



Chapter 31 Atomic Physics

31.1 Early Model of the Atom

31.2 The Spectrum of Atomic Hydrogen

31.3 Bohr's Model of the Hydrogen Atom

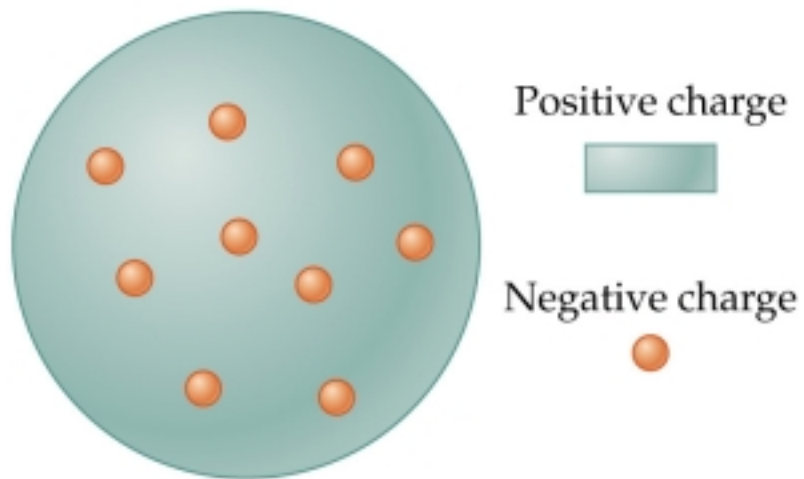
31.4 de Broglie Waves and the Bohr Model

31.5 The Quantum Mechanical Hydrogen
Atom

31.6 Multielectron Atoms and the Periodic
Table

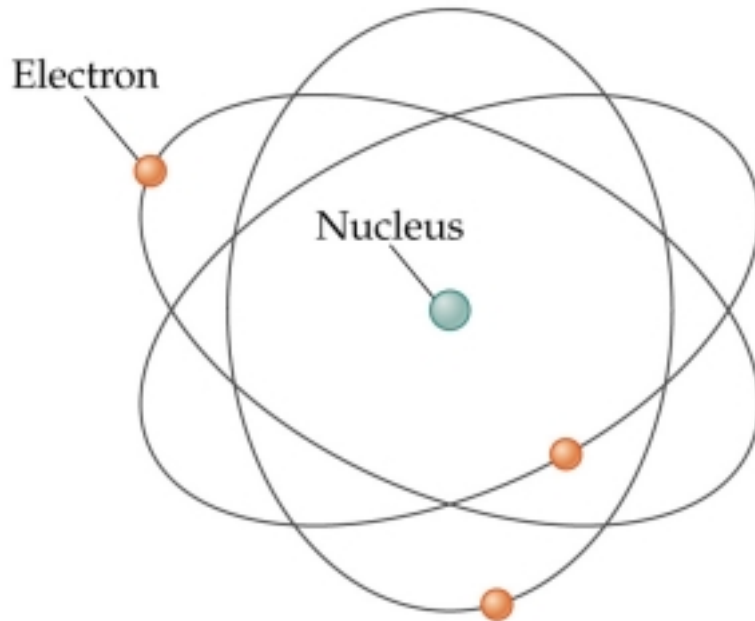
**** The idea that all matter is composed of atoms is fundamental to our modern view of the world. It has given us a firm basis for understanding the properties of solid, liquids, and gases. This understanding has led to a host of useful devices, one of the most famous being the laser. The laser beams arise because atoms generate light as they undergo transitions from a higher to a lower atomic energy level. Our venture into the atomic world begins with the concept of the nuclear atom and continues with the Bohr model of the hydrogen atom. This model introduces many basic atomic features, including the notions of discrete energy states and quantum numbers. We will see, however, that quantum mechanics has displaced the Bohr model and provides a more complete description of the atom. Following an overview of quantum mechanics, the Pauli exclusion principle will be discussed. This principle is important for explaining how electrons are arranged in complex atoms and how the elements are ordered in the periodic table.

Figure 31-1 The plum pudding model of an atom



- The model of an atom proposed by J. J. Thomson consists of a uniform positive charge, which accounts for most of the mass of an atom, with small negatively charged electrons scattered throughout, like raisins in a pudding.

Figure 31-2 The solar system model of an atom



- Ernest Rutherford proposed that an atom is like a miniature solar system, with a massive, positively charged nucleus orbited by lightweight negatively charged electrons.

