THE JOY OF CHEMISTRY

The Amazing Science of Familiar Things

CATHY COBB &
MONTY L. FETTEROLF

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We dedicate this work, with all of our love, to Neda Jo Fetterolf, a model of strength, resilience, and courage.

See the blind beggar dance, the cripple sing,
The sot a hero, lunatic a king;
The starving chemist in his golden views
Supremely blest, the poet in his muse.

—Alexander Pope, *An Essay on Man*, 1734

The sight of the planet through a telescope is worth all the course on astronomy: the shock of the electric spark in the elbow outvalues all the theories; the taste of the nitrous oxide, the firing of an artificial volcano, are better than volumes of chemistry.

—Ralph Waldo Emerson, Essays: Second Series, 1844

APOLOGIA

Chemistry books usually fall into one of two categories—textbooks or children's books—which, in essence, excludes the general reader. There are some excellent works that address this deficit, including *Chemistry Connections*: *The Chemical Basis of Everyday Phenomena* by Kerry K. Karukstis and Gerald R. Van Hecke and *The Genie in the Bottle* by Joe Schwarcz.[1] Still, we have chosen to add our own effort to the mix, believing it to represent a slightly different approach. We have attempted to write a fireside chemistry: a chemistry book for the armchair scholar, to be perused in the comfort of one's own home and enlivened with straightforward demonstrations that can be carried out in the kitchen or garage. We have written this book because when we first started our chemistry studies, these were the types of books we loved. This book was written by chemists, but it is not written for chemists. It is written for the students we used to be and the scholars within us all. We hope the reader enjoys the following pages as much as we did the writing.

PREFACE

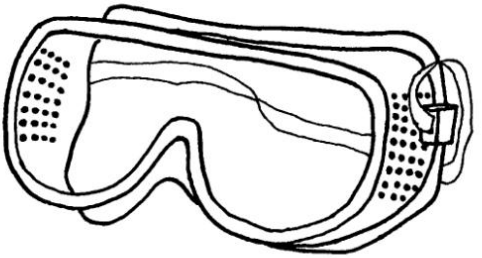
I am beginning chemistry, a most unusual study. I've never seen anything like it before. Molecules and Atoms are the material employed, but I'll be in a position to discuss them more definitely next month.
—Jean Webster, *Daddy-Long-Legs*, 1912

Given the present state of Western society, it may be difficult to believe that sex and cooking were once subjects to be approached with trepidation. But when Irma Rombauer wrote the *The Joy of Cooking* in 1931 and Alex Comfort wrote *The Joy of Sex* in 1972, they were addressing very real societal concerns. For Rombauer it was clarification of the kitchen, and for Comfort it was the demystification of the boudoir. Nowadays, chemistry can, for some, be perceived as a likewise worrisome affair. So perhaps it is time to discover the joy of chemistry.

Within these pages, we will explore the magic inherent in chemistry—from the fascination of fall foliage and fireworks, to the functioning of smoke detectors and computers, to the fundamentals of digestion and combustion—and we'll illustrate these precepts with hands-on

experiments using familiar materials. Laboratories and calculators are not required to enjoy the beauty of chemistry: the concepts can be explained in terms of everyday experience and confirmed with materials from your own closet.

Chemistry is often called the central science because it is central to our understanding of the physical and biological world and central to our common concerns, from medicine to politics to economics. Therefore, the principles of chemistry should resonate with our perceptions and experience. Chemistry can be a fascinatingly familiar science that appeals to our intuition, logic, and—if we're willing to get down and dirty—our sense of enjoyment, too. Our purpose here is to partake in the delights of chemistry and discover a little something in the process—about cars, the cosmos, and even crime.



A FEW NECESSARY WORDS ON SAFETY

While reading a textbook on chemistry, I came upon the statement "nitric acid acts upon copper" . . . I had seen a bottle marked "nitric acid" on a table . . . I was willing to sacrifice one of the few copper cents then in my possession. . . .

A greenish blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the performance became dark red. A great colored cloud arose . . . picking (the penny) *up . . . I learned another fact—nitric acid not only acts upon copper but it acts upon fingers. . . . I drew my fingers across my trousers and another fact was discovered. Nitric acid also acts upon trousers.*

—Ira Remsen, *My First Hundred Years*, ca. 1970

The precedence for books demonstrating home chemistry has been well established, but so is the precedence for a safety discussion at the beginning.[1] True, we are all reasonable, responsible adults here, but so was the hero of the above quote. Therefore, a few necessary words on safety could not hurt. *The Joy of Cooking* offers advice on food preserva-

tion and *The Joy of Sex* offers precautions against disease, so why shouldn't chemists have their warnings, too? Cooking, chemistry, and sex share comparable dangers for the unprotected and unprepared. Let's review the basics.

If you don't own a pair of safety glasses already, go now to the hardware store and buy a pair. Wear them while doing the demonstrations in this book.

Read the safety precautions embedded in the demonstrations and take them to heart.

Resist the urge to get creative. Random mixtures of household chemicals can create some pretty nasty brews. *Never mix ammonia with bleach!*

Work away from sources of sparks, which can include, interestingly enough, cell phones or other electric appliances.

Pour waste into the toilet, not the kitchen sink or bathroom sink.

Keep the amounts to a minimum. The difference between a bang and a pop is a matter of magnitude.

Know the location of the nearest working shower, fire extinguisher, and phone, but you should know these things anyway, as a normal functioning household presents enough hazards to warrant these precautions.

Keep everything away from children. This is an *adult* chemistry set. Keep things away from pets, too.

You will be directed to use kitchen utensils in many of the demonstrations. Don't assume that this means that the chemicals are nontoxic. Use plasticware and paper plates whenever possible (you will be told when only glassware will do) and throw the paper and plastic away after use. Even inexpensive paper plates will work for many of the demonstrations if they have a thin plastic coating. If you have to use glass or ceramic utensils, wash them thoroughly and separately before any contemplated reuse with food, or, better, maintain a separate set dedicated to chemical demonstrations. You will probably do most of the demonstrations in the basement, kitchen, or bathroom, but keep the chemicals well separated from toothbrushes, foodstuffs, or counters that will come in contact with food or toothbrushes. You may cover your counters with newspaper for added protection.

Microwave ovens vary, so treat the microwave times given in the

demonstrations as approximate. Get to know your own microwave oven and use mitts to remove hot items. Microwave ovens can also heat unevenly, so be aware that solutions heated in the microwave may be at the boiling point without showing bubbles. When in doubt, wait a minute after opening the door and use mitts. And, as always, be certain that you remove any metal, such as spoons, from items to be inserted in the microwave. This is chemistry, not cooking, but that does not change the physics of the microwave. Any resulting sparks may be exciting, but they are dangerous and can easily destroy a microwave.

The chemical procedures presented here have been carefully chosen to maximize your pleasure and minimize any risk, but please do not assume this is true of all chemical reactions or all chemical demonstrations. Again, experimentation is best left to the professional experimentalists and more complicated demonstrations to those trained in chemical demonstration.

Don't use any of the chemicals we discuss in a poultice or as a suppository or anything else that people were wont to do in the past or are doing in the present. Don't, under any circumstances, eat the chemicals in these demonstrations or drink them or bathe in them or splash them in your eyes. Don't expose them to flames or sparks (unless it's part of the demonstration). Protect your clothes or wear old clothes. Protect your hands. Wear rubber gloves when handling chemicals and wear yard-work gloves when handling steel wool. Treat all chemicals in this book as you would treat chlorine laundry bleach, gasoline, insecticides, and other things you know to be hazardous.

In other words, follow the same types of precautions offered in any children's chemistry set. In fact, you might go out and buy such a chemistry set. You can say it's for your children, or you can order it online. The chemistry kit will have the safety glasses you need and a few chemicals and a few experiments. It won't have the in-depth explanations, which is what we offer here, but it might be amusing and could save some trips to the hardware store.

SHOPPING LIST AND SOLUTIONS

Shopping List

In alphabetical order, here is a list of ingredients that will be used in the demonstrations. Most items can be purchased at a hardware store or the grocery store. If they cannot be purchased locally, they can be purchased online.

adhesive tape
alum (can be found with the spices in a grocery store)
aluminum foil
aluminum wire (can be purchased at a hardware store)
ammonia (household, clear)
analgesic stomach tablets, fizzing variety
aquarium pH-lowering solution that lists "sulfuric acid" and perhaps "water" as its only ingredients (at least three 37-milliliter bottles)
aquarium pH-testing kit (freshwater, should say "bromothymol blue" in the ingredients list)
aspirin

baking soda (*not* baking powder)
balloons
bar magnets (small)
basting syringe
black-light lightbulb and fixture
bleach (household)
candles
canola oil
chalk
citric acid (sold as "sour salt" in gourmet food shops or can be purchased online)
clear, carbonated soda (usually a lemon-lime soda will do)
clothes pins (plastic or wooden)
coat hanger (metal or plastic)
copper pot scrubber that is all copper mesh (can be found at hardware or grocery stores)
copper wire, at least two feet (60 centimeters), preferably uninsulated, but insulated can be used
corks, rubber or natural, that will fit into the end of a twenty-ounce (590-mL) plastic soda bottle
cornstarch
cottage cheese (purchase as needed)
cream of tartar (usually available on baking aisle or with spices in grocery store)
D-cell battery and battery holder (the holder is not absolutely necessary)
dropper with squeeze bulb (commonly called an eyedropper)
fingernail polish remover (the kind that contains acetone)
flexible gloves (the type used to protect the hands while washing dishes or when using household cleaners are okay, but they tend to fit poorly; surgical gloves can be purchased in a drugstore and are better)
food coloring (the liquid drops, not the squeezable gel)
galvanized nails (zinc-coated nails available in hardware stores)
garden gloves (thick cloth or leather)
gas-reducing digestive aid (the kind taken before meals to reduce intestinal gas)

glycerin (also called glycerol, from drug counter)
hydrogen peroxide
kitchen matches
lemon
liquid laundry bluing
lye (generally sold as drain cleaner, but do not get the kind that contains aluminum metal)
masking tape and permanent markers
metal spoons (two that can be discarded after use)
milk (whole milk, not skimmed or 2 percent, purchase as needed)
mineral oil
mozzarella (real and imitation, purchase as needed)
paper clips, metal
paper plates with coating
paper towels
more paper towels
plastic bags (the type that can be sealed by pressing the top together, in the sandwich size and in the gallon size)
plastic cups (clear and disposable)
plastic funnel
plastic soda bottles (clear, twenty fluid ounces [590 milliliter], at least ten)
plastic soda bottle (clear, 67.6 fluid ounces [2 liter])
plastic spoons
plastic wrap
purple cabbage (can be chopped up and frozen until needed)
rubbing alcohol
safety glasses (can be purchased from a scientific or educational supply store or hardware store)
salt petre, also called sodium nitrate (can be purchased from a cooking-supply store and some sophisticated grocery stores or online)
small plastic jars with caps (half-pint water bottles work well)
small, clear plastic containers to use as test tubes, or actual test tubes (which can be purchased at a scientific or educational supply store)
small juice glasses (three or four)

sponge (artificial or natural)
steel nails
steel wool (fine grade)
straws
super glue (the kind that contains cyanoacrylate)
swimming pool pH-testing kit (should say "phenol red" in the ingredients list)
table salt
table sugar
tall drinking glasses (three or four)
tea bag of brown tea
thermometer, range up to at least 220°F (105°C) (most meat thermometers will do)
tincture of iodine
vinegar
voltmeter (able to measure DC voltage in the one-to-three volt range)
whole milk
wooden kitchen matches

SOLUTIONS

1. Copper sulfate

Caution: Some fumes will be generated during this preparation, so work in an open area as you would if you were working with solvents, varnish, or paint thinner. The fumes are nitrogen-oxygen compounds such as those present in smog.

a) Put on the safety glasses. If you have a copper-mesh pot scrubber, use this in place of the copper wire. If not, strip ten inches (30 centimeters) of copper wire if it is insulated (use as is if it is not) and coil it so it lies flat in the bottom of a plastic cup that can hold at least two cups (500 milliliters) of liquid. If you have to cut it to strip off the insulation, that is all right. The wire can be in smaller pieces that add up to about ten inches (30 centimeters). If you are using a copper pot scrubber, put the pot scrubber in the bottom of the glass. The

cup can be clear plastic or clear glass, but the cup should not be reused for drinking in either case.

b) Add a third of a cup (80 milliliters) of the aquarium pH-lowering solution until the copper is well covered or emersed. This solution is acidic, so you may want to use gloves. Be careful, in any case, and do not get it on your hands or clothing. If you get it on your hands, rinse your hands thoroughly with water.

c) Add about four plastic teaspoons (about 8 grams) of salt petre (sodium nitrate) crystals. Swirl to mix. Loosely cover the cup with plastic wrap and set it aside. The solution should be ready in twelve hours or so. Although the reaction is slow, you can tell it is proceeding by the stream of small bubbles coming from the copper surface. If all has gone well, the solution will be an intense blue. There will be copper and sodium nitrate crystals left over in the bottom of the cup.

d) Using your funnel, pour off the blue liquid into a plastic bottle that can be capped. Take care not to pour any leftover crystals or copper into the bottle. Leave the leftover crystals and copper in the cup. You may want to use gloves because the solution is still acidic and you want to avoid getting it on your hands. If you do get it on your hands, rinse your hands with water.

e) Cap and label the bottle with the date, your name, and "copper sulfate in sulfuric acid." Add "Caution: Corrosive!" Store the solution at room temperature in a secure location, well away from children and pets.

f) Leave any undissolved copper and crystals in the bottom of the cup. When more copper solution is needed, you can simply replenish the ingredients in this cup. Cover the cup with plastic wrap, label it, and store it in a secure location, away from children and pets.

2. Iron acetate

a) Put on your safety glasses. Pour two cups of vinegar into a clear cup. (The cup can be plastic, but it doesn't have to be.)

