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THE DISAPPEARING SPOON

And Other True Tales of Madness, Love, and the History of the World from the Periodic Table of the Elements

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THE
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Geography Is Destiny

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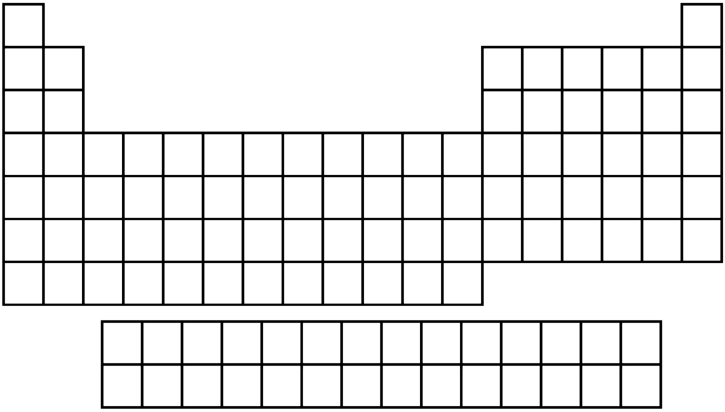
When most people think of the periodic table, they remember a chart hanging on the front wall of their high school chemistry class, an asymmetric expanse of columns and rows looming over one of the teacher's shoulders. The chart was usually enormous, six by four feet or so, a size both daunting and appropriate, given its importance to chemistry. It was introduced to the class in early September and was still relevant in late May, and it was the one piece of scientific information that, unlike lecture notes or textbooks, you were encouraged to consult during exams. Of course, part of the frustration you might remember about the periodic table could flow from the fact that, despite its being freely available to fall back on, a gigantic and fully sanctioned cheat sheet, it remained less than frickin' helpful.

On the one hand, the periodic table seemed organized and honed, almost German engineered for maximum scientific utility. On the other hand, it was such a jumble of long numbers, abbreviations, and what looked for all the world like computer error messages ($[Xe]6s^24f^15d^1$), it was hard not to feel anxious. And although the periodic table obviously had something to do with other sciences, such as biology and physics, it wasn't clear what exactly. Probably the biggest frustration for many students was that the people who *got* the periodic table, who could really unpack how it worked, could pull so many facts from it with such dweeby nonchalance. It was the same irritation colorblind people must feel when the fully sighted find sevens and nines lurking inside those parti-colored dot diagrams—crucial but hidden information that never quite resolves itself into coherence. People remember the table with a mix of fascination, fondness, inadequacy, and loathing.

Before introducing the periodic table, every teacher should strip away all the clutter and have students just stare at the thing, blank.



What does it look like? Sort of like a castle, with an uneven main wall, as if the royal masons hadn't quite finished building up the left-hand side, and tall, defensive turrets on both ends. It has eighteen jagged columns and seven horizontal rows, with a "landing strip" of two extra rows hanging below. The castle is made of "bricks," and the first non-obvious thing about it is that the bricks are not interchangeable. Each brick is an *element*, or type of substance (as of now, 112 elements, with a few more pending, make up the table), and the entire castle would crumble if any of those bricks didn't sit exactly where it does. That's no exaggeration: if scientists determined that one element somehow fit into a different slot or that two of the elements could be swapped, the entire edifice would tumble down.

Another architectural curiosity is that the castle is made up of different materials in different areas. That is, not all the bricks are made of the same substance, nor do they have the same characteristics. Seventy-five percent of the bricks are metals, which means most elements are cold, gray solids, at least at temperatures human beings are used to. A few columns on the eastern side contain gases. Only two elements, mercury and bromine, are liquids at room temperature. In between the metals and gases, about where Kentucky sits on a U.S. map, lie some hard-to-define elements, whose amorphous nature gives them interesting properties, such as the ability to make acids billions of times stronger than anything locked up in a chemical supply room. Overall, if each brick was made of the substance it represented, the castle of the elements would be a chimera with additions and wings from incongruent eras, or, more charitably, a Daniel Libeskind building, with seemingly incompatible materials grafted together into an elegant whole.

