3.5 Atomic Structure and the Periodic Table

As the study of chemistry progressed from the seventeenth through the nineteenth centuries, it became evident to scientists that Earth is composed of a great many elements, each with very different properties. Scientists realized that the structure of matter is much more complicated than the simple model of earth, air, fire, and water suggested by ancient philosophers. As more elements were separated and identified, chemists learned more about their properties. Gradually, chemists began to notice patterns in the properties of elements.

The Periodic Table

UNIT TASK **BOOKMARK**

Use the information in this section as you work on the Unit Task on page 268.

Have you ever wondered why elements in the periodic table are arranged the way they are? (Refer to the periodic table in Appendix B1). The elements of the periodic table are arranged according to the way electrons arrange themselves around the nuclei of atoms. Electron arrangement determines the chemical behaviour of every element. After reading this section, you will be able to look at the position of an element in the periodic table and predict its electron arrangement. You will also be able to apply the quantum mechanical model of the atom to explain the periodicity of chemical properties.

Understanding how electrons are arranged in atoms has allowed engineers to make new products and develop new technologies. For example, lithium (**Figure 1(a)**) is a very reactive element found in Group 1 of the periodic table. Lithium's reactivity is due to the arrangement of its electrons and their location relative to the nucleus. When lithium is exposed to water, it reacts noticeably (**Figure 1(b)**). Scientists have learned to maximize the potential of lithium by using it in batteries for cellphones, MP3 players, laptop computers, radios, cameras, pacemakers, hybrid and electric cars, and many other devices (**Figure 1(c)**).

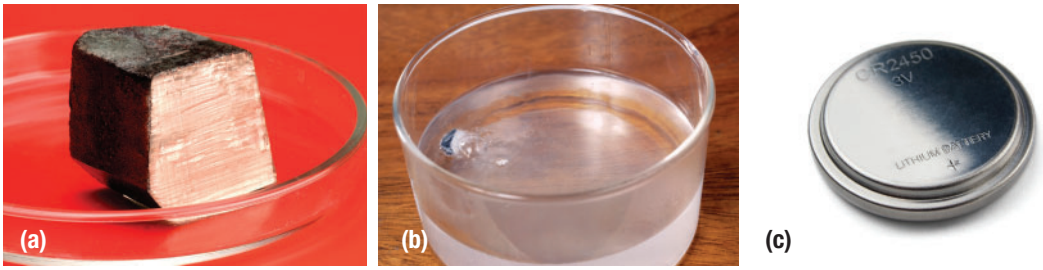


Figure 1 (a) The lithium atom is reactive. (b) Lithium reacts readily with water. (c) Lithium is used in many batteries because of its reactivity.

Multi-electronic Atoms

The Bohr–Rutherford model of the atom provides only a limited explanation of how electrons are configured in the atom. Bohr–Rutherford diagrams are useful for the first 20 elements, up to calcium. The quantum mechanical atomic model, with its four quantum numbers, not only describes all atoms in the periodic table, but also allows us to make theoretical predictions about atoms and their chemical properties.

To see how the quantum mechanical model applies to multi-electronic atoms (atoms with more than 1 electron), consider helium, which has 2 protons in its nucleus and 2 electrons in the 1*s* orbital.

Three energy contributions must be considered in the description of the helium atom: (1) the kinetic energy of the electrons as they move about the nucleus, (2) the potential energy of attraction between the nucleus and the electrons, and (3) the potential energy of repulsion between the 2 electrons. Even though the helium atom can be described in terms of the quantum mechanical model, the Schrödinger wave equation cannot be solved because the repulsions between electrons cannot be calculated exactly. This problem is called the electron correlation problem.

The electron correlation problem occurs with all multi-electronic atoms. To treat these systems using the quantum mechanical model, it is necessary to make approximations. The most common approximation is to treat each electron as if it were moving in a field of charge that is the net result of the nuclear attraction and the average repulsions of all the other electrons. For example, a sodium atom has 11 electrons: 10 electrons in the first and second energy shells and 1 electron in the third shell (**Figure 2**). Now consider the single outermost electron and the forces acting on it. The outermost electron is attracted to the positively charged nucleus, but it is also repelled by the 10 electrons in the first and second energy shells. The net effect is that the outer electron is not bound to the nucleus as tightly as it would be if the other electrons were not present. In essence, it is screened, or shielded, from the nuclear charge by the repulsions of the other electrons.

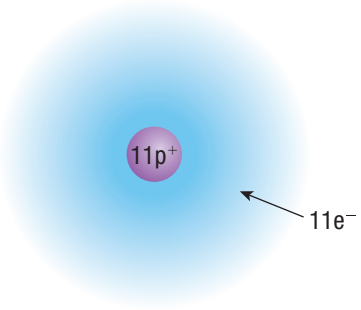


Figure 2 Sodium has 11 electrons and 11 protons.

The orbitals of multi-electronic atoms have the same general shapes as the orbitals for hydrogen, but their sizes and energy values are different. These differences occur because of the interplay between nuclear attraction and the repulsions from other electrons.

One important difference between multi-electronic atoms and the hydrogen atom is that, for hydrogen, all the orbitals of a given principal quantum level (shell) have the same energy. For example, in a hydrogen atom, electrons in the 2*s* and 2*p* orbitals possess the same energy. In multi-electronic atoms, for a given principal quantum level, *n*, the energies of electrons in the different orbitals vary as follows:

$$E_{ns} < E_{np} < E_{nd} < E_{nf}$$

In other words, electrons in a particular quantum shell fill orbitals in order of increasing energy: *s*, *p*, *d*, and then *f*. The reason for this has to do with the probability distributions of these orbitals. For example, notice in **Figure 3** that an electron in a 2*p* orbital has its maximum probability closer to the nucleus than does an electron in the 2*s* orbital. You might predict that a 2*p* orbital electron has lower energy than a 2*s* orbital electron. However, notice the small increase in electron density that occurs in the 2*s* orbital very near the nucleus. This means that for a small but very significant time, the 2*s* electron is closer to the nucleus than the 2*p* electron. This effect, called "penetration," causes an electron of a 2*s* orbital to be attracted to the nucleus more strongly than an electron of a 2*p* orbital. To summarize, electrons in the 2*s* orbital possess less energy than electrons in the 2*p* orbital of a multi-electronic atom. The probability distributions can also be represented by the probability profiles in **Figure 4**.

