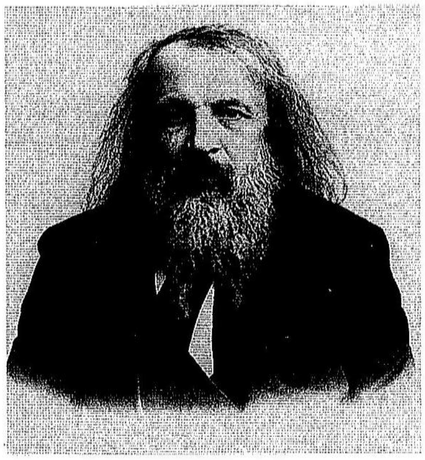
THE PERIODIC TABLE

Although the world is varied and complex, everything in it—air, water, rocks, living tissue, and the almost infinite number of other objects and materials around us—is actually made up of only a limited number of chemical elements. We know today that only 91 such elements exist naturally on the Earth. They range from hydrogen, the lightest element, to uranium, the heaviest. Actually, several more elements do exist, but these have to be made artificially in laboratories.

The basic components of each chemical element are atoms. The atoms of an element consist of three kinds of particles: protons, neutrons, and electrons. Protons and neutrons exist at the core, or nucleus, of the atom. One of the important ways in which these two kinds of particles differ from one another is that each proton carries a single, positive electric charge, whereas a neutron carries no electric charge. Electrons, which are much smaller than either protons or neutrons, each carry a single negative electric charge. Electrons are present at some distance away from the nucleus of the atom and travel rapidly around it in complex paths known as orbits. Under normal circumstances, the number of electrons orbiting around the nucleus of a particular atom is exactly equal to the number of protons in the nucleus of the atom, so that the overall positive electric charge provided by its protons is exactly balanced by the overall negative charge provided by the electrons orbiting its nucleus.

The unique properties of each of the chemical elements are determined by their number of neutrons, protons, and electrons. Besides determining the properties of a pure chemical element, the neutron, proton, and electron content of its atoms also determines its behavior in relation to other chemical elements. Although each element behaves differently and has different properties from all of the others, the atoms of different elements can combine with one another to form clusters of atoms called molecules. It is this combination of atoms that accounts for the enormous variety of chemical substances that can be found in nature and created by modern technology.

When scientists first tried to describe the physical and chemical properties of the elements and chemical compounds, which are formed by the combination of atoms of different elements, they soon became buried under a mountain of seemingly unconnected facts. Many early scientists recognized the need to organize this information, and they attempted to discover some sort of order or pattern that could simplify what seemed to them an overwhelming array of chemical facts. The solution to the



The Russian chemist Dmitry Mendeleyev created his periodic table (below) in 1869 while preparing a chemistry textbook for his students.

problem was the so-called periodic table of the chemical elements.

The modern periodic table is based primarily on the work of the Russian chemist Dmitry Ivanovich Mendeleyev (1834–1907) and the German physicist Julius Lothar Meyer (1830–1895). Working independently, both of these scientists developed similar periodic tables within a few months of each other in 1869. Mendeleyev, however, is usually given the credit for having developed the periodic table because he managed to publish his work first.

Mendeleyev, who was a professor of chemistry at the University of St. Petersburg, developed the periodic table while preparing a chemistry textbook for his students. As part of this project, he had written down the properties of the elements on cards. While sorting through these cards he noticed that when the elements were arranged in order of their weight, similar chemical properties occurred repeatedly at regular intervals. Mendeleyev used this observation to construct his periodic table. He placed the elements in horizontal rows, according to their weight, with the lightest in each row at the left end of the row and the heaviest at the right, and with one row beneath the other, so that all elements with similar properties fell into vertical columns.

Much of Mendeleyev's success depended on his placing elements with similar properties in the same group, even though this left occasional gaps in the table. He reasoned, correctly, that the elements that belonged in the gaps had not yet been discovered. The locations of the gaps enabled Mendeleyev to predict with remarkable accuracy the properties of these yet-to-be found elements.

