The Language of Physics

Rutherford model of the atom
A planetary model of the atom wherein the negative electron orbits about the positive nucleus in a circular orbit. The orbiting electron is an accelerated charge and should radiate energy. As the electron radiates energy it loses energy and should spiral into the nucleus. Therefore, the Rutherford model of the atom is not correct (p. 966).

Bohr theory of the atom
A revised Rutherford model, wherein the electron can be found only in an orbit for which the angular momentum is quantized in multiples of $\hbar$. The consequence of the quantization postulate is that the electron can be considered as a standing matter wave in the electron orbit. Because standing waves do not transmit energy, the electron does not radiate energy while in its orbit and does not spiral into the nucleus. The Bohr model is thus stable. Bohr then postulated that when the electron jumps from a higher energy orbit to a lower energy orbit, a photon of light is emitted. Thus, the spectral lines of the hydrogen atom should be discrete, agreeing with experimental results. However, the Bohr theory could not explain the spectra from multielectron atoms and it is not, therefore, a completely accurate model of the atom (p. 969).

Quantum mechanical model of the atom
This model arises from the application of the Schrödinger equation to the atom. The model says that the following four quantum numbers are necessary to describe the electron in the atom: (1) the principal quantum number n, which quantizes the energy of the electron; (2) the orbital quantum number l, which quantizes the magnitude of the orbital angular momentum of the electron; (3) the magnetic quantum number m_l, which quantizes the direction of the orbital angular momentum of the electron; and (4) the spin quantum number s, which quantizes the spin angular momentum of the electron (p. 975).

Zeeman effect
When an atom is placed in an external magnetic field a torque acts on the orbital magnetic dipole moment of the atom giving it a potential energy. The energy of the electron depends on the magnetic quantum number as well as the principal quantum number. For a particular value of n, there are multiple values of the energy. Hence, instead of a single spectral line associated with a transition from the nth state to the ground state, there are many spectral lines depending on the value of m_l. Thus, a single spectral line has been split into several spectral lines (p. 982).

Pauli exclusion principle
No two electrons in an atom can exist in the same quantum state. Hence, no two electrons can have the same quantum numbers (p. 988).

Summary of Important Equations

Atomic number
$$Z = \frac{A}{2} \qquad (32.5)$$

Distance of closest approach to nucleus
$$r_0 = \frac{2kZe^2}{\text{KE}} \qquad (32.7)$$

Relative size of atom
$$r_a = 10{,}000\, r_n \qquad (32.8)$$

Radius of nucleus
$$R = R_0 A^{1/3} \qquad (32.9)$$

Bohr Theory of the Hydrogen Atom

Angular momentum is quantized
$$L = mvr = n\hbar \qquad (32.25)$$

Orbital radius
$$r_n = \frac{\hbar^2}{kme^2} n^2 \qquad (32.31)$$
$$r_n = r_1 n^2 \qquad (32.32)$$

Orbital velocity
$$v_n = \frac{ke^2}{n\hbar} \qquad (32.33)$$
$$v_n = \frac{v_1}{n} \qquad (32.34)$$

Electron energy
$$E_n = -\frac{mk^2e^4}{2n^2\hbar^2} \qquad (32.35)$$
$$E_n = -\frac{E_1}{n^2} \qquad (32.37)$$

Einstein's relation
$$h\nu = E_i - E_f \qquad (32.38)$$

Frequency of emitted photon
$$\nu = \frac{E_1}{h}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \qquad (32.39)$$

Wavelength of emitted photon
$$\frac{1}{\lambda} = \frac{E_1}{hc}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \qquad (32.40)$$

Quantum Mechanical Theory of the Hydrogen Atom

Principal quantum number
$$n = 1, 2, 3, \ldots \qquad (32.42)$$

Orbital quantum number
$$l = 0, 1, 2, \ldots, (n-1) \qquad (32.43)$$

Magnetic quantum number
$$m_l = 0, \pm 1, \pm 2, \ldots, \pm l \qquad (32.44)$$

Electron energy
$$E_n = -\frac{k^2e^4m}{2\hbar^2}\frac{1}{n^2} \qquad (32.45)$$

Angular momentum
$$L = \sqrt{l(l+1)}\,\hbar \qquad (32.46)$$

z-component of angular momentum
$$L_z = m_l\hbar \qquad (32.47)$$

Direction of **L**
$$\theta = \cos^{-1}\frac{m_l}{\sqrt{l(l+1)}} \qquad (32.49)$$

