INTERNATIONAL BESTSELLER!

C 6

Carbon

N 7

Nitrogen

O 8

Oxygen

"This glorious book is more than just a guide to the elements; it will fundamentally deepen your appreciation of the substances that make up our world."
–Oliver Sacks

Si 14

Silicon

P 15

Phosphorus

S 16

Sulfur

Cl 17

Chlorine

18

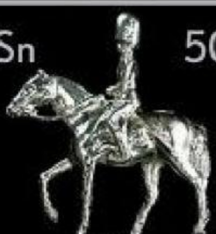

Argon

Ge 32

Germanium

THE Elements

A Visual Exploration of Every Known Atom in the Universe

Sn 50

Tin

Sb 51

Antimony

Te 52

Tellurium

I 53

Iodine

54

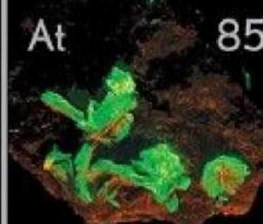

Xenon

THEODORE GRAY

Photographs by Theodore Gray and Nick Mann

Po 84

Polonium

At 85

Astatine

Rn 86

Radon

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There is not anything which returns to nothing, but all things return dissolved into their elements.

–Lucretius, De Rerum Natura, 50 BC

The periodic table is the universal catalog of everything you can drop on your foot. There are some things, such as light, love, logic, and time, that are not in the periodic table. But you can't drop any of those things on your foot.

The earth, this book, your foot—everything tangible—is made of elements. Your foot is made mostly of oxygen, with quite a bit of carbon joining it, giving structure to the organic molecules that define you as an example of carbon-based life. (And if you're not a carbon-based life-form: Welcome to our planet! If you have a foot, please don't drop this book on it.)

Oxygen is a clear, colorless gas, yet it makes up three-fifths of the weight of your body. How can that be?

Elements have two faces: their pure state, and the range of chemical compounds they form when they combine with other elements. Oxygen in pure form is indeed a gas, but when it reacts with silicon they become together the strong silicate minerals that compose the majority of the earth's crust. When oxygen combines with hydrogen and carbon, the result can be anything from water to carbon monoxide to sugar.

Oxygen atoms are still present in these compounds, no matter how unlike pure oxygen the substances may appear. And the oxygen atoms can always be extracted back out and returned to pure gaseous form.

But (short of nuclear disintegration) each oxygen atom can never itself be broken down or taken apart into something simpler. This property of indivisibility is what makes an element an element.

In this book I try to show you both faces of every element. First, you will see a great big photograph of the pure element (whenever that is physically possible). On the facing page you will see examples of the ways that element lives in the world—compounds and applications that are especially characteristic of it.

Before we get to the individual elements, it's worth looking at the periodic table as a whole to see how it is put together.

The periodic table, this classic shape, is known the world over. As instantly recognizable as the Nike logo, the Taj Mahal, or Einstein's hair, the periodic table is one of our civilization's iconic images.

The basic structure of the periodic table is determined not by art or whim or chance, but by the fundamental and universal laws of quantum mechanics. A civilization of methane-breathing pod-beings might advertise their pod-shoes with a square logo, but their periodic table will have recognizably the same logical structure as ours.

Every element is defined by its atomic number, an integer from 1 to 118 (so far—more will no doubt be discovered in due time). An element's atomic number is the number of protons found in the nucleus of every atom of that element, which in turn determines how many electrons orbit around each of those nuclei. It's those electrons, particularly the outermost "shell" of them, that determine the chemical properties of the element. (Electron shells are described in more detail on page 12)

The periodic table lists the elements in order by atomic number. The sequence skips across gaps in ways that might seem quite arbitrary, but that of course are not. The gaps are there so that each vertical column contains elements with the same number of outer-shell electrons.

And that explains the most important fact about the periodic table: Elements in the same column tend to have similar chemical properties.

Let's look at the major groups in the periodic table, as defined by the arrangement of columns.

The very first element, hydrogen, is a bit of an anomaly. It's conventionally placed in the leftmost column, and it does share some chemical properties with the other elements in that column (principally the fact that in compounds, it normally loses one electron to form an H^+ ion, just as sodium, element 11, loses one electron to form Na^+). But hydrogen is a gas, while the other elements in the first column are soft metals. So some presentations of the periodic table isolate hydrogen in a category all its own.