Put on garden gloves (a must) and then tear off about a cubic inch (20 milliliters) of fine steel wool. This is about as much as you can tear off in one healthy pinch. Put the steel wool in the vinegar, and cover the cup with plastic wrap. Label the cup "iron acetate" and add your name and the date. The solution will be ready for use after sitting overnight. Save the solution with the steel wool still in it, as it will be used in a couple of demonstrations.

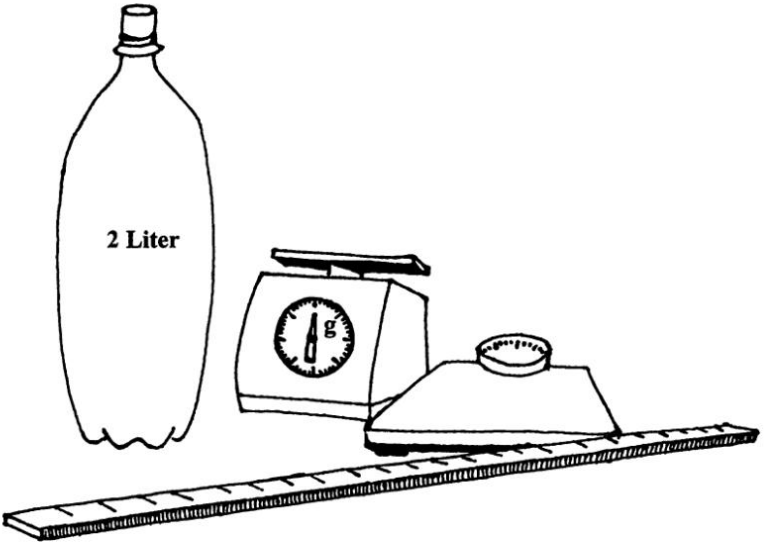
b) Store the solution at room temperature in a secure location, well away from children and pets.

3. Purple-cabbage indicator

a) Chop up about a half cup (120 milliliters) of purple cabbage, put it in a sandwich bag with a half cup (120 milliliters) of water. Heat in the microwave for approximately thirty seconds. The bag and contents should be very warm but not hot enough to scald or boil. After warming, the water should be shaded purple by the cabbage, and a bit of squeezing and manual crushing of the cabbage may make it darker.

b) The resulting solution will have a strong odor, which will get stronger on aging. To avoid overly unpleasant smells, the indicator can be made up only as needed, frozen if made up ahead of time, or can be made from dark purple plums if these are available.

But chemistry can be a smelly business!



A MOMENT FOR METRIC

Metric units (grams, centimeters, milliliters), as opposed to traditional US units (ounces, inches, cups), are preferred for scientific measurements. There is a valiant effort in the United States to move toward the use of metric units in everyday applications; however, most people in the United States still use traditional units for common purposes. Throughout this book, US units will be used, followed by their approximate metric equivalent in parentheses. Though no doubt metric units are now second nature to many, they are foreign to as many more, so a set of comparisons is offered here. Bear in mind that a centimeter is about the width of a small paper clip, a penny weighs about two grams, a meter is a little longer than a yard, and a milliliter is (very roughly) a quarter teaspoon. You may also want to keep in mind that an ounce is a measure of mass, but a fluid ounce is a measure of volume—which reminds us of the complexity of US units. So even if metric is not second nature now, don't worry. Anyone who has mastered the subtleties of US units shouldn't have much trouble with metric.

Some useful approximate US/metric conversions are given below. In the demonstrations, approximate metric equivalents will be provided with measurements in US units.

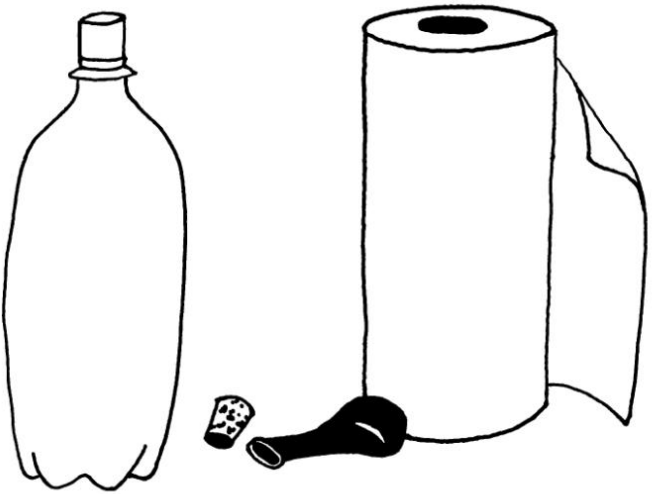
1 inch = 2.5 centimeters	1 pound = 0.5 kilogram
1 quart = 0.9 liter	1 ounce = 28.4 grams
1 tablespoon = 15 milliliters	1 teaspoon = 5 milliliters
1 cup = 236.7 milliliters	1 fluid ounce = 29.5 milliliters

For clarity's sake we will avoid abbreviations, but for the sake of completeness some common abbreviations are listed below.

cup (c)	gram (g)	liter (L)
meter (m)	milliliter (mL)	ounce (oz)
pound (lb)	tablespoon (tbsp)	teaspoon (tsp)

Scientists also normally use the Celsius temperature scale, but we will report temperatures in both degrees Celsius and the traditional US Fahrenheit temperature scale. For convenience, a comparison of some common temperatures is given below.

	room temperature	boiling water	frozen water	normal body temperature
Fahrenheit	68°	212°	32°	98.6°
Celsius	20°	100°	0°	37°



Introductory Demonstration: Bottle Rocket and Oobleck

Their faces were as rosy as if they had been exposed to the roaring flames of an oven; their voices resounded in loud accents; their words escaped like a champagne cork driven out by carbonic acid.

—Jules Verne, *Around the Moon*, ca. 1870

Let's begin with a bang. A bottle rocket.

Put on your safety glasses. Tear a paper towel into a strip about four inches wide and seven to eight inches long. Take a rounded plastic spoonful of baking soda (about 2 teaspoons or 10 milliliters) and spread it down the length of the paper-towel strip, in the center. Leave an inch or two at the end of the paper towel free of baking soda. Roll the paper towel lengthwise, as if rolling a cigarette, making sure that the baking soda doesn't get close to the ends of the towel. The roll should be small enough to go through the end of a plastic soda bottle. Once it is rolled, twist the ends so that the powder does not come out. When you are done, your paper towel roll should look like a cigar with twisted ends.

Remove the label from a twenty-fluid-ounce (590-milliliter) clear plastic soda bottle, rinse it out, and shake out any excess water. Find a cork that fits snuggly into the bottle opening and make certain it seals the bottle well without needing to jam it in. As extra insurance, pull a balloon over the cork (put the cork *inside* the balloon) so that you have a rubber-coated cork that will make a nice tight seal with the top of the soda bottle.

Pour a half cup (120 milliliters) of vinegar and a half cup (120 milliliters) of water into the bottle and swirl to mix. A two-liter soda bottle may be substituted if you use twice as much vinegar and water. Add enough purple-cabbage indicator so that the resulting solution is bright pink, but do not add more than one tablespoon.

Find a location, preferably outside, where it will not matter if a cork shoots upward with respectable force. Put the soda bottle with vinegar on the selected spot, drop in the twisted tube of baking soda, and rapidly and firmly cork the bottle. Stand the bottle upright so the cork will fire straight up in the air and out of the way of children, pets, or neighbors. Stand back. There should be a few second's delay while the vinegar solution penetrates the paper towel, then there should be a vigorous foaming and a satisfying bang as the cork shoots out. At this point the solution in the bottle will have changed from pink to blue or violet (depending on the amount of baking soda used). If the reaction appears to die down after a minute instead of popping the cork, you can swirl the bottle or you can push at the bottom of the cork to dislodge it as it was probably jammed too tightly into the opening. It may take a few attempts to get it right, but it's worth it.

The reaction that pops the cork is

$$\text{baking soda} + \text{vinegar} \rightarrow \text{carbon dioxide} + \text{water}$$

Carbon dioxide is a gas and tends to occupy a larger space than the solution. The expanding gas builds up a force sufficient to expel the cork from the mouth of the bottle.

Gas production is a good indication that a chemical reaction has taken place. Color change is also an indication of a chemical reaction; however, in this case the reaction could have proceeded just as well without the purple-cabbage indicator and without changing color. The purple-cab-

bage indicator is added to show that the reaction is an acid-base reaction, which we will discuss shortly.

The defining feature of a chemical reaction is that you start with one type of material or materials (a sour, pink liquid and a salty solid in this case) and end up with another type of material (a violet, watery liquid and a gas). There are other signs that can indicate a chemical reaction has occurred, such as a change in temperature, the production of light, or the formation of a solid that settles out of a solution—and in future demonstrations we will see all these and more. But to understand how the processes occur, we need a solid background in the basic structure of matter, and that is where we begin.

That is, after one more demonstration and this time with a splat—from oobleck.

Start with a cup (240 milliliters) of cornstarch in a small bowl and add a third cup (80 milliliters) water. Using a metal spoon (a plastic spoon might break), stir the water into the cornstarch. The mixture will quickly become thick and difficult to stir, but persevere. The exact amounts of water and cornstarch needed will depend on the brand of cornstarch, so you may have to add a little water or sprinkle in a little more cornstarch. You want to end up with something approximately the consistency of wet cement or corn-bread batter, but slightly dryer. Cutting through the mixture with the spoon helps the mixing process. When you are done, you have what science teachers have come to call *oobleck*, after the Dr. Seuss book *Bartholomew and the Oobleck*.[1]

Try smacking the spoon down on the oobleck, and then try picking up the oobleck in your hands. It feels fairly solid when you pick it up and squeeze it, but if you hold it in the palm of your hand, it appears to "melt" into a puddle and you can drip it between your fingers. Oobleck behaves in this manner—not quite solid, not quite liquid—because cornstarch traps water, somewhat like a sponge, but the structure of starch is less rigid so it can flow with its water load.

Try rolling up a pinch of the oobleck into a ball and throwing it on the floor. It should bounce a bit then hit the floor with a splat and puddle. Try slicing through the oobleck with a knife. It cuts somewhat like cheese, appearing more solid than liquid. Try pulling the oobleck apart quickly and then slowly. When the oobleck is separated quickly, it behaves more like a solid; when it is pulled apart slowly, it behaves more like a liquid.

Although the properties of oobleck seem quite different from the cornstarch and the water from which it was formed, the formation of oobleck is not a chemical reaction; it is a physical change. The cornstarch has trapped the water, but the water can be removed by the simple process of letting the oobleck dry out (though it may take several hours). The *physical* properties of the cornstarch and water have changed, but the *chemical* properties have not.

Changing a physical property is something like changing a person's appearance. You can cut your hair or change your clothes, but you are basically the same person as when you started. The change can be reversed by simply letting your hair grow out or by changing back into your former attire. A physical change for a material such as water might be to boil or freeze it or divide it into separate cups. Chemical changes, however, are a fundamental change of one type of material into another. A chemical change results in a new material with new physical and chemical properties. The bottle rocket represents a chemical change because the vinegar and baking soda turned into new materials: carbon dioxide and water. No amount of drying will recover the baking soda.

Because most chemical reactions also involve physical change, we will be discussing both, but our emphasis will be on the chemistry. As chemists, of course, we like a good splat, but nothing beats a bang!

INTRODUCTION

It Begins with a Bang . . . and a Splat

It's chemistry, brother, chemistry! There's no help for it, your reverence, you must make way for chemistry.

—Fyodor Dostoevsky, *Brothers Karamazov*, 1880

I must go to the laboratory and look into a little matter of acids and salts and alkalis. I've burned a hole as big as a plate in the front of my chemistry apron, with hydrochloric acid. If the theory worked, I ought to be able to neutralize that hole with good strong ammonia, oughtn't I?

—Jean Webster, *Daddy-Long-Legs*, 1912

We lead off with these quotes because they illustrate two observations. First, if one judges by the number of references to chemistry in Western literature of the late nineteenth and early twentieth centuries, it would appear that chemical literacy was once more common. From Arthur Conan Doyle to Emily Dickinson, many of these earlier authors displayed a friendly acquaintance with the principles of chemistry and assumed a familiarity on the part of their readers, too. Such allusions to chemistry, unfortu-

nately, seem to be less frequent in contemporary literature, presumably because of a lack of chemical fluency on the part of the authors and the readership. It is our belief that this situation might be alleviated by providing more in the way of public access to chemistry, which is the objective of this book.

The second observation illustrated by these two particular quotes is that some people, past and present, are naturally drawn to chemistry—and some are not. Why the difference? Is there a variation in brain cells? Is there an undiscovered chemistry gene? Maybe. But probably not. A better explanation may be found in the following abbreviated biographies of chemistry Nobel Prize winners and obvious aficionados of chemistry.

