**Answer to Essential Question 28.1:** (a) In general, four energy levels give six photon energies. One way to count these is to start with the highest level, $-12$ eV. An electron starting in the $-12$ eV level can drop to any of the other three levels, giving three different photon energies. An electron starting at the $-15$ eV level can drop to either of the two lower levels, giving two more photon energies. Finally, an electron can drop from the $-21$ eV level to the $-31$ eV level, giving one more photon energy (for a total of six). (b) The minimum photon energy corresponds to the $3$ eV difference between the $-12$ eV level and the $-15$ eV level. The maximum photon energy corresponds to the $19$ eV difference between the $-12$ eV level and the $-31$ eV level.

**28-2 Models of the Atom**

It is amazing to think about how far we have come, in terms of our understanding of the physical world, in the last century or so. A good example of our progress is how much our model of the atom has evolved. Let’s spend some time discussing the evolution of atomic models.

**Ernest Rutherford probes the plum-pudding model**

J. J. Thomson, who discovered the electron in 1897, proposed a plum-pudding model of the atom. In this model, electrons were thought to be embedded in a ball of positive charge, like raisins are embedded in a plum pudding. Ernest Rutherford (1871 - 1937) put this model to the test by designing an experiment that involved firing alpha particles (helium nuclei) at a very thin film of gold. The experiment was carried out in Rutherford’s lab by Hans Geiger and Ernest Marsden. If the plum-pudding model was correct, the expectation was that the alpha particles should make it through the ball of spread-out positive charge with very little deflection. For the most part, this was the case; however, a small fraction of the alpha particles were deflected through large angles, with some even being deflected through $180^\circ$. Rutherford made a famous statement about this, which was “It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.” Through a careful analysis of the results, Rutherford determined that the positive charge of the atom was not spread out throughout the volume occupied by the atom, but was instead concentrated in a tiny volume, orders of magnitude smaller than that of the atom, which we now call the **nucleus**.

**Niels Bohr provides a theoretical framework for Rydberg’s equation**

The next major advance in our understanding of the atom came from the Danish physicist, Niels Bohr, who incorporating ideas from Rutherford and quantum ideas. In Bohr’s model, electrons traveled in circular orbits around a central nucleus, similar to the way planets travel around the Sun. It is important to understand that the Bohr model does not reflect reality, but it provides a basis for our understanding of the atom. With an analysis based on principles of physics we have discussed earlier in this book, such as the attractive force between charged particles, Bohr was able to show that the quantized energy levels in hydrogen were completely consistent with Rydberg’s equation (see Equation 28.1 in section 28-1) for the wavelengths of light emitted by hydrogen. In Bohr’s model, the angular momenta of the electrons are also quantized, a result we also accept today. Where the Bohr model breaks down is in the electron orbits. In the Bohr model, the electrons are confined to planar orbits of very particular radii. This is not at all the modern view of the atom, which we understand using quantum mechanics.

**The modern view of the atom**

Over the course of the 20th century, many people, Bohr included, contributed to furthering our understanding of the atom. We will spend some time in Chapter 29 exploring the nucleus, so for the moment let us focus our attention on the electrons in the atom. As far as the electrons are concerned, the nucleus can be thought of as a tiny ball of positive charge.
In the Bohr model of the atom, the electrons are found only in certain orbits, with the radii of the orbits being quantized, so an electron will never be found at other distances from the nucleus. Our modern understanding is rather different. Now, we talk about the probability of finding the electron at a particular distance from the nucleus. For an electron in the **ground state** (the lowest-energy state) of the hydrogen atom, for instance, the Bohr model states that the electron is a distance of $5.29 \times 10^{-11}$ m from the nucleus (this is known as the **Bohr radius**). The modern view of where the electron in hydrogen’s ground state is located is illustrated by Figure 28.4. Even in the modern view, the most likely place to find this ground-state electron is at a distance of one Bohr radius from the nucleus. However, as the graph shows, the electron can be found at any distance from the nucleus, aside from right at the nucleus or infinitely far away.

**Figure 28.4**: A graph of the probability, per unit length, of finding the electron in the hydrogen ground state at various distances from the nucleus. The total area under the curve, when the curve is extended to infinity, is 1 — the electron is 100% likely to be found between $r = 0$ and $r = \infty$.

Graphs like that in Figure 28.4 come from solving the Schrödinger equation, which is essentially conservation of energy applied to the atom. Solutions to the Schrödinger equation are called wave functions. The square of a wave function gives the probability of finding a particle in a particular location. The Schrödinger equation is named for the Austrian physicist Erwin Schrödinger (1887 – 1961), who shared the 1933 Nobel Prize in Physics for his contributions to quantum mechanics.

