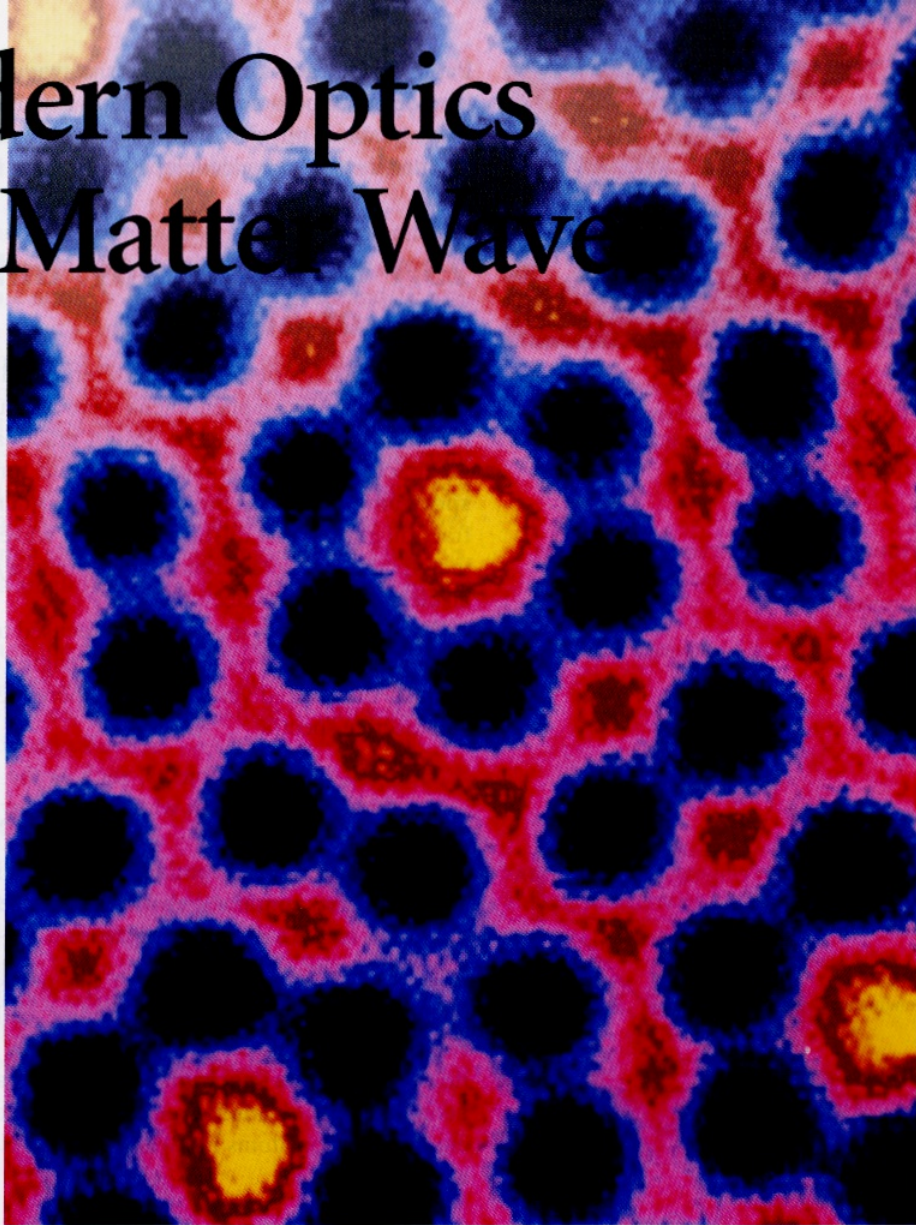


Modern Optics and Matter Waves

This image from a scanning tunneling microscope shows individual silicon atoms at the surface of a silicon crystal.



▶ Looking Ahead

The goal of Chapter 25 is to explore the limits of the wave and particle models. In this chapter you will learn to:

- Understand how light and x rays are used to study atoms and solids.
- Use the photon model of light.
- Recognize the experimental evidence for the wave nature of matter.
- Understand that energy quantization is a consequence of the wave-like properties of matter.

