TOPIC 6 Modern Physics

Vocabulary

absorption spectrum	excited state	quantized
antimatter	ground state	quantum
antiparticle	hadron	quantum theory
antiquark	ionization potential	quark
atom	lepton	spectral line
atomic spectrum	meson	Standard Model of Particle Physics
baryon	neutrino	stationary state
bright-line spectrum	nucleus	strong nuclear force
emission spectrum	photon	universal mass unit
energy level	Planck's constant	
energy-level diagram	positron	

Wave-Particle Duality of Energy and Matter

Light, a form of electromagnetic radiation, can be represented as a wave propagated by an interchange of energy between periodically varying electric and magnetic fields. Waves of electromagnetic energy are identified by their frequency, wavelength, amplitude, and velocity. In addition, electromagnetic radiation exhibits the phenomena of diffraction, interference, and the Doppler effect, which are readily explained by a wave model of light.

Waves Have a Particle Nature

The wave model of light, however, cannot explain other phenomena such as interactions of light with matter. In these interactions, light—or other electromagnetic radiation—acts as if it is composed of particles possessing kinetic energy and momentum. For example, when light strikes matter, some of the light's momentum is transferred to the matter. In the latter part of the nineteenth century it was discovered that light having a frequency above some minimum value and incident on certain metals caused electrons to be emitted from the metal. This phenomenon, called the photoelectric effect, could not be explained by a wave model of light. Albert Einstein explained the phenomenon using quantum theory developed by Max Planck.

Quantum Theory

Quantum theory assumes that electromagnetic energy is emitted from and absorbed by matter in discrete amounts or packets. Each packet of electromagnetic energy emitted or absorbed is called a **quantum** (plural, quanta) of energy. The amount of energy E of each quantum is directly proportional to the frequency f of the electromagnetic radiation. The proportionality constant between the energy of a quantum and its frequency is called **Planck's constant,** h. Thus, the energy of a quantum is given by this formula.

$$E = hf$$

The energy E is in joules, the frequency f is in hertz, and Planck's constant h is a universal constant equal to 6.63×10^{-34} joule · second (J · s). The small energy values of quanta are often expressed in electronvolts, eV (1 eV = 1.60×10^{-19} J).

The quantum, or basic unit, of electromagnetic energy is called a **photon.** Although a photon is a massless particle of light, it carries both energy and momentum. The energy of a photon can be found using the previous equation. For light in a vacuum, $f = c/\lambda$ the energy of a photon can also be described in this way.

$$E_{photon} = hf = \frac{hc}{\lambda}$$ R

The equation states that the energy of a photon is directly proportional to its frequency and inversely proportional to its wavelength.

SAMPLE PROBLEM

The energy of a photon is 2.11 electronvolts.

(a) Determine the energy of the photon in joules.
(b) Calculate the frequency of the photon.
(c) Identify the color of light associated with the photon.

SOLUTION: Identify the known and unknown values.

Known	*Unknown*
$E = 2.11$ eV	$E = ?$ J
$h = 6.63 \times 10^{-34}$ J · s	$f = ?$ Hz
1 eV = 1.60×10^{-19} J	color = ?

1. Convert electronvolts to joules using the relationship 1 eV = 1.60×10^{-19} J.

$$2.11\ \text{eV}\left(\frac{1.60 \times 10^{-19}\ \text{J}}{1\ \text{eV}}\right) = 3.38 \times 10^{-19}\ \text{J}$$

2. Solve the formula $E_{photon} = hf$ for frequency f.

$$f = \frac{E_{photon}}{h}$$

3. Substitute the known values and solve.

$$f = \frac{3.38 \times 10^{-19}\ \text{J}}{6.63 \times 10^{-34}\ \text{J} \cdot \text{s}} = 5.10 \times 10^{14}\ \text{Hz}$$

4. According to the electromagnetic spectrum chart found in the *Reference Tables for Physical Setting/Physics*, a frequency of 5.10×10^{14} Hz corresponds to yellow light.