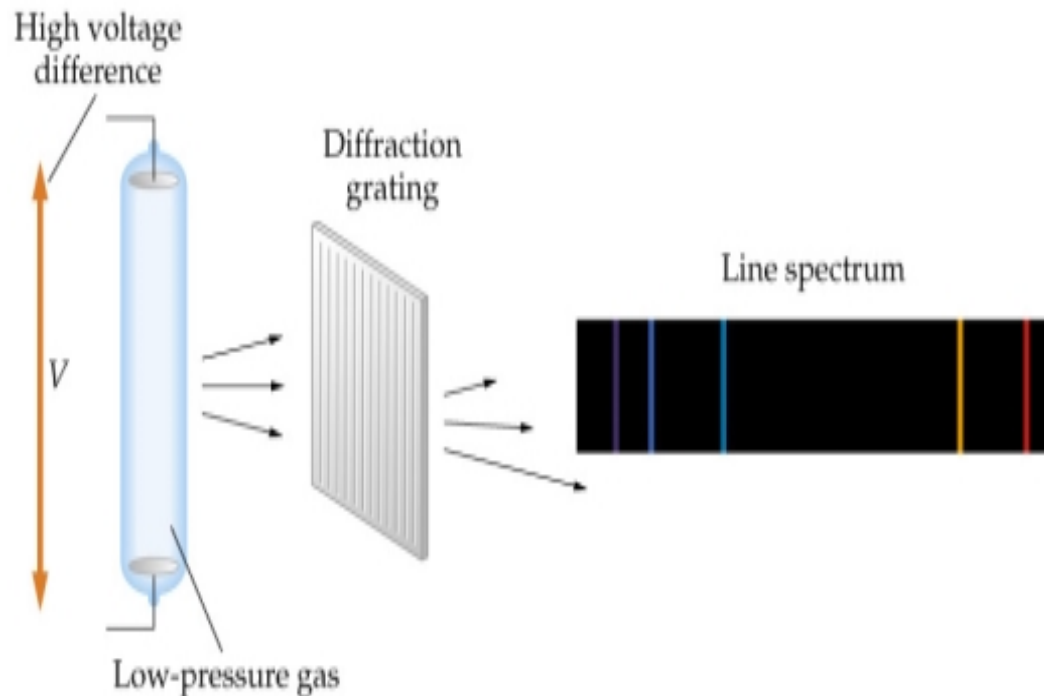
**** Nuclear atom

The idea of *nuclear atom* originated in 1911, as a result of experiments by *Ernest Rutherford*, in which alpha particles were scattered by a thin metal foil. The phrase "nuclear atom" refers to the fact that an atom consists of a small, positively charged nucleus surrounded at relatively large distance by a number of electrons, whose total negative charge equals the positive nuclear charge when the atom is electrically neutral.

Line spectrum

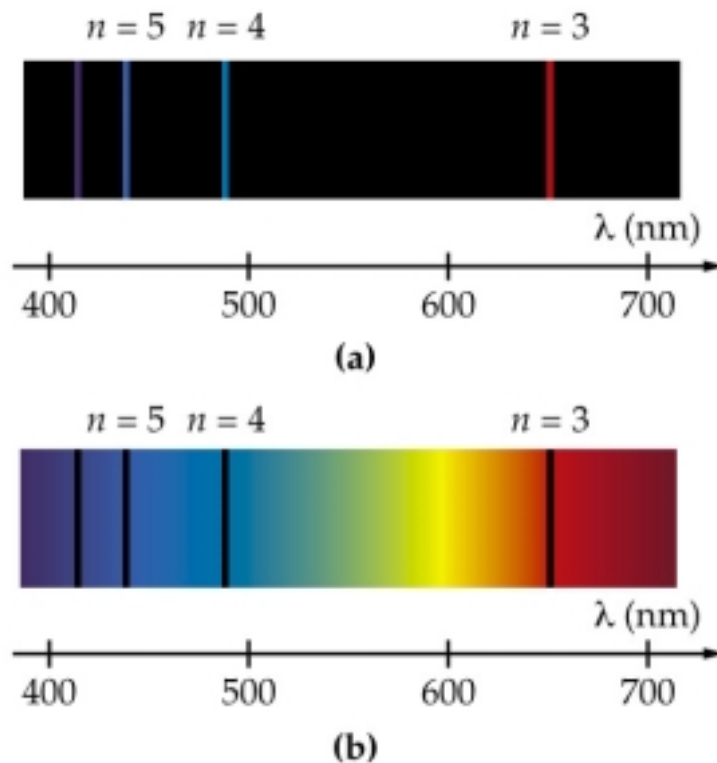
A *line spectrum* is a series of discrete electromagnetic wavelengths emitted by the atoms of a low-pressure gas that is subjected to a sufficiently high potential difference. Certain groups of discrete wavelengths are referred to as "series." The line spectrum of atomic hydrogen includes, among others, the *Lyman series*, the *Balmer series*, and the *Paschen series* of wavelength.

