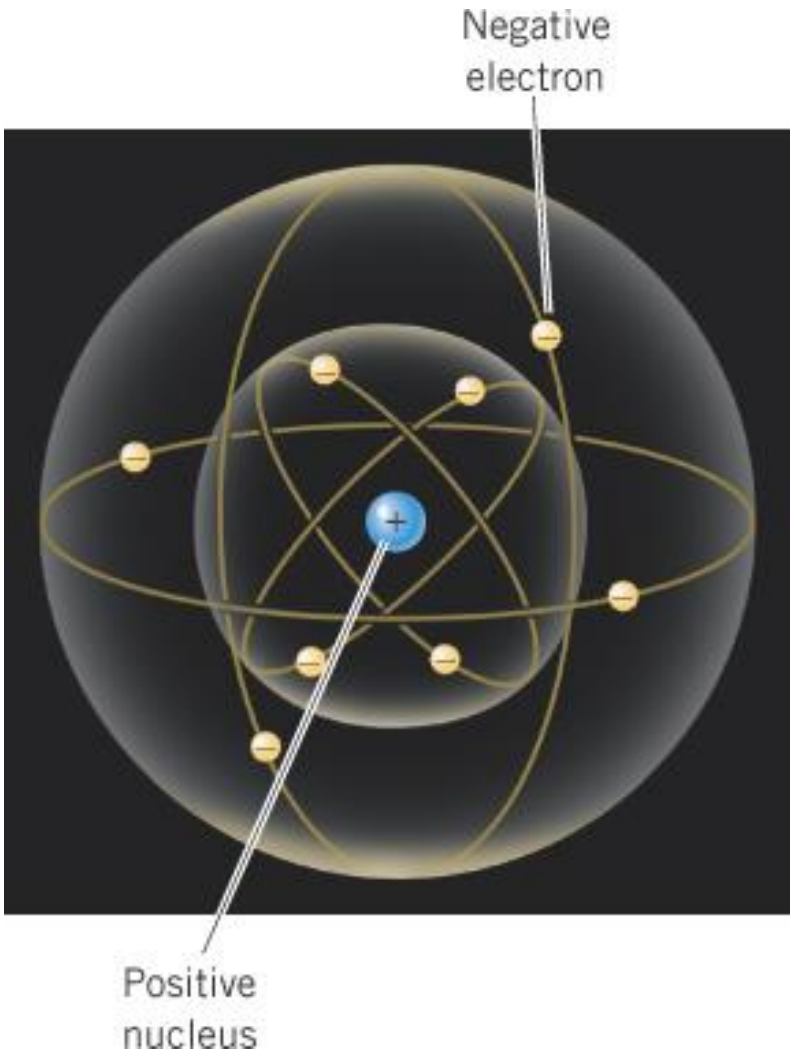
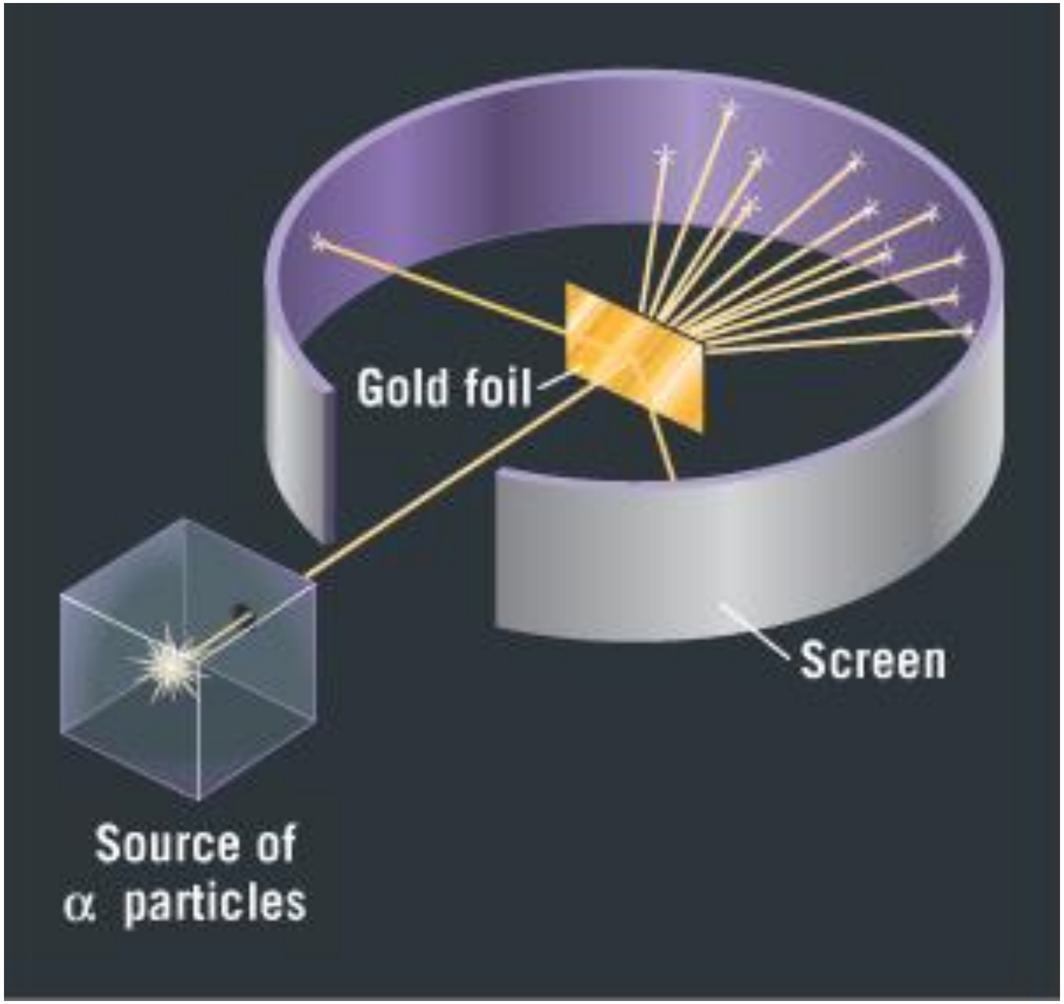
Chapter 30

The Nature of the Atom

30.1 Rutherford Scattering and the Nuclear Atom



In its natural state, an atom is electrically neutral.

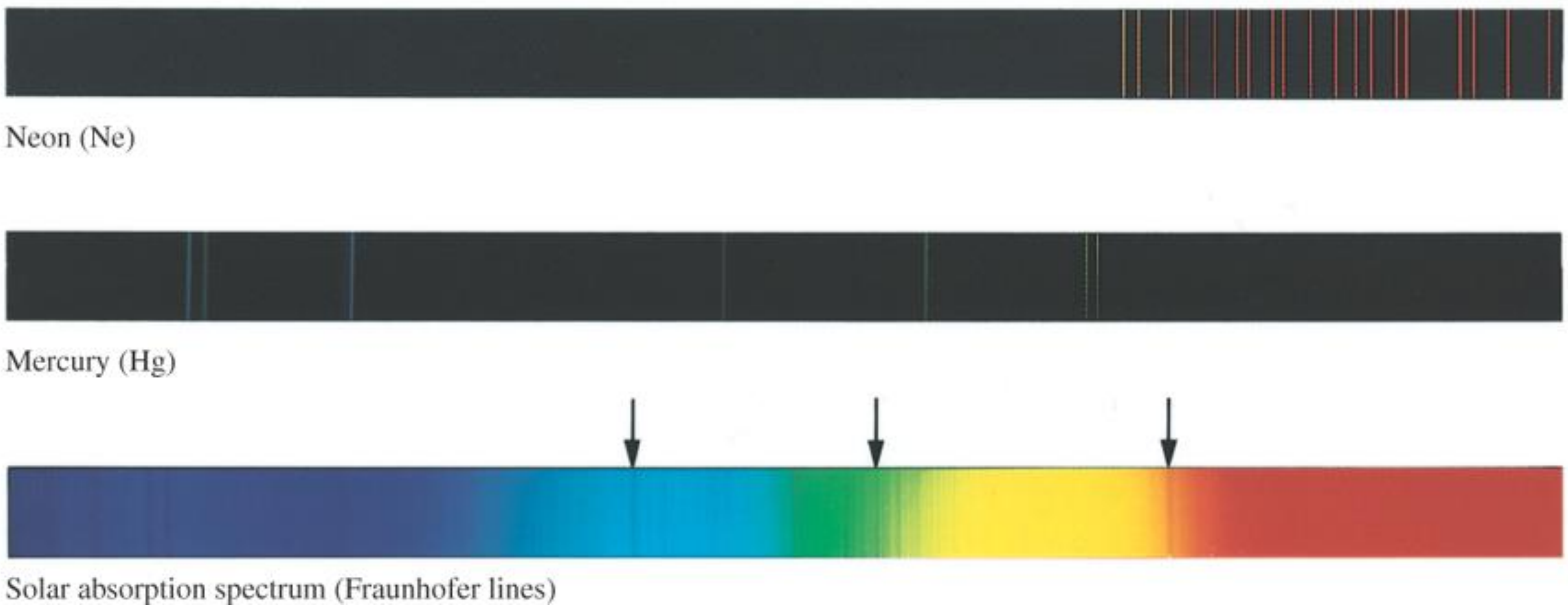


A Rutherford scattering experiment.

Conceptual Example 1 **Atoms are Mostly Empty Space**

In the planetary model of the atom, the nucleus (radius = 10^{-15}m) is analogous to the sun (radius = $7x10^{8}$m). Electrons orbit (radius = 10^{-10}m) the nucleus like the earth orbits (radius = $1.5x10^{11}$m) the sun. If the dimensions of the solar system had the same proportions as those of the atom, would the earth be closer to or farther away from the sun than it actually is?

The individual wavelengths emitted by two gases and the continuous spectrum of the sun.