The reason for lingering over the blueprints of the castle walls is that the coordinates of an element determine nearly everything scientifically interesting about it. For each element, its geography is its destiny. In fact, now that you have a sense of what the table looks like in outline, I can switch to a more useful metaphor: the periodic table as a map. And to sketch in a bit more detail, I'm going to plot this map from east to west, lingering over both well-known and out-of-the-way elements.

First up, in column eighteen at the far right-hand side, is a set of elements known as the noble gases. *Noble* is an archaic, funny-sounding word, less chemistry than ethics or philosophy. And indeed, the term "noble gases" goes back to the birthplace of Western philosophy, ancient Greece. There, after his fellow Greeks Leucippus and Democritus invented the idea of atoms, Plato minted the word "elements" (in Greek, *stoicheia*) as a general term for different small particles of matter. Plato—who left Athens for his own safety after the death of his mentor, Socrates, around 400 BC and wandered around writing philosophy for years—of course lacked knowledge of what an element really is in chemistry terms. But if he had known, he no doubt would have selected the elements on the eastern edge of the table, especially helium, as his favorites.

In his dialogue on love and the erotic, *The Symposium*, Plato claimed that every being longs to find its complement, its missing half. When applied to people, this implies passion and sex and all the troubles that accompany passion and sex. In addition, Plato emphasized throughout his dialogues that abstract and unchanging things are intrinsically more noble than things that grub around and interact with gross matter. This explains why he adored geometry, with its idealized circles and cubes, objects perceptible only to our reason. For nonmathematical objects, Plato developed a theory of "forms," which argued that all objects are shadows of one ideal type. All trees, for instance, are imperfect copies of an ideal tree, whose perfect "tree-ness" they aspire to. The same with fish and "fish-ness" or even cups and "cup-ness." Plato believed that these forms were not merely theoretical but actually existed, even if they floated around in an empyrean realm beyond the direct perception of humans. He would have been as shocked as anyone, then, when scientists began conjuring up ideal forms on earth with helium.

In 1911, a Dutch-German scientist was cooling mercury with liquid helium when he discovered that below –452°F the system lost all electrical resistance and became an ideal conductor. This would be sort of like cooling an iPod down to hundreds of degrees below zero and finding that the battery remained fully charged no matter how long or loud you played music, until infinity, as long as the helium kept the circuitry cold. A Russian-Canadian team pulled an even neater trick in 1937 with pure helium. When cooled down to –456°F, helium turned into a superfluid, with exactly zero viscosity and zero resistance to flow—perfect fluidness. Superfluid helium defies gravity and flows uphill and over walls. At the time, these were flabbergasting finds. Scientists often fudge and pretend that effects like friction equal zero, but only to simplify calculations. Not even Plato predicted someone would actually find one of his ideal forms.

Helium is also the best example of "element-ness"—a substance that cannot be broken down or altered by normal, chemical means. It took scientists 2,200 years, from Greece in 400 BC to Europe in 1800 AD, to grasp what elements really are, because most are too changeable. It was hard to see what made carbon *carbon* when it appeared in thousands of compounds, all with different properties. Today we would say that carbon dioxide, for instance, isn't an element because one molecule of it divides into carbon and oxygen. But carbon and oxygen *are* elements because you cannot divide them more finely without destroying them. Returning to the theme of *The Symposium* and Plato's theory of erotic longing for a missing half, we find that virtually every element seeks out other atoms to form bonds with, bonds that mask its nature. Even most "pure" elements, such as oxygen molecules in the air (O_2), always appear as composites in nature. Yet scientists might have figured out what elements are much sooner had they known about helium, which has never reacted with another substance, has never been anything but a pure element.*

Helium acts this way for a reason. All atoms contain negative particles called electrons, which reside in different tiers, or energy levels, inside the atom. The levels are nested concentrically inside each other, and each level needs a certain number of electrons to fill itself and feel satisfied. In the innermost level, that number is two. In other levels, it's usually eight. Elements normally have equal numbers of negative electrons and positive particles called protons, so they're electrically neutral. Electrons, however, can be freely traded between atoms, and when atoms lose or gain electrons, they form charged atoms called ions.