Figure 31-3 The line spectrum of an atom



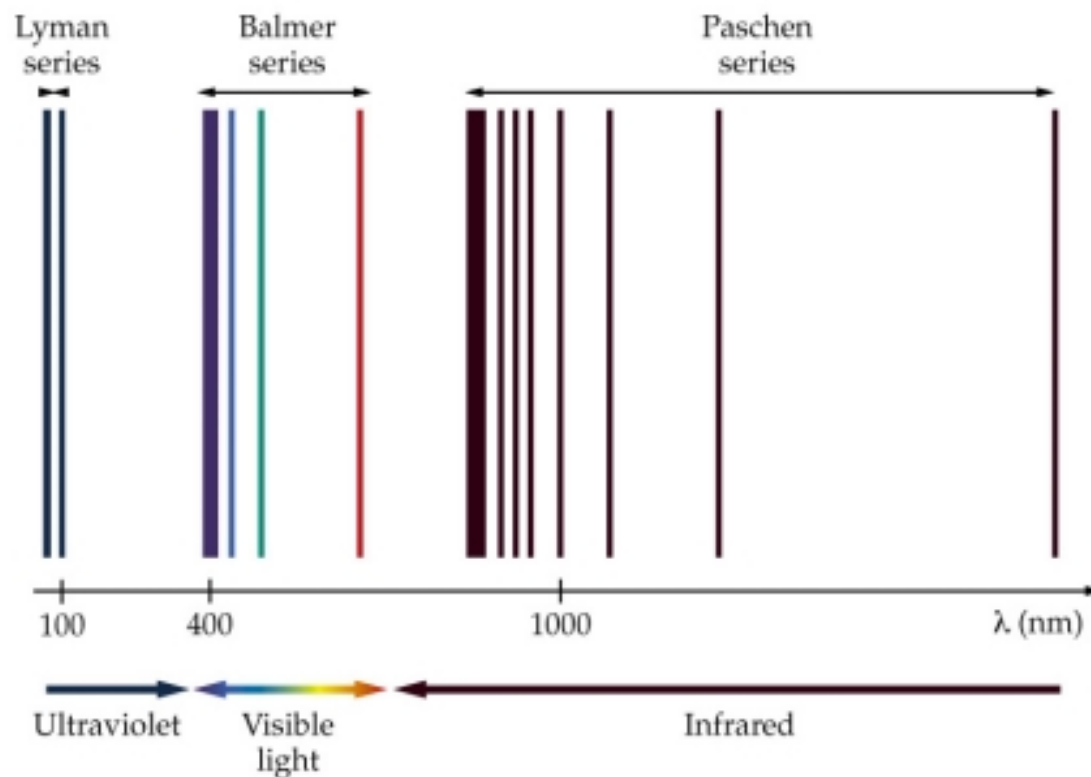
- The light given off by individual atoms, as in a low-pressure gas, consists of a series of discrete wavelengths corresponding to different colors.

Figure 31-4 The line spectrum of hydrogen



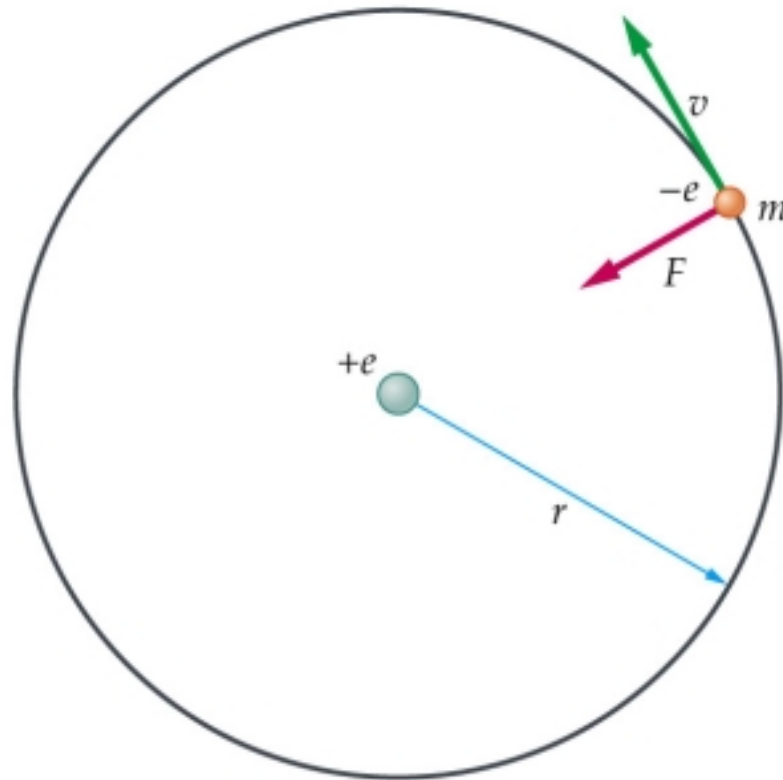
- The emission (a) and absorption (b) spectra of hydrogen. Note that the wavelengths absorbed by hydrogen (dark lines) are the same as those emitted by hydrogen (colored lines). The location of these lines is predicted by the Balmer series (Equation 31-1) with the appropriate values of n .

Figure 31-5 The Lyman, Balmer, and Paschen series of spectral lines



- The first three series of spectral lines in the spectrum of hydrogen. The shortest wavelengths appear in the Lyman series. There is no upper limit to the number of series in hydrogen, or to the wavelengths that can be emitted.

Figure 31-6 A Bohr orbit



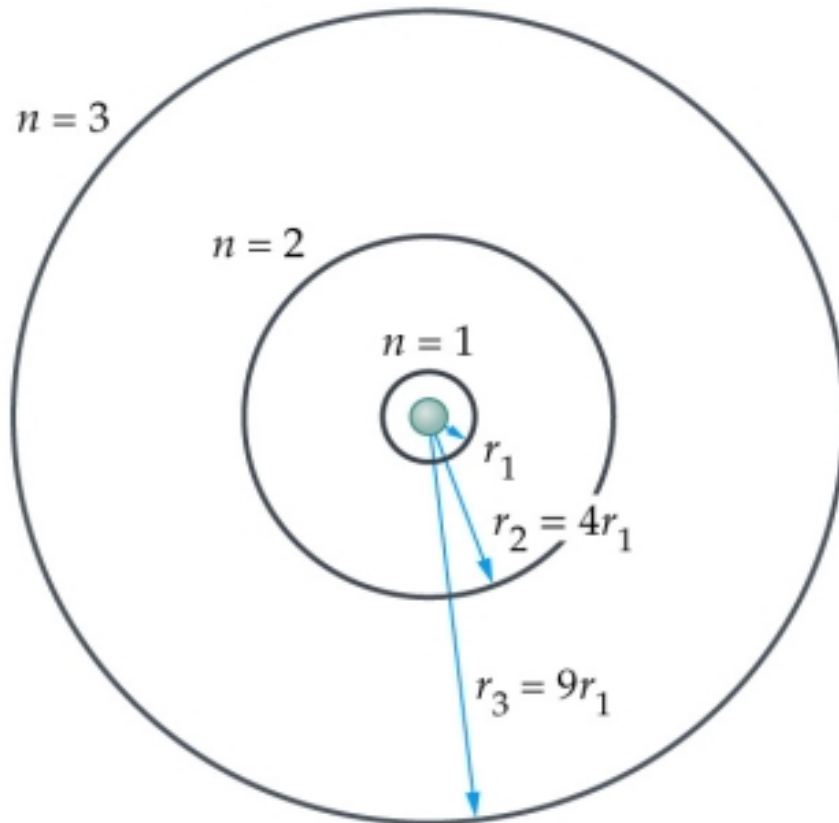
- In the Bohr model of hydrogen, electrons orbit the nucleus in circular orbits. The centripetal acceleration of the electron, v^2/r , is produced by the Coulomb force of attraction between the electron and the nucleus.