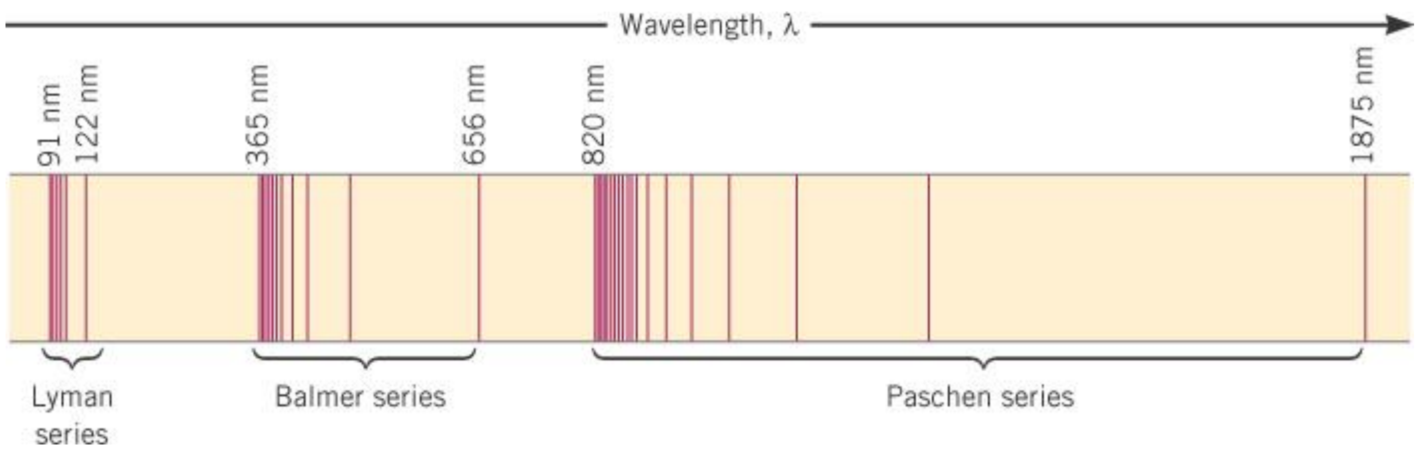


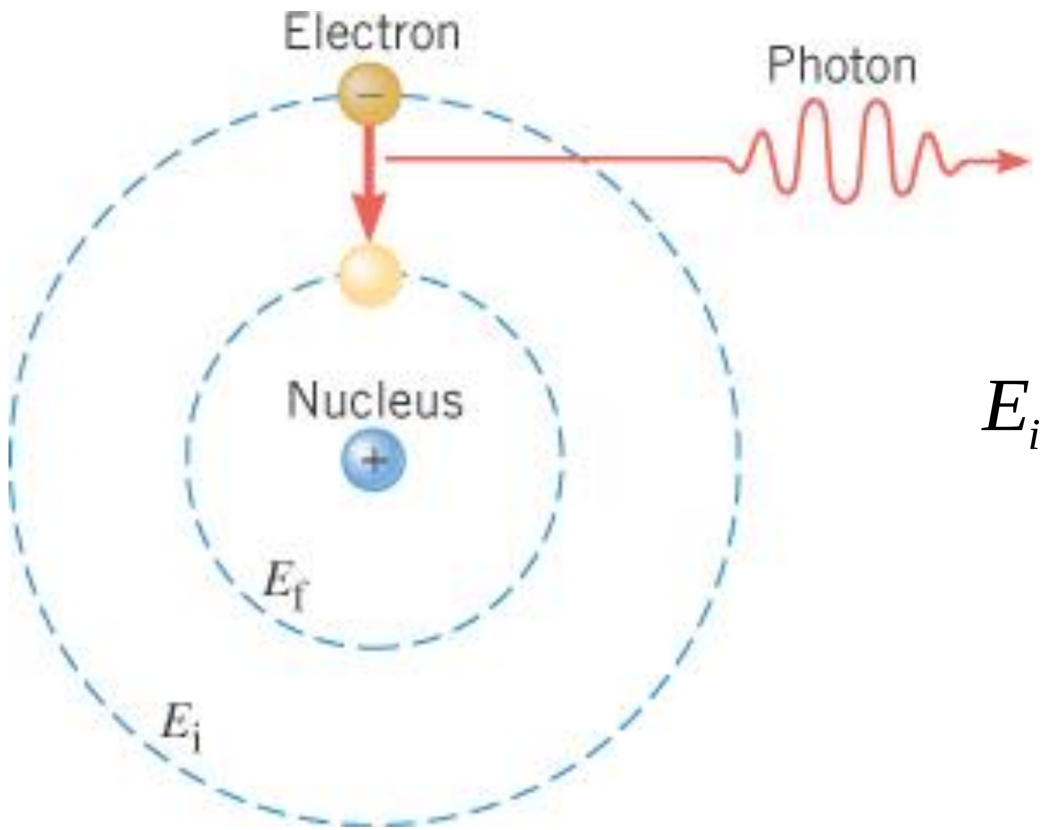
The Line Spectrum of Hydrogen

Lyman series $$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right) \qquad n = 2, 3, 4, \ldots$$

Balmer series $$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5, \ldots$$

Paschen series $$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \qquad n = 4, 5, 6, \ldots$$

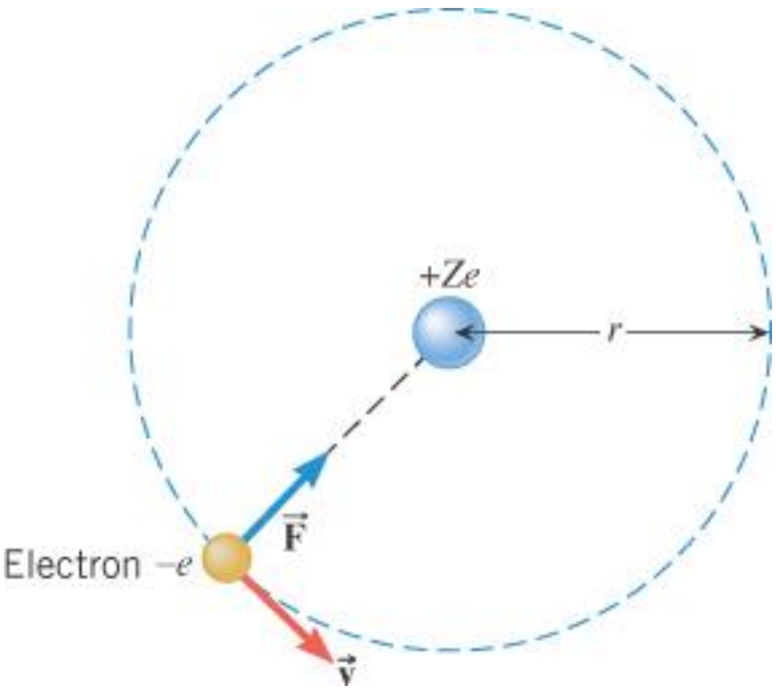




$$E_i - E_f = hf$$

In the Bohr model, a photon is emitted when the electron drops from a larger, higher-energy orbit to a smaller, lower energy orbit.

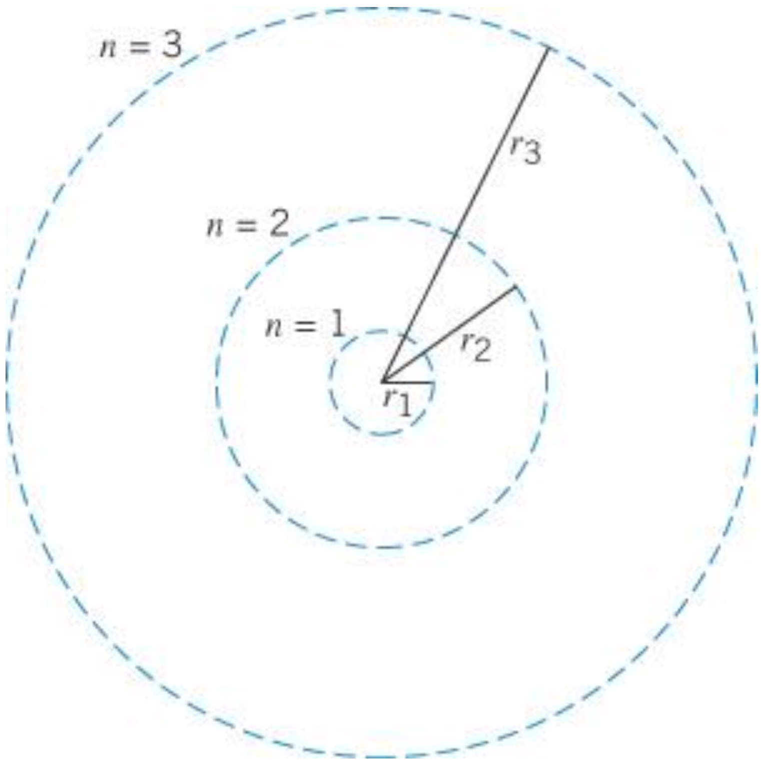
THE ENERGIES AND RADII OF THE BOHR ORBITS



$$E = \text{KE} + \text{EPE}$$

$$= \tfrac{1}{2}mv^2 - \frac{kZe^2}{r}$$

$$= -\frac{kZe^2}{2r}$$



Angular momentum is quantized.

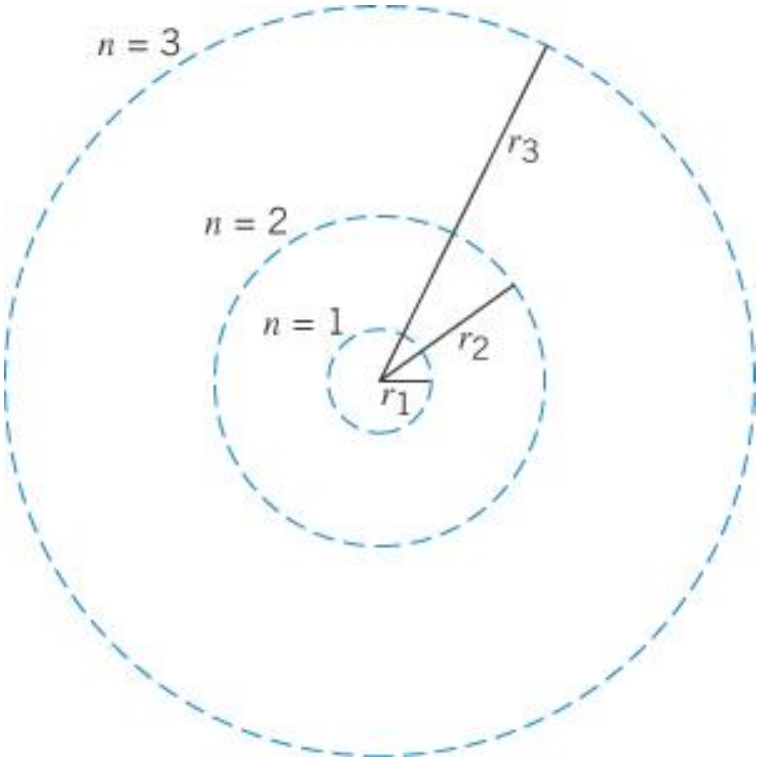
$$L_n = mv_n r_n = n\frac{h}{2\pi} \qquad n = 1, 2, 3, \ldots$$

$$r_n = \left(\frac{h^2}{4\pi^2 mke^2}\right)\frac{n^2}{Z} \qquad n = 1, 2, 3, \ldots$$

Radii for Bohr orbits

$$r_n = \left(5.29\times10^{-11}\,\mathrm{m}\right)\frac{n^2}{Z} \qquad n = 1, 2, 3, \ldots$$

30.3 The Bohr Model of the Hydrogen Atom



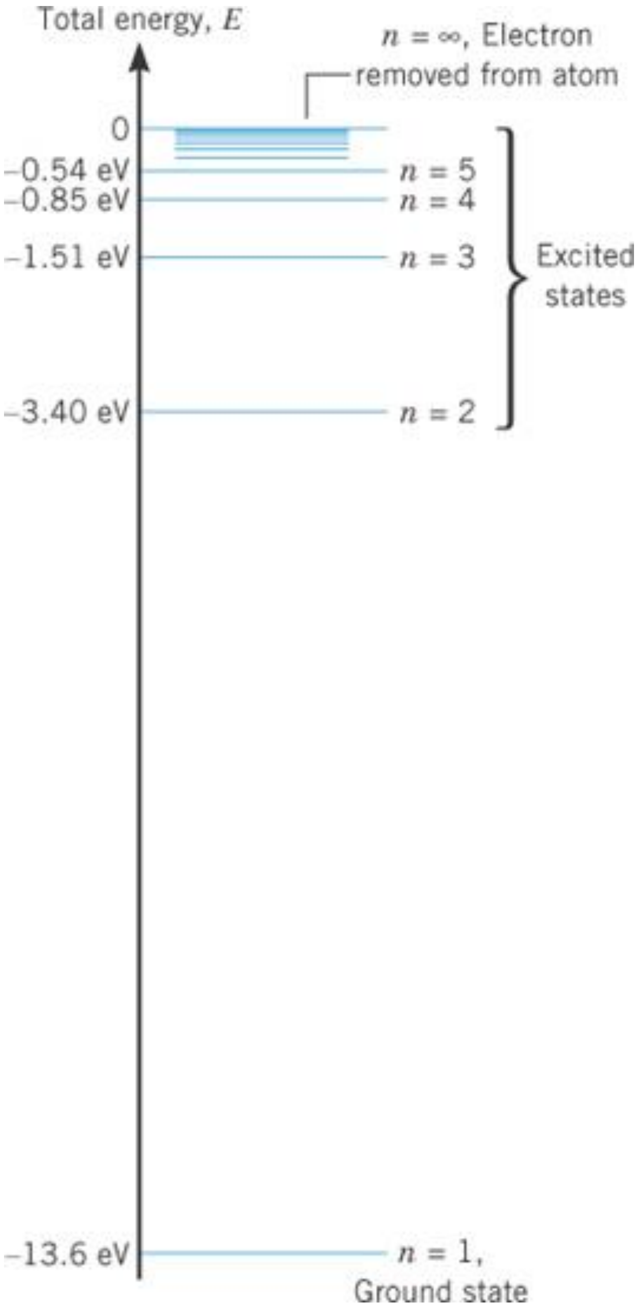
$$E_n = -\left(\frac{2\pi^2 mk^2e^4}{h^2}\right)\frac{Z^2}{n^2} \quad n = 1, 2, 3, \ldots$$

