28–5 Quantum-Mechanical View of Atoms

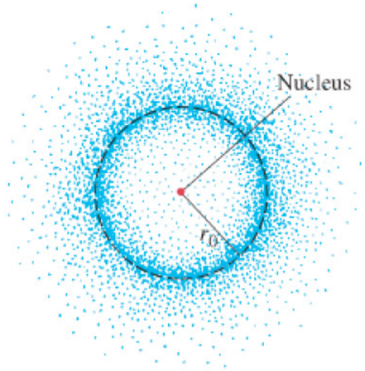

FIGURE 28–6 Electron cloud or "probability distribution" for the ground state of the hydrogen atom. The dashed circle represents the Bohr radius. (This is a 2-dimensional slice through the atom that includes the nucleus.)

At the beginning of this Chapter, we discussed the limitations of the Bohr theory of atomic structure. Now we examine the quantum-mechanical theory of atoms, which is far more complete than the old Bohr theory. Although the Bohr model has been discarded as an accurate description of nature, nonetheless, quantum mechanics reaffirms certain aspects of the older theory, such as that electrons in an atom exist only in discrete states of definite energy, and that a photon of light is emitted (or absorbed) when an electron makes a transition from one state to another. But quantum mechanics is a much deeper theory, and has provided us with a very different view of the atom. According to quantum mechanics, electrons do not exist in well-defined circular orbits as in the Bohr theory. Rather, the electron (because of its wave nature) might be thought of as spread out in space as a "cloud." The size and shape of the electron cloud can be calculated for a given state of an atom. For the ground state in the hydrogen atom, the electron cloud is spherically symmetric, as shown in Fig. 28–6. The electron cloud roughly indicates the "size" of an atom. But just as a cloud may not have a distinct border, atoms do not have a precise boundary or a well-defined size. Not all electron clouds have a spherical shape, as we shall see later in this Chapter.

The electron cloud can be interpreted from either the particle or the wave viewpoint. Remember that by a particle we mean something that is localized in space—it has a definite position at any given instant. By contrast, a wave is spread out in space. The electron cloud, spread out in space as in Fig. 28–6, is a result of the wave nature of electrons. Electron clouds can also be interpreted as

Probability distributions

probability distributions for a particle. If you were to measure the position of an electron in a hydrogen atom 500 different times, the majority of the results would show the electron at points where the probability is high (dark area in Fig. 28–6). Only occasionally would the electron be found where the probability is low.

28–6 Quantum Mechanics of the Hydrogen Atom; Quantum Numbers

We now look more closely at what quantum mechanics tells us about the hydrogen atom. Much of what we say here also applies to more complex atoms, which are discussed in the next Section.

Quantum mechanics is a much more sophisticated and successful theory than Bohr's. Yet in a few details they agree. Quantum mechanics predicts the same basic energy levels (Fig. 27–27) for the hydrogen atom as does the Bohr theory. That is,

$$E_n = -\frac{13.6\text{ eV}}{n^2}, \qquad n = 1, 2, 3, \cdots,$$

where n is an integer. In the simple Bohr theory, there was only one quantum number, n. In quantum mechanics, it turns out that four different quantum numbers are needed to specify each state in the atom:

Principal quantum number, n

(1) The *quantum number, n*, from the Bohr theory is found also in quantum mechanics and is called the **principal quantum number**. It can have any integer value from 1 to ∞. The total energy of a state in the hydrogen atom depends on n, as we saw above.

Orbital quantum number, l $0 \le l \le n-1$

(2) The **orbital quantum number**, l, is related to the magnitude of the angular momentum of the electron; l can take on integer values from 0 to $(n-1)$. For the ground state, $n = 1$, l can only be zero.† For $n = 3$, l can be 0, 1, or 2.

†Contrast this with the Bohr theory, which assigned $l = 1$ to the ground state (Eq. 27–11).