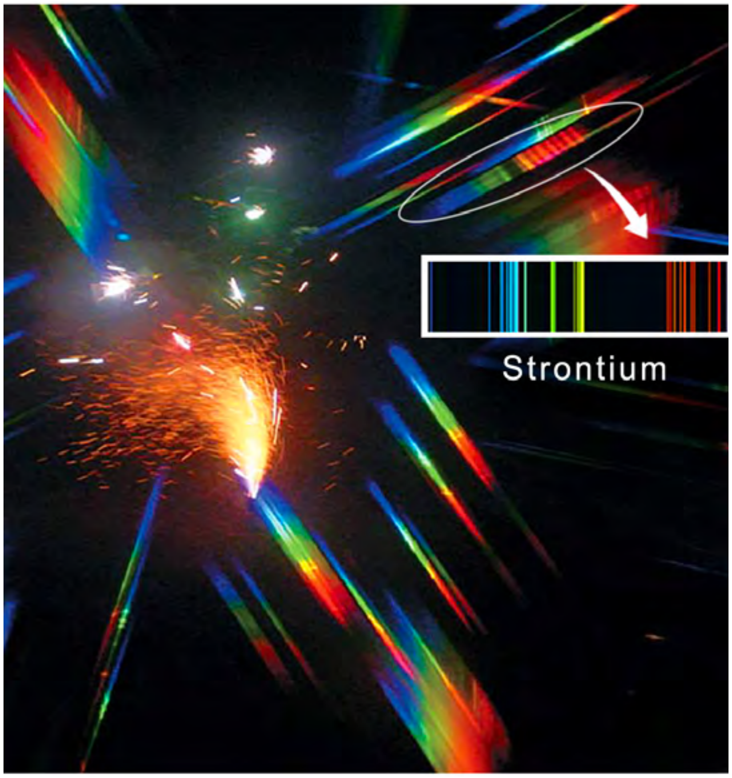


Chapter 4

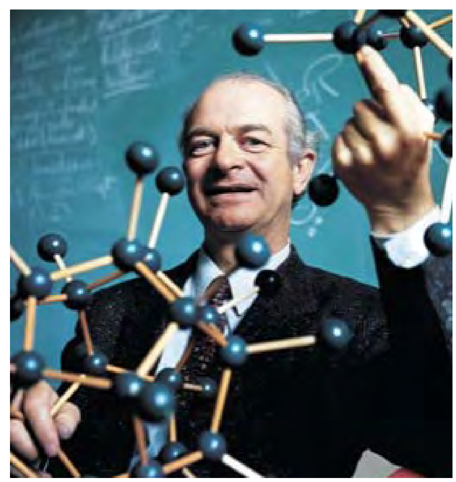
Subatomic Particles

THE MAIN IDEA

Atoms are made of electrons, protons, and neutrons

4.9 The Shell Model and Periodic Table

Chemistry is the study of how atoms combine to form new materials. But how exactly does this happen? In order to answer this question, we first need to understand a bit about the properties of the atoms themselves. For this, we turn to the shell model, made popular by the noted chemistry and two-time Nobel laureate Linus Pauling (1901-1994). This model, described in detail in the previous section, is similar to Bohr's planetary model in that it shows electrons restricted to particular distances from the nucleus. The shell model, however, is a bit more sophisticated because it incorporates the wave nature of electrons.



Two-time Nobel laureate Linus Pauling (1901–1994) was an early proponent of teaching beginning chemistry students a shell model, from which the organization of the periodic table could be described. In 1954 Pauling won the Nobel Prize in Chemistry for his research into the nature of the chemical bond. In 1962 he was awarded the Nobel Peace Prize for his campaign against the testing of nuclear bombs, which was introducing massive amounts of radioactivity into the environment.

First, let's take some time to review the three main points of the previous section:

Point 1

According to the shell model, electrons behave as though they are arranged in a series of concentric shells. A shell is a region of space around the atomic nucleus within which electrons may reside. An important aspect of this model is there are at least seven shells and each shell can hold only a limited number of electrons. As was shown in Figure 4.32, the innermost shell can hold 2 electrons, the second and third shells 8 each, the fourth and fifth shells 18 each, and the sixth and seventh shells 32 each.