**** Bohr model (atom)

The *Bohr model* applies to atoms or ions that have only a single electron orbiting a nucleus containing Z protons. This model assumes that the electron exists in circular orbits that are called *stationary orbits*, because the electron does not radiate electromagnetic waves while in them. According to this model, a photon is emitted only when an electron changes from a higher energy orbit (level) to a lower one. The model also assumes that the orbital angular momentum L_n of the electron can only have values that are integer multiples of Planck's constant divided by 2π ,

$L_n = n(h/2\pi)$; $n = 1, 2, 3, \dots$ With the assumption above, it can be shown that the n th Bohr orbit has a radius (in meters) of $r_n = (5.29 \times 10^{-11} \text{ m})(n^2/Z)$ and that the total energy associated with this orbit is $E_n = -(13.6 \text{ eV})(Z^2/n^2)$.

Figure 31-7 The first three Bohr orbits



- The first Bohr orbit has a radius $r_1 = 5.29 \times 10^{-11} \text{ m}$. The second and third Bohr orbits have radii
 $r_2 = 2^2 r_1 = 4 r_1$,
and $r_3 = 3^2 r_1 = 9 r_1$,
respectively.

Figure 31-8 Energy-level diagram for the Bohr model of hydrogen



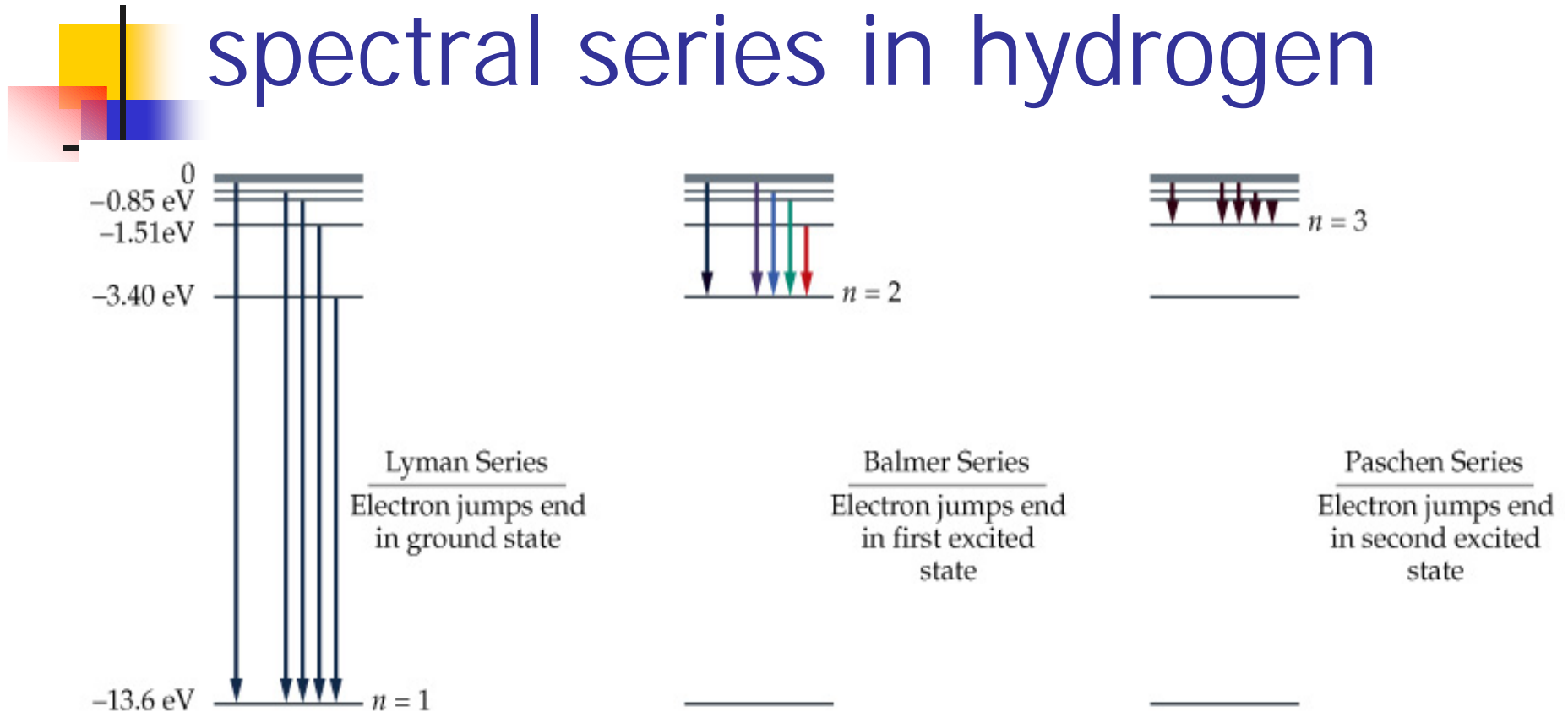
-13.6 eV ————— $n = 1$ Ground state

- The energy of the ground state of hydrogen is -13.6 eV . Excited states of hydrogen approach zero energy. Note that the difference in energy from the ground state to the first excited state is $3/4$ (13.6 eV), and the energy difference from the first excited state to the zero level is only $1/4$ (13.6 eV).

Ionization energy

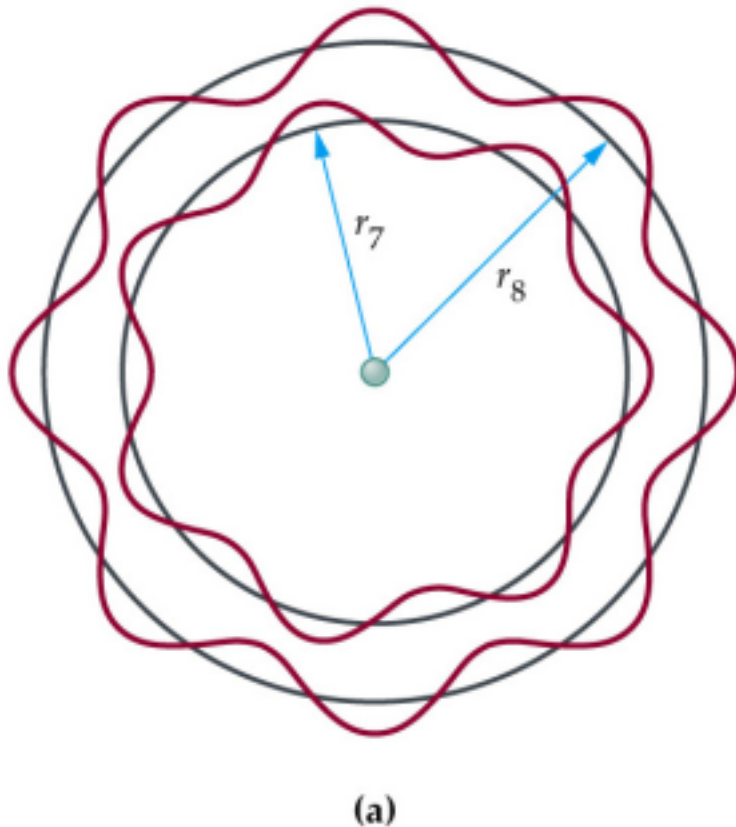
The *ionization energy* is the energy needed to remove an electron completely from an atom. The Bohr model predicts that the wavelengths comprising the line spectrum emitted by a hydrogen atom are given by Equation 31.2.

Figure 31-9 The origin of spectral series in hydrogen



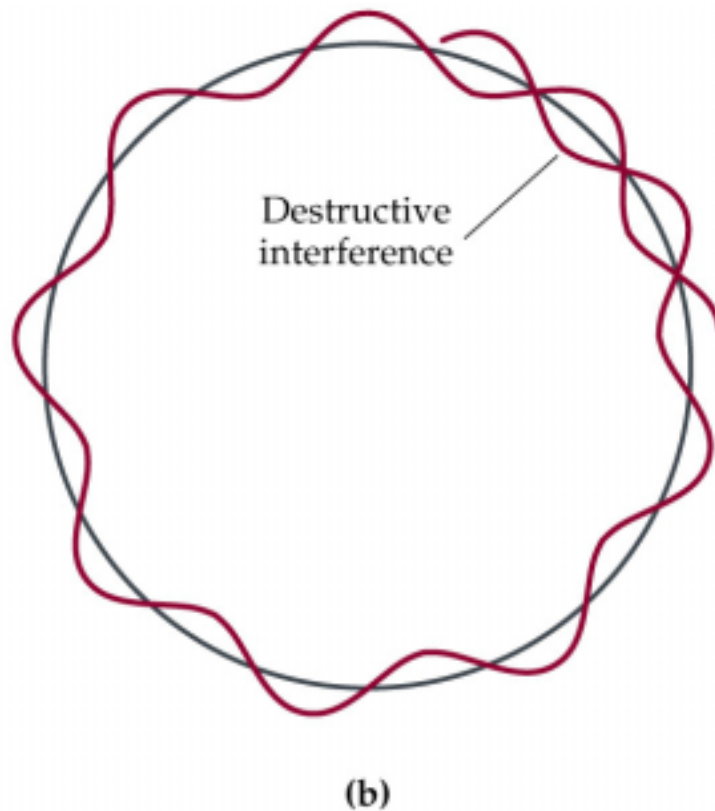
Each series of spectral lines in hydrogen is the result of electrons jumping from an excited state to a particular lower level. For the Lyman series the lower level is the ground state. The lower level for the Balmer is the first excited state ($n = 2$), and the lower level for the Paschen series is the second excited state ($n = 3$).

