 become, after taking the logarithms of them, -2 , 2, and 4. Voilà! A wide range of numbers reduced to a small range of numbers that are a bit easier to work with.

Assuming you used pH paper when I asked you to, you discovered the pH values of a few common liquids. The pH of lemon juice is about 2.3, the pH of vinegar is about 3.0, the pH of orange juice is about 3.5, and the pH of ammonia is about 11.0. Keep in mind that the higher the pH, the smaller the hydronium ion concentration and the greater the hydroxide ion concentration. The usual range of pH values chemists use is from 0.0 (highly acidic) to 14.0 (highly basic), even though values lower (and even negative) and higher are possible. A pH of 7.0 is considered neutral. And remember that pH is a logarithmic scale. That means that a pH of 5 is 10 times more acidic than a pH of 6 and 100 times more acidic than a pH of 7.

Before going on, I want to make a distinction between *strong* acids and bases and *concentrated* acids and bases. Hydrochloric acid is a strong acid because when it’s placed in water, its separation into H_3O^+ ions and Cl^- ions is pretty much complete. Substances that readily donate protons are strong acids. However, you could have a solution of hydrochloric acid that has very little HCl dissolved in water or a lot of HCl dissolved in water. In other words, you could have a low *concentration* of a *strong* acid that isn’t all that dangerous. If you’ll recall, we measure concentration in molarity, so a 0.001 M solution of hydrochloric acid (0.001 moles of HCl per liter of solution) is much weaker overall (and would have a higher pH) than a 0.1 M solution of hydrochloric acid (0.1 moles of HCl per liter of solution), even though you would consider them both to be “strong” acids. And what I just said about acids applies to bases. The moral is to pay just

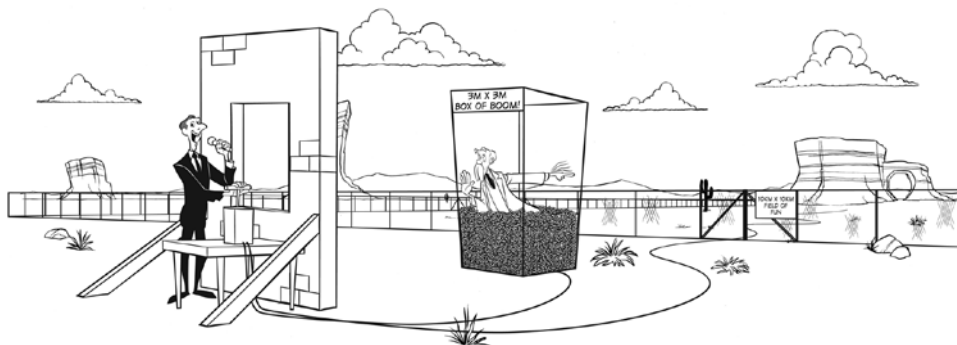
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as much attention to the molarity of an acid or base as you do to whether the acid or base is considered to be strong or weak.

Let's try an analogy to illustrate the difference between strong and weak on one hand and concentrated or not concentrated on the other hand. Suppose someone is going to set off firecrackers in your vicinity, and you want to be as safe as possible so as not to incur physical damage. There are tiny firecrackers and really big firecrackers. All other things being equal, you might choose to be in the vicinity of tiny firecrackers. That would be analogous to choosing a weak acid over a strong acid. But let's say this person is going to set off 100 fireworks, and there are two options—they'll be set off in a 3×3 meter closed room or in a 10×10 kilometer field. The different locations correspond to different concentrations. 100 firecrackers in a small room versus 100 firecrackers in a huge open field. The first option is a high concentration of firecrackers and the second is a low concentration of firecrackers. The difference in concentration is big enough that you could choose the large firecrackers (strong acid) in the large field (low concentration) and be safer than if you chose the small firecrackers (weak acid) in the small room (high concentration). So, the strength of an acid or base is important, but so is the concentration.



"That's correct!! You want to be near the small firecrackers. But you're going to have to move faster, Professor, if you want to make it to the next round."



"Our contestant is signaling he wants to be enclosed in the 3 m by 3 m case and not the 10 km by 10 km field. My, he certainly is intrepid, isn't he."

