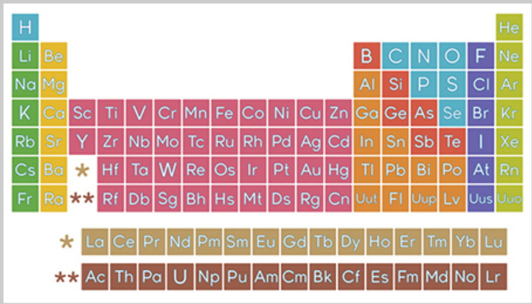
CHEMISTRY
CHALLENGES AND SOLUTIONS

Unit 4: *Organizing Atoms and Electrons—The Periodic Table*



Unit Overview

As scientists discovered more and more chemical elements, they began developing systems to organize the elements by their chemical properties, leading to the modern periodic table. Through its organization, the periodic table makes clear the underlying chemical and physical trends among the elements. These characteristics—reactivity, atomic radius, electronegativity, and density—are linked to the distribution of electrons around the nucleus. The periodic table—undoubtedly the most important and useful document in chemistry—is being continually updated even today as scientists strive to create new manmade elements in laboratories.

by Louisa Morrison

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Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 1: Introduction

The periodic table is a blueprint for all the known elements in the universe. Everything that makes up the food we eat, the air we breathe, and the ground we stand on can be found within the seven rows and 18 columns of the periodic table. The periodic table is widely used and scientifically approved, and no chemist can do work without one. Impressively, even with the major shift in thinking following the development of quantum mechanics, the general structure of the periodic table has not changed much since 1869. (Figure 4-1) It allows those with the appropriate knowledge to unlock the secrets of the elements, including their average mass, relative reactivity, and internal organization. How can this one document tell us so much?

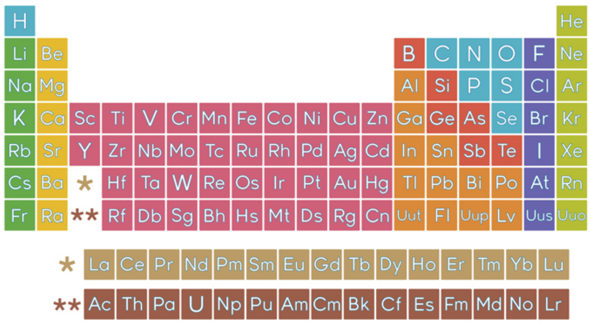

Figure 4-1. The Periodic Table

As of 2012, the periodic table contains 118 confirmed chemical elements. Of these elements, 114 have been officially recognized and named by the International Union of Pure and Applied Chemistry (IUPAC). The elements that start with "Uu" are not officially recognized.

The table is organized by atomic number, which is the number of protons in the nucleus. We can organize the periodic table this way because all atoms of a specific element have the same number of protons. In fact, if the number of protons were to change, the atom's identity would change. In addition to the atomic number, there are many patterns within the periodic table. In this unit, we will discuss some of those patterns, including electron configuration, size, and reactivity. We also will examine how these patterns allow us to predict the formation of compounds.

The Last Natural Element

The last stable natural element, rhenium, was discovered in 1925 in Berlin by Ida Tacke (1896–1978) and Walter Noddack (1893–1960).

Molybdenite. © Wikimedia Commons, CC License 3.0. Author: John Chapman, 22 April 2008.

It was known at the time that element number 75 should exist, but it had not yet been found. Tacke and Noddack found rhenium as a contaminant in a sample of a platinum ore and in several other minerals such as columbite and molybdenite. From almost three-fourths of a ton of molybdenite, they were able to isolate a gram of pure rhenium. Rhenium is the least abundant stable element on Earth, and thus no pure samples can be found in nature. One common use of rhenium is as filaments in flash bulbs for photography. It is also a very potent catalyst for many modern industrial processes. Since rhenium's discovery, no new stable elements have been discovered in nature.

In addition to chemical properties, the periodic table reveals the cultural history of each element. Some elements have countries in their names (francium, polonium, and germanium), while others bear the names of notable scientists (einsteinium, nobelium, and curium). The discoverers of these elements named them in honor of home countries or to commemorate other influential scientists. The periodic table contains a wealth of information, once we know how to interpret it.

The naming of elements turned out to be one of the last battlegrounds of the Cold War. In the 1960s, research groups in the Soviet Union and the United States simultaneously claimed discovery of elements 104 and 105. The International Union of Pure and Applied Chemistry, or IUPAC, the global organization charged with establishing standard practices for chemists, is responsible for handing out naming rights to labs in honor of discovering the element. Only after the end of the Cold War, in 1997, did IUPAC settle the dispute between the Americans and Soviets. Americans won the rights to name element 104 after Ernest Rutherford (rutherfordium), and the Russians won the rights to name element 105 (dubnium) after Dubna, the town in Russia where the element was discovered.

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Section 2: History of the Periodic Table

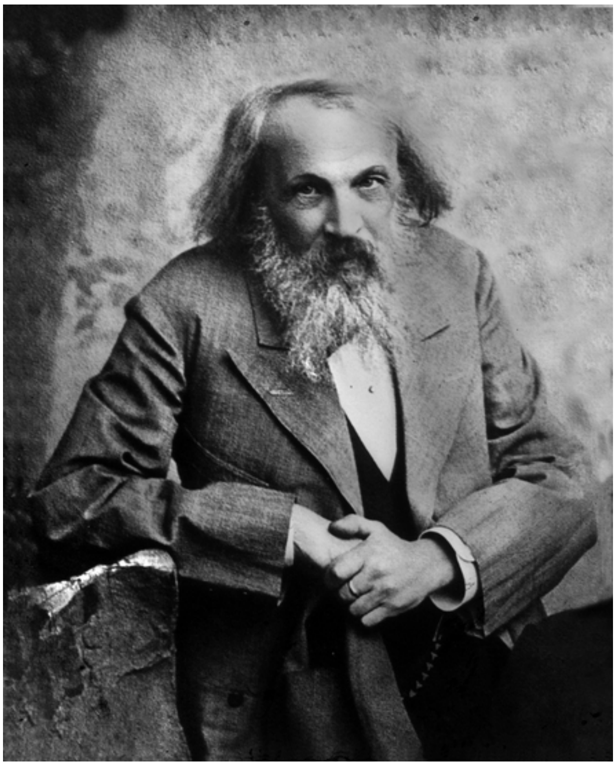
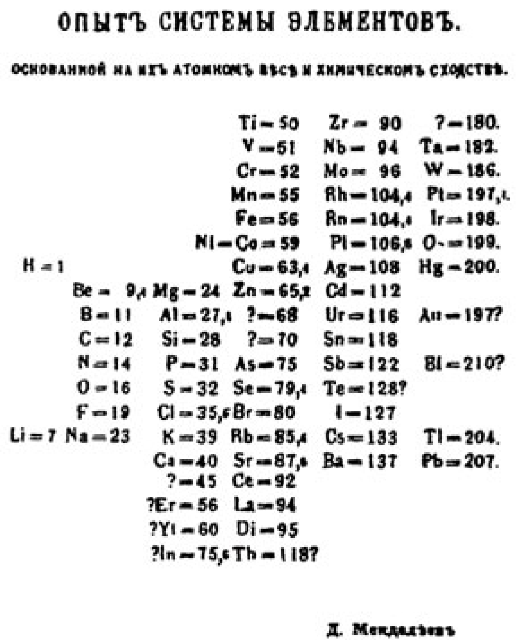


Figure 4-2. Dmitri Mendeleev Created the Periodic Table That We Still Use Today

Born in Siberia, Russia, Mendeleev was the youngest of 14 children. Mendeleev's parents owned a glass factory. After his father died, Mendeleev and his family moved to St. Petersburg. There, Mendeleev became a chemistry professor and went on to create the periodic table.