◀ Looking Back

The material in this chapter depends on the wave model of light. Please review:

- Section 22.1 Models of light.
- Sections 22.2 and 22.3 Double-slit interference and diffraction gratings.

The scanning tunneling microscope is one of the most important inventions of the late 20th century. For the first time, we can “see” the structure of materials at the atomic level. The scanning tunneling microscope works by exploiting the wave properties of electrons.

Wave properties? Aren't electrons particles?

Perhaps not. Our journey through physics has brought us to about 1890, a little over a century ago. The physics of particles and waves was well understood by then, and it seemed that Newtonian physics would soon succeed in explaining all the phenomena of nature in terms of the particle and wave models. But trouble was on the horizon. Discoveries made during the last decade of the 19th century and the opening years of the 20th century refused to yield to a Newtonian analysis.

At the heart of the crisis was a breakdown of the basic particle and wave models. As physicists probed more deeply into the nature of light, they began to make observations that couldn't be reconciled with the wave model. Sometimes, as you will see, light refuses to act like a wave and seems more like a collection of particles. Even more troubling experiments found that electrons sometimes behave like waves.

These discoveries eventually led to a radical new theory of light and matter called *quantum physics*. We will return to study quantum physics more thoroughly in



Some modern spectrometers are small enough to hold in your hand. (The rainbow has been superimposed to show how it works.)

FIGURE 25.1 A diffraction spectrometer for the accurate measurement of wavelengths.

Part VII. Our goal in this chapter, as we conclude our study of waves, is to use our knowledge of particles and waves to examine some of the experimental evidence that led to quantum physics. By doing so, we will discover the limits of the wave and particle models that we've developed.

25.1 Spectroscopy: Unlocking the Structure of Atoms

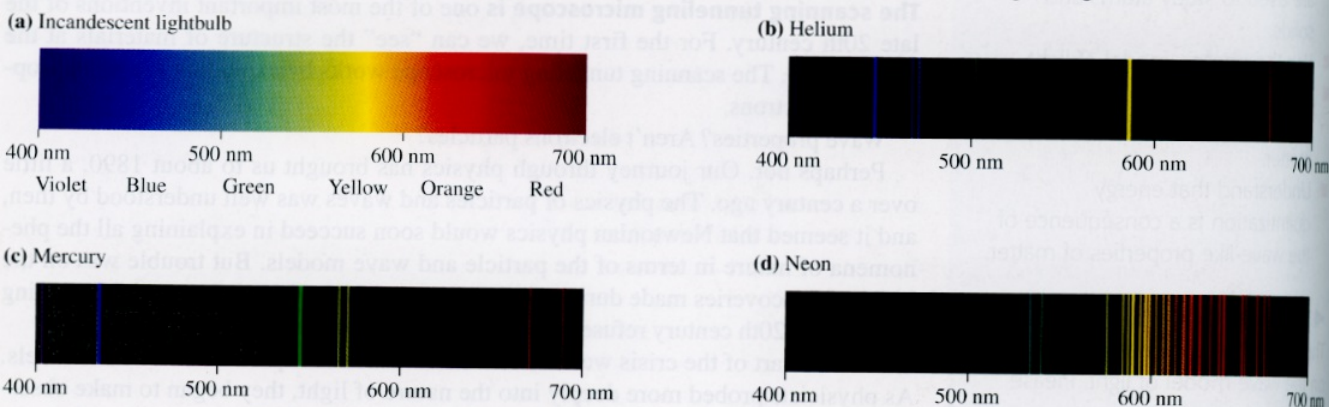
The basic discoveries of the interference and diffraction of light were made early in the 19th century. These phenomena were well understood by the end of the century, and the knowledge was used to design practical tools for measuring wavelengths with great accuracy. The primary instrument for measuring the wavelengths of light is a **spectrometer**, such as the one shown in **FIGURE 25.1**. The heart of a spectrometer is a diffraction grating that diffracts different wavelengths of light at different angles. A lens then focuses the interference fringes onto a *photographic plate* or (more likely today) an electronic array detector.