Linus Pauling (1901–1994), son of an Oregon pharmacist, shared a friend's chemistry set when he was fourteen years old.[2]

Mario Molina (1943–), who worried about atmospheric chemistry as his work carried him over the globe, had an aunt who helped him set up a laboratory in the bathroom of his boyhood home in Mexico City.[3]

Dorothy Hodgkin (1910–1994), a British chemist with an archaeologist father and a botanist mother, grew crystals with a soil-chemist test kit when she was ten years old.[4]

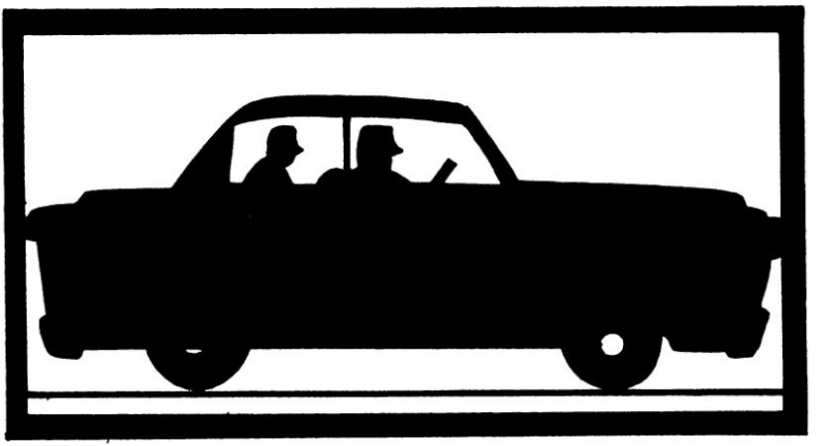
As a young Texan, Robert Curl (1933–) used up the chemicals in the chemistry set his parents bought him for Christmas, so he found a sympathetic pharmacist to help him resupply. His parents rethought the wisdom of the gift when, in using his new stock, he permanently altered the porcelain top on his mother's stove.[5]

Not much in the way of a shared gene pool, but perhaps some other traits in common. For one, these successful chemists (as well as many others) seem to have first encountered chemistry as an amusement. Does this mean that to appreciate chemistry as an adult you should choose parents who give you a chemistry set as a child? No. It is never too late to indulge yourself in the joy of chemistry. Each of our budding Nobel chemists also had a strong curiosity, and if you have picked up this book, then you do, too—and it is our hope to gratify this curiosity. We are proud to present herein a virtual adult chemistry set complete with demonstrations. The critical difference with our approach and that of most chemistry sets, however, is that the results are explained and the demonstrations are integrated into an account of the science of chemistry. A description of the expected outcome will be given with each of the demonstrations, as well as an explanation of the principles being illustrated, so the reader may, in

fact, choose to enjoy this book without performing all, or any, of the procedures given in the demonstrations. But the temptation to join in the fun—with a bang and a splat—may be difficult to resist.

The above short bios might also imply that the successful chemist must have access, as a child, to a mentor or companion with whom to discuss the chemical arts. Does this mean you have to maintain a discussion group to appreciate the principles of chemistry? Not necessarily. We have provided a sort of virtual companion in the "For Example" sections, which present topical essays germane to the chemical theory being discussed. We include these to help the reader make intuitive connections with the concepts of chemistry and to dispel the notion that chemistry is somehow removed from everyday experience.

Take, for instance, our mundane experience of driving cars. What does chemistry have to do with cars? As it turns out, quite a lot.



Introductory Example: Chemistry and Cars

Throughout this book, we will discuss chemistry in situations as diverse as cooking and solving crime. By way of introduction, our first "For Example" is going to take a quick walk through many of the topics that will be covered in these pages and illustrate them using an icon of modern technology: the automobile.

The *Joy of Chemistry* consists of two parts: Part 1 presents the fundamental principles of chemistry, and part 2 is a survey of some of the specialized fields of chemistry in which these principles are applied. Chap-

ters 1 and 2 of part 1 introduce atomic structure with the chemist's best friend: the periodic table. Atomic structure, as reflected by the periodic table, determines the makeup of chemical compounds—such as those that supply the structure and energy for the petroleum-based internal combustion engine. The type of reaction that is responsible for the power of this engine is the same type of reaction that supplies metal for its construction—oxidation-reduction reactions—which we'll describe in chapter 3. The connection of chemistry with automobile exhaust is made in chapter 4, which examines acid-base chemistry. The chemistry that clogs up the radiator is touched on in chapter 5, in the discussion of displacement reactions. The properties of the many materials that make up an automobile are a consequence of chemical bonding, discussed in chapter 6. The control and predictability of the various chemical reactions that take place in a car are governed by the principles of chemical reactions, the topic of chapter 7. The properties of automotive lubricants are explained in terms of intermolecular forces, which are the subject of chapter 8. The concentration of mixtures, the topic of chapter 9, concerns systems such as the choke, the radiator, and even the battery. The explosive reaction that changes gasoline to gas—and the gas properties that push the piston—are topics for chapters 10 and 11, which describe the gas phase of matter and reactions involving this phase.

Soaps and surfactants, useful for washing and waxing cars, come within the scope of chapter 12, which describes solid-state and surface chemistry. Chapter 13, on thermodynamics, delves into the operation of car-powered refrigerators. Metal fatigue is a topic for chapter 14, which investigates phase changes for pure materials and mixtures. And after reading chapter 15, a discussion of chemical equilibrium, the reader will be better positioned to understand why virtually all the gasoline turns to gas. The colligative properties of solutions, explored in chapter 16, can be applied to the salting of icy roads and the use of antifreeze in the radiator. The treatment of chemical kinetics in chapter 17 will relate to engine knock and catalytic converters. The discussions of photochemistry and electrochemistry in chapter 18 will shed light on automobile paint and chrome plating, the finishing touch.

In part 2 we will take a look at the specialized areas of chemistry and see that here, too, are intimate connections to daily living—and daily driving. The organic chemists of chapter 1 of part 2, who have turned the rich petroleum stew into plastics and polymers for interiors, gadgets,

hoses, and tires, are today facing a new challenge: the development of comparable materials from a new feedstock, either the biomass or recyclables—an exciting and interesting challenge for the creative chemical scientist. As we will see in chapter 2 of part 2, the inorganic chemists have more of the periodic table to choose from and have every bit as much challenge. Consideration of their work will lead us to a discussion of how fuel cells may help us in the future—and how radiator cleaners work right now.

The concerns of the biochemist, as outlined in chapter 3 of part 2, address the automobile's most dangerous component—the one located directly behind the wheel. But biochemists are also concerned with the sources and management of alternative materials harvested from the biosphere. In chapter 4 of part 2, we consider the efforts of analytical chemists and find that these intrepid souls draw from all of the above described disciplines. It is the job of the analytical chemist to certify the consistency and quality of many materials, including materials for cars, and in some instances, the consistency and quality of human life—as we will see when we consider the contributions of this field to crime-solving as well as cars. In chapter 5, we peer into our crystal ball to see the future of chemistry: a universe of new sources and substances. These new materials will no doubt lead to remarkable innovations—and, of course, new cars!

Ready? Let's open the hood and take a look. . . .

PART I

INTRODUCTION

Theory, Octaves, and Scales

Not unlike music and literature, chemistry is described in terms of its elements and has a theory based on fundamental principles. And, as with music and literature, there is much in chemistry that is art as well. Looking at nature is like looking in a mirror, and we interpret what we see in light of our experience and from our own perspective. The principles of chemistry represent an attempt to draw from nature a system and, as such, must remain a bit blurry about the edges and able to accommodate exceptions to the rule. Yet, over the years, there has evolved a substantial repertoire of physical models and methods that allows us to describe, understand, and predict the behavior of a considerable body of materials with a respectable range of properties. So we'll begin with the basics of this theory. For instance, we all know that chemists use scales, but did you know that chemists think in octaves, too? In part 1 of *Joy of Chemistry*, we'll learn why.

Demonstration 1: Water Witch

For days he made ceaseless calculations, but they were all wondrously unsatisfactory. . . . He finally concluded that the only way to prove himself was to go into the blaze . . . to watch his legs to discover their merits and faults. . . . To gain it, he must have blaze, blood, and danger, even as a chemist requires this, that, and the other.

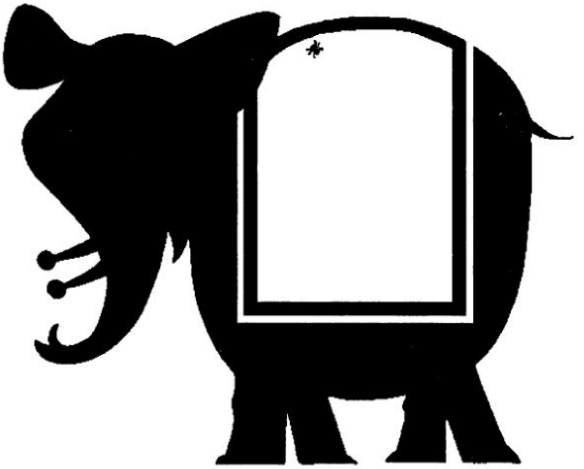
—Stephen Crane, *Red Badge of Courage*, ca. 1895

Take a plastic spoon and rub it in your hair or on a sweater until the spoon acquires a static charge, as evidenced by the attraction of the spoon for the hair or fibers on the sweater. Turn on a faucet so that there is a very thin stream of water. Hold the spoon close to the water and you will see the stream of water bend. What happened? Electron transfer.

Friction will cause electrons to transfer from one material to another if one of the materials has a stronger attraction for electrons. Most plastics have a stronger attraction for electrons than hair and clothing, so the direction of transfer was probably from your hair or sweater to the spoon, allowing the spoon to acquire a negative charge.[1] It could be a positive

charge if the transfer were in the other direction, but the effect on the stream of water would be the same. Water is made up of atoms of hydrogen and oxygen, which have negative electrons in clouds surrounding positively charged nuclei. For reasons to be explored shortly, the electron cloud around water has an uneven distribution of electrons. One end of the water molecule is more negative than the other. Opposite charges attract, and equivalent charges repel. When close to the negative electric field of the plastic spoon, the negative ends of the water molecules shift away from the spoon. The positive ends shift toward the spoon, which draws the water and the spoon together.

Find this idea of electrons bewitching? Read on.



CHAPTER 1

Electrons and Atoms, Elephants and Fleas

Our science is sensual. . . .

—Ralph Waldo Emerson, ca. 1830

In the above epigraph, Emerson was not intimating that science is sexy; he meant that we base our scientific theory on the input of our senses: touch, taste, hearing, sight, and smell. But when it comes to the theory of atoms, our senses fail. We can't see an individual atom, we can't taste one, we can't feel one, we can't hear one, and we can't smell just one. If someone hits us on the head with a single atom of tungsten (fairly hefty as far as atoms are concerned), we remain blissfully unaware.

The reason we are so insensitive to single atoms is that a single atom is extremely small. Some ten million individual atoms would have to line up to span the length of a grain of rice. The parts that compose atoms are smaller yet. The nucleus, or center of an atom, is made up of protons and neutrons, and the radius of a proton is on the order of a femtometer, which is a quadrillionth of a meter, or a millionth of a billionth of a meter. Mighty small. The protons and neutrons have a mass of about a septillionth of a gram (a trillionth of a trillionth), which is diminutive in anybody's book. The electrons are about two thousand times less massive

than the protons and neutrons. Electrons are to protons as fleas are to an elephant—the proton being the elephant. When we calculate the mass of an elephant, we don't add on the mass of the elephant's fleas, and when we calculate the mass of an atom, we don't add the mass of the electrons.

So the question naturally arises: If atomic particles are so small we can't use our senses to detect them, how do we know they are there? By inference. Humankind learned long ago that the input of the senses can be flawed. Optical, olfactory, and tactile illusions abound. So to discover the nature of those parts of the world that cannot be smelled, touched, and seen, people have learned to look at secondary effects and infer their causes. The concept was well captured by the venerated scientist Ernest Rutherford in the advice he gave to James Chadwick when Chadwick was looking for evidence for the neutron. Rutherford advised,

> How could you find the Invisible Man in Piccadilly Circus? . . . [B]y the reactions of those he pushed aside.[2]

Similarly, J. J. Thomson, making inferences from his own work and the work of others, declared the existence of electrons in 1897. By consensus, electrons were assigned a negative charge. Thomson was able to determine the amount of charge on a given mass of electrons by bending a beam of electrons in a magnetic field.[3]

This interaction of electrons with a magnetic field could cause a television picture to distort in the presence of a magnet. The electron beam in the cathode ray tube (CRT or television tube), which causes the phosphorescence on the screen, bends in a magnetic field. Of course, one should not bring a magnet up to a TV screen unless one has a dispensable TV because the interaction could do permanent damage to the electronics. But if one has a dispensable TV, it is an interesting effect to witness.

However, J. J. Thomson did not irrefutably establish the particulate nature of matter. It remained until 1909 for Jean Perrin to provide the definitive evidence for atoms, which he did by measuring the motion of microscopic pollen particles suspended in water. His detailed observations of this Brownian motion (named after the botanist Robert Brown) could be explained if it were assumed they were being buffeted about by moving atoms. His observations convinced the scientific community of the validity of the atomic model. Of course, they had been using the

model successfully before Perrin, but it was nice to have such elegant confirmation.[4]

In 1910 Ernest Rutherford realized that atoms must be composed of a central, dense nucleus surrounded by a lot of empty space. He fired some atomic-sized particles at an ultrathin gold foil and found that most of the particles passed through the foil, but a few bounced back. Always able to turn an interesting phrase, Rutherford commented, "It was . . . as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."[5]

Rutherford suggested that this nucleus at the center of the atom was composed of densely packed positively charged particles. Soon after, Henry Moseley, before his early death at Gallipoli in World War I, supplied experimental evidence for these particles, the protons.[6] The other particles in the nucleus, the neutrons, proved a bit harder to pin down because they have no charge. But James Chadwick, taking Rutherford's advice, finally confirmed their existence in 1932. Chadwick measured the rebound of certain radiation from nitrogen and helium and found it corresponded to a neutral particle with about the same mass as a proton.[7]

It took until the 1930s to discover all the pieces of the atom, which shows how difficult it was. Even so, laying out all the pieces still didn't solve the puzzle. There are other features of the atom—besides its size—that boggle the imagination. For one, there is the density of the nucleus. The density of a substance is the mass of a given volume. For instance, a bushel of feathers and a bushel of pebbles have the same volume but a definitely different mass. The bushel of pebbles is heavier; therefore, the bushel of pebbles is denser. We mentioned that Rutherford found that atoms have densely packed nuclei, but we didn't say how dense. It turns out that a nucleus, for all its diminutive proportions, packs some whopping million trillion grams per cubic centimeter. Compare that, say, to the density of lead, which is about eleven grams per cubic centimeter. The difference is that a nucleus consists of closely packed particles, but an atom is mostly empty space. A comparison can be made with Earth and the sun. If the nucleus of a lead atom were Earth, then the distance to the nearest electron would be about the distance to the sun. That is to say, most of the atom is just empty space. If you were to pack just nuclei into a container, you would be packing in solid marbles of matter. When you pack lead into a container, you are, in essence, packing in bubbles of empty space.