Orbital magnetic dipole moment

$$\mu_l = -\frac{e}{2m_e}\mathbf{L} \qquad (32.57)$$

$$\mu_l = \frac{eL}{2m_e}$$

$$= \frac{e}{2m_e}\sqrt{l(l+1)}\,\hbar \qquad (32.58)$$

Potential energy of a dipole in an external magnetic field

$$\text{PE} = -\mu_l B \cos\theta \qquad (32.59)$$

$$\text{PE} = m_l \frac{e\hbar}{2m_e} B \qquad (32.62)$$

Zeeman Effect

Splitting of energy state

$$E' = E + m_l \frac{(e\hbar)B}{2m_e} \qquad (32.66)$$

Splitting of spectral lines

$$\lambda_+ = \frac{c}{\nu_+}$$

$$\lambda_+ = \frac{c}{\nu_0 + e\hbar B/2m_e h} \qquad (32.70)$$

in an external

$$\lambda = \frac{c}{\nu_0} = \frac{ch}{\frac{3}{4}E_1} \qquad (32.65)$$

magnetic field

$$\lambda_- = \frac{c}{\nu_-}$$

$$\lambda_- = \frac{c}{\nu_0 - e\hbar B/2m_e h} \qquad (32.71)$$

Selection rules for transitions

$$\Delta l = \pm 1$$

$$\Delta m_l = 0, \pm 1 \qquad (32.72)$$

Spin quantum number

$$s = \tfrac{1}{2} \qquad (32.73)$$

Spin angular momentum

$$S = \sqrt{s(s+1)}\,\hbar \qquad (32.74)$$

$$S = \tfrac{1}{2}\sqrt{3}\,\hbar \qquad (32.75)$$

z-component of spin

$$S_z = m_s\hbar \qquad (32.76)$$

$$S_z = \pm\frac{\hbar}{2} \qquad (32.77)$$

Spin magnetic dipole moment

$$\mu_s = -2.0024\left(\frac{e}{2m_e}\right)\mathbf{S} \qquad (32.78)$$

Potential energy

$$\text{PE} = -\mu_s B \cos\theta \qquad (32.79)$$

of electron

$$\text{PE} = +2\frac{e}{2m_e}SB\cos\theta \qquad (32.80)$$

due to spin

$$\text{PE} = \pm\frac{e\hbar}{2m_e}B_{so} \qquad (32.81)$$

Spin-orbit splitting

$$E_1 = E_0 + \frac{e\hbar}{2m_e}B_{so} \qquad (32.82)$$

of energy state

$$E_1 = E_0 - \frac{e\hbar}{2m_e}B_{so} \qquad (32.83)$$

Questions for Chapter 32

1. Discuss the differences among (a) the plum pudding model of the atom, (b) the Rutherford model of the atom, (c) the Bohr theory of the atom, and (d) the quantum mechanical theory of the atom.

†2. Discuss the effect of the uncertainty principle and the Bohr theory of electron orbits.

†3. How can you use spectral lines to determine the chemical composition of a substance?

4. If you send white light through a prism and then send it through a tube of hot hydrogen gas, what would you expect the spectrum to look like when it emerges from the hydrogen gas?

†5. In most chemical reactions, why are the outer electrons the ones that get involved in the reaction? Is it possible to get the inner electrons of an atom involved?

6. When an atom emits a photon of light what does this do to the angular momentum of the atom?

7. Explain how the Bohr theory can be used to explain the spectra from singly ionized atoms.

†8. Discuss the process of absorption of light by matter in terms of the atomic structure of the absorbing medium.

†9. Rutherford used the principle of scattering to "see" inside the atom. Is it possible to use the principle of scattering to "see" inside a proton and a neutron?

†10. How can you determine the chemical composition of a star?

†11. How can you determine if a star is approaching or receding from you? How can you determine if it is a small star or a very massive star?

Problems for Chapter 32

Section 32.1 The History of the Atom

1. Find the potential energy of two α particles when they are brought together to a distance of 1.20×10^{-15} m, the approximate size of a nucleus.

2. Find the potential energy of an electron and a proton when they are brought together to a distance of 5.29×10^{-11} m to form a hydrogen atom.

3. A silver nucleus is bombarded with 8-MeV α particles. Find (a) the maximum radius of the silver nucleus and (b) the more probable radius of the silver nucleus.

4. Estimate the radius of a nucleus of $^{238}_{92}U$.

5. How many electrons, protons, and neutrons are there in a silver atom?

6. How many electrons, protons, and neutrons are there in a uranium atom?

7. Find the difference in the orbital radius of an electron in the Rutherford atom if the electron is initially in an orbit of 5.29×10^{-11} m radius and the atom radiates 2.00 eV of energy.