What's important to know is that atoms fill their inner, lower-energy levels as full as possible with their own electrons, then either shed, share, or steal electrons to secure the right number in the outermost level. Some elements share or trade electrons diplomatically, while others act very, very nasty. That's half of chemistry in one sentence: atoms that don't have enough electrons in the outer level will fight, barter, beg, make and break alliances, or do whatever they must to get the right number.

Helium, element two, has exactly the number of electrons it needs to fill its only level. This "closed" configuration gives helium tremendous independence, because it doesn't need to interact with other atoms or share or steal electrons to feel satisfied. Helium has found its erotic complement in itself. What's more, that same configuration extends down the entire eighteenth column beneath helium—the gases neon, argon, krypton, xenon, and radon. All these elements have closed shells with full complements of electrons, so none of them reacts with anything under normal conditions. That's why, despite all the fervid activity to identify and label elements in the 1800s—including the development of the periodic table itself—no one isolated a single gas from column eighteen before 1895. That aloofness from everyday experience, so like his ideal spheres and triangles, would have charmed Plato. And it was that sense the scientists who discovered helium and its brethren on earth were trying to evoke with the name "noble gases." Or to put it in Plato-like words, "He who adores the perfect and unchangeable and scorns the corruptible and ignoble will prefer the noble gases, by far, to all other elements. For they never vary, never waver, never pander to other elements like hoi polloi offering cheap wares in the marketplace. They are incorruptible and ideal."

The repose of the noble gases is rare, however. One column to the west sits the most energetic and reactive gases on the periodic table, the halogens. And if you think of the table wrapping around like a Mercator map, so that east meets west and column eighteen meets column one, even more violent elements appear on the western edge, the alkali metals. The pacifist noble gases are a demilitarized zone surrounded by unstable neighbors.

Despite being normal metals in some ways, the alkalis, instead of rusting or corroding, can spontaneously combust in air or water. They also form an alliance of interests with the halogen gases. The halogens have seven electrons in the outer layer, one short of the octet they need, while the alkalis have one electron in the outer level and a full octet in the level below. So it's natural for the latter to dump their extra electron on the former and for the resulting positive and negative ions to form strong links.

This sort of linking happens all the time, and for this reason electrons are the most important part of an atom. They take up virtually all an atom's space, like clouds swirling around an atom's compact core, the nucleus. That's true even though the components of the nucleus, protons and neutrons, are far bigger than individual electrons. If an atom were blown up to the size of a sports stadium, the proton-rich nucleus would be a tennis ball at the fifty-yard line. Electrons would be pinheads flashing around it—but flying so fast and knocking into you so many times per second that you wouldn't be able to enter the stadium: they'd feel like a solid wall. As a result, whenever atoms touch, the buried nucleus is mute; only the electrons matter.*

One quick caveat: Don't get too attached to the image of electrons as discrete pinheads flashing about a solid core. Or, in the more usual metaphor, don't necessarily think of electrons as planets circling a nucleic sun. The planet analogy is useful, but as with any analogy, it's easy to take too far, as some renowned scientists have found out to their chagrin.

Bonding between ions explains why combinations of halogens and alkalis, such as sodium chloride (table salt), are common. Similarly, elements from columns with two extra electrons, such as calcium, and elements from columns that need two extra electrons, such as oxygen, frequently align themselves. It's the easiest way to meet everyone's needs. Elements from nonreciprocal columns also match up according to the same laws. Two ions of sodium (Na^+) take on one of oxygen (O^{-2}) to form sodium oxide, Na_2O. Calcium chloride combines as $CaCl_2$ for the same reasons. Overall, you can usually tell at a glance how elements will combine by noting their column numbers and figuring out their charges. The pattern all falls out of the table's pleasing left-right symmetry.

