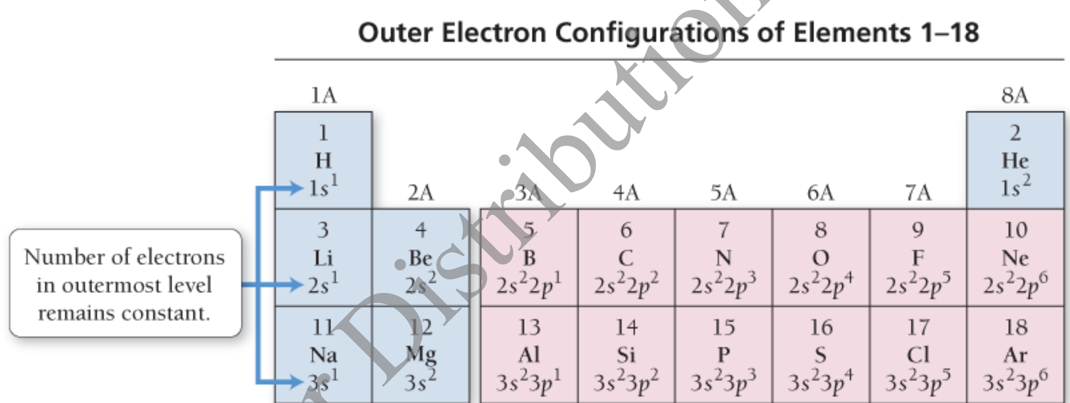
Pearson

3.4: Electron Configurations, Valence Electrons, and the Periodic Table

Key Concept Video **Writing an Electron Configuration Based on an Element's Position on the Periodic Table**

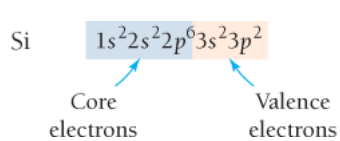
Recall from Section 3.2 that Mendeleev arranged the periodic table so that elements with similar chemical properties lie in the same column. We can begin to make the connection between an element's properties and its electron configuration by superimposing the electron configurations of the first 18 elements onto a partial periodic table, as shown in Figure 3.9. As we move to the right across a row (which is also called a period), the orbitals fill in the correct order. With each subsequent row, the highest principal quantum number increases by one. Notice that as we move down a column, *the number of electrons in the outermost principal energy level (highest n value) remains the same*. The key connection between the macroscopic world (an element's chemical properties) and the particulate world (an atom's electronic structure) lies in these outermost electrons.

Figure 3.9 **Outer Electron Configurations of the First 18 Elements in the Periodic Table**



An atom's **valence electrons** are the most important in chemical bonding. *For main-group elements, the valence electrons are those in the outermost principal energy level*. For transition elements, we also count the outermost *d* electrons among the valence electrons (even though they are not in an outermost principal energy level). The chemical properties of an element depend on its valence electrons, which are instrumental in bonding because they are held most loosely (and are therefore the easiest to lose or share). We can now see *why* the elements in a column of the periodic table have similar chemical properties: *They have the same number of valence electrons*.

We distinguish valence electrons from all the other electrons in an atom, which we call **core electrons**. The core electrons are those in *complete* principal energy levels and those in *complete d* and *f* sublevels. For example, silicon, with the electron configuration $1s^22s^22p^63s^23p^2$ has four valence electrons (those in the $n = 3$ principal level) and ten core electrons:



Example 3.3 Valence Electrons and Core Electrons

Write the electron configuration for Ge. Identify the valence electrons and the core electrons.

SOLUTION

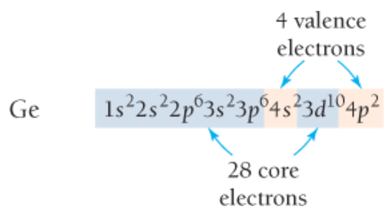
To write the electron configuration for Ge, determine the total number of electrons from germanium's atomic number (32) and distribute them into the appropriate orbitals.

Ge $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

Because germanium is a main-group element, its valence electrons are those in the outermost principal energy level. For germanium, the $n = 1, 2,$ and 3 principal levels are complete (or full), and the $n = 4$ principal level is outermost. Consequently, the $n = 4$ electrons are valence electrons and the rest are core electrons.