The other elements of the first column, not counting hydrogen, are called the *alkali metals*, and they are all fun to throw into a lake. Alkali metals react with water to release hydrogen gas, which is highly flammable. When you throw a large enough lump of sodium into a lake, the result is a huge explosion a few seconds later. Depending on whether you took the right precautions, this is either a thrilling and beautiful experience or the end of your life as you have known it when molten sodium sprays into your eyes, permanently blinding you.

Chemistry is a bit like that: powerful enough to do great things in the world, but also dangerous enough to do terrible things just as easily. If you don't respect it, chemistry bites.

The elements of the second column are called the *alkali earth metals*. Like the alkali metals, these are relatively soft metals that react with water to liberate hydrogen gas. But where the alkali metals react explosively, the alkali earths are tamer—they react slowly enough that the hydrogen does not spontaneously ignite, allowing calcium (20), for example, to be used in portable hydrogen generators.

The wide central block of the periodic table is known as the *transition metals*. These are the workhorse metals of industry—the first row alone is a veritable who's who of common metals. All the transition metals except mercury (80) are fairly hard, structurally sound metals. (And so, in fact, is mercury, if you cool it enough. Mercury freezes into a metal remarkably like tin, element 50.) Even technetium (43), the lone radioactive element in this block, is a sturdy metal like its neighbors. It's just not one you'd want to make a fork out of—not because it wouldn't work, but because it would be very expensive and would slowly kill you with its radioactivity.

The transition metals as a whole are relatively stable in air, but some do oxidize slowly. The most notable example is of course iron (26), whose tendency to rust is by far our most destructive unwanted chemical reaction. Others, such as gold (79) and platinum (78), are prized for their extreme resistance to corrosion.

The two empty spots in the lower left corner are reserved for the lanthanide and actinide series of elements, highlighted on page 11. According to the logic of the periodic table, a fourteen-element-wide gap should appear between the second and third columns, with the elements of the lanthanide and actinide groups inserted in that gap. But because this would make the periodic table impractically wide, the convention is to close up that gap and display the rare earths in two rows at the bottom.

The lower left triangle here is known as the *ordinary metals*, though in reality most of the metals that people think of as ordinary are in fact transition metals in the previous group. (By now you may have noticed that the great majority of elements are metals of one sort or another.)

The upper right triangle is known as the *nonmetals*. (The next two groups, halogens and noble gases, are also not metals.) The nonmetals are electrical insulators, while all metals conduct electricity at least to some extent.

Between the metals and nonmetals is a diagonal line of fence-sitters known as the *metalloids*. These are, as you might expect from the name, somewhat like metal and somewhat not like metal. In particular they conduct electricity, but not very well. The metalloids include the semiconductors that have become so important to modern life.

The fact that this line is diagonal violates the general rule that elements in a given vertical column share common characteristics. Well, it’s only a general rule—chemistry is too complicated for any rule to be absolutely hard and fast. In the case of the metal-to-nonmetal boundary, several factors compete with each other to determine whether an element falls into one camp or the other, and the balance drifts toward the right as you move down the table.

The seventeenth (second-to-last) column is called the *halogens*, and its members are a pretty nasty lot in pure form. All the elements of this column are highly reactive, violently smelly substances. Pure fluorine (9) is legendary for its ability to attack nearly anything; chlorine (17) was used as a poison gas in World War I. But in the form of compounds such as fluoridated toothpaste and table salt (sodium chloride), the halogens are tamed for domestic use.

The very last column is the *noble gases*. Noble is used here in the sense of "above the business of the common riffraff." Noble gases almost never form compounds with each other or with any other elements. Because they are so inert, the noble gases are often used to shield reactive elements, since under a blanket of noble gas there's nothing for the reactive element to react with. If you buy sodium from a chemical supplier, it will come in a sealed container filled with argon (18).

These two groups are known collectively as the *rare earths*, despite the fact that some of them are not rare at all. The top row, starting with lanthanum (57), is known as the *lanthanides*; and you will not be surprised to learn that the bottom row, starting with actinium (89), is known as the *actinides*.

As you will read when you get to lutetium (71), the lanthanides are especially notorious for being chemically similar to each other. Some are so similar that people argued for years whether they were really separate elements at all.

All the actinides are radioactive, with uranium (92) and plutonium (94) being the most famous. Adding the actinides to the standard layout of the periodic table can be blamed on Glenn Seaborg, largely because he was responsible for discovering so many new elements in this range that a new row became necessary. (Although new elements have been discovered by many people, Seaborg is the only one forced to invent a row to display all of his discoveries.)

Now that we have seen the periodic table as a whole and in parts, we're ready to start our journey through the wild, beautiful, up-and-down, fun, and terrifying world of the elements.

