***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°. 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 20$^{th}$ 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.

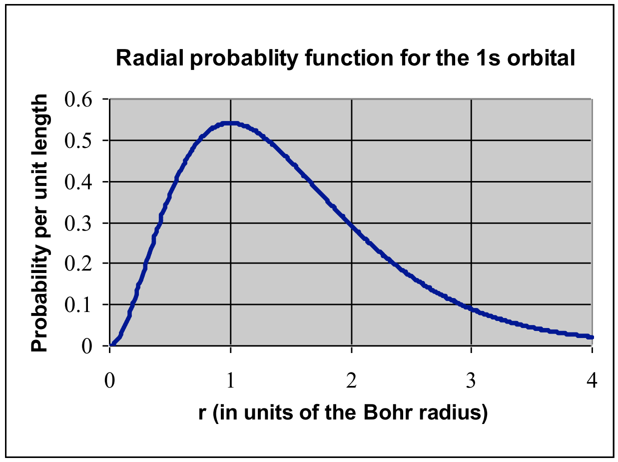


**Figure 28.4**: A graph of the probability, per unit length, of finding the electron in the hydrogen ground state at various distances from the nucleus. The total area under the curve, when the curve is extended to infinity, is 1 – the electron is 100% likely to be found between $r = 0$ and $r =$ infinity.

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.