Learning Objectives

• Relate the electron configurations of the elements to the shape of the periodic table.
• Determine the expected electron configuration of an element by its place on the periodic table.
• Identify the valence electrons and the valence electron count for an atom
• Determine the expected electron configuration for a cation or anion.

In Chapter 3, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure 1. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations: the shape of the periodic table mimics the filling of the subshells with electrons.

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure 2).
The next two electrons, for Li and Be, would go into the 2s subshell. Figure \(\PageIndex{3}\) shows that these two elements are adjacent on the periodic table.

![Figure \(\PageIndex{3}\): The 2s Subshell. In Li and Be, the 2s subshell is being filled.](image1)

For the next six elements, the 2p subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure \(\PageIndex{4}\)).

![Figure \(\PageIndex{4}\): The 2p Subshell. For B through Ne, the 2p subshell is being occupied.](image2)

The next subshell to be filled is the 3s subshell. The elements when this subshell is being filled, Na and Mg, are back on the left side of the periodic table (Figure \(\PageIndex{5}\)).

![Figure \(\PageIndex{5}\): The 3s Subshell. Now the 3s subshell is being occupied.](image3)

Next, the 3p subshell is filled with the next six elements (Figure \(\PageIndex{6}\)).
Fig. 8.5.6: The 3p subshell. Next, the 3p subshell is filled with electrons.

Instead of filling the 3d subshell next, electrons go into the 4s subshell (Figure \(\PageIndex{7}\)).

Figure \(\PageIndex{7}\): The 4s Subshell. The 4s subshell is filled before the 3d subshell. This is reflected in the structure of the periodic table.

After the 4s subshell is filled, the 3d subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure \(\PageIndex{8}\)).

Figure \(\PageIndex{8}\): The 3d Subshell. The 3d subshell is filled in the middle section of the periodic table.

And so forth. As we go across the columns of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the s subshells are being occupied. Because of this,
the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** are the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure \(\PageIndex{9}\) shows the blocks of the periodic table.

![Image of periodic table blocks](image)

**Valence Electrons**

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called **core electrons**.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

- **H**: \(1s^1\)
- **Li**: \(1s^2 2s^1\)
- **Na**: \([Ne]3s^1\)
- **K**: \([Ar]4s^1\)
- **Rb**: \([Kr]5s^1\)
- **Cs**: \([Xe]6s^1\)

They all have a similar electron configuration in their valence shells: a single \(s\) electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar
chemistry—and they do. The organization of electrons in atoms explains not only the shape of the periodic table but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the s and p blocks. In the d and f blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure \(\PageIndex{10}\). It is in the fourth column of the p block. This means that its electron configuration should end in a \(p^4\) electron configuration. Indeed, the electron configuration of Se is \([Ar]4s^23d^{10}4p^4\), as expected.

![Figure \(\PageIndex{10}\): Selenium on the Periodic Table](image)

Example \(\PageIndex{1}\):

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure \(\PageIndex{11}\)).

![Figure \(\PageIndex{11}\): Various Elements on the Periodic Table](image)

1. Ca
2. Sn
Solution

1. Ca is located in the second column of the s block. We would expect that its electron configuration should end with $s^2$. Calcium's electron configuration is [Ar]4s$^2$.

2. Sn is located in the second column of the p block, so we expect that its electron configuration would end in $p^2$. Tin's electron configuration is [Kr]5s$^2$4d$^10$5p$^2$.

Exercise \(\PageIndex{1}\)

From the element's position on the periodic table, predict the valence shell electron configuration for each atom. Figure \(\PageIndex{11}\).

a. Ti
b. Cl

Answer

a. [Ar]4s$^2$3d$^2$

b. [Ne]3s$^2$3p$^5$

\[\text{Figure } \PageIndex{6}: \text{This version of the periodic table shows the outer-shell electron configuration of each element. Note that down each group, the configuration is often similar.}\]

Electron Configurations of Ions

We have seen that ions are formed when atoms gain or lose electrons. A cation (positively charged ion) forms when one or more electrons are removed from a parent atom. For main group elements, the electrons that were added last are the first electrons removed. For transition metals and inner transition metals, however, electrons in the s orbital are easier to remove than the d or f electrons, and so the highest ns electrons are lost, and then the \((n-1)d\) or \((n-2)f\) electrons are
removed. An anion (negatively charged ion) forms when one or more electrons are added to a parent atom. The added electrons fill in the order predicted by the Aufbau principle.

Example \(\PageIndex{2}\): Predicting Electron Configurations of Ions

What is the electron configuration and orbital diagram of:

a. Na
b. P

Example \(\PageIndex{2}\): Predicting Electron Configurations of Ions

What is the electron configuration and orbital diagram of:

a. Na
b. P

c. Al

d. Fe

e. Sm

Solution

First, write out the electron configuration for each parent atom. We have chosen to show the full, unabbreviated configurations to provide more practice for students who want it, but listing the core-abbreviated electron configurations is also acceptable.

Next, determine whether an electron is gained or lost. Remember electrons are negatively charged, so ions with a positive charge have lost an electron. For main group elements, the last orbital gains or loses the electron. For transition metals, the last s orbital loses an electron before the d orbitals.

a. Na: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^1\). Sodium cation loses one electron, so Na\(^+\): 1s\(^2\)2s\(^2\)2p\(^6\).

b. P: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^3\). Phosphorus trianion gains three electrons, so P\(^3-\): 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\).

c. Al: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^1\). Aluminum dication loses two electrons Al\(^2+\): 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\).

d. Fe: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^6\). Iron(II) loses two electrons and, since it is a transition metal, they are removed from the 4s orbital Fe\(^2+\): 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3d\(^6\).

e. Sm: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^10\)4p\(^6\)5s\(^2\)4d\(^10\)5p\(^6\)6s\(^2\)4f\(^6\). Samarium trication loses three electrons. The first two will be lost from the 6s orbital, and the final one is removed from the 4f orbital. Sm\(^3+\): 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^10\)4p\(^6\)5s\(^2\)4d\(^10\)5p\(^6\)6s\(^2\)4f\(^5\).

Exercise \(\PageIndex{2}\)

a. Which ion with a +2 charge has the electron configuration 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4d\(^5\)?

b. Which ion with a +3 charge has this configuration?

Answer a

Tc\(^2+\)
Answer b

Ru$^{3+}$

Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.