**Quantum tunneling**

If you throw a tennis ball against a solid wall, the ball will never make it to the far side of the wall unless you give it enough kinetic energy to pass over the top of the wall. Such rules do not apply to quantum particles. If the wave function of a quantum particle extends through a barrier to the far side, then there is some probability of finding the particle on the far side of the barrier. Even if the particle’s energy is insufficient to carry it over the barrier, the particle will eventually be found on the far side of the barrier. This process, of passing through a barrier, is known as **quantum tunneling**.

Quantum tunneling is exploited in scanning tunneling microscopes (STM’s), in which a very sharp tip is scanned over the surface. By measuring the rate at which electrons tunnel across the gap between this tip and the surface, a two-dimensional picture of the surface can be created. The chapter-opening picture shows an image of a quantum corral. First, a corral of iron atoms was created on a surface by pulling the atoms into place with an STM. The STM was then used to scan the surface, to visualize electrons trapped inside the corral. Note how the electrons are wave-like in this situation, and nor particle-like.

**Related End-of-Chapter Exercises: 5 – 7.**

**Essential Question 28.2**: Return to the graph in Figure 28.4. Is the ground-state electron in hydrogen more likely to be found at a position closer than 1 Bohr radius from the nucleus, or farther than 1 Bohr radius from the nucleus? Using the graph, estimate the relative probability of finding the electron in these two ranges.
Answer to Essential Question 28.2: The area under the curve for the region beyond 1 Bohr radius from the nucleus is clearly larger than the area for the region less than 1 Bohr radius away from the nucleus – thus, the electron is more likely to be found farther than 1 Bohr radius away. A reasonable estimate for the ratio of the areas is 2:1. In other words, the electron, when in the ground state of hydrogen, is approximately twice as likely to be found at a radius farther than 1 Bohr radius from the nucleus than it is to be found at a distance less than one Bohr radius.

28-3 The Quantum Mechanical View of the Atom

In the Bohr model of the atom, only one quantum number, \( n \), was used. With that single quantum number, Bohr came up with expressions for the energies of quantized energy levels, quantized angular momenta of the electrons, and quantized radii of the electron orbits. Our modern view of the atom, applying the equations of quantum mechanics, is a little more complicated, requiring four quantum numbers to completely specify the various electron energy states. These quantum numbers are:

1. \( n \), the principal quantum number. To a first approximation, the energies of the quantized energy levels in an atom with an atomic number \( Z \) and just one electron are given by:

\[
E_n = \left(\frac{13.6 \text{ eV}}{n^2}\right) Z^2, \quad \text{(Eq. 28.3: Energies of the electron levels)}
\]

where \( n = 1, 2, 3, \ldots \)

2. \( \ell \), the orbital quantum number. This quantum number quantizes the magnitude of the electron’s orbital angular momentum (this is somewhat analogous to the angular momentum associated with a planet’s orbit around the Sun). The magnitude of the orbital angular momentum is given by:

\[
L = \sqrt{\ell(\ell + 1)} \frac{\hbar}{2\pi}, \quad \text{(Eq. 28.4: Orbital angular momentum)}
\]

where \( \ell = 0, 1, 2, \ldots n - 1 \).

3. \( m_\ell \), the magnetic quantum number. This quantum number quantizes the direction of the electron’s angular momentum (this is known as space quantization, and is illustrated in Figure 28.5). Conventionally, we say that this quantum number defines the \( z \)-component of the angular momentum. For a given value of the orbital quantum number, the different values of \( m_\ell \) give electron states of the same energy unless a magnetic field is present, in which case the states have different energies (this is called the Zeeman effect).

\[
L_z = m_\ell \frac{\hbar}{2\pi}, \quad \text{where } m_\ell = -\ell, -\ell + 1, \ldots \ell - 1, \ell.
\]

(Eq. 28.5: \( z \)-component of the orbital angular momentum)

Figure 28.5: The allowed directions of the orbital angular momentum, when \( \ell = 2 \). The half-circle has a radius of \( \sqrt{6} \) units. The units on both axes are angular momentum units of \( \hbar/(2\pi) \).
4. \( m_s \), the spin quantum number. The concept of electron spin is somewhat analogous to the angular momentum associated with a planet’s rotation about its own axis. The electron has two possible spin states, which we refer to as spin up and spin down. The two possible \( z \)-components of the spin angular momentum, \( S_z \), are given by:

\[
S_z = m_s \frac{\hbar}{2\pi}, \quad \text{(Eq. 28.6: Spin angular momentum)}
\]

where \( m_s = \pm \frac{1}{2} \).