In the early 19th century, a scientist named Johann Döbereiner (1780–1849) noticed that strontium's atomic mass fell exactly between the atomic masses of calcium and barium. Investigating this further, Döbereiner found that other elements followed this pattern, and so he began to group elements into "triads." Other chemists built on Döbereiner's work to observe that elements with similar atomic masses also exhibited similar properties. By the mid-19th century, chemists across the world rushed to organize the elements into a chart that would help them make sense of what they were observing in their experiments. Out of the early crop of periodic tables, Russian chemistry professor Dmitri Mendeleev's table emerged as the defining document of the elements. (Figure 4-2)

The most remarkable aspect of Mendeleev's periodic table, and the reason why his name has endured, is that Mendeleev left spaces open for elements that were not yet discovered. Later discoveries proved that his predictions were correct. Elements, including germanium and boron, fit neatly into Mendeleev's table. It is also important to note that not only did Mendeleev predict the existence of these missing elements, but he also properly predicted their physical and chemical properties. Perhaps most remarkable of all, Mendeleev did all this without knowing the structure of the atom.



ОПЫТЪ СИСТЕМЫ ЭЛЕМЕНТОВЪ.

ОСНОВАННОЙ НА ИХЪ АТОМНОМЪ ВѢСѢ И ХИМИЧЕСКОМЪ СХОДСТВѢ.

			Ti = 50	Zr = 90	? = 180.
			V = 51	Nb = 94	Ta = 182.
			Cr = 52	Mo = 96	W = 186.
			Mn = 55	Rh = 104,4	Pt = 197,4
			Fe = 56	Rn = 104,4	Ir = 198.
			Ni = Co = 59	Pl = 106,6	Os = 199.
H = 1			Cu = 63,4	Ag = 108	Hg = 200.
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112	
	B = 11	Al = 27,4	? = 68	Ur = 116	Au = 197?
	C = 12	Si = 28	? = 70	Sn = 118	
	N = 14	P = 31	As = 75	Sb = 122	Bi = 210?
	O = 16	S = 32	Se = 79,4	Te = 128?	
	F = 19	Cl = 35,5	Br = 80	I = 127	
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133	Tl = 204.
		Ca = 40	Sr = 87,6	Ba = 137	Pb = 207.
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75,6	Th = 118?		

Д. Менделѣевъ

Figure 4-3. Mendeleev's Table

Dmitri Mendeleev organized the elements by atomic weight for his table. He wrote all the elements on cards, moving them around. This table was originally published in 1871 in his textbook, *The Principles of Chemistry*. Mendeleev called his table "the periodic table" because of the patterns he found throughout his layout.

Because the proton, the neutron, and the electron had not yet been discovered, Mendeleev initially sorted the elements by atomic mass. One of the main patterns Mendeleev observed was the repeating pattern of reactivity with oxygen. In groups of eight, he observed that as mass increased, so did the number of oxygen atoms that would react with a given element. While Mendeleev's table was mostly successful, a few elements did not seem to make sense in the positions he assigned. (Figure 4-3)

It turns out that some of the imperfections in the Mendeleev periodic table stemmed from using atomic mass as his way of ordering the elements. It would take until 1913, when British physicist Henry Moseley (1887–1915) began experiments exploring the newly discovered nucleus of the atom. In his experiment, Moseley shot a stream of electrons at atoms of different elements. Each element gave off a different X-ray signature, and this X-ray signature was related to its position on the periodic table. From his data, he was able to determine a whole number, *Z*, for each different element; Moseley called this the "atomic number." Based on his findings, Moseley proposed that the elements should be ordered based on atomic number, and that elements with the atomic numbers 43, 61, 72, and 75 had yet to be discovered. We now know that the atomic number is the number of protons, which were discovered shortly after Moseley's death by his mentor, Ernest Rutherford (1871–1937). Henry Moseley's work provided amazing insight into the periodic table. Unfortunately, Moseley was killed at age 27 in the Battle of Gallipoli in World War I, before he was able to reap the rewards of his discovery.

The last major change to the periodic table was in the 1940s, when Glenn T. Seaborg (1912–1999), a chemist at the University of California, Berkeley, moved heavy elements to the bottom of the table. On the periodic table we use today, there are two rows and 14 columns separate from the rest of the elements. Seaborg recognized that these two series, called the "lanthanide series" (also known as the "rare Earth elements") and the "actinide series," belonged next to the transition metal block of the periodic table. Because this would make the periodic table too long to fit on a sheet of paper, Seaborg moved these blocks to the bottom of the table. Seaborg's colleague, Darleane Hoffman (born 1926), tells this story in the accompanying video.

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Section 3: A Tour of the Periodic Table

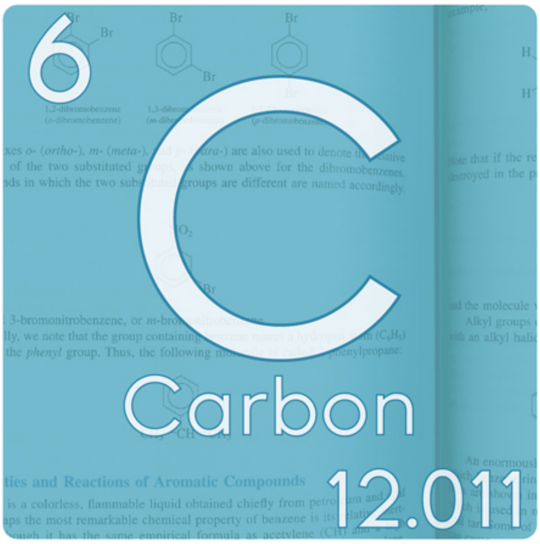

Figure 4-4. Representative Box from the Periodic Table

There are usually four main components in each box on the periodic table: atomic number, average atomic mass, element symbol, and element name. The box for carbon is shown here. In this example, the atomic number, 6, is in the top left corner. The average atomic mass, 12.011, is in the bottom right corner. The element symbol, C, is in the center, and the element name, carbon, is directly below the symbol. There are slight variations in the exact location of each component depending on the version of the table. A quick way to tell the atomic number and average atomic mass apart is that the atomic number is always a whole number, and the average atomic mass is usually written with at least a few decimal places of precision, unless it is an unstable radioactive element.

The periodic table is made up of boxes for each unique element arranged in 18 columns, called "groups" or "families," and seven rows, called "periods." As of 2012, the official periodic table released by IUPAC contains 114 total named elements with flerovium and livermorium as the most recent additions. Every bit of matter in the universe is made of the elements on the periodic table. The naturally occurring elements are all found in the first 92 boxes. Amazingly, by the mid-20th century, scientists were able to produce new elements in the laboratory by combining two naturally occurring elements in a nuclear reactor. Unit 4 video illustrates how scientists at Lawrence Livermore National Laboratory in California combine the nuclei of two naturally occurring elements to produce larger and larger elements. In theory, the table is limitless and only depends on how well scientists are able to create new elements in the laboratory.

Each box on the periodic table tells us important information about the element it represents. Nearly all periodic tables will have these four core pieces of information in each box: the name of the element, the atomic number of the element, the one- or two-letter symbol for the element, and the atomic mass of the element. (Figure 4-4) Other periodic tables may try to cram even more information about their properties into the box. This can serve two purposes: one is to have an easy location for collecting information about the elements; the other is to visualize the patterns of the elements on the periodic table. However, different periodic table designers will put that information in different locations inside the box.

