Chapter Thirteen
MODERN PHYSICS

KEY IDEAS

Energy is not continuous; it occurs in a series of discrete bundles called quanta. Einstein's explanation of the photoelectric effect introduced the concept that light energy is quantized in units called photons. The energy of a photon is directly related to its frequency. Compton's experiments verified the photon nature of light. De Broglie proposed that matter such as electrons possess wavelike characteristics, and diffraction experiments verified this theory.

Rutherford established the nuclear model of the atom by analyzing the results when alpha particles were scattered by metallic foils. Bohr refined the model of hydrogen by proposing that the energy levels in the atom are quantized and that electrons can make only discrete transitions within these levels. The currently accepted model of the atom is based on the idea of probability, and each electron is viewed as a cloud rather than as a specific point in space.

KEY OBJECTIVES

At the conclusion of this chapter you will be able to:

- Define the term *quantum of energy*, and relate this term to Planck's constant.
- Describe the photoelectric effect and Einstein's explanation of it.
- Solve problems using Einstein's photoelectric equation.
- Explain how the Compton effect supports the photon theory of light.
- Calculate the momentum of a photon, given its frequency or wavelength.
- Describe Rutherford's experiments involving the scattering of alpha particles by metallic foils and the model of the atom he proposed as a result of those experiments.
- Explain why the Rutherford model did *not* provide a complete picture of the atom.
- State the hypotheses that Bohr used in developing his model of the hydrogen atom.
- Define the terms *ground state, excited state,* and *stationary state* as they apply to the Bohr model.

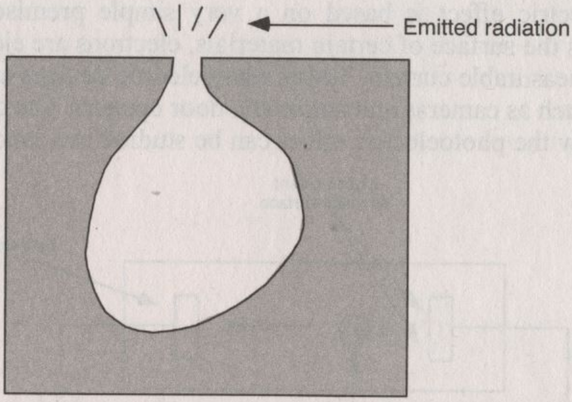
- Describe how Bohr was able to explain the existence of line spectra.
- Define the term *ionization*, and use the energy level diagrams for hydrogen and mercury to calculate the energies involved in various electron transitions.
- Define the term *electron cloud*, and state why the cloud model was needed to provide a more nearly complete picture of the atom.

13.1 INTRODUCTION

The title of this chapter is not completely accurate since the theories and discoveries we describe originated between 1895 and 1930. In this time period, two major scientific theories were advanced: quantum physics and Einstein's theory of relativity. Although both theories revolutionized the sciences and technology, they presented various aspects of nature in ways that defied the familiar, "commonsense" view of the world.

🍎 13.2 BLACK-BODY RADIATION AND PLANCK'S HYPOTHESIS

When a solid is heated, it emits a variety of electromagnetic radiation. As the temperature of the solid is increased, the radiation shifts toward shorter wavelengths, a change that explains why heated solids begin to glow red, then orange, and finally white. To study this radiation effectively, physicists used a device called a *cavity radiator*, also known as an *ideal black body*. This device is a hollow solid with a small opening drilled in one of the walls, as shown in the following diagram.



When the solid was heated, it was found that that the radiation emitted *through the opening* depended only on the temperature of the solid, not on

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the material of which the cavity radiator was made. The origin of the radiation was presumed to be the atoms of the solid. The heating caused the atoms to oscillate (just as springs do), and then the energy was released into the cavity as electromagnetic radiation. However, there was serious disagreement between the theory and the experimental results.