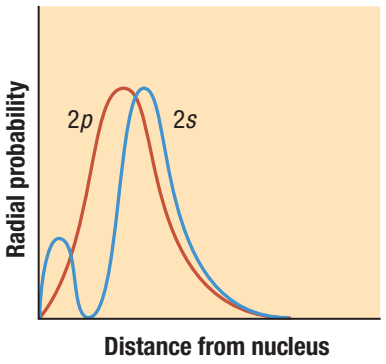


Figure 3 The probability profile of 2*s* orbital electrons and 2*p* orbital electrons

Figure 4 The probability profiles of (a) the 2*s* orbital and (b) the 2*p* orbital

The same rule applies to the other principal quantum levels. The relative energies of the electrons in the $n = 3$ orbitals are $E_{3s} < E_{3p} < E_{3d}$. In general, the more effectively electrons of an orbital penetrate the shielding electrons, the lower the energy of the electrons in that orbital.

The relative energies of the orbitals of the various sublevels are shown in **Figure 5**.

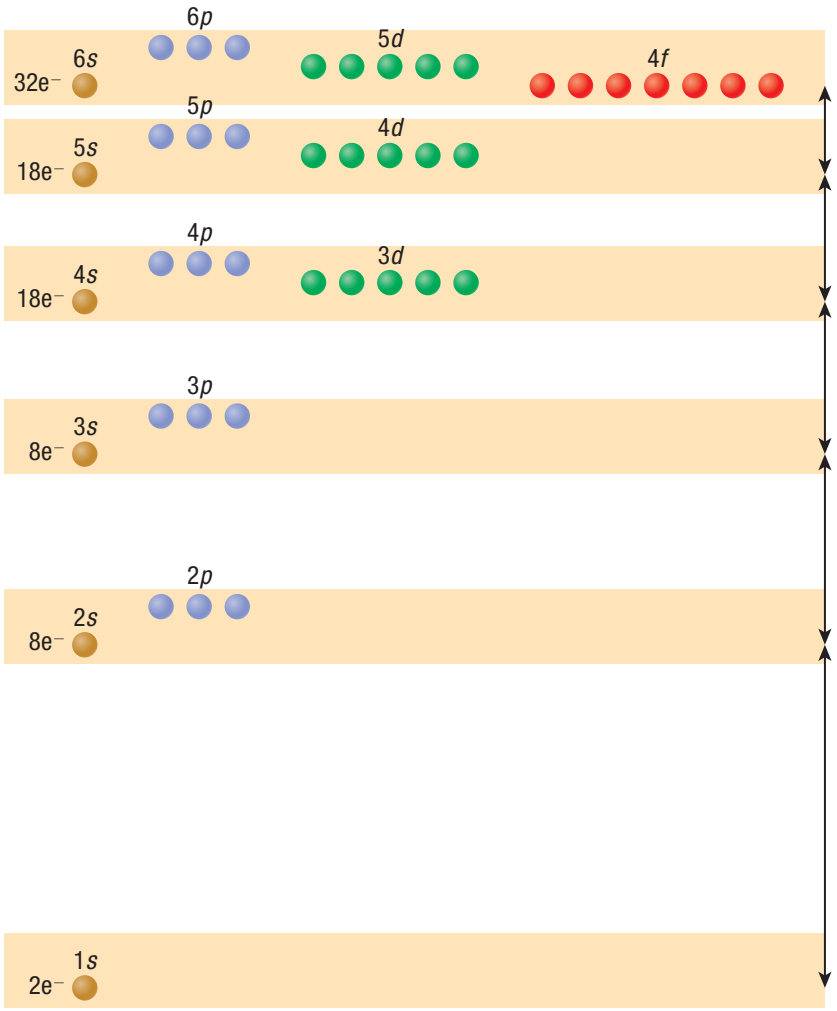


Figure 5 Relative energies of electron orbitals. Each orbital (dot) can contain up to 2 electrons.

electron configuration the location and number of electrons in the electron energy levels of an atom

aufbau principle the theory that an atom is "built up" by the addition of electrons, which fill orbitals starting at the lowest available energy orbital before filling higher energy orbitals (for example, 1*s* before 2*s*)

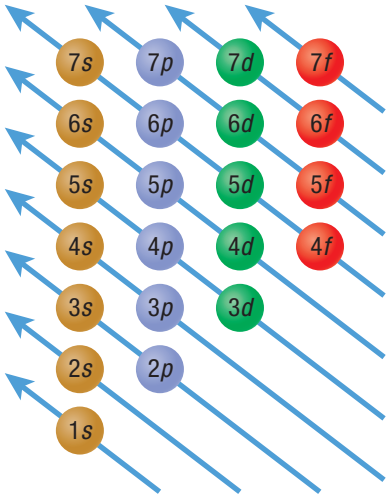


Figure 6 An aufbau diagram shows how electrons are added to each orbital, beginning with the 1*s* orbital (bottom left), until no electrons remain (top right).

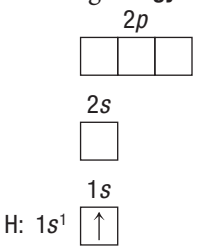
The Aufbau Principle and Writing Electron Configurations

You can apply the quantum mechanical model to show how the electron arrangements in atoms in the ground state account for the organization of the periodic table. The location and number of electrons in the energy levels of an atom or ion is called **electron configuration**.