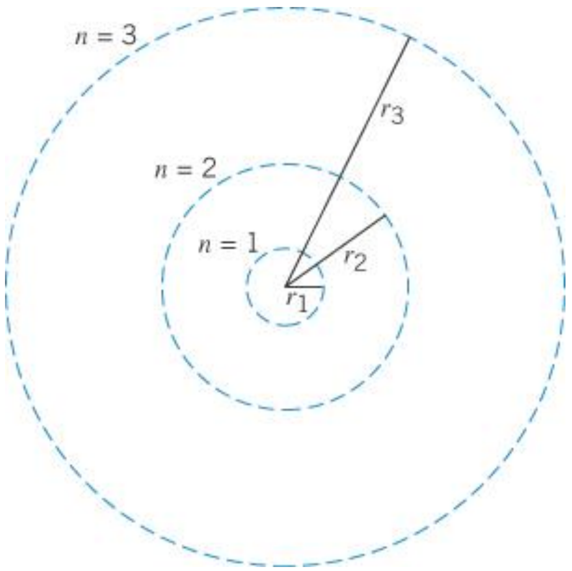
Bohr energy levels

$$E_n = -\left(2.18\times10^{-18}\,\text{J}\right)\frac{Z^2}{n^2} \quad n = 1, 2, 3, \ldots$$

$$E_n = -\left(13.6\,\text{eV}\right)\frac{Z^2}{n^2} \quad n = 1, 2, 3, \ldots$$

ENERGY LEVEL DIAGRAMS





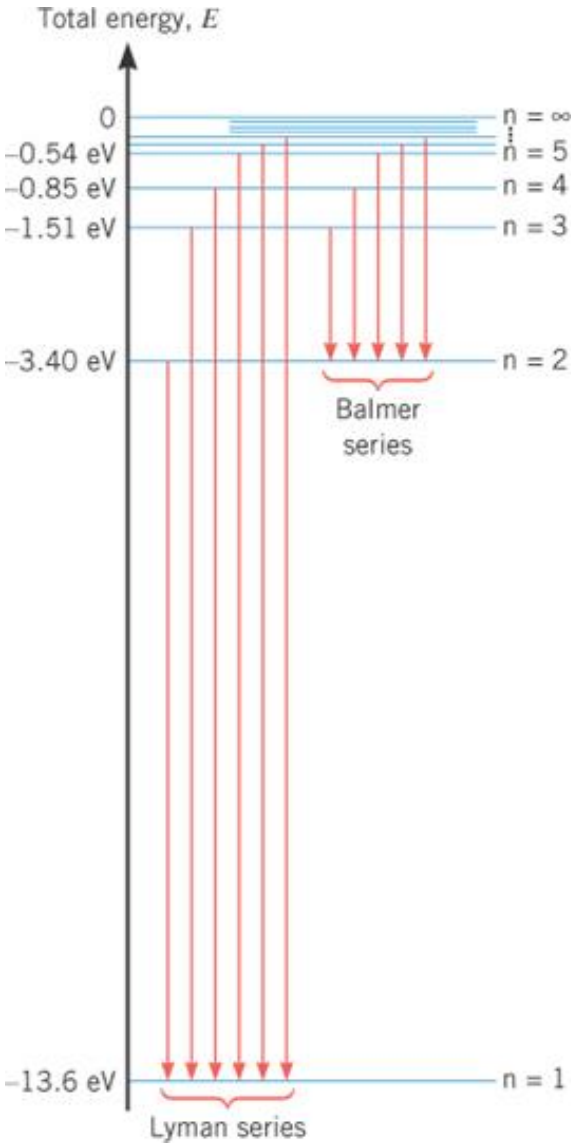
Example 3 The Ionization Energy of Li^{2+}

Li^{2+} is a lithium atom (*Z=3*) with only one electron. Obtain the ionization energy of Li^{2+}.

$$E_n = -(13.6\,\text{eV})\frac{Z^2}{n^2} = -(13.6\,\text{eV})\frac{3^2}{1^2} = -122\,\text{eV}$$

$$\text{Ionization energy} = +122\,\text{eV}$$

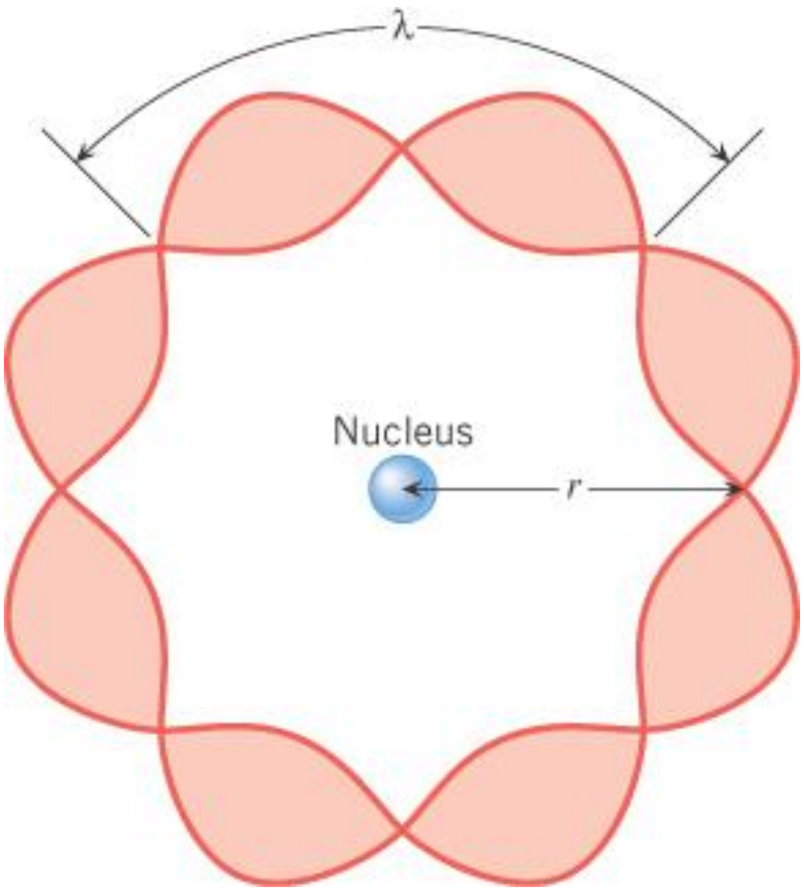
THE LINE SPECTRA OF THE HYDROGEN ATOM



$$\frac{1}{\lambda} = \frac{2\pi^2 m k^2 e^4}{h^3 c}\left(Z^2\right)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$n_i, n_f = 1, 2, 3, \ldots \qquad n_i > n_f$$

30.4 De Broglie's Explanation of Bohr's Assumption About Angular Momentum



$$2\pi r = n\lambda \quad n = 1, 2, 3, \ldots$$

De Broglie suggested standing particle waves as an explanation for Bohr's angular momentum assumption.

30.5 The Quantum Mechanical Picture of the Hydrogen Atom

Quantum mechanics reveals that four different quantum numbers are required to describe each state of the Hydrogen atom.

1. The principal quantum number *n*. This number determines the total energy of the atom and can have only integer values.

$$n = 1, 2, 3, \ldots$$

2. The orbital quantum number *l*. This number determines the orbital angular momentum of the electron integer values.

$$\ell = 1, 2, 3, \ldots, (n-1)$$

angular momentum

$$L = \sqrt{\ell(\ell+1)}\,\frac{h}{2\pi}$$

3. The magnetic quantum number m_l. This number determines the effect of a magnetic field on the energy of the atom.

$$m_\ell = -\ell, \ldots -2, -1, 0, 1, 2, \ldots, +\ell$$

$$L_z = m_\ell \frac{h}{2\pi}$$

z component of the angular momentum

4. The spin quantum number m_s. This number is needed because the electron has an intrinsic property called spin.

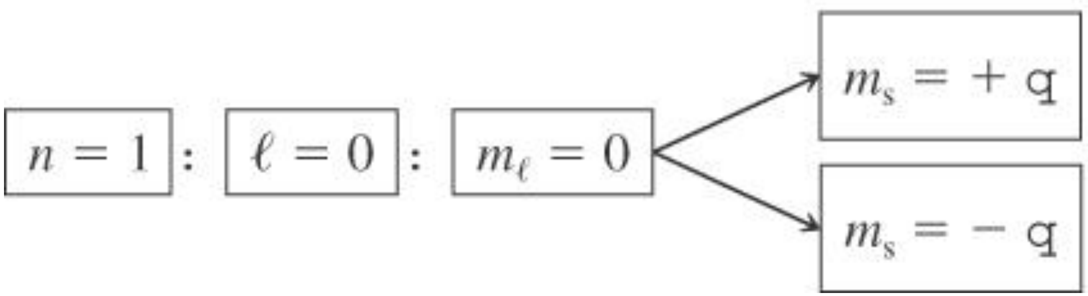
$$m_s = +\tfrac{1}{2} \quad \text{or} \quad -\tfrac{1}{2}$$

Table 30.1 Quantum Numbers for the Hydrogen Atom

Name	Symbol	Allowed Values
Principal quantum number	n	1, 2, 3, . . .
Orbital quantum number	ℓ	0, 1, 2, . . . , $(n - 1)$
Magnetic quantum number	m_ℓ	$-\ell, \ldots, -2, -1, 0, +1, +2, \ldots, +\ell$
Spin quantum number	m_s	$-\frac{1}{2}, +\frac{1}{2}$

Example 5 The Bohr Model Versus Quantum Mechanics

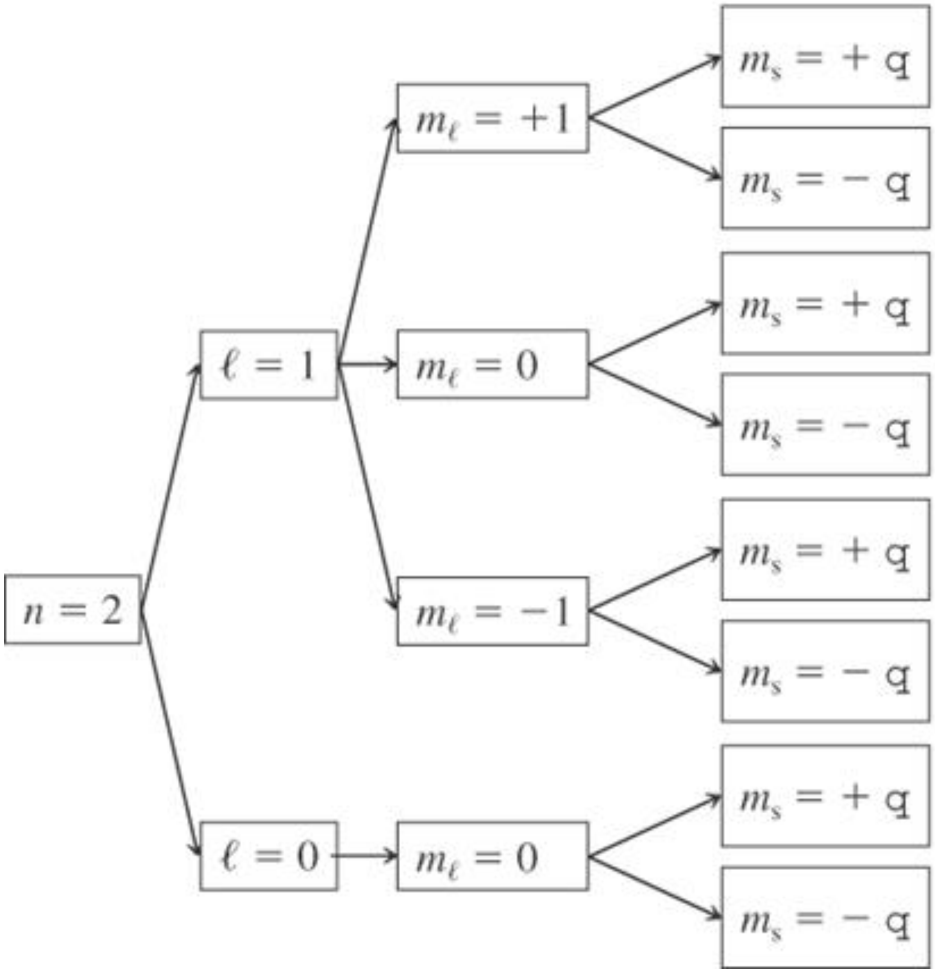
Determine the number of possible states for the hydrogen atom when the principal quantum number is (a) n=1 and (b) n=2.