This is all there is. From here to Timbuktu, and including Timbuktu, everything everywhere is made of one or more of these elements. The infinite variety of combinations and recombinations that we call chemistry starts and ends with this short and memorable list, the building blocks of the physical world.

Almost everything you see in this book is sitting somewhere in my office, except that one thing the FBI confiscated and a few historical objects. I had a great time collecting these examples of the vibrant diversity of the elements, and I hope you have as much fun reading about them.

See you at hydrogen!

How the Periodic Table Got Its Shape

Hang on tight, we're going to explain quantum mechanics in one page. (If you find this section too technical, feel free to skim it—there isn't going to be a quiz at the end.)

Every element is defined by its atomic number, the number of positively charged protons in the nucleus of every atom of that element. These protons are matched by an equal number of negatively charged electrons, found in "orbits" around the nucleus. I say "orbits" in quotes because the electrons are not actually moving around their orbits like planets around a star. In fact, you can't really speak of them as moving at all.

Instead, each electron exists as a probability cloud, more likely to be in one place than another, but not actually *in* any one place at any given time. The figures below show the various three-dimensional shapes of the probability clouds of electrons around a nucleus.

The first type, called an "s" orbital, is totally symmetrical—the electron is not any more likely to be in one direction than another. The second type, called a "p" orbital, has two lobes, meaning the electron is more likely to be found on one side or the other of the nucleus, and less likely to be found in any direction in between.

While there is only one "s"-type orbital, there are three "p" types, with lobes pointing in the three orthogonal directions (x, y, z) of space. Similarly there are five different types of "d" orbitals and seven different types of "f" orbitals, with increasing numbers of lobes. (You may think of these shapes as a bit like three-dimensional standing waves.)

Each shape of orbital can exist in multiple sizes, for example the 1s orbital is a small sphere, 2s is a larger sphere, 3s is larger still, and so forth. The energy required for an electron to be in any given orbital increases as the orbit becomes bigger. And all else being equal, electrons will always settle into the smallest, lowest-energy orbit.

So do all the electrons in an atom normally sit together in the lowest-energy 1s orbital? No, and here we come to one of the most fundamental discoveries in the early history of quantum mechanics: No two particles can ever exist in exactly the same quantum state. Because electrons have an internal state known as "spin," which can be either up or down, it turns out that exactly two electrons can reside in a given orbital—one with spin up and one with spin down.

Hydrogen has only one electron, so it sits in the 1s orbital. Helium has two, and they both fit into 1s, filling it to its capacity of two. Lithium has three, and since there is no room in 1s anymore, the third electron is forced to sit in the higher-energy 2s orbital. And so on—the orbitals are filled one at a time in order of increasing energy.

Look at the Electron Filling Order diagram on the right side of any element page in this book, and you'll see a graph of the possible orbitals from 1s to 7p, with a red bar indicating which ones are filled with electrons (7p is the orbital of highest energy occupied by electrons of any known element). The exact order in which orbitals are filled turns out to be surprisingly subtle and complex, but you can watch it happen as you flip through the pages of this book. Pay particular attention around gadolinium (64)—if you think you've got it figured out, your confidence might be shaken by what happens there.

It is this filling order that determines the shape of the periodic table. The first two columns represent electrons filling “s” orbitals. The next ten columns are electrons filling the five “d” orbitals. The final six columns are electrons filling the three “p” orbitals. And last but not least, the fourteen rare earths are electrons filling the seven “f” orbitals. (If you’re asking yourself why helium, element 2, is not above beryllium, element 4, congratulations—you’re thinking like a chemist rather than a physicist. Eric Scerri’s book, referenced in the bibliography, is a good start toward answering such questions.)

s orbital

p orbitals

d orbitals

f orbitals

Elemental

Eve you need to know. Nothing you do

NAVIGATION TABLE
The mini table on every element page has one highlighted yellow square to show you where that element is located on the periodic table. The colors divide the table into the groups described on the preceding pages.

ATOMIC WEIGHT
An element's atomic weight (not to be confused with its atomic number) is the average weight per atom in a typical sample of the element, expressed in "atomic mass units," or amu. The amu is defined as 1/12 the mass of a ^{12}C atom. Roughly speaking, one amu is the mass of one proton or one neutron, and thus an element's atomic weight is approximately equal to the total number of protons and neutrons in its nucleus.

However, you will notice that the atomic weights of some elements fall well between whole integers. When typical samples of an element contain two or more naturally occurring isotopes, the averaging of isotopic weights explains the fractional amu. (Isotopes are explained in more detail under protactinium, element 91; the basic idea is that an element's isotopes all have the same number of protons, and thus the same chemistry, but differ in the numbers of neutrons in their nuclei).