Photon-Particle Collisions

The photoelectric effect demonstrates that when a photon in the visible light range is incident on a metal surface, the photon's energy is completely absorbed and transferred to the emitted electron. However, when X-ray photons, which have much higher frequencies and energies than photons of visible light, strike a metal surface, not only are electrons ejected but electromagnetic radiation of lower frequency is also given off.

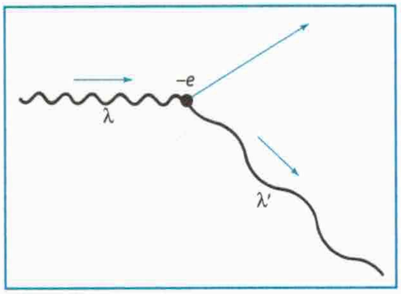


Figure 6-1. A collision of an X-ray photon and an electron in an atom: Besides the electron ejected from the atom, a photon of lower energy (longer wavelength) is also emitted (scattered) by the atom. The energy transferred to the electron equals the difference in energy between the incident photon and the scattered photon. The vector sum of the momentum of the electron and the scattered photon also equals the momentum of the incident photon.

When an X-ray photon and an electron collide, some of the energy of the photon is transferred to the electron and the photon recoils with less energy. Less energy means that the photon has lower frequency. Figure 6-1 illustrates this phenomenon.

Both energy, a scalar quantity, and momentum, a vector quantity, are conserved in this interaction, just as they are in collisions between particles. The incident photon loses energy and momentum, while the electron gains energy and momentum. Photons in a vacuum always travel at the speed of light. Thus, the momentum of a photon depends only on its wavelength or frequency.

Particles Have a Wave Nature

Just as radiation has both wave and particle characteristics, matter in motion has wave as well as particle characteristics. The wavelengths of the waves associated with the motion of ordinary objects, such as a thrown baseball, are too small to be detected. But the waves associated with the motion of particles of atomic or subatomic size, such as electrons, can produce diffraction and interference patterns that can be observed. Diffraction and interference phenomena provide evidence for the wave nature of particles.

Review Questions

1. In which part of the electromagnetic spectrum does a photon have the least energy?

(1) gamma rays
(2) microwaves
(3) visible light
(4) ultraviolet

2. The energy of a photon varies inversely with its

(1) frequency
(2) momentum
(3) speed
(4) wavelength

3. Compared to the frequency and wavelength of a photon of red light, a photon of blue light has a

(1) lower frequency and shorter wavelength
(2) lower frequency and longer wavelength
(3) higher frequency and shorter wavelength
(4) higher frequency and longer wavelength

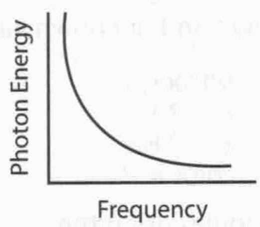
4. A photon has an energy of 8.0×10^{-19} joule. What is this energy expressed in electronvolts?

(1) 5.0×10^{-38} eV
(2) 1.6×10^{-19} eV
(3) 8.0×10^{-19} eV
(4) 5.0 eV

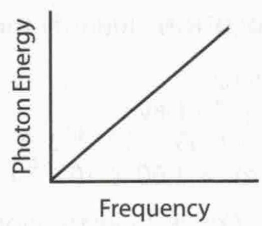
5. The slope of a graph of photon energy versus photon frequency represents

(1) Planck's constant
(2) the mass of a photon
(3) the speed of light
(4) the speed of light squared

6. Which graph best represents the relationship between photon energy and photon frequency?



(1)



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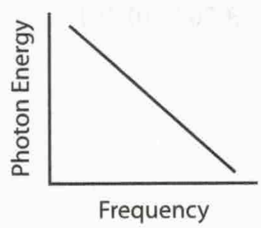
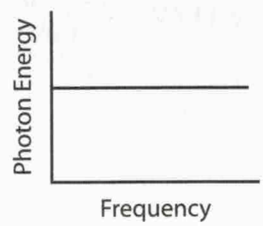


(2)



(4)