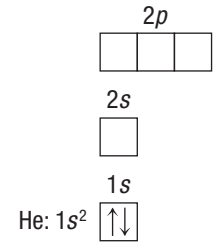
As you move across a row of the periodic table from left to right, atoms are arranged in order of increasing atomic number (proton number). Note that, in a neutral atom, the atomic number is also equal to the number of electrons. The result is that, as you move across the periodic table from left to right, the electron number of the elements also increases. The electron configuration of an atom can be determined using the **aufbau principle**, which hypothesizes that an atom is "built up" by progressively adding electrons. Furthermore, it states that as electrons are added, they assume their most stable condition (energy orbital) by filling the lowest available energy orbitals before filling higher energy orbitals. **Figure 6** shows how an aufbau diagram can be used as a tool for determining the electron configuration of an atom. The word *aufbau* is German for "building up."

The aufbau principle can be demonstrated by examining the electron configuration of the elements, moving from left to right along the periodic table. The following examples illustrate this. The hydrogen atom has 1 electron, which occupies the 1*s* orbital in its ground state. The electron configuration for hydrogen is written as $1s^1$, which can be represented by the following **energy-level diagram**, or **orbital diagram:**

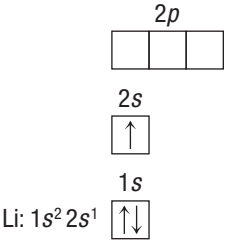
energy-level diagram (orbital diagram) a diagram that represents the relative energies of the electrons in an atom



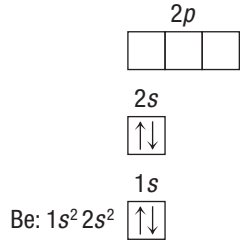
The next element, helium, has 2 electrons. According to the Pauli exclusion principle, since 2 electrons with opposite spins can occupy an orbital, the electrons in helium are in the 1*s* orbital with opposite spins. Helium therefore has a $1s^2$ configuration:



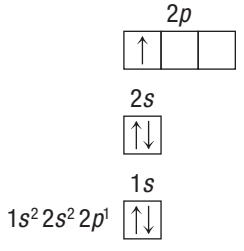
The atom of the element lithium has 3 electrons, 2 of which can go into the 1*s* orbital, which is now full. The third electron occupies the next lowest energy orbital ($n = 2$), the 2*s* orbital, so lithium has a $1s^2 2s^1$ configuration:



The atom of the next element, beryllium, has 4 electrons, which occupy the 1*s* and 2*s* orbitals:



The boron atom has 5 electrons, 4 of which occupy the 1*s* and 2*s* orbitals. The fifth electron goes into the second type of orbital with $n = 2$, the 2*p* orbital:

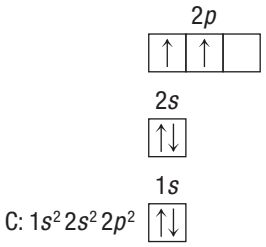


All the electrons in the 2*p* orbitals have the same energy, so it does not matter which 2*p* orbital the fifth electron occupies. By convention, we write it in the left-hand orbital.

Hund's rule a rule stating that in a particular set of orbitals of the same energy, the lowest energy configuration for an atom is the one with the maximum number of unpaired electrons allowed by the Pauli exclusion principle; unpaired electrons are represented as having parallel spins

For an atom with unfilled orbitals, the most stable energy level (lowest energy) is achieved when electrons occupy separate orbitals with parallel spins. An atom of carbon, the next element, has 6 electrons. Two electrons occupy the 1*s* orbital, 2 electrons occupy the 2*s* orbital, and 2 electrons occupy 2*p* orbitals. Since there are three 2*p* orbitals with the same energy, the electrons occupy separate 2*p* orbitals. This configuration is summarized by **Hund's rule,** which states that in a particular set of orbitals that have the same energy, the lowest energy configuration for an atom is the one with the maximum number of unpaired electrons allowed by the Pauli exclusion principle. In simple terms, this means that before any 2 electrons occupy an orbital in a subshell, other orbitals in the same subshell must first each contain 1 electron. Electrons filling a subshell will have parallel spin before the shell starts filling up with electrons having the (after the first orbital gains a second electron).

The electron configuration for the carbon atom could be written as $1s^2 2s^2 2p_x^1 2p_y^1$ to indicate that the electrons occupy separate 2*p* orbitals. However, the configuration is usually written as $1s^2 2s^2 2p^2$, and it is understood that the electrons are in different 2*p* orbitals. The energy-level diagram for carbon is



Note that the unpaired electrons in the 2*p* orbitals are shown with parallel spins.

The higher-energy orbitals have some anomalies in the order of filling orbitals. For example, in Figure 5 on page 162, the energy of the 4*s* orbital is lower than the energy of the 3*d* orbitals. Therefore, the 4*s* orbital must be filled before the 3*d* orbitals. When building the electron configuration of atoms, it is important to strictly follow this order. Use an aufbau diagram to write electron configurations in order of increasing energy. To do this you will also need to know the number of orbitals in each type of subshell: $s = 1$, $p = 3$, $d = 5$, and $f = 7$.

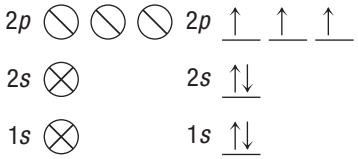
Procedure for Writing an Electron Configuration

1. Use the periodic table to determine the number of electrons in the atom or ion.
2. Assign electrons by main energy level and then by sublevel, using an energy-level diagram or an aufbau diagram.
3. Distribute electrons into orbitals that have the same energy according to Hund's rule.
4. Fill each sublevel before starting with the next sublevel. Continue until all electrons are assigned.
   - For anions (negatively charged ions), add an appropriate number of additional electrons.
   - For cations (positively charged ions), remove an appropriate number of electrons.