Point 2

A series of seven such concentric shells accounts for the seven periods of the periodic table. Furthermore, the number of elements in each period is equal to the shell's capacity for electrons. The first shell, for example, has the capacity for only two electrons. That's why we find only two elements, hydrogen and helium, in the first period, as was shown in Figure 4.33. Hydrogen is the element whose atoms have only one electron. This one electron resides within the first shell, which is the shell closest to the nucleus. Each helium atom has two electrons, both of which are also within the first shell, which is thus filled to its maximum capacity. Similarly, the second and third shells each have the capacity for eight electrons, so eight elements are found in both the second and third periods.

Point 3

The electrons of the outermost occupied shell in any atom are directly exposed to the external environment and are the first to interact with other atoms. Most notably, they are the ones that participate in chemical bonding, as we will discuss in Chapter 6. The electrons in the outermost shell, therefore, are quite important.

The quality of a song depends on the arrangement of musical notes. In a similar fashion, the properties of an element depend on the arrangements of electrons in its atoms, especially the outer-shell electrons. Look carefully at Figure 4.33 and you'll see that the outer-shell electrons of atoms above and below one another on the periodic table (within the same group) are similarly organized. This explains why elements of the same group have similar properties—a concept first introduced in Section 3.3.

CONCEPT **CHECK**

Do atoms really consist of shells that look like those depicted in Figure 4.32?

CHECK YOUR ANSWER

No. The shell model is not a depiction of the "appearance of an atom." Rather, it is a conceptual model that allows us to account for observed behavior. An atom, therefore, does not actually contain a series of concentric shells; it merely behaves as though it does.

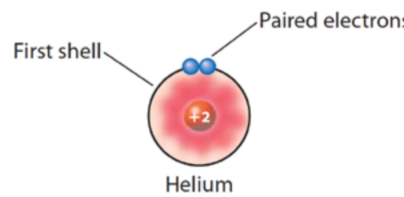
Effective Nuclear Charge

Recall from Section 3.3 that properties of elements gradually change across the periodic table. The atoms of elements toward the upper right of the periodic table, for example, tend to be smaller than the atoms of elements toward the lower left (see Figure 3.15). Knowing this, you can reasonably predict that selenium atoms (atomic number 34) are smaller than calcium atoms (atomic number 20), which is indeed the case.

Recall that such a gradual change in moving across the periodic table is called a *periodic trend*. Underlying most periodic trends are two important concepts: *inner-shell shielding* and *effective nuclear charge*.

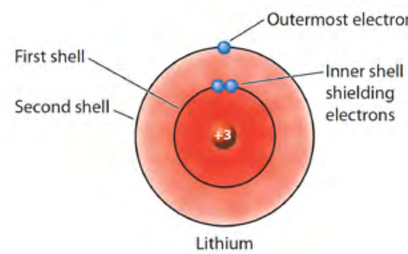
Consider the two electrons within the first shell of a helium atom. As shown in **Figure 4.34**, both of these electrons have an unobstructed "view" of the nucleus. Their attractions for this atomic nucleus are therefore the same. The situation gets more complicated for atoms beyond helium, which have more than one shell occupied by electrons. In these cases, inner-shell electrons weaken the attraction between outer-shell electrons and the nucleus. Imagine, for example, you are that second-shell electron in the lithium atom shown in **Figure 4.35**. Looking toward the nucleus, what do you sense? Not just the attractive force of the positively charged nucleus. You also feel the effect of the two electrons in the first shell—which, because of their negative charge, exert a repulsive force. Thus, the two inner electrons have the effect of weakening your electrical attraction to the nucleus. This is **inner-shell shielding**—inner-shell electrons shield electrons farther out from some of the attractive pull exerted by the positively charged nucleus.

Because inner-shell electrons diminish the attraction outer-shell electrons have for the nucleus, the nuclear charge sensed by outer-shell electrons is always less than the actual charge of the nucleus. This diminished nuclear charge experienced by outer-shell electrons is called the **effective nuclear charge** and is abbreviated Z* (pronounced zee-star), where Z stands for the nuclear charge and the asterisk indicates that this charge appears to be less than it actually is. A rough estimate of the Z* for any electron can be calculated by subtracting the total number of inner-shell electrons from the nuclear charge. For lithium (atomic number 3), the total number of inner-shell electrons, 2, is subtracted from the charge of the nucleus, +3, to give a calculated effective nuclear charge of +1. The second-shell electron of lithium, therefore, senses a nuclear charge of about +1, which is much less than the actual nuclear charge of +3. For most elements, subtracting the total number of inner-shell electrons from the nuclear charge provides a convenient estimate of the effective nuclear charge, as **Figure 4.36** illustrates.



