When the constant in Eq. 27–16 is evaluated with $Z = 1$, it is found to have the measured value of the Rydberg constant, $R = 1.0974 \times 10^7\,\text{m}^{-1}$ in Eq. 27–9, in accord with experiment (see Problem 53).

The great success of Bohr's model is that it gives an explanation for why atoms emit line spectra, and accurately predicts the wavelengths of emitted light for hydrogen. The Bohr model also explains absorption spectra: photons of just the right wavelength can knock an electron from one energy level to a higher one. To conserve energy, only photons that have just the right energy will be absorbed. This explains why a continuous spectrum of light entering a gas will emerge with dark (absorption) lines at frequencies that correspond to emission lines (Fig. 27–21c).

Absorption lines explained

The Bohr theory also ensures the stability of atoms. It establishes stability by decree: the ground state is the lowest state for an electron and there is no lower energy level to which it can go and emit more energy. Finally, as we saw above, the Bohr theory accurately predicts the ionization energy of 13.6 eV for hydrogen. However, the Bohr model was not so successful for other atoms, and has been superseded as we shall discuss in the next Chapter. We discuss the Bohr model because it *was* an important start and because we still use the concept of stationary states, the ground state, and transitions between states. Also, the terminology used in the Bohr model is still used by chemists and spectroscopists.

EXAMPLE 27–12 **Wavelength of a Lyman line.** Use Fig. 27–27 to determine the wavelength of the first Lyman line, the transition from $n = 2$ to $n = 1$. In what region of the electromagnetic spectrum does this lie?

APPROACH We use Eq. 27–10, $hf = E_u - E_l$, with the energies obtained from Fig. 27–27 to find the energy and the wavelength of the transition. The region of the electromagnetic spectrum is found using the EM spectrum in Fig. 22–8.

SOLUTION In this case, $hf = E_2 - E_1 = \{-3.4\,\text{eV} - (-13.6\,\text{eV})\} = 10.2\,\text{eV} = (10.2\,\text{eV})(1.60 \times 10^{-19}\,\text{J/eV}) = 1.63 \times 10^{-18}\,\text{J}$. Since $\lambda = c/f$, we have

$$\lambda = \frac{c}{f} = \frac{hc}{E_2 - E_1} = \frac{(6.63 \times 10^{-34}\,\text{J}\cdot\text{s})(3.00 \times 10^8\,\text{m/s})}{1.63 \times 10^{-18}\,\text{J}} = 1.22 \times 10^{-7}\,\text{m},$$

or 122 nm, which is in the UV region of the EM spectrum, Fig. 22–8. See also Fig. 27–23.

NOTE An alternate approach would be to use Eq. 27–16 to find λ, and it gives the same result.

EXAMPLE 27–13 **Wavelength of a Balmer line.** Determine the wavelength of light emitted when a hydrogen atom makes a transition from the $n = 6$ to the $n = 2$ energy level according to the Bohr model.

APPROACH We can use Eq. 27–16 or its equivalent, Eq. 27–9, with $R = 1.097 \times 10^7\,\text{m}^{-1}$.

SOLUTION We find

$$\frac{1}{\lambda} = (1.097 \times 10^7\,\text{m}^{-1})\left(\frac{1}{4} - \frac{1}{36}\right) = 2.44 \times 10^6\,\text{m}^{-1}.$$

So $\lambda = 1/(2.44 \times 10^6\,\text{m}^{-1}) = 4.10 \times 10^{-7}\,\text{m}$ or 410 nm. This is the fourth line in the Balmer series, Fig. 27–22, and is violet in color.