The periodic table, electron shells, and orbitals

The Bohr model and atomic orbitals. Using an element's position in the periodic table to predict its properties, electron configuration, and reactivity.

Introduction

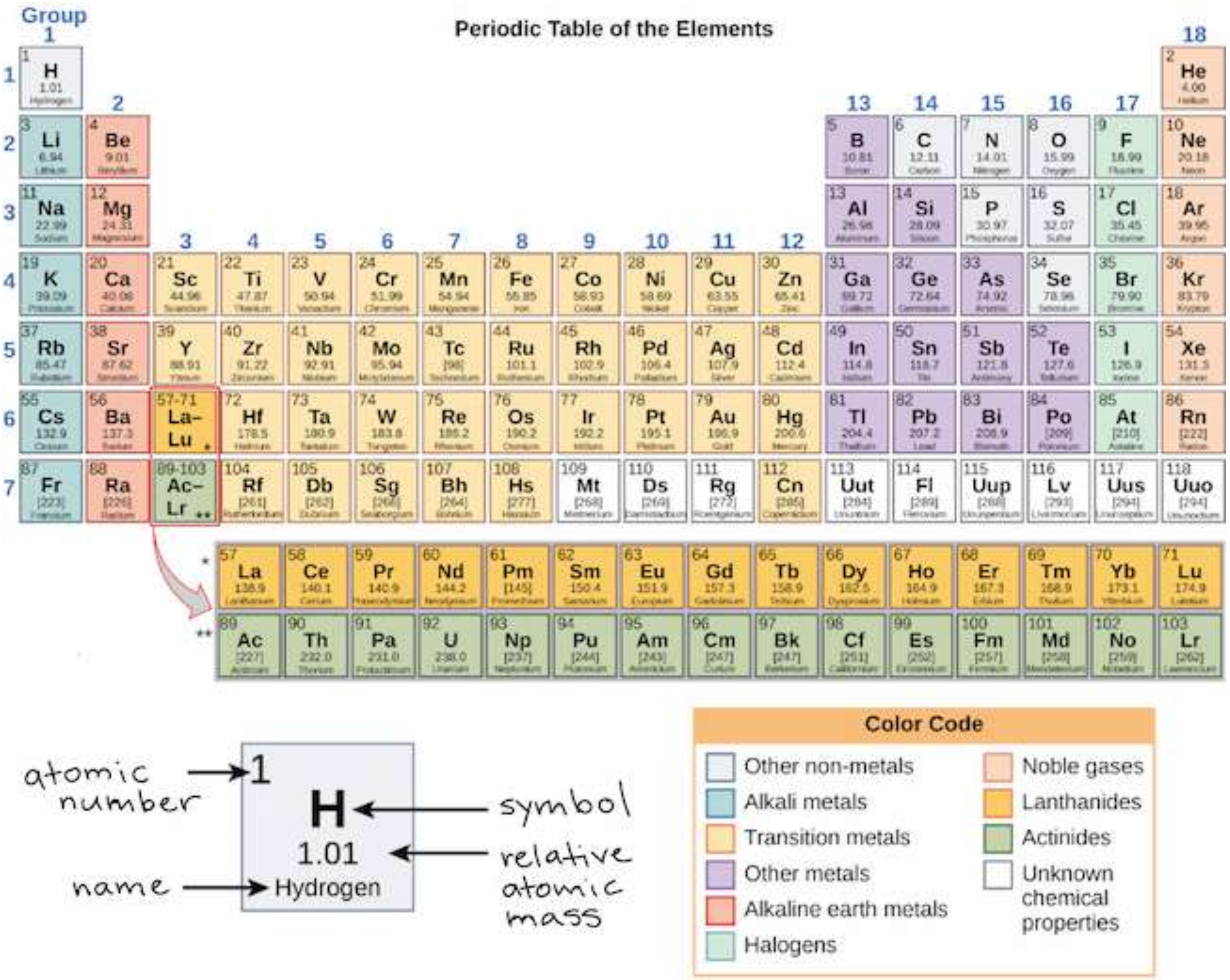
At some point in your chemistry education, you may have been introduced to the song “The Elements,” in which Tom Lehrer does a rapid-fire musical rendition of all the elements' names. Like me, you may even have been offered the opportunity to memorize this song for extra credit. If so, it’s possible that you still remember the names of all the elements, which is an impressive feat—not to mention a fun trick to pull out at cocktail parties.

If you’ve memorized the names of the elements, does that mean you’ll never need a periodic table again? Well ... probably not. That’s because the periodic table isn’t just a big bucket that holds all of the elements. Instead, it’s more like a filing system. The position of each element in the table gives important information about its structure, properties, and behavior in chemical reactions. Specifically, an element’s position in the periodic table helps you figure out its electron configuration, how the electrons are organized around the nucleus. Atoms use their electrons to participate in chemical reactions, so knowing an element’s electron configuration allows you to predict its reactivity—whether, and how, it will interact with atoms of other elements.
In this article, we’ll look in more detail at the periodic table, how atoms organize their electrons, and how this allows us to predict the reactivity of elements.