7. A photon of green light has a frequency of 6.0×10^{14} hertz. The energy associated with this photon is

(1) 1.1×10^{-48} J
(2) 6.0×10^{-34} J
(3) 5.0×10^{-7} J
(4) 4.0×10^{-19} J

8. Calculate the energy of a photon having a wavelength of 4.00×10^{-7} meter in air.

9. An X-ray photon collides with an electron in an atom, ejecting the electron and emitting another photon. During the collision there is a conservation of
   (1) momentum only
   (2) energy only
   (3) both momentum and energy
   (4) neither momentum nor energy

10. Experiments performed with light indicate that light exhibits
   (1) particle properties only
   (2) wave properties only
   (3) both particle and wave properties
   (4) neither wave nor particle properties

11. A photon of light carries
   (1) energy, but not momentum
   (2) momentum, but not energy
   (3) both energy and momentum
   (4) neither energy nor momentum

Early Models of the Atom

An **atom** is the smallest particle of an element that retains the characteristics of the element. Models for the structure of the atom have evolved over centuries as scientists have developed more sophisticated methods and equipment for studying particles that are too small to be detected by the unaided eye.

Thomson's Model

Over 100 years ago, J. J. Thomson discovered that electrons are relatively low-mass, negatively charged particles present in atoms. Because he knew that atoms are electrically neutral, Thomson concluded that part of the atom must possess a positive charge equal to the total charge of the atom's electrons. Thomson proposed a model in which the atom consists of a uniform distribution of positive charge in which electrons are embedded, like raisins in plum pudding.

Rutherford's Model

Less than two decades later, Ernest Rutherford proposed a different model of the atom. He performed experiments in which he directed a beam of massive, positively charged particles, traveling at approximately 5% of the speed of light, at extremely thin gold foil. Rutherford postulated that if an atom was like those described in Thomson's model, there would be only small net Coulomb forces on a positively charged particle as it passed through or near a gold atom in the foil, and the particle would pass through the foil relatively unaffected. However, he found that, although nearly all the positively charged particles were not deflected from a straight-line path through the gold foil, a small number of particles were scattered at large angles.

To explain the large angles of deflection of those few particles, Rutherford theorized that the massive, energetic, positively charged particles must have collided with other even more massive positively charged particles. Assuming that atoms are symmetrical, he concluded that this concentration of mass and positive charge in the atom, which he called the nucleus, is located at the atom's center. From the relative number of deflected particles, he calculated that the nucleus is only about $\frac{1}{10,000}$ the diameter of the average atom.

Based on the results of these scattering experiments, Rutherford described an atom as being similar to a miniature solar system. The tiny nucleus at the center of the atom contains all the positive charge of the atom and virtually all of its mass. The nucleus is surrounded by enough electrons to balance the positive charge of the nucleus and make the atom electrically neutral. The electrons move in orbits around the nucleus and are held in orbit by Coulomb forces of attraction between their negative charges and the positive charge of the nucleus.

In Rutherford's model, the electrons orbiting the nucleus accelerate due to a change in direction of motion. Rutherford knew that these accelerated charges should radiate electromagnetic energy, lose kinetic energy and momentum in the process, and spiral rapidly to the nucleus. The radiated electromagnetic energy would increase in frequency and produce a continuous spectrum. This expected behavior is contradicted by the observed bright-line spectrum that is characteristic of each element. (Bright-line spectra will be discussed later in this topic.)

The Bohr Model of the Hydrogen Atom

About two years later, Niels Bohr attempted to explain why electrons in atoms can maintain their positions outside the nucleus rather than spiral into the nucleus and cause the atom to collapse. Bohr developed a model of the hydrogen atom based on these assumptions:

- All forms of energy are **quantized,** that is, an electron can gain or lose kinetic energy only in fixed amounts, or quanta.
- The electron in the hydrogen atom can occupy only certain specific orbits of fixed radius and no others.
- The electron can jump from one orbit to a higher one by absorbing a quantum of energy in the form of a photon.
- Each allowed orbit in the atom corresponds to a specific amount of energy. The orbit nearest the nucleus represents the smallest amount of energy that the electron can have. The electron can remain in this orbit without losing energy even though it is being constantly accelerated toward the nucleus by the Coulomb force of attraction.