Fig 31.10a de Broglie wavelengths and Bohr orbits



- Bohr's condition that the angular momentum of an allowed orbit must be an integer n times $h / 2\pi$ is equivalent to the condition that n de Broglie wavelengths must fit into the circumference of an orbit. (a) de Broglie waves for the $n = 7$ and $n = 8$ orbits.

Fig 31.10b de Broglie wavelengths and Bohr orbits



Bohr's condition that the angular momentum of an allowed orbit must be an integer n times $h / 2\pi$ is equivalent to the condition that n de Broglie wavelengths must fit into the circumference of an orbit. (b) If an integral number of wavelengths do not fit the circumference of an orbit, the result is destructive interference.

**** Quantum number

Because the various electron orbits and the associated energies have only discrete values, one uses integer numbers to identify each orbit and its corresponding energy. Such integer number is called a *quantum number*. Quantum mechanics describes the hydrogen atom in terms of four quantum numbers: (1) *the principal quantum number* n , which can have the integer values $n = 1, 2, 3, \dots$; (2) *the orbital quantum number* l , which can have values $l = 0, 1, 2, \dots, (n - 1)$; (3) *the magnetic quantum number* m_l , which can have positive and negative integer values $m_l = -l, \dots, -1, 0, +1, \dots, +l$, and (4) *the spin quantum number* m_s , which can be either $m_s = +\frac{1}{2}$ or $m_s = -\frac{1}{2}$. According to quantum mechanics, an electron does not reside in a circular orbit but, rather, has some probability of being found at various distances from the nucleus.

Figure 31-17 Maximum number of electrons in an energy level

- The maximum number of electrons that can occupy a given energy level in a multielectron atom is $2(2l + 1)$. Occupancy by more than this number of electrons would violate the Pauli exclusion principle.

