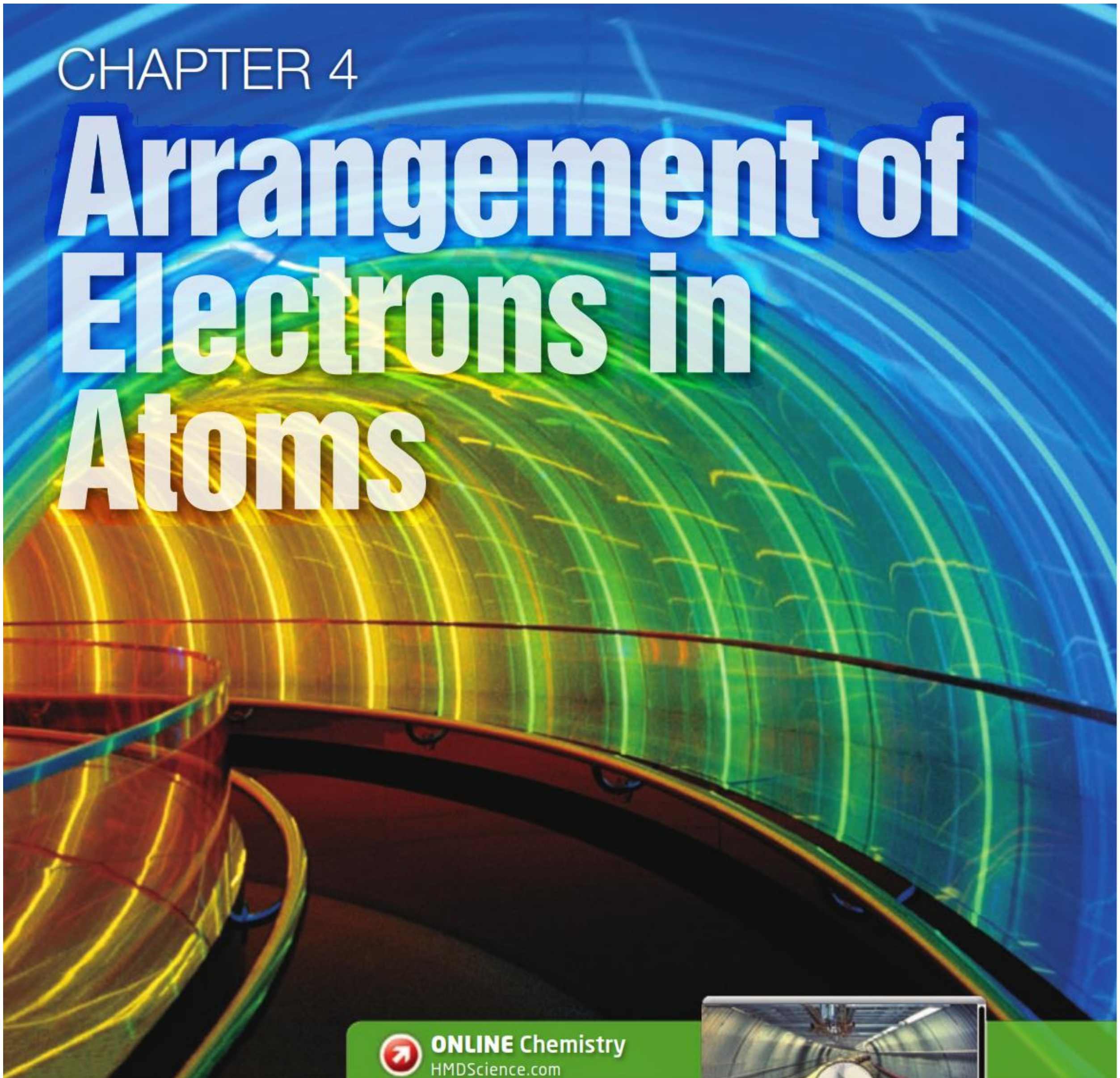
CHAPTER 4

Arrangement of Electrons in Atoms



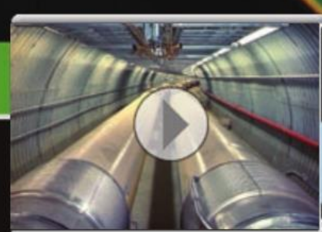
ONLINE Chemistry
HMDScience.com

SECTION 1
The Development of a New Atomic Model

SECTION 2
The Quantum Model of the Atom

SECTION 3
Electron Configurations

ONLINE LABS include:
Flame Tests



PREMIUM CONTENT

Why It Matters Video
HMDScience.com

Atoms

(b) ©Silvestre Machado/SuperStock; (br) DOE

The Development of a New Atomic Model

SECTION 1

Main Ideas

- Light has characteristics of both particles and waves.
- When certain frequencies of light strike a metal, electrons are emitted.
- Electrons exist only in very specific energy states for atoms of each element.
- Bohr's model of the hydrogen atom explained electron transition states.

Key Terms

electromagnetic radiation
electromagnetic spectrum
wavelength
frequency
photoelectric effect
quantum
photon
ground state
excited state
line-emission spectrum
continuous spectrum

