

A Brief History of the Periodic Table

You can find some version of the periodic table in every chemistry classroom or lab in the world. Without a doubt, nothing else quite like it exists in any other academic discipline.

"The periodic table of the elements is one of the most powerful tools in science; a single document that merges our knowledge of chemistry."

The story of the periodic system for classifying the elements can be traced back over 200 years. Throughout its long history, the periodic table has been disputed, altered and improved as science has progressed and as new elements have been discovered. But despite the dramatic changes that have taken place in science over the past century, namely the development of the theories of relativity and quantum mechanics, there has been no revolution in the basic nature of the periodic table. In some instances, new findings initially appeared to call into question the theoretical foundations of the periodic table, but each time scientists eventually managed to incorporate the results while preserving the table's fundamental purpose. Remarkably, the periodic table is noted both for its historical roots and for its modern relevance.

The term "periodic" reflects the fact that the elements show patterns in their chemical properties in certain regular intervals; were it not for the simplification provided by this chart, students of chemistry would need to learn the properties of all 116 known elements. Fortunately, the periodic table allows chemists to function by mastering the properties of a handful of typical elements; all the others fall into groups or families with similar chemical properties.

The discovery of the periodic law for classifying the elements represents the culmination of a number of scientific developments, rather than a sudden brainstorm on the part of one individual. Yet historians typically consider one event as marking the formal birth of the modern periodic table: on February 17, 1869, a Russian professor of chemistry; Dimitri Mendeleev; completed the first of his numerous periodic charts while writing a textbook of chemistry. It included 63 known elements arranged according to increasing atomic weight; Mendeleev also left spaces for as yet undiscovered elements for which he predicted atomic weights. Prior to Mendeleev's discovery, however; other scientists had been actively developing some kind of organizing system to describe the elements.

In 1787, for example, French chemist **Antoine Lavoisier**, working with Antoine Fourcroy, Louis-Bernard Guyton de Morveau and Claude-Louis Berthollet, **devised a list of the 33 elements known at the time**. Yet such lists were simply one-dimensional representations. The power of the modern table lies in its two- or even three-dimensional display of all the known elements (and even the ones yet to be discovered) in a logical system of precisely ordered rows and columns.

In an early attempt to organize the elements into a meaningful array, German chemist **Johann Dobereiner (1817)** **pointed out that many of the known elements could be arranged by their similarities into groups of three, which he called triads**. Dobereiner singled out triads of the elements lithium, sodium and potassium as well as chlorine, bromine and iodine. He noticed that if the three members of a triad were ordered according to their atomic weights, the properties of the middle element fell in between those of the first and third elements. For example, lithium, sodium and potassium all react vigorously with water. But lithium, the lightest of the triad, reacts more mildly than the other two, whereas the heaviest of the three, potassium, explodes violently. In addition, Dobereiner showed that the atomic weight of the middle element is close to the average of the weights for the first and third members of the triad. Dobereiner's work encouraged others to search for correlations between the chemical properties of the elements and their atomic weights. One of those who pursued the triad approach further during the 19th century was Peter Kremers of Cologne, who suggested that certain elements could belong to two triads placed perpendicularly. Kremers thus broke new ground by comparing elements in two directions, a feature that later proved to be an essential aspect of Mendeleev's system.

In 1857 French chemist Jean-Baptiste-Andre Dumas turned away from the idea of triads and focused instead on devising a set of mathematical equations that could account for the increase in atomic weight among several groups of chemically similar elements. But as chemists now recognize, any attempt to establish an organizing pattern based on an element's atomic weight will not succeed, because atomic weight is not the fundamental property that characterizes each of the elements.

Periodic Properties

The crucial characteristic of Mendeleev's system was that it illustrated a periodicity, or repetition, in the properties of the elements at certain regular intervals. This feature had been observed previously in an arrangement of elements by atomic weight devised by French geologist Alexandre-Emile Beguyer de Chancourtois (1862). The system relied on a fairly intricate geometric configuration; de Chancourtois positioned the elements according to increasing atomic weight along a spiral inscribed on the surface of a cylinder and inclined at 45 degrees from the base. The first full turn of the spiral coincided with the element oxygen, and the second full turn occurred at sulfur. Elements that lined up vertically on the surface of the cylinder tended to have similar properties, so this arrangement succeeded in capturing

some of the patterns that would later become central to Mendeleev's system. Yet for a number of reasons, de Chancourtois's system did not have much effect on scientists of the time. His original article failed to include a diagram of the table, the system was too complicated, and the chemical similarities among elements were not displayed very well.