The question that may immediately spring to mind is this: If matter is composed of tiny bubbles, then why don't we collapse into puddles? An unsatisfying answer is that the electrons stay where they are because they are in orbit around the nucleus. But, one might object, if the electrons have a negative charge and the protons have a positive charge, shouldn't these opposite charges attract and shouldn't the electrons come crashing into the proton?

As it turns out, this is no trivial question.

In the early 1900s, the noted physicist Niels Bohr helped to answer this question by showing that an electron could remain at a distance from a proton as long as it kept moving. An analogy could again be made with Earth and the sun. Gravity does pull these two bodies together, but Earth, by moving in its orbit, can keep falling toward the sun but never fall into it. A similar picture can be painted for the electron. The electron can be thought of as circling the nucleus and thereby being pulled toward the nucleus, but not into it.

The reader has no doubt noticed that the above paragraphs contain quite a few qualifying phrases, such as "Niels Bohr . . . helped to explain" instead of "Niels Bohr explained" and "an electron can be thought of" rather than "an electron is." This approach is necessary because the analogies are just that—analogies—and they work only at a rather crude level. The analogies fail, and fail quickly, when any degree of precision is required. We can have no exact analogy because the physics at the atomic level is quite different from the physics we experience in our relatively gargantuan, everyday, macroscopic world.

So it is with this disclaimer in mind that we present the structure of the atom as it is currently understood. We first note the main features that are of common knowledge. The nucleus, or center of the atom, is composed of positively charged particles called protons and uncharged, neutral particles called neutrons. Electrons, as is customarily acknowledged, are the negatively charged particles that reside in "orbits" around the nucleus.

Here the term *orbit* is being used metaphorically, not literally. Though the common picture is to show electrons orbiting, satellite-like, around the nucleus, the space occupied by the electrons cannot be clearly delineated. The best we can do is to describe a sort of fuzzy region of proba-

bility in which we believe the electron may be found. To emphasize this difference, we call the space occupied by electrons around a nucleus an *orbital*, not an orbit. These orbitals can be spherical, a sort of dumbbell shape, or intricate structures of loops, doughnuts, and lobes. Once again, however, there is nothing in common experience that is really quite like them, so the analogies are not perfect.

The problem only becomes worse when you include more than one electron in the discussion, which is true for every element beyond hydrogen. Electrons are charged particles, and charged particles tend to be attracted to each other if they are of opposite charge and repelled from one another if they are of the same charge. A fairly good model for this behavior is found in the behavior of magnets. Opposite poles of magnets repel, while like poles attract. In an atom, the situation is much murkier because there are not just magnetlike interactions of one electron to one electron, or one electron to one proton, but a sea of negative and positive charges that are all interacting. The Nobel physicist Enrico Fermi once likened the situation to boats bobbing in a harbor. We know intuitively that the motion of one boat will influence the motion of all the others, and vice versa, but in ways so intertwined and convoluted that the final motion of just one boat cannot be predicted at any one time.

This problem is called the *three-body problem* by the people who study such things (the theoreticians of quantum mechanics). When you have two particles in motion that attract each other, you can describe the situation with an equation. But when you have three particles, and there are attractions and repulsions, and all these particles are in motion, there are too many things going on for one neat equation. The problem is one of clouds: a cloud exists and we can point to it and measure it, but to predict in advance just where it will be and what form it will take is not possible. There are too many factors, too many variables, many of which are unknown or unknowable. This problem is at the heart of the probability approach to atomic structure.

But luckily you don't have to know the position of every cloud to predict the weather.[8] Based on the probability approach, the theory of quantum mechanics is able to accurately interpret and predict many of the properties of atoms and molecules and how they interact. Scientists have also been able to understand and work with another interesting beastie: the *ion*. An ion is an atom or molecule that has lost some electrons or

gained some extra electrons, as our plastic spoon gained some electrons in the "Water Witch" demonstration. The fact that there are too few or too many electrons means that the positive charge from the protons is unbalanced and the ion has a net positive or negative charge. In the case of the plastic spoon, electrons were moved by friction, and the spoon acquired a net negative charge. Other charged species are able to perform other wonderful tricks. For instance, the miracle of photocopy machines.

For Example: Protons and Photocopies

Okay, some people might consider it an exaggeration to call a photocopy machine a miracle, but for anyone who has dealt with mimeograph machines, miracle is almost too weak. Photocopy machines are nearly as integral to the information age as computers and satellite cell phones. For all the photocopiers' modern-age convenience and programming—as well as features that allow one to copy, collate, stack, staple, and punch holes—the technology underlying them is actually fairly straightforward. The basic principle behind the operation of photocopiers is that static charges attract.

The parts of a photocopier are familiar to everyone who has ever used one to any extent because anyone who has ever used one to any extent has had a paper jam and had to open the copier to remove the jammed paper. If you have somehow avoided this experience, then simply go to a machine and open it to familiarize yourself with the interior workings. Integral to the photocopier is a rotating drum, a movable light source, a black powdery mess called toner, a heat source, and an elaborate system of rollers.

The first step in the photocopy process is for the drum to acquire a uniform static electric charge. The method used is slightly fancier than the

method by which the spoon acquired its static charge in the "Water Witch" demonstration, but just slightly. With this static charge, the drum can now attract toner, just as the spoon in the "Water Witch" attracted water. If this were the last step, however, then the paper produced would be a uniform black color, which is not the promised copy. To accomplish the copy, a very bright light passes under the paper to be copied.

Light is used because, as we will see, light and matter routinely interact: camera film reacts with light to form an image; digital sensors record light levels in a digital camera; photosensors detect the presence of a person in a light beam and trigger a door to open. When the light strikes a dark part of the paper, it is absorbed, but when the light strikes a white part of the paper, the light is reflected onto the drum. The interior material of the drum is photoconductive, which means light will cause the interior to eject an electron. Wherever an electron is ejected from the interior of the drum, it neutralizes the static charge at the surface of the drum in just that one spot. The rotation of the drum is synchronized with the movement of the light under the paper so the flat image can be transferred to the curved surface of the drum.

The light-exposed drum then turns past the toner, and toner is attracted to the still-charged portions of the drum. A piece of paper with a static charge is now passed over the surface of the drum and attracts the toner away from the drum. The paper is heated to fix the toner onto the paper, and the new copy is presented.

If the principle behind the operation of photocopiers—the attraction of materials to a static charge—is so basic and well understood, why did it take so long to produce a practical photocopier? The answer, as with many such innovations, is that the fundamental concept existed long before the materials necessary to implement the idea. Materials science is a discipline all to its own because of the virtually infinite variety of properties that the elements, as well as the substances derived from the elements, can display. Consider, for instance, the variation in the behavior of aluminum and copper (both of which are used in electrical wiring) and steel nails and galvanized nails (both of which are hit on the head). In the next chapter, we lay the differences on the table—the periodic table, that is.

Demonstration 2: Coppers and Robbers

> *You're capable of anything, you and Osmond. I don't mean Osmond by himself, and I don't mean you by yourself. But together . . . like some chemical combination.*
>
> —Henry James, *Portrait of a Lady*, ca. 1880

Be careful with this demonstration. The solution you will be making is caustic and can damage clothing, surfaces, eyes, and skin. If any is spilled, rinse the area immediately with plenty of water. If any should get in your eyes, rinse your eyes with copious amounts of water and seek medical attention. For this demonstration, it is imperative to have proper eye protection.

Put on your safety glasses and protective gloves. For the first part of this demonstration, take two clear plastic cups and add tap water to a depth of one inch (2.5 centimeters). Cut one-inch (2.5-centimeter) strands from the aluminum wire and the copper wire. Remove the insulation on the wires, if there is any. Place the aluminum wire in one cup and the copper wire in the other. Using a plastic spoon, add a half teaspoon (2 milliliters)

of the lye specified in "Shopping List and Solutions"—the crystal drain cleaner without added aluminum metal. Swirl the cups gently to mix. There will be some small bubbles as the lye dissolves.

After about fifteen or twenty seconds, small streams of tiny bubbles should be seen coming from the aluminum wire but not from the copper wire.

Copper and aluminum wire seem similar. Both are used in electrical wiring. But the aluminum visibly dissolves in this caustic solution, and the copper wire does not. The bubbles you observed rising from the dissolving aluminum are hydrogen, which explains why thin strips of aluminum foil are added to some lye-based drain cleaners: to provide agitation to help break up clogs.

Keep your safety glasses on for the second part of this demonstration. Take the copper sulfate solution prepared as outlined in "Shopping List and Solutions" and pour it into a plastic cup to a quarter-inch (0.5-centimeter) or less depth. Place a steel nail and a galvanized nail in the cup, taking care that they do not touch. Both the steel nail and the galvanized nail acquire a copper coating in the copper solution, but the steel nail has a more rapid reaction. Moreover, its coating of copper will look more like copper than the coating on the galvanized nail. Steel nails are mostly iron, and galvanized nails are steel nails that have a corrosion-resisting zinc coating. Both the nails perform the physical function of fastening boards, but their chemical makeup, as determined by the positions of iron and zinc on the periodic table, accounts for the difference in chemical behaviors seen here.

CHAPTER 2

Periodically Speaking

That army could not recover anywhere. Since the battle of Borodino and the pillage of Moscow it had borne within itself, as it were, the chemical elements of dissolution.

—Leo Tolstoy, *War and Peace*, ca. 1864

When taking a picture at a family reunion, a photographer may ask members of immediate families to stand together and may arrange them by height or generation. A fastidious photographer might also position the residents of the West Coast on one side and the East Coast on the other and have the Northerners stand a bit to the back and dwellers of the South in front. Such arrangements would take some time and effort but would also ensure that future generations could discern a good deal of information about any one particular subject by finding his or her position in the photo. The periodic table is similar to such a family-reunion photo: once the principle of the arrangement is understood, information about a particular element can be obtained by simply knowing its position in the table.

The clever photographer who arranged the subjects for the periodic table was Dmitri Mendeleev. Born in Serbia in 1834, he would have

understood the need for organization at a family reunion, as he was the youngest of fourteen children. He also would have understood the benefit of organization by region because his mother, seeing his talent, used her limited resources to relocate him to Moscow for his education. After securing a teaching position, following the manner of teachers of all time, Mendeleev began searching for a way to organize his material. His subject was the chemical behavior of the elements, and the pedagogical device he came up with grew into the periodic table.[1]

To organize the elements, Mendeleev wrote down all their observed behaviors on cards and arranged the cards into groups with similar behavior. Then he added a second level of organization and began arranging the cards according to their characteristic mass.

The terms *mass* and *weight* are often used interchangeably, but there is a subtle yet important difference. The weight of an object is a measure of how strongly gravity pulls on the object. Because the gravity of the moon is less than the gravity of Earth, objects weigh less on the moon. Mass, however, is measured on a balance, relative to some standard. The statue of "blind justice," used to symbolize the rule of law, is a blindfolded woman holding such a balance. To measure the mass of an object, the object is placed on one side of the balance, and standards of known mass are placed on the other side until the sides are level. Because gravity affects the standards the same way it affects the object being measured, the measured mass would be the same on Earth as on the moon.

By the time of Mendeleev it was known that elements combined with each other in set ratios, the way ingredients are combined in a recipe. So just as a recipe for cake might call for one egg and one cup milk, a recipe for table salt might call for one part sodium and one part chlorine. The mass of "one part sodium" was taken as the characteristic mass of sodium, and it was this characteristic mass for each element that Mendeleev used for his second layer of organization.

Other investigators had arranged the elements in groups according to their characteristic mass, but Mendeleev took the bold step of keeping groups together that had similar chemical behavior, even when it meant leaving gaps in the table. The genius of Mendeleev's insight was quickly revealed when others found the missing elements and fit them into the blanks in his table.

Several decades later, when other investigators deciphered the structure of the atom, they found that trends in atomic structure followed the arrangement in the periodic table, revealing the connection between atomic structure and chemical behavior. The concept may seem obvious in hindsight: the behavior of everything from fish to flashlights is determined by how they are put together. But the underlying structure of the atoms—electrons, protons, neutrons—was unknown when Mendeleev constructed his table.

The modern periodic table, considerably expanded since Mendeleev's time, is shown in figure 1.2.1. Many of the symbols may be familiar, such as H for hydrogen and O for oxygen, but others may not. Some elements such as tungsten (W) and sodium (Na) derive their symbols from alternate names for these elements. Tungsten's *W* is from the German *Wolfram.* Sodium's *Na* is from the Latin *natrium.* The numbers above the symbols are the atomic numbers (a concept to be elaborated upon shortly). Listings of the elements by name, by symbol, and by atomic number are given in the appendix.

1																	2
H																	He
3	4											5	6	7	8	9	10
Li	Be											B	C	N	O	F	Ne
11	12											13	14	15	16	17	18
Na	Mg											Al	Si	P	S	Cl	Ar
19	20	21									22	23	24	25	26	27	28
K	Ca	Sc									Ti	V	Cr	Mn	Fe	Co	Ni
37	38	39									40	41	42	43	44	45	46
Rb	Sr	Y									Zr	Nb	Mo	Tc	Ru	Rh	Pd
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72
Cs	Ba	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Hf
87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104
Fr	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Rf

29	30	31	32	33	34	35	36										
Cu	Zn	Ga	Ge	As	Se	Br	Kr										
47	48	49	50	51	52	53	54										
Ag	Cd	In	Sn	Sb	Te	I	Xe										
73	74	75	76	77	78	79	80	81	82	83	84	85	86				
Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn				
105	106	107	108	109													
Db	Sg	Bh	Hs	Mt													

Figure 1.2.1. The modern periodic table.

A more conventional, condensed representation of the periodic table, without atomic numbers, is shown in figure 1.2.2.

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac†	Rf	Db	Sg	Bh	Hs	Mt	▪	▪	▪						

*Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
†Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 1.2.2. The conventional periodic table.