∧ Figure 4.34

The two electrons in a helium atom have equal exposure to the nucleus; hence, they experience the same degree of attraction, represented by the pink shading in the space between the nucleus and the shell boundary.



∧ Figure 4.35

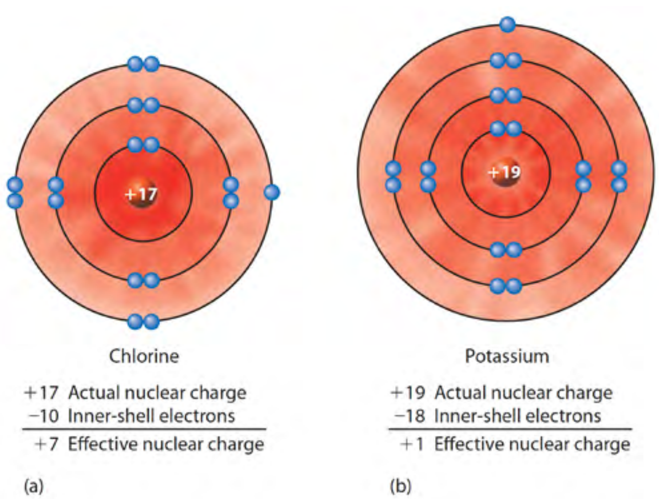
Lithium's two first-shell electrons shield the second-shell electron from the nucleus. The nuclear attraction, again represented by pink shading, is less intense in the second shell.

Why Atoms toward the Upper Right Are Smaller

From left to right across any row of the periodic table, the atomic diameters get *smaller*. Let's look at this trend from the point of view of effective nuclear charge. Consider lithium's outermost electron, which experiences an effective nuclear charge of about +1. Then look across period 2 to neon, in which each outermost electron experiences an effective nuclear charge of about +8, as **Figure 4.37** shows. Because the outer-shell electrons in neon experience a greater attraction to the nucleus, they are pulled in closer to the nucleus. So, neon, although nearly three times as massive as lithium, has a considerably smaller diameter.

READING CHECK

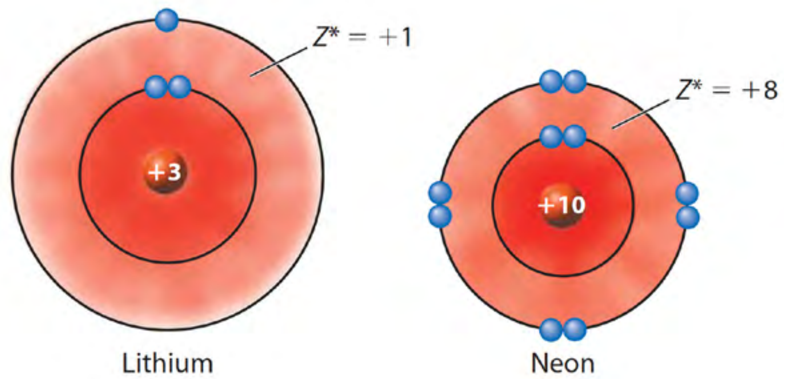
How is it possible to make a rough estimate of the effective nuclear charge experienced by an electron?



< Figure 4.36

A chlorine atom has three occupied shells. The ten electrons of the inner two shells shield each of the seven electrons of the third shell from the +17 nucleus. The third-shell electrons therefore experience an effective nuclear charge of about 17 - 10 = + 7. (b) In a potassium atom, the fourth-shell electron experiences an effective nuclear charge of about 19 - 18 = + 1.

Figure 4.37 >
Lithium's outermost electron experiences an effective nuclear charge of about +1, while those of neon experience an effective nuclear charge of about +8. As a result, the outer-shell electrons in neon are closer to the nucleus and the diameter of the neon atom is smaller than the diameter of the lithium atom.