State			
n	ℓ	m_ℓ	m_s
1	0	0	$+q$
1	0	0	$-q$

q = ½

30.5 The Quantum Mechanical Picture of the Hydrogen Atom



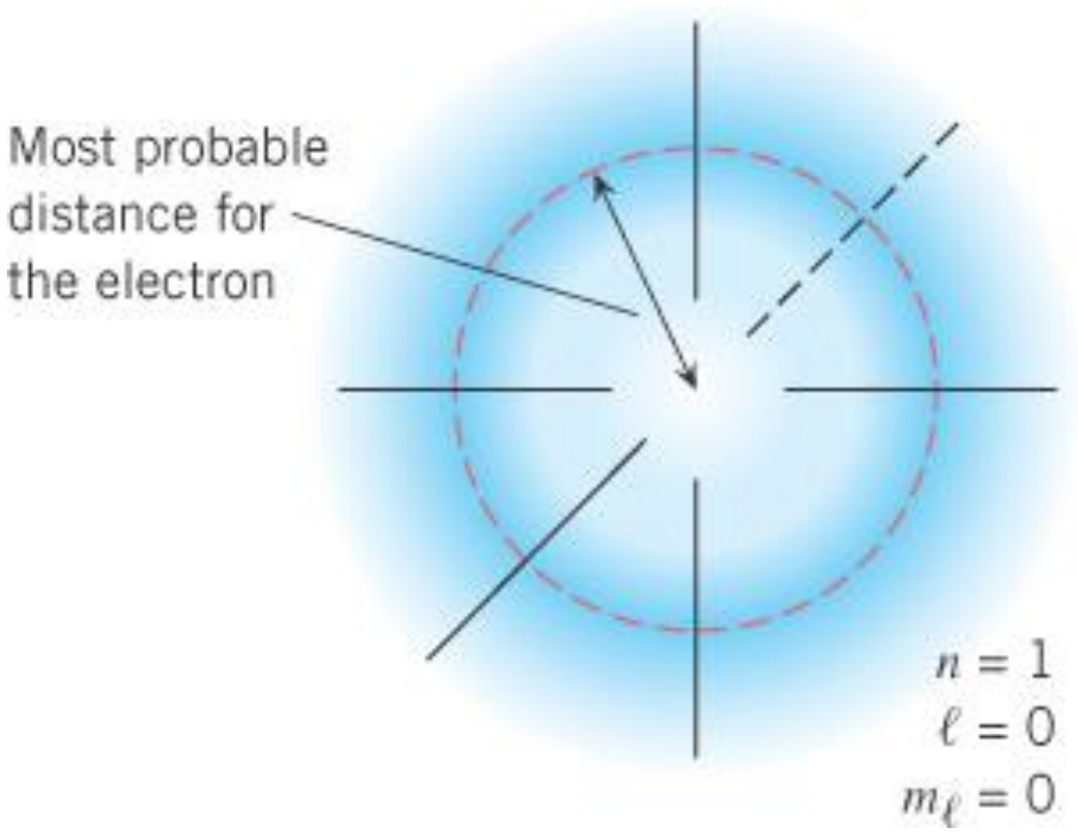
State			
n	ℓ	m_ℓ	m_s
2	1	+1	+ q
2	1	+1	− q
2	1	0	+ q
2	1	0	− q
2	1	−1	+ q
2	1	−1	− q
2	0	0	+ q
2	0	0	− q

q = ½

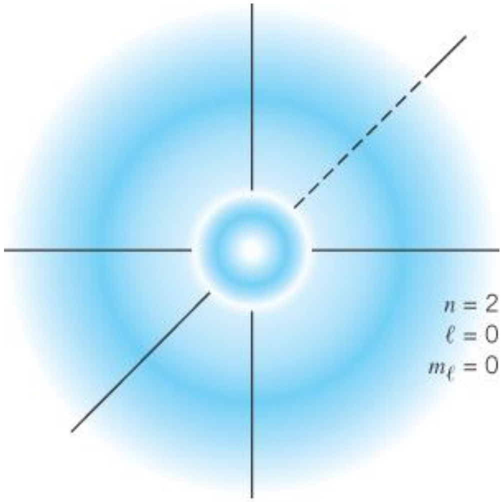
Conceptual Example 6 The Bohr Model Versus Quantum Mechanics

Consider two hydrogen atoms. There are no external magnetic fields present, and the electron in each atom has the same energy. According to the Bohr model and to quantum mechanics, is it possible for the electrons in these atoms (a) to have zero orbital angular momentum and (b) to have different angular momenta?

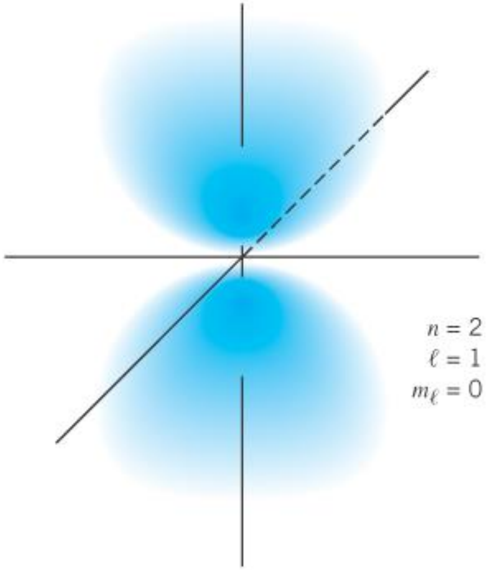
30.5 The Quantum Mechanical Picture of the Hydrogen Atom



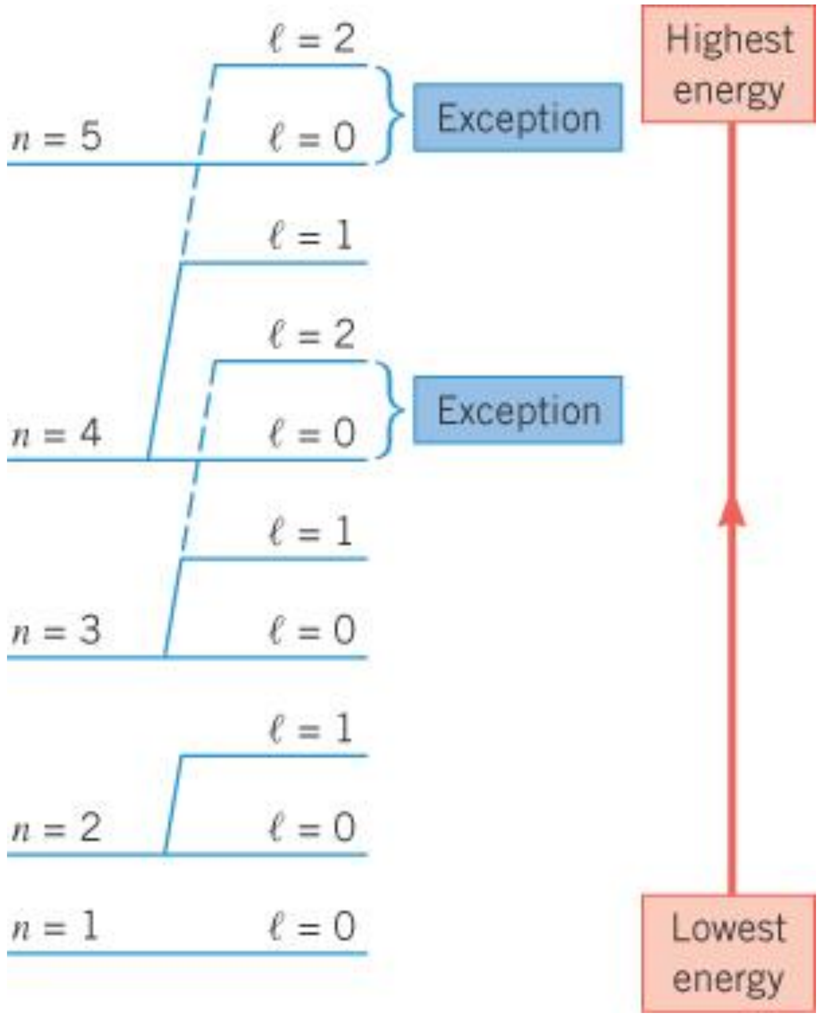
30.5 The Quantum Mechanical Picture of the Hydrogen Atom



(a)



(b)



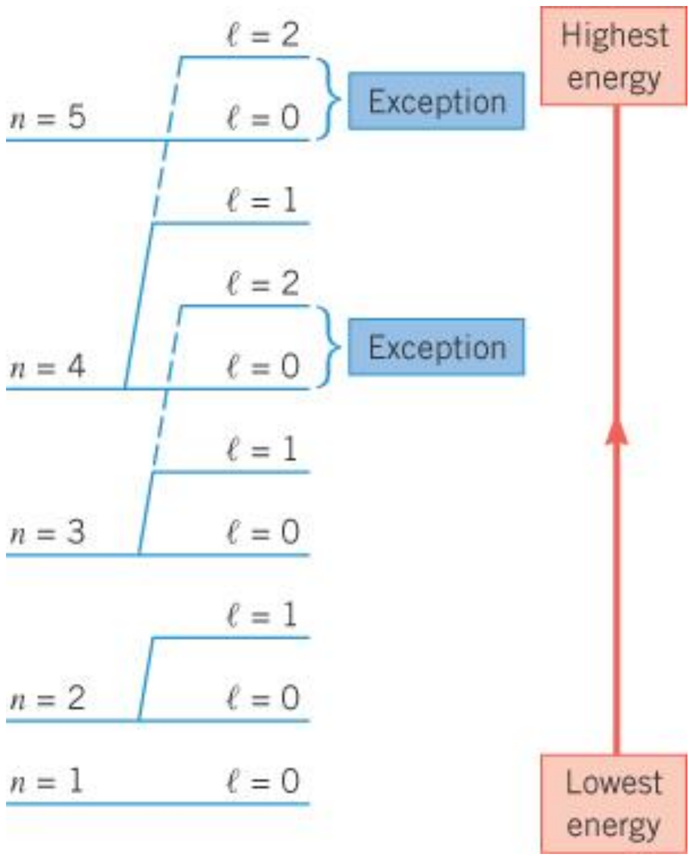
Generally, the energy increases with increasing n. There are exceptions to the general rule.

THE PAULI EXCLUSION PRINCIPLE

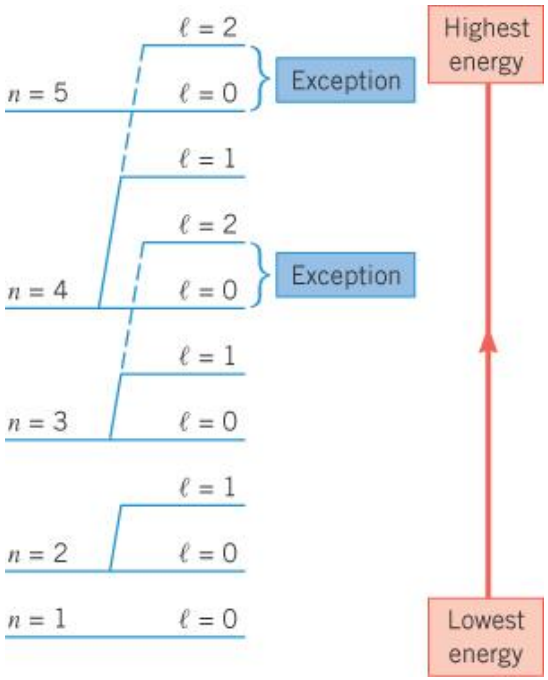
No two electrons in an atom can have the same set of values for the four quantum numbers.