Unfortunately, not all of the periodic table is so clean and neat. But the raggedness of some elements actually makes them interesting places to visit.

* * *

There's an old joke about a lab assistant who bursts into a scientist's office one morning, hysterical with joy despite a night of uninterrupted work. The assistant holds up a fizzing, hissing, corked bottle of green liquid and exclaims he has discovered a universal solvent. His sanguine boss peers at the bottle and asks, "And what is a universal solvent?" The assistant sputters, "An acid that dissolves all substances!"

After considering this thrilling news — not only would this universal acid be a scientific miracle, it would make both men billionaires — the scientist replies, "How are you holding it in a glass bottle?"

It's a good punch line, and it's easy to imagine Gilbert Lewis smiling, perhaps poignantly. Electrons drive the periodic table, and no one did more than Lewis to elucidate how electrons behave and form bonds in atoms. His electron work was especially illuminating for acids and bases, so he would have appreciated the assistant's absurd claim. More personally, the punch line might have reminded Lewis how fickle scientific glory can be.

A wanderer, Lewis grew up in Nebraska, attended college and graduate school in Massachusetts around 1900, and then studied in Germany under chemist Walther Nernst. Life under Nernst proved so miserable, for legitimate and merely perceived reasons, that Lewis returned to Massachusetts for an academic post after a few months. That, too, proved unhappy, so he fled to the newly conquered Philippines to work for the U.S. government, taking with him only one book, Nernst's *Theoretical Chemistry*, so he could spend years rooting out and obsessively publishing papers on every quibbling error.*

Eventually, Lewis grew homesick and settled at the University of California at Berkeley, where, over forty years, he built

Berkeley's chemistry department into the world's best. Though that may sound like a happy ending, it wasn't. The singular fact about Lewis is that he was probably the best scientist never to win the Nobel Prize, and he knew it. No one ever received more nominations, but his naked ambition and a trail of disputes worldwide poisoned his chances of getting enough votes. He soon began resigning (or was forced to resign) from prestigious posts in protest and became a bitter hermit.

Apart from personal reasons, Lewis never secured the Nobel Prize because his work was broad rather than deep. He never discovered one amazing thing, something you could point to and say, Wow! Instead, he spent his life refining how an atom's electrons work in many contexts, especially the class of molecules known as acids and bases. In general, whenever atoms swap electrons to break or form new bonds, chemists say they've "reacted." Acid-base reactions offer a stark and often violent example of those swaps, and Lewis's work on acids and bases did as much as anyone's to show what exchanging electrons means on a submicroscopic level.

Before about 1890, scientists judged acids and bases by tasting or dunking their fingers in them, not exactly the safest or most reliable methods. Within a few decades, scientists realized that acids were in essence proton donors. Many acids contain hydrogen, a simple element that consists of one electron circling one proton (that's all hydrogen has for a nucleus). When an acid like hydrochloric acid (HCl) mixes with water, it fissures into H^+ and Cl^-. Removing the negative electron from the hydrogen leaves just a bare proton, the H^+, which swims away on its own. Weak acids like vinegar pop a few protons into solution, while strong acids like sulfuric acid flood solutions with them.

Lewis decided this definition of an acid limited scientists too much, since some substances act like acids without relying on hydrogen. So Lewis shifted the paradigm. Instead of saying that H^+ splits off, he emphasized that Cl^- absconds with its electron. Instead of a proton donor, then, an acid is an electron thief. In contrast, bases such as bleach or lye, which are the opposites of acids, might be called electron donors. These definitions, in addition to being more general, emphasize the behavior of electrons, which fits better with the electron-dependent chemistry of the periodic table.