DENSITY

The density of an element is defined as the idealized density of a hypothetical flawless single crystal of the absolutely pure element. This can never be realized exactly in practice, so the densities are generally calculated from a combination of the atomic weight and x-ray crystallographic measurements of the spacing of atoms in crystals. The density is given in units of grams per cubic centimeter.

ATOMIC RADIUS

The density of a material depends on two things: how much each atom weighs, and how much space each atom takes up. The atomic radius shown for each element is the calculated average distance to the outermost electrons from the nucleus in picometers (trillionths of a meter). The diagrams are merely schematic—they represent all the electrons in their respective electron shells, with the overall size matching the size of the atom, but the position of individual electrons is not to scale, nor do electrons actually exist as sharp points spinning around the atom. The dashed blue reference circle shows the radius of the largest of all atoms, cesium (55).

CRYSTAL STRUCTURE

The crystal structure diagram shows the arrangement of atoms (the unit cell that is repeated to form the whole crystal) when the element is in its most common pure crystalline form. For elements that are normally gas or liquid, this is the crystal form they take on when they are cooled enough to freeze solid.

ELECTRON FILLING ORDER

This diagram shows the order in which electrons fill the available atomic orbitals, which are explained in detail on the preceding page.

ATOMIC EMISSION SPECTRUM

When atoms of a given element are heated to very high temperatures, they emit light of characteristic wavelengths, or colors, which correspond to the differences in energy levels between their electron orbitals. This diagram shows the colors of these lines, each one corresponding to a particular energy-level difference, arranged into a spectrum from the barely visible red at the top to the nearly ultraviolet at the bottom.

STATE OF MATTER

This temperature scale in degrees Celsius shows the range of temperatures over which the element is solid, liquid, or gas. The boundary between solid and liquid is the melting point, while the boundary between liquid and gas is the boiling point. Twist the pages of the book to spread the edges of the pages out, and you will see a graph of the melting and boiling points, which shows very pronounced trends across the periodic table.

Hydrogen

Stars shine because they are transmuting vast amounts of hydrogen into helium. Our sun alone consumes six hundred million tons of hydrogen per second, converting it into five hundred and ninety-six million tons of helium. Think about it: Six hundred million tons *per second*. Even at *night*.

And where does the other four million tons per second go? It's converted into energy according to Einstein's famous formula, $E{=}mc^2$. About three-and-a-half-pounds-per-second's worth finds its way to the earth, where it forms the light of the dawn rising, the warmth of a summer afternoon, and the red glow of a dying day.

The sun's ferocious consumption of hydrogen sustains us all, but hydrogen's importance to life as we know it begins closer to home. Together with oxygen it forms the clouds, oceans, lakes, and rivers. Combined with carbon (6), nitrogen (7), and oxygen (8), it bonds together the blood and body of all living things.

Hydrogen is the lightest of all the gases—lighter even than helium—and much cheaper, which accounts for its ill-advised use in early airships such as the *Hindenburg*. You may have heard how well that went, though in fairness the people died because they fell, not because they were burned by the hydrogen, which in some ways is less dangerous to have in a vehicle than, say, gasoline.

Hydrogen is the most abundant element, the lightest, and the most beloved by physicists because, with only one proton and one electron, their lovely quantum mechanical formulas actually work exactly on it. Once you get to helium with two protons and two electrons, the physicists pretty much throw up their hands and let the chemists have it.

By weight, 75 percent of the visible universe is hydrogen. Ordinarily it is a colorless gas, but vast quantities of it in space absorb starlight, creating spectacular sights such as the Eagle Nebula, seen here by the Hubble Space Telescope.

Tritium (^3H) luminous key chain, illegal in the U.S. because it is deemed a “frivolous” use of this strategic material.

Tritium watches, on the other hand, are legal in the U.S.

The orange-red glow of an oxygen-hydrogen flame.

The sun works by turning hydrogen into helium.

The mineral scolecite, $CaAl_2Si_3O_{10} \cdot 3H_2O$, from Puna, Jalgaon, India.

The inside of a high-speed thyratron, a type of electronic switch filled with a small amount of hydrogen gas.

Helium

Helium is named for the Greek god of the sun, Helios, because the first hints of its existence were dark lines in the spectrum of sunlight that could not be explained by the presence of any elements known at the time.

It might seem a paradox that an element common enough to fill party balloons with was the first element to be discovered in space. The reason is that helium is one of the noble gases, so named because they do not interact with the common riffraff of elements, remaining inert and aloof to nearly all chemical bonding. Because it does not interact, helium could not easily be detected by conventional wet chemical methods.