Note: In this book, we always write electron configurations with the orbitals in the order of filling. However, it is also common to write electron configurations in the order of increasing principal quantum number. The electron configuration of germanium written in order of increasing principal quantum number is:

$$\text{Ge} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$$

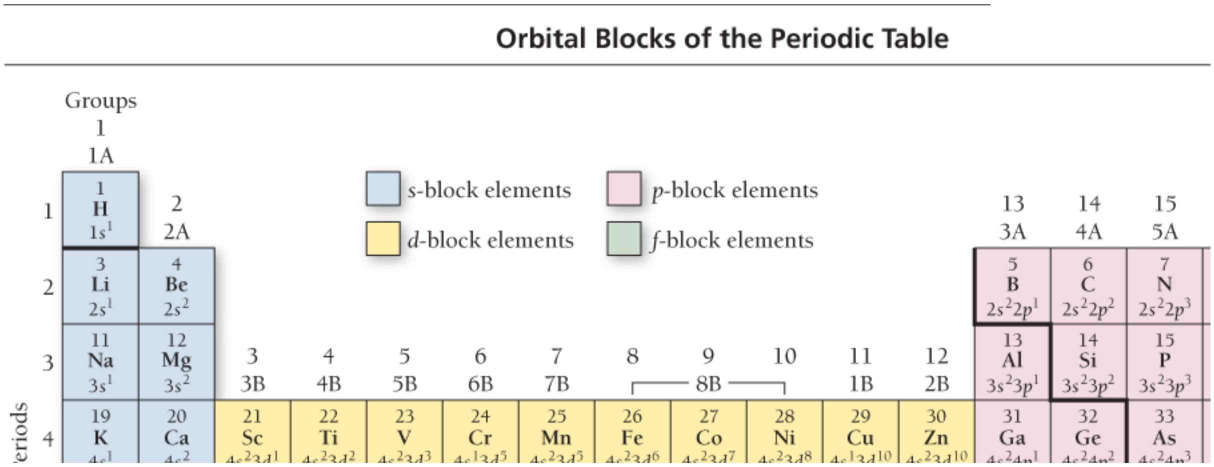


FOR PRACTICE 3.3 Write the electron configuration for phosphorus. Identify its valence electrons and core electrons.

Orbital Blocks in the Periodic Table

A pattern similar to what we saw for the first 18 elements exists for the entire periodic table, as shown in Figure 3.10. Note that, because of the filling order of orbitals, the periodic table can be divided into blocks representing the filling of particular sublevels. The first two columns on the left side of the periodic table constitute the *s* block, with outer electron configurations of ns^1 (group 1A) and ns^2 (group 2A). The six columns on the right side of the periodic table constitute the *p* block, with outer electron configurations of $ns^2\,np^1$, $ns^2\,np^2$, $ns^2\,np^3$, $ns^2\,np^4$, $ns^2\,np^5$, and $ns^2\,np^6$. Together, the *s* and *p* blocks constitute the *main-group* elements. The *transition* elements consitute the *d* block, and the lanthanides and actinides (also called the inner transition elements) constitute the *f* block. (For compactness, the *f* block is typically printed below the *d* block instead of being embedded within it.)

Figure 3.10 **The s, *p*, *d*, and *f* Blocks of the Periodic Table**



5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^24d^1$	40 Zr $5s^24d^2$	41 Nb $5s^14d^4$	42 Mo $5s^14d^5$	43 Tc $5s^24d^5$	44 Ru $5s^14d^7$	45 Rh $5s^14d^8$	46 Pd $4d^{10}$	47 Ag $5s^14d^{10}$	48 Cd $5s^24d^{10}$	49 In $5s^25p^1$	50 Sn $5s^25p^2$	51 Sb $5s^25p^3$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^25d^1$	72 Hf $6s^25d^2$	73 Ta $6s^25d^3$	74 W $6s^25d^4$	75 Re $6s^25d^5$	76 Os $6s^25d^6$	77 Ir $6s^25d^7$	78 Pt $6s^15d^9$	79 Au $6s^15d^{10}$	80 Hg $6s^25d^{10}$	81 Tl $6s^26p^1$	82 Pb $6s^26p^2$	83 Bi $6s^26p^3$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^26d^1$	104 Rf $7s^26d^2$	105 Db $7s^26d^3$	106 Sg $7s^26d^4$	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc

Lanthanides	58 Ce $6s^24f^15d$	59 Pr $6s^24f^3$	60 Nd $6s^24f^4$	61 Pm $6s^24f^5$	62 Sm $6s^24f^6$	63 Eu $6s^24f^7$	64 Gd $6s^24f^75d^1$	65 Tb $6s^24f^9$	66 Dy $6s^24f^{10}$	67 Ho $6s^24f^{11}$	68 Er $6s^24f^{12}$
Actinides	90 Th $7s^26d^2$	91 Pa $7s^25f^26d^1$	92 U $7s^25f^36d^1$	93 Np $7s^25f^46d^1$	94 Pu $7s^25f^6$	95 Am $7s^25f^7$	96 Cm $7s^25f^76d^1$	97 Bk $7s^25f^9$	98 Cf $7s^25f^{10}$	99 Es $7s^25f^{11}$	100 Fm $7s^25f^{12}$

> Helium is an exception. Even though it lies in the column with an outer electron configuration of $ns^2\ np^6$, its electron configuration is simply $1s^2$.