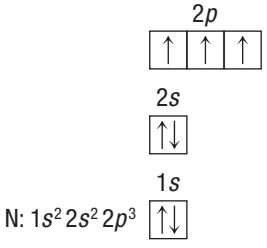
LEARNING **TIP**

Energy-level Diagrams

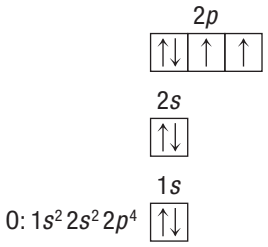
You may sometimes see energy-level diagrams expressed in different ways. The method used in this text (arrows in square boxes) is probably the most common. Two alternative ways to draw energy-level diagrams are shown below:



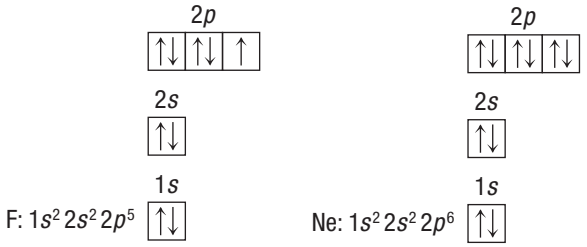
The electron configuration for the nitrogen atom (7 electrons) is $1s^2 2s^2 2p^3$. The 3 electrons in the 2*p* orbitals occupy separate orbitals with parallel spins:



The configuration for the oxygen atom, which has 8 electrons, is $1s^2 2s^2 2p^4$. One of the $2p$ orbitals is now occupied by a pair of electrons with opposite spins, as required by the Pauli exclusion principle:



The energy-level diagrams and electron configurations for atoms of the elements fluorine (9 electrons) and neon (10 electrons) are as follows:



For neon, the orbitals $n = 1$ and $n = 2$ are now completely filled.

In atoms of sodium, Na(s), the first 10 electrons occupy the $1s$, $2s$, and $2p$ orbitals, so the eleventh electron must occupy the first orbital with $n = 3$, the $3s$ orbital. The electron configuration for sodium atoms is $1s^2 2s^2 2p^6 3s^1$. To avoid writing the increasing string of inner-level electrons, a shorthand using noble gas configurations is used. In this abbreviated form of the electron configuration, the nearest noble gas preceding the element is notated in brackets, and the electron configuration is continued from that point forward. For example, the noble gas configuration for sodium is $[Ne]3s^1$, where [Ne] represents the electron configuration of neon, $1s^2 2s^2 2p^6$.

The atom of the next element, magnesium, has the electron configuration $1s^2 2s^2 2p^6 3s^2$, or $[Ne]3s^2$. Atoms of the next 6 elements, aluminum through argon, have configurations obtained by filling the $3p$ orbitals one electron at a time.

Consider an analogy for the aufbau principle and the process of filling orbitals. Imagine that the atom is a concert hall (**Figure 7**). The stage is the nucleus, and the seating area represents the space where electrons most likely exist. The placement of electrons in orbitals in order of increasing energy follows a similar pattern to filling the seats closest to the stage first. The principal quantum number (which tells you the energy level) is like the seating section. The secondary quantum number (which tells you the subshell) is like the row. The magnetic quantum number (which gives specific information on the orientation of the orbital) is your seat. In this analogy, each seat is a loveseat. It can seat two people (left and right). The loveseat analogy is not totally appropriate, though. For the fourth quantum number—the spin quantum number—one person would have to be sitting on his or her head for the analogy to work completely.

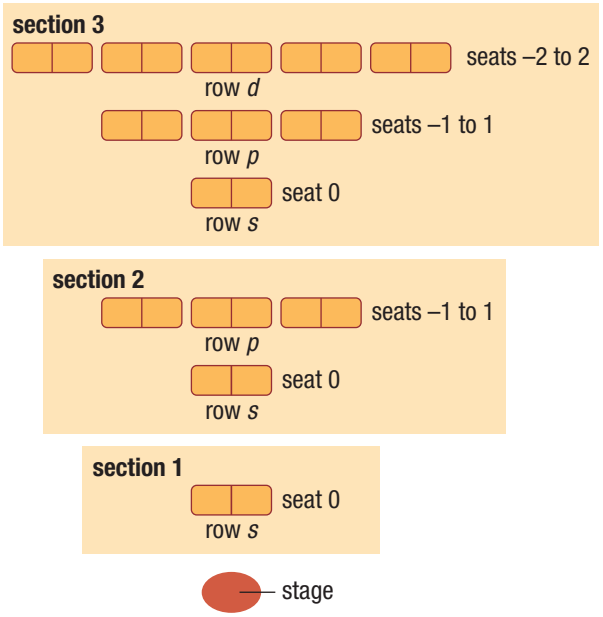


Figure 7 The atom can be thought of as analogous to a concert hall: the process of filling seats is similar to filling orbitals with electrons.

Tutorial 1 Writing Electron Configurations

In this tutorial, you will practise writing electron configurations of atoms and ions in different ways, and use electron configurations to identify an element.

Sample Problem 1: Writing a Full Electron Configuration

Write the full electron configuration for the sulfur atom.

Solution

Step 1. Determine the total number of electrons in the atom.

Sulfur has 16 electrons.

Step 2. Assign electrons in order of main energy levels and sublevels using the aufbau principle. Remember that each orbital can hold 2 electrons and the various subshells have specific numbers of orbitals (there is 1 *s* orbital, 3 *p* orbitals, 5 *d* orbitals, and so on).

1*s* has 2 electrons, 2*s* has 2 electrons, 2*p* has 6 electrons, 3*s* has 2 electrons, and 3*p* will fill 4 of 6 orbitals.

Statement: The electron configuration for sulfur is $1s^2 2s^2 2p^6 3s^2 3p^4$.

Sample Problem 2: Identifying an Element

Identify the element that has atoms with the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

Solution

Count the number of electrons.