The representation of the periodic table in figure 1.2.2 has all the same elements as the table in figure 1.2.1, but it requires the user to remember that the elements from cerium (Ce) to lutetium (Lu) and thorium (Th) to lawrencium (Lr) actually follow lanthanum (La) and actinium (Ac), respectively. In the conventional periodic table, these elements are shown as detached rows for typesetting convenience.

But whichever form one chooses to consult, the striking feature of the periodic table is its orderliness. In a world of shape-shifting clouds and randomly branching trees and no-two-look-alike cells or snowflakes, the order of the periodic table, rigid as an orthodox liturgy, may seem a bit unnatural. However, this primary order, dissolving through complexity into disorder (a discussion reserved for our chapter on thermodynamics), forms the fundamental basis for our varied world.

The first and most definitive pattern is in the number of protons in the nucleus of each element: hydrogen, the first element listed, has one proton in its nucleus. Helium, the second element (reading from left to right), has two protons in the nucleus. After helium, dropping down to the second row and starting again from left to right, lithium has three protons in the nucleus, and the next, beryllium, has four. Proceeding straight across, jumping the gap,

boron is next and has five protons in its nucleus. The pattern continues: carbon has six, nitrogen has seven, oxygen has eight, and on it goes. The number of protons in the nucleus is called the *atomic number* and is the number given along with the element symbol in figure 1.2.1. Complexity starts to creep in, however, as soon as the mass of the elements is considered. The atomic numbers count up the elements in nice, neat whole numbers. But a look at the appendix tells us that the atomic *masses* for the elements are not nice, neat, whole numbers. The irregular nature of the atomic masses is a result of the fact that atoms are built from protons and neutrons. There is a basic theme that describes how they are built, but there are also variations on the theme.

You have to start somewhere, and when it came to mass, scientists started with the proton. They assigned the proton a mass of one *atomic mass unit*, or *amu*. Because they are so very close in mass, for many purposes the neutron is assumed to have a mass of one amu, too. The number of protons in an atom of an element is always the same and equal to its atomic number. If an atom has seventeen protons, then its atomic number is 17, and it is chlorine. If an atom is chlorine, it has seventeen protons. But the mass of a given atom of an element can vary because the mass of an atom is determined by the number of protons *and* the number of neutrons, and the number of neutrons can vary. For instance, each chlorine atom has seventeen protons, but some chlorine atoms have eighteen neutrons, and some have twenty neutrons. The total mass for the atoms of chlorine with eighteen neutrons will be seventeen amu plus eighteen amu, or thirty-five amu. The total mass for the atoms with twenty neutrons will be seventeen amu plus twenty amu, or thirty-seven amu. When there are atoms of an element that have different masses, these atoms are called *isotopes*. The atomic mass for a given element is the average mass of its isotopes.

The number of protons in an atom is sort of like the characteristic that distinguishes a human as male. If a person has this characteristic, then he is male. If an atom has a certain number of protons, then that decides what element it is. But different males can have different masses. We could easily find the average mass of a sample of males by adding all their individual masses and dividing by the number in the sample. The same procedure is followed for elements on the periodic table. As can be seen in the appendix, the mass for chlorine is given as 35.45 amu, which is not the mass of either isotope. It is an average over the natural abundance of isotopes. However, it is the number of protons that determines the element, so even though isotopes may

have different numbers of neutrons and different masses, it is the atomic number that tells the number of protons, and the number of protons determines the element. The number that increases as you go across the periodic table is the atomic number, and this is the same as the number of protons.

But if the number of protons in the nucleus were the only basis of classification, then why would the table have its curious shape? The answer is that the order of the table goes deeper than the number of protons.

In biology, organisms are classified by kingdom, phylum, class, order, family, genus, and species. In chemistry, elements are also classified, though with fewer layers. What concerns us here is the classification of elements into what are termed different *species*. The number of protons determines the element, so atomic number could be thought of as the "genus." If an atom has two protons, then it is helium. If an atom is helium, then it has two protons. That doesn't change. The number of electrons in an atom of an element, however, can vary, and each variation can be thought of as a "species." In the neutral species of an atom, the number of protons is the same as the number of electrons.

Protons carry one unit of positive charge, so for an atom to be neutral, that positive charge must be offset by a negatively charged electron. Therefore, in addition to giving the number of protons in a particular element, the periodic table also tells us the number of electrons in an uncharged atom of the element. For instance, hydrogen has one proton, so a neutral atom also of hydrogen has one electron. Carbon has six protons in its nucleus, so it has six electrons in an electrically neutral atom.

The electrons occupy orbitals, as we previously outlined, and now we are going to add that these orbitals are arranged around the nucleus in *shells*. The shells are nested, as sketched in figure 1.2.3, and each one is able to accommodate more electrons. The number of electrons each shell can hold is determined by a mathematical progression: the first shell holds two electrons, the second shell holds eight electrons, the third shell holds eighteen electrons, and the fourth shell holds thirty-two electrons. A shell's capacity for holding electrons increases in a regular pattern: the square of the shell number times two. The first shell, shell 1, holds one squared times two, or two electrons. The second shell, shell 2, holds two squared times two, or eight electrons, and so on.

Here we have represented the shells as neat layers in nested spheres, sort of like a jawbreaker with many layers of flavor. But in truth, these

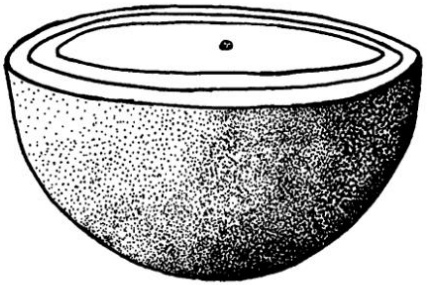


Figure 1.2.3. A pictorial representation of shells around an atom. Not to scale!

shells are made of subshells, and if we were to represent the subshells in our model, we would end up with something that looked more like a sliced-open onion. In an onion, the layers overlap and become interwoven, and that is true for the subshells of atoms, too. But this is a layer of the onion that we don't presently need. Theoretically, once filled with electrons, the shells are spherical, so we will consider them spherical here.[2]

The rows of the periodic table are called the *periods* (hence periodic table). The position that an element occupies in a period is determined by the number of electrons in its outermost shell. Each time the atomic number increases by one, the number of protons in the nucleus increases by one, and the new element requires one more electron for a neutral atom. The electrons arrange themselves in shells around the nucleus. Each time a shell is filled, a row is finished. A filled shell is represented on the periodic table as a filled row, as shown in figure 1.2.4.

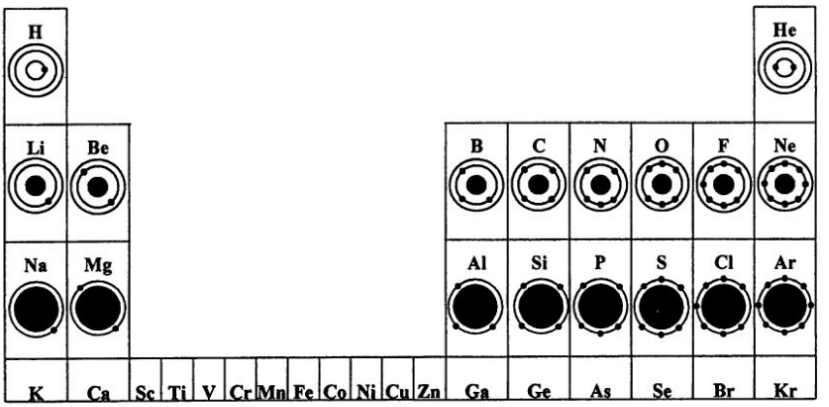


Figure 1.2.4. Each time the atomic number increases by one, the number of protons in the nucleus increases by one, and the new element requires one more electron for a neutral atom. The electrons arrange themselves in shells around the nucleus. Each time a shell is filled, a row is finished. Helium, neon, and argon each finish out a row, and each has a filled shell.

The shells, therefore, determine the shape of the periodic table, but there is more to the story than that. The existence of shells has another important effect: the occupancy or partial occupancy of a shell dictates the chemical behavior, or the *chemical reactivity*, of the various elements. The vertical columns of the periodic table are groups of elements that share a similar condition of occupancy of their outermost shell and are therefore members of the same family, but families that share a common chemistry rather than a common ancestry.

For instance, helium, neon, and argon each have filled shells, and this vertical column is commonly called the noble gases because of their dignified, inert, unreactive nature. Their lack of chemical reactivity is typified by helium, which is so unreactive that we use helium balloons for decorations at parties. Hydrogen, on the other hand, is very reactive. Hydrogen balloons at a party might result in entertainment more explosive than we would care for.

The chemical reactions of nitrogen and phosphorus are similar because they share the same number of electrons in their outer shell (five). The reactivity of oxygen resembles the reactivity of sulfur because of their shared outer-shell occupancy (six). This outer-shell occupancy of an atom is called its *valence*. Carbon has a valence of four (with four electrons in its outer shell), and its chemistry shares some similarities with silicon, which also has a valence of four. Silicon, germanium, tin, and lead, which have the same valence, have all been used in various proportions to form semiconductors, interesting and important materials that we will investigate later when we discuss chemical bonding.

Oxygen reacts with hydrogen in a ratio of one to two, as does sulfur, selenium, and tellurium. It is this chemical reactivity, the ability of elements to join chemically into compounds, that is responsible for transformation of the elements of the cosmos into the materials of Earth and ultimately the materials of life. Although we have to put off our detailed discussion of the complex molecules of life until we have a firmer foundation, we already have enough information to discuss two of the most important chemicals of life and decidedly life's necessary precursors: salt and water. The first life formed in the salty oceans, as evidenced by the salty solutions of our cells, and is always present in the salty character of our bodies.

Normally in the English-speaking world when we think of salt, we

think of table salt, sodium chloride, NaCl. If there is a second-most familiar chemical formula, second to H_2O, that is, it must be sodium chloride, the formula for table salt. Sodium chloride and potassium chloride share so many properties that potassium chloride is used as a substitute for table salt. They have similar formulas, too, in that sodium chloride is formed from one part sodium and one part chlorine, and potassium chloride is formed from one part potassium and one part chlorine. But this ratio of one to one is not true for all salts. For instance, calcium chloride has a ratio of two parts chlorine to one part calcium. Why does sodium behave one way and calcium another? The periodic table speaks to that issue, too.

To begin with, the periodic table can be broken down into two major sections, the metals and the nonmetals. The metals are to the left of the zigzag staircase line that starts on the right between B (boron) and Al (aluminum) and ends between Po (polonium) and At (astatine). (With the exception of hydrogen, which is considered a nonmetal.) The metals Cu (copper), Ni (nickel), and Zn (zinc) are familiar and do a good job representing what we expect from metals: they are hard at room temperature, they can be bent and molded, and they conduct heat and electricity. The nonmetals are a bit harder to pin down with a set of generic characteristics, but the properties of oxygen and carbon are familiar enough to establish that these elements are *not* metallic in their behavior.

The division between metal and nonmetals is rather fuzzy, so the elements in the immediate vicinity of the zigzag staircase line are called *metalloids*, which means they don't fit either definition exactly. Aluminum (Al) finds itself in this position, and though we generally tend to think of aluminum as a metal, it is an odd metal because it is very light (that is why it is chosen for engine blocks in race cars) and because it can appear grainy and gray rather than sleek and shiny. Aluminum foil is malleable and shiny but changes drastically in the charcoal fire of a barbecue, as do aluminum cans. So aluminum acts like a metal and like a nonmetal. It is a metalloid.

A salt is a compound formed from a metal and a nonmetal; thus, sodium, a metal, reacts with chlorine, a nonmetal, to form sodium chloride, the salt we know as table salt. The unvarying recipe for sodium chloride is one sodium atom with one chlorine atom, and the basis for this strict ratio is again shells, as we will now explain.

The mass of an atom is determined by the number of protons and neu-

trons in its nucleus, as we mentioned earlier. The electrons, you will recall, add virtually nothing to the mass, as fleas on an elephant add virtually nothing to the mass of the elephant. But just as fleas, despite their size, influence the behavior of the elephant, the electrons, despite their small mass, determine the chemistry of the elements. When an electron is gained or lost, the atom experiences an imbalance of charge and becomes an *ion*, which can be considered a new species of that element. An atom of an element with more electrons than protons is a negatively charged species, a negative ion. An atom of an element with fewer electrons than protons is a positively charged species, a positive ion. An ion is a reactive species, seeking, as it were, an entity of the opposite charge with which to combine.

It might seem at first that any amount of charge could be gained or lost and any ion could be formed, but it turns out that there are certain predictable amounts of charge that each element prefers to handle—sort of a comfort level for ionization, if you will. This comfort level can be predicted from the periodic table.

The first shell (the first row) is filled with two electrons; the second shell needs eight; the third shell holds eighteen; and so on. We said that the noble gases were unreactive, and it turns out that they also have filled shells. The first and most fundamental principle of the *shell theory of atoms*—the theory that chemical behavior depends on the number of electrons in shells—is that atoms like to have fully filled shells. A demanding motivation for chemical reactions is for atoms to fill or empty their electron load, by whatever means available, until they reach this happy state. Thus, chlorine (Cl), with seven outer electrons, will tend to acquire an electron so that it can have eight electrons and fill its shell. Sodium (Na) will shed its outer electron so that it can have just a filled shell of eight. In fact, the number *eight* is so often the choice for stable electron arrangements, we refer to all these stable, filled-outer-shell arrangements as *octets*. Hence musicians have their octaves, and chemists have their octets. But for chemists the concept is a bit fuzzier. For hydrogen, for instance, a filled outer shell consists of two electrons, so its "octet" is two. Actually a duet.