The Rutherford model of the atom was an improvement over previous models, but it was incomplete. It did not explain how the atom's negatively charged electrons are distributed in the space surrounding its positively charged nucleus. After all, it was well known that oppositely charged particles attract each other. So what prevented the negative electrons from being drawn into the positive nucleus?

In the early twentieth century, a new atomic model evolved as a result of investigations into the absorption and emission of light by matter. The studies revealed a relationship between light and an atom's electrons. This new understanding led directly to a revolutionary view of the nature of energy, matter, and atomic structure.

MAIN IDEA

Light has characteristics of both particles and waves.

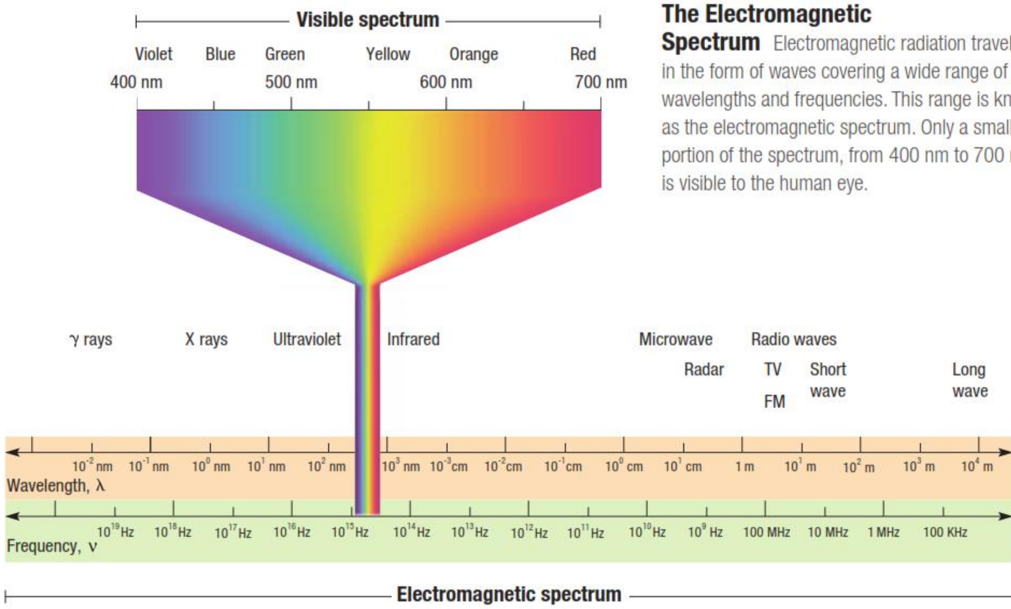
Before 1900, scientists thought light behaved solely as a wave. This belief changed when it was later discovered that light also has particle-like characteristics. Still, many of light's properties can be described in terms of waves. A quick review of these wavelike properties will help you understand the basic theory of light as it existed at the beginning of the twentieth century.