Metals

More than two-thirds of elements are metals, including sodium, lead, uranium, iron, and zinc. They tend to be solid at room temperature, have shiny or reflective surfaces, and are excellent conductors of heat and electricity. Metals are also highly **malleable**, which means that we can hammer them into new shapes; this is how gold can be hammered out into thin sheets called "gold leaf." Other properties of metals come over large ranges; for example, metals can have very different densities, which is the amount of mass they have per unit volume. Sodium has such a low density that it floats in water, whereas the metal lead is one of the densest of all the elements; it quickly sinks in water. A small piece of lead would be extremely heavy to hold. (Figure 4-5)

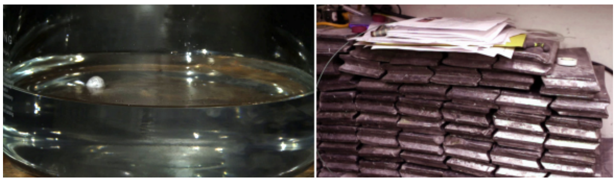

Figure 4-5. Metals Can Have a Wide Range of Densities

On the left, a small piece of sodium floats in water. On the right, one bar of lead weighs 42 pounds.

Non-Metals

Non-metals vary greatly in their physical properties. For example, some non-metals are gases at room temperature, some are solids, and one, bromine, is a liquid. Non-metals that are solid tend to be brittle, dull, and poor conductors of heat and electricity. Some gaseous non-metals, including hydrogen, oxygen, and nitrogen, occur as **diatomic elements**, which means they are found in nature in pairs: H_2, O_2, and N_2. This contrasts with the noble gases like argon and neon, which are monatomic and exist just as Ar or Ne atoms in nature. All non-metals range greatly in color, from reddish bromine to yellow sulfur. The non-metallic elements are essential to life on Earth: Carbon, hydrogen, oxygen, and nitrogen comprise most organic molecules, often with smaller amounts of phosphorus and sulfur.

Semi-Metals or Metalloids

Semi-metals or metalloids are the elements that comprise the boundary between metals and non-metals, so they include the typical properties of metals and non-metals. When we see semi-metals like silicon or antimony, they look almost metallic, but are brittle. Semi-metals have low conductivity at cooler temperatures, but high conductivity at warmer temperatures. Because of this, the most common semi-metal, silicon, is used in electronic devices, as it strictly controls the flow of electricity.

The Periodic Table as Art

The periodic table has many possible shapes and arrangements. Rebecca Kamen, a professor of art at Northern Virginia Community College, produced a 3D installation of the periodic table entitled, *Divining Nature: An Elemental Garden*. Each tower or atomic flower form represents an element, with the petals as the energy levels of the electrons.

Rebecca Kamen's artwork that represents the periodic table as a garden. © Angie Seckinger.

Kamen found inspiration for her work by looking at rare manuscripts provided by the American Philosophical Society, the Chemical Heritage Foundation, and the National Library of Medicine. She has spoken at research institutions, including the Marine Biological Laboratory, National Institutes of Health, Brown University, and Harvard University, about the intersection between art, philosophy, and science. Her work demonstrates the limitless possibilities of the arrangement of elements and the beauty that can be found in science.

Let's consider three common elements: carbon (a non-metal), silicon (a semi-metal), and copper (a metal). On any given day, we probably see each of these in their elemental forms, or at least we're very close to them. Pure elemental carbon comes in two major forms in nature: diamond and graphite. Diamonds are one of the hardest materials in nature; we can see them in jewelry and drill bits. Graphite, on the other hand, is the technical name for pencil lead. In addition to being a common writing tool, it is an industrial lubricant. Pure silicon is around us all the time, as it is found in the chips inside electronics, including computers and cell phones. Lastly, copper metal, with its reddish brown color, is also found in many electronics, as it is the most common material in wires, because it conducts electricity so well. We also see it on the outer coating of pennies.

Glossary

Diatomic elements

The elements that only are found in nature as pairs covalently bound together, for example oxygen (O_2) and hydrogen (H_2).

Malleable

Describes matter that with increased heat or pressure can have its shape changed, as if it had been hit by a hammer.

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Section 4: Atomic Mass, Atomic Number, and Carbon-12

Athletes, Artificial Steroids, and Carbon-14

Different isotopes of carbon have different properties, which is very useful for studying certain biological phenomena. Unlike carbon-12 and carbon-13, which are stable isotopes, carbon-14 is radioactive. This means that its nucleus is unstable and it eventually decays into nitrogen-14. (Radioactive decay will be covered in detail in Unit 12). This property allows scientists to determine if athletes are taking artificial steroids.

The human body naturally produces steroids, such as testosterone. All living things contain carbon-14, so these naturally produced steroids will contain carbon-14. Artificial steroids are made from petrochemicals, which come from oil and coal mining. Oil and coal have been in the earth for millions if not billions of years, so all the carbon-14 that was once in the oil and coal has since decayed. Thus, synthetic steroids will not have any carbon-14, only carbon-12 and carbon-13.

After collecting blood and isolating the testosterone, mass spectrometry can show if carbon-14 is not present, thereby proving the use of synthetic steroids. However, what happens if an athlete takes testosterone from another animal rather than synthetic testosterone? Unfortunately, while that is also frowned upon by sporting organizations, the carbon-14 atoms will still be present, and this test will not help discover the banned behavior.

How do we weigh something as tiny as an atom? All of our standard weight measurements, like grams or pounds, are far too heavy for atomic measurements. To solve this problem, scientists picked an atom of carbon-12 as the standard against which all other atoms could be weighed. This standard assumes that an atom of carbon-12 has a mass of exactly 12 atomic mass units (u). This is a 12 followed by an infinite number of zeroes. Following this standard, one atomic mass unit is 1/12 the mass of a carbon-12 atom. What does it mean to be carbon-12? This is a carbon atom with a mass number of 12, meaning it has 6 protons and 6 neutrons, or 12 heavy particles.

However, if we look at the periodic table, carbon's average atomic mass is 12.011, slightly over 12. This is because the periodic table usually represents average atomic mass. As previously mentioned, the number of protons in an atom is the key to an element's identity. The number of neutrons can change without changing the element's identity. The different forms of an element based on the number of neutrons in the nucleus are called "**isotopes**." Carbon-12 is only one isotope of the element carbon, with 6 protons and 6 neutrons. There are also carbon-13, with 6 protons and 7 neutrons, and carbon-14, with 6 protons and 8 neutrons. Because these isotopes have more neutrons than carbon-12, they are heavier.

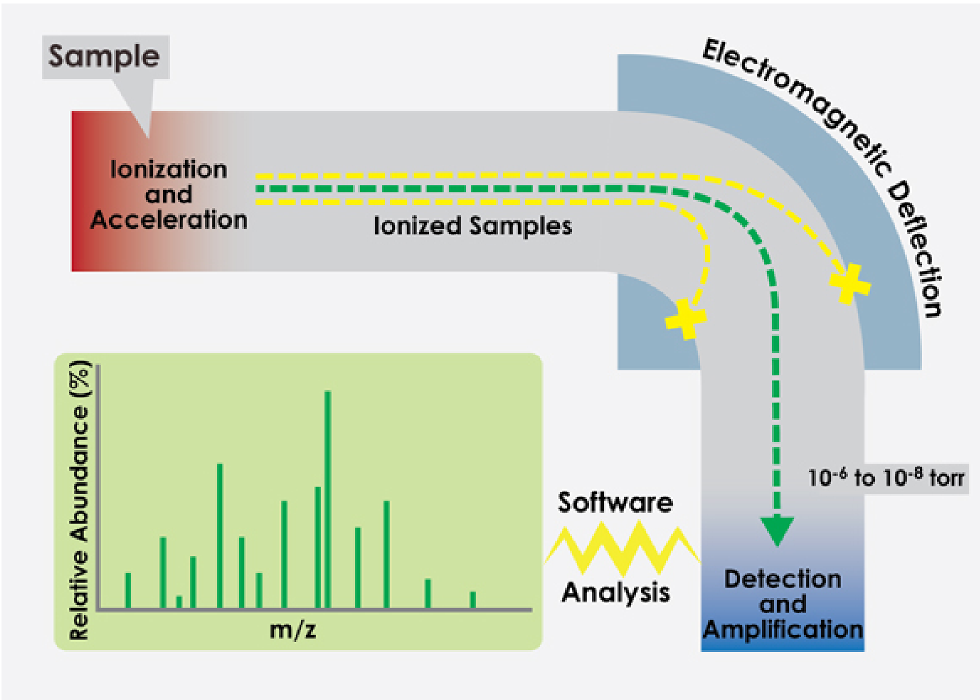
Table 4-1. ***The Relative Abundance of Carbon Isotopes***