Shedding an electron gives sodium a net charge of positive one, which is shown as a superscript + on the symbol for the element, Na^+. Of course, sodium could also add seven electrons to fill its octet, but atoms like to do

what is easiest. Losing one electron is easier than taking on seven more. Adding an electron gives chlorine a net charge of negative one, which is shown as a superscript – on the symbol for the element, Cl^-. The equal and opposite charges of the sodium ion and the chloride ion mean that the two will seek each other out and combine in a one-to-one ratio, NaCl.

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac†	Rf	Db	Sg	Bh	Hs	Mt	•	•	•						

*Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
†Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 1.2.5. If sodium, Na, loses an electron, it will have a filled shell. If chlorine gains an electron, it will have a filled shell. With one fewer electron, sodium becomes Na^+. With one extra electron, chlorine becomes Cl^-. Na^+ and Cl^- can come together to form NaCl.

It works the same way for potassium chloride. Potassium (with the symbol K for the Latin *kalium*) is right below sodium on the periodic table. Potassium can attain a blissful state of eight electrons in its outer shell by losing one electron, which leaves one extra, unchecked proton and a positive charge. This one positive charge matches with the negative charge of the chloride ion in its happiest state, and the result is KCl, a compound able to substitute for table salt. Consulting the periodic table, we see that calcium realizes a filled-shell configuration when it loses two electrons and acquires a plus-two charge, Ca^{2+}, so, predictably, it will join with two chloride ions, 2 Cl^-, to form the salt $CaCl_2$.

$$Ca^{2+} + 2\ Cl^- \rightarrow CaCl_2$$

So there you have it in a nutshell. Or an atomic shell, if you wish.

The idea can be extended to materials formed from nonmetals, too. For instance, the formula H_2O can be explained in terms of atomic shells. There are some subtleties regarding the difference between the bonding in sodium chloride and the bonding in water that we will explore later when we delve more into chemical bonding, but the rearrangement of electrons to fill shells is still our basic premise.

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac†	Rf	Db	Sg	Bh	Hs	Mt	▪	▪	▪						

*Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
†Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 1.2.6. The number of valence (outer-shell) electrons for hydrogen and oxygen can be determined from their position in the periodic table. Hydrogen has an atomic number of 1, so it has one proton and one valence electron. Oxygen has a total of eight electrons, but two of these are in a filled shell. Only the outer six electrons are found in the valence shell.

The number of valence electrons (outer-shell electrons) brought into the mix by hydrogen is one, as evidenced by its position as the first element on the periodic table. Hydrogen craves a filled shell, such as that enjoyed by helium, which means it needs another electron. Similarly, the number of valence electrons brought in by oxygen is six, as evidenced by its position as the sixth element in period two, reading from left to right across the periodic table. However, oxygen is not satisfied, either. True happiness (stability) for oxygen would be to have a filled shell, as in the case of neon. To fill its shell, however, oxygen needs two more electrons. Hydrogen and oxygen fulfill their mutual needs by sharing electrons, but in this particular mélange, two hydrogen atoms are required to fill the needs of one oxygen

atom. When the two hydrogen atoms share their two lone electrons with oxygen, oxygen ends up with eight valence electrons. In the exchange, each hydrogen is able to gain the use of one electron from oxygen and fill its own shell . . . and destiny. The resulting arrangement is water, H_2O.

There can be no question that we need water for life, and there is also no question that we need air. But not every component of air. Air is mostly nitrogen gas, which exists as the *diatomic* (two-atom) molecule N_2. Mixed in with nitrogen is a substantial amount of oxygen gas, which exists as the diatomic molecule O_2. The natural existence of nitrogen as a diatomic can also be explained by the desire to have filled shells. An atom of nitrogen has five electrons in its outermost shell. In diatomic nitrogen, each shares three electrons with its fellow nitrogen—which gives both the use of eight electrons—to fill their shells.

The reason the atmosphere is much more enriched in nitrogen than oxygen is chemical. Nitrogen is much less chemically active; that is, it participates in fewer types of chemical reactions than oxygen does. In addition, nitrogen gas is very stable in the atmosphere, exposed to solar radiation, whereas oxygen is not. But to say that nitrogen gas is less reactive than oxygen gas is not to say the element of nitrogen is unreactive. It is, in fact, very reactive, as evidenced by the number of materials and mixtures that have nitrogen as a major constituent.

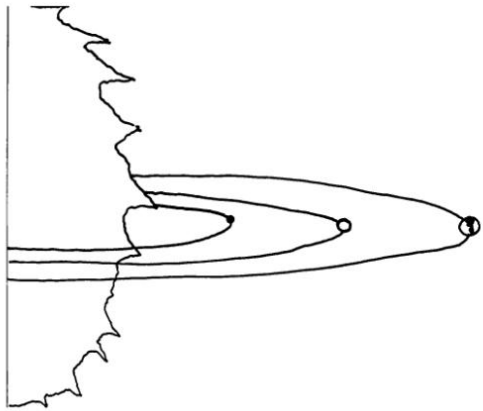
For instance, nitrogen is a principle component of gunpowder, yet nitrogen is also an essential component of amino acids and necessary for life as we know it. Nitrogen compounds give smog its red-brown color, nitric acid is found in acid rain, and liquid nitrogen is used to remove warts. The same nitrites that preserve food cause cancer in mice. Nitrogen compounds are responsible for nitrogen-rich, fertile soil and also for the growth of diaper-rash bacteria. Some compounds of nitrogen are powerful cleaning agents, and others are the source of the fishy odor of degrading fish.

The laughing gas that enables painless tooth extraction is an oxide of nitrogen, but TNT stands for trinitrotoluene, a less constructive compound. Physiologically active nitrogen compounds such as caffeine, nicotine, morphine, codeine, and quinine have shaped human history. The chemistry that threatens life with nitrogen-based fertilizer bombs is shared with the chemistry that protects life through the cushion of nitrogen air bags. So destructive and so nurturing at the same time. How can this be?

To understand the virtuosity of nitrogen, we need to take another look

at the periodic table. From nitrogen's position we can see that there are two ways it can achieve a filled shell: it can gain three electrons to move forward to a filled shell, as noted above, or it can also lose five electrons to go backwards to a filled shell. As it turns out, there are some intermediate situations, such as losing only three electrons, acceptable to nitrogen, too; therefore, nitrogen has several personalities and is able to undergo many metamorphoses. We will continue to learn about the versatility of nitrogen—and the rest of the periodic table—in the pages to come.

For Example: Elements of Diversity



Now that we've been introduced to the periodic table, let's take a short guided tour. Many of the elements on the periodic table, such as iron (Fe, from the Latin name for iron, *ferrum*), nitrogen (N), phosphorus (P), oxygen (O), and carbon (C), are familiar. Others, such as technetium (Tc) and lanthanum (La), generally are not. We are familiar with hydrogen and oxygen because these elements are evident in the air we breathe and the dirt we walk on, but no matter how we pant or how we dig, we do not expect to ever inhale one atom of lawrencium (Lr) or unearth one gram of seaborgium (Sg). The unfamiliarity of these elements is not just based on difficulty in pronunciation and spelling (though that could be argued), but on cosmology and physics. The reason has to do with something called the Big Bang.

According to the Big Bang theory, our universe was born with an unimaginably huge explosion (hence the name Big Bang). The debris of the Big Bang first condensed into the lighter elements, hydrogen and helium, because they are relatively simple: hydrogen is composed of one proton and one electron, and helium is composed of two protons, two neutrons, and two electrons. Only when enough of these atoms had gathered together and the mass started to compress under its own gravity did these nuclei begin

to fuse into the nuclei of heavier atoms. The heaviest nucleus that could have been formed in the first stars by the *fusion* process, the fusing together of nuclei, is iron. To contain the elements it has, our sun must have begun its fusion process with some heavier nuclei already present, which means it is probably a third-generation star—the result of the collected dust from the explosions of two previous stars. But even so, a look at the relative abundance of the elements in the sun, as shown in figure 1.2.7, reveals that the lightest elements, hydrogen and helium, are still the clear winners.

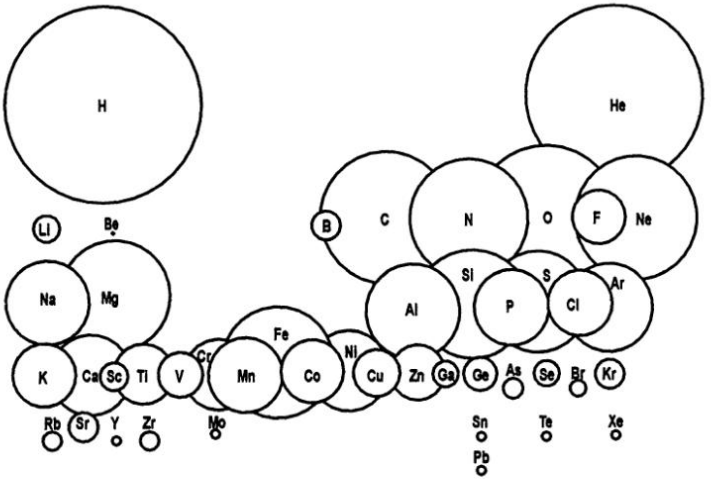


Figure 1.2.7. The abundance of elements in the sun. Adapted with permission from Steven I. Dutch, "Periodic Table of Elemental Abundance," *Journal of Chemical Education* 76 (1999): 356–58.

In figure 1.2.7, the abundances are displayed with a *logarithmic scale*: the changes in the radius of the circle for the element represent changes of orders of magnitude, or powers of ten, like a decibel scale. A sound registering 20 on the decibel scale is ten times as loud as a sound registering 10; a sound registering 30 is one hundred times greater than a sound registering 10. On this periodic table, it is the radius of the circle around the element that represents the multiple of 10.

We use a logarithmic scale for the abundance of the elements because there is so much difference between them. An attempt to compare abundances without this device would be a bit like trying to picture Godzilla and a mouse in the same movie frame (which may be the reason there is no *Godzilla versus Mighty Mouse* movie). A skyscraper can be shown

along with Godzilla, but not a mouse. We will not concern ourselves with exact numbers here; we just need a feel for the relative abundances of the elements, and these are nicely represented by the circles.

After hydrogen and helium, there is a jump to carbon because intermediate nuclei are too unstable to form in the fusion process. Helium has two protons in its nucleus and carbon has six, so a carbon nucleus can be made from the fusion of three helium nuclei, a process that requires the extreme heat and density of massive stars. After that there is a marked preference for nuclei with an even number of protons in the nucleus, oxygen has eight protons, neon has ten protons, magnesium has twelve protons, and so on. This preference can be explained by the theories of nuclear physics—but that is a topic for another time. We note that the elements essential to life, hydrogen, carbon, nitrogen, oxygen, sodium, magnesium, phosphorus, calcium, iron, and so forth, are all there in the sun, too.

Sometimes the flotsam of the Big Bang and exploding stars can condense into a planet rather than another star, which is probably what happened in the case of our planet, Earth. From the abundance of elements on our Earth's crust shown in figure 1.2.8, it can be seen that the material of Earth is essentially that of the sun, depleted in the more volatile elements such as helium and hydrogen.

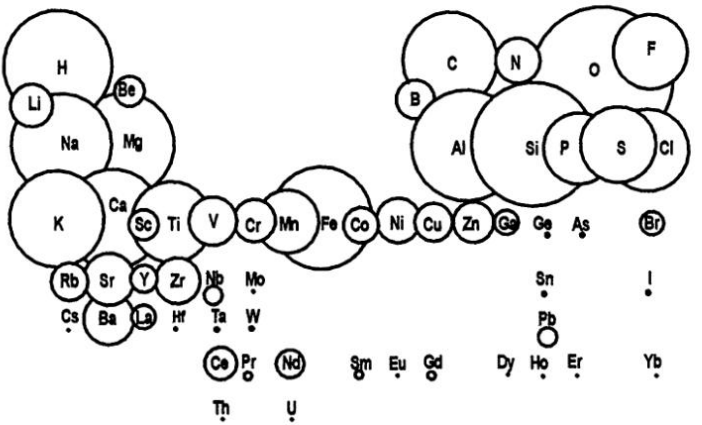


Figure 1.2.8. The abundance of elements in Earth's crust. Adapted with permission from Steven I. Dutch, "Periodic Table of Elemental Abundance," *Journal of Chemical Education* 76 (1999): 356–58.

The relative wealth of aluminum and silicon compared to iron and nickel shows a shift in favor of the lighter elements because we are

looking at Earth's crust: in the upheavals that have occurred over the eons, the heavier elements have sifted toward the interior and the lighter elements have risen to the top.

In figure 1.2.9, we have compared the top forty elements in Earth's crust with the top ten elements in the human body.[3] Reassuringly for most people, though perhaps discouraging for theorists of extraterrestrial seeding, it can be seen that the abundance of elements in the human body is remarkably similar to the abundance of elements on Earth, which rather nails down our evolutionary origin. In order of decreasing amounts, the human body is composed of oxygen, carbon, hydrogen, nitrogen, calcium, phosphorus, potassium, sulfur, sodium, and chlorine. Other elements are present, but in quantities of less than 0.1 percent.

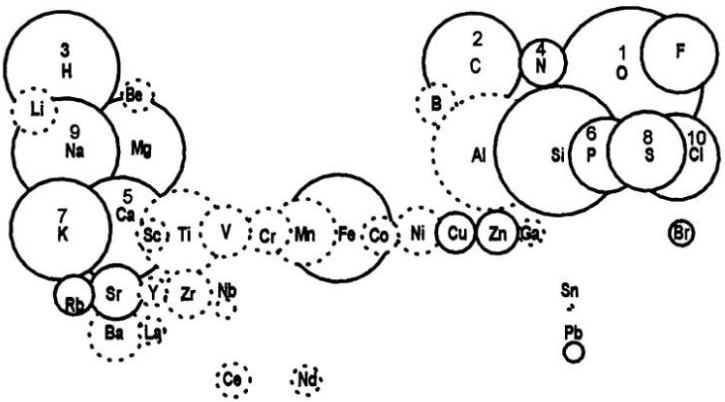
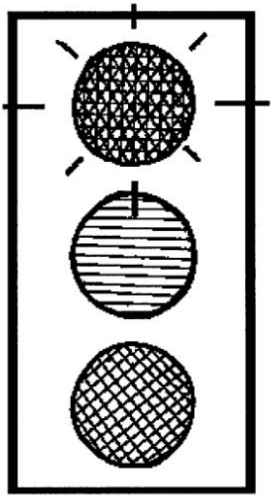


Figure 1.2.9. Comparison of the top forty elements in Earth's crust with the top ten elements in the body. The elements found in the body are numbered to show their abundance. Oxygen is the most abundant (1), carbon is second (2), etc.