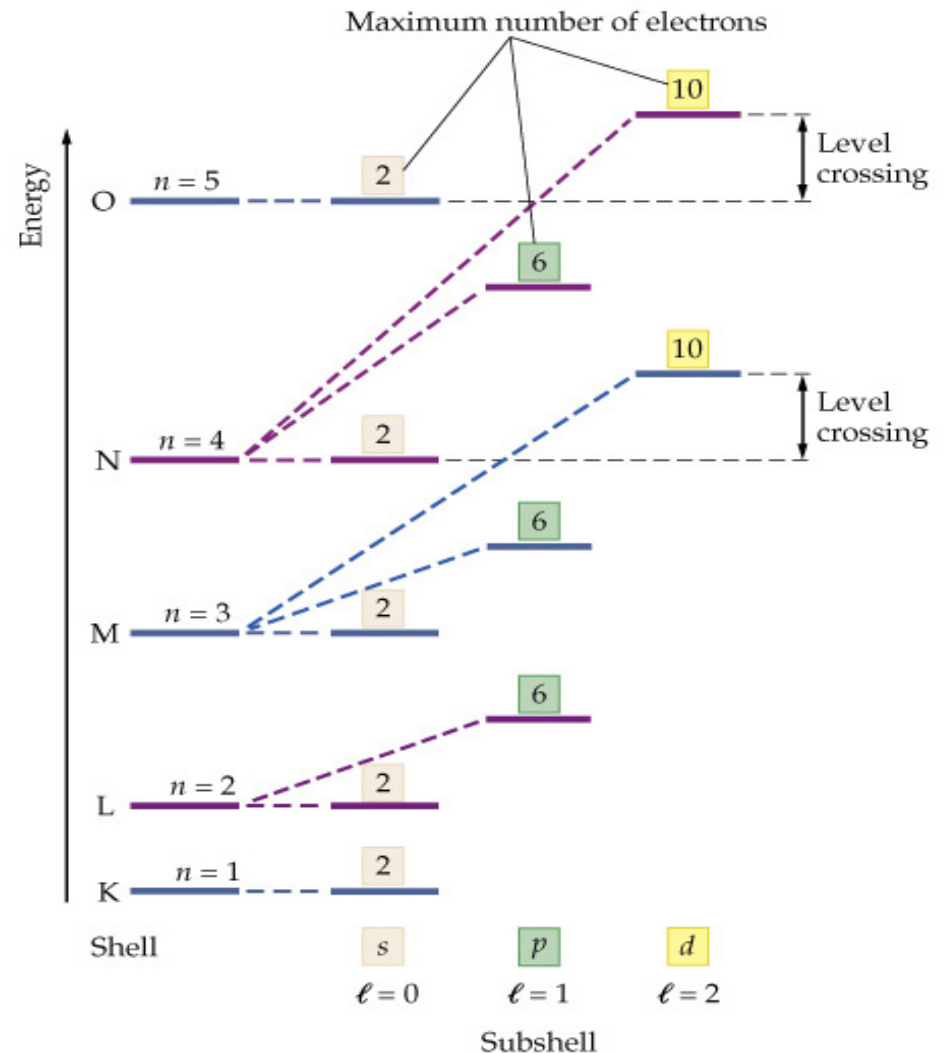
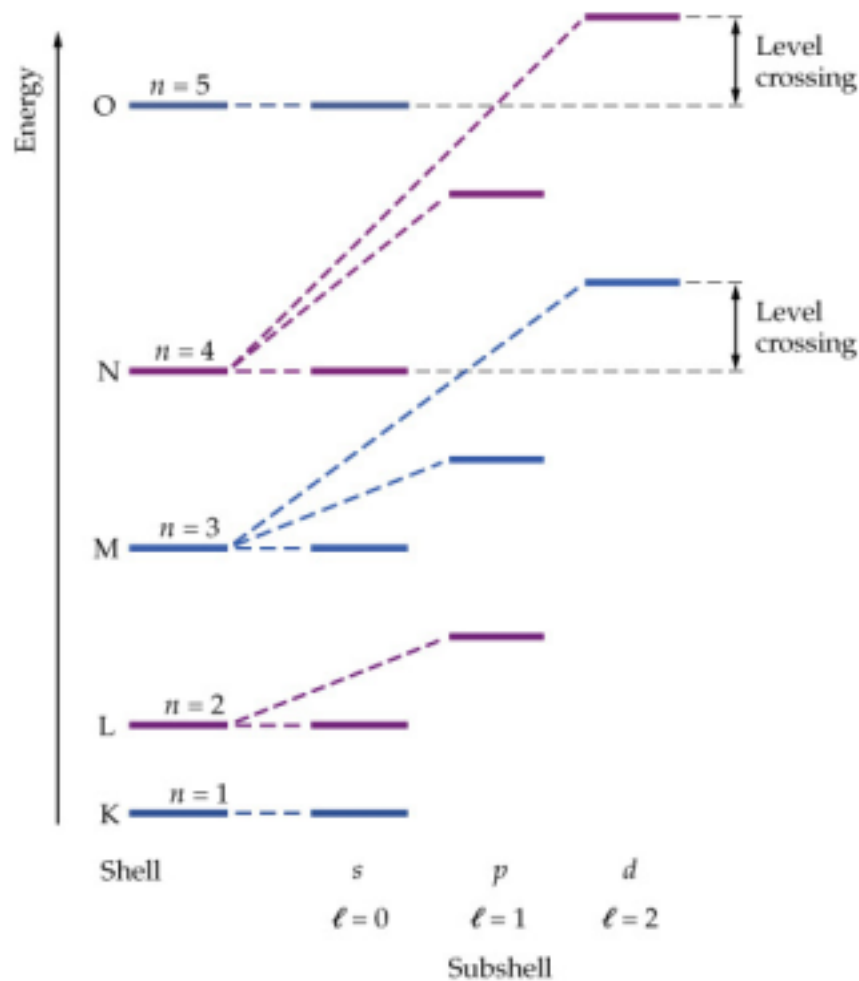
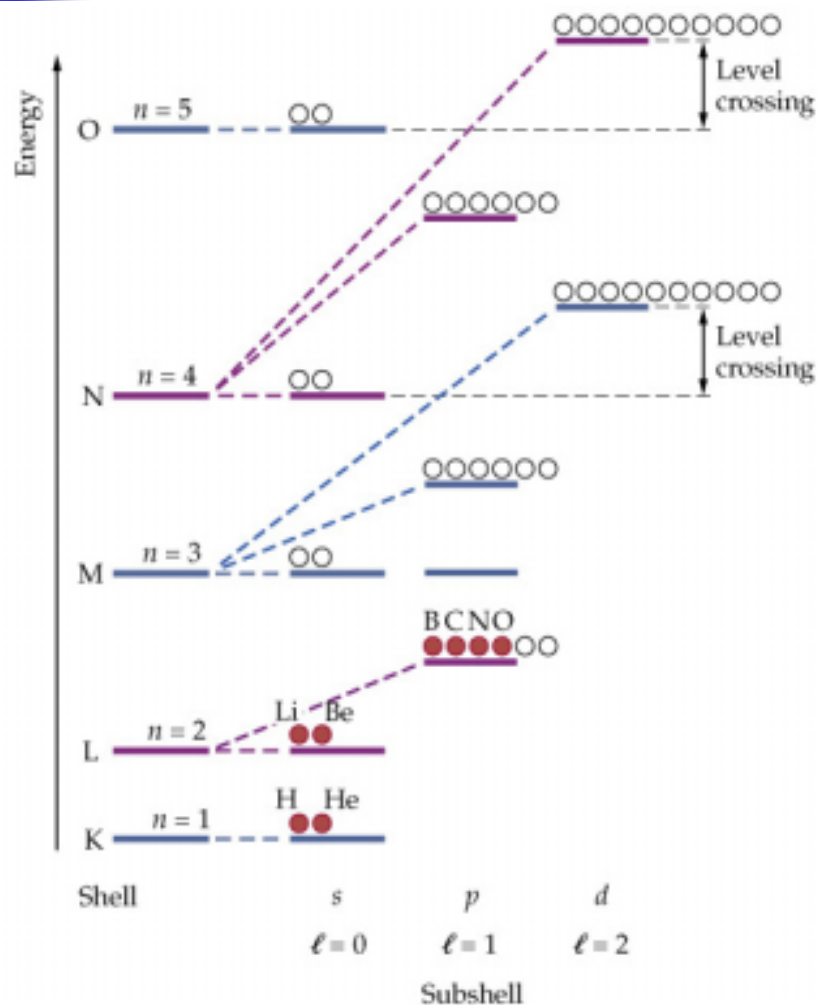