The Wave Description of Light

Visible light is a kind of electromagnetic radiation, which is a form of energy that exhibits wavelike behavior as it travels through space. Other kinds of electromagnetic radiation include X-rays, ultraviolet and infrared light, microwaves, and radio waves. **Together, all the forms of electromagnetic radiation form the electromagnetic spectrum.** The electromagnetic spectrum is represented in **Figure 1.1** on the next page. All forms of electromagnetic radiation move at a constant speed of 3.00×10^8 meters per second (m/s) through a vacuum and at slightly slower speeds through matter. Because air is mostly empty space, the value of 3.00×10^8 m/s is also light's approximate speed through air.

The significant feature of wave motion is its repetitive nature, which can be characterized by the measurable properties of wavelength and frequency. **Wavelength (λ) is the distance between corresponding points on adjacent waves.** The unit for wavelength is a distance unit.

FIGURE 1.1



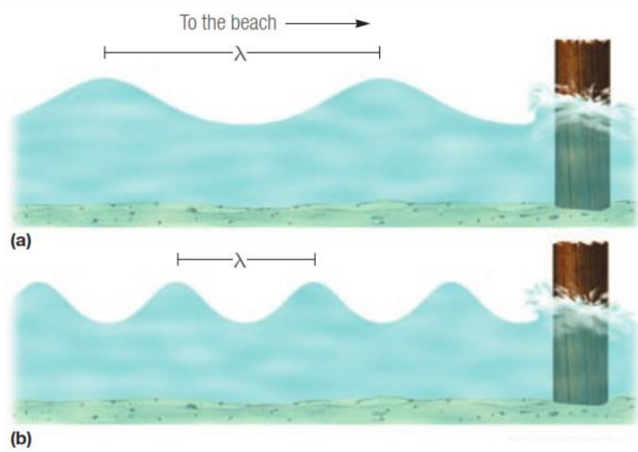
The Electromagnetic Spectrum Electromagnetic radiation travels in the form of waves covering a wide range of wavelengths and frequencies. This range is known as the electromagnetic spectrum. Only a small portion of the spectrum, from 400 nm to 700 nm, is visible to the human eye.

Depending on the type of electromagnetic radiation, it may be expressed in meters, centimeters, or nanometers, as shown in **Figure 1.1**.

Frequency (ν) is defined as the number of waves that pass a given point in a specific time, usually one second. Frequency is expressed in waves/second. One wave/second is called a hertz (Hz), named for Heinrich Hertz, who was a pioneer in the study of electromagnetic radiation. **Figure 1.2** illustrates the properties of wavelength and frequency for a familiar kind of wave, a wave on the surface of water. Frequency and wavelength are mathematically related to each other.

FIGURE 1.2

Wavelength and Frequency
The distance between two corresponding points on one of these water waves, such as from crest to crest, is the wave's wavelength, λ. We can measure the wave's frequency, ν, by observing how often the water level rises and falls at a given point, such as at the post.



CRITICAL THINKING
Analyze Which waves have a greater wavelength, those in (a) or those in (b)? Which have the greater frequency? Explain each answer.

For electromagnetic radiation, the mathematical relationship between frequency and wavelength is written as follows:

$$c = \lambda\nu$$

In the equation, c is the speed of light (in m/s), λ is the wavelength of the electromagnetic wave (in m), and ν is the frequency of the electromagnetic wave (in s^{-1}). Because c is the same for all electromagnetic radiation, the product $\lambda\nu$ is a constant. Consequently, as the wavelength of light decreases, its frequency increases, and vice versa.