The table that Mendeleyev developed is in many ways similar to the one we use today. One of the main differences is that Mendeleyev's table lacks the column containing the elements helium through radon. In Mendeleyev's time none of the elements in this column had yet been found because they are relatively rare and because they show no tendency to undergo chemical reactions. Occasionally, Mendeleyev was forced to switch the order of his elements to make the table come out right, placing elements with greater atomic weights ahead of those with smaller weights. The atomic weight of an element is the average weight of all the atoms that form the element. It is mostly determined by the total number of protons and neutrons the atoms contain. Electrons are so much lighter than these

xviii PRINCIPLES OF CHEMISTRY

PERIODIC SYSTEM OF THE ELEMENTS IN GROUPS AND SERIES.

Series	GROUPS OF ELEMENTS 0	I	II	III	IV	V	VI	VII	VIII
1	—	Hydrogen H 1·008	—	—	—	—	—	—	
2	Helium He 4·0	Lithium Li 7·03	Beryllium Be 9·1	Boron B 11·0	Carbon C 12·0	Nitrogen N 14·04	Oxygen O 16·00	Fluorine F 19·0	
3	Neon Ne 19·9	Sodium Na 23·05	Magnesium Mg 24·3	Aluminium Al 27·0	Silicon Si 28·4	Phosphorus P 31·0	Sulphur S 32·06	Chlorine Cl 35·45	
4	Argon Ar 38	Potassium K 39·1	Calcium Ca 40·1	Scandium Sc 44·1	Titanium Ti 48·1	Vanadium V 51·4	Chromium Cr 52·1	Manganese Mn 55·0	Iron Cobalt Nickel Fe Co Ni (Cu) 55·9 59 59
5		Copper Cu 63·6	Zinc Zn 65·4	Gallium Ga 70·0	Germanium Ge 72·3	Arsenic As 75	Selenium Se 79	Bromine Br 79·95	
6	Krypton Kr 81·8	Rubidium Rb 85·4	Strontium Sr 87·6	Yttrium Y 89·0	Zirconium Zr 90·6	Niobium Nb 94·0	Molybdenum Mo 96·0	—	Ruthenium Rhodium Palladium Ru Rh Pd (Ag) 101·7 103·0 106·5
7		Silver Ag 107·9	Cadmium Cd 112·4	Indium In 114·0	Tin Sn 119·0	Antimony Sb 120·0	Tellurium Te 127	Iodine I 127	
8	Xenon Xe 128	Cæsium Cs 132·9	Barium Ba 137·4	Lanthanum La 139	Cerium Ce 140	—	—	—	— — —
9		—	—	—	—	—	—	—	
10	—	—	—	Ytterbium Yb 173	—	Tantalum Ta 183	Tungsten W 184	—	Osmium Iridium Platinum Os Ir Pt (Au) 191 193 194·9
11		Gold Au 197·2	Mercury Hg 200·0	Thallium Tl 204·1	Lead Pb 206·9	Bismuth Bi 208	—	—	
12	—	—	Radium Rd 224	—	Thorium Th 232	—	Uranium U 239		
	HIGHER SALINE OXIDES								
	R	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
	HIGHER GASEOUS HYDROGEN COMPOUNDS								
					RH_4	RH_3	RH_2	RH	

nuclear particles that they contribute very little to the weight of an atom. Apparently, listing elements in order of their atomic weights did not always work. It was not until the beginning of the 20th century, with the knowledge gained about the structure of the atom, that the correct way of ordering the elements was discovered and the present periodic table was formulated.

THE NUCLEAR ATOM

The key event that led to the modern understanding of the atom was the discovery that atoms are made up of electrons, protons, and neutrons. Thus, despite its name, which derives from the Greek word for "indivisible," the atom could indeed be divided into smaller components.

In April 1897, Joseph John Thompson, professor of physics and director of the Cavendish Laboratory at Cambridge University in England, announced the discovery of the electron. Thompson reported that this tiny particle had a negative electric charge and a mass of about one two-thousandth of that of the lightest atom. Thompson's momentous discovery of a particle of matter smaller than the atom so startled his colleagues that many thought he had been "pulling their legs." It was no joke, however, and in Thompson's own words: "The production of electrons essentially involves the splitting up of the atom, [with] a part of the mass of the atom getting free and becoming detached from the original atom—that part being one or more electrons."

