The wavelength of light emitted is 654 nm.

$\lambda = \dfrac{c}{f}$

$c = 3.00 \times 10^{8}$ m/s

$f = \dfrac{c}{\lambda}$

$$= \dfrac{\dfrac{3.00 \times 108\ \cancel{m}}{1\ s}}{6.54 \times 10^{-7}\ \cancel{m}}$$

$$f = \dfrac{4.59 \times 10^{14}}{1\ s} = 4.59 \times 10^{14}\ \text{Hz}$$

The frequency of the light emitted is 4.59×10^{14} Hz.

$E = hf$

$h = 6.63 \times 10^{-34}$ J/Hz

$$E = \dfrac{6.63 \times 10^{-34}\ \text{J}}{1\ \cancel{\text{Hz}}} \times 4.59 \times 10^{14}\ \cancel{\text{Hz}}$$

$E = 3.0 \times 10^{-19}$ J

The electron's energy difference between the second and third levels is 3.0×10^{-19} J.

(e)

---- $n = 4$

---- $n = 3$ $\Delta E_{4 \to 2} = 4.1 \times 10^{-19}$ J

---- $n = 2$ $\Delta E_{3 \to 2} = 3.0 \times 10^{-19}$ J

1 e– $n = 1$

1 p^{+}

H

hydrogen atom

$(4.1 \times 10^{-19}\ \text{J}) - (3.0 \times 10^{-19}\ \text{J}) = 1.1 \times 10^{-19}$ J

The energy difference between hydrogen atom electron energy levels $n = 4$ and $n = 3$ will be 1.1×10^{-19} J.

14. The most likely assumption would seem to be that theories would be advanced that would try to describe electron arrangements and energies for atoms more complex than hydrogen.

3.5 QUANTUM NUMBERS

PRACTICE

(Page 182)

Understanding Concepts

1. Bohr and Sommerfeld both used observations of line spectra.
2. Bohr proposed circular electron orbits for hydrogen, while Sommerfeld proposed several elliptical orbits.

3.

Table 3 Sommerfeld's Electron Energy Sublevels

Primary energy level	Principal quantum number, *n*	Possible secondary quantum numbers, *l*	Number of sublevels per primary level
1	1	0	1
2	2	0,1	4
3	3	0,1,2	9
4	4	0,1,2,3	16

4. For any principal quantum number, n, the highest possible value of l is n–1.
5. For any principal quantum number, n, the possible values of l include all of the integers from 0 to n–1.

SECTION 3.5 QUESTIONS

(Page 184)

Understanding Concepts

1. The main kind of evidence used comes from atomic line spectra, particularly the splitting of lines.
2. The first quantum number describes the main energy level; the second quantum number describes small energy level steps within the main energy level corresponding to different shapes of "orbits"; the third quantum number describes the orientation in space of the electron "orbits"; and the fourth quantum number describes the "spin" of electrons.
3. (a) For l = 0, 1, 2, and 3, there are 0, 3, 5, and 7 possible values of m_l, respectively.
   (b) Each number is the next greater odd integer (or 2l + 1 for all ls except l = 0).
   (c) From the answer to (b), the number of possible values for m_l for l = 4 must be 9 (the next odd integer).
4. The fourth quantum number is m_s, and it is necessary to explain magnetic properties of atoms.
5.

Table 4 Summary of Quantum Numbers

(n)	(0 to n – 1)	(–l to +l)	(+1/2 or –1/2)
4	0	0	+1/2, –1/2
	1	-1, 0, +1	+1/2, –1/2
	2	-2, -1, 0, +1, +2	+1/2, –1/2
	3	-3, -2, -1, 0, +1, +2, +3	+1/2, –1/2

6. It takes four quantum numbers to describe fully an electron in an atom. An example listing labels and values of each quantum number might be n = 2, l = 1, m_l = –1, and m_s = +1/2 . This might describe an electron in a hydrogen atom in an "excited" state.
7. For each principal quantum number from n = 1 to n = 3 (see Table 4), there can be 2, 8, and 18 different electron descriptions.
8. In the development of scientific knowledge, empirical knowledge usually comes first. Examples from this section are the investigation of bright line spectra and of magnetic effects upon these spectra—both of which preceded the theory that attempts to explain them in terms of atomic structure.

3.6 ATOMIC STRUCTURE AND THE PERIODIC TABLE

PRACTICE

(Page 191)

Understanding Concepts

1. The aufbau principle states that electrons occupy lowest energy orbitals first. The Pauli exclusion principle states that no more than two electrons (of opposite spin) may occupy the same orbital, and Hund's rule states that electrons are not paired within sublevel orbitals until each sublevel orbital has at least one electron.