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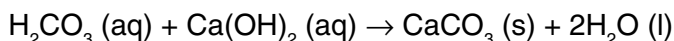
Even more things to do before you read even more science stuff

We covered what acids and bases are, so it's time for reactions. For this activity, you'll need some calcium hydroxide (about 0.1 M concentration) dissolved in distilled water (get the powder from a science chemical supplier)¹⁰, a straw, a small amount of *clear* carbonated soda (club soda or lemon-lime soda), a pair of safety goggles, and your breath. Pour a small amount of the calcium hydroxide solution (the liquid you made, not the powder) in a clear glass and notice what color it is (clear, yes?). Use the straw to blow bubbles into the liquid (possible splashback is why you have goggles). Notice any change in the calcium hydroxide (there *should* be a change). Rinse out the glass and pour a small amount of clear carbonated soda into it. By definition, this is clear. Then pour a small amount of calcium hydroxide into the glass and notice any change.

Even more science stuff

If all went well, cloudiness happened. When you blow bubbles into clear calcium hydroxide, it turns cloudy. When you pour clear calcium hydroxide into clear carbonated soda, the solution turns cloudy. The explanation for both of these is that you observed an acid-base reaction. In each case, the cloudiness came from the interaction of an acid—carbonic acid (H_2CO_3)—and a base—calcium hydroxide. A carbonated soda already has carbonic acid in it. When you blow bubbles into calcium hydroxide, the carbon dioxide (CO_2) from your breath interacts with the water in the calcium hydroxide solution to form H_2CO_3 . Here's the balanced equation for the interaction of carbonic acid and calcium hydroxide:

carbonic acid + calcium hydroxide $\rightarrow$ calcium carbonate + water



Quick review. How did I know the formulas for carbonic acid, calcium hydroxide, and calcium carbonate? Yeah, I could have looked them up, but I didn't. I simply used the Periodic Table along with the knowledge that OH^- has a negative charge (an extra electron) and that CO_3^{2-} has a -2 charge (two extra electrons). Also, if you don't know how I balanced the equation by using two waters, it might be worth reviewing how to balance equations.

Calcium carbonate is an insoluble salt (doesn't dissolve in water), which is why things turned cloudy. In fact, all acid-base reactions have the same result.

¹⁰ The concentration of your dissolved calcium hydroxide isn't critical. About a teaspoon of powder in a cup of distilled water will do the trick. You can also buy solutions of calcium hydroxide so you don't have to bother dissolving any powder into distilled water.

You add the acid and base together, and you get a salt (there are many kinds of salt that can be produced) plus water. The production of salt isn't always easy to see, though. For example, if you mix vinegar (an acid) and ammonia (a base), you won't see any cloudiness. The salt produced in this case (ammonium acetate) stays dissolved in the solution.

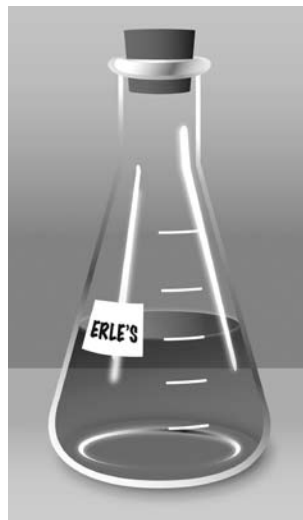
Now, you might be thinking that we could figure out this reaction without the whole notion of acids, bases, and acid-base reactions. You would be right. You could simply look at the relative attractions for electrons of the atoms and compounds involved and know that certain bonds are going to be stronger than others.¹¹ You could also know that carbonic acid is very likely to give up a proton and calcium hydroxide is very likely to accept a proton in this reaction, without classifying them as acids and bases. The bottom line, though, is that there are so many acid-base reactions that it's useful for them to have a category all by themselves. Simply knowing that you are starting with acids and bases tells you what will happen in the reaction without going through all that other stuff.

And even more things to do before you read even more science stuff

You'll need a few special chemicals for this one. The first is methylene blue, which is an acid-base indicator. You just need a tiny amount (maybe 10 drops) of this. The second is glucose (also known as dextrose anhydride), which you can get as a powder in small amounts from a science supply outlet. The third is potassium hydroxide, which you can get as a powder or solution. You'll also need a large glass container with a tight-fitting lid. To make it look all chemistry-like, people traditionally use a 500 ml Erlenmeyer flask with a rubber stopper (see Figure 4.10), but that's not absolutely necessary. Just make sure there's no danger of spilling what you end up with because it can irritate your skin.

Now, if I wanted to be mean, I'd give you the molarities of the solutions you need and have you use what you learned in Chapter 2 to make up the solutions. Because I'm a nice guy, though, here's what to do.