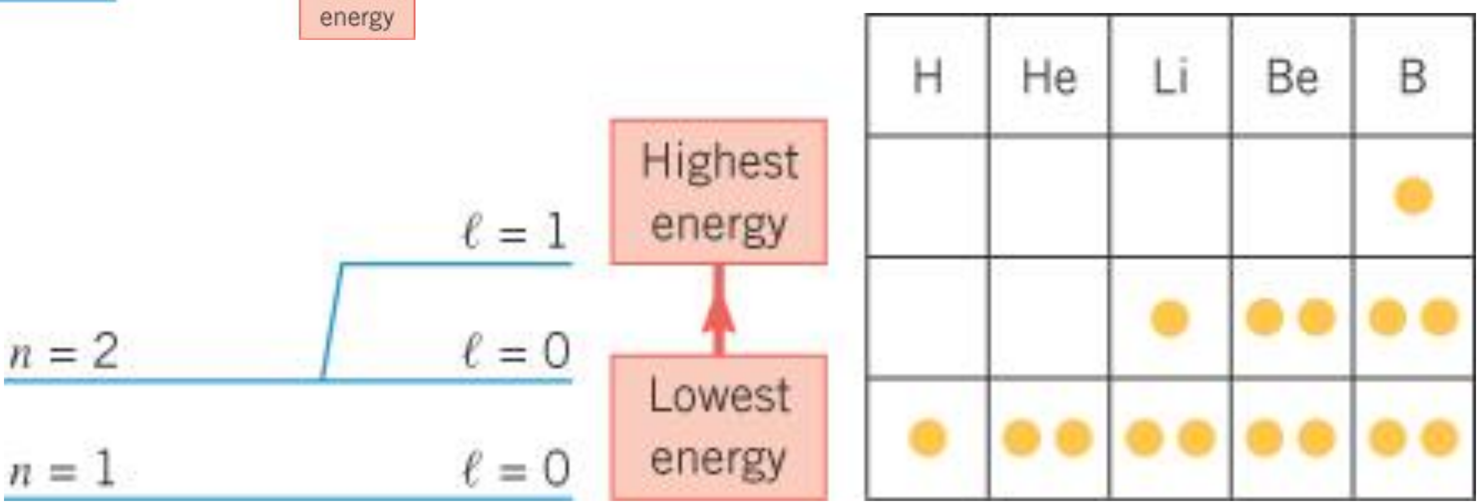
Example 8 Ground States of Atoms

Determine which of the energy levels in the figure are occupied by the electrons in the ground state of hydrogen, helium, lithium, beryllium, and boron.



30.6 The Pauli Exclusion Principle and the Periodic Table of the Elements





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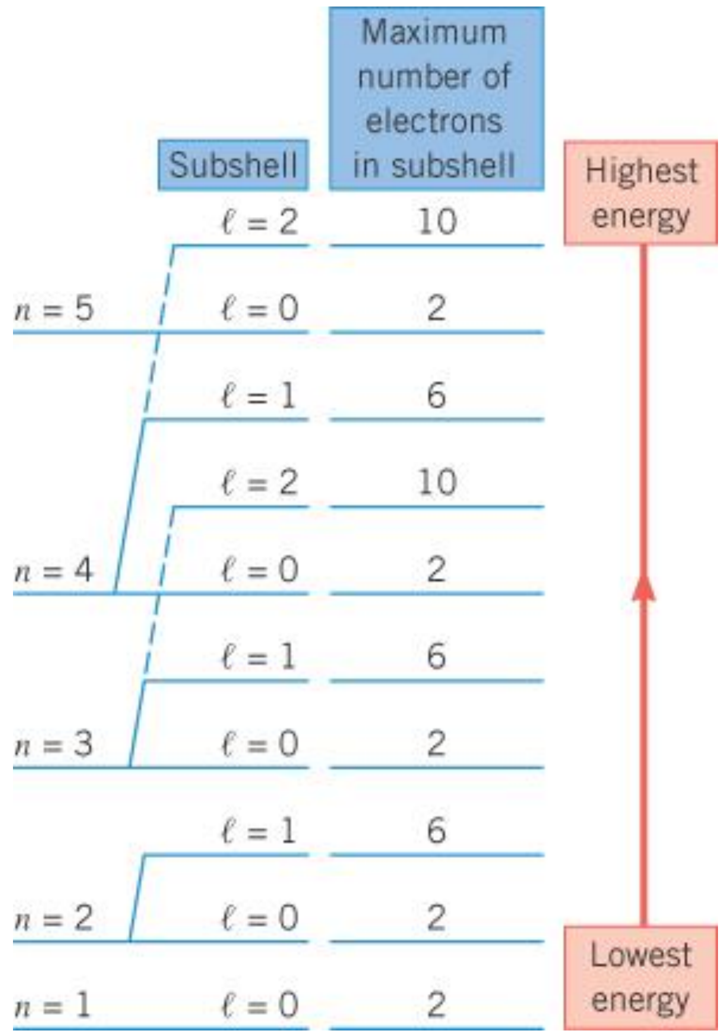


30.6 The Pauli Exclusion Principle and the Periodic Table of the Elements

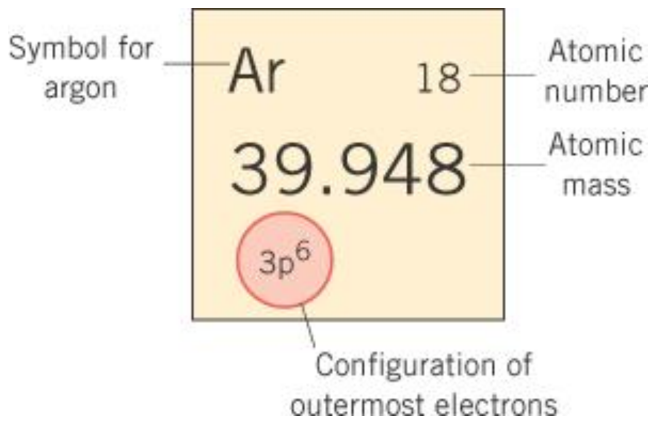


Table 30.2 The Convention of Letters Used to Refer to the Orbital Quantum Number

Orbital Quantum Number ℓ	Letter
0	s
1	p
2	d
3	f
4	g
5	h

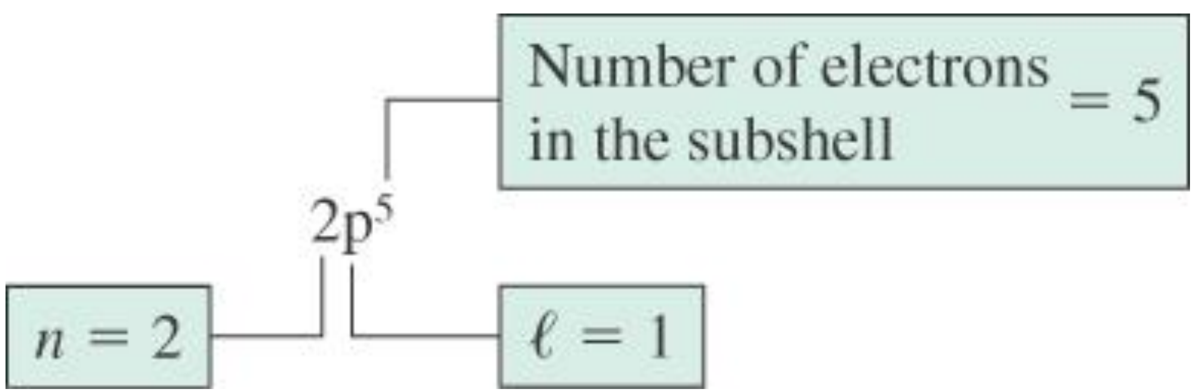
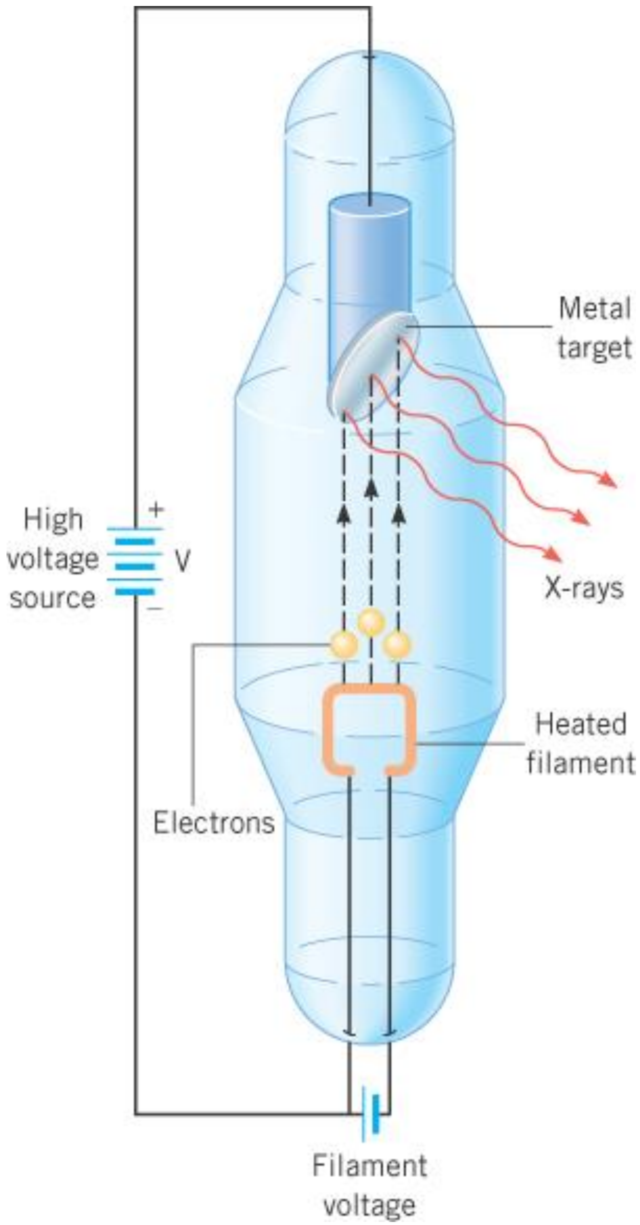


Table 30.3 Ground-State Electronic Configurations of Atoms

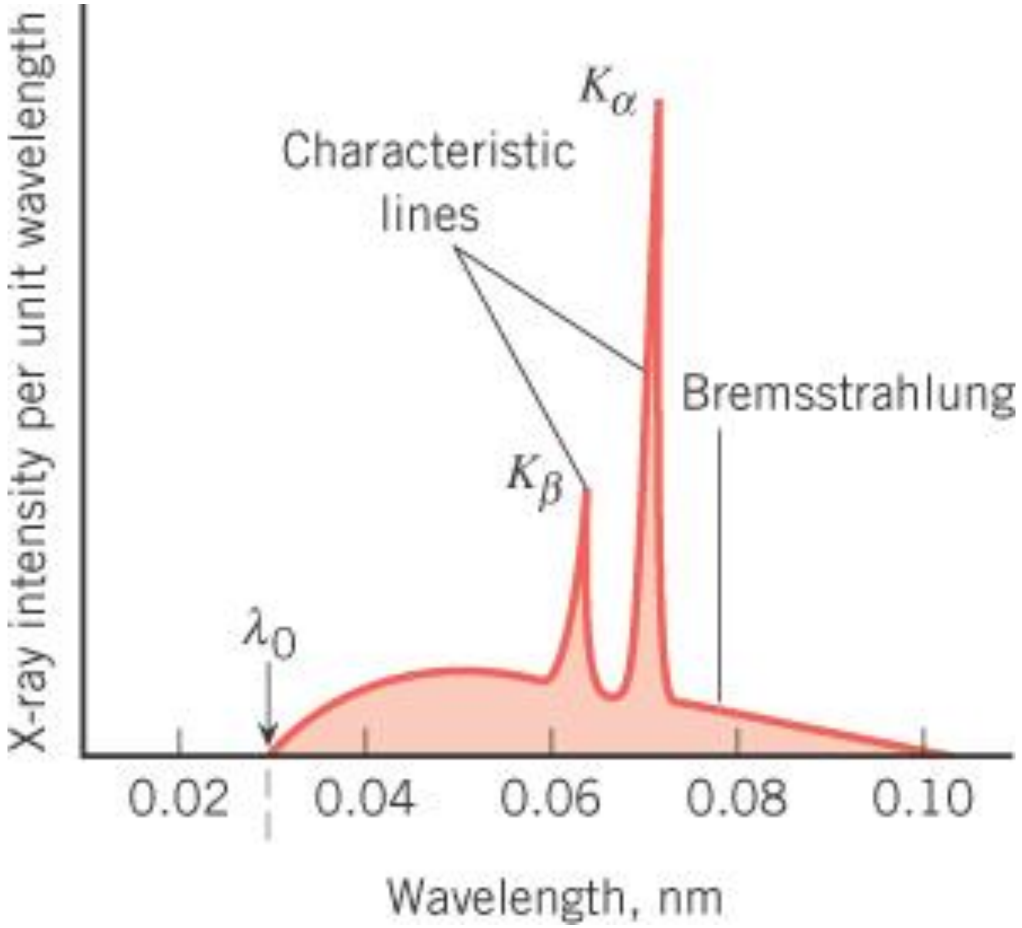
Element	Number of Electrons	Configuration of the Electrons
Hydrogen (H)	1	$1s^1$
Helium (He)	2	$1s^2$
Lithium (Li)	3	$1s^2\ 2s^1$
Beryllium (Be)	4	$1s^2\ 2s^2$
Boron (B)	5	$1s^2\ 2s^2\ 2p^1$
Carbon (C)	6	$1s^2\ 2s^2\ 2p^2$
Nitrogen (N)	7	$1s^2\ 2s^2\ 2p^3$
Oxygen (O)	8	$1s^2\ 2s^2\ 2p^4$
Fluorine (F)	9	$1s^2\ 2s^2\ 2p^5$
Neon (Ne)	10	$1s^2\ 2s^2\ 2p^6$
Sodium (Na)	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium (Mg)	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminum (Al)	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