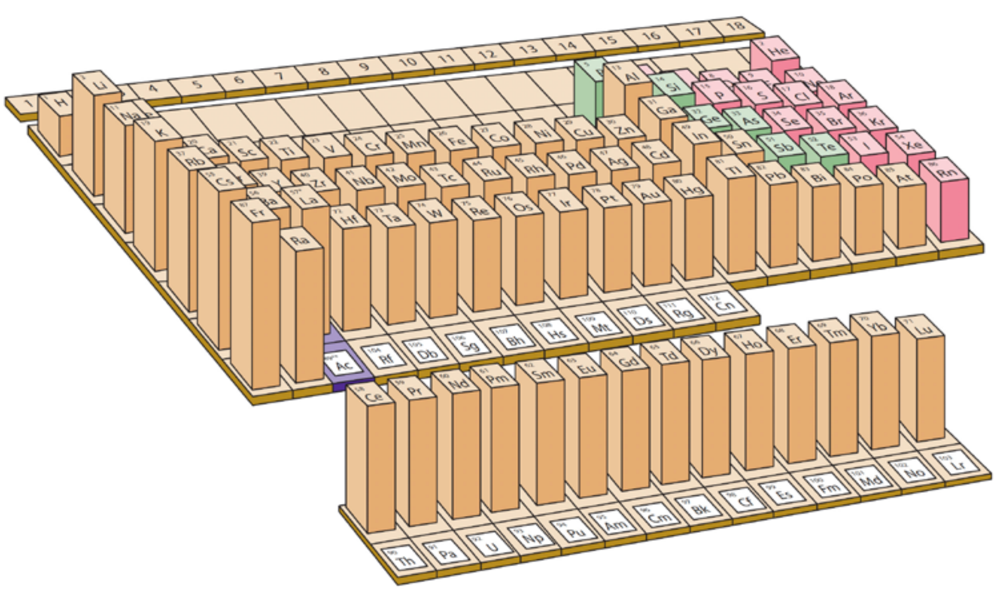


In general, across any period from left to right, atomic diameters become smaller because of an increase in effective nuclear charge. Look back at Figure 4.33 and you will see this trend illustrated for the first three periods. In addition, **Figure 4.38** shows relative atomic diameters obtained from experimental data. Note that there are some exceptions to this trend, especially between groups 12 and 13. These exceptions can be explained by probing further into the shell model than we need to for our purposes. Moving down a group, atomic diameters get larger because of an increasing number of occupied shells. Whereas lithium has a small diameter because it has only two occupied shells, francium (atomic number 87) has a much larger diameter because it has seven occupied shells.

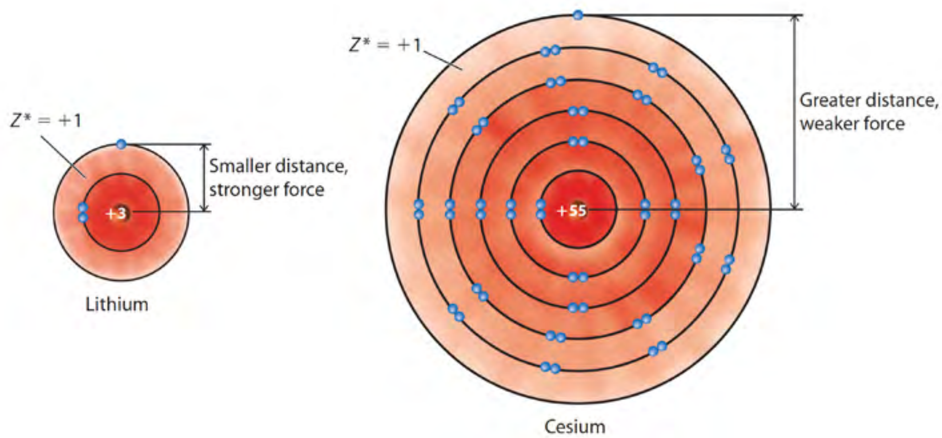
The Smallest Atoms Have the Most Strongly Held Electrons

How strongly electrons are bound to an atom is another property that changes gradually across the periodic table. In general, the trend is that the smaller the atom, the more tightly bound are its electrons.

As discussed earlier, effective nuclear charge increases in moving from left to right across any period. Thus, not only are atoms toward the right smaller, but their electrons are held more strongly. It takes about four times as much energy to remove an outer electron from a neon atom, for example, than to remove the outer electron from a lithium atom.



< Figure 4.38
This chart shows the relative atomic sizes, indicated by height. Note that atomic size generally decreases in moving to the upper right of the periodic table.



< Figure 4.39
In both lithium and cesium, the outermost electron has a calculated effective nuclear charge of +1. The outermost electron in a cesium atom, however, is not held as strongly to the nucleus because of its greater distance from the nucleus.

