Answer to Essential Question 28.3: If $n = 3$, then the orbital quantum number ℓ can take on 3 possible values, 0, 1, and 2. If $\ell = 0$, $m_\ell = 0$, and there are two possible spin states – thus, there are 2 states with $\ell = 0$. If $\ell = 1$, $m_\ell = -1$, 0, or $+1$, each with two possible spin states – thus, there are 6 states with $\ell = 1$. If $\ell = 2$, $m_\ell = -2, -1, 0, +1$, or $+2$, each with two possible spin states – thus, there are 10 states with $\ell = 2$. That is a total of 18 states with $n = 3$.

28-4 The Pauli Exclusion Principle

In Section 28-3, we discussed the fact that it takes four quantum numbers to completely specify the state of an electron. In this section and in Section 28-5, we examine how these four quantum numbers determine the structure of the periodic table. To understand how these four numbers can determine a structure as complex as the periodic table, we begin with the Pauli exclusion principle, which is a simple statement with far-reaching consequences.

The Pauli exclusion principle: No two electrons can simultaneously occupy the same electron state in an atom. In other words, no two electrons in an atom can simultaneously have the same set of four quantum numbers.

The exclusion principle is named for the Austrian physicist Wolfgang Pauli (1900–1958), shown in the picture in Figure 28.7. Pauli was awarded the Nobel Prize in Physics in 1945 for the exclusion principle.
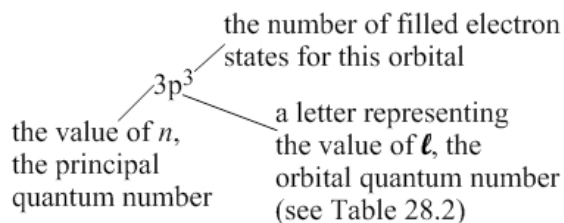

Figure 28.7: A photograph of Wolfgang Pauli, a colorful character who came up with the exclusion principle. Photo credit: Fermilab employees, via Wikimedia Commons (a public-domain image).

With the exclusion principle in mind, let's examine the ground-state (lowest energy) configurations of various atoms. In general, equation 28.3, in which the energy of an electron state is determined solely by the principal quantum number, is inadequate. The energy is also determined by the electron's orbital angular momentum. In general, for a given value of the principal quantum number, n, the higher the value of the orbital quantum number, ℓ, the higher the energy of that state. Let's begin by looking at the ground state configurations of a few elements in the periodic table. These are shown in Table 28.1. An explanation of what the number-letter-number notation means is given in Figure 28.8.

Atomic number	Name (symbol)	Ground-state configuration
2	Helium (He)	$1s^2$
6	Carbon (C)	$1s^2\, 2s^2\, 2p^2$
10	Neon (Ne)	$1s^2\, 2s^2\, 2p^6$
15	Phosphorus (P)	$1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^3$
20	Calcium (Ca)	$1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2$
26	Iron (Fe)	$1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^4$

Table 28.1: Ground-state configurations for selected atoms.

$3p^3$

the number of filled electron states for this orbital

the value of n, the principal quantum number

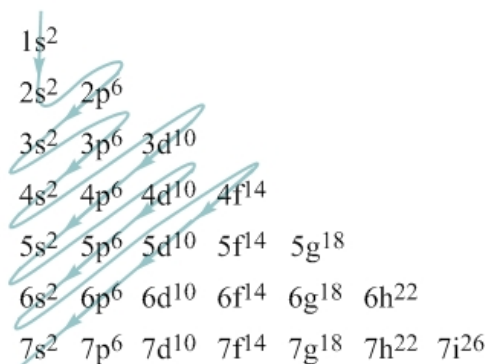a letter representing the value of ℓ, the orbital quantum number (see Table 28.2)

Figure 28.8: Explaining the number-letter-number notation.

Referring to Table 28.1, we see that helium, with only two electrons, can have both electrons in the $n = 1$ shell, which can only fit two electrons. For atoms that have more than 2 electrons, two electrons are in the $n = 1$ level, and then the others are in states that have larger n values. The $n = 2$ level can fit eight electrons (two with $\ell = 0$ and six with $\ell = 1$). Thus, for elements up to and including neon, which has 10 electrons in its ground-state configuration, the electrons are in the $n = 1$ and $n = 2$ levels. Beyond this, electron state with higher n values come into play.

The layout of the ground-state configurations in Table 28.1 needs further explanation. This is provided by Table 28.2, which shows what value of ℓ the various letters correspond to, and by Figure 28.9, which shows how the various orbitals compare to one another, in terms of their energy level. As you wind your way through Figure 28.9 from top to bottom, the energy of a particular orbital increases. This figure shows the order in which electrons fill states in the ground-state configuration, because the ground-state configuration minimizes the total energy of the electrons.

Value of ℓ	Letter (stands for)
0	s (sharp)
1	p (principal)
2	d (diffuse)
3	f (fundamental)
4	g (letter after f)
5	h (letter after g)
6	i (letter after h)

Table 28.2: Letters for various ℓ values.

$1s^2$
$2s^2$ $2p^6$
$3s^2$ $3p^6$ $3d^{10}$
$4s^2$ $4p^6$ $4d^{10}$ $4f^{14}$
$5s^2$ $5p^6$ $5d^{10}$ $5f^{14}$ $5g^{18}$
$6s^2$ $6p^6$ $6d^{10}$ $6f^{14}$ $6g^{18}$ $6h^{22}$
$7s^2$ $7p^6$ $7d^{10}$ $7f^{14}$ $7g^{18}$ $7h^{22}$ $7i^{26}$

Figure 28.9: The order in which electrons fill the different subshells.

Terminology

An **atomic orbital** is both the mathematical function that describes where an electron can be, and the corresponding region in space. It takes three quantum numbers (n, ℓ, and m_ℓ) to define an orbital. Because there are two possible values of the spin quantum number, each atomic orbital can contain two electrons.

A **subshell** is the set of orbitals with the same values of n and ℓ.

A **shell** is the set of orbitals with the same value of n.

Figure 28.9 is consistent with what is known as the **aufbau principle**, derived from the German word *aufbauprinzip*, which means "building-up principle." The basic rule is that subshells are filled in order of lowest $n + \ell$ values. In cases of equal $n + \ell$ values, the subshells with the lower n value are filled first. The diagonal lines in Figure 28.9 represent subshells that have equal $n + \ell$ values, with those with the lower n values coming first.

Related End-of-Chapter Exercises: 10, 11, 57.

Essential Question 28.4: What is the minimum atomic number an atom should have before $4d^x$, where x represents a positive integer, appears in the atom's ground-state configuration? What is the range of values x can take on?