Isotopes	Protons	Neutrons	Natural Abundance
Carbon-12	6	6	98.9%
Carbon-13	6	7	1.1%
Carbon-14	6	8	Trace

Mass Spectrometry

Mass spectrometry, or "mass spec" for short, is a useful tool to identify molecules based on their mass-to-charge ratio. A sample in a laboratory or from a crime scene is vaporized and ionized. The sample is then passed through magnetic plates on its way to a detector. The magnetic plates separate based on size because smaller particles of equivalent charge will pass through the field created by the plates more quickly. The data appears as a bar graph, with the height of the peak calibrated to the abundance of the sample. Different molecules create unique patterns of lines, allowing scientists to distinguish between compounds.

The spectrum in the bottom left corner of the image is an example of the type of data produced by mass spec. Each peak represents the abundance of a particular fragment at a specific mass-to-charge ratio. Compounds make distinctive patterns of peaks.



Schematic of Mass Spectrometry. © Science Media Group.

Average atomic mass is calculated based on the natural abundance of each isotope of the element. It is a weighted average. For example, if 99.6% of all nitrogen atoms weigh 14.0031 atomic mass units (u) and 0.4% weigh 15.0001 u, then the average atomic mass is (0.996)(14.0031) + (0.004)(15.0001) = 14.007 u, which is the value listed on the periodic table. The higher-numbered elements have average atomic mass numbers in parentheses because there is only one isotope that has been measured.

Glossary

Isotopes

Forms of an element that have the same number of protons, but different numbers of neutrons.

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Section 5: The Orbital Structure of the Atom

In section 9 of Unit 3, we were briefly introduced to the concept of energy levels and **orbitals**. Energy levels and orbitals tell us how and where an atom places electrons. Electrons can have different energies, which impacts the behavior of every atom in terms of its properties and its reactivity with other atoms.

When Niels Bohr and others first studied hydrogen, they considered that electrons "orbited" around the nucleus like planets around a sun. While we now know that electrons are not actually orbiting, the term orbital has stuck around to describe energy levels inside an atom. Bohr gave numbers to the different energy levels (1, 2, 3, and so on), but each of those energy levels can be further subdivided. So before we go any further, let's talk about these specific energy levels in an atom.

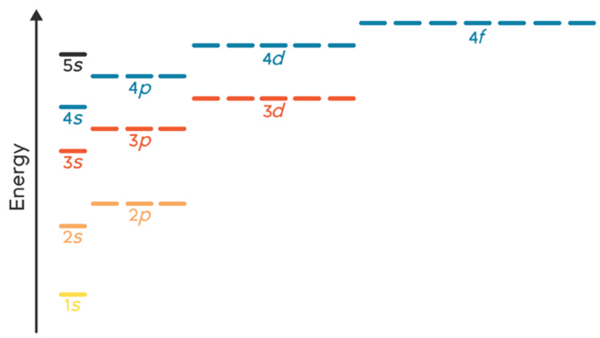

Figure 4-6. The Energy Level Diagram for an Atom

This figure shows the first four shells in a multi-electron atom with their relative energies. Each shell has been labeled a different color. A shell is a set of subshells that all have the same number. Within a given shell, all of the subshells are close to each other in energy, but not exactly the same. Roughly, each shell's subshells are lower in energy than the subshells of the next higher number. Note that the subshells are lined up over each other from right to left, just to make the diagram easier to read.

A set of energy levels that all share the same number is referred to as a "shell." Each shell is made up of a number of subshells. Subshells contain orbitals, which can each hold up to two electrons. Different types of subshells hold different numbers of orbitals. The first four types of subshells are called *s*, *p*, *d*, and *f*. These come from old terms *sharp*, *principle*, *diffuse*, and *fundamental*. What is most important to know about subshells is how many orbitals each contains. In summary, to find where an electron could be in an atom, look in an orbital, within a subshell, within a shell. Table 4-2 summarizes the four common subshells and their sizes.

Table 4-2. ***Summary of Each Type of Subshell***

Type of Subshell	Number of Orbitals in This Subshell	Maximum Electrons This Subshell Can Hold
s	1	2
p	3	6
d	5	10
f	7	14

Now, not every shell has every type of subshell in it. For example, if we look at Figure 4-6, we can see that the first shell only has an *s* subshell, and an *s* subshell is only a single orbital. As the shell numbers increase, each new shell has one new subshell. In Figure 4-6, we can see that the 4th shell has four subshells: 4*s*, 4*p*, 4*d*, and 4*f*, each which is made up of a different number of individual orbitals.

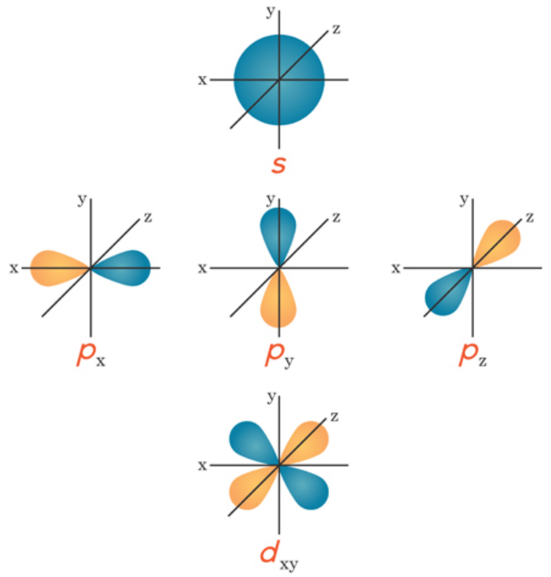

Figure 4-7. Shapes of Individual Orbitals

This figure shows a collection of specific orbitals that can be found in an atom. The shapes are overlaid on a set of Cartesian coordinate axes. First is the *s* orbital, which represents the spherical area where an electron can be. The *s* subshell only contains an *s* orbital. Then, there are all three orbitals that can be found in a *p* subshell: the p_x, the p_y, and the p_z, each of which looks like a dumbbell along the axis after which it is named. Just for comparison, one of the five *d* orbitals in a *d* subshell is shown.

In addition to being at different energies, the subshells actually represent different volumes of space where the electrons are likely to be found. (Figure 4-7) For example, an *s* orbital is always spherical in shape, indicating that an electron in an *s* orbital is likely to be found close to the nucleus. However, the other subshells are much more complicated. A *p* subshell has three different orbitals in it, each of which looks more like a dumbbell; this means that the electron is much more limited in the space where it can be found. Specifically for the *p* subshell, each of those orbitals actually points in a different direction from the other two. By the time we get up to the *d* and *f* subshells, there are five or seven different types of orbitals and they look quite complicated with lots of lobes pointing out into space. Why do we care about the shapes of the orbitals? In Unit 5, as we begin to discuss how atoms bond together, the shapes of the orbitals play a role.