Each wavelength in the light is focused to a different position on the detector, producing a distinctive pattern of wavelengths called the **spectrum** of the light. Spectroscopists discovered very early that there are two types of spectra, continuous spectra and discrete spectra:

- Hot, self-luminous objects, such as the sun or an incandescent lightbulb, emit a *continuous spectrum* in which a rainbow is formed by light being emitted at every possible wavelength.
- In contrast, the light emitted by a gas discharge tube (such as those used to make neon signs) contains only certain discrete, individual wavelengths. Such a spectrum is called a **discrete spectrum**.

FIGURE 25.2 shows examples of spectra as they would appear on the photographic plate of a spectrometer. Each bright line, called a **spectral line**, represents *one* specific wavelength present in the light emitted by the source. A discrete spectrum is sometimes called a **line spectrum** because of its appearance on the plate. You can see that a neon light has its familiar reddish-orange color because nearly all of the wavelengths emitted by neon atoms fall within the wavelength range 600–700 nm that we perceive as orange and red.

FIGURE 25.2 Examples of spectra in the visible wavelength range 400–700 nm.



Two important conclusions had been established by the end of the 19th century:

1. The light emitted by atoms in a gas discharge tube has a discrete spectrum.
2. Every element in the periodic table has its own unique spectrum.

The fact that each element emits a unique spectrum means that atomic spectra can be used as “fingerprints” to identify elements. Consequently, atomic spectroscopy is the basis of many contemporary technologies for analyzing the composition of unknown materials, monitoring air pollutants, and studying the atmospheres of the earth and other planets.

You know from chemistry that an element’s *atomic number* specifies the number of protons and electrons within an atom. Hydrogen, with atomic number 1, has one electron and one proton, while neon, at atomic number 10, has 10 electrons and 10 protons. It was soon recognized that an atom’s internal structure determines the wavelengths the atom emits. If we only knew how to “decode” an element’s spectrum, we would be able to determine the trajectories of the electrons within the atom.

Despite heroic attempts by some of the best scientists of the late 19th century, Newtonian mechanics and the (then) new theory of electromagnetism were completely unable to provide an explanation of atomic spectra or atomic structure. Not only did they fail to predict why one element’s spectrum should differ from another, these classical theories predicted that atomic electrons should spiral into the nucleus, destroying the atoms and the universe in a small fraction of a second! This prediction is obviously incorrect.

Physics from the time of Newton through the mid-19th century had been spectacularly successful. But the physics of particles and waves was unable to explain the puzzle of discrete spectra. The first hint of a new direction in which to turn was made in 1885 by a Swiss school teacher named Johann Balmer.

Balmer and the Hydrogen Atom

Balmer was intrigued by the spectrum of hydrogen. Hydrogen is the simplest atom, with one electron orbiting a proton, and it also has the simplest atomic spectrum. The *visible spectrum* of hydrogen, between 400 nm and 700 nm, consists of only four spectral lines. The wavelengths are given in Table 25.1. Physicists felt certain that such a simple spectrum must have a simple and straightforward explanation.

In 1885, Johann Balmer found by trial and error that the four wavelengths in the visible spectrum of hydrogen could be represented by the simple formula

$$\lambda = \frac{91.18 \text{ nm}}{\left(\frac{1}{2^2} - \frac{1}{n^2}\right)} \quad n = 3, 4, 5, 6 \quad (25.1)$$

Balmer’s formula calculated the measured wavelengths with much better than 0.1% accuracy. Not only was his formula accurate, it was *simple*, in keeping with expectations that the hydrogen spectrum should have a simple explanation.

Balmer knew only the four *visible* wavelengths shown in Table 25.1, but an obvious question to ask was whether Equation 25.1 also predicts wavelengths for $n = 7$, 8, 9, and so on. The prediction for $n = 7$ is $\lambda = 397.1 \text{ nm}$, an ultraviolet wavelength. Spectroscopists were just beginning to extend their craft into the ultraviolet and infrared regions of the spectrum, and it was soon confirmed that Balmer’s formula does, indeed, work for *all* values of n .