Note also that *the number of columns in a block corresponds to the maximum number of electrons that can occupy the particular sublevel of that block*. The *s* block has two columns (corresponding to one *s* orbital holding a maximum of two electrons); the *p* block has six columns (corresponding to three *p* orbitals with two electrons each); the *d* block has ten columns (corresponding to five *d* orbitals with two electrons each); and the *f* block has 14 columns (corresponding to seven *f* orbitals with two electrons each).

Except for helium, *the number of valence electrons for any main-group element is equal to its lettered group number*. For example, we know that chlorine has seven valence electrons because it is in group number 7A.

Lastly, note that, for main-group elements, *the row number in the periodic table is equal to the number (or n value) of the highest principal level*. For example, because chlorine is in row 3, its highest principal level is the $n = 3$ level.

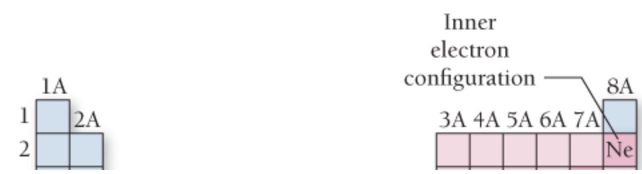
> Recall from Section 3.2 that main-group elements are those in the two far-left columns (groups 1A and 2A) and the six far-right columns (groups 3A–8A) of the periodic table.

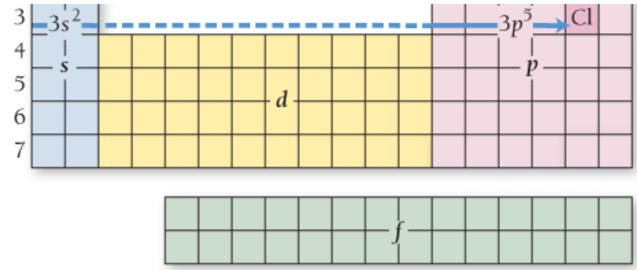
Summarizing Periodic Table Organization

- The periodic table is divisible into four blocks corresponding to the filling of the four quantum sublevels (*s*, *p*, *d*, and *f*).
- The lettered group number of a main-group element is equal to the number of valence electrons for that element.
- The row number of a main-group element is equal to the highest principal quantum number of that element.

Writing an Electron Configuration for an Element from Its Position in the Periodic Table

The organization of the periodic table allows us to write the electron configuration for any element based on its position in the periodic table. For example, suppose we want to write an electron configuration for Cl. The *inner electron configuration* of Cl is that of the noble gas that precedes it in the periodic table, Ne. So we represent the inner electron configuration with [Ne]. We obtain the *outer electron configuration*—the configuration of the electrons beyond the previous noble gas—by tracing the elements between Ne and Cl and assigning electrons to the appropriate orbitals, as shown here. Remember that the highest *n* value is indicated by the row number (3 for chlorine).





We begin with [Ne], then add in the two 3*s* electrons as we trace across the *s* block, followed by five 3*p* electrons as we trace across the *p* block to Cl, which is in the fifth column of the *p* block. The electron configuration is:

$$\text{Cl} \quad [\text{Ne}]\ 3s^2\ 3p^5$$

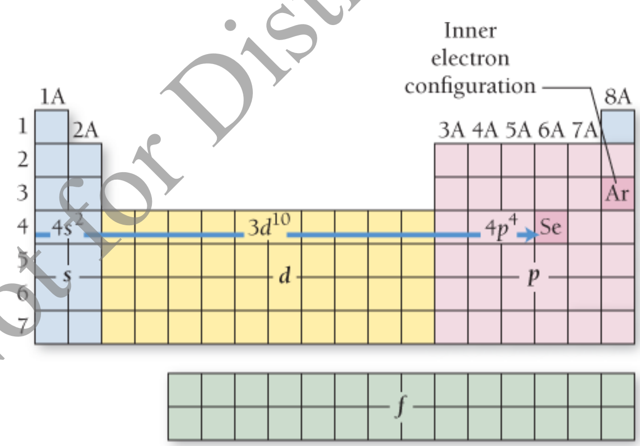
Notice that Cl is in column 7A and therefore has seven valence electrons and an outer electron configuration of $ns^2\ np^5$.