Some of the electrons in an atom are much more important in how they affect the overall properties. We call these the **valence electrons**. In general, the valence electrons are the electrons in the outermost shell (the shell with the highest number), but that's a bit of an oversimplification. Thus, in order to determine how to predict the atom's chemical behavior, we have to first place electrons in their orbitals, and create an **electron configuration**. The following section examines the rules that will help us organize these electrons.

Glossary

Electron configuration

The arrangement of electrons in the orbitals of an atom.

Orbitals

The probable location and energy state of electrons.

Valence electrons

In general, the electrons in the outermost shell (the shell with the highest number). For most common elements, these are the *s* and *p* electrons.

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Section 6: Electron Configurations

Keeping track of all the electrons in an atom can be an intimidating task. In order to organize and inventory where all the electrons in an atom are, we use something called the "electron configuration." The placement of electrons in an atom dictates how each atom behaves, what compounds it will form, and how reactive it is. For example, full sets of orbitals confer stability, and unpaired electrons are highly reactive. In order to get to the point where we can understand and use electron configurations, first we need to understand some of the basic rules that electrons follow in orbitals.

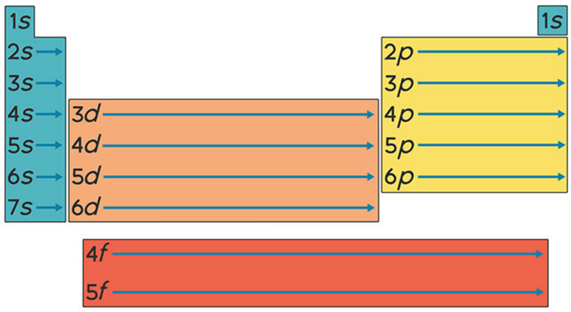

Figure 4-8. Orbital Blocks on the Periodic Table

The periodic table can be divided into blocks based on the location of the outermost electrons in the atom—called "valence electrons." The first two columns are the *s* block. Columns 11–18 are the *p* block, 3–10 are the *d* block, and the bottom two rows are the *f* block. The blocks, in addition to the period or row numbers, can be used to determine the electron configuration of any element.

Aufbau Principle

The Aufbau Principle states that electrons will fill the lowest energy orbital first. The word *Aufbau* is the German word meaning "filling." Roughly, the orbitals that are closest in proximity to the nucleus are the lowest in energy. Every atom starts by filling the 1*s* orbital first. However, the actual order of the orbitals is empirical, which means it must be experimentally determined. So how do we know the order of all these orbitals? Conveniently, the periodic table is laid out in this order from left to right, as can be seen in Figure 4-8. The order of the subshell energies can be transcribed from the periodic table and is represented as a list here:

1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s* 4*d* 5*p* 6*s* 4*f* 5*d* 6*p* 7*s*…

As we can see, the layout and design of the periodic table provide a shortcut for filling orbitals based on the rows and orbital blocks. The first two columns of the periodic table are known as the *s* block; the middle columns are the *d* block; the bottom rows are the *f* block; and the far right columns are the *p* block. These four letters represent the outermost subshells of each element. For the main group elements, each row is a new shell. For example, the third row is the beginning of the third shell, which contains a 3*s* and a set of 3*p* orbitals. Or if we look in the sixth row, we'll see these elements are filling the 6*s* and 6*p*. The *d* and *f* blocks do not quite follow the shell pattern. This is just part of the quirk of nature that all of the subshells in a shell are not at the exact same energy, and sometimes orbitals from a different shell can be lower in energy. For example, 3*d* is higher in energy than 4*s*.

Hund's Rule

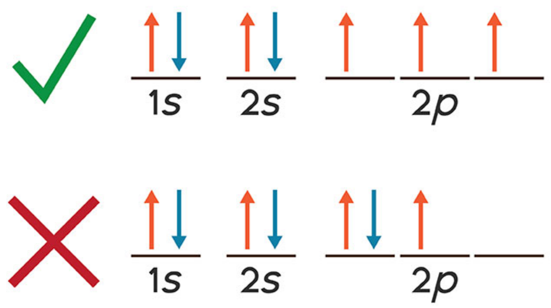

Figure 4-9. Hund's Rule

In this example for nitrogen, the electrons in the 2*p* subshell should be placed in each of the three available orbitals before doubling up. Electrons repel each other, and this method accounts for electrons naturally maximizing the space between themselves.

A German physicist, Friedrich Hund (1896–1997), developed a rule for filling a set of orbitals that all have equal energies. The *p*, *d*, and *f* subshells have multiple orbitals at the same energy. Every time there is a *p* subshell, it comes as a set of three orbitals, each of which can hold a maximum of two electrons. In Figure 4-6 of Section 5 of this Unit, lines represent orbitals. Note that 2*p*, 3*p*, and 4*p* subshells are drawn as three orbitals at the same energy level. Any *d* or *f* orbital will be drawn with 5 or 7 orbitals, respectively.

Hund's Rule can also be thought of as the Bus Rule: Imagine getting onto a city bus. There are many empty double seats. They slowly fill up with one person in each. Riders generally choose empty double seats before sitting down next to a stranger. Just like an electron, wouldn't we sit in the empty seat?

Pauli Exclusion Principle

In orbital notation, arrows pointing up or down represent electrons. An arrow that points up is representing a "spin up" electron, and an arrow that points down represents a "spin down" electron. There is another fundamental property of subatomic particles that is called "**spin**." However, nothing is actually spinning on the electron; that's just the name of this property. An electron can have only two possible spin values, up or down.

Electrons in an atom must follow the Pauli Exclusion Principle, which states that multiple electrons in the same orbital cannot have the same spin. Since there are only two types of spin, each orbital can only hold, at most, two electrons (one that is spin up and one that is spin down). In Figure 4-9, the three electrons in the 2*p* orbital are all of the same type of spin, one in each orbital. The next electron would be pointed down to represent the other type of spin, and placed in any one of the 2*p* orbitals to avoid violating the Pauli Exclusion Principle.

In summary, there are three basic principles for arranging electrons in an atom. First, fill orbitals from lowest energy to highest; second, put one electron into each orbital of a subshell before pairing up electrons in the same orbital; and third, place, at most, two electrons, with opposite spins in any given orbital. The electron configurations of the first 12 elements are shown in Table 4-3.

Table 4-3. ***Electron Configurations for the First 12 Elements***

Element	Electron Configuration	Element	Electron Configuration
Hydrogen	$1s^1$	Nitrogen	$1s^22s^22p^3$
Helium	$1s^2$	Oxygen	$1s^22s^22p^4$
Lithium	$1s^22s^1$	Fluorine	$1s^22s^22p^5$
Beryllium	$1s^22s^2$	Neon	$1s^22s^22p^6$
Boron	$1s^22s^22p^1$	Sodium	$1s^22s^22p^63s^1$
Carbon	$1s^22s^22p^2$	Magnesium	$1s^22s^22p^63s^2$

Noble Gas Configuration

As a shortcut, we can use noble gas configuration. The noble gases represent a full shell of electrons, which is why they do not react readily with other elements. Instead of writing out $1s^22s^22p^63s^2$ for magnesium, we can replace $1s^22s^22p^6$ with [Ne]. The noble gas configuration for magnesium would be [Ne]$3s^2$. The inner electrons represented by [Ne] are lower in energy and more stable than the two electrons in the $3s^2$ orbitals. Those electrons would be the ones involved in reactions.

Valence Electrons and Metals

In Section 5, we learned that valence electrons are the electrons that dictate atomic behavior. For the main group and non-metallic elements, it is easy to tell how many valence electrons an element has just by looking at the periodic table. As we move from left to right, each column has one more valence electron. By the time we reach the final column of noble gases, all the orbitals are full.