Ernest Rutherford, the distinguished New Zealand physicist who had been a pupil of Thompson's and who was a professor of physics at Cambridge University, supplied the next step toward the modern understanding of the atom in 1911. He discovered that the atom had a nucleus and that one of the important particles that occupied the nucleus was the positively charged proton.

As a probe for his study of the atom, Rutherford used the newly discovered phenomenon of radioactivity. Radioactive atoms, like uranium and radium, are unstable, and their nuclei spontaneously disintegrate. One of the products of this disintegration is a massive, positively charged particle called an alpha particle. At the time of its discovery, Rutherford did not know that the alpha particle was the nucleus of a helium atom, consisting of two protons and two neutrons. He therefore used the first letter of the Greek alphabet, *alpha,* to identify this particle and distinguish it from the other products given off by radioactive atoms.

Because the atom was too small to observe directly, Atoms are normally electrically neutral, with an equal number of electrons and protons. This means, for example, that carbon, with an atomic number of 6, has six protons in its nucleus and six electrons outside the nucleus.

ISOTOPES

Moseley's experiments demonstrated that what distinguishes one element from another is its atomic number, the number of protons in the nucleus of its atoms, not its atomic weight, which is a measure of the total number of protons and neutrons in the nucleus. The correct way of ordering the elements in the periodic table was, therefore, by their atomic number, and not, as Mendeleyev had thought, by their atomic weight.

Although all the atoms of a given element have the same number of protons, they can have different numbers of neutrons. This explains, for example, why there are three different species of the element hydrogen. Ordinarily, a hydrogen atom has a lone proton in its nucleus and no neutrons. A heavier form of hydrogen, called deuterium, also has a single proton in the nucleus but contains a neutron as well. A still heavier form of hydrogen, known as tritium, has two neutrons in addition to the proton. These three species are called isotopes of the element hydrogen. Yet even though a deuterium atom, because of its extra neutron, weighs twice as much as an ordinary hydrogen atom, its chemical behavior is similar to that of hydrogen, indicating that the number of protons in the nucleus is what determines the behavior of each element.

Like hydrogen, the majority of the elements have isotopes. Some elements have only two isotopes, while others can have as many as eight or nine. It is a remarkable fact that the relative percentage, or abundance, of each of the various isotopes of any element is the same all over the Earth.

The existence of isotopes also explains why the atomic weight is an unreliable indicator of the position of an element in the periodic table. For any element, the atomic weight really measures the average weight of a mixture of its different isotopes. On this basis, it is possible for an element like argon, which has an atomic number of 18, to exist in a mixture of isotopes that have a greater average atomic weight than that of potassium, whose atomic number is 19. The atomic weight in the periodic table is often a number with a decimal fraction. The atomic weight of calcium, Ca, for example, is 40.08. The nucleus of the calcium atom cannot contain 0.08 of a neutron. This number is derived by averaging the weights of the isotopes of calcium that occur together, some of which have more than 20 neutrons in the nucleus.

THE MODERN PERIODIC TABLE

The modern statement of the periodic law is that the chemical and physical properties of the elements vary in a periodic way with their atomic numbers.

The modern periodic table is arranged very much like Mendeleyev's table. The elements are arranged in rows called periods; in each period the elements are arranged in order of their atomic number. The periods are numbered from 1 to 8 from the top row to the bottom row. Below the main body of the table are two long rows of 14 elements each. One of these long groups follows lanthanum ($Z = 57$) and is known as the lanthanides. The other group follows actinium ($Z = 89$) and is known as the actinides. These elements actually belong in the main body of the table but are too long to fit conveniently into it.