Figure 4.10



¹¹ How we do this is discussed in the first *Stop Faking It!* chemistry book. It involves the concept of electronegativity (a measure of affinity of atoms for electrons) and comparison of the numbers associated with electronegativity.

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- Do the following preparations within ten or fifteen minutes of doing the activity. Leave things around for a day, and it won't work.
- Add about 8 grams of potassium hydroxide (KOH) to about 300 ml (that's milliliters) of distilled water in the flask and swirl it to mix thoroughly.
- Add about 10 grams of glucose (dextrose) to this solution and allow it to dissolve.
- Add 6 to 8 drops of methylene blue to the flask and swirl things around again.
- Let everything rest until the solution becomes clear.

Now for the fun part. Stopper the flask and shake up the solution inside. It should turn blue. Let it sit for a bit and it should clear up. Shake it again, and it should turn blue again. You can repeat quite a few times before there's no longer a color change.

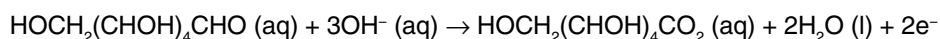
And even more science stuff

First, the quick and dirty explanation of what happened. When you shake up the solution, you dissolve oxygen from the air in the container into the solution. The oxygen combines with the molecules in the solution, resulting in a change to the blue color. Then there's a spontaneous reaction in which oxygen is released back into the air, and the solution changes to clear.

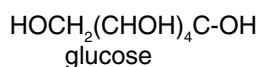
Now, that's a pretty general explanation, so obviously we need more detail. There are a fair number of chemicals involved, so I'll just focus on the important ones. When you first mix things together, a couple of things happen. The first is listed below.



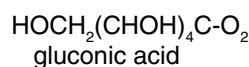
Using chemical formulas, this equation looks like



Don't get flustered by how complicated the molecules are. We're just going to look at certain parts of the equation to understand what happens to those parts. The hydroxide ions in this equation come from the potassium hydroxide you added to the solution. Also, note that e^- is just the symbol for an electron. Let's look at just the glucose and the gluconic acid molecules. I'll write them to emphasize what's happening at the far right side of each.

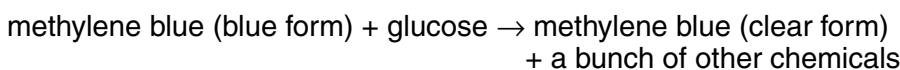


and

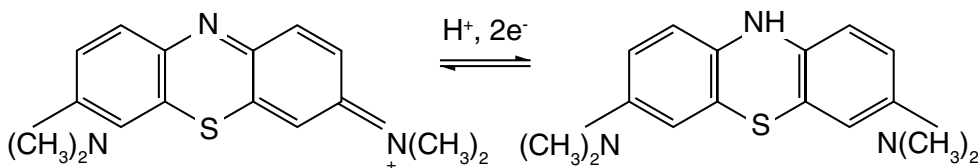


The only difference between these two molecules is that glucose has an OH on the end and gluconic acid has an O₂ on the end. We know that OH⁻ ions have an overall charge of -1, so, because the entire molecule is neutral, we can think of the rest of the molecule as having a +1 charge. For gluconic acid, we need to look at the O₂. Oxygen atoms need two electrons to fill their outer shells (refer to the Periodic Table), so when oxygen combines with other molecules, they share two electrons and those shared electrons spend more time near the oxygen.¹² Because there are two oxygen atoms (O₂) instead of one, there are four electrons (two per oxygen atom) spending more time near the O₂. *In a sense*, then, the O₂ at the end of gluconic acid has a -4 charge. Because this molecule is neutral, that means that *in a sense*, the rest of the molecule has a +4 charge. I emphasize the words *in a sense* because we're not talking about an ionic bond, in which electrons have jumped from one atom to another. We're just focusing on where the electrons spend more of their time and assigning plus and minus charges based on this observation.¹³ Anyway, what's happened is our glucose molecule, in becoming a gluconic acid molecule, has gone from a +1 "charge" to a +4 "charge." Because the molecule has become more positive, it's as if the molecule has *lost electrons* (losing negative electrons increases the positive charge). When a molecule loses electrons, we say it has undergone **oxidation**. So we would say that the glucose has been oxidized. In this case, the molecule actually added an oxygen molecule, and that's what historically led to this kind of reaction being called an oxidation, but oxidation can take place without oxygen being involved.