In 1900, German physicist Max Planck reported a startling discovery: the experimental results could be explained precisely if it was assumed that the atomic "oscillators" can have only certain energies given by this equation:

PHYSICS CONCEPTS

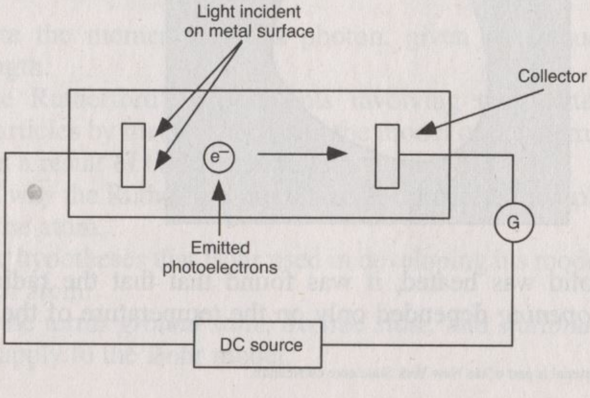
$$E = nhf$$

where E is the energy of the atomic oscillator, f is its frequency of oscillation, n is an integer (1, 2, 3, . . .) and h, known as *Planck's constant*, has the value of 6.6×10^{-34} joule · second. Each of these discrete values of energy is known as a *quantum* of energy. As a result, the radiation emitted from the cavity opening is also restricted to certain values.

This idea was revolutionary because physicists had always assumed that a particle could take on or emit any value of energy. (In fact, Planck himself was not comfortable with the concept of quantized energy!) As we will see in the next section, Planck's idea was reinforced in 1905 by German-American physicist Albert Einstein, who used it to explain the nature of light in an experiment involving the *photoelectric effect.*

13.3 THE PHOTOELECTRIC EFFECT

The photoelectric effect is based on a very simple premise: When light energy strikes the surface of certain materials, electrons are ejected, creating a small but measurable current. Today, photoelectric devices are in wide use in products such as cameras and automatic door openers. The diagram below illustrates how the photoelectric effect can be studied in a laboratory.



As monochromatic light strikes a metal surface, a potential difference causes the ejected photoelectrons (e^-) to move onto a collector. The galvanometer (G) measures the current that is produced as a result. The arrangement illustrated above allows for changes in the (1) metal surface, (2) potential difference, (3) light intensity, and (4) color of the light used.

Such photoelectric experiments yielded results that did not agree with the idea that light behaves as a wave. For example, certain colors of light did not eject electrons, regardless of the intensity of the incident light. According to wave theory, very bright light (of any color) should possess a great deal of energy and be able to eject electrons from a metal surface. It was found that, when electrons are ejected, the *color* of the light determines how energetic the electrons are; the intensity determines the *rate* at which the electrons are ejected (i.e., the current).

The potential difference can also be made to *oppose* the movement of the electrons. There is a value, known as the *stopping potential*, that will stop even the most energetic electrons from reaching the collector. If we measure the value of the stopping potential (V_0), we can calculate the kinetic energy of the fastest electrons (KE_{max}). Since energy (work) is equal to potential difference multiplied by charge, it follows that:

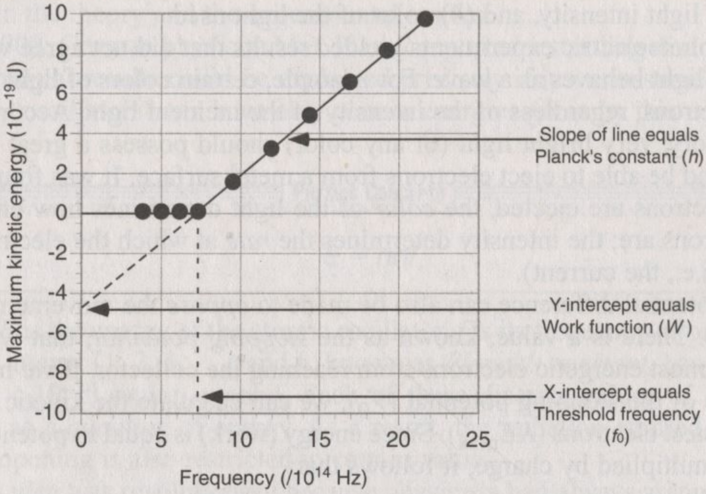
PHYSICS CONCEPTS

$$KE_{max} = V_o e$$

PROBLEM

In a photoelectric experiment, the stopping potential is 5.0 volts. Calculate the maximum kinetic energy of the photoelectrons.