The periodic table

By convention, elements are organized in the **periodic table**, a structure that captures important patterns in their behavior. Devised by Russian chemist Dmitri Mendeleev (1834–1907) in 1869, the table places elements into columns—**groups**—and rows—**periods**—that share certain properties. These properties determine an element’s physical state at room temperature—gas, solid, or liquid)—as well as its **chemical reactivity**, the ability to form chemical bonds with other atoms.

In addition to listing the atomic number for each element, the periodic table also displays the element’s relative atomic mass, the weighted average for its naturally occurring isotopes on earth. Looking at hydrogen, for example, its symbol, H appear, as well as its atomic number of one—in the upper left-hand corner—and its relative atomic mass of 1.01.



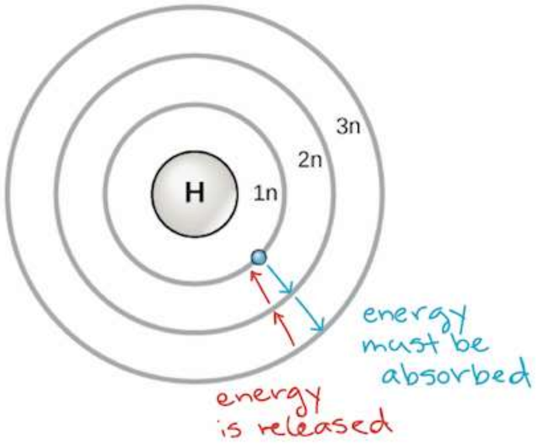
The periodic table of the elements

Image credit: modified from OpenStax Biology

Differences in chemical reactivity between elements are based on the number and spatial distribution of their electrons. If two atoms have complementary electron patterns, they can react and form a chemical bond, creating a molecule or compound. As we will see below, the periodic table organizes elements in a way that reflects their number and pattern of electrons, which makes it useful for predicting the reactivity of an element: how likely it is to form bonds, and with which other elements.

Electron shells and the Bohr model

An early model of the atom was developed in 1913 by the Danish scientist Niels Bohr (1885–1962). The Bohr model shows the atom as a central nucleus containing protons and neutrons, with the electrons in circular electron shells at specific distances from the nucleus, similar to planets orbiting around the sun. Each electron shell has a different energy level, with those shells closest to the nucleus being lower in energy than those farther from the nucleus. By convention, each shell is assigned a number and the symbol n—for example, the electron shell closest to the nucleus is called 1n. In order to move between shells, an electron must absorb or release an amount of energy corresponding exactly to the difference in energy between the shells. For instance, if an electron absorbs energy from a photon, it may become excited and move to a higher-energy shell; conversely, when an excited electron drops back down to a lower-energy shell, it will release energy, often in the form of heat.

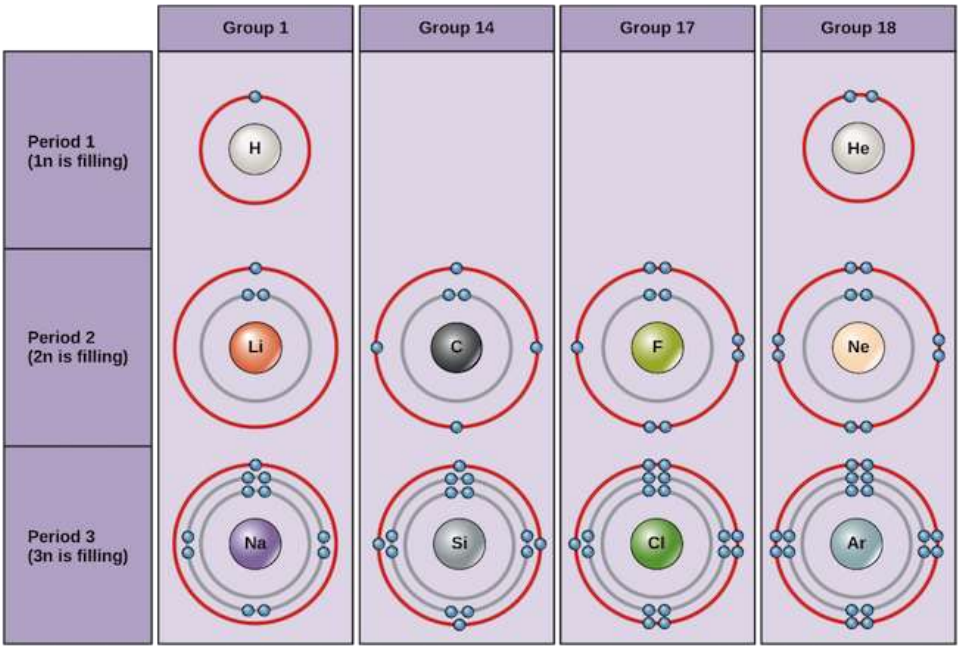


Bohr model of an atom, showing energy levels as concentric circles surrounding the nucleus. Energy must be added to move an electron outward to a higher energy level, and energy is released when an electron falls down from a higher energy level to a closer-in one.