Several other researchers put forward their own versions of a periodic table during the 1860s. Using newly standardized values for atomic weights, English chemist John Newlands suggested in 1864 that when the elements were arranged in order of atomic weight, any one of the elements showed properties similar to those of the elements eight places ahead and eight places behind in the list-a-feature that Newlands called "the law of octaves". In his original table, Newlands left empty spaces for missing elements, but his more publicized version of 1866 did not include these open slots. Other chemists immediately raised objections to the table because it would not be able to accommodate any new elements that might be discovered. In fact, some investigators openly ridiculed Newlands's ideas. At a meeting of the Chemical Society in London in 1866, George Carey Foster of University College London asked Newlands whether he had considered ordering the elements alphabetically, because any kind of arrangement would present occasional coincidences. As a result of the meeting, the Chemical Society refused to publish Newlands's paper. Despite its poor reception, Newlands's work was the first time anyone used a sequence of ordinal numbers (in this case, one based on the sequence of atomic weights) to organize the elements. In this respect, Newlands anticipated the modern organization of the periodic table, which is based on the sequence of so-called atomic numbers.

The Modern Periodic Table

Around the same time, Mendeleev assembled his own periodic table, chemist Julius Lothar Meyer of Breslau University in Germany, in the process of revising his chemistry textbook, produced a periodic table that turned out to be remarkably similar to Mendeleev's famous 1869 version; although Meyer failed to classify all the elements correctly. But, the table did not appear in print until 1870 because of a publisher's delay, a factor that contributed to a bitter dispute for priority that ensued between Meyer and Mendeleev.

Unlike his predecessors, Mendeleev had sufficient confidence in his periodic table to use it to predict several new elements and the properties of their compounds. He also corrected the atomic weights of some already known elements. Interestingly, Mendeleev admitted to having seen certain earlier tables, such as those of Newlands, but claimed to have been unaware of Lothar Meyer's work when developing his chart. Although the predictive aspect of Mendeleev's table was a major advance, it seems to have been overemphasized by historians, who have generally suggested that Mendeleev's table was accepted especially because of this feature. These scholars have failed to notice that the citation from the

Royal Society of London that accompanied the Davy Medal (which Mendeleev received in 1882) makes no mention whatsoever of his predictions. Instead Mendeleev's ability to accommodate the already known elements may have contributed as much to the acceptance of the periodic system as did his striking predictions. Although numerous scientists helped to develop the periodic system, Mendeleev receives most of the credit for discovering chemical periodicity because he elevated the discovery to a law of nature and spent the rest of his life boldly examining its consequences and defending its validity.

Defending the periodic table was no simple task. Its accuracy was frequently challenged by subsequent discoveries. One notable occasion arose in 1894, when William Ramsay of University College London and Lord Rayleigh (John William Strutt) of the Royal Institution in London discovered the element argon; over the next few years, Ramsay announced the identification of four other elements helium, neon, krypton and xenon-known as the noble gases. The name "noble" derives from the fact that all these gases seem to stand apart from the other elements, rarely interacting with them to form compounds. As a result, some chemists suggested that the noble gases did not even belong in the periodic table. Mendeleev or anyone else had not predicted these elements, and only after six years of intense effort could chemists and physicists successfully incorporate the noble gases into the table. In the new arrangement, an additional column was introduced between the halogens (the gaseous elements fluorine, chlorine, bromine, iodine and astatine) and the alkali metals (lithium, sodium, potassium, rubidium, cesium and francium).

A second point of contention surrounded the precise ordering of the elements. Mendeleev's original table positioned the elements according to atomic weight, but in 1913 Dutch amateur theoretical physicist Anton van den Broek suggested that the ordering principle for the periodic table lay instead in the nuclear charge of each atom. Physicist Henry Moseley, working at the University of Manchester, tested this hypothesis, also in 1913, shortly before his tragic death in World War I. Moseley began by photographing the x-ray spectrum of 12 elements, 10 of which occupied consecutive places in the periodic table. He discovered that the frequencies of features called K-lines in the spectrum of each element were directly proportional to the squares of the integers representing the position of each successive element in the table. As Moseley put it, here was proof that "there is in the atom a fundamental quantity, which increases by regular steps as we pass from one element to the next". This fundamental quantity first referred to as atomic number in 1920 by Ernest Rutherford, who was then at the University of Cambridge, is now identified as the number of protons in the nucleus.

Moseley's work provided a method that could be used to determine exactly how many empty spaces remained in the periodic table. After this discovery, chemists turned to using atomic number as the fundamental ordering principle for the periodic table, instead of atomic weight. This change resolved many of the lingering problems in the arrangement of the elements. For example, when iodine and

tellurium were ordered according to atomic weight (with iodine first), the two elements appeared to be incorrectly positioned in terms of their chemical behavior. When ordered according to atomic number (with tellurium first), however, the two elements were in their correct positions.