The graph below represents an experiment that relates the color (i.e., the frequency) of the incident light to the maximum kinetic energy of the photoelectrons.



For any photoemissive material tested, the graph is a straight line whose slope is Planck's constant. There is a frequency, called the *threshold frequency* (f_0) at which the maximum kinetic energy of the photoelectrons is zero. The energy of a photon at this frequency is just sufficient to break the bonds holding the electron to the surface of the material. This energy, known as the *work* function (W), is the y-intercept of the graph. At frequencies *less* than the threshold value, photons cannot produce a photoelectric effect.

Einstein predicted these results before there was any experimental verification. He theorized that light behaved as if it were a collection of particles called *photons*. The energy of each photon depends on the frequency of the light. The energy is given by the following equation:

PHYSICS CONCEPTS

$$E = h_f \text{ or substituting } \frac{c}{\lambda} \text{ for } f$$

$$E = \frac{h_f}{\lambda}$$

Einstein summarized his theory by means of his *photoelectric equation*:

PHYSICS CONCEPTS

$$KE_{max} = hf - W$$

When a photon with energy hf strikes a surface, a part of its energy (the work function, W) frees the electron from its bonds. The remainder of the energy $(hf - W)$ gives the electron its kinetic energy (KE_{max}).

PROBLEM

A photon with a frequency of 8.0×10^{14} hertz strikes a photoemissive surface whose work function is 1.7×10^{-19} joule. Calculate (a) the maximum kinetic energy of the ejected photoelectrons and (b) the threshold frequency of the surface.

SOLUTION

(a) $KE_{max} = hf - W$

$= (6.6 \times 10^{-34}$ [illegible] Hz) $- 1.7 \times 10^{-19}$ J

$= 3.6 \times 10^{-19}$ [illegible]

(b) The threshold frequency [illegible] ency at which the maximum kinetic energy of the photo[illegible] n. Applying the photoelectric equation gives

$KE_{max} = hf - W$

$0 = hf_0 - W$

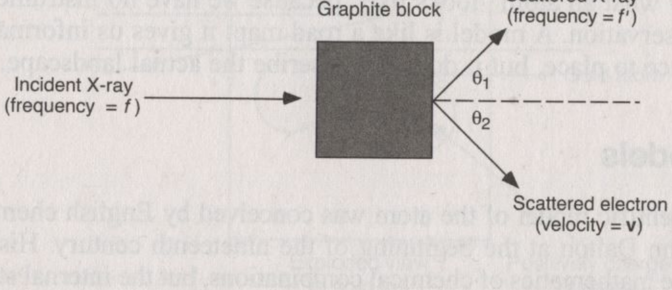
$W = hf_0$

1.7×10^{-19} J $= (6.6 \times 10^{-34}$ [illegible]

$f_0 = 2.6 \times 10^{14}$ H[illegible]

13.4 THE COMPTON EFFECT

The photon theory of light was strongly supported by the experiments of Arthur Compton, a U.S. physicist. When Compton bombarded a block of graphite with X-rays of known frequency, he discovered that both electrons and X-rays emerged from the block, as shown in the diagram.



Compton observed that the scattered X-rays had lower frequencies than the incident X-rays, and he recognized that both energy and momentum were conserved in these collisions. Compton used the following relationship for the magnitude of the momentum of a photon:

PHYSICS CONCEPTS

$$p = \frac{E}{c} = \frac{hf}{c} = \frac{h}{\lambda}$$

PROBLEM
Calculate the momentum of an X-ray photon whose wavelength is 1.0×10^{-10} meter.

SOLUTION

$$p = \frac{h}{\lambda} = \frac{6.6 \times 10^{-34}\ \text{J·s}}{1.0 \times 10^{-10}\ \text{m}} = 6.6 \text{ [illegible]}$$

This relationship demonstrates t[illegible] (p), the particle property of light, wavelength (λ), and the w[illegible] of light, are inversely related: High-frequency light (ultraviolet, [illegible] radiation) behaves more like particles and less like waves; lo[illegible] (radio waves, microwaves, infrared radiation) behaves more [illegible] less like particles.