There are, however, some interesting exceptions. Silicon, which is present in every grain of sand, tread on daily by every creature, and gripped firmly by every root, is absent from the body's top ten. Aluminum, also conspicuous in Earth's crust, not only is missing from the body's top ten but also is being investigated for its links with neurological damage.[4] Iodine doesn't make the top forty for abundance in Earth's crust, but if it is absent from the body, people suffer goiters. From this comparison we see that it is not just physical presence that decides the importance of an element but its chemical reactivity, too—so into that realm we proceed.



Demonstration 3: Stop-and-Go Chemistry

Our earthly fire also consumes more or less rapidly according as the object which it attacks is more or less combustible so that human ingenuity has even succeeded in inventing chemical preparations to check or frustrate its action.

—James Joyce, *Portrait of the Artist as a Young Man*, ca. 1915

Don the safety glasses. Ladle two tablespoons (30 milliliters) of iron acetate solution (the vinegar solution with dissolved steel wool described in the "Shopping List and Solutions") onto a plastic-coated paper plate. Add a teaspoon (5 milliliters) of household ammonia. The resulting solution should turn from a light orange-brown to a dark green. Now drop in a teaspoon (5 milliliters) of hydrogen peroxide. The resulting solution is a deep blood-clot red.

What happened? Iron, like nitrogen, can tolerate different numbers of electrons associated with its nucleus. The *oxidation state* of an element in a given environment—be it a free element, an ion, or in a compound—is an expression of the element's electron load in that environment. In the

current demonstration, iron is changing between two different oxidation states: the *ferric ion* (Fe^{3+}, or iron with a positive three charge, which means it is missing three electrons) and the *ferrous ion* (Fe^{2+}, or iron with a positive two charge, which means it is missing two electrons). The agent causing the change in the iron is the peroxide.

ferrous iron + hydrogen peroxide → ferric iron + water + oxygen

Iron is changed from the ferrous ion to the ferric ion because the peroxide attracts the electrons away from the iron. Hydrogen peroxide uses the extra electrons to change into water and oxygen gas, which bubbles out of the solution during the reaction. The compound formed between ferrous iron and ammonia is green. When the peroxide changes the ferrous iron to the ferric iron, the ammonia, still in solution and unchanged, forms a red compound with ferric iron. Reactions such as these, where elements gain or lose electrons, are called *reduction* (if they gain electrons) or *oxidation* (if they lose electrons). Because the reactions must happen in tandem, the combined result is called a *redox reaction*, a class of reactions of current and historical importance.

CHAPTER 3

Reason, Reactions, and Redox

They first broke the ore into little pieces, and cleansed them with the hand. . . . Then coal and ore were arranged in heaps and in successive layers. . . . In this way . . . the coal would be transformed into carbonic acid, then into oxide of carbon, its use being to reduce the oxide of iron, that is to say, to rid it of the oxygen.

—Jules Verne, *Mysterious Island*, ca. 1870

Doubtless a vigorous error vigorously pursued has kept the embryos of truth a-breathing: the quest of gold being at the same time a questioning of substances, the body of chemistry is prepared for its soul, and Lavoisier is born.

—George Eliot, *Middlemarch*, ca. 1871

Like a fickle lover, an electron will flirt with, or even flit to, any nucleus it finds more attractive. In the last demonstration, "Stop-and-Go Chemistry," we saw electrons leave iron quite cavalierly in favor of hydrogen peroxide. True, the electrons were content to remain with iron as long as there were no better prospects in the vicinity, but in the presence of hydrogen peroxide, away they went. In other words, the degree of association, lover with beloved, electron with nucleus, depends on the environment. The ferrous ion is iron missing two electrons; the ferric ion is iron missing three electrons. The green ferrous ammonia compound became a red ferric ammonia compound because hydrogen peroxide acts like an electron sponge.

When an element loses electrons, it is said to be *oxidized*. In the case of our stoplight reaction, the moniker makes sense: oxygen is acting on the iron, taking its electrons. However, for other reactions, *oxidation* is a misnomer: oxygen does not have to be involved for oxidation to take place. For instance, with safety glasses in place, take a metal paper clip (not plastic coated), unwind it part way, and put the unwound end in your copper sulfate solution. In five minutes or so, the end of the paper clip in solution will have acquired a shiny copper coating. In this particular case, copper ions in solution are gaining electrons lost from the metal of the paper clip and turning into copper metal, an exchange of electrons that occurs without the intervention of oxygen.

When an element gains an electron, it is said to be *reduced*. The first quotation at the beginning of this chapter, the quote from Jules Verne, describes just such a process: the gleaning of iron from iron ore. Metal smelting was one of the first organized and controlled chemical reactions (after, perhaps, cooking, pottery, painting, and fermentation). Metal smelting is the process for which the gaining of electrons, or *reduction*, is named. The process is called reduction because the metal smelted from the ore weighs less than the beginning ore; that is, it is reduced in weight.

Most metals exist in nature combined with one or more nonmetals, such as oxygen, sulfur, or chlorine. They are compounds in ores, not pure metals. In the smelting process, the metal is separated from the nonmetal by heating the ore with a material that has a stronger attraction for the nonmetal than the metal. In the case of the process described by Jules Verne, the nonmetal is oxygen and the second material with a stronger attraction

for oxygen is carbon, provided in the form of charcoal. The oxides of carbon, such as carbon dioxide, are mostly gases that will blow away, leaving the pure metal behind. Along with the understanding of atomic structure came the realization that the metal had gained electrons in the process of being reduced from ore to metal, so *reduction* became the name for gaining electrons. Charcoal's experience is called *oxidation* because the carbon combines with the nonmetal to carry it off, and many times this is oxygen. At the atomic level the carbon loses electrons, so losing electrons came to be known as oxidation. The defining characteristic of a redox reaction is that the element being reduced gains electrons and the element being oxidized loses electrons, a convention that has been immortalized in the mnemonic "LEO the lion says GER," which stands for "Lose Electrons Oxidation" and "Gain Electrons Reduction." Whenever there is oxidation there must be reduction; that is, the reaction pair is always found together: you can't have something losing electrons without the electrons going someplace else.

The copper sulfate and iron acetate solutions made from the recipes in the "Shopping List and Solutions" are examples of redox reactions. In the copper sulfate solution, the copper metal loses two electrons to become copper ion, Cu^{2+}, so it is oxidized. The fumes are from the nitrogen/ oxygen compounds in salt petre being reduced to volatile nitrogen/oxygen compounds. In the iron acetate solution, the iron in the steel wool is oxidized to Fe^{2+} (recall that the symbol for iron is Fe from *ferrum*). The electrons given up by iron are taken up by the hydrogen ion in vinegar to form the hydrogen gas bubbles that you probably observed in the iron acetate solution.

The fact that most metals exist in nature in a combined form (combined with nonmetals) can be satisfactorily demonstrated by a tour of a beach or a rocky field. One finds sand and rocks in all shades and textures, but unless one is very lucky, no chunks of pure metal. Not that it is impossible to find pure metal. Even nuggets of gold, as evidenced by the California gold rush, are possible to find. But such occurrences are rare, which helps give gold its value. The desire to transmute more common metals into precious gold was one of the goals of alchemy—an early, unsystematic practice of chemistry that flourished in Alexandrian Egypt and later in Europe during the Middle Ages. Based on colorful transformations such as the ones witnessed in our "Stop-and-Go" demonstration,

alchemists believed that if they could find the right recipe, they would be able to transform other metals into gold. George Eliot alluded to this quest in the quotation from *Middlemarch* at the beginning of this chapter. As she implies, no alchemist ever found a way to make gold. But the effort led to many fundamental revelations, such as the discovery of oxygen as a pure element. This discovery was made by several scientists at about the same time, including Antoine Lavoisier, the French chemist referred to by George Eliot as the "soul" of chemistry. But while Lavoisier may have been born to be the soul of chemistry, he died on the guillotine during the French Revolution. There may be joy in chemistry, but the same cannot always be said for politics.

As our quick trot from prehistoric chemistry to the French Revolution might indicate, paired reduction-oxidation reactions, or redox reactions, have played a continuing role in the history of humankind. In our one-sentence survey of prehistoric chemistry, we did not include an extremely important redox reaction—the making of fire—but we include it now because combustion, it turns out, is one of the most famous, and infamous, redox reactions.

For Example: Firefighters and the Chemistry of Combustion

Combustion is a fascinating reaction: from the gentle light of a candle, to the comfort of the hearth, to the welcoming warmth of a campfire. But these are cases where combustion is controlled. In the cases of uncontrolled combustion, the fascination turns to fear, primal and pure. Thankfully, there are those who routinely face this fear to protect others from the danger. A brief consideration of combustion chemistry hints at the magnitude of their challenge.

The primary combustion reaction is a straightforward redox reaction. A compound made primarily of hydrogen and carbon, a *hydrocarbon*, is usually the fuel. When hydrocarbons mix with oxygen, they can react to produce carbon dioxide (CO_2) and water and heat.

$$\text{Hydrocarbon} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} + \text{heat}$$

In the reaction, the carbon is oxidized (literally gaining oxygen this time) and the oxygen gas gains enough electrons to attract hydrogen and form water, so it is reduced.

Gasoline is a hydrocarbon. The reactions in the cylinders of cars are combustion reactions in which hydrocarbons and oxygen are combined, as in the above equation, so car exhaust, the reaction product, consists primarily of carbon dioxide and water. Sugar and starch are hydrocarbons, oxygen is the gas we breathe in, so heat-producing metabolism can be thought of as a combustion reaction. That metabolism is a form of combustion can be demonstrated by blowing out a candle: the rush of oxygen-poor and carbon dioxide–rich air both cools the flame and deprives it of oxygen. Wood, paper, and dry grass are mainly hydrocarbons; they make excellent fuels and dangerous fires. The consumption of oxygen by fire and the production of carbon dioxide make suffocation as much of a hazard to the firefighter as thermal injury.

That oxygen is necessary for combustion can be demonstrated by covering a candle with a nonplastic clear glass. After all the oxygen in the glass is used up, the candle goes out. Showing that carbon dioxide is a product of combustion is a bit more complicated and requires safety glasses and gloves, but it can be done. Take two similar nonplastic (real-glass) glasses and to one add four teaspoons (20 milliliters) of water. Add two drops of the fish-tank indicator suggested for purchase in the "Shopping List and Solutions." Placing both glasses on a sheet of white paper for viewing, you should see that the one with water has a faint green-blue color.

Light the candle and get a good flame going. Invert the empty glass over the candle. Allow the flame to burn out. Take the glass off the candle and immediately pour the water-indicator solution from the other glass into the glass that was over the candle. Quickly cover the mouth of the receiving glass with your palm and swirl the solution a bit so it can mix with the gases from the candle flame. Now put this glass back on the

white paper and observe the color. The CO_2 in the flame's gases should have changed the indicator from green-blue to yellow.

How do we know that it was CO_2 that caused the change? Take a clean glass and prepare another solution with four teaspoons water (20 milliliters) and two drops of fish-tank indictor. Now blow up a balloon. Holding the balloon pinched shut, put the balloon on one end of a straw and put the other end of the straw down into the indicator solution. When you ease open the neck of the balloon, freshly exhaled CO_2 from your lungs will bubble quickly through the solution and you should see the indicator once more turn from green-blue to yellow.

That water is a product of combustion can be demonstrated with an ice cube and a nonplastic glass or a nonplastic saucer. Invert the glass over the candle flame for a minute with the ice cube balanced on top of the glass, on the outside. Moisture from the flame will condense on the inside of the glass. Alternatively, place the ice cube on the saucer and then hold the saucer over the flame without touching the flame. Moisture will condense on the bottom.

Moisture can sometimes be observed on a metal spoon held over a flame, but water will not be the only material observed on the spoon. There will also be soot. Combustion is usually not the clean, simple reaction outlined above. Combustion is usually a complicated conglomerate of many reactions in which fragments of hydrocarbons, or soot, are produced. Some of the most interesting behaviors of combustion reactions are a result of incomplete combustion and more complicated reactions.

For instance, fire can spread, almost like a living thing, in pulsating waves. Part of the propagation is due to falling or flying debris, but because flames are burning gases, these gases and fire can travel anywhere the wind goes. To show that the flame contains burning gases, try the following experiment. Find an area that is free of stray breezes and clear the area of any flammable material (including loose hair and clothing). Put on your safety glasses. Light a candle with a wooden match and then hold the match aside, still burning. Blow out the candle and quickly bring the burning match close to the candlewick, but without touching the wick, as shown in figure 1.3.1. The flame will jump from the match to the candle as it races back through the combustion gases. You may have to try it a couple of times to get it right, but it should work.

Jumping through the air and spreading by falling debris aren't the

only ways fires can propagate. Differences in air pressure are associated with fires because they consume oxygen and produce heat. The consumption of oxygen creates a partial vacuum at the source of the flame, which is filled by air coming in from the sides. The heated air above the fire causes an updraft. All forces together conspire to create considerable convection currents and pressure differences. Just as pressure differences in the atmosphere can produce tornadoes, pressure differences in a fire can cause what is called a *blowup*: a whirl of fire that can travel at nearly the speed of the wind.