However, there are a lot of metallic elements on the table, and their valence electrons are a bit more complicated. Because they are what we call *f*-block or *d*-block elements, their valence electrons are actually a combination of their highest-level *s* orbital as well as some other electrons. The *d*-block elements, which we call the "transition metals," have their valence electrons in both the *s* orbital and the *d* orbitals. The *f*-block elements, the lanthanides and actinides, have their valence electrons in their outermost *s* orbital and in their *f* orbitals as well. The chemistry of metals is very complicated and will be revisited in Units 11 and 13.

Glossary

Spin

A quantum mechanical property associated with an electron that can be plus or minus one-half.

Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 7: Effective Nuclear Charge and Size

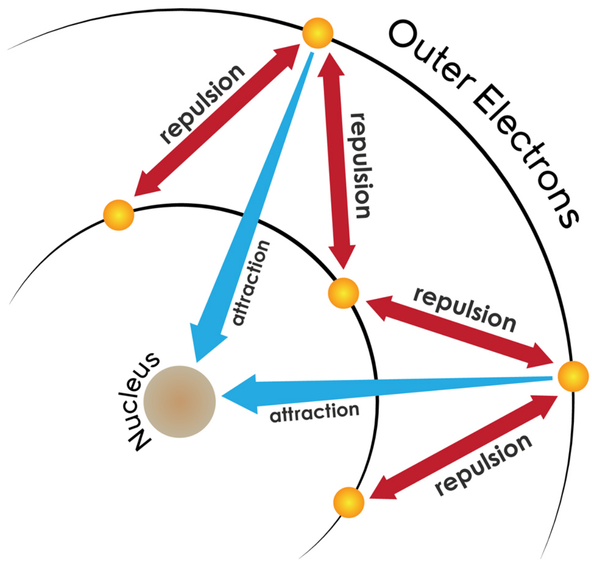

Figure 4-10. Effective Nuclear Charge

The outermost electrons in the atom control the size and reactivity of an atom. These outer electrons in an atom feel both the pull from the protons in the nucleus and the repulsion from the inner electrons. The balance between these attractive and repulsive forces is known as "effective nuclear charge." We also consider the inner electrons to be "shielding" the outer electrons from some of the attractive force of the nucleus.

An electron in an atom feels two forces. One comes from the nucleus, where positively charged protons attract the negatively charged electrons. The second force comes from the other electrons, which repel each other because of their like charges. Effective nuclear charge is one way of expressing this balance between the attractive and repulsive forces of electrons when we think about the outermost valence electrons in an atom. (Figure 4-10)

To calculate effective nuclear charge, we have to take the total number of electrons into account. Electrons in outer energy shells feel a weaker pull from the nucleus than the inner electrons because the inner electrons "shield" some of the positive force coming from the protons in the nucleus. This means that the outermost electrons feel the weakest pull from the nucleus, and the inner electrons feel the most pull. We can calculate a numerical value for effective nuclear charge (Z_{eff}) by subtracting the number of shielding, or inner, electrons from the atomic number. From left to right on the periodic table, Z_{eff} increases because the number of shielding electrons within each row stays constant while the number of protons increases. Z_{eff} is one of the factors in the size of an atom. If an atom has a strong pull on its outer electrons, the atom as a whole will be smaller.

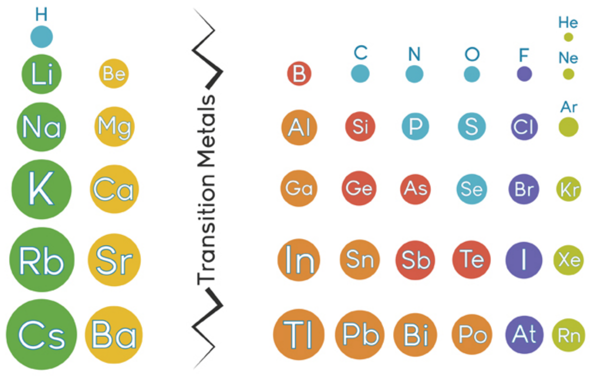

Figure 4-11. The Trend in Atomic Radii

The circles we see above represent atomic radii. To see the trend more easily, the image above only contains the main group elements. The radii increase from top to bottom and decrease from left to right. The increase going down each group is due to additional shells of electrons, which also decreases Z_{eff}. This makes each element in a group much larger than the one above it. The decrease going across a period is due to a gradual increase in the effective nuclear charge (Z_{eff}). The nucleus has a stronger hold on its electrons, so the radius decreases.

Despite knowing Z_{eff}, it is very difficult to determine the size of an atom, which is the size of the electron cloud around the nucleus of an atom. But because we can never know the exact location of an electron, we cannot actually specify the size of an atom. One way to get around this problem is to assume that the atom is a perfect sphere. We can then determine an atomic radius by measuring the nucleus to the edge of the spherical cloud of electrons. There are a variety of experimental and theoretical methods to determine or calculate the atomic radius for an element. Using these radii, going from left to right across the periodic table, the atomic radius decreases because effective nuclear charge increases. (Figure 4-11)

Another determinant of atomic size is electron shells. Each row on the periodic table starts a new shell for the electrons. Each subsequent shell is much farther from the nucleus, so atomic radius increases down a column on the periodic table.

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Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 8: Ionization Energy and Ionic Radius

Absorption of Strontium in the Aftermath of Chernobyl

Elements in the same group react in a similar manner. For example, every element in the second group forms a +2 ion, including calcium, which is a main structural component of bones and teeth. Because elements in the same group behave in similar ways, health problems can occur if we are exposed to a different element from the same group instead of calcium. Strontium is another alkaline earth element below calcium on the periodic table. Unfortunately, in the case of nuclear disasters, the radioactive strontium-90 isotope is often formed.

Chernobyl Plant © Wikimedia Commons, Public Domain.

On April 26, 1986, one of the reactors exploded at a nuclear power plant in Chernobyl in the Ukraine. When the plant exploded, it sent radioactive isotopes, including strontium-90, into the air. Weather patterns blew the radioactive cloud across Europe. Because of its chemical similarities to calcium, plants absorbed the strontium-90 thinking it was calcium. Animals ate the plants, and humans ate the plants and animals. Again, due to the similarity between strontium and calcium, bones absorbed the strontium-90, causing terrible health problems. Once inside the body, strontium-90 atoms can cause cellular damage, leading to bone cancer or leukemia. One effective treatment for strontium-90 exposure is to take calcium supplements to out-compete the strontium for positions in the bone. Unfortunately, once strontium-90 is in the bone, it is there forever. To this day, there is a 19-mile radius around the site of the accident in Chernobyl where humans are forbidden to live because of health concerns.

Ions

An atom in its neutral state has an equal number of positively charged protons and negatively charged electrons. When the atom gains or loses electrons, the net charge of the entire atom changes. An atom with a charge is called an "ion." If the atom loses an electron, there are more protons than neutrons, and the atom has a net positive charge. A positively charged atom is called a "**cation**." Metals commonly become cations. An atom that gains electrons and has a net negative charge is called an "**anion**."

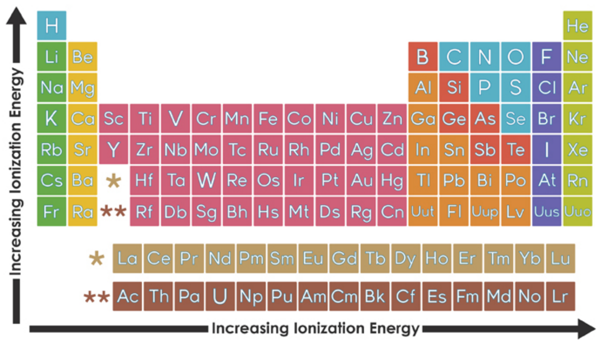

Figure 4-12. The Trend in Ionization Energy

Ionization energy increases across a period. Note the highest values for each row correspond to the noble gas. Ionization energy decreases down a group. This trend corresponds to the trend in atomic radii. Electrons that are closer to the nucleus are more difficult to remove.