There are 28 electrons.

The number of electrons corresponds to the number of protons in a neutral atom. Therefore, the number of protons is 28 so the element is nickel (Ni).

Statement: The element with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ is nickel.

Sample Problem 3: Writing a Shorthand Electron Configuration

Write the shorthand electron configuration (noble gas configuration) for the chlorine atom.

Solution

Determine the noble gas immediately preceding chlorine, which is neon (Ne). Add the additional electrons to this noble gas.

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

Statement: The shorthand electron configuration (noble gas configuration) for chlorine is $[Ne]3s^2 3p^5$.

Sample Problem 4: Electron Configurations for Ions

Write the electron configuration for the calcium ion Ca^{2+}.

Solution

Follow the same process as you did for a neutral atom, but add or remove electrons as required.

The electron configuration for a calcium atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$. The calcium ion, Ca^{2+}, has a +2 charge, so remove 2 electrons. The electron configuration for the calcium ion is $1s^2 2s^2 2p^6 3s^2 3p^6$.

Statement: The electron configuration for the calcium ion Ca^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6$.

Practice

1. Write the full electron configuration for an atom of the element francium. T/I C
   [ans: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^6 7s^1$]
2. Identify the element with the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ T/I C [ans: K]
3. Write the shorthand electron configuration (noble gas configuration) for an atom of silicon. T/I C [ans: $[Ne]3s^2 3p^2$]
4. Write the electron configuration for the magnesium ion Mg^{2+}. T/I C [ans: $1s^2 2s^2 2p^6$]

Explaining the Periodic Table

Electron configuration can help explain the structure of the periodic table and periodic trends. **Figure 8** summarizes the electron configurations of the first 18 elements by giving the number of electrons in the type of orbital occupied last. Note that an important pattern is developing: the elements in the same group have the same valence electron configuration. Recall that **valence electrons** are the electrons in the outermost shell (outermost principal quantum level) of an atom. For example, the valence electrons in a nitrogen atom are the 2*s* and 2*p* electrons, a total of 5 electrons. In a sodium atom, the valence electron is the electron in the 3*s* orbital, and so on. Valence electrons are the most important electrons to chemists because they are involved in bonding, as you will read in the next chapter. Remember that Mendeleev originally placed the elements in groups based on similarities in chemical properties. Now you know the reasons for these similarities. Elements with the same valence electron configuration show similar chemical behaviour. Chemical similarities allow you to draw conclusions about electron configurations. WEB LINK

valence electron an electron in the outermost principal quantum level of an atom

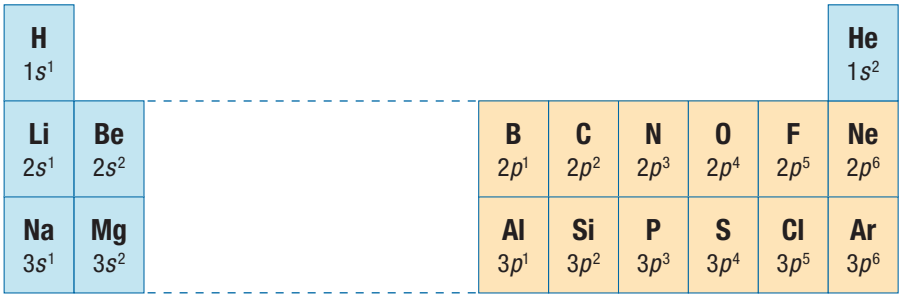


Figure 8 The valence electron configurations for the first 18 elements

The nineteenth element of the periodic table is potassium. Recall that the 3*p* orbitals in an argon atom are fully occupied. Consequently, you might expect the additional electron in a potassium atom to occupy a 3*d* orbital, since for $n = 3$, the orbitals are 3*s*, 3*p*, and 3*d*. However, experimentation has shown that the properties of potassium are very similar to those of lithium and sodium. Potassium, like sodium, is very reactive with oxygen, so it must be carefully stored away from air. This property of potassium is evidence that the outermost electron in the potassium atom occupies the 4*s* orbital, rather than one of the 3*d* orbitals. The electron configuration of a potassium atom is shown below:

K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or $[\text{Ar}]4s^1$.

The electron configuration of the atom of the next element, calcium, is Ca: $[\text{Ar}]4s^2$.

The next element, scandium, begins a series of 10 elements (scandium through zinc) called the **transition metals**, which are elements whose highest-energy electrons are in *d* orbitals. The electron configuration of an atom of a transition metal is obtained by adding electrons to the five 3*d* orbitals. The configuration of an atom of scandium is Sc: $[\text{Ar}]4s^2 3d^1$, that of titanium is Ti: $[\text{Ar}]4s^2 3d^2$, and that of vanadium is V: $[\text{Ar}]4s^2 3d^3$.

transition metal an element whose highest-energy electrons are in *d* orbitals

The expected electron configuration of atoms of the next element, chromium, is $[\text{Ar}]4s^2 3d^4$. However, the observed configuration is Cr: $[\text{Ar}]4s^1 3d^5$. The chromium atom is an exception to the aufbau principle. One explanation for this anomaly is provided by experimental evidence that indicates that unfilled subshells are less stable than half-filled and filled subshells, and that unfilled subshells have higher energy. It is less important for *s* orbitals to be filled or half-filled compared with *d* orbitals. In the chromium atom, an *s* electron moves to the *d* subshell and creates two half-filled *s* and *d* subshells: $4s^2 3d^4$ becomes $4s^1 3d^5$. This movement of electrons creates an overall energy state that is lower and therefore more stable.