Now on to the next equation, and remember that we're just looking at what happens when you first mix the chemicals.



Now let's just focus on what happens to the methylene blue and represent it as the molecular structure rather than just formulas. Again, don't get disturbed by the complexity of the molecule. We'll only focus on small parts of it.



¹² Recall that this happens in water (H₂O) molecules. The shared electrons spend more time around the oxygen atom, making the molecule polar.

¹³ The "charges" I'm talking about here are referred to in chemistry circles as **oxidation states**. Just thought I'd mention it in case you come across the term somewhere.

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That two-way arrow simply means that this reaction is reversible—it can occur in either direction. In comparing the two molecules, you can see that the one on the left has an N at the top center, while the one on the right has an NH. Also, the plus sign on the lower right of the left molecule has disappeared on the right molecule. The symbols above the arrows indicate how this happened. We have added an H^+ ion and also added two electrons to the molecule on the left. These come from the glucose molecule in the word equation I wrote first. One of the added electrons makes the NH at the top neutral (remember, we added an H^+ ion), and the other gets rid of the plus sign at the lower right. Because we added two negatively charged electrons and only one positively charged H^+ ion, the molecule has gained one electron overall. When a molecule gains an electron, we refer to that as a **reduction**.



Topic: Redox Reactions

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The reaction I have just described is an example of a special class of chemical reactions known as **oxidation-reduction** reactions, or **redox** reactions for short. When an atom or molecule *loses* electrons in a reaction, we say the substance is oxidized. When an atom or molecule *gains* electrons in a reaction, we say the substance is reduced.¹⁴ In a redox reaction, at least one atom or molecule gains electrons, and at least one atom or molecule loses electrons.

What happens when a substance is reduced or oxidized can be confusing, but there's an easy way to remember it. Just think of a lion, and remember that **LEO** says **GER**. This stands for **L**oss of **E**lectrons is **O**xidation and **G**ain of **E**lectrons is **R**eduction. And no, I didn't come up with that. I'm not that clever.

As with our previous classifications of reactions, classifying reactions as oxidation-reduction is simply a tool to help our understanding. In that sense, there's nothing new at all in this chapter. Chemicals get together and do their thing. If we can place groups of reactions in separate classifications, however, we can better predict what might happen in other circumstances. That said, I should let you know that there are still many other classifications of reactions that chemists use. The three reactions in this chapter aren't the only ones, but they are major ones and illustrate the classification process. And if you think I've done something wrong by not including all the different types of reactions, I ask you to remember the purpose of this book, which is to help you develop a deep understanding of basic concepts. Don't think of this as a comprehensive resource for all your chemistry needs.

¹⁴ As I stated earlier, the reason for the term oxidation-reduction has to do with the fact that these kinds of reactions were at first thought to involve only the combining of oxygen with other molecules, as happens in this reaction. But this class of reactions is much more general and doesn't necessarily include oxygen.

Chapter summary

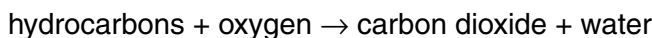
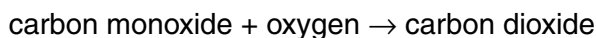
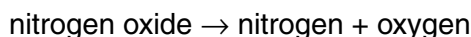
- A combustion reaction is one in which oxygen combines with a hydrocarbon to produce carbon dioxide and water.
- A catalyst is a substance that speeds up or improves the efficiency of a chemical reaction. The catalyst itself is unchanged as a result of the reaction.
- Activation energy is the energy input required to get a chemical reaction to proceed in a certain direction. A catalyst effectively lowers the activation energy for a reaction.
- An acid is any chemical that, when dissolved in water, increases the concentration of hydronium ions. Alternatively, an acid is anything that donates protons in a reaction.
- A base is any chemical that, when dissolved in water, increases the concentration of hydroxide ions. Alternatively, a base is anything that accepts protons in a reaction.
- pH is a logarithmic scale that measures the concentration of hydronium ions in a solution. The scale ranges from 0 (highly acidic) to 14 (highly basic).
- There is a difference between the strength of an acid or base and the concentration of an acid or base. The strength describes how completely the acid or base donates or accepts protons. The concentration tells how much of an acid or base a solution contains.
- Combining acids and bases always results in the production of a salt and water.
- When an atom or molecule loses electrons during a reaction, we say it is oxidized. When an atom or molecule gains electrons during a reaction, we say it is reduced.
- Oxidation-reduction (redox) reactions involve at least one substance being oxidized and at least one substance being reduced. Keeping track of the oxidation and reduction is a useful way to analyze what happens in a chemical reaction.