Although Lewis laid this theory out in the 1920s and 1930s, scientists are still pushing the edge of how strong they can make acids using his ideas. Acid strength is measured by the pH scale, with lower numbers being stronger, and in 2005 a chemist from New Zealand invented a boron-based acid called a carborane, with a pH of –18. To put that in perspective, water has a pH of 7, and the concentrated HCl in our stomachs has a pH of 1. But according to the pH scale's unusual accounting methods, dropping one unit (e.g., from 4 to 3) boosts an acid's strength. So moving from stomach acid, at 1, to the boron-based acid, at –18, means the latter is ten billion billion times stronger. That's roughly the number of atoms it would take to stack them to the moon.

There are even worse acids based on antimony, an element with probably the most colorful history on the periodic table.* Nebuchadnezzar, the king who built the Hanging Gardens of Babylon in the sixth century BC, used a noxious antimony-lead mix to paint his palace walls yellow. Perhaps not coincidentally, he soon went mad, sleeping outdoors in fields and eating grass like an ox. Around that same time, Egyptian women were applying a different form of antimony as mascara, both to decorate their faces and to give themselves witchlike powers to cast the evil eye on enemies. Later, medieval monks—not to mention Isaac Newton—grew obsessed with the sexual properties of antimony and decided this half metal, half insulator, neither one thing nor the other, was a hermaphrodite. Antimony pills also won fame as laxatives. Unlike modern pills, these hard antimony pills didn't dissolve in the intestines, and the pills were considered so valuable that people rooted through fecal matter to retrieve and reuse them. Some lucky families even passed down laxatives from father to son. Perhaps for this reason, antimony found heavy work as a medicine, although it's actually toxic. Mozart probably died from taking too much to combat a severe fever.

Scientists eventually got a better handle on antimony. By the 1970s, they realized that its ability to hoard electron-greedy elements around itself made it wonderful for building custom acids. The results were as astounding as the helium superfluids. Mixing antimony pentafluoride, SbF_5, with hydrofluoric acid, HF, produces a substance with a pH of –31. This superacid is 100,000 billion billion billion times more potent than stomach acid and will eat through glass, as ruthlessly as water through paper. You couldn't pick up a bottle of it because after it ate through the bottle, it would dissolve your hand. To answer the professor in the joke, it's stored in special Teflon-lined containers.

To be honest, though, calling the antimony mix the world's strongest acid is kind of cheating. By themselves, SbF_5 (an electron thief) and HF (a proton donor) are nasty enough. But you have to sort of multiply their complementary powers together, by mixing them, before they attain superacid status. They're strongest only under contrived circumstances. Really, the strongest solo acid is still the boron-based carborane ($HCB_{11}Cl_{11}$). And this boron acid has the best punch line so far: It's simultaneously the world's strongest *and gentlest* acid. To wrap your head around that, remember that acids split into positive and negative parts. In carborane's case, you get H^+ and an elaborate cagelike structure formed by everything else ($CB_{11}Cl_{11}^-$). With most acids it's the negative portion that's corrosive and caustic and eats through skin. But the boron cage forms one of the most stable molecules ever invented. Its boron atoms share electrons so generously that it practically becomes helium, and it won't go around ripping electrons from other atoms, the usual cause of acidic carnage.

So what's carborane good for, if not dissolving glass bottles or eating through bank vaults? It can add an octane kick to gasoline, for one thing, and help make vitamins digestible. More important is its use in chemical "cradling." Many chemical reactions involving protons aren't clean, quick swaps. They require multiple steps, and protons get shuttled around in millionths of billionths of seconds—so quickly scientists have no idea what really happened. Carborane, though, because it's so stable and unreactive, will flood a solution with protons, then freeze the molecules at crucial intermediate points. Carborane holds the intermediate species up on a soft, safe pillow. In contrast, antimony superacids make terrible cradles, because they shred the molecules scientists most want to look at. Lewis would have enjoyed seeing this and other applications of his work with electrons and acids, and it might have brightened the last dark years of his life. Although he did government work during World War I and made valuable contributions to chemistry until he was in his sixties, he was passed over for the Manhattan Project during World War II. This galled him, since many chemists he had recruited to Berkeley played important roles in building the first atomic bomb and became national heroes. In contrast, he puttered around during the war, reminiscing and writing a wistful pulp novel about a soldier. He died alone in his lab in 1946.