The atoms of each of the next four elements, manganese through nickel, have the expected configurations:

Mn: $[Ar]4s^2 3d^5$ Co: $[Ar]4s^2 3d^7$

Fe: $[Ar]4s^2 3d^6$ Ni: $[Ar]4s^2 3d^8$

The electron configuration of the copper atom is expected to be $[Ar]4s^2 3d^9$. However, the observed configuration is Cu: $[Ar]4s^1 3d^{10}$. Copper is another exception to the aufbau principle. In the copper atom, an *s* electron moves to the *d* subshell and creates a half-filled *s* subshell and a filled *d* subshell, which makes it more stable.

The atom of the next element, zinc, has the expected configuration: Zn: $[Ar]4s^2 3d^{10}$.

The configurations of elements in the first row of the transition metals are shown in **Figure 9**. The elements gallium through krypton have configurations that correspond to filling the 4*p* orbitals.

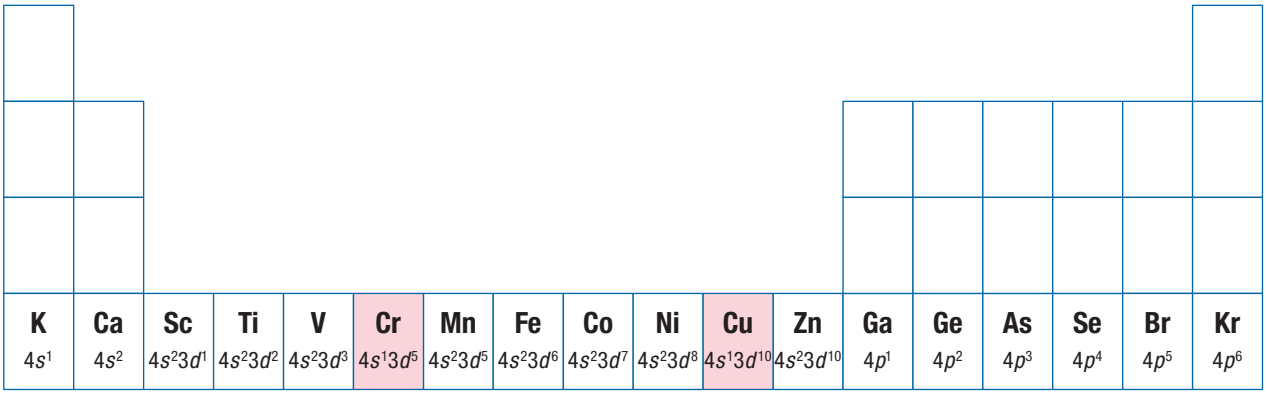


Figure 9 Valence electron configurations for potassium through krypton. The fourth period transition metals (scandium through zinc) have the general configuration $[Ar]4s^2 3d^n$, except for chromium and copper (highlighted in pink).

Figure 10 shows which orbitals in the periodic table are the last to be filled.

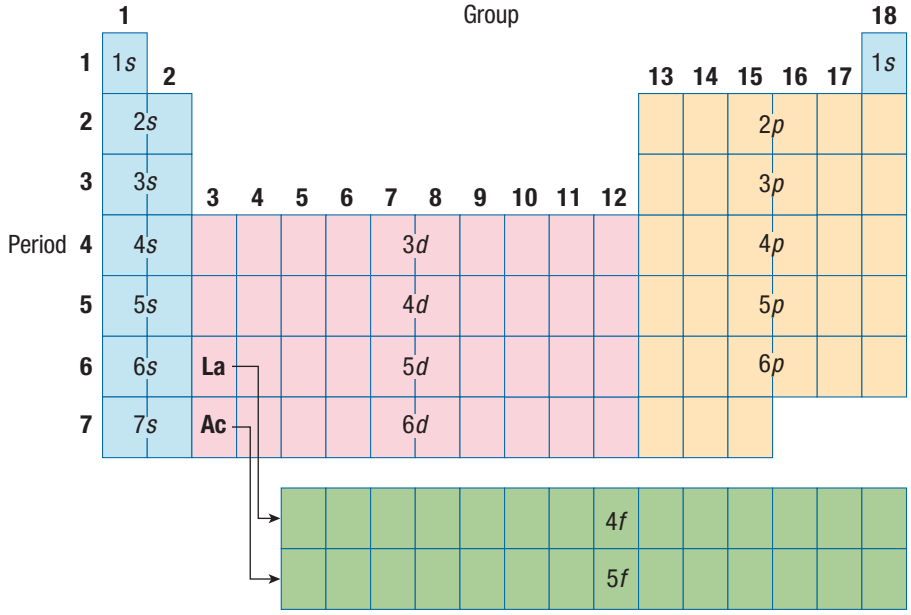


Figure 10 In the atoms of a period, the $(n + 1)s$ orbital fills before the *nd* orbital. The group labels indicate the number of valence electrons (*ns* plus *np* electrons) of atoms of the elements in each group.

Note the following additional points (with reference to Figure 10):

1. The $(n + 1)s$ orbitals always fill before the nd orbitals. For example, the 5*s* orbitals fill before the 4*d* orbitals in atoms of the period 5 transition metals (yttrium through cadmium). This order of filling can be explained by the penetration effect: the 4*s* orbital penetrates closer to the nucleus, so it has a lower energy than the 4*d* orbital (see Figure 5, p. 162).
2. After lanthanum ($[Xe]6s^2 5d^1$), there is a group of 14 elements called the lanthanide series, or the lanthanides. This series of elements corresponds to filling the seven 4*f* orbitals in their atoms. Similarly, after actinium ($[Rn]7s^2 6d^1$), there are the 14 elements that make up the actinide series.
3. In the lanthanide and actinide series, since the 5*d* and 4*f*, and the 6*d* and 5*f* orbitals are close in energy, the increased stability associated with empty, half-filled and fully filled orbitals affects the electron configurations.
4. The group labels 1 to 18 indicate the total number of valence electrons for the atoms in these groups. For example, atoms of all the elements in Group 15 have the configuration ns^2np^3. (The *d* electrons fill one period later and are usually not counted as valence electrons.) The elements in Groups 1 to 18 are often called the main group, or **representative elements**. The atom of every member in each group has the same valence electron configuration.

representative elements those elements in the main blocks of the periodic table, which are Groups 1 to 18 (the *s* and *p* blocks)

The quantum mechanical model explains the arrangement of elements in the periodic table. This model allows you to understand why the elements in a group have similar chemistry: they all have atoms with the same valence electron configuration. Only the principal quantum number of the valence orbitals changes in a particular group.