Applications

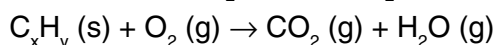
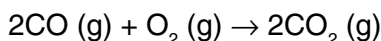
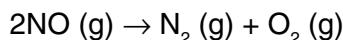
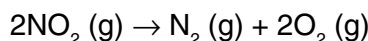
1. Most of us use a catalyst reaction almost every day. Cars have **catalytic converters** that help remove harmful substances before they escape out the exhaust pipe. The reason engines produce harmful substances is that the combustion, or burning, in the engine is never complete. After the engine does its thing, the following harmful substances remain: nitrogen oxide,

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nitric oxide, carbon monoxide, and unburned hydrocarbons.¹⁵ To get rid of these bad substances, we'd like the following reactions to occur:

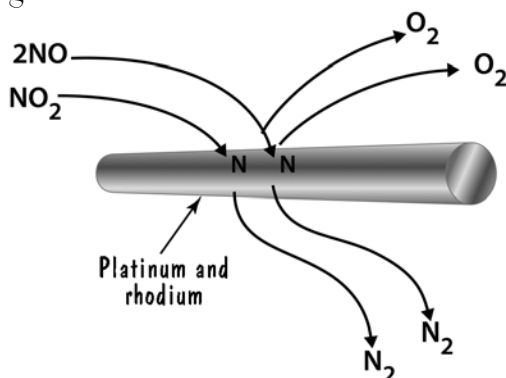


The balanced chemical equations for these reactions are below. Note that C_xH_y is the generic term for hydrocarbons that can contain any number of carbon and hydrogen atoms (hence the x and y subscripts rather than numbers). Because we don't have exact numbers for x and y , the last equation isn't balanced.



Okay. These are the reactions we want to happen, and they do happen somewhat as the byproducts of the engine leave the exhaust. The problem is that they don't happen all that quickly, leaving bad stuff heading out into the atmosphere. That's why we use catalysts that lower the activation energies for these reactions and help them proceed at a rapid rate. A catalytic converter is inserted in the exhaust system of a car so that it receives the bad stuff after it's left the engine. The catalytic converter contains a ceramic core coated with very expensive metals—platinum, rhodium, and palladium—and has two sections. The first section deals with nitrogen oxide and nitric oxide. This first part has platinum and rhodium covering the ceramic. These metals rip nitrogen atoms off the nitrogen oxide and nitric oxide. The leftover oxygen atoms form oxygen molecules, and the nitrogen atoms held by the metals combine to form harmless nitrogen. Figure 4.11 illustrates the process.

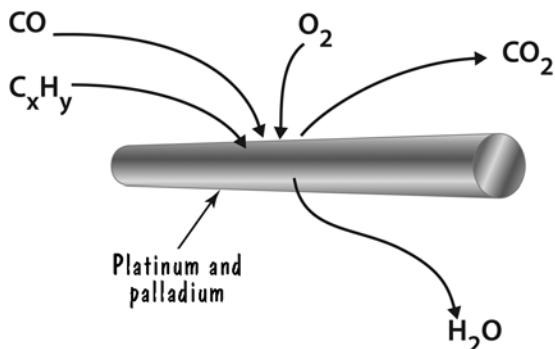
Figure 4.11



¹⁵ Hydrocarbons are any molecules that contain carbon and hydrogen.

The second part of the catalytic converter deals with the carbon monoxide and the unburned hydrocarbons. Here the ceramic is coated with platinum and palladium. These metals attract carbon monoxide, hydrocarbons, and oxygen. By trapping these molecules, the metals make it easier for them to get together in a chemical reaction. Figure 4.12 illustrates this.

Figure 4.12



As with all catalyst reactions, the catalysts (the expensive metals) help make it easier for a chemical reaction to take place while not changing themselves in the process. The metals get the other atoms together, or rip molecules apart, while staying on the ceramic. This is a good thing because you want a catalytic converter to last a while.