Figure 31-15 Energy levels in multielectron atoms



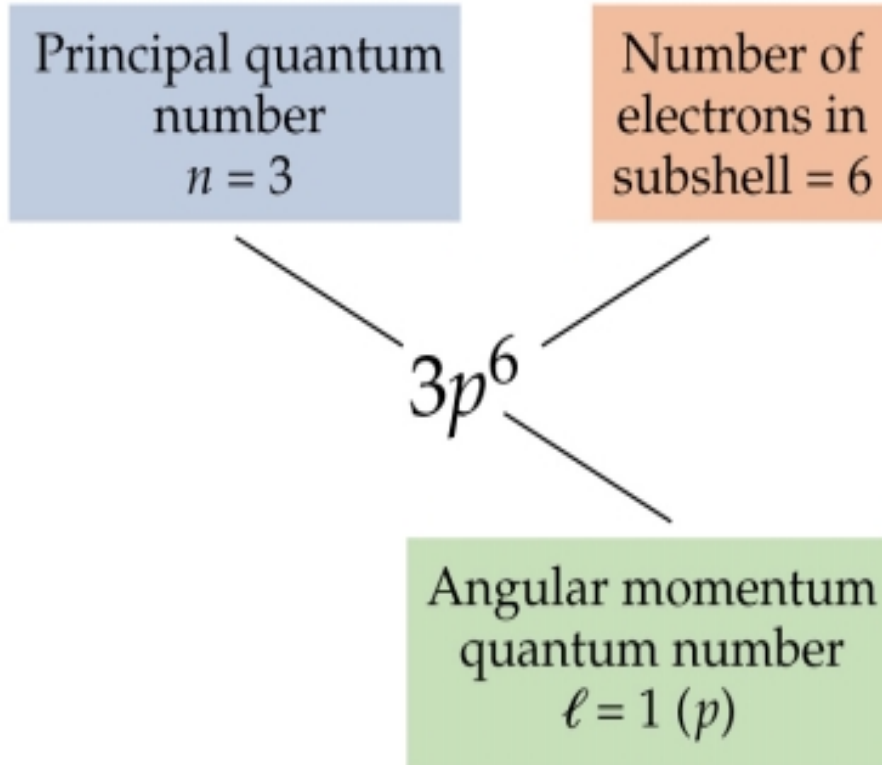
- In multi-electron atoms the energy increases with n for fixed l and increases with l for fixed n . Note the possibility of energy level crossing. For example, the $n = 3, l = 2$ energy level is higher than the $n = 4, l = 0$ energy level.

Figure 31-18 First eight elements of the periodic table



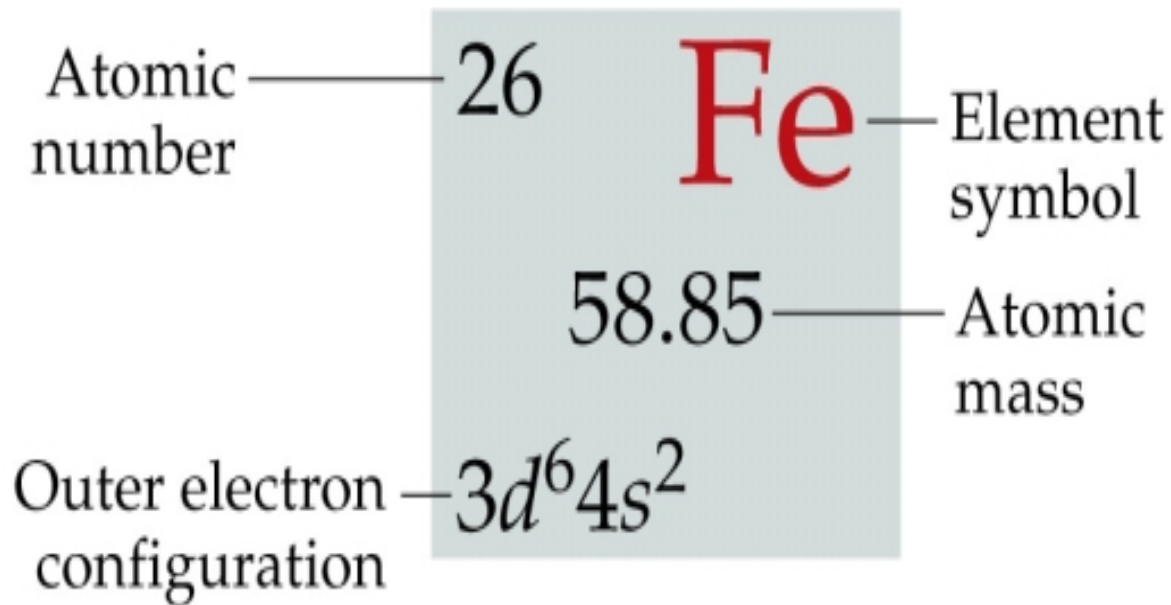
- The elements hydrogen and helium fill the $n = 1, \ell = 0$ level; the elements lithium and beryllium fill the $n = 2, \ell = 0$ level; the elements boron, carbon, nitrogen, and oxygen fill four of the six available states in the $n = 2, \ell = 1$ level.

Figure 31-19 Designation of electronic configuration



- The example shown here indicates six electrons in the $n = 3$, $\ell = 1$ (p) energy level.

Figure 31-20 Designation of elements in the periodic table



- An explanation of the various entries to be found for each element in the periodic table of Appendix D. The example shown here is for the element iron (Fe).