In this text, when an electron configuration is given, the orbitals are listed in the order in which they fill. It is important to be able to write the electron configuration of atoms of each of the main group elements. If you understand how the periodic table is organized, it is not necessary to memorize the order in which the orbitals fill. Review Figure 10 to ensure you understand the correspondence between the orbitals and the periods and groups.

Predicting the electron configurations of the atoms of the transition metals (3*d*, 4*d*, and 5*d* elements), the lanthanides (4*f* elements), and the actinides (5*f* elements) is more difficult because there are many exceptions. Familiarize yourself with the electron configurations of the atoms chromium and copper, the two exceptions in the first-row transition metals, because these elements are often encountered: Cr: $[Ar]4s^1 3d^5$ and Cu: $[Ar]4s^1 3d^{10}$.

UNIT TASK **BOOKMARK**

Use the information in this section as you work on the Unit Task on page 268.

Tutorial 2 Writing Electron Configurations Using the Periodic Table

In this tutorial, you will practise writing electron configurations using the periodic table instead of the energy-filling order.

Sample Problem 1: Writing Electron Configurations

Give the electron configurations for the atoms of sulfur, S; cadmium, Cd; hafnium, Hf; and radium, Ra, using the periodic table (Appendix B1). Use Figure 10 on page 168 as an aid.

Solution

Sulfur is element 16 and resides in period 3, where the 3*p* orbitals of the atoms are being filled. Since sulfur is the fourth of the 3*p* elements in Figure 10, its atom must have 4 3*p* electrons. Its configuration is

S: $1s^2 2s^2 2p^6 3s^2 3p^4$ or [Ne]$3s^2 3p^4$

Cadmium is element 48 and is located in period 5 at the end of the 4*d* transition series. Since it is the tenth element in the series, the cadmium atom has 10 electrons in the 4*d* orbitals, in addition to the 2 electrons in the 5*s* orbital. Its electron configuration is

Cd: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$ or [Kr]$5s^2 4d^{10}$

Hafnium is element 72 and is found in period 6. Note that it occurs just after the lanthanide series. Thus, the 4*f* orbitals are already filled in the hafnium atom. Hafnium is the second member of the 5*d* transition series and its atom has two 5*d* electrons. The electron configuration of the hafnium atom is

Hf: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^2$

or [Xe]$6s^2 4f^{14} 5d^2$

Radium is element 88 and is in period 7. Thus, the radium atom has 2 electrons in the 7*s* orbital. Its electron configuration is

Ra: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2$

or [Rn]$7s^2$

Practice

1. Give the full electron configurations for the elements titanium, Ti(s), and aluminum, Al(s), using the periodic table in Appendix B1. Modify these configurations to show the shorthand noble gas form. T/I C

   [ans: Ti: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$ or [Ar]$4s^2 3d^2$; Al: $1s^2 2s^2 2p^6 3s^2 3p^1$ or [Ne]$3s^2 3p^1$]

2. Write the shorthand configuration for the following: $_{35}$Br; $_{43}$Tc; $_{83}$Bi, $_{59}$Pr; $_{96}$Cm T/I C

   [ans: $_{35}$Br: [Ar]$4s^2 3d^{10} 4p^5$; $_{43}$Tc: [Kr]$5s^2 4d^5$; $_{83}$Bi: [Xe]$6s^2 4f^{14} 5d^{10} 6p^3$; $_{59}$Pr: [Xe]$6s^2 5d^1 4f^2$; $_{96}$Cm: [Rn]$7s^2 6d^1 5f^7$]

Explaining Ion Charges

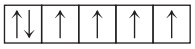
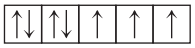
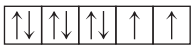
The ability of transition metals to form multiple ions can now be explained. The electron configuration for the neutral cadmium atom is Cd: [Kr]$5s^2 4d^{10}$. It shows that the cadmium atom has 12 electrons in its outer orbitals. Cadmium forms a 2+ ion. If the outer two 5*s* electrons were removed, the neutral cadmium atom would be left with a full 4*d* orbital and would be relatively stable: Cd^{2+}: [Kr]$4d^{10}$. Alternatively, the cadmium atom could give up 10 electrons and leave filled 5*s* orbitals. (However, this possibility is very unlikely.)

Lead, Pb, is an example of how 2+ and 4+ ions form. The neutral electron configuration for the lead atom is Pb: [Xe]$6s^2 4f^{14} 5d^{10} 6p^2$. It shows full 6*s*, 4*f*, and 5*d* orbitals and a partially filled 6*p* orbital (2 out of 6 electrons). The lead atom could lose these two 6*p* electrons fairly easily, forming a 2+ lead ion. To form a 4+ ion, the lead atom would have to lose 4 electrons from its 6*s* and 6*p* orbitals.

Explaining Magnetism

Table 1 shows three strongly magnetic elements and the electron configurations of their atoms. Based on the evidence in Table 1, you might speculate that magnetism is caused by the presence of several unpaired electrons in an atom. However, if you examine the periodic table more closely, you will notice that the elements ruthenium, rhodium, and palladium—which are in the same group as iron, nickel, and cobalt—are only weakly magnetic. These elements also have several unpaired electrons. Therefore, having several unpaired electrons only partially explains the strong magnetic properties of iron, nickel, and cobalt. There must be other properties that can explain why these metals are so strongly magnetic.