As a replacement for hydrogen in airships, helium, which is completely nonflammable, has much to recommend it. The main problem is that it's a lot more expensive, and provides somewhat less lift. Anyone want to go for a ride in the low bid model?

The helium we use today is extracted from natural gas as it comes out of the ground. But unlike all other stable elements, it was not deposited there when the earth was formed. Instead it was created over time by the radioactive decay of uranium (92) and thorium (90). These elements decay by alpha particle emission, and "alpha particle" is simply the physicist's name for the nucleus of a helium atom. So when you fill a party balloon, you're filling it with atoms that just a few tens or hundreds of millions of years ago were random protons and neutrons in the nuclei of large radioactive atoms. That, frankly, is weird. Though not as weird as the way lithium messes with your mind.

Ordinarily a colorless, inert gas, helium glows creamy pale peach when an electric current runs through it.

Pure helium is an invisible gas, as in this antique sample ampoule.

A characteristic helium peach-colored glow is visible through the open side of this helium-neon laser. The laser light coming out the front is neon red.

Disposable helium tanks are available in party supply stores but often contain added oxygen to prevent suffocation if inhaled by children.

Helium-filled latex party balloons don’t last long as this tiny atom escapes rapidly. Metalized Mylar balloons last days instead of hours.

Lithium

Lithium is a very soft, very light metal. So light that it floats on water, a feat matched by only one other metal, sodium (11). While floating on water, lithium will react with that water, releasing hydrogen gas at a steady, moderate rate. (The real excitement in this department begins with sodium.)

Despite its reactive nature, lithium is widely used in consumer products. Lithium metal inside lithium-ion batteries powers countless electronic devices, from pacemakers to cars, including the laptop on which I am typing this text. Lithium-ion batteries pack tremendous power into not much weight, in part because of lithium's low density. Lithium stearate is also used in the popular lithium grease found on cars, trucks, and mechanics.

People who pay attention to these things have noticed an interesting fact: There's only one place in the world with a really large amount of easily recoverable lithium. If electric cars based on lithium-ion batteries ever become very widespread, you might want to keep an eye on Bolivia.

The lithium ion has another trick up its sleeve: It keeps some people on an even emotional keel. For reasons that are only vaguely understood, a steady dose of lithium carbonate (which dissolves into lithium ions in the body) smoothes out the highs and lows of bipolar disorder. That a simple element could have such a subtle effect on the mind is testimony to how even a phenomenon as complex as human emotion is at the mercy of basic chemistry.

Lithium is soft, reactive, and helps keep things in balance. Beryllium is, well, let's just say *different*.

Lithium is soft enough to cut with hand shears, which leave marks such as you see on this sample of the pure metal.

Lithium carbonate pills control mood swings.

The mineral elbaite, $Na(LiAl)_3Al_6(BO_3)_3Si_6O_{18}(OH)_4$, from Minas Gerais, Brazil.

Common lithium grease contains lithium stearate to improve performance.

Lithium batteries can be exotic, like the pacemaker battery above, or common, like this standard AA-sized disposable lithium cell.

Beryllium

Beryllium is a light metal (though three and a half times the density of lithium, it's still significantly less dense than aluminum, element 13). Where lithium is soft, low-melting, and reactive, beryllium is strong, melts at a high temperature, and is notably resistant to corrosion.

These properties, combined with its high cost and poisonous nature, account for the unique niche beryllium has carved out for itself: missile and rocket parts, where cost is no object, where strength without weight is king, and where working with toxic materials is the least of your worries.

Beryllium has other fancy applications. It is transparent to x-rays, so it's used in the windows of x-ray tubes, which need to be strong enough to hold a perfect vacuum, yet thin enough to let the delicate x-rays out. A few percent of it alloyed with copper (29) forms a high-strength, nonsparking alloy used for tools deployed around oil wells and flammable gases, where a spark from an iron tool could spell disaster, in great big flaming red letters.

In keeping with the sport of golf's tendency to use high-tech materials out of a desperate hope that they may help get the ball where it's supposed to go, beryllium copper is also used in golf-club heads. Needless to say, it doesn't help any more than the manganese bronze or titanium (22) used for the same purpose.

Combining beauty with brawn, the mineral beryl is a crystalline form of beryllium aluminum cyclosilicate. You may be more familiar with the green and blue varieties of beryl, which are known as emerald and aquamarine.

Beryllium: A debonair, James Bond–style metal able to launch rockets one minute and charm the ladies the next. Then there's boron.