13.5 MODELS OF THE ATOM

The impact of modern physics is most evident in the development of the atomic model of matter. Though the concept of the atom goes back to ancient Greece and Rome, it was not until the twentieth century that the structure of the atom was reasonably well understood. We use the term *atomic model* to indicate that we are trying to describe the key features of the atom. We really do not know what an atom "looks like" because we have no instruments for its direct observation. A model is like a road map: it gives us information to get from place to place, but it does *not* describe the actual landscape.

Early Models

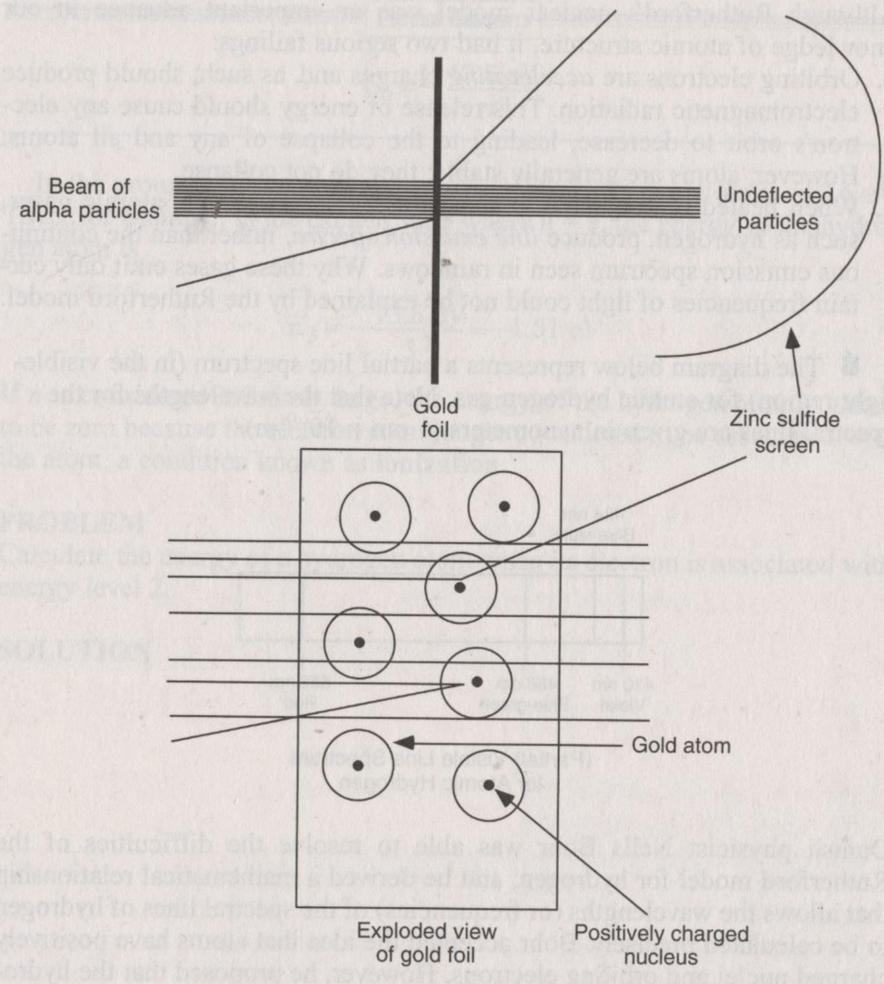
The first scientific model of the atom was conceived by English chemist and physicist John Dalton at the beginning of the nineteenth century. His model explained the mathematics of chemical combinations, but the internal structure of the atom remained a mystery. At the beginning of the twentieth century, the discoveries of Antoine Henri Becquerel, the Curies, Frédéric Joliot-Curie, Iréne Joliot-Curie, and J.J. Thomson led to the idea that an atom is constructed of positively and negatively charged particles, most notably the electron.

Rutherford Model

British physicist Ernest Rutherford and his two assistants, Hans Geiger and Ernest Marsden, bombarded thin metallic foils, such as gold, with massive positively charged particles known as *alpha particles*. They counted scintillations of scattered alpha particles on a zinc sulfide screen. They observed the following:

1. Most of the particles passed through the foils without being deflected.
2. A very small number of particles were deflected through large angles, some approaching 180°.