Understanding the Atom

The periodic table inspired the work not only of chemists but also of atomic physicists struggling to understand the structure of the atom. In 1904, while working at Cambridge, physicist J. J. Thomson (who also proved the existence of the electron) developed a model of the atom, paying close attention to the periodicity of the elements. He proposed that the atoms of a particular element contained a specific number of electrons arranged in concentric rings. Furthermore, according to Thomson, elements with similar configurations of electrons would have similar properties; Thomson's work thus provided the first physical explanation for the periodicity of the elements. Although Thomson imagined the rings of electrons as lying inside the main body of the atom, rather than circulating around the nucleus as is believed today, his model does represent the first time anyone addressed the arrangement of electrons in the atom, a concept that pervades the whole of modern chemistry.

Danish physicist Niels Bohr, the first to bring quantum theory to bear on the structure of the atom, was also motivated by the arrangement of the elements in the periodic system. In Bohr's model of the atom, developed in 1913, electrons inhabit a series of concentric shells that encircle the nucleus. Bohr reasoned that elements in the same group of the periodic table might have identical configurations of electrons in their outermost shell and that the chemical properties of an element would depend in large part on the arrangement of electrons in the outer shell of its atoms. Bohr's model of the atom also served to explain why the noble gases lack reactivity: noble gases possess full outer shells of electrons, making them unusually stable and unlikely to form compounds. Indeed, most other elements form compounds as a way to obtain full outer electron shells. More recent analysis of how Bohr arrived at these electronic configurations suggests that he functioned more like a chemist than has generally been credited. Bohr did not derive electron configurations from quantum theory but obtained them from the known chemical and spectroscopic properties of the elements.

In 1924, another physicist Austrian-born Wolfgang Pauli, set out to explain the length of each row, or period, in the table. As a result, he developed the Pauli exclusion principle, which states that no two electrons can exist in exactly the same quantum state, which is defined by what scientists call quantum numbers. The lengths of the various periods emerge from experimental evidence about the order of electron-shell filling and from the quantum-mechanical restrictions on the four quantum numbers that electrons can adopt. The modifications to quantum theory made by Werner Heisenberg and Erwin Schrodinger in the mid-1920s yielded quantum mechanics in essentially the form used to this day.

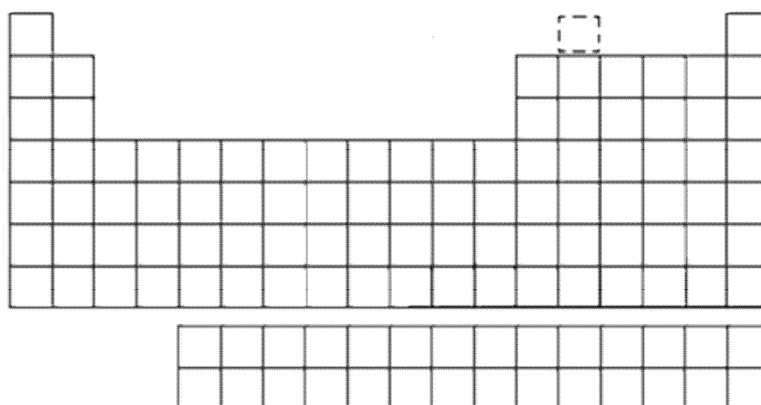
But the influence of these changes on the periodic table has been rather minimal. Despite the efforts of many physicists and chemists, quantum mechanics cannot explain the periodic table any further. For example, it cannot explain from first principles the order in which electrons fill the various electron shells. The electronic configurations of atoms, on which our modern understanding of the periodic table is based, cannot be derived using quantum mechanics (this is because the fundamental equation of quantum mechanics, the Schrodinger equation, cannot be solved exactly for atoms other than hydrogen). As a result, quantum mechanics can only reproduce Mendeleev's original discovery by the use of mathematical approximations-it cannot predict the periodic system.

The Periodic law

The periodic law was the basis for creating the periodic table but Mendeleev had no idea about the nucleus, proton or neutron. Therefore, it was natural that he would support the periodicity of the elements on the atomic mass rather than the positive charge (atomic number) alone. It was Moseley who stated the modern version we use today.

Modern Periodic Law: The physical and chemical properties of the elements are periodic functions of the atomic number (positive charge).

Notes



“To deny, to believe, and to doubt absolutely -- this is for man what running is for a horse”. - Blaise Pascal