This pure broken crystal of refined beryllium ordinarily would be melted down and turned into strong, lightweight parts for missiles and spacecraft.

Beryllium oxide high-voltage insulator.

Beryllium copper nonsparking gas-valve wrench.

Complex beryllium missile gyroscope.

Beryllium foil windows mounted in an x-ray tube.

A large aquamarine beryl ($Be_3Al_2Si_6O_{18}$) from the author's father's extensive collection.

Beryllium copper golf club.

Boron

Poor boron—with a name like that, how can it get any respect? It doesn't help that boron's most commonly found in borax, the laundry aid. But boron is more glamorous than you might think.

Combine boron (5) with nitrogen (7), and you get crystals similar to those of their average, carbon (6), the element that forms diamond. Cubic boron nitride crystals are very nearly as hard as diamond, but much less expensive to create and more heat resistant, making them popular abrasives for industrial steelworking.

Recent theoretical calculations indicate that the alternate wurtzite-crystal form of boron nitride, as yet never created in single-crystal form, might actually be harder than diamond under certain conditions, and for certain technical definitions of "hard." Unseating diamond from its long reign as the hardest known material would be quite a coup, but for the time being "wurtzite" boron nitride's only accomplishment is causing an annoying footnote you now have to put next to any claim that diamond is the hardest known substance.

Boron carbide, also one of the hardest known substances, even has a genuine secret-agent application: Granules of it poured into the oil-fill hole of an internal combustion engine will destroy the engine by irreparably scoring the cylinder walls. Of slightly less interest to the CIA is the fact that boron is critical in cross-linking the polymers that gives Silly Putty its amazing ability to be both soft and moldable in your hand, yet hard and bouncy when you throw it against the wall.

But while boron is not quite the frump you might expect from its name, it's really not in the same league as carbon.

Boron is rarely seen in pure form, as in these polycrystalline lumps. While extremely hard, boron is too brittle in pure form to have any practical applications.

Cubic boron nitride is used in machine tool inserts for cutting hardened steel.

Silly Putty®.

Boron carbide engine sabotage solution.

Boric acid was recommended for everything from eye washing to ant poison.

Carbon

Carbon is *the* most important element of life, period. Sure, there are many others without which life would not exist, but from the spiral backbone of DNA to the intricate rings and streamers of the steroids and proteins, carbon is the element whose unique properties tie it all together. The very term "organic compound" refers exclusively to chemicals containing carbon.

Not content to be the foundation of all life on earth, carbon also forms diamond, the hardest known substance (at least for now; challengers are discussed under boron, element 5). But contrary to popular belief, diamonds are not particularly rare, nor are they unusually beautiful, nor are they forever: all three are myths created by the DeBeers diamond company. Diamonds would cost a tenth as much but for DeBeers's monopoly control. Cubic zirconia or crystalline silicon carbide are just as pretty. And at high enough temperatures, diamonds burn up into nothing but carbon dioxide.

If I were writing these words twenty-five years or so ago, I would probably have been doing it with carbon. The "lead" in pencils is actually graphite, a form of carbon, and has been since the 16th-century discovery in the English Lake District of the great mine at Borrowdale, the first source of pure graphite.

Carbon atoms like to form sheets, like a honeycomb with a carbon atom at each corner. Stack the sheets and you have graphite. Fold them into a sphere and you have a C_{60} "buckyball," named for Buckminster Fuller who invented the geodesic dome. Roll the sheets into tubes and you have the strongest material known to science: carbon nanotubes.

Carbon has now become a focus of political controversy centered on the fact that our civilization is pumping carbon dioxide back into the atmosphere at about 100,000 times the rate it was put away by the dinosaurs and their swamps. Interestingly, the situation with nitrogen is exactly reversed.

A diamond is forever, unless you heat it too much, in which case it burns up into carbon dioxide gas.

A “Congo cube,” natural cheap polycrystalline diamond clusters.

Coal (roughly speaking C_nH_{2n}) carvings are found everywhere that coal is.

A block of graphite (pure carbon) from the first atomic pile, described under fermium, element 100.

Copper-clad graphite welding eletrodes are available in any welding shop

Computer model of C_{60} "bucky ball."

Tiny industrial diamonds embedded in this steel disk turn it into a powerful grinding wheel.

Coal as you buy it for heating and blacksmithing.

Nitrogen

At the same time that modern civilization has been pumping carbon dioxide into the atmosphere, we've been pulling out nitrogen and eating it.