Image credit: modified from OpenStax Biology

Atoms, like other things governed by the laws of physics, tend to take on the lowest-energy, most stable configuration they can. Thus, the electron shells of an atom are populated from the inside out, with electrons filling up the low-energy shells closer to the nucleus before they move into the higher-energy shells further out. The shell closest to the nucleus, 1n, can hold two electrons, while the next shell, 2n, can hold eight, and the third shell, 3n, can hold up to eighteen.
The number of electrons in the outermost shell of a particular atom determines its reactivity, or tendency to form chemical bonds with other atoms. This outermost shell is known as the **valence shell**, and the electrons found in it are called **valence electrons**. In general, atoms are most stable, least reactive, when their outermost electron shell is full. Most of the elements important in biology need eight electrons in their outermost shell in order to be stable, and this rule of thumb is known as the **octet rule**. Some atoms can be stable with an octet even though their valence shell is the 3n shell, which can hold up to 18 electrons. We will explore the reason for this when we discuss electron orbitals below.

Examples of some neutral atoms and their electron configurations are shown below. In this table, you can see that helium has a full valence shell, with two electrons in its first and only, 1n, shell. Similarly, neon has a complete outer 2n shell containing eight electrons. These electron configurations make helium and neon very stable. Although argon does not technically have a full outer shell, since the 3n shell can hold up to eighteen electrons, it is stable like neon and helium because it has eight electrons in the 3n shell and thus satisfies the octet rule. In contrast, chlorine has only seven electrons in its outermost shell, while sodium has just one. These patterns do not fill the outermost shell or satisfy the octet rule, making chlorine and sodium reactive, eager to gain or lose electrons to reach a more stable configuration.



Bohr diagrams of various elements

Image credit: OpenStax Biology

Electron configurations and the periodic table

Elements are placed in order on the periodic table based on their atomic number, how many protons they have. In a neutral atom, the number of electrons will equal the number of protons, so we can easily determine electron number from atomic number. In addition, the position of an element in the periodic table—its column, or group, and row, or period—provides useful information about how those electrons are arranged.

If we consider just the first three rows of the table, which include the major elements important to life, each row corresponds to the filling of a different electron shell: helium and hydrogen place their electrons in the 1n shell, while second-row elements like Li start filling the 2n shell, and third-row elements like Na continue with the 3n shell. Similarly, an element's column number gives information about its number of valence electrons and reactivity. In general, the number of valence electrons is the same within a column and increases from left to right within a row. Group 1 elements have just one valence electron and group 18 elements have eight, except for helium, which has only two electrons total. Thus, group number is a good predictor of how reactive each element will be:

- Helium (He), neon (Ne), and argon (Ar), as group 18 elements, have outer electron shells that are full or satisfy the octet rule. This makes them highly stable as single atoms. Because of their non-reactivity, they are called the **inert gases** or **noble gases**.
- Hydrogen (H), lithium (Li), and sodium (Na), as group 1 elements, have just one electron in their outermost shells. They are unstable as single atoms, but can become stable by losing or sharing their one valence electron. If these elements fully lose an electron—as Li and Na typically do—they become positively charged ions: Li^+ and Na^+.
- Fluorine (F) and chlorine (Cl), as group 17 elements, have seven electrons in their outermost shells. They tend to achieve a stable octet by taking an electron from other atoms, becoming negatively charged ions: F^-, start superscript, minus, end superscript and Cl^-.