30.7 X-Rays



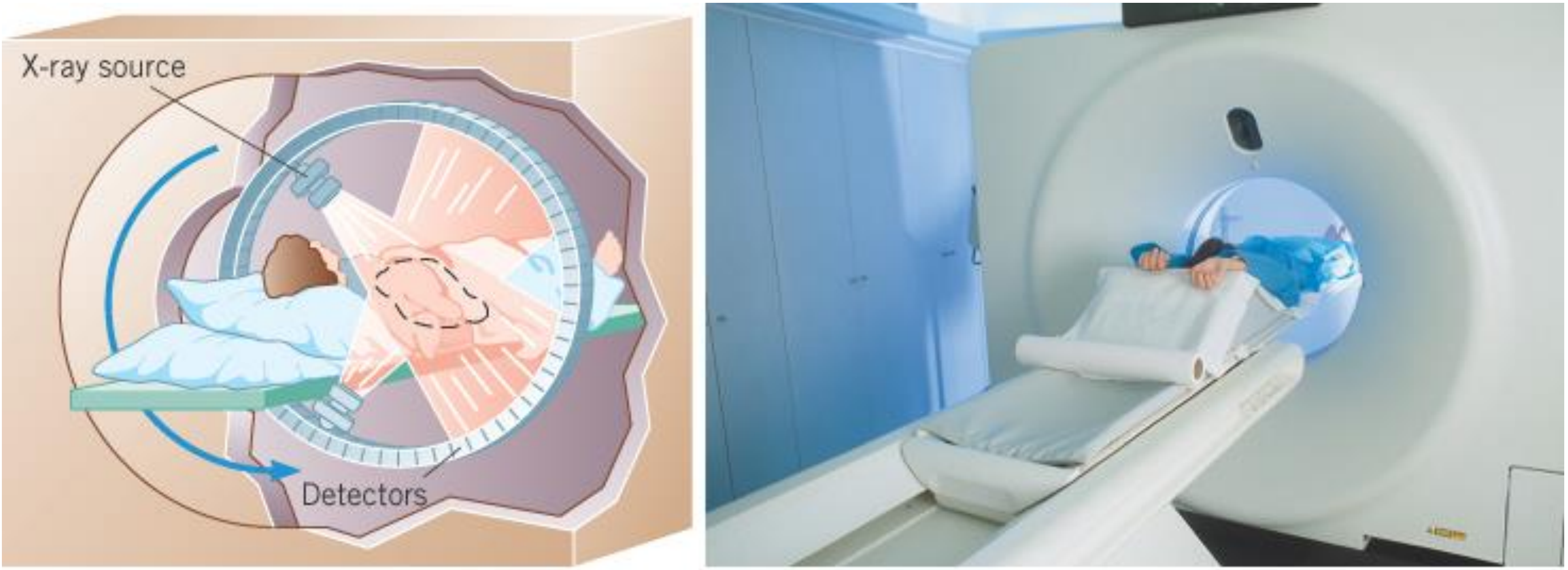
Electrons are emitted from a heated filament and accelerated through a large voltage.

When they strike the target, X-rays are emitted.

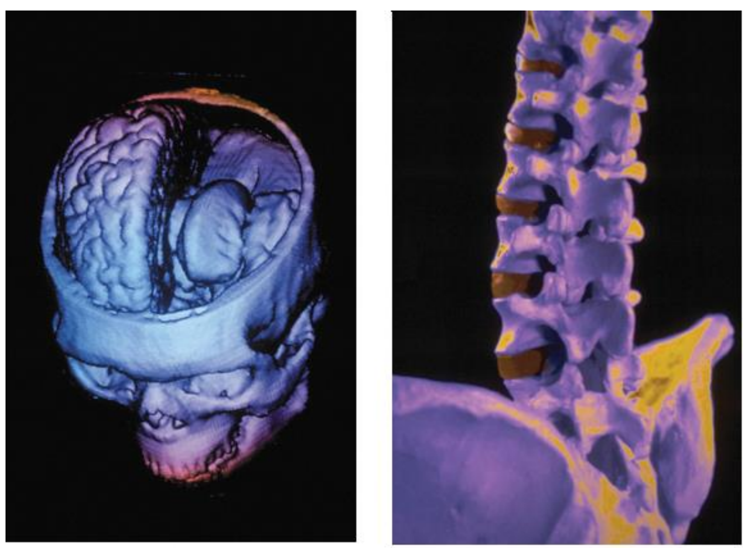


The sharp peaks are called ***characteristic X-rays*** because they are characteristic of the target material.

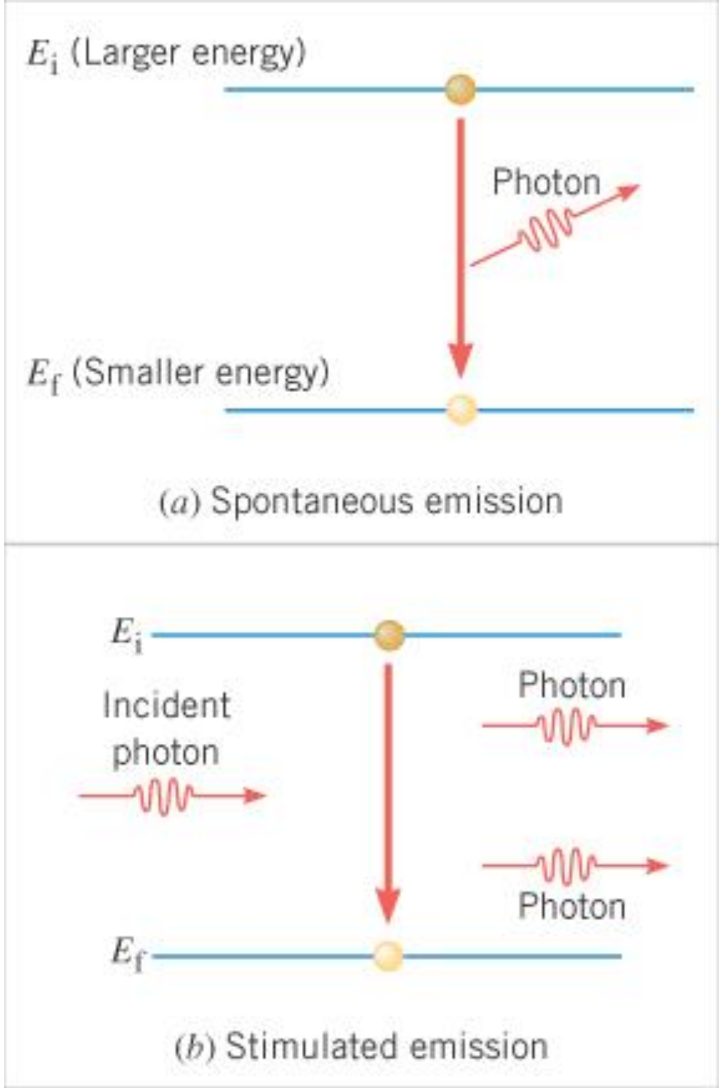
30.7 X-Rays



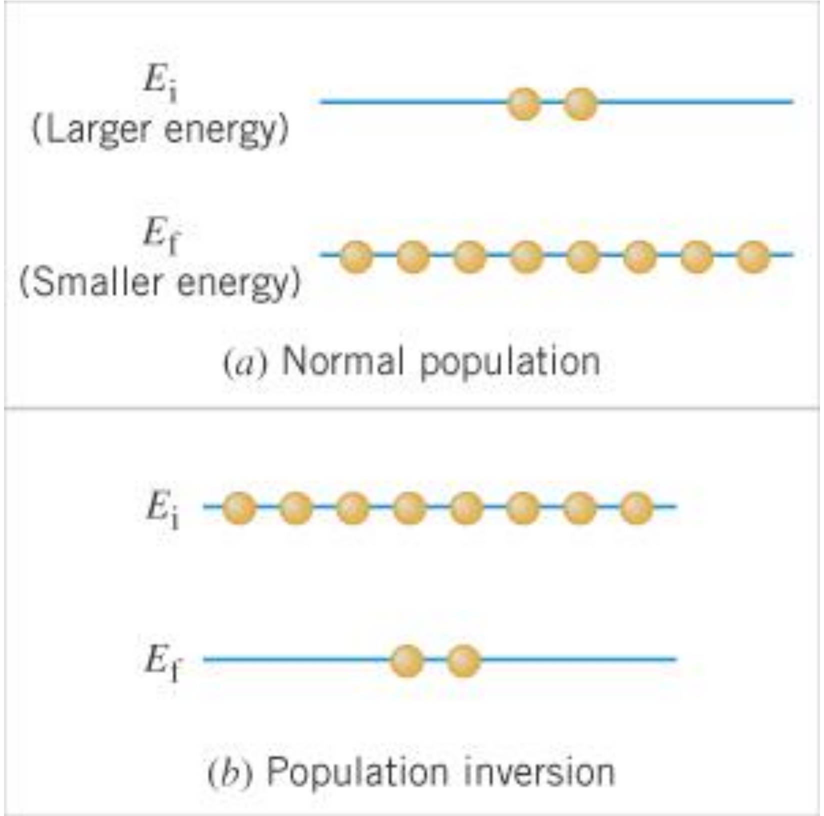
(a) (b)



(a) (b)



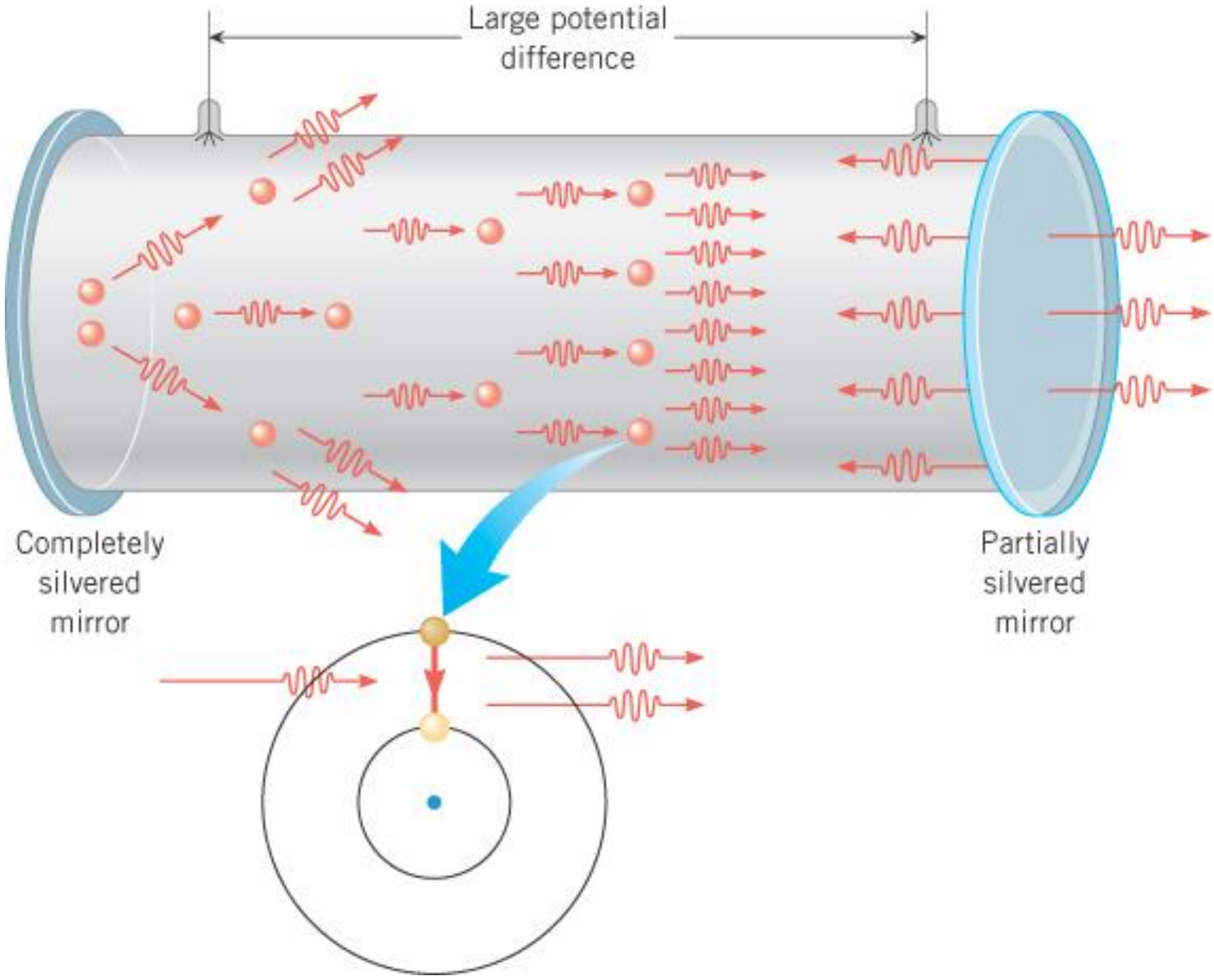
Spontaneous emission versus stimulated emission.