Nitrogen as N_2 in the air is inert and largely useless, but when it's converted to a more reactive form, such as ammonia (NH_3), it becomes a vital fertilizer. Only some plants, beans for example, aided by microorganisms residing in their roots, are able to draw the nitrogen they need directly from the air. This is one reason that, before the advent of cheap nitrogen fertilizer, corn, which cannot "fix" nitrogen, was alternated in the fields with beans or alfalfa, which leave the soil with more nitrogen than it started with.

Just before World War I, Fritz Haber invented a practical process for converting nitrogen from the air into ammonia, one of the most important discoveries in human history. Ammonia fertilizer now feeds a third of the world (the rest being fed mainly by phosphate fertilizers). His work with chlorine (17) was less benevolent, as you can read about under that element.

And since plant growth absorbs carbon dioxide from the air, nitrogen fertilization even helps, at least a bit, with alleviating the effects of global warming.

Liquid nitrogen is a cheap and readily available cryogenic cooling liquid. With a boiling point of -196°C it is cold enough to freeze almost anything. It is used to preserve biological samples, to amuse children by freezing and shattering flowers, and occasionally to make ice cream in record time.

There's a lot of nitrogen around: Over 78 percent of the atmosphere is nitrogen. What's the other 22 percent? Most of it is the oxygen we need to breathe.

A Dewar flask filled with boiling liquid nitrogen at -196°C (-320°F).

Nitrogen-gas canister for a wine-preservation gadget. The claim of 100% purity is suspect: Nothing is ever 100%.

The mineral nitratine ($NaNO_3$)

Silicon nitride (Si_3N_4) is so hard it is used to make cutting tools, such as this milling bit insert.

Silicon nitride (Si_3N_4) ceramic ball bearing for very expensive skateboards.

Nitroglycerine ($C_3H_5N_3O_9$) pills for angina.

Oxygen

If carbon (6) is the foundation of life, then oxygen is the fuel. Oxygen's ability to react with just about any organic compound is what drives the processes of life. Combustion with oxygen also drives your car, your furnace, and if you work for NASA, your rockets. (Actually, the term "fuel" usually refers to the thing that is burned by an "oxidizer," so I'm speaking metaphorically when I say oxygen is the fuel of life. Technically speaking, oxygen is the oxidizer of life.)

The fact that you can light and burn wood, paper, or gasoline has less to do with what those things are made of, and more to do with the fact that our atmosphere is over 21 percent oxygen, providing a ready source of highly reactive oxidizer. Jet airplanes can travel great distances with far less fuel than a comparable rocket would require, because unlike jets that travel in air, rockets must function in the vacuum of space and must therefore carry their oxygen supply with them.

Concentrated into liquid form, oxygen goes from being gently life-giving to life-threateningly fierce. It's fair to say that the real power for most rockets comes not from the fuel they burn, but from their oxygen supply. The Saturn V moon rocket, for example, ran on kerosene. (Yes, we made it to the moon on diesel fuel.) But it wasn't the kerosene that was special, it was the ten cubic yards *per second* of liquid oxygen the Saturn V consumed at full thrust.

Given how intense oxygen is, it might surprise you to learn that it is the most abundant element on earth, accounting for nearly half the weight of the earth's crust and 86 percent of the weight of the oceans. But the crust and the oceans are made not of pure oxygen but of its compounds, and as we will learn from fluorine, the fiercer the element, the more stable its compounds.

At -183°C, oxygen is a beautiful pale blue liquid.

An emergency oxygen generator for aircraft use: Because when worse comes to worst, the one thing you need more than anything else is oxygen.

Disposable oxygen tanks for hobby brazing, and as a refreshing pick-me-up, hold very little oxygen.

In element collections, pure oxygen can only be represented by a seemingly empty bottle.

High-pressure portable oxygen tank for use by medics.

The mineral apophyllite, $KCa_4Si_8O_{20}(F,OH)\cdot 8H_2O$ + $KCa_4Si_8O_{20}(OH,F)\cdot 8H_2O$.

Fluorine

Fluorine is among the most reactive of all the elements. Blow a stream of fluorine gas at almost anything, and it will burst into flame. That includes things not normally thought of as flammable, such as glass. Interestingly, the more reactive an element is, the more stable are its compounds.

When we say fluorine is highly reactive, we mean that a large amount of energy is released when it combines with other elements. The resulting compounds are very stable because the same large quantity of energy must be put back in if you want to tear them apart. This energy must be supplied by some yet-more-reactive substance, of which, in the case of fluorine, there are precious few.