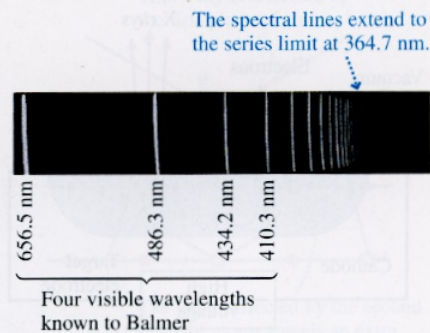
Balmer’s formula predicts a *series* of spectral lines of gradually decreasing wavelength, converging to the *series limit* wavelength of 364.7 nm as $n \rightarrow \infty$. Although there are an infinite number of spectral lines in this series, their intensities rapidly get weaker as n increases until, for large values of n , they blur together and cannot be resolved. This series of spectral lines is now called the **Balmer series**. FIGURE 25.3 shows a photograph of the Balmer series of hydrogen in which the series limit is quite obvious.

TABLE 25.1 Wavelengths of visible lines in the hydrogen spectrum*

656.46 nm
486.27 nm
434.17 nm
410.29 nm

*Wavelengths in vacuum.

FIGURE 25.3 The Balmer series of hydrogen as seen on the photographic plate of a spectrometer.



With the success of Balmer's formula, it was natural to ask what happens if the 2^2 in Equation 25.1 is changed to 1^2 or 3^2 or m^2 . It was easy to calculate that all spectral lines in the series with 1^2 , if they existed, would have fairly extreme ultraviolet wavelengths, while all those in the series with 3^2 would be in the infrared. Spectroscopists accepted the challenge and went to work developing the techniques for infrared and ultraviolet spectroscopy.

The $m = 1$ series was discovered by Theodore Lyman and is called the Lyman series. The $m = 3$ series, found by Louis Paschen, is called the Paschen series. They confirmed, beyond doubt, that Balmer's formula could be generalized to

$$\lambda = \frac{91.18 \text{ nm}}{\left(\frac{1}{m^2} - \frac{1}{n^2}\right)} \quad \begin{cases} m = 1 & \text{Lyman series} \\ m = 2 & \text{Balmer series} \\ m = 3 & \text{Paschen series} \\ \vdots & \end{cases} \quad (25.2)$$

$$n = m + 1, m + 2, \dots$$

As spectroscopists acquired ever more data, it became increasingly clear that Equation 25.2 could predict *every* line in the hydrogen spectrum, from the extreme ultraviolet to the far infrared.

Surely Balmer's success was not a mere coincidence. There must be some underlying meaning to his formula. But what? Balmer did not present a *theory*. He simply said, "Here's a formula that accurately calculates the wavelengths in the hydrogen spectrum." In effect, Balmer's formula was a challenge. Any successful theory of atoms must be able to *derive* Equation 25.2 from the basic laws and principles of the theory. It was 30 years before a theory was proposed that could meet this challenge.

It is particularly striking that Equation 25.2 depends on two *integers*. Hydrogen atoms simply do not emit wavelengths for $m = 1.6$ or for $n = 3.4$. This must tell us *something* important about the structure of the hydrogen atom. Newtonian mechanics does not deal in such "discrete" quantities. Masses, forces, velocities, and energies can take on any value at all; they are not restricted to having only some values but not others.

However, we have seen one exception to this: standing waves. Standing waves exist for only certain frequencies and wavelengths that are described by an *integer* called the mode number. Could there, somehow, be a connection between standing waves and the structure of atoms? To answer this question, we must probe yet deeper into the nature of light and matter.

25.2 X-Ray Diffraction

In 1895, the German physicist Wilhelm Röntgen made a remarkable discovery. The late 19th century was a period in which the technology of vacuum tubes was being perfected, and Röntgen was studying how electrons (called *cathode rays* at the time) traveled through a vacuum. He sealed an electron-producing cathode and a metal target electrode into a vacuum tube, such as shown in **FIGURE 25.4**. A high voltage pulled electrons from the cathode and accelerated them to very high speeds before they struck the target. Röntgen and others had done similar experiments previously, but one day he happened by chance to have left a sealed envelope containing film near the vacuum tube. He was later surprised to discover that the film had been exposed even though it had never been removed from the envelope. This serendipitous discovery was the beginning of the study of x rays.

FIGURE 25.4 Röntgen's x-ray tube.