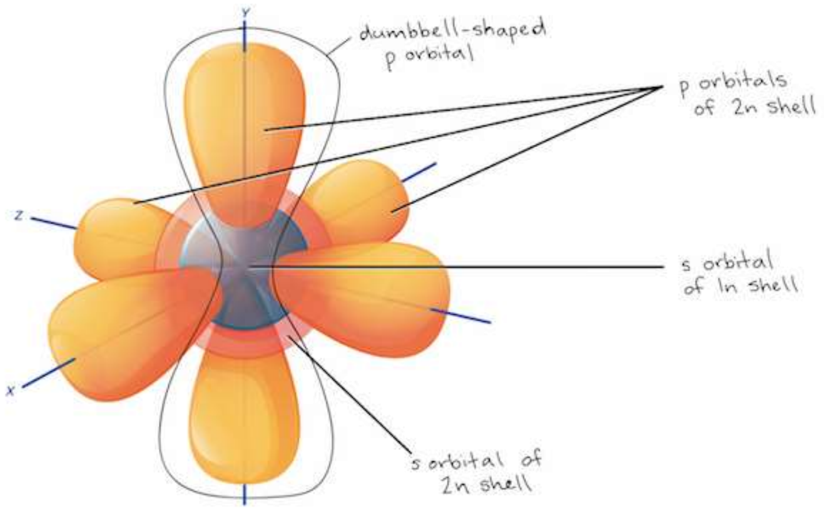
- Carbon (C), as a group 14 element, has four electrons in its outer shell. Carbon typically shares electrons to achieve a complete valence shell, forming bonds with multiple other atoms.

Thus, the columns of the periodic table reflect the number of electrons found in each element's valence shell, which in turn determines how the element will react.

Subshells and orbitals

The Bohr model is useful to explain the reactivity and chemical bonding of many elements, but it actually doesn't give a very accurate description of how electrons are distributed in space around the nucleus. Specifically, electrons don't really circle the nucleus, but rather spend most of their time in sometimes-complex-shaped regions of space around the nucleus, known as **electron orbitals**. We can't actually know where an electron is at any given moment in time, but we can mathematically determine the volume of space in which it is most likely to be found—say, the volume of space in which it will spend 90% of its time. This high-probability region makes up an orbital, and each orbital can hold up to two electrons.

So, how do these mathematically defined orbitals fit in with the electron shells we saw in the Bohr model? We can break each electron shell down into one or more subshells, which are simply sets of one or more orbitals. Subshells are designated by the letters ssss, pppp, dddd, and ffff, and each letter indicates a different shape. For instance, ssss subshells have a single, spherical orbital, while pppp subshells contain three dumbbell-shaped orbitals at right angles to each other. Most of organic chemistry—the chemistry of carbon-containing compounds, which are central to biology—involves interactions between electrons in ssss and pppp subshells, so these are the most important subshell types to be familiar with. However, atoms with many electrons may place some of their electrons in dddd and ffff subshells. Subshells dddd and ffff have more complex shapes and contain five and seven orbitals, respectively.



3D diagram of circular 1s and 2s orbitals and dumbbell-shaped 2p orbitals. There are three 2p orbitals, and they are at right angles to each other.

Image credit: modified from OpenStax Biology

The first electron shell, 1n, corresponds to a single 1s orbital. The 1s orbital is the closest orbital to the nucleus, and it fills with electrons first, before any other orbital. Hydrogen has just one electron, so it has a single spot in the 1s orbital occupied. This can be written out in a shorthand form called an **electron configuration** as $1s^1$ where the superscripted 1 refers to the one electron in the $1s$ orbital. Helium has two electrons, so it can completely fill the $1s$ orbital with its two electrons. This is written out as $1s^2$ referring to the two electrons of helium in the $1s$ orbital. On the periodic table, hydrogen and helium are the only two elements in the first row, or period, which reflects that they only have electrons in their first shell. Hydrogen and helium are the only two elements that have electrons exclusively in the $1s$ orbital in their neutral, non-charged, state. The second electron shell, 2n, contains another spherical s orbital plus three dumbbell-shaped p orbitals, each of which can hold two electrons. After the $1s$ orbital is filled, the second electron shell begins to fill, with electrons going first into the $2s$ orbital and then into the three p orbitals. Elements in the second row of the periodic table place their electrons in the 2n shell as well as the 1n shell. For instance, lithium (Li) has three electrons: two fill the $1s$ orbital, and the third is placed in the $2s2s2s2$, s orbital, giving an electron configuration of $1s^2\ 2s^1$. Neon (Ne), on the other hand, has a total of ten electrons: two are in its innermost $1s$ orbital and eight fill the second shell—two each in the $2s$ and three p orbitals, $1s^2\ 2s^2\ 2p^6$. Because its 2n shell is filled, it is energetically stable as a single atom and will rarely form chemical bonds with other atoms.

The third electron shell, 3n, also contains an s orbital and three p orbitals, and the third-row elements of the periodic table place their electrons in these orbitals, much as second-row elements do for the 2n shell. The 3n shell also contains a d orbital, but this orbital is considerably higher in energy than the $3s$ and $3p$ orbitals and does not begin to fill until the fourth row of the periodic table. This is why third-row elements, such as argon, can be stable with just eight valence electrons: their s and p subshells are filled, even though the entire 3n shell is not. While electron shells and orbitals are closely related, orbitals provide a more accurate picture of the electron configuration of an atom. That’s because orbitals actually specify the shape and position of the regions of space that electrons occupy.