The diagrams below represent the results of Rutherford's experiments.



On the basis of these experiments, Rutherford drew the following conclusions:

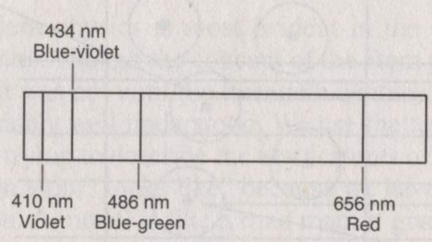
1. Most of the atom is empty space.
2. Most of the mass of the atom is concentrated in a dense, positively charged *nucleus*. The negative electrons orbit the nucleus much as the planets orbit the Sun.
3. According to calculations using Coulomb's law, the interaction of alpha particles with a very small, massive nucleus would result in it following a hyberbolic path and produce the observed angular pattern of scattering.

Bohr Model

Although Rutherford's nuclear model was an important advance in our knowledge of atomic structure, it had two serious failings:

1. Orbiting electrons are *accelerating* charges and, as such, should produce electromagnetic radiation. This release of energy should cause any electron's orbit to decrease, leading to the collapse of any and all atoms. However, atoms are generally stable; they do not collapse.
2. When heated or subjected to a high potential difference, atomic gases, such as hydrogen, produce *line emission spectra,* rather than the continuous emission spectrum seen in rainbows. Why these gases emit only certain frequencies of light could not be explained by the Rutherford model.

The diagram below represents a partial line spectrum (in the visible-light region) for atomic hydrogen gas. Note that the wavelengths for the spectral lines are given in nanometers (1 nm = 10^{-9} m).



(Partial) Visible Line Spectrum
for Atomic Hydrogen

Danish physicist Neils Bohr was able to resolve the difficulties of the Rutherford model for hydrogen, and he derived a mathematical relationship that allows the wavelengths (or frequencies) of the spectral lines of hydrogen to be calculated precisely. Bohr accepted the idea that atoms have positively charged nuclei and orbiting electrons. However, he proposed that the hydrogen atom possesses distinct orbits. At any instant, the single electron of hydrogen can be associated with one and only one of these orbits. Each orbit gives the hydrogen atom a specific amount of energy. (For this reason, Bohr called the orbits *energy levels.*) Energy levels are identified by integers: 1, 2, 3, . . . , *n*. Today these integers are known as *quantum numbers*.

If the electron is on the first energy level, the atom is said to be in the **ground state**. On any other level, the atom is in an **excited state**. No energy is either emitted or absorbed by the atom when it is in any single state, known as a *stationary state*. Rather, energy is *emitted* (as photons) when the atom goes from a higher energy to a lower energy state, and energy is *absorbed* (as photons) when the atom goes from a lower energy to a higher energy state.

Bohr's relationship for the energy levels in a hydrogen atom is as follows:

PHYSICS CONCEPTS

$$E_n = -\frac{13.6 \text{ eV}}{n^2}$$

In the ground state ($n = 1$), the energy of the hydrogen atom is –13.6 eV. Then, for example, in the excited state where $n = 3$, the energy of the hydrogen atom is

$$E_3 = -\frac{13.6 \text{ eV}}{3^2} = -1.51 \text{ eV}$$

If *n* is considered infinitely large, the energy of the hydrogen atom is taken to be zero because the electron is no longer considered to be associated with the atom, a condition known as **ionization**.

PROBLEM

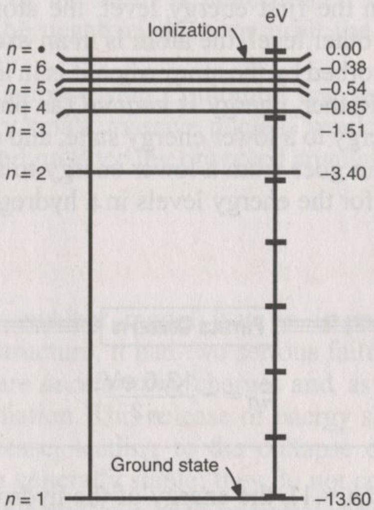
Calculate the energy of a hydrogen atom when its electron is associated with energy level 2.