When the electron is in any particular orbit, it is said to be in a **stationary state.** Each stationary state represents a specific amount of energy and is called an **energy level.** The successive energy levels of the hydrogen atom are assigned integral numbers, denoted by $n = 1$, $n = 2$, etc. When the electron is in the lowest energy level ($n = 1$), the atom is said to be in the **ground state.** When the electron is in any level above $n = 1$, the atom is said to be in an **excited state.**

Energy Levels Any process that raises the energy level of electrons in an atom is called <u>excitation</u>. Excitation can be the result of absorbing the energy of colliding particles of matter, such as electrons, or of photons of electromagnetic radiation. A photon's energy is absorbed by an electron in an atom only if the photon's energy corresponds exactly to an energy-level difference possible for the electron. Excitation energies are different for different elements.

Atoms rapidly lose the energy of their various excited states as their electrons return to the ground state. This lost energy is in the form of photons (radiation) of specific frequencies, which appear as spectral lines in the characteristic spectrum of each element. A **spectral line** is a particular frequency of absorbed or emitted energy characteristic of an atom.

Ionization Potential An atom can absorb sufficient energy to raise an electron to an energy level such that the electron is essentially removed from the atom and an ion is formed. The energy required to remove an electron from an atom to form an ion is called the atom's **ionization potential.** An atom in an excited state requires a smaller amount of energy to become an ion than does an atom in the ground state.

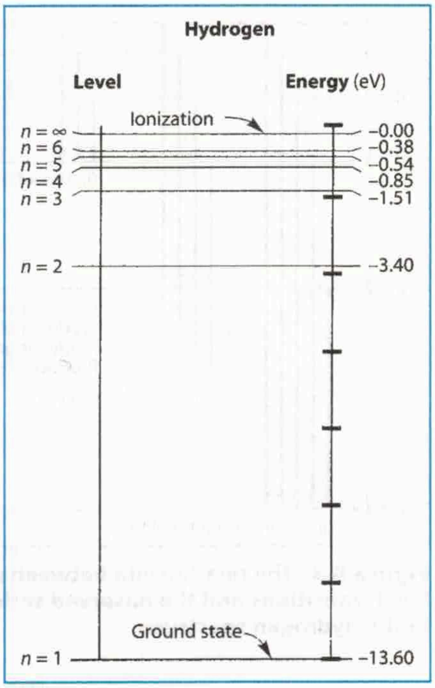


Figure 6-2. Energy levels for the hydrogen atom

Figure 6-2 shows the energy-level diagram for the hydrogen atom. An **energy-level diagram** is one in which the energy levels of a quantized system are indicated by distances of horizontal lines from a zero energy level. The energy level of an electron that has been completely removed from the atom ($n = \infty$) is defined to be 0.00 eV. Thus, all other energy levels have negative values. As an electron moves closer to the nucleus, the energy associated with the electron becomes smaller. Because an electron in the ground state has the lowest energy, its energy has the largest negative value. The *Physics Reference Tables for Physical Setting/Physics* contain energy level diagrams for hydrogen and mercury.

Limitations of Bohr's Model Although Bohr's model explained the spectral lines of hydrogen, it could not predict the spectra or explain the electron orbits of elements having many electrons. Nevertheless, Bohr's model with its quantized energy levels set the stage for future atomic models.

The Cloud Model

Bohr's model of the atom has been replaced by the cloud model. In this model, electrons are not confined to specific orbits. Instead, they are spread out in space in a form called an electron cloud. The electron cloud is densest in regions where the probability of finding the electron is highest. Complicated equations describe the shape, location, and density of each electron cloud in an atom. Each cloud corresponds to a particular location for an electron. By incorporating the cloud model into the Rutherford-Bohr model, scientists have been able to construct accurate models of the electron arrangements for all the elements.

Atomic Spectra

When the electrons in excited atoms of an element in the gaseous state return to lower energy levels, they produce a specific series of frequencies of electromagnetic radiation called the **atomic spectrum** of the element. Each element has a characteristic spectrum that differs from that of every other element. Thus, the spectrum can be used to identify the element, even when the element is mixed with other elements.