There's general agreement that after smoking twenty-some cigars per day for forty-plus years, Lewis died of a heart attack. But it was hard not to notice that his lab smelled like bitter almonds—a sign of cyanide gas—the afternoon he died. Lewis used cyanide in his research, and it's possible he dropped a canister of it after going into cardiac arrest. Then again, Lewis had had lunch earlier in the day—a lunch he'd initially refused to attend—with a younger, more charismatic rival chemist who had won the Nobel Prize and served as a special consultant to the Manhattan Project. It's always been in the back of some people's minds that the honored colleague might have unhinged Lewis. If that's true, his facility with chemistry might have been both convenient and unfortunate.

In addition to reactive metals on its west coast and halogens and noble gases up and down its east coast, the periodic table contains a "great plains" that stretches right across its middle—columns three through twelve, the transition metals. To be honest, the transition metals have exasperating chemistry, so it's hard to say anything about them generally—except be careful. You see, heavier atoms like the transition metals have more flexibility than other atoms in how they store their electrons. Like other atoms, they have different energy levels (designated one, two, three, etc.), with lower energy levels buried beneath higher levels. And they also fight other atoms to secure full outer energy levels with eight electrons. Figuring out what counts as the outer level, however, is trickier.

As we move horizontally across the periodic table, each element has one more electron than its neighbor to the left. Sodium, element eleven, normally has eleven electrons; magnesium, element twelve, has twelve electrons; and so on. As elements swell in size, they not only sort electrons into energy levels, they also store those electrons in different-shaped bunks, called shells. But atoms, being unimaginative and conformist, fill shells and energy levels in the same order as we move across the table. Elements on the far left-hand side of the table put the first electron in an s-shell, which is spherical. It's small and holds only two electrons—which explains the two taller columns on the left side. After those first two electrons, atoms look for something roomier. Jumping across the gap, elements in the columns on the right-hand side begin to pack new electrons one by one into a p-shell, which looks like a misshapen lung. P-shells can hold six electrons, hence the six taller columns on the right. Notice that across each row near the top, the two s-shell electrons plus the six p-shell electrons add up to eight electrons total, the number most atoms want in the outer shell. And except for the self-satisfied noble gases, all these elements' outer-shell electrons are available to dump onto or react with other atoms. These elements behave in a logical manner: add a new electron, and the atom's behavior should change, since it has more electrons available to participate in reactions.

Now for the frustrating part. The transition metals appear in columns three through twelve of the fourth through seventh rows, and they start to file electrons into what are called d-shells, which hold ten electrons. (D-shells look like nothing so much as misshapen balloon animals.) Based on what every other previous element has done with its shells, you'd expect the transition metals to put each extra d-shell electron on display in an outer layer and for that extra electron to be available for reactions, too. But no, transition metals squirrel their extra electrons away and prefer to hide them beneath other layers. The decision of the transition metals to violate convention and bury their d-shell electrons seems ungainly and counterintuitive — Plato would not have liked it. It's also how nature works, and there's not much we can do about it.

There's a payoff to understanding this process. Normally as we move horizontally across the table, the addition of one electron to each transition metal would alter its behavior, as happens with elements in other parts of the table. But because the metals bury their d-shell electrons in the equivalent of false-bottomed drawers, those electrons end up shielded. Other atoms trying to react with the metals cannot get at those electrons, and the upshot is that many metals in a row leave the same number of electrons exposed. They therefore act the same way chemically. That's why, scientifically, many metals look so indistinguishable and act so indistinguishably. They're all cold, gray lumps because their outer electrons leave them no choice but to conform. (Of course, just to confuse things, sometimes buried electrons do rise up and react. That's what causes the slight differences between some metals. That's also why their chemistry is so exasperating.)