2. The reactions that take place in the two parts of a catalytic converter are referred to as the *reduction* reaction and the *oxidation* reaction. Can you figure out which is which? Aside from the fact that carbon atoms latch onto oxygen atoms, which indicates an oxidation, you can figure this out by looking at the “charge” (which we call an oxidation state) on the nitrogen atoms and the carbon atoms in each reaction. First, consider the NO molecule. There is a double bond between the atoms, and the electrons in those bonds spend more time around the oxygen than they do around the nitrogen. This is because oxygen has a greater affinity for electrons than nitrogen. This means that we can think of the nitrogen in NO as having a slightly positive charge or oxidation state. As with our earlier examples of oxidation and reduction, it’s not that the nitrogen is positively charged and the oxygen negatively charged, but in the molecule they sort of act that way. One product of the reaction is N₂. In this molecule, each nitrogen atom is bonded to another nitrogen atom, so the electrons in the bonds don’t prefer one atom over another. Thus, in N₂, the “charge” on each nitrogen is zero. So, nitrogen has effectively *gained* electrons in the reaction (going from a plus charge to zero charge). This means that nitrogen is *reduced*, and the reactions that take place in the first part of a catalytic converter are known as *reduction reactions*. Note, though, that in labeling these as reduction reactions, we’re only focusing on what happens to the nitrogen atoms. While the nitrogen is being reduced, the oxygen is being oxidized—you can’t have a reduction of one substance without a corresponding oxidation of another substance.

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Okay, what about the second part of the catalytic converter? Once again, we just have to look at one of the reactions, such as hydrocarbons adding to oxygen to produce carbon dioxide and water. Let's just look at the carbon atoms. In a hydrocarbon, the carbon atoms are bonded to one another or to hydrogen atoms. When carbon and hydrogen atoms get together, the different affinities for electrons dictate that the electrons will be more attracted to the carbon atoms, so the carbon atoms will have a negative "charge" in hydrocarbons. After the reaction, carbon atoms are bound to oxygen atoms (in CO_2). Again, affinities for electrons tell us what will happen. Electrons spend more time around the oxygen atoms, so the carbon atoms in CO_2 have a positive "charge." Carbon goes from a negative "charge" to a positive "charge."¹⁶ This process involves a loss of electrons, so carbon is *oxidized* in this process. Therefore, the reactions that take place in the second part of the catalytic converter are called *oxidation reactions*.

3. Molecular formulas for acids and bases often are written so it's obvious that they are acids or bases. For acids, the hydrogen ion (proton) that the molecule readily loses is placed at the beginning of the formula rather than lumped with the rest of the hydrogen atoms in the molecule. For example, acetic acid (vinegar) is $\text{HC}_2\text{H}_3\text{O}_2$, with that first H (or sometimes more than one H) being the one that jumps off easily. Lactic acid is $\text{HC}_3\text{H}_5\text{O}_3$, and citric acid is $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$. Two structures that easily accept protons, and thus make certain molecules bases, are NH and NH_2 . For easy identification, many bases are written with these structures tacked onto the end of the molecular formula. For example, we write the bases dimethylamine, methylamine, and aniline as $(\text{CH}_3)_2\text{NH}$, CH_3NH_2 , and $\text{C}_6\text{H}_5\text{NH}_2$.
4. Acids, bases, and acid-base reactions play a big part in the functioning of the human body. For one thing, your body is pretty sensitive to changes in pH. Anything too acidic or too basic will generally be a problem. Even though your stomach relies on acids to break down foods, too much acid in your stomach can hurt. To counteract this, you take a base called an antacid (duh!). Because bases snarf up hydronium ions, antacids make your stomach less acidic (higher pH). Of course, you have to be careful about what bases you put in your stomach. Ammonia is a pretty strong household base, but it's much too strong to put in your body. When my father had an upset stomach, he always swallowed baking soda and water—a base that is quite a bit stronger than your usual antacid. That usually solved his problem quickly but violently. I don't recommend it!

¹⁶ Throughout this application, I have put the word *charge* in quotation marks. To be absolutely correct, I should use the term *oxidation state*. I just want to say once again that these atoms and molecules carry a charge only *in a sense* because while the atoms and molecules are connected, the shared electrons spend more time around one than the other.

Chapter 4

5. Sometimes chemical equations are difficult to balance just by trying to balance the number of each kind of atom on both sides of the equation. If the equation is an oxidation-reduction reaction, there's a surefire method for balancing the equation. I won't give a detailed example, but I can tell you that the method involves looking at the oxidation states (that "charge" that atoms and molecules have, sort of, when combined) of the atoms and molecules and balancing the electrons lost with the electrons gained. It's a nifty trick that works every time, so if you find yourself with chemical equations that you just can't balance by inspection, you might want to look up the method and learn it.

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