Moving down any group, the effective nuclear charge—as calculated by subtracting the charge of inner-shell electrons from the charge of the nucleus—stays the same. The effective nuclear charge for all group 1 elements, for example, works out to +1. Because of their greater number of shells, however, elements toward the bottom of a group (vertical column) are larger. The electrons in the outermost shell are therefore *farther* from the nucleus by an appreciable distance. From physics we learn that the electric force weakens rapidly with increasing distance. As **Figure 4.39** illustrates, an outer-shell electron in a larger atom, such as cesium, is not held as tightly as an outer-shell electron in a smaller atom, such as lithium. As a consequence, the energy needed to remove the outer electron from a cesium atom is about half the energy needed to remove the outer electron from a lithium atom. This is true even though they both have a calculated Z^* of +1. So, we see that our calculated Z^* values provide only a rough estimate of the nuclear pull on electrons. The effect of distance between a shell and the nucleus also needs to be considered.

The combination of the increasing effective nuclear charge from left to right and the increasing number of shells from top to bottom creates a periodic trend in which the electrons in atoms at the upper right of the periodic table are held most strongly and the electrons in atoms at the lower left are held least strongly. This is reflected in **Figure 4.40**, which shows **ionization energy**, the amount of energy needed to pull an electron away from an atom. The greater the ionization energy, the greater the attraction between the nucleus and its outermost electrons.

How strongly an atomic nucleus is able to hold on to the outermost electrons in an atom plays an important role in determining the atom's chemical behavior. What do you suppose happens when an atom that holds its outermost electrons only weakly comes into contact with an atom that has a very strong pull on its outermost electrons? As we explore in Chapter 6, the atom that pulls strongly may remove one or more electrons from the other atom. The result is that the two atoms become chemically bonded. So, the shell model not only gives us insight into the workings of the periodic table, it also helps us to understand the heart of chemistry, which is the study of how new materials are created by the bonding of atoms.

CHEMICAL CONNECTIONS

How is the weather connected to an atom?

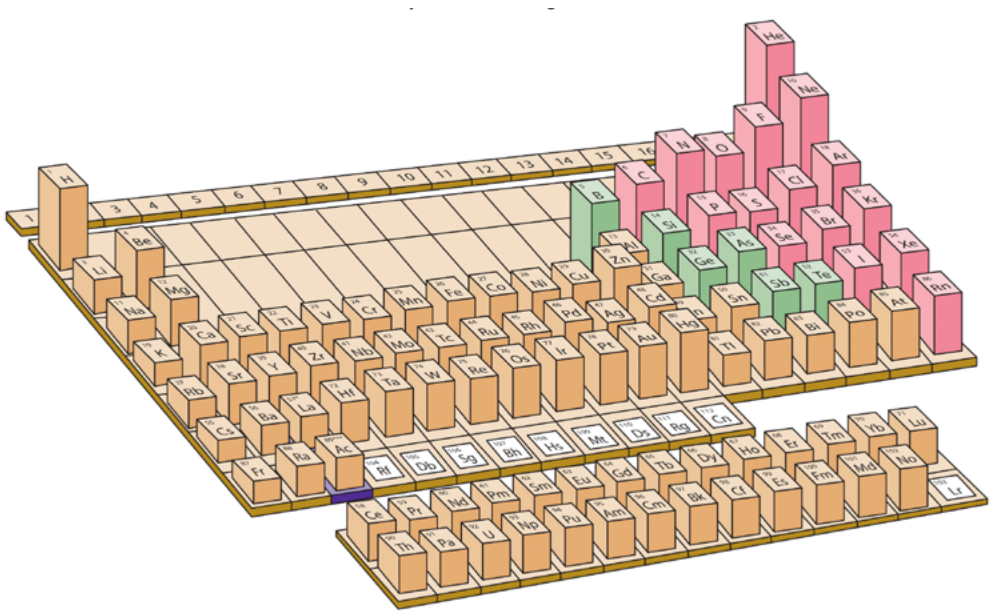


Figure 4.40 >
This chart shows the trends in ionization energy. The attraction an atomic nucleus has for its outermost electrons is indicated by height. Note that atoms at the upper right tend to have the greatest ionization energy and those at the lower left have the least.

CONCEPT **CHECK**

Which loses one of its outermost electrons more easily: a francium, Fr, atom (atomic number 87) or a helium, He, atom (atomic number 2)?

CHECK YOUR ANSWER

A francium, Fr, atom loses electrons much more easily than does a helium, He, atom. Why? Because a francium atom's outer electrons are not held as tightly by its nucleus, which is buried deep beneath many layers of shielding electrons.