F-shell elements are similarly messy. F-shells begin to appear in the first of the two free-floating rows of metals beneath the periodic table, a group called the lanthanides. (They're also called the rare earths, and according to their atomic numbers, fifty-seven through seventy-one, they really belong in the sixth row. They were relegated to the bottom to make the table skinnier and less unwieldy.) The lanthanides bury new electrons even more deeply than the transition metals, often two energy levels down. This means they are even more alike than the transition metals and can barely be distinguished from one another. Moving along the row is like driving from Nebraska to South Dakota and not realizing you've crossed the state line.

It's impossible to find a pure sample of a lanthanide in nature, since its brothers always contaminate it. In one famous case, a chemist in New Hampshire tried to isolate thulium, element sixty-nine. He started with huge casserole dishes of thulium-rich ore and repeatedly treated the ore with chemicals and boiled it, a process that purified the thulium by a small fraction each time. The dissolving took so long that he could do only one or two cycles per day at first. Yet he repeated this tedious process fifteen thousand times, by hand, and winnowed the hundreds of pounds of ore down to just ounces before the purity satisfied him. Even then, there was still a little cross-contamination from other lanthanides, whose electrons were buried so deep, there just wasn't enough of a chemical handle to grasp them and pull them out.

Electron behavior drives the periodic table. But to really understand the elements, you can't ignore the part that makes up more than 99 percent of their mass—the nucleus. And whereas electrons obey the laws of the greatest scientist never to win the Nobel Prize, the nucleus obeys the dictates of probably the most unlikely Nobel laureate ever, a woman whose career was even more nomadic than Lewis's.

Maria Goeppert was born in Germany in 1906. Even though her father was a sixth-generation professor, Maria had trouble convincing a Ph.D. program to admit a woman, so she bounced from school to school, taking lectures wherever she could. She finally earned her doctorate at the University of Hannover, defending her thesis in front of professors she'd never met. Not surprisingly, with no recommendations or connections, no university would hire her upon her graduation. She could enter science only obliquely, through her husband, Joseph Mayer, an American chemistry professor visiting Germany. She returned to Baltimore with him in 1930, and the newly named Goeppert-Mayer began tagging along with

Mayer to work and conferences. Unfortunately, Mayer lost his job several times during the Great Depression, and the family drifted to universities in New York and then Chicago.

Most schools tolerated Goeppert-Mayer's hanging around to chat science. Some even condescended to give her work, though they refused to pay her, and the topics were stereotypically "feminine," such as figuring out what causes colors. After the Depression lifted, hundreds of her intellectual peers gathered for the Manhattan Project, perhaps the most vitalizing exchange of scientific ideas ever. Goeppert-Mayer received an invitation to participate, but peripherally, on a useless side project to separate uranium with flashing lights. No doubt she chafed in private, but she craved science enough to continue to work under such conditions. After World War II, the University of Chicago finally took her seriously enough to make her a professor of physics. Although she got her own office, the department still didn't pay her.

Nevertheless, bolstered by the appointment, she began work in 1948 on the nucleus, the core and essence of an atom. Inside the nucleus, the number of positive protons — the atomic number — determines the atom's identity. In other words, an atom cannot gain or lose protons without becoming a different element. Atoms do not normally lose neutrons either, but an element's atoms can have different numbers of neutrons — variations called isotopes. For instance, the isotopes lead-204 and lead-206 have identical atomic numbers (82) but different numbers of neutrons (122 and 124). The atomic number plus the number of neutrons is called the atomic weight. It took scientists many years to figure out the relationship between atomic number and atomic weight, but once they did, periodic table science got a lot clearer.

Goeppert-Mayer knew all this, of course, but her work touched on a mystery that was more difficult to grasp, a deceptively simple problem. The simplest element in the universe, hydrogen, is also the most abundant. The second-simplest element, helium, is the second most abundant. In an aesthetically tidy universe, the third element, lithium, would be the third most abundant, and so on. Our universe isn't tidy. The third most common element is oxygen, element eight. But why? Scientists might answer that oxygen has a very stable nucleus, so it doesn't disintegrate, or "decay." But that only pushed the question back—why do certain elements like oxygen have such stable nuclei?