**Visualizing the wave functions**

Figure 28.6 shows a common way to visualize the various electron wave functions for hydrogen. Remember that there is only one electron in the hydrogen atom, but the electron has an infinite number of states that it can choose between – some of the lower-energy states are shown below. When we look at pictures like this, we again interpret them in terms of probability. The brighter the picture is at a particular point, the more likely it is to find the electron at that point. Although the pictures are two-dimensional, the wave functions are three-dimensional.

The image in the top left of Figure 28.6 corresponds to the graph in Figure 28.4. At first glance, they appear to contradict one another. For instance, the graph in Figure 28.4 shows that the probability of finding the electron at \( r = 0 \) is zero, while the top left image in Figure 28.6 has the brightest (highest probability) spot right in the center, where \( r = 0 \). The probability of finding the electron at any particular point, not just the center point, is zero, because a point has no volume. It makes more sense to discuss the probability of finding the electron within a particular volume. Both the graph in Figure 28.4 and all six diagrams below are best interpreted that way. From that perspective, the diagram at the upper left and the graph in Figure 28.4 are consistent.

**Figure 28.6**: A visualization of the wave functions (or, atomic orbitals) for the \( n = 1, 2, \) and 3 levels for hydrogen. The brighter it is in a particular region, the more likely it is to find the electron in that region. Conversely, there is no chance of finding the electron in a region that is black. The numbers on the right side of the diagram show values of \( n \), the principal quantum number, for each row. The letters across the top represent the values of the orbital quantum number, \( \ell \), for a particular column. Only one wave function is shown in the first row because, when \( n = 1 \), the only possible value of the orbital quantum number is \( \ell = 0 \). Similarly, there are only two images in the second row because there are two allowed \( \ell \) values when \( n = 2 \). The wave functions are positive in the red areas and negative in the blue areas. In all cases shown, \( m_\ell = 0 \).

**Related End-of-Chapter Exercises**: 8, 9, 19 – 23.

**Essential Question 28.3**: How many different electron states are there that have \( n = 3 \)?
Answer to Essential Question 28.3: If \( n = 3 \), then the orbital quantum number \( \ell \) 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, \( \ell \), 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.

<table>
<thead>
<tr>
<th>Atomic number</th>
<th>Name (symbol)</th>
<th>Ground-state configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Helium (He)</td>
<td>1s^2</td>
</tr>
<tr>
<td>6</td>
<td>Carbon (C)</td>
<td>1s^2 2s^2 2p^2</td>
</tr>
<tr>
<td>10</td>
<td>Neon (Ne)</td>
<td>1s^2 2s^2 2p^6</td>
</tr>
<tr>
<td>15</td>
<td>Phosphorus (P)</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^3</td>
</tr>
<tr>
<td>20</td>
<td>Calcium (Ca)</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^6 4s^2</td>
</tr>
<tr>
<td>26</td>
<td>Iron (Fe)</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4</td>
</tr>
</tbody>
</table>

**Table 28.1:** Ground-state configurations for selected atoms.

**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 $\ell$ 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.

### Table 28.2: Letters for various $\ell$ values.

<table>
<thead>
<tr>
<th>Value of $\ell$</th>
<th>Letter (stands for)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>s (sharp)</td>
</tr>
<tr>
<td>1</td>
<td>p (principal)</td>
</tr>
<tr>
<td>2</td>
<td>d (diffuse)</td>
</tr>
<tr>
<td>3</td>
<td>f (fundamental)</td>
</tr>
<tr>
<td>4</td>
<td>g (letter after f)</td>
</tr>
<tr>
<td>5</td>
<td>h (letter after g)</td>
</tr>
<tr>
<td>6</td>
<td>i (letter after h)</td>
</tr>
</tbody>
</table>

**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 ($n$, $\ell$, and $m_\ell$) 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 $\ell$.

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 *aufbau-prinzip*, 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?
Answer to Essential Question 28.4: The atom should have an atomic number of at least 39. Following the twists and turns of Figure 28.9, we can count that there are 38 electron states up to and including the 5s subshell. Thus, the 39th electron should go into the 4d subshell. Because the 4d subshell can contain as many as 10 electrons, the \( x \) in \( 4d^x \) is an integer between 1 and 10.

