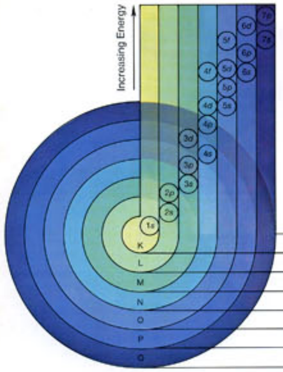
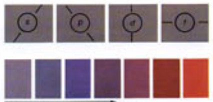
The Periodic Table of the Elements





Increasing Electronegativity

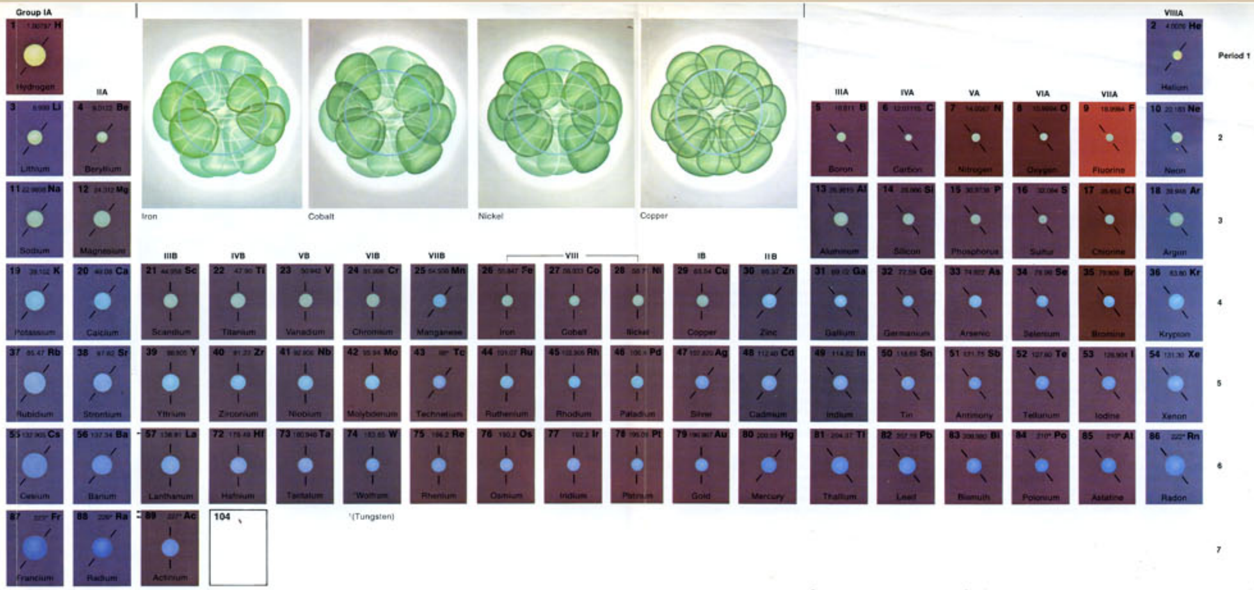
Electronegativity Scale. The increase in color intensity from left to right denotes the relative increase in electronegativity. The gray color on the extreme left indicates values of electronegativity that are as yet undetermined. The gray color to its right indicates that the element is not electronegative (as in Group VIIIA, the noble gases).

Principal Quantum Number	Shell Capacity (Electrons)	Orbital Types	Number of Electron Pairs
1	2	1s	1
2	8	2s, 2p	1, 3
3	18	3s, 3p, 3d	1, 3, 5
4	32	4s, 4p, 4d, 4f	1, 3, 5, 7
5	32	5s, 5p, 5d, 5f	1, 3, 5, 7
6	18	6s, 6p, 6d	1, 3, 5
7	8	7s, 7p	1, 3

This color wheel is the visual key to the table. Successive shells are shown from K to Q. These shells are related to principal quantum numbers and to other characteristics shown in the table to the right. The spheres in the periodic table that depict relative atomic sizes are color-keyed back to this wheel. The sphere color indicates which shell is being filled by electrons. The graph above the wheel is an energy level diagram, which shows the order of orbital filling. The various orbitals within each shell are indicated.

The sequence for the filling of the *s* and *p* orbitals from lithium to neon is shown below. Here, electrons are indicated symbolically. The *p* orbitals are considerably less complex than the *d* orbitals.

The sequence of electron filling for the transition elements is shown above the main table. Orbitals that possess an electron pair are depicted as a darker green, whereas those that possess only one electron per orbital are depicted as a lighter green. The sequence of iron, cobalt, and nickel shows that there is an outer 4s orbital (blue) which already has two electrons in it while the complex inner 3d orbitals are being filled. In copper, all the 3d orbitals are filled while there is only one electron in the 4s orbital. It should be noted that these images represent the average probability of finding an electron occupying the volume in space for a specific orbital.



*denotes most stable or best known isotope

Lithium

Beryllium

Boron

Carbon

Nitrogen

Oxygen

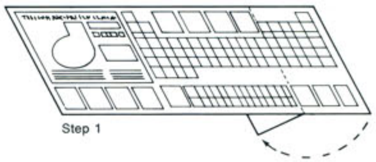
Fluorine

Neon

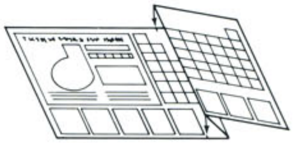
The Periodic Table of the Elements

The periodic table condenses a great store of information about physics and chemistry that enables us to understand various relationships among different elements. This information is based on the periodic law, which states that the properties of elements are periodic functions of their atomic numbers. It is this periodicity that provides a basis for organizing the chemical and physical properties of the elements.