MAIN IDEA

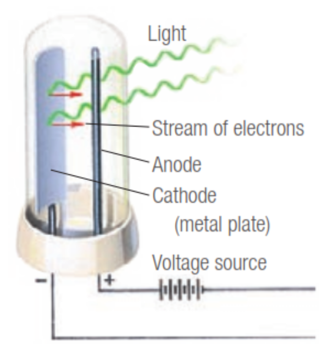
When certain frequencies of light strike a metal, electrons are emitted.

In the early 1900s, scientists conducted two experiments involving interactions of light and matter that could not be explained by the wave theory of light. One experiment involved a phenomenon known as the photoelectric effect. **The photoelectric effect refers to the emission of electrons from a metal when light shines on the metal,** as illustrated in **Figure 1.3**.

The mystery of the photoelectric effect involved the frequency of the light striking the metal. For a given metal, no electrons were emitted if the light's frequency was below a certain minimum—regardless of the light's intensity. Light was known to be a form of energy, capable of knocking loose an electron from a metal. But the wave theory of light predicted that light of any frequency could supply enough energy to eject an electron. Scientists couldn't explain why the light had to be of a minimum frequency in order for the photoelectric effect to occur.

FIGURE 1.3

The Photoelectric Effect Electromagnetic radiation strikes the surface of the metal, rejecting electrons from the metal and causing an electric current.



The Particle Description of Light

The explanation of the photoelectric effect dates back to 1900, when German physicist Max Planck was studying the emission of light by hot objects. He proposed that a hot object does not emit electromagnetic energy continuously, as would be expected if the energy emitted were in the form of waves. Instead, Planck suggested that the object emits energy in small, specific packets called quanta. **A quantum of energy is the minimum quantity of energy that can be lost or gained by an atom.** Planck proposed the following relationship between a quantum of energy and the frequency of radiation.

Quantum of Energy $E = h\nu$

In the equation, E is the energy, in joules, of a quantum of radiation, ν is the frequency, in s^{-1}, of the radiation emitted, and h is a fundamental physical constant now known as Planck's constant; $h = 6.626 \times 10^{-34}$ J • s.

In 1905, Albert Einstein expanded on Planck's theory by introducing the radical idea that electromagnetic radiation has a dual wave-particle nature. Light exhibits many wavelike properties, but it can also be thought of as a stream of particles. Each particle carries a quantum of energy.

Einstein called these particles photons. **A photon is a particle of electromagnetic radiation having zero mass and carrying a quantum of energy.** The energy of a photon depends on the frequency of the radiation.

$$E_{photon} = h\nu$$

Einstein explained the photoelectric effect by proposing that electromagnetic radiation is absorbed by matter only in whole numbers of photons. In order for an electron to be ejected from a metal surface, the electron must be struck by a single photon possessing at least the minimum energy required to knock the electron loose. According to the equation $E_{photon} = h\nu$, this minimum energy corresponds to a minimum frequency. If a photon's frequency is below the minimum, then the electron remains bound to the metal surface. Electrons in different metals are bound more or less tightly, so different metals require different minimum frequencies to exhibit the photoelectric effect.

MAIN IDEA

Electrons exist only in very specific energy states for every atom of each element.

When current is passed through a gas at low pressure, the potential energy of the gas atoms increases. **The lowest energy state of an atom is its ground state. A state in which an atom has a higher potential energy than it has in its ground state is an excited state.** There are many possible excited states, each with a unique energy, but only one ground state energy for atoms of a given element. When an excited atom returns to its ground state or a lower energy excited state, it gives off the energy it gained in the form of electromagnetic radiation. The production of colored light in neon signs, as shown in **Figure 1.4**, is a familiar example of this process.

FIGURE 1.4

Light Emission Excited neon atoms emit light when electrons in higher energy levels fall back to the ground state or to a lower-energy excited state.

