For $n = 2$, l can be 0 or 1. The distribution for $n = 2$, $l = 0$ is shown in Fig. 28–9a, and it is seen to differ from that for the ground state (Fig. 28–6), although it is still spherically symmetric. For $n = 2$, $l = 1$, the distributions are not spherically symmetric as shown in Figs. 28–9b (for $m_l = 0$) and 28–9c (for $m_l = +1$ or -1).

Although the spatial distributions of the electron can be calculated for the various states, it is difficult to measure them experimentally. Most of the experimental information about the atom has come from a careful examination of the emission spectra under various conditions.

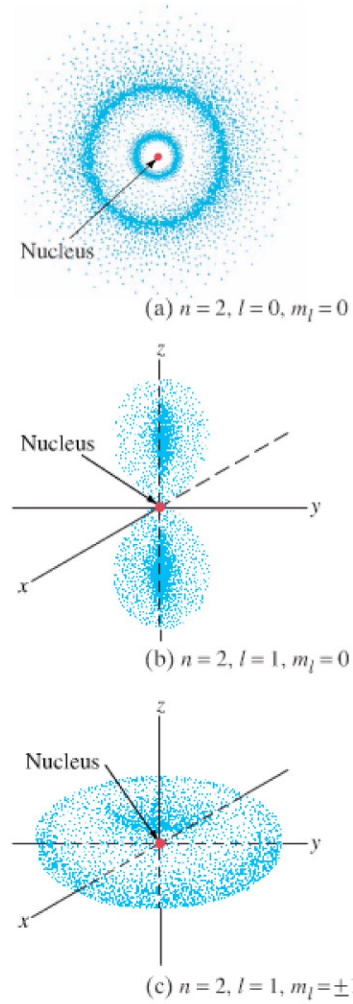

FIGURE 28–9 Electron cloud, or probability distribution, for $n = 2$ states in hydrogen.

* Selection Rules: Allowed and Forbidden Transitions

Another prediction of quantum mechanics is that when a photon is emitted or absorbed, transitions can occur only between states with values of l that differ by exactly one unit:

$$\Delta l = \pm 1.$$

According to this **selection rule**, an electron in an $l = 2$ state can jump only to a state with $l = 1$ or $l = 3$. It cannot jump to a state with $l = 2$ or $l = 0$. A transition such as $l = 2$ to $l = 0$ is called a **forbidden transition**. Actually, such a transition is not absolutely forbidden and can occur, but only with very low probability compared to **allowed transitions**—those that satisfy the selection rule $\Delta l = \pm 1$. Since the orbital angular momentum of an H atom must change by one unit when it emits a photon, conservation of angular momentum tells us that the photon must carry off angular momentum. Indeed, experimental evidence of many sorts shows that the photon can be assigned a spin angular momentum of $1\hbar$.

28–7 Complex Atoms; the Exclusion Principle

We have discussed the hydrogen atom in detail because it is the simplest to deal with. Now we briefly discuss more complex atoms, those that contain more than one electron, and whose energy levels can be determined experimentally from an analysis of the emission spectra. The energy levels are *not* the same as in the H atom, since the electrons interact with each other as well as with the nucleus. Each electron in a complex atom still occupies a particular state characterized by the same quantum numbers n, l, m_l, and m_s. For atoms with more than one electron, the energy levels depend on both n and l.

The number of electrons in a neutral atom is called its **atomic number**, Z; Z is also the number of positive charges (protons) in the nucleus, and determines what kind of atom it is. That is, Z determines most of the properties that distinguish one atom from another.

Quantum mechanics in the years after 1925 proved successful also in dealing with complex atoms. The mathematics becomes very difficult, however, since in multi-electron atoms, each electron is not only attracted to the nucleus but is repelled by the other electrons.

To understand the possible arrangements of electrons in an atom, a new principle was needed. It was introduced by Wolfgang Pauli (1900–1958; Fig. 28–2) and is called the **Pauli exclusion principle**. It states:

Pauli exclusion principle

No two electrons in an atom can occupy the same quantum state.

Thus, no two electrons in an atom can have exactly the same set of the quantum numbers n, l, m_l, and m_s. The Pauli exclusion principle† forms the basis not only for understanding complex atoms, but also for understanding molecules and bonding, and other phenomena as well.

†The exclusion principle applies to identical particles whose spin quantum number is a half-integer ($\frac{1}{2}$, $\frac{3}{2}$, and so on), including electrons, protons, and neutrons; such particles are called **fermions** (after Enrico Fermi who derived a statistical theory describing them). The exclusion principle does not apply to particles with integer spin (0, 1, 2, and so on), such as the photon and π meson, all of which are referred to as **bosons** (after Satyendranath Bose, who derived a statistical theory for them).