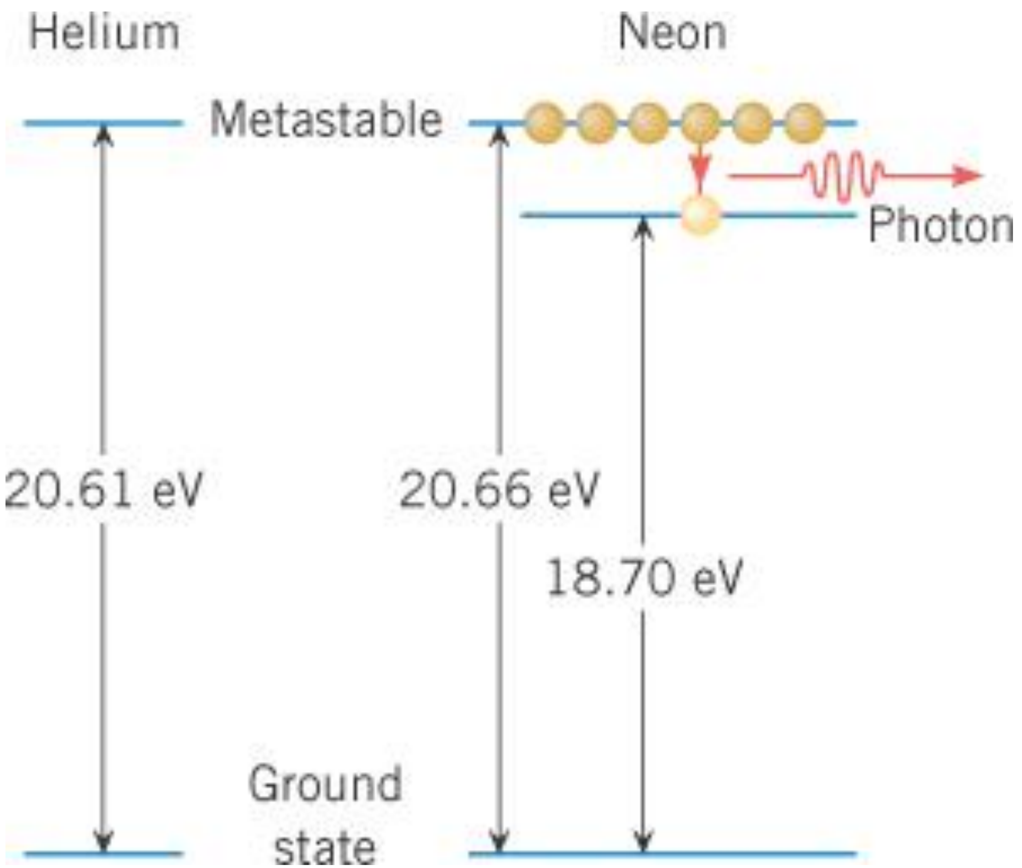
(a) Normal population

(b) Population inversion

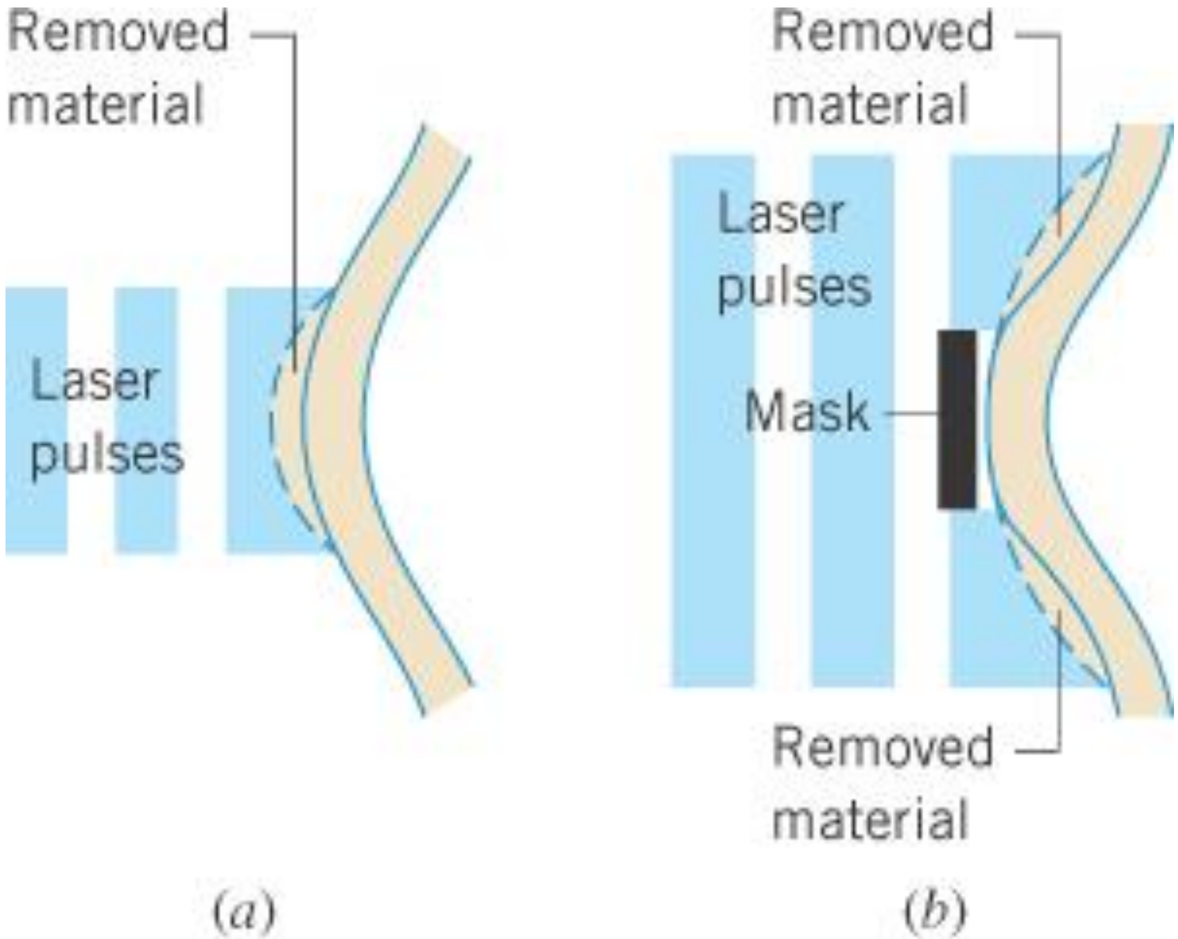
An external energy source populates the higher level with electrons.