The most famous highly stable fluorine compound is Teflon, which was discovered quite by accident. So many important chemicals have been discovered by accident that one has to wonder what a bunch of bumblers chemists are. Or maybe they are just exceptionally good at spotting serendipity when it ruins their day. Teflon was discovered when its unexpected formation completely ruined an attempt to create the first chlorofluorocarbon refrigerants, which have now been banned as ozone-depleting menaces. Not a bad trade, I'd say.

Teflon is almost completely resistant to chemical attack, and coincidentally also very slippery, which makes it useful in everything from nonstick pans to acid-storage bottles. Fluorine is important primarily because of the stable compounds it forms, while neon forms no stable compounds whatsoever.

Fluorine is a pale yellow gas that reacts violently with virtually everything, including glass. This pure quartz ampoule probably held it for a while anyway.

Fluoride supplement tablets.

A 37-pound cylinder of solid Teflon®.

Teflon® stopcock in a laboratory burette.

Beautiful purple fluorite with hydrocarbon impurities that tint the center yellow.

Teflon® suture with single-use needle.

Gore-Tex® Teflon-based fabric.

Fluoride toothpaste.

Gore-Tex® industrial filter bag.

Teflon® non-stick frying pan.

Neon

Neon is literally up in lights. As in, up there, in those lights, there is neon. So close is the association between the element and its most common application that Times Square and Las Vegas are described as being "awash with neon."

Unlike "platinum" credit cards that contain no platinum, some "neon" lights —the orange-red ones—really do contain neon. When a high-voltage electric discharge is run though a tube filled with low-pressure neon, the gas glows bright orange-red in a fuzzy line down the center of the tube. (Any other color, and it's not neon. And if you see a tube where the light comes from an opaque coating on the inside surface of the glass, rather than from inside the tube itself, you've got yourself a mercury vapor or krypton tube with a phosphor coating.)

Oliver Sacks, in his delightful book *Uncle Tungsten*, describes walking through Times Square with a pocket spectroscope, enchanted by the great variety of spectral lines he could see. That's another way to tell a genuine neon light—by its unique spectrum, unlike that of any other element or phosphor.

Helium-neon lasers were the first continuous-beam lasers in commercial use, and while they have been replaced in many applications by incredibly cheap laser diodes, HeNe lasers remain an important application for this element. There are very few things you can do with neon that don't rely in one way or another on the light it emits when stimulated with electricity. That neon has so few applications is masked by the fact that neon lights are so vivid and so widespread they make it seem like an important element, even though it would be one of the least missed.

The least reactive of all the elements, neon completely refuses to react with any others. That's something you definitely can't say about sodium, as we jump back to the left side of the periodic table.

Neon signs really are made with neon, like this Ne tube. An electric current runs through it, creating the light.

A tiny indicator light, no more than 1/8 inch across, glows from applied 120V AC.

Pure neon is an invisible gas, seen here in an antique sample ampoule.

Several thousand volts illuminate this neon sculpture in the shape of a Hilbert fractal.

Sodium

Sodium is the most explosive, and the best tasting, of all the alkali metals (the elements from the first column of the periodic table).

Explosive because if you throw it into water, it rapidly generates hydrogen gas, which seconds later ignites with a tremendous bang, throwing flaming sodium in all directions. (The other alkali metals react similarly with water, but sodium, overall, creates the most attractive explosions and is thus favored by mischief makers the world over for throwing into lakes and rivers.)

Best tasting because, together with chlorine (17), it forms sodium chloride, or table salt, widely considered the tastiest of the alkali metal chloride salts. Potassium chloride is sold as a salt substitute for people on a low-sodium diet, but it adds a bitter metallic note to its saltiness. Rubidium chloride and cesium chloride are less salty and more metallic in taste, while lithium chloride produces a burning sensation followed by an oily metallic aftertaste.

Pure sodium metal is used in large quantities in the chemical industry as a reducing agent, and while it might seem like a really bad idea, liquid sodium is used to move heat from the reactor core to the steam turbines in some nuclear reactors (yes, there have been spectacular sodium leaks). Closer to home, yellowish sodium vapor lamps create more light per unit of electricity than nearly any other type, while making people under them look dead.

Sodium is used only for its chemical properties. The next element, magnesium, is very useful both for its chemical and its structural properties.

These soft, silvery sodium chunks were cut with a knife and stored under oil. In air they turn white in seconds; exposed to water, they generate hydrogen gas and explode in flaming balls of molten sodium.

The mineral sodalite ($Na_4Al_3Si_3O_{12}Cl$).

Sodium hydroxide, whose traditional name is lye, is commonly sold as a drain opener.