Ionization Energy

Many of the chemical reactions that take place involve electrons moving from one atom to another, creating ions. Thus, it is important to understand the patterns of how easy or hard it is to make an ion in order to understand the reactivity of the elements. To understand the first part of this pattern, we need to look at the ionization energies for the atoms. Ionization energy is the amount of energy it takes to remove an electron from an atom when it is in the gas phase. We use the gas phase because the atom is freely floating and easy to isolate. Ionization energy is the lowest at the left and bottom of the periodic table, where the atoms are the largest. (Figure 4-12) This is because electrons that are far away from the nucleus are higher in energy and are lost much more easily. Also, larger atoms tend to lose these electrons easily because the electrons feel less pull from the nucleus. For example, it takes less energy to remove an electron from radon, which is at the bottom of Group 18, than helium, which is at the top of Group 18.

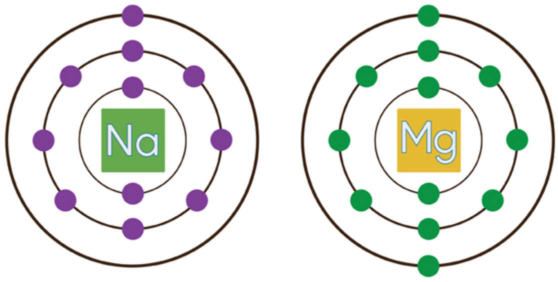

Figure 4-13. Models of Sodium and Magnesium

Sodium has one valence electron, whereas magnesium has two. The first ionization energy is going to be lower for sodium than for magnesium because sodium wants to get rid of its one valence electron. However, the second ionization energy will be much lower for magnesium than the second ionization energy for sodium. This is because sodium is already stable and has the full shell electron configuration of a noble gas, but magnesium still has one more loosely held electron in its outermost shell.

The ionization energy just described is called "first ionization energy," because it refers to the energy required to remove the first electron from an atom. Second or third ionization energies, the energy it takes to remove a second or third electron, vary depending on how many valence electrons are in the atom. The second ionization energy for sodium will be much higher than that for magnesium. Sodium loses its one valence electron rather easily because without that electron, it has a stable, full outer shell. Once an atom has a full outer shell, like the noble gases, it takes a lot more energy to remove electrons. By comparison, magnesium has two valence electrons, so its first and second ionization energies would both be relatively small. Removing the second valence electron from magnesium causes the ion to have a full outer shell, which is a stable configuration. Magnesium's third ionization energy would be quite high and comparable to sodium's second ionization energy. (Figure 4-13)

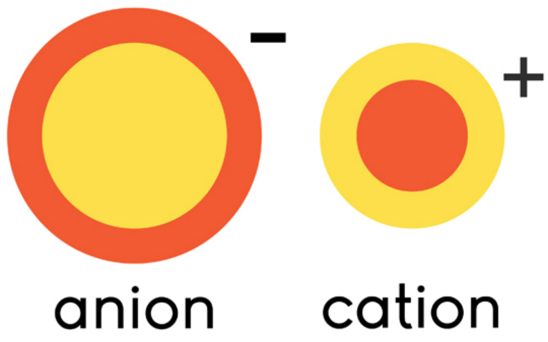

Figure 4-14. Ionic Radius—Anion (left) and Cation (right)

In the image above, the yellow circle represents the original atomic radius. On the left, the atom has gained an electron. There is one more electron in the electron cloud. It has an overall negative charge, and the radius increases. The orange outline represents the change in size. On the right, the atom has lost an electron. There is an overall positive charge, and the nucleus has a tighter hold on the electrons. There is one less electron in the electron cloud, and the radius decreases. The solid orange circle represents the smaller radius.

Ionic Radius

Electron gains or losses affect the size of an atom. When an atom loses an electron, there is a net positive charge, causing each electron to feel a greater pull of the nucleus, and there is also a decrease in the number of electrons making up the cloud. The greater net pull causes the radius to decrease. When an atom gains an electron, there is a net negative charge, causing each electron to feel a weaker pull from the nucleus, and there is now an additional electron in the electron cloud. This causes the radius to increase. (Figure 4-14)

Glossary

Anion

A negatively charged atom or molecule.

Cation

A positively charged atom or molecule.

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Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 9: Forming Compounds

At the heart of chemistry is the reactivity of atoms. All atoms aim to become more stable by gaining, losing, or sharing electrons in order to have full outer shells. There are two major types of bonds that achieve this goal: ionic and covalent.

Ionic Bonds

Ionic bonds are interactions between oppositely charged atoms, or ions. In ionic bonds, usually a metal gives up electrons to a non-metal, forming a cationic metal and an anionic non-metal. The general term for compounds with ionic bonds is "salt." The compound that we know as table salt is a compound of sodium and chlorine, called "sodium chloride." Many other compounds we interact with daily are salts, such as calcium oxide (a concrete precursor also known as "quicklime") and tin (II) fluoride (found in toothpaste).

To investigate ionic bonds further, let's look at sodium chloride (NaCl). Sodium, a metal, had one valence electron ([Ne] $3s^1$). To become stable, it would be easiest for sodium to give one electron away. On the other side, chlorine needs one electron to fill up its outermost shell with eight electrons and become stable ([Ne] $3s^2 3p^5$). Sodium gives up its electron to chlorine, so sodium becomes a cation and chlorine an anion. After the transfer of electrons, sodium has a positive charge and chlorine has a negative charge, and now both atoms have stable full-shell electron configurations of [Ne] and [Ar]. Electrostatic interactions cause the ions to attract, or bond. Even though each atom in sodium chloride is more stable as an ion, the overall goal for a molecule is to be neutral; all of the charges of the ions must perfectly cancel each other out.

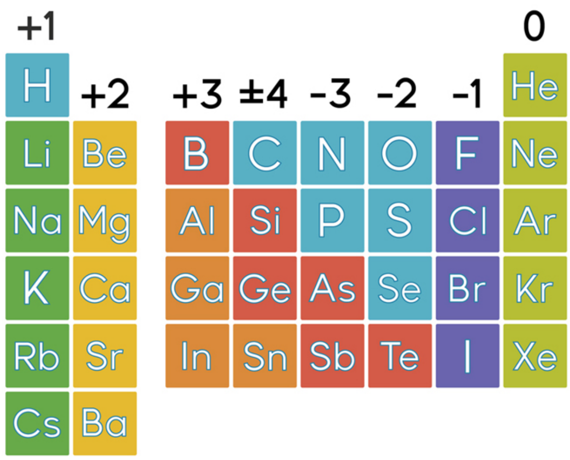

Figure 4-15. The Charges of the Eight Main Group Families

The column number relates to the charge of the elements. Each element in a given column has the same number of valence electrons, so they also tend to form the ions of the same oxidation state. For example, all of the elements in Group 2 have two valence electrons and tend to lose both to form a +2 ion. This is another pattern on the periodic table that relates directly to electron configurations.

Another pattern on the periodic table, which works very well for main group elements, is the charge that their ions like to form. The charge on an element, positive or negative, is referred to as the "oxidation state" of the atom. For the main group elements, those in columns 1, 2, and 13-18, the position of the element conveys the charge it will make. (Figure 4-15) Elements in Group 1 make a +1 ion, whereas elements in Group 17 elements make -1 ions. In each case, the elements are gaining and losing electrons to try to fill up or empty out their valence shell to have a stable noble gas configuration.