The vertical columns of the periodic table are also called groups. There has been some disagreement about how these should be numbered. In one commonly used system the groups are labeled with Roman numerals and divided into A-groups and B-groups. The International Union of Pure and Applied Chemistry (IUPAC), an international body of scientists responsible for setting standards in chemistry, has officially adopted a system in which the groups are simply numbered in sequence from left to right, using Arabic numerals from 1 to 18. Thus, Group VIIA in the old Roman numeral system is Group 17 in the IUPAC system. The scientific world has still not achieved uniformity in the system used for the periodic table, and most chemists in the United States prefer the more traditional system. We will also use the traditional system in this book, but reference to the periodic table shown on page 6 will quickly translate the heading for a particular column into the IUPAC scheme.

All of the elements within a group have similar chemical properties and are sometimes referred to as families of elements. The elements in the A-groups, or longer groups, are known as representative elements. The elements in the B-groups are called transition elements.

Many of the groups of elements in the periodic table have acquired common names. For example, the elements in group IA, with the exception of hydrogen, are called the alkali metals. The

Rutherford's brilliant idea was to use alpha particles as projectiles, firing them at atoms and observing how they scattered. This was like firing bullets at a sealed box and deducing the contents of the box by seeing how the bullets bounced. His target atoms were at first gold atoms contained in very thin sheets of gold foil. Gold was used because it is possible to make gold foil that is very thin, often thinner than fine paper. Rutherford observed that although most of the alpha particles passed right through the target, many were deflected at very large angles. Some were even deflected backward, as if they had hit a stone wall. Rutherford was so astonished by this that he compared it to "firing a 15-inch shell at a piece of tissue paper and having it come back and hit the gunner." Because most of the alpha particles went right through the foil, he reasoned that the atom was mainly empty space but that it must contain a small, heavy, positively charged core that was capable of repelling and scattering the projectiles fired at it. Rutherford called this massive core the nucleus of the atom.

After Rutherford's discovery of the nucleus, it became obvious that the nucleus of hydrogen, the lightest of the atoms, must play a fundamental role in the structure of all atoms. In 1920, he proposed to call this particle the proton, the name by which it has been known ever since.

Finally, in 1932, the British physicist Sir James Chadwick, who also worked at the Cavendish Laboratory in Cambridge, discovered that yet another particle existed in the nucleus of atoms. This new particle was the neutron. It has a mass close to the mass of the proton, but it has no electric charge.

These fundamental discoveries, coupled with the work of a brilliant young English physicist named Henry Moseley, ultimately led to the reason for Mendeleyev's success with the periodic table. Moseley, just before World War I, had been investigating the X rays given off by various elements. X rays are a very penetrating form of radiation usually produced by accelerating electrons to high speeds and then abruptly stopping them by having them smash into a metal target. The collision causes the target to give off X rays. When different elements are used as targets, the X rays have different properties. Each element has its own set of characteristic X rays. They are almost like a fingerprint of the element. Moseley was able to relate the properties of the X rays to the number of protons contained in the element. He discovered that every element had a different number of protons in the nucleus. The number of protons came to be called the atomic number of the element, represented by the letter Z, and it was always a whole number.

elements in Group IIA are called the alkaline-earth metals, and those in Group VIIA are called the halogens.

What causes this periodic behavior of the elements? Why do the elements within a particular group have similar chemical behavior? The reason is that atoms are attracted to each other by electric forces. The atomic number, the number of positively charged protons in the nucleus, determines how many negatively charged electrons are contained in the atoms of a particular element, and it is the electrons that determine how elements behave and react with one another. The chemical behavior of an element is determined by the way in which the electrons orbiting the nucleus are structured. It was the new quantum physics, developed in the early 20th century by the Danish physicist Niels Bohr, the German physicist Werner Heisenberg, and the Austrian physicist Erwin Schrödinger, that put forward the idea of a complex arrangement of electron orbits or "energy levels" as a way of explaining the bonding properties of elements.