SOLUTION

$$E_n = -\frac{13.6 \text{ eV}}{n^2}$$

$$E_2 = -\frac{13.6 \text{ eV}}{2^2} = -3.40 \text{ eV}$$

The diagram below is a graphic representation of the energy levels of the hydrogen atom:



Energy Levels of the Hydrogen Atom

When a hydrogen atom changes from energy state E_{initial} to energy state E_{final}, the amount of energy emitted or absorbed by the atom is determined as follows:

PHYSICS CONCEPTS

$$\Delta E = E_{\text{final}} - E_{\text{initial}} = \frac{-13.6\ \text{eV}}{n^2_{\text{final}}} - \frac{-13.6\ \text{eV}}{n^2_{\text{initial}}}$$

which can be simplied to

$$\Delta E = 13.6\ \text{eV}\left(\frac{1}{n^2_{\text{initial}}} - \frac{1}{n^2_{\text{final}}}\right)$$

If ΔE is a *negative* number, the energy is emitted; if it is positive, the energy is absorbed. The frequency of the photon that is emitted (or absorbed) can be calculated from the familiar relationship $\Delta E = hf$.

PROBLEM

Calculate the energy of the photon that is emitted when a hydrogen atom changes from energy state $n = 3$ to $n = 2$.

SOLUTION

We have already calculated the … $E_3 = -1.51$ eV, $E_2 = -3.40$ eV

$$\Delta E = E_{final} - \ldots E_3$$

$$= (-3.40 \ldots \text{eV}) = -1.89 \text{ eV}$$

A 1.89-eV photon is emitted … atom in changing from $n = 3$ to $n = 2$. This corresponds to … the hydrogen spectrum shown above.

The visible line spectrum … rogen is known as the *Balmer series*. It consists of electron t… ergy level 2 from higher levels (3, 4, 5, . . .). Transitions dow… l 1 yield a line spectrum in the ultraviolet region, while transi… 4, or 5 yield line spectra in the infrared region.

PROBLEM

How much energy is needed to ionize a hydrogen atom in the ground state?

SOLUTION

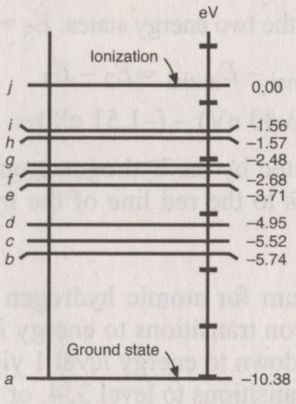
We must raise the energy level … atom from level 1 to "infinity."

$$\Delta E = E_\infty - E_1 = \ldots \text{eV}) = +13.6 \text{ eV}$$

Thus, 13.6 eV must be absorb… atom in the ground state in order to ionize it. This energy … *ionization energy* (or *ionization potential*) of the hydroge…

Unfortunately, the Bohr mo… ssful in explaining atoms that have more than one electron.

The diagram on the le… presents a portion of the line spectrum of mercury. Notice … of the lines is more complex and less regular than for hydro… model *cannot* be applied to this kind of atom. Nevertheless, w… late the energy of an electron transition between states bec… ere given the energy values (labeled *a* through *j*) associate…



A Few Energy Levels for the Mercury Atom

The equation we use in this case is simplified to the following:

PHYSICS CONCEPTS

$$E = E_i - E_f$$

Cloud Model

The current model of the atom is able to explain the structure of all of the atoms in the Periodic Table. This model is based on a mathematical area of physics known as *quantum mechanics* or *wave mechanics*. Quantum mechanics does not place electrons in specific orbits; rather, it indicates the *probability* that an electron will be in a region of space near the nucleus. The most probable regions of the electron's location define what is known as the **electron cloud**. The energy levels of the Bohr model are subdivided into other levels termed *sublevels* and *orbitals*. Together, these define the energy and structure of a particular atom.

13.6 ENERGY AND MASS

The term mc^2 is the total energy of an object (most familiarly stated as $E = mc^2$); it is the sum of the object's kinetic energy and rest energy. We state, without showing the mathematics, that at speeds that are low compared to the speed of light the kinetic energy of an object reduces to the familiar form:

$$KE = \frac{1}{2} mv^2$$