Large potential
difference
Completely
silvered
mirror
Partially
silvered
mirror

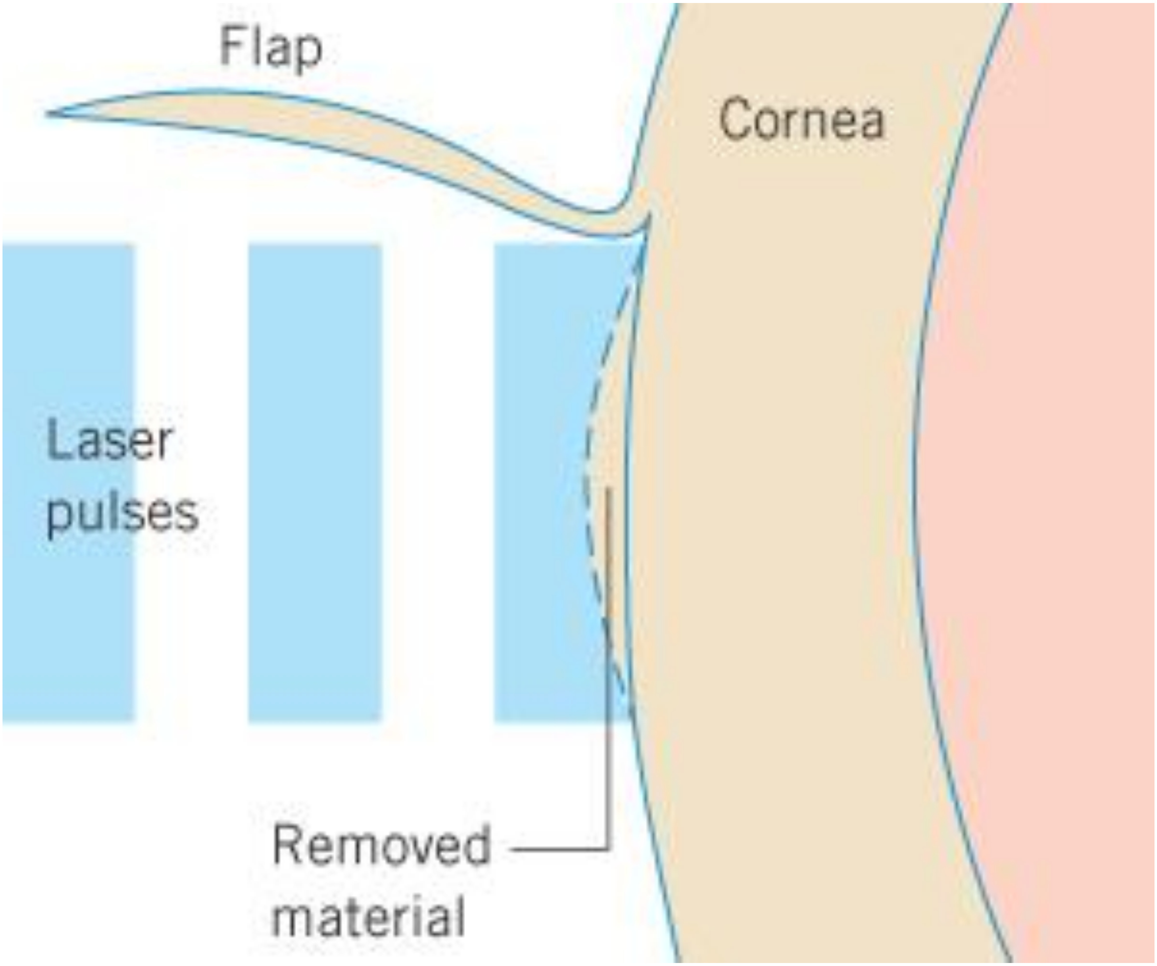
30.8 The Laser



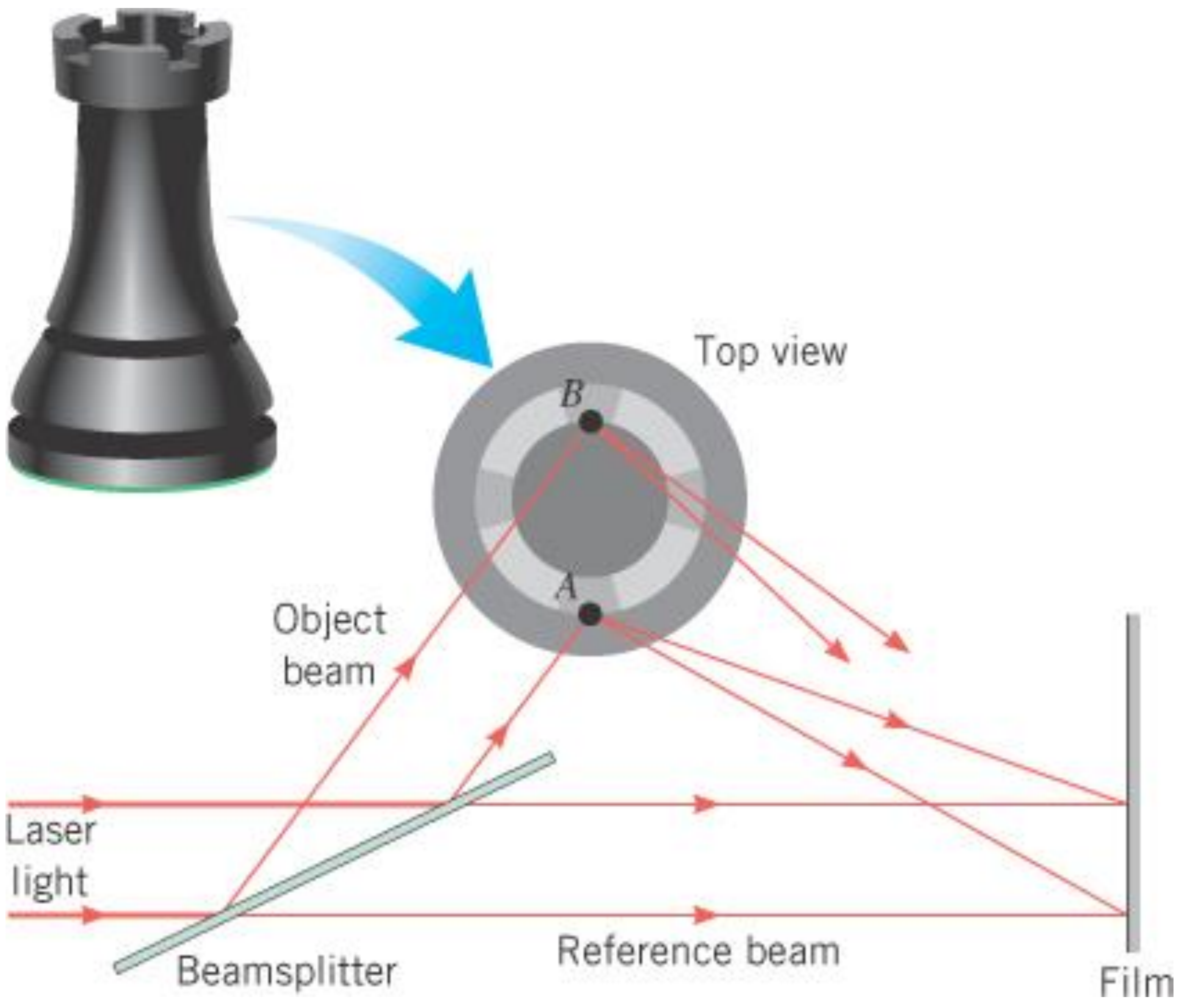
30.9 Medical Applications of the Laser



Lasers being used to change the shape of the cornea.


Flap
Cornea
Laser
pulses
Removed
material

30.10 Holography



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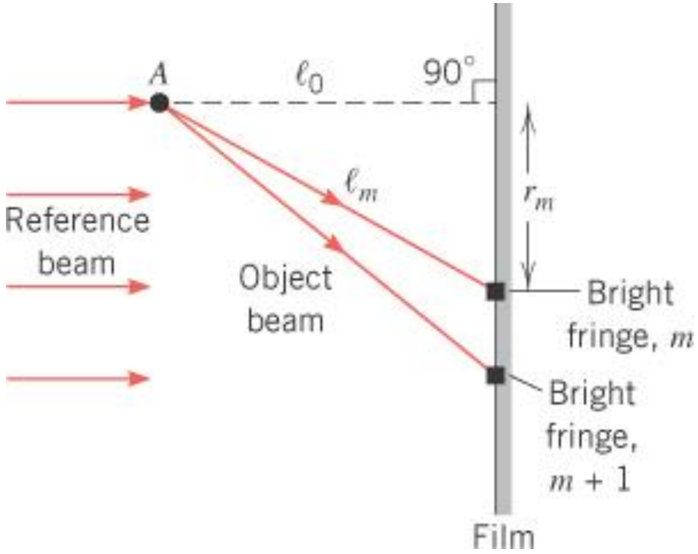
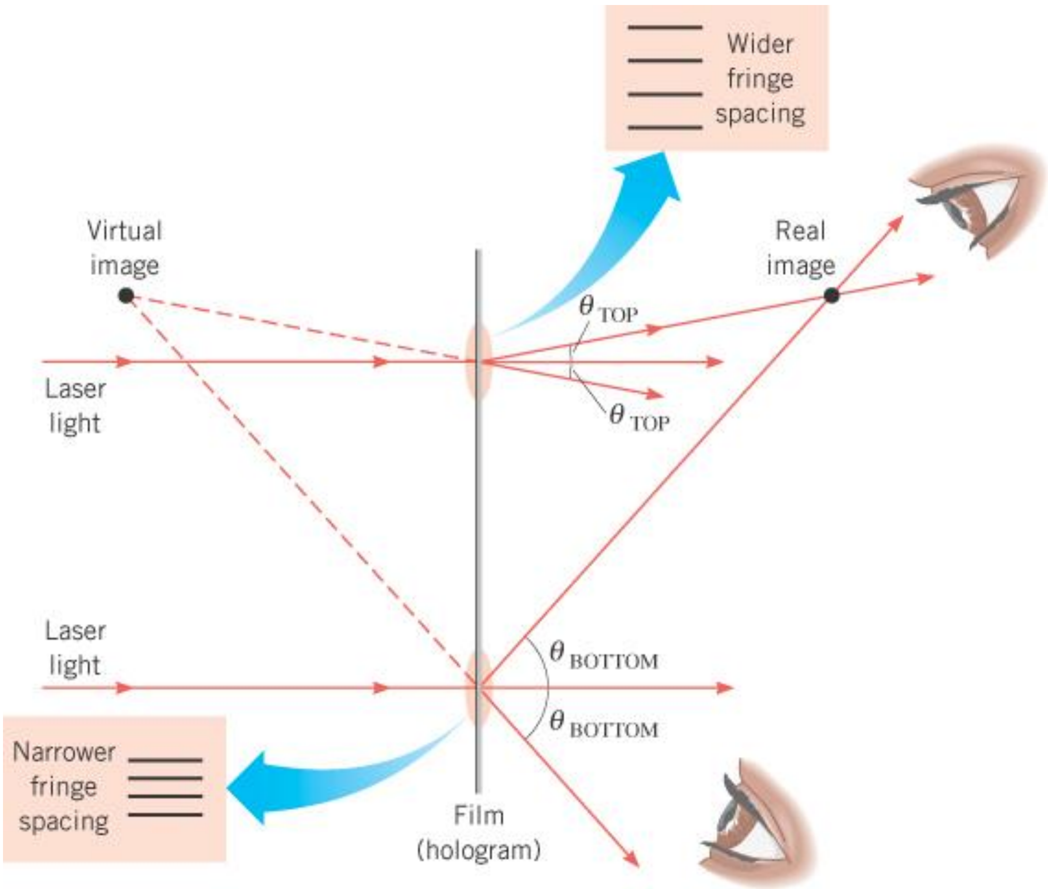
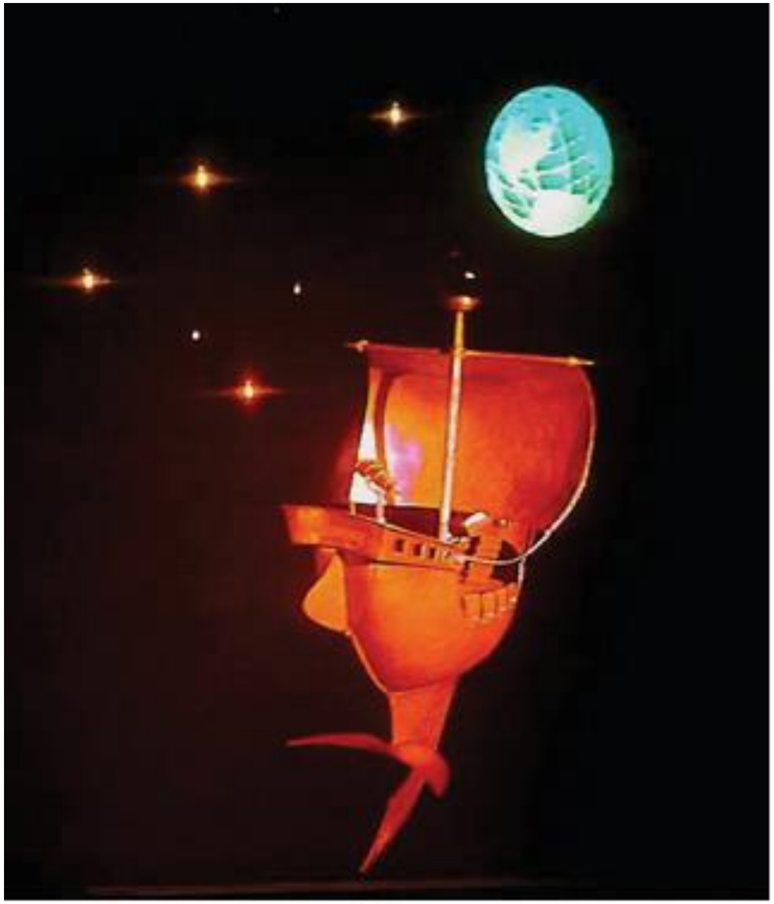
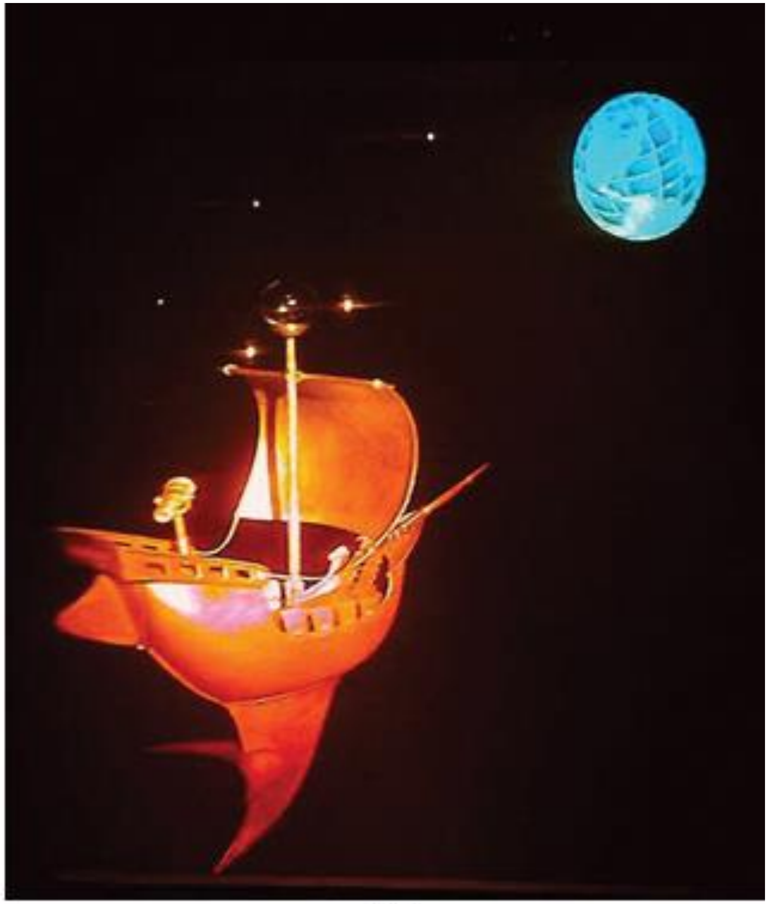




30.10 Holography



(*a*)



(*b*)