This new quantum physics, based on the idea that matter has properties resembling those of waves, tells us that the electrons in an atom are restricted to certain orbitals. These orbitals, which vaguely resemble the orbits of the planets of our solar system around the sun, are often referred to as shells. The inner shells, closest to the nucleus, are the most stable, and the electrons in these shells are closely held by the attractive force of the nuclear protons. If an electron absorbs energy, it jumps to the next outer orbital. If an electron in an outer orbital gives off energy, it drops to the next inner orbital. Electrons in the outer shells are relatively loosely bound to the nucleus. These electrons may be attracted to other atoms or they may become energetic enough to separate from the atom altogether, leaving behind an atom with a net positive charge that will attract electrons belonging to other atoms.

Strict rules govern how many electrons can occupy any particular shell of an atom. For example, two electrons will fill the first shell closest to the nucleus, whereas eight can occupy the next shell, slightly farther out from the nucleus, and eighteen can occupy the shell beyond this. Because each major shell contains various subshells, the exact electron configuration of an atom can become quite complex. The distribution of the electrons in the outer shell of the atom, the one farthest from the nucleus, is the important one, however, because these are the electrons that are exposed to other atoms when the atoms react.

Atoms with similar outer-shell configurations have similar chemical properties. Chemists call the outer shell the valence shell, and the electrons that occupy it are known as the valence electrons.

The term *valence* is derived from the Latin word *valent,* which means "strength." The valence electrons determine the chemical "strength" of atoms—their reactivity, or how strongly and in what way they will bind with other atoms. Elements in the same group in the periodic table have the same number of electrons in their outer shells and are therefore said to have the same valence electron configuration. As a result, the chemical and physical properties of the elements in this group will be similar. As the inner shells of an atom become filled with electrons, its outer shell takes on a specific valence configuration that is determined by the rules that govern how many electrons can occupy a particular shell. It is this regularity in the number of electrons that occupy the outer shell that accounts for the periodic behavior shown by the elements as the atomic number increases. Other properties, such as the size of an atom, are also determined by the number of shells it contains. For example, the radius of the atoms of the elements in a particular group in the periodic table tends to increase from the top of the group to the bottom.

Elements whose shells are completely full are extremely

HOW ELECTRONS OCCUPY ATOMIC SHELLS

	Scientific name	Permitted subshells	Maximum electrons in subshell	Maximum electrons in shell
shell closest to nucleus	K	1s	2	2
next shell farther out	L	2s	2	8
		2p	6	
next outer shell	M	3s	2	18
		3p	6	
		3d	10	
next outer shell	N	4s	2	32
		4p	6	
		4d	10	
		4f	14	

stable and seem to react with almost nothing else. The elements of Group VIIIA, for example, the so-called noble gases, all have complete shells and are the most chemically inert elements that exist. A complete shell of electrons is so energetically stable that atoms with incomplete shells will tend to react with other atoms in a manner that will complete these shells. In other words, atoms react in order to attain a "noble gas" configuration. In moving from left to right across a horizontal row, or period, within the periodic table, there is a transition from elements that are metals to those that are nonmetals. Metals, which generally have few electrons in their outermost valence shells, tend to lose electrons when they react, so that they reach a state in which they have fewer shells, all of which are completely filled with electrons. Nonmetals, whose valence shells are almost completely filled, tend to accept electrons to fill these shells and stabilize their configuration. In both cases, the tendency is to assume a completely filled valence shell, approximating that of a noble gas.

The periodic table, then, is a map of the way in which electrons arrange themselves in the atoms of a particular element. As you go down a column within a group, all the elements of that group have the same number of valence electrons. As you go across a row, from left to right, electrons are being added to a shell. The ability to predict the chemical behavior of an element, based on the row and column in which it is found, makes the periodic table an indispensable reference tool for scientists. Open a chemistry textbook and the chances are that there will be a periodic table, often in bright colors, printed on the inside cover of the book. Its constant use by chemists emphasizes the central role the periodic table plays in making sense out of what otherwise might be a chaotic jumble of facts about the elements and their many molecular combinations.

The chemical group of each of the elements described in the sections that follow is listed directly below the chemical symbol of the element. The similarity of chemical properties of elements in the same group should be as apparent to you as it was to Mendeleyev or to any chemist who uses the periodic table for information and research.