FIGURE 1.5

Emission-Line Spectra Excited hydrogen atoms emit a pinkish glow, as is shown in this diagram. When the visible portion of the emitted light is passed through a prism, it is separated into specific wavelengths that are part of hydrogen's emission-line spectrum. The line at 397 nm is in the ultraviolet and is not visible to the human eye.

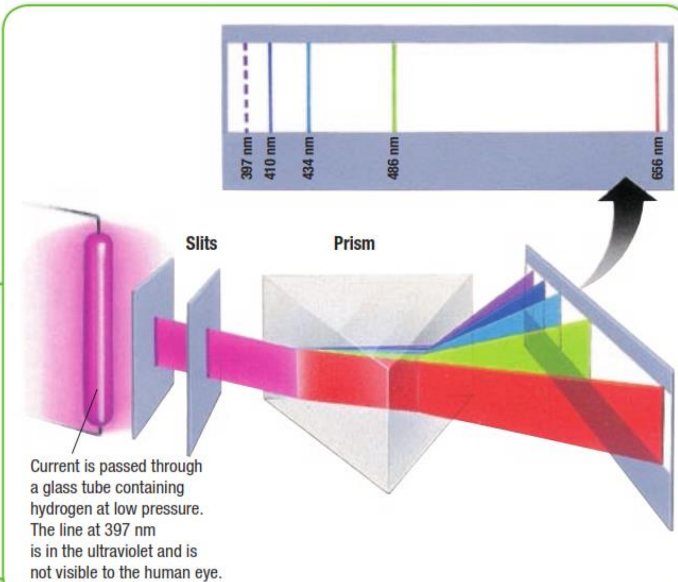
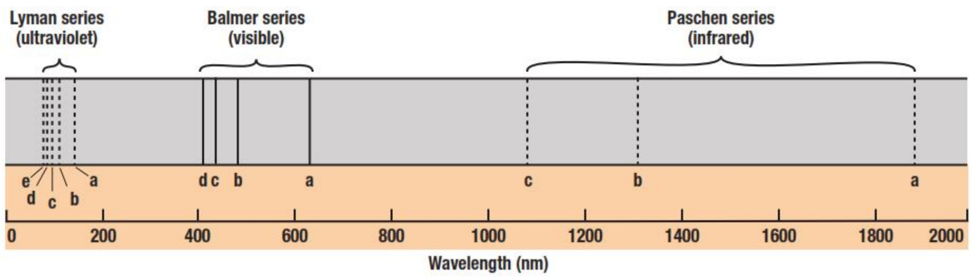


FIGURE 1.6

Explaining Energy-Levels A series of specific wavelengths of emitted light makes up hydrogen's emission-line spectrum. The letters below the lines label hydrogen's various energy-level transitions. Niels Bohr's model of the hydrogen atom provided an explanation for these transitions.



When investigators passed electric current through a vacuum tube containing hydrogen gas at low pressure, they observed the emission of a characteristic pinkish glow. **When a narrow beam of the emitted light was shined through a prism, it was separated into four specific colors of the visible spectrum. The four bands of light were part of what is known as hydrogen's emission-line spectrum.** The production of hydrogen's emission-line spectrum is illustrated in **Figure 1.5** (on the previous page). Additional series of lines were discovered in the ultraviolet and infrared regions of hydrogen's emission-line spectrum. The wavelengths of some of the spectral series are shown in **Figure 1.6**. They are known as the Lyman, Balmer, and Paschen series, after their discoverers.

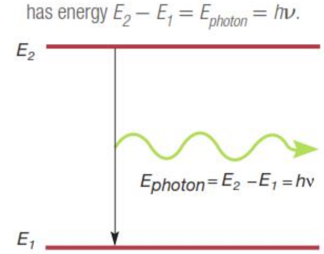
Classical theory predicted that the hydrogen atoms would be excited by whatever amount of energy was added to them. **Scientists had thus expected to observe the emission of a continuous range of frequencies of electromagnetic radiation, that is, a continuous spectrum.** Why had the hydrogen atoms given off only specific frequencies of light? Attempts to explain this observation led to an entirely new atomic theory called *quantum theory*.