Section 32.2 The Bohr Theory of the Atom

†8. An electron is in the third Bohr orbit. Find (a) the radius, (b) the speed, (c) the energy, and (d) the angular momentum of the electron in this orbit.

9. The orbital electron of a hydrogen atom moves with a speed of 5.459×10^5 m/s. (a) Determine the value of the quantum number n associated with this electron. (b) Find the radius of this orbit. (c) Find the energy of the electron in this orbit.

†10. An electron in the third Bohr orbit drops to the ground state. Find the angular momentum of the electron in (a) the third Bohr orbit, and (b) the ground state. (c) Find the change in the angular momentum of the electron. (d) Where did the angular momentum go?

11. At what temperature will the average thermal speed of a free electron equal the speed of an electron in its second Bohr orbit?

12. The lifetime of an electron in an excited state of an atom is about 10^{-8} s. How many orbits will an electron in the $2p$ state execute before falling back to the ground state?

13. Show that the ratio of the speed of an electron in the first Bohr orbit to the speed of light is equal to 1/137. This ratio is called the *fine-structure constant.*

14. Find the radius of the first Bohr orbit of an electron in a singly ionized helium atom.

15. Find the angular momentum of an electron in the third Bohr orbit and the second Bohr orbit. How much angular momentum is lost when the electron drops from the third orbit to the second orbit?

Section 32.3 The Bohr Theory and Atomic Spectra

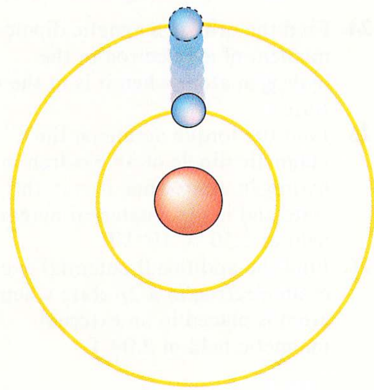
16. An electron in the third Bohr orbit drops to the second Bohr orbit. Find (a) the energy of the photon emitted, (b) its frequency, and (c) its wavelength.

17. An electron in the third Bohr orbit drops to the ground state. Find (a) the energy of the photon emitted, (b) its frequency, and (c) its wavelength.

18. Calculate the wavelength of the first two lines of the Paschen series.

Section 32.4 The Quantum Mechanical Model of the Hydrogen Atom

19. Find the angular momentum of the electron in the quantum mechanical model of the hydrogen atom when it is in the $2p$ state. How much angular momentum is lost when the electron drops to the $1s$ state?



20. Find the angular momentum of the electron in the quantum mechanical model of the hydrogen atom when it is in the $3d$ state. (a) How much angular momentum is lost when the electron drops to the $1s$ state? (b) How much energy is lost when the electron drops to the $1s$ state?

21. An orbital electron is in the $5d$ state. (a) Find the energy of the electron in this state. (b) Find the orbital angular momentum of the electron. (c) Compute all possible values of the z-component of the orbital angular momentum of the electron. (d) Determine the largest value of θ, the angle between the orbital angular momentum vector and the z-axis.

22. Find the angle θ that the angular momentum vector makes with the z-axis when the electron is in the $3d$ state.
23. Find (a) the z-component of the angular momentum of an electron when it is in the $2p$ state and (b) the angle that **L** makes with the z-axis.

Section 32.5 The Magnetic Moment of the Hydrogen Atom

24. Find the orbital magnetic dipole moment of an electron in the hydrogen atom when it is in the $4f$ state.
25. Find the torque acting on the magnetic dipole of an electron in the hydrogen atom when it is in the $3d$ state and is in an external magnetic field of 2.50×10^{-3} T.
26. Find the additional potential energy of an electron in a $2p$ state when the atom is placed in an external magnetic field of 2.00 T.

†27. Find (a) the total energy of an electron in the three $2p$ states when it is placed in an external magnetic field of 2.00 T, (b) the energy of the photons emitted when the electrons fall to the ground state, (c) the frequencies of the spectral lines associated with these transitions, and (d) the wavelengths of their spectral lines.

Section 32.7 Electron Spin

28. Calculate the magnitude of the spin magnetic dipole moment of the electron in a hydrogen atom.
29. Find the potential energy associated with the spin magnetic dipole moment of the electron in a hydrogen atom when it is placed in an external magnetic field of 2.50×10^{-3} T.

Section 32.8 The Pauli Exclusion Principle and the Periodic Table of the Elements

30. How many electrons are necessary to fill the N shell of an atom?
31. Write the electron configuration for the chemical element potassium.
32. Write the electron configuration for the chemical element iron.
33. Enumerate the quantum states (n, l, m_l, m_s) of each of the orbital electrons in the element $^{40}_{20}Ca$.