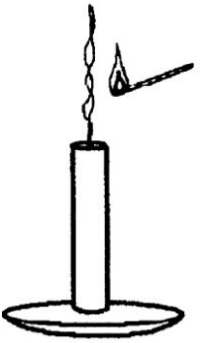
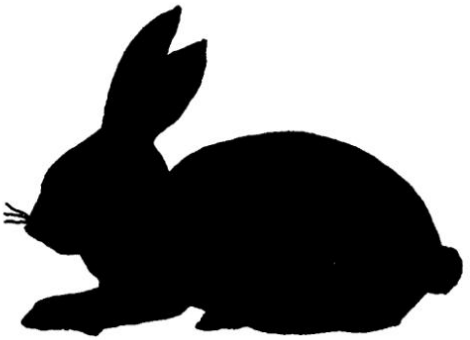


Figure 1.3.1. Bring a lit match up to a recently extinguished candle wick and the flame will "jump" through the burning gases and relight the wick.

In his excellent monograph *Young Men and Fire*, Norman Mclean states that these natural phenomena are poorly understood because there have been few reported direct observations: to compile a report the observer would have to be a survivor and a survivor would be someone who did not wait around to observe.[1] As Mclean details, however, one person, a parachuting firefighter, or *firejumper* as they are called, managed to survive such a whirling fire. He burned a bare patch ahead of the blowup and buried his face in the ashes as the blowup passed over, an event that took five minutes by his watch. He reported being picked up off the ground several times by the vortex but was able to maintain minimal breathing because of oxygen that remained close to the ground. That oxygen was necessary for his bodily combustion—the redox reaction that had to continue if he were to return to tell the story.

As pervasive as redox reactions are, however, they are certainly not all there is to chemistry. Another very common and important type of reaction is found in the acid-base reactions that occur in the air, on skin, in stomachs, in sewers, and even in swimming pools and fish tanks. The fish-tank indicator that we used to detect carbon dioxide is an acid-base indicator. It worked to detect carbon dioxide because carbon dioxide dissolved in water turns water slightly acidic. Intrigued? Read on.



Demonstration 4: Purple-Cabbage Indicator

Tragedy is like strong acid—it dissolves away all but the very gold of truth.
—D. H. Lawrence, *The Letters of D. H. Lawrence,* ca. 1920

It has been known since at least the 1600s, and probably much earlier, that some highly colored plants and flowers, such as red cabbage, contain a dye that will assume various colors in different chemical environments. Such a dye is called an *indicator* because its color indicates its environment. The purple-cabbage indicator is used to indicate whether a solution is acidic or basic. Briefly, acid solutions are sour solutions such as lemon juice and vinegar. Basic solutions are caustic and bitter like soapy water and household ammonia.

To see how this works, put on your safety glasses and then add about a tablespoon (15 milliliters) of purple-cabbage indicator (prepared as outlined in the "Shopping List and Solutions") to a cup (240 milliliters) of water and observe the color. It should be light lavender. Now make up a saturated solution of baking soda in water. A saturated solution is one in which no more solid can be made to dissolve. Baking soda is not very sol-

uble in water, so it should not take more than two tablespoons (30 milliliters) of baking soda per cup of water (240 milliliters). The solution is saturated when there is baking soda left on the bottom of the glass that will not dissolve. Add purple-cabbage indicator to this mixture and it will turn blue. Baking soda dissociates into two ions in water, the sodium ion and the *hydrogen carbonate ion*, otherwise known as *bicarbonate*. The sodium ion is neutral, but the bicarbonate is a base. Purple-cabbage indicator turns blue in baking soda solution; therefore, blue indicates a basic solution when using purple-cabbage indicator.

Add some purple-cabbage indicator to a sample of vinegar. Vinegar is an acid. The vinegar turns the cabbage indicator bright pink, so pink indicates an acid solution when using purple-cabbage indicator.

Using one-cup (240-milliliter) samples of saturated baking soda solution, vinegar, ammonia, and water, you can set up the following banks of solutions and observe the colors that result when you add purple-cabbage, swimming-pool, or fish-tank indicator.

Swimming-pool phenol red indicator	
water	reddish orange
vinegar	yellow
ammonia	reddish violet
baking soda	pink
Fish-tank bromothymol blue indicator	
water	greenish blue
vinegar	pale yellow
ammonia	pale blue
baking soda	blue
Purple-cabbage indicator	
water	purple
vinegar	shocking pink
ammonia	green
baking soda water	teal blue

Now take some of the clear, carbonated soda that was a suggested purchase in the "Shopping List and Solutions" and add the purple-cabbage indicator. You should observe a pale pink solution. Carbonated soda

is slightly acidic because carbon dioxide dissolves in water to form carbonic acid. Now take the baking soda/purple-cabbage solution and slowly and carefully pour the blue liquid that is over the undissolved baking soda (a process called *decanting* by chemist and wine connoisseur alike) into the carbonated soda/purple-cabbage solution. Eventually the solution should have a color very close to that for indicator in pure water. Why? Because acid *neutralizes* base and vice versa. What's neutral? Water. What was the old-time remedy for stomach acid? Baking soda.

Cheers!

CHAPTER 4

The Basic Stuff

"Ask Asia for a cup of sour cream, then your cakes will be light without much soda, which I don't like," was the first order. Demi tore downstairs, and returned with the cream, also a puckered-up face, for he had tasted it on his way, and found it so sour that he predicted the cakes would be uneatable. Mrs. Jo took this occasion to deliver a short lecture from the stepladder on the chemical properties of soda, to which Daisy did not listen, but Demi did, and understood it, as he proved by the brief but comprehensive reply: "Yes, I see, soda turns sour things sweet, and the fizzling up makes them light. Let's see you do it, Daisy."

—Louisa May Alcott, *Little Men*, 1871

In the above quote, Demi's sour-sweet reaction is an acid-base reaction. Baking soda is a base (as we found in the purple-cabbage demonstration), and sour cream is an acidic mixture: sour cream is made sour by the acid products of bacterial action. In the bottle rocket demonstration, we saw another acidic solution, vinegar, encounter baking soda and experi-

ence an explosive reaction. This same type of reaction, on a more modest scale, is what makes Daisy's cakes light:

acid + baking soda → carbonic acid → carbon dioxide (a gas) + water

Acid reacts with baking soda, and their product, carbonic acid, quickly turns into carbon dioxide and water. Carbon dioxide is the gas that blew the cork out of the bottle rocket and that blows bubbles into our fluffy little cakes. The reason you were cautioned to buy baking soda in the "Shopping List and Solutions," rather than baking powder, is because baking powder contains an acidic compound such as cream of tartar. Cream of tartar acts as a *leavening agent*; that is, when it dissolves in water with baking soda they come together and react to form bubbles.

Not all acid-base reactions create such volatile products. Vinegar is an acid, as indicated by the pink color of its solution with purple-cabbage indicator, and ammonia is a base, as indicated by the blue color of its solution with the purple-cabbage indicator. But when vinegar and ammonia are poured together without added indicator, there are none of the usual signs of a chemical reaction: no new fumes, no frothing, no heat, no light, no color change, no solid formation. If it weren't for the sour odor of vinegar and the pungent odor of ammonia, it would appear that two clear, watery liquids were being poured together. However, if a teaspoon (5 milliliters) of purple-cabbage indicator is added to a cup (240 milliliters) of each before pouring them together, the result will be a change in the color of the solution. Indicators are used to show evidence of a chemical reaction when there may be no other obvious sign.

So what are acids and bases? Vinegar is actually a dilute solution of acetic acid in water, about a 5 percent solution, but it rather nicely displays the characteristic properties of acids: they are sour, they turn purple-cabbage indicator red or pink, and they react with bases to form water. A solution of sodium bicarbonate nicely displays several of the characteristics of basic solutions: it tastes bitter, it turns purple-cabbage indicator blue, and it reacts with acids to form water. The last property, listed for both acid and base, the ability to react with each other, is really the defining property because acid-base reactions, like redox reactions, occur in tandem: one substance acts as an acid and one substance acts as a base. Acid neutralizes base and base neutralizes acid.

People have understood the notion of neutralization—the negation of the properties of one substance by another—since the earliest pharmacopeias prescribed ground-up, toasted egg shell as a comfort for stomach problems. Egg shell is principally calcium carbonate, the same material that makes up chalk, and this material, being a base, can cancel excess stomach acid, the acid that is produced naturally in the stomach to digest food. Basic materials are also called *alkaline* materials, partly from tradition—the word *alkali* is Arabic in origin, as are *alchemy, alcohol, algebra, alcove,* and *algorithm*—and partly to avoid confusion: calling something a basic material sounds too much like calling it a fundamental material rather than something that reacts with an acid.

Many examples of acid-base reactions can be found in cooking, such as the soda–sour cream reaction in *Little Men*. In addition, the unfortunate result of acid rain (the formation of which we will discuss later) is that the acid in the rain reacts with the carbonates found in limestone and marble, which causes the deterioration of statues, some of which had managed to survive without corrosion for thousands of years before the advent of the industrial age. (But before one completely condemns the industrial age, it should be remembered that the bacteria of Black Death, smallpox, and syphilis also managed to survive for thousands of years before modern technology brought them to bay.) This ability of acid rain to dissolve marble brings up another property common to all acids and bases: they are corrosive.

The corrosive property is addressed in the quotation from the story *Daddy Long Legs* given in the introduction and repeated here for convenience.

> *I must go to the laboratory and look into a little matter of acids and salts and alkalis. I've burned a hole as big as a plate in the front of my chemistry apron, with hydrochloric acid. If the theory worked, I ought to be able to neutralize that hole with good strong ammonia, oughtn't I?*
>
> —Jean Webster, *Daddy-Long-Legs*, 1912

The joke is that base neutralizes acid, so the chemistry student supposes that an application of base will reverse the effects of the acid and restore the destroyed material of the apron. It is interesting to note that readers in the early 1900s were supposed to understand the joke without explanation—and that first-year chemistry students in the early 1900s were

encouraged to handle chemicals sufficiently active to eat a hole in a chemistry apron. But the concept of acidity is not exactly unfamiliar in our twenty-first century. We speak of an acid tongue, an acid smile, an acid voice, or acid criticism. Acid solutions are part of our kitchens: vinegar is an acetic acid solution, lemon juice is a citric acid solution, and carbonated sodas contain carbonic acid. Basic solutions and substances are also part of our common experience: ammonia is a basic solution, as is the lye used to unclog drains. Many times we use the older terms—alkali or alkaline—to describe basic materials, such as "alkaline batteries" or "adjusting the alkalinity of the swimming pool." Alkalosis, the dangerous condition of blood made too basic by the lack of carbon dioxide, can be caused by hyperventilation.

In fact, the entire human body can function well only within a very narrow range of alkalinity and acidity. But before you toss out the hot sauce for health's sake, be assured that the body has devised a method for maintaining the blood at the alkalinity level necessary for proper functioning even if vinegar is used in the salad dressing or seltzer water is consumed. The blood is *buffered*: it is a solution of significant quantities of both weak acids and weak bases so that small additions of acid or base will not cause a notable shift in alkalinity. A good analogy for a buffer might be a punching bag: an unanchored opponent might be sent reeling with a right jab, but the anchored punching bag absorbs the blow and comes back for more. A buffered solution can absorb the blows of added acid or base and bounce back to nearly its original condition. The situation can be demonstrated with another bodily fluid that is buffered: milk.

Take two of the small clear containers suggested in the "Shopping List and Solutions" and add a half cup (120 milliliters) of whole milk to each. It is necessary to use *whole* milk because cows care little about calories and provide buffers that work with the intact, not skimmed, substance. Take two more of the small clear containers and add a half cup (120 milliliters) of water to these. To all four samples add several drops of the phenol red indicator from the swimming pool test kit suggested in the "Shopping List and Solutions." The purple-cabbage indicator does not work as well in this demonstration because you have to add too much to see a vivid color and the excess liquid dilutes the milk and compromises the buffering action. Adding several drops of swimming-pool-test-kit pH indicator should give both the milk and the water a pale yellow tinge.

Now add a drop of baking soda solution (the liquid over the undissolved solid) to one of the samples of water and one of the samples of milk. If you have managed to obtain some eyedroppers, these work best, but if not, you can dip a straw in the solution and cover the end with your finger to capture a drop. Stir the solutions. The baking soda–water solution should turn a shocking pink, but the milk should stay pale yellow.

The best comparison is made by putting the containers on a white sheet of paper and looking directly down into the container. There is a considerable difference in the color of the water-indicator solution with and without baking soda. With and without baking soda, the indicator-milk solutions look the same. The milk has buffered, or moderated, the effect of the sodium bicarbonate.

In the blood, it is carbon dioxide that provides the buffer. Carbonated soda is a solution of carbonic acid formed when water is saturated with carbon dioxide, and in the purple-cabbage demonstration we saw that carbonated soda is slightly acidic. Our blood is infused with carbon dioxide, the end product of metabolism, and this carbon dioxide forms carbonic acid in the blood. As we saw in our purple-cabbage demonstration, sodium bicarbonate, baking soda, is a base. Bicarbonate, formed by removing just one hydrogen in carbonic acid, is produced by the kidneys to act as a base in the blood. So carbon dioxide is used to form both blood acid and blood base. When the blood is too acidic, the bicarbonate base sops up, neutralizes, the excess acid. When the blood is too alkaline, carbonic acid neutralizes the excess base. But as can be demonstrated by pouring baking soda solution directly into the indicator-milk solution, buffers do not have an unlimited ability to absorb acid or base: the milk will soon assume a pink shade of its own. In blood, if the balance tips too far one way or the other, the conditions of alkalosis or acidosis can occur, which are very serious medical conditions.

The blood is not the only system in the body that has to have a fairly stable acid-base environment. Buffered aspirin is aspirin that has been altered so that it will not be too acidic and upsetting to the stomach. Many soaps and shampoos are specifically blended so that they do not differ significantly from the acid quality of the skin or hair. If there were a difference in acid quality, they could cause an acid-base reaction when applied. An acid-base reaction on the skin or in the eyes could cause damage to sensitive cells, which is why you are careful to use gloves and eye protection while doing these demonstrations.