Whenever an excited hydrogen atom falls to its ground state or to a lower-energy excited state, it emits a photon of radiation. The energy of this photon ($E_{photon} = h\nu$) is equal to the difference in energy between the atom's initial state and its final state, as illustrated in **Figure 1.7**. The fact that hydrogen atoms emit only specific frequencies of light indicated that the energy differences between the atoms' energy states were fixed. This suggested that the electron of a hydrogen atom exists only in very specific energy states.

In the late nineteenth century, a mathematical formula that related the various wavelengths of hydrogen's emission-line spectrum was discovered. The challenge facing scientists was to provide a model of the hydrogen atom that accounted for this relationship.

FIGURE 1.7

Mathematical Model of Energy States When an excited atom with energy E_2 falls back to energy E_1, it releases a photon that has energy $E_2 - E_1 = E_{photon} = h\nu$.



MAIN IDEA

Bohr's model of the hydrogen atom explained electron transition states.

The puzzle of the hydrogen-atom spectrum was solved in 1913 by the Danish physicist Niels Bohr. He proposed a hydrogen-atom model that linked the atom's electron to photon emission. According to the model, the electron can circle the nucleus only in allowed paths, or *orbits*. When the electron is in one of these orbits, the atom has a definite, fixed energy. The electron—and therefore the hydrogen atom—is in its lowest energy state when it is in the orbit closest to the nucleus. This orbit is separated from the nucleus by a large empty space where the electron cannot exist. The energy of the electron is higher when the electron is in orbits that are successively farther from the nucleus.

The electron orbits, or atomic energy levels, in Bohr's model can be compared to the rungs of a ladder. When you are standing on a ladder, your feet are on one rung or another. The amount of potential energy that you possess corresponds to standing on the first rung, the second rung, and so forth. Your energy cannot correspond to standing between two rungs because you cannot stand in midair. In the same way, an electron can be in one orbit or another, but not in between.

How does Bohr's model of the hydrogen atom explain the observed spectral lines? While in a given orbit, the electron is neither gaining nor losing energy. It can, however, move to a higher-energy orbit by gaining an amount of energy equal to the difference in energy between the higher-energy orbit and the initial lower-energy orbit. When a hydrogen atom is in an excited state, its electron is in one of the higher-energy orbits. When the electron falls to a lower energy level, a photon is emitted, and the process is called *emission*. The photon's energy is equal to the energy difference between the initial higher energy level and the final lower energy level. Energy must be added to an atom in order to move an electron from a lower energy level to a higher energy level. This process is called *absorption*. Absorption and emission of radiation in Bohr's model of the hydrogen atom are illustrated in **Figure 1.8**. The energy of each absorbed or emitted photon corresponds to a particular frequency of emitted radiation, $E_{photon} = h\nu$.

Based on the different wavelengths of the hydrogen emission-line spectrum, Bohr calculated the allowed energy levels for the hydrogen atom. He then related the possible energy-level changes to the lines in the hydrogen emission-line spectrum. The five lines in the Lyman series, for example, were shown to be the result of electrons dropping from energy levels E_6, E_5, E_4, E_3, and E_2 to the ground-state energy level E_1.

WHY IT MATTERS

Fireflies S.T.E.M.

What kinds of reactions produce light? In this chapter, you are learning how excited atoms can produce light. In parts of the United States, summer is accompanied by the appearance of fireflies, or lightning bugs. What makes them glow? A bioluminescent chemical reaction that involves luciferin, luciferase (an enzyme), adenosine triphosphate (ATP), and oxygen takes place in the firefly and produces the characteristic yellow-green glow. Unlike most reactions that produce light, bioluminescent reactions do not generate energy in the form of heat.