Example 3.4 Writing Electron Configurations from the Periodic Table

Refer to the periodic table to write the electron configuration for selenium (Se).

SOLUTION

The atomic number of Se is 34. The noble gas that precedes Se in the periodic table is argon, so the inner electron configuration is [Ar]. Obtain the outer electron configuration by tracing the elements between Ar and Se and assigning electrons to the appropriate orbitals. Begin with [Ar]. Because Se is in row 4, add two 4*s* electrons as you trace across the *s* block (n = row number). Next, add ten 3*d* electrons as you trace across the *d* block (n = row number − 1). (See explanation after example.) Lastly, add four 4*p* electrons as you trace across the *p* block to Se, which is in the fourth column of the *p* block (n = row number).



$$\text{Se} \quad [\text{Ar}]\ 4s^2\ 3d^{10}\ 4p^4$$

FOR PRACTICE 3.4 Refer to the periodic table to write the electron configuration of bismuth (Bi).

FOR MORE PRACTICE 3.4 Refer to the periodic table to write the electron configuration for iodine (I).

Interactive Worked Example 3.4 Writing Electron Configurations from the Periodic Table

The Transition and Inner Transition Elements

The electron configurations of the transition elements (*d* block) and inner transition elements (*f* block) exhibit trends that differ somewhat from those of the main-group elements. As we move to the right across a row in the *d* block, the *d* orbitals fill as shown here:

4	21 **Sc** $4s^2 3d^1$	22 **Ti** $4s^2 3d^2$	23 **V** $4s^2 3d^3$	24 **Cr** $4s^1 3d^5$	25 **Mn** $4s^2 3d^5$	26 **Fe** $4s^2 3d^6$	27 **Co** $4s^2 3d^7$	28 **Ni** $4s^2 3d^8$	29 **Cu** $4s^1 3d^{10}$	30 **Zn** $4s^2 3d^{10}$
5	39 **Y** $5s^2 4d^1$	40 **Zr** $5s^2 4d^2$	41 **Nb** $5s^1 4d^4$	42 **Mo** $5s^1 4d^5$	43 **Tc** $5s^2 4d^5$	44 **Ru** $5s^1 4d^7$	45 **Rh** $5s^1 4d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^1 4d^{10}$	48 **Cd** $5s^2 4d^{10}$

Note that *the principal quantum number of the d orbitals that fill across each row in the transition series is equal to the row number minus one*. In the fourth row, the 3*d* orbitals fill; in the fifth row, the 4*d* orbitals fill; and so on. This happens because, as we discussed in Section 3.3, the 4*s* orbital is generally lower in energy than the 3*d* orbital (because the 4*s* orbital more efficiently penetrates into the region occupied by the core electrons). The result is that the 4*s* orbital fills before the 3*d* orbital, even though its principal quantum number $(n = 4)$ is higher.

Keep in mind, however, that the 4*s* and the 3*d* orbitals are extremely close to each other in energy so that their relative energy ordering depends on the exact species under consideration; this causes some irregular behavior in the transition metals. For example, in the first transition series of the *d* block, the outer configuration is $4s^2\ 3d^x$ with two exceptions: Cr is $4s^1\ 3d^5$ and Cu is $4s^1\ 3d^{10}$. This behavior is related to the closely spaced 3*d* and 4*s* energy levels and the stability associated with a half-filled (as in Cr) or completely filled (as in Cu) sublevel. Actual electron configurations are determined experimentally (through spectroscopy) and do not always conform to the general pattern. Nonetheless, the patterns we have described allow us to accurately predict electron configurations for most of the elements in the periodic table.

As we move across the *f* block (the inner transition series), the *f* orbitals fill. For these elements, the principal quantum number of the *f* orbitals that fill across each row is the row number *minus two*. (In the sixth row, the 4*f* orbitals fill, and in the seventh row, the 5*f* orbitals fill.) In addition, within the inner transition series, the close energy spacing of the 5*d* and 4*f* orbitals sometimes causes an electron to enter a 5*d* orbital instead of the expected 4*f* orbital. For example, the electron configuration of gadolinium is $[\mathrm{Xe}]\ 6s^2\ 4f^2\ 5d^1$ (instead of the expected $[\mathrm{Xe}]\ 6s^2\ 4f^8$).