The element helium was found on the Sun before it was isolated on Earth. Spectral lines of the Sun's corona were studied during a solar eclipse. The lines were not previously reported for any known element, so the new element was named helium from the Greek word for sun, *helios.*

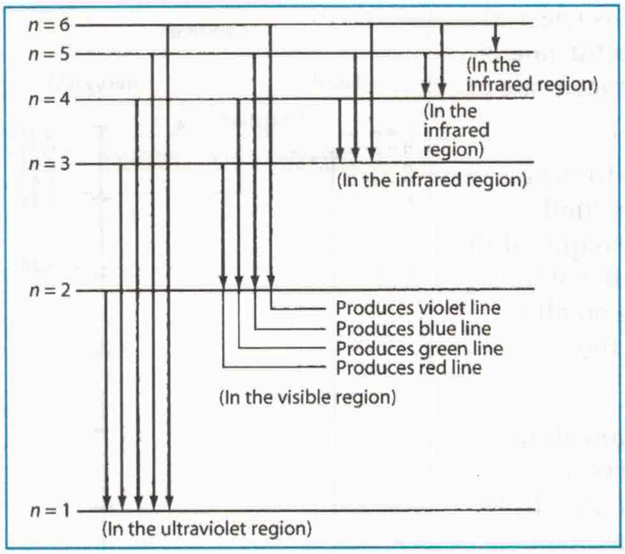


Figure 6-3. The relationship between possible energy level transitions and the observed series of frequencies in the hydrogen spectrum

Emission (Bright-Line) Spectra

Energy levels in an atom, introduced by Bohr, provided an explanation for atomic spectra. When an electron in an atom in an excited state falls to a lower energy level, the energy of the emitted photon is equal to the difference between the energies of the initial and final states. That is

$$E_{photon} = E_i - E_f$$

where E_i is the initial energy of an electron in the higher energy level and E_f is the final energy of the electron in the lower energy level. Each energy difference between two energy levels corresponds to a photon having a specific frequency. A specific series of frequencies, characteristic of the element, is produced when the electrons of its atoms in excited states fall to lower energy levels and the atoms return to lower states or to the ground state. When these emitted frequencies are viewed in a spectroscope, the frequencies appear as a series of bright lines against a dark background and, therefore, are called a **bright-line spectrum** or an **emission spectrum.** In Figure 6-3 the energy emissions producing various series of lines in the ultraviolet, visible light, and infrared regions are indicated in the energy-level diagram for hydrogen.

Absorption Spectra

As explained earlier, an atom can absorb only photons having energies equal to specific differences in its energy levels. The frequencies and wavelengths of these absorbed photons are exactly the same as those of the photons emitted when electrons lose energy and fall between the same energy levels. If the atoms of an element are subjected to white light, which consists of all the visible frequencies, the atoms will selectively absorb the same frequencies that they emit when excited. The absorbed frequencies appear as dark lines in the otherwise continuous white-light spectrum. This series of dark lines, resulting from the selective absorption of particular frequencies in the white-light spectrum of an atom, is called an **absorption spectrum.** An atom will absorb a photon only if the photon possesses the exact amount of energy required to raise the atom to one of its possible excited states.

Review Questions

12. The lowest energy state of an atom is called its
(1) ground state
(2) ionized state
(3) initial energy state
(4) final energy state

13. Which electron transition in the hydrogen atom results in the emission of a photon with the greatest energy?
(1) $n = 2$ to $n = 1$
(2) $n = 3$ to $n = 2$
(3) $n = 4$ to $n = 2$
(4) $n = 5$ to $n = 3$

14. What is the minimum energy required to ionize a hydrogen atom in the $n = 3$ state?
(1) 13.60 eV
(2) 12.09 eV
(3) 5.52 eV
(4) 1.51 eV

15. Which photon energy could be absorbed by a hydrogen atom that is in the $n = 2$ state?
(1) 0.66 eV
(2) 1.51 eV
(3) 1.89 eV
(4) 2.40 eV