Unlike most of her contemporaries, Goeppert-Mayer saw a parallel here to the incredible stability of noble gases. She suggested that protons and neutrons in the nucleus sit in shells just like electrons and that filling nuclear shells leads to stability. To an outsider, this seems reasonable, a nice analogy. But Nobel Prizes aren't won on conjectures, especially those by unpaid female professors. What's more, this idea ruffled nuclear scientists, since chemical and nuclear processes are independent. There's no reason why dependable, stay-at-home neutrons and protons should behave like tiny, capricious electrons, which abandon their homes for attractive neighbors. And mostly they don't.

Except Goeppert-Mayer pursued her hunch, and by piecing together a number of unlinked experiments, she proved that nuclei do have shells and do form what she called magic nuclei. For complex mathematical reasons, magic nuclei don't reappear periodically like elemental properties. The magic happens at atomic numbers two, eight, twenty, twenty-eight, fifty, eighty-two, and so on. Goeppert-Mayer's work proved how, at those numbers, protons and neutrons marshal themselves into highly stable, highly symmetrical spheres. Notice too that oxygen's eight protons and eight neutrons make it doubly magic and therefore eternally stable—which explains its seeming overabundance. This model also explains at a stroke why elements such as calcium (twenty) are disproportionately plentiful and, not incidentally, why our bodies employ these readily available minerals.

Goeppert-Mayer's theory echoes Plato's notion that beautiful shapes are more perfect, and her model of magic, orb-shaped nuclei became the ideal form against which all nuclei are judged. Conversely, elements stranded far between two magic numbers are less abundant because they form ugly, oblong nuclei. Scientists have even discovered neutron-starved forms of holmium (element sixty-seven) that give birth to a deformed, wobbly "football nucleus." As you might guess from Goeppert-Mayer's model (or from ever having watched somebody fumble during a football game), the holmium footballs aren't very steady. And unlike atoms with misbalanced electron shells, atoms with distorted nuclei can't poach neutrons and protons from other atoms to balance themselves. So atoms with misshapen nuclei, like that form of holmium, hardly ever form and immediately disintegrate if they do.

The nuclear shell model is brilliant physics. That's why it no doubt dismayed Goeppert-Mayer, given her precarious status among scientists, to discover that it had been duplicated by male physicists in her homeland. She risked losing credit for everything. However, both sides had produced the idea independently, and when the Germans graciously acknowledged her work and asked her to collaborate, Goeppert-Mayer's career took off. She won her own accolades, and she and her husband moved a final time in 1959, to San Diego, where she began a real, paying job at the new University of California campus there. Still, she never quite shook the stigma of being a dilettante. When the Swedish Academy announced in 1963 that she had won her profession's highest honor, the San Diego newspaper greeted her big day with the headline "S.D. Mother Wins Nobel Prize."

But maybe it's all a matter of perspective. Newspapers could have run a similarly demeaning headline about Gilbert Lewis, and he probably would have been thrilled.

Reading the periodic table across each row reveals a lot about the elements, but that's only part of the story, and not even the best part. Elements in the same column, latitudinal neighbors, are actually far more intimately related than horizontal neighbors. People are used to reading from left to right (or right to left) in virtually every human language, but reading the periodic table up and down, column by column, as in some forms of Japanese, is actually more significant. Doing so reveals a rich subtext of relationships among elements, including unexpected rivalries and antagonisms. The periodic table has its own grammar, and reading between its lines reveals whole new stories.

SHORTLIST

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2011

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- *The Disappearing Spoon* by Sam Kean (Little, Brown and Company (USA), Doubleday (UK))
- *The Wavewatcher's Companion* by Gavin Pretor-Pinney (Bloomsbury)
- *Massive: The Missing Particle That Sparked the Greatest Hunt in Science* by Ian Sample (Basic Books)
- *The Rough Guide to The Future* by Jon Turney (Rough Guides)

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