Table 1 Electron Configurations of Ferromagnetic Elements

Ferromagnetic element	Electron configuration	*d*-orbital filling	Pairing of *d* electrons
Fe	$[Ar]4s^2 3d^6$		1 pair; 4 unpaired
Co	$[Ar]4s^2 3d^7$		2 pairs; 3 unpaired
Ni	$[Ar]4s^2 3d^8$		3 pairs; 2 unpaired

Scientists have found that iron, nickel, and cobalt consist of small, dense packs of atoms called domains. The magnetic fields of atoms in a domain align in the same direction, even though the domains themselves align randomly (**Figure 11(a)**). However, in the presence of an outside magnetic field, the magnetic fields of most of the domains become aligned with the outside field and, hence, with one another (**Figure 11(b)**). As a result, many groups of atoms become aligned in the same direction and the metal becomes a permanent magnet. They stay magnetized even when the external magnetic field is removed. This is because their magnetic properties are based on the alignment of their atoms. The mechanism by which elements form permanent magnets is called **ferromagnetism**. Iron, nickel, and cobalt are the most well-known ferromagnetic elements.

ferromagnetism the very strong magnetism commonly exhibited by materials that contain nickel, iron, and cobalt; applies to a collection of atoms

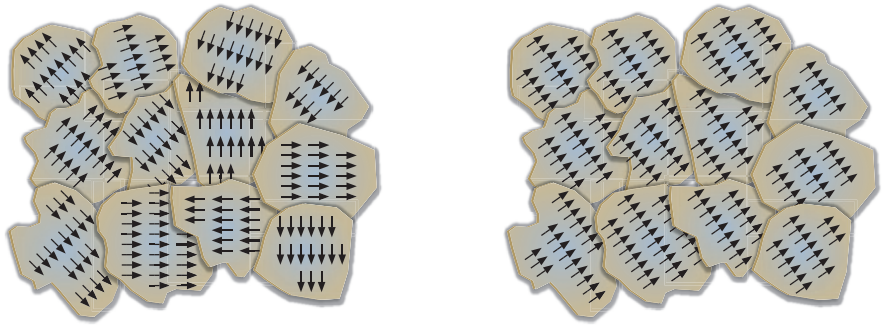


Figure 11 (a) In the absence of an external magnetic field, the magnetic domains are randomly aligned. (b) In the presence of a magnetic field, the domains align with the field. Once aligned, these domains stay aligned until they are disturbed.

Iron, cobalt, and nickel display strong magnetic properties, but there are other elements, such as aluminum and platinum, that display weak but measurable magnetic fields. These fields arise due to the presence of unpaired electrons in the atom. Hund's rule tells us that unpaired electrons in an atom all have the same spin. This spinning generates a weak magnetic field, a property called **paramagnetism**. This magnetic field is not usually noticed but is detectable when the element interacts with a strong magnetic field.

paramagnetism the weak attraction of a substance to a magnet; applies to individual atoms

Investigation 3.5.1

Paramagnetism (page 182)
Some elements exhibit magnetic properties, whereas some elements display magnetic properties only when exposed to a magnetic field. This controlled experiment will give you an opportunity to identify paramagnetic materials.

3.5 Review

Summary

- The electron configuration of an atom describes the energies and locations of the electrons in the atom.
- The aufbau principle states that each successive electron added to an atom occupies the lowest-energy orbital available to it.
- Hund's rule states that orbitals of equal energy fill with 1 unpaired electron each before electrons begin to pair up.
- Valence electrons—the electrons in the outermost principal quantum level of an atom—determine many of the chemical properties of elements.
- Electron configurations can be simplified by writing shorthand configurations.
- The characteristic properties and positions of elements in each of the *s*, *p*, and *d* sections of the periodic table are related to their electron configurations.

Questions

1. (a) Four blocks of elements in a periodic table refer to various subshells being filled. What are the four blocks and the corresponding orbitals?
   (b) What information from the periodic table helps you derive an electron configuration from the position of an element in the periodic table?
   (c) State the aufbau principle and Hund's rule.
   (d) Explain how the aufbau principle and Hund's rule help us determine the order of orbital filling. K/U
2. Draw an outline of the periodic table. Shade the *s* block in red, the *p* block in blue, the *d* block in green, and the *f* block in yellow. Include a key. K/U C
3. For elements 1 to 36, there are two exceptions to the filling order predicted by the periodic table. K/U T/I
   (a) Write the electron configurations for these two elements, and indicate how many unpaired electrons are present.
   (b) Briefly explain why these unpredicted electron configurations form.
4. The elements Si, Ga, As, Ge, Al, Cd, S, and Se are all used in the manufacture of various semiconductor devices. Write the expected full electron configurations for the atom of each of these elements. T/I C
5. The elements Cu, O, La, Y, Ba, Tl, and Bi are all found in high-temperature ceramic superconductors. Write the shorthand noble gas configurations for the atom of each of these elements. T/I C
6. Use the periodic table to predict which orbital the last electron enters for each of the following atoms: T/I
   (a) Zn
   (b) I
   (c) Ba
7. Using only the periodic table, write the expected ground-state electron configuration for the atom of
   (a) the third element in Group 15
   (b) element number 116
   (c) an element with 3 unpaired 5*d* electrons
   (d) the halogen with electrons in the 6*p* atomic orbitals K/U T/I C
8. Write the full electron configuration for the atom of each of the following: K/U T/I C
   (a) the lightest halogen
   (b) an alkali metal with a full 3*p* orbital
   (c) the Group 13 element in the same period as Sn
   (d) the non-metallic elements in Group 14
9. Draw the energy-level diagram for the atom of each of the following elements: T/I C
   (a) boron
   (b) silicon
   (c) mercury
10. Use electron configurations to explain
    (a) why arsenic can acquire charges of -3 and $+3$ but not $+2$
    (b) the $+2$ charge on the lead(II) ion
    (c) the $+1$ charge of the silver ion K/U T/I C