**6.3 | LINE SPECTRA AND THE BOHR MODEL**

The work of Planck and Einstein paved the way for understanding how electrons are

arranged in atoms. In 1913, the Danish physicist Niels Bohr (\_ **FIGURE 6.8**) offered a

theoretical explanation of *line spectra,* another phenomenon that had puzzled scientists

during the nineteenth century.

**Line Spectra**

A particular source of radiant energy may emit a single wavelength, as in the light from

a laser. Radiation composed of a single wavelength is *monochromatic*. However, most

common radiation sources, including lightbulbs and stars, produce radiation containing

many different wavelengths and is *polychromatic*. A **spectrum** is produced when

radiation from such sources is separated into its component wavelengths, as shown in

\_ **FIGURE 6.9**. The resulting spectrum consists of a continuous range of colors—violet

merges into indigo, indigo into blue, and so forth, with no blank spots. This rainbow

of colors, containing light of all wavelengths, is called a

**continuous spectrum**. The most familiar

example of a continuous spectrum is the

rainbow produced when raindrops or

mist acts as a prism for sunlight.

Not all radiation sources produce a

continuous spectrum. When a high voltage

is applied to tubes that contain different

gases under reduced pressure, the gases emit different colors of light (\_ **FIGURE 6.10**).





The light emitted by neon gas is the familiar red-orange glow of many “neon” lights,

whereas sodium vapor emits the yellow light characteristic of some modern streetlights.

When light coming from such tubes is passed through a prism, only a few wavelengths

are present in the resultant spectra (\_ **FIGURE 6.11**). Each colored line in such spectra

represents light of one wavelength. A spectrum containing radiation of only specific

wavelengths is called a **line spectrum**.

When scientists first detected the line spectrum of hydrogen in the mid-1800s, they

were fascinated by its simplicity. At that time, only four lines at wavelengths of 410 nm

(violet), 434 nm (blue), 486 nm (blue-green), and 656 nm (red) were observed

(Figure 6.11). In 1885, a Swiss schoolteacher named Johann Balmer showed that

the wavelengths of these four lines fit an intriguingly simple formula that relates

the wavelengths to integers. Later, additional lines were found in the ultraviolet

and infrared regions of hydrogen’s line spectrum. Soon Balmer’s equation was

extended to a more general one, called the *Rydberg equation,* which allows us to

calculate the wavelengths of all the spectral lines of hydrogen:

[6.4]

\_ **FIGURE 6.11 Line spectra of hydrogen and neon.**

\_ **FIGURE 6.10 Atomic emission of**

**hydrogen and neon.** Different gases emit

light of different characteristic colors when

an electric current is passed