FIGURE 1.8

Absorption and Emission **(a)** Absorption and **(b)** emission of a photon by a hydrogen atom according to Bohr's model. The frequencies of light that can be absorbed and emitted are restricted because the electron can only be in orbits corresponding to the energies E_1, E_2, E_3, and so forth.

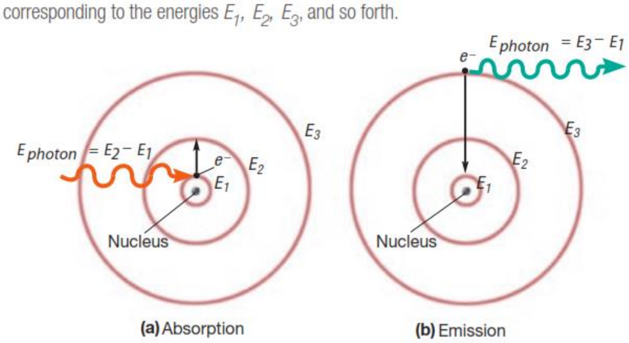
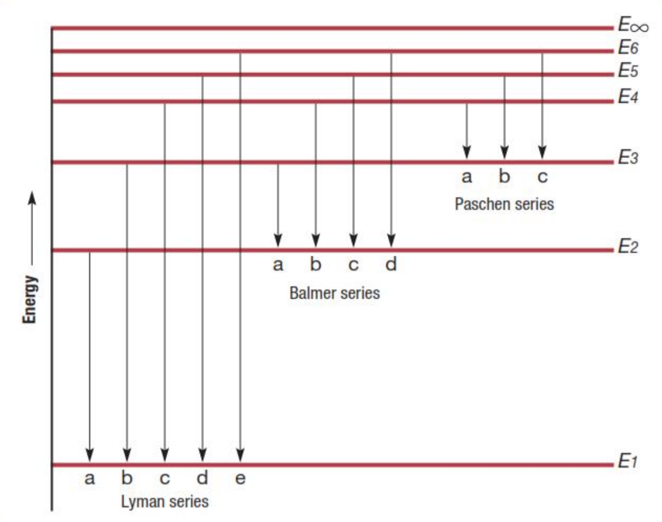


FIGURE 1.9

Electron Energy Transitions This energy-state diagram for a hydrogen atom shows some of the energy transitions for the Lyman, Balmer, and Paschen spectral series. Bohr's model of the atom accounted mathematically for the energy of each of the transitions shown.

CRITICAL THINKING

Explain Why might spectra of atoms with more than one electron be difficult to explain using Bohr's model?



Bohr's calculated values agreed with the experimentally observed values for the lines in each series. The origins of three of the series of lines in hydrogen's emission-line spectrum are shown in **Figure 1.9.**

Bohr's model of the hydrogen atom explained observed spectral lines so well that many scientists concluded that the model could be applied to all atoms. It was soon recognized, however, that Bohr's approach did not explain the spectra of atoms with more than one electron. Nor did Bohr's theory explain the chemical behavior of atoms.

SECTION 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. What was the major shortcoming of Rutherford's model of the atom?

2. Write and label the equation that relates the speed, wavelength, and frequency of electromagnetic radiation.

3. Define the following:
- **a.** electromagnetic radiation
- **b.** wavelength
- **c.** frequency
- **d.** quantum
- **e.** photon

4. What is meant by the dual wave-particle nature of light?

5. Describe Bohr's model of the hydrogen atom.

Critical Thinking

6. INTERPRETING GRAPHICS Use the diagram in **Figure 1.9** to answer the following:
- **a.** Characterize each of the following as absorption or emission: an electron moves from E_2 to E_1; an electron moves from E_1 to E_3; and an electron moves from E_6 to E_3.
- **b.** Which energy-level change above emits or absorbs the highest energy? the lowest energy?