Covalent Bonds

Some atoms do not lose and gain electrons to obtain a full outer shell. They can also share electrons with other atoms. In covalent bonds, two atoms will share the same electron to fill their electron shells; but the atoms do not always share the electron equally. Covalent bonds are much more common in chemistry and occur between two or more non-metals. Unit 5 will focus on this type of bonding, which includes molecules like the nitrogen found in air (N_2) and water (H_2O).

Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 10: Electronegativity

Sometimes atoms in covalent bonds will share electrons equally, but sometimes one atom will have a stronger hold on a shared electron. The measure of how much pull an atom has on electrons in a covalent bond is called "electronegativity." If there is no difference between the atoms' electronegativity values, then the bond is perfectly covalent and considered non-polar, like a bond between two nitrogen atoms in N_2. If there is unequal sharing, the bond is considered polar. An example of this is a bond between oxygen and hydrogen in water (H_2O). Hydrogen's electronegativity is lower than oxygen, so the shared electrons will spend more time around the oxygen atom. (Figure 4-16)

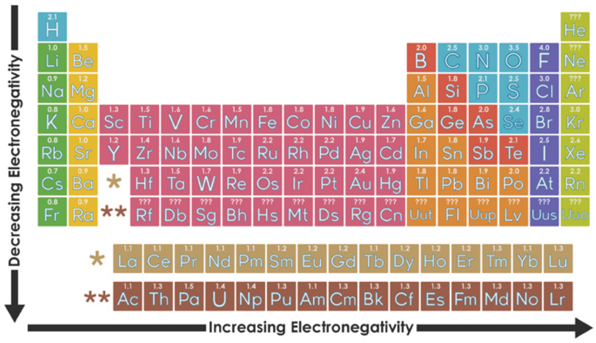

Figure 4-16. The Trend in Electronegativity

This diagram shows the Pauling Scale, named after its creator, Linus Pauling (1901–1994), for the electronegativities of the elements. Note how the electronegativity increases across a period and decreases down a group. Electronegativity is a measure of how much an atom "hogs" electrons in a covalent bond. Note that some of the noble gases do not have electronegativity values. Noble gases do not make bonds, so they do not have electronegativity values.

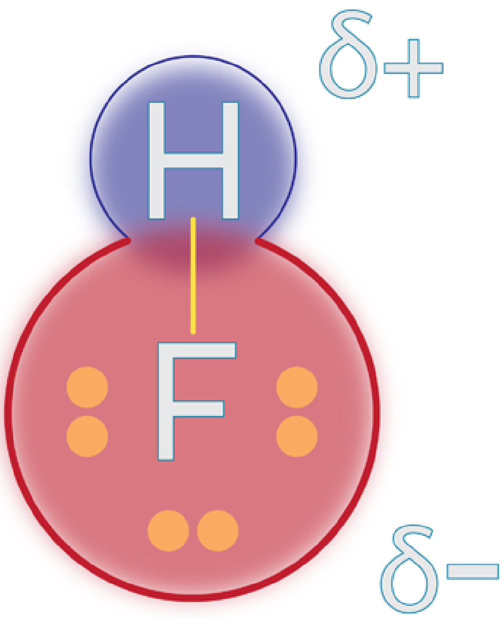

Figure 4-17. A Polar Covalent Bond

Fluorine is much more electronegative than hydrogen, so the shared electrons spend more time around fluorine than hydrogen. Because of this unequal sharing, fluorine has a partial negative charge, and hydrogen has a partial positive charge.

In a polar covalent bond, the atom that has the higher electronegativity will obtain a slight negative charge because the electrons spend more time around that atom. The atom that has the lower electronegativity will obtain a slight positive charge. In Figure 4-17, the lower case letter delta, δ, represents a partial charge. Fluorine is more electronegative than hydrogen, so F has the partial negative charge. The chemical hydrogen fluoride (HF) is quite toxic to humans, but it is the perfect chemical for an artisan who needs to etch glass. These partial charges have implications for properties like boiling point, freezing point, and the viscosity of compounds. These properties and their relationship to bond polarity will be discussed in more detail in Unit 5.

Unit 4: *Organizing Atoms and Electrons—The Periodic Table*

Section 11: Naming Compounds

When we look at the ingredients on the back of a shampoo bottle, cleaning fluid, or even a candy bar, there are many long words that are difficult to pronounce. Believe it or not, there is a method behind the madness. Each of those ingredients is written in chemical nomenclature. It is actually much simpler than it seems at first because there are specific rules to follow when naming chemical compounds.

Ionic Compound Naming

When naming ionic compounds, the name of the cation metal always comes first and always remains the same as the element name. Sodium stays sodium; zinc stays zinc. The non-metal, or anion, is modified to end in -ide. Chlorine becomes chloride, sulfur becomes sulfide, and so on.

Examples:
$AlCl_3$ = aluminum chlor*ide*
Na_2S = sodium sulf*ide*
K_2O = potassium ox*ide*

One exception to these guidelines is an ionic compound involving **polyatomic ions**, which are groups of non-metal atoms that in a compound have a positive or negative charge. Some common examples are listed in Table 4-4 below:

Table 4-4. ***The Names and Formulas of Some Common Polyatomic Ions***

Name	Formula
Ammonium	NH_4^+
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}
Carbonate	CO_3^{2-}
Bicarbonate	HCO_3^-
Nitrate	NO_3^-
Permanganate	MnO_4^-

The subscripts confer the number of each element in the polyatomic ion. For example, there are one nitrogen atom and four hydrogen atoms in ammonium with an overall charge of +1. When forming ionic compounds with polyatomic ions, the subscript remains untouched as seen in Table 4-5. When there is more than one polyatomic ion in a compound, parentheses are used to maintain the integrity of the unit. NH_4^+ must be treated as one unit.

Table 4-5. ***Naming Compounds with Polyatomic Ions***

Cation	Anion	Formula	Name
NH_4^+	Cl^-	NH_4Cl	Ammonium Chloride
NH_4^+	NO_3^-	NH_4NO_3	Ammonium Nitrate
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$	Ammonium Sulfate
Na^+	HCO_3^-	$NaHCO_3$	Sodium Bicarbonate

Most transition metals can form more than one ion of differing oxidation states. For example, copper can form Cu^+ and Cu^{2+}. When naming ionic compounds involving transition metals, chemists use Roman numerals to communicate the specific charge on the metal. So, copper ions can be paired with chloride ions that have a negative one charge (Cl^-), but the overall compound must be neutrally balanced. Copper(I) chloride has the formula of CuCl, whereas copper(II) chloride has the formula of $CuCl_2$. Note how the copper(II) ion needs two chloride ions to balance the charge.

Covalent Compound Naming

When naming covalent compounds made from just two elements, we use Greek prefixes, like *mono*-, *di*-, *tri*-, and so on, to communicate how many of each atom are in the molecule. Common examples here include dinitrogen tetroxide (N_2O_4) and carbon dioxide (CO_2). In this method of naming compounds, the most electronegative element is always listed second.

Glossary

Polyatomic ions

Charged molecules with more than one atom.

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Section 12: Conclusion

The periodic table is a powerful tool for predicting the properties of elements based on their electrons, and electron configurations are a great tool for understanding how electrons behave. Atoms lose, gain, and share electrons to obtain a full outer shell in order to maximize their stability. Additionally, there are regular patterns, or trends, in the properties of atoms, such as atomic size and ionization energy. Once we understand the hidden secrets of the periodic table, we can start to explain and predict the formation of compounds. In Unit 5, we will learn more about covalent bonds, and how the shape and properties of molecules affect the physical and chemical characteristics of substances.

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