When elements are listed together on the basis of increasing atomic number, elements having similar chemical and physical properties appear at regular intervals in the listing. More than a century ago Dmitri Mendeleev worked out a scheme for showing the periodicity of the elements, but his scheme was based on atomic mass, not on atomic number. The large foldout table presented here is a modified version of Mendeleev's original layout.



Step 1



Step 2

To distinguish the transition elements from the representative families, fold along the leading edge of Group IIIA (Step 1). After the fold has been made, bring the two short line marks together (Step 2).

Conceptual Design: Payson R. Stevens

Traditional tables have relied on the printed word to describe periodic relationships. In contrast, this version provides a series of visual reference points for the main elemental properties in order to emphasize this periodicity. The large table lists elements with similar properties in vertical columns, or *groups*; elements having dissimilar properties are listed in horizontal rows, or *periods*. The arrangement shows the order in which atoms lose or gain electrons, which is called their *order of activity*. It is on the basis of this ordering that we classify elements as "metals," "non-metals" (including inert gases), and "semi-metals" (or *transition elements*). When the large table is folded as indicated, the transition elements are separated from the other, more representative elements.

The more active the element, the more intense the color used to shade the square for that element in the table. From this shading it becomes evident that the most active metal (the one with the strongest tendency to lose electrons in chemical reactions) falls in the lower left-hand corner of the table, and that the most active nonmetal (the one with the strongest tendency to acquire electrons) falls in the upper right-hand corner. This gradation is referred to as an "electronegativity" scale, and it provides a qualitative way of describing the strength with which atoms bond to each other. Generally, the greater the difference between electronegativities of two elements, the greater the strength of the bond between them.

What is the physical basis for the periodicity seen in this table? Each atom has only certain allowed energy levels, which are designated by n—a whole number (or *principal quantum number*) in a set starting with $n = 1$. But within each set of n are further divisions. For example, 1*s* stands for the lowest energy level $n = 1$; and 2*s* and 2*p* stands for the two energy levels corresponding to $n = 2$. These divisions contain "hidden" energy states. For example, every *s* level contains two states; every *p* level contains six states; every *d* level, ten states, and so on. According to the Pauli exclusion principle, there can be no more than one electron in each of these states. On the average, *all* of the electrons with the same principal quantum number n are roughly the same distance away from the nucleus, and all have similar energy values. Such electrons are said to belong to the same electron *shell*.

The schematic color wheel is a visual key to the successive ordering of these shells. Each shell is given a letter (K, L, M, N, O, . . .) to correspond with a principal quantum number (1, 2, 3, 4, 5, . . .). The table above the wheel shows the order of filling for the orbitals, based on quantum mechanics. Each *s*, *p*, *f*, and *d* subshell is indicated by a slash mark. Using this slash mark, the table can be examined in an orderly way. The group A elements (or representative families) can be divided into two *s* and *p* blocks on the left- and right-hand side of the table. The transition elements show *d*-orbital filling, and the inner transition elements, *f*-subshell filling. Usually in these last two groups, inner *d* and *f* subshells are being filled even though electrons already occupy orbital positions in the next outer subshell. Some of the outer subshells are associated with lower energies, and the atom can achieve a more stable configuration by filling these shells first.

The energy an electron has in a given shell also depends somewhat on the *orbital quantum number* l, which describes the *magnitude* of the electron's angular (or rotational) momentum. The smaller the orbital quantum number, the closer to the nucleus the electron is likely to be, and the lower its total energy. Like linear momentum, angular momentum is a vector quantity and to fully describe it we must give its *direction* as well. The orientatior of these orbitals affects the way atoms combine into molecules. An electron subshell consists of all the electrons having the same principal quantum number n and the same orbital quantum number l.

When a given shell or subshell contains all the electron it can hold, it is said to be *closed*. And whether a shell or subshell is "open" or closed relates to the order of activity found in the periodic table. When an atom has only closed shells or subshells, its electric charge is distributed uniformly. It does not attract additional electrons nor can its electrons be detached very easily. (This behavior is characteristic of the "passive" inert gases.) When an atom has only one electron in its outermost shell it tends to lose that electron because the negative electro is far away from the attractive pull of the positive charge of the nucleus. (Alkali metals fall in this category.) And when an atom lacks only one electron in an otherwise filled shell or subshell, it tends to acquire that electron because its nuclear attraction is not completely "sealed off" by the innermost electrons. (This behavior is characteristic of the voracious halogens.)

These patterns are quite important, because the electron configuration of an element determines its properties Interpretation of these configurations enables, for example, an understanding of crystal structure, electric properties or conductivity, oxidation state, relative atomic size and electronegativity. One example is the relationship between the valence number of an element (a measure of the combining power of elements with each other) and the position of the element within the table. There is a strong similarity between elements in the same vertical group: they all have the same valence and exhibit similar chemical properties even though they have different quantum numbers. There is even some resemblance between the properties of elements in different groups but having the same number of valence electrons.

When you look at the relative atomic *size* of the elements, another pattern emerges. As atomic number increases within a group, there is an increase in atomic size. But within a given horizontal row, or period, atomic size decreases with an increase in atomic number. This behavior is also explainable in terms of the changes that take place in electronic structure. Within a group, an increase in atomic number means there is an increase in the principal quantum number and a corresponding increase in the distance between the outermost electrons and the nucleus. Within a given period, the principal quantum number remains the same, but the increased nuclear charge exerts a stronger force on the outermost electrons and keeps them closer to the nucleus.

These are just a few examples that show how the electronic structure of the elements is related to their chemicɛ and physical properties. Once these basic relationships are understood, the table may also be used to explore the properties that matter possesses, in both elemental and compound forms.