28-5 Understanding the Periodic Table

In the previous two sections, we have laid the groundwork for understanding the periodic table. An idealized periodic table is laid out in Figure 28.10. Table 28.2 and Figure 28.9 are repeated here from Section 28-4 so you can more easily see the connection between the diagram that goes with the aufbau principle (Figure 28.9) and the periodic table.

<table>
<thead>
<tr>
<th>Value of ( \ell )</th>
<th>Letter (stands for)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>s (sharp)</td>
</tr>
<tr>
<td>1</td>
<td>p (principal)</td>
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<tr>
<td>2</td>
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<tr>
<td>3</td>
<td>f (fundamental)</td>
</tr>
<tr>
<td>4</td>
<td>g (letter after f)</td>
</tr>
<tr>
<td>5</td>
<td>h (letter after g)</td>
</tr>
<tr>
<td>6</td>
<td>i (letter after h)</td>
</tr>
</tbody>
</table>

Table 28.2: Letters for various \( \ell \) values.

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

Figure 28.10 shows an idealized periodic table, showing the state of the last few electrons in each element’s ground-state configuration. Atoms with completely filled subshells, or half-filled subshells, are particularly stable, and atoms with similar last-electron states generally have similar chemical properties. Note that the true ground-state configurations of elements with atomic numbers 105 and above are not yet known.

In reality, there are some deviations from the idealized behavior shown in Figure 28.10. These deviations can be quite instructive. Some of those deviations include:

- In column 11, the last two terms in the ground-state configurations for copper (Cu), silver (Ag), and gold (Au) are actually \( 4s^1\ 3d^{10} \), \( 5s^1\ 4d^{10} \), and \( 6s^1\ 5d^{10} \), respectively. Each of these configurations ends with a half-full s subshell and a full d subshell, rather than a full s subshell and an almost-full d subshell, demonstrating that full and half-full subshells are particularly stable (lower energy) states.

- In column 6, the last two terms in the ground-state configurations for chromium (Cr) and molybdenum (Mo) are actually \( 4s^1\ 3d^{5} \) and \( 5s^1\ 4d^{5} \), respectively. Each of these configurations ends with a half-full subshell and a half-full d subshell, rather than a full s subshell and partly-filled d subshell, demonstrating that half-full subshells are particularly stable (lower energy) states.

- The last two terms in the ground-state configurations for gadolinium (Gd, element 64) and curium (Cm, element 96) are actually \( 4f^7\ 5d^1 \) and \( 5f^7\ 6d^1 \), respectively. Each of these configurations has a half-full f subshell and one electron in a d subshell, rather than partly-filled f subshell, again demonstrating that half-full subshells are particularly stable (lower energy) states.
Figure 28.10: This figure lays out an idealized version of the periodic table, to show the predicted term in the ground-state configuration for each element. When viewed this way, one gets a better understanding of why the table is laid out as it is. In a given column, for instance, the elements have similar chemical properties because they have similar electron configurations. In the 17th column, for instance, these elements (known as halogens) have similar properties because in their ground-state configurations, each of the elements has 5 electrons in a p orbital, that p orbital being the highest-energy orbital that contains electrons for those elements. For most of the elements, the last term in their ground-state configurations match what is shown here, but some elements differ from the idealized version shown here, as detailed on the previous page. Also, note that helium is generally shown in the 18th column, because it is just as unreactive as the other noble gases, but you could make a strong argument for helium belonging at the top of column 2, based on its ground-state configuration.


Essential Question 28.5: In Figure 28.10, the final term in the ground-state configuration of the element with the chemical symbol Po (element 84) is shown to be 6p^4. What is the common name for this element? What is the complete ground-state configuration for this element, showing all the terms?
**Answer to Essential Question 28.5**: The common name for the element with the chemical symbol Po is polonium (not potassium, which has a chemical symbol of K). To write out the complete ground-state configuration of polonium, we can wind our way through Figure 28.9, starting from the top: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁶ 6s² 4f⁴ 5d¹⁰ 6p⁴.

**28-6 Some Applications of Quantum Mechanics**

**Lasers**

The word laser is an acronym, coming from the phrase light amplification by the stimulated emission of radiation. In our modern world, there are many applications of lasers. Such applications include bar-code readers in stores; surgery, particularly in eye surgery (including LASIK, Laser-Assisted In Situ Keratomileusis), where it is absolutely critical to do precision cutting; laser pointers, laser printers, and CD and DVD players; and fiber-optic communications.

The vast majority of lasers manufactured worldwide are diode lasers, in which light is produced from carefully constructed layers of semiconductors. In this section, we will focus on a different kind of laser, the helium-neon (or HeNe) laser that you may have seen in class or even used in a physics lab experiment. It is quite common to do demonstrations or experiments showing diffraction or interference of light using HeNe lasers.

As the name suggests, a helium-neon laser contains a mixture of helium and neon gas, at a relatively low pressure, with many more helium atoms than neon atoms. Coincidentally, the difference in energy between two of the electron energy levels in helium is almost the same as the difference in energy between two of the electron energy levels in neon—this is why these two elements are used. A high-voltage electrical discharge through the gas will excite helium atoms from their ground state to one of the \( n = 2 \) states, requiring an energy difference of 20.61 eV. These \( n = 2 \) states are metastable, which means that the electron will not immediately drop down to the ground state—it will remain in the excited state for a while.

With an extra 0.05 eV worth of kinetic energy, the excited helium atoms, when they collide with neon atoms that are in the ground state, can transfer 20.66 eV of energy to the neon atoms, just what is required to boost an electron in neon from the ground state to one of the \( n = 3 \) states (specifically, the 3s state). This level is also metastable, but electrons in some of these excited states will spontaneously drop down to the 2p state, emitting a photon of 632.8 nm (in air), corresponding to the wavelength of light emitted by a typical red HeNe laser. These photons interact with the excited neon atoms, which encourages them also to make the 3s to 2p transition. This part of the process is the stimulated emission that is part of what laser stands for. An energy-level diagram for the HeNe laser is shown in Figure 28.11.

**Figure 28.11**: An energy-level diagram for the HeNe laser. Because the energy difference between the two lowest energy levels in helium almost exactly equals the energy difference between the \( n = 1 \) and \( n = 3 \) levels in neon, excited helium atoms can transfer energy to ground-state neon atoms via collisions. Note that, despite what it looks like in the diagram, the \( E_1 \) levels in helium and neon are at completely different energies. What is critical, however, is that the difference between two of the helium levels almost exactly matches the difference between two of the neon levels.
Once the light has been created, by the process of stimulated emission inside the laser, a laser beam must then be created. The photons and the low-pressure mixture of helium and neon are contained in a tube that is typically 15 – 50 cm in length. At one end of the tube is a highly reflective mirror, while at the other end of the tube (the end that the beam emerges from) is a mirror that reflects most of the light, but which allows a little light (about 1%) to pass through. Generally, the photons emerge after bouncing back and forth many times between the mirrors, resulting in a beam of light that has very little spread.

The design of the mirrors inside the laser is also interesting, because the mirrors exploit thin-film interference for a wavelength corresponding to the wavelength of light emitted by the laser. There are several different electron transitions associated with helium and neon, so a HeNe laser can actually emit several different wavelengths (a different wavelength for each transition between electron energy levels). By adjusting the thin films on the mirrors inside the laser, the laser can be optimized for the emission of the standard 632.8 nm red light, or a different wavelength. The first HeNe laser ever made, in the 1960’s, for instance, emitted ultraviolet light with a wavelength of 1150 nm.

**Fluorescence and phosphorescence**

Two more applications of quantum mechanics, again associated with electron energy levels and the photons that are emitted when electrons make a transition from a higher-energy state to a lower-energy state, are fluorescence and phosphorescence. These two phenomena are similar, in that exposing a fluorescent or phosphorescent material to (usually) ultraviolet light will excite electrons from lower-energy states to higher-energy states. When the electrons drop back toward the lower-energy state, however, they do so by dropping down a smaller step in energy, to an intermediate level. If this smaller energy causes photons to be emitted in the visible spectrum, we can then see them. In fluorescent materials, the phenomenon is present only when the light source that excites the upward transitions is present. In phosphorescent materials, it takes a long time, on average, before the excited electrons make a transition to a lower level. Because the electron transitions occur over a long time period, visible light continues to be emitted long after the light source that excites the upward transitions is removed. A photograph of a variety of fluorescent minerals is shown in Figure 28.12.

**Figure 28.12:** A variety of fluorescent minerals, photographed while they are being exposed to ultraviolet light. In general, these minerals are much less colorful when they are not fluorescing. Image credit: Hannes Grobe, via Wikimedia Commons.

**Related End-of-Chapter Exercises:** 12, 27.

**Essential Question 28.6:** Consider the information given in Figure 28.11. (a) If the $n = 3$ to $n = 2$ transition in neon produces a photon with a wavelength of 632.9 nm (in vacuum), what is the difference in energy between these two levels? (b) What is the difference in energy between the $n = 2$ and $n = 1$ levels in neon?