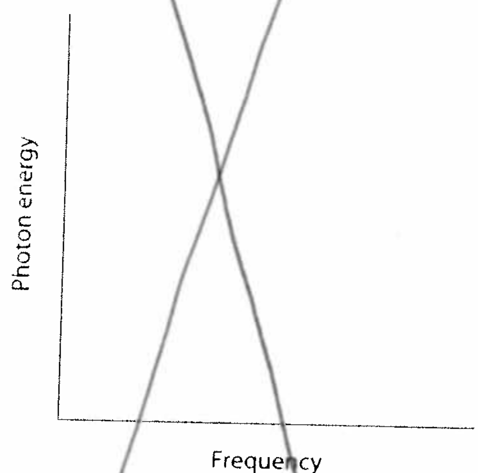
Review Questions

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 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. On the axes that follow, sketch a line to represent the relationship between photon energy and frequency for a series of photons.



5. A photon of green light has a frequency of approximately 6.0×10^{14} hertz. The energy associated with this photon is approximately (1) 1.1×10^{-48} J (2) 6.0×10^{-34} J (3) 5.0×10^{-7} J (4) 4.0×10^{-19} J

6. Determine the energy of a photon having a wavelength of 4.00×10^{-7} meter.

7. 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

8. A gamma photon collides with an electron at rest. During the interaction, the momentum of the photon (1) decreases (2) increases (3) remains the same

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. According to the quantum theory of light, light energy is carried in discrete units called (1) protons (2) photons (3) photoelectrons (4) quarks

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

Just 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 one-tenth 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 an atom are assigned integral numbers, denoted by $n = 1, n = 2$, etc. When the electron is in the lowest energy level ($n = 1$), it is said to be in the **ground state.** For a hydrogen atom, an electron in any level above the ground state is said to be in an **excited state.**

ENERGY LEVELS Any process that raises the energy level of electrons in an atom is called excitation. 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 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.

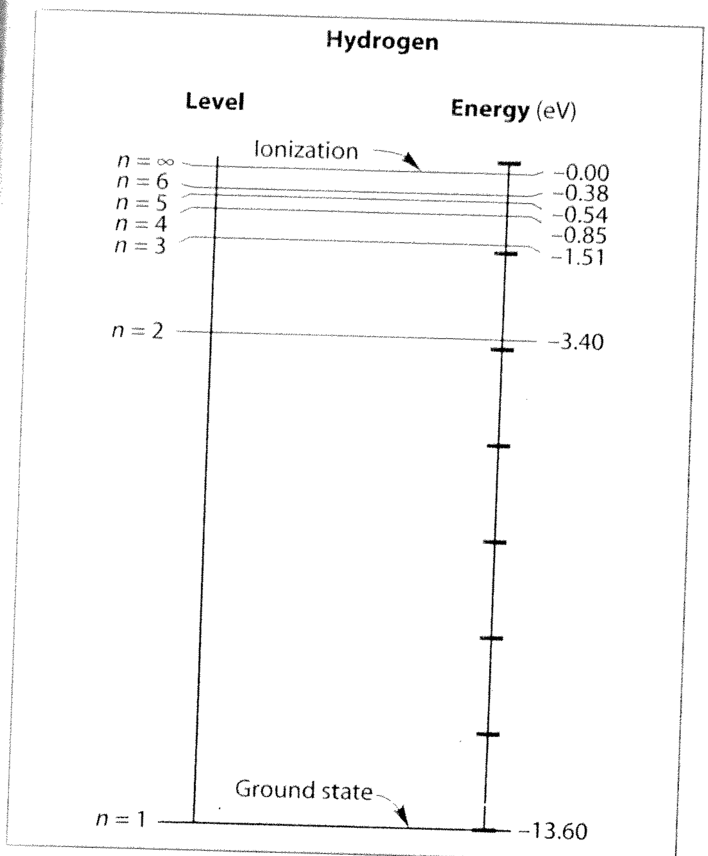


Figure 6-2. **Energy levels for the hydrogen atom**

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 states, 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*.

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.

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

E_i is the initial energy of the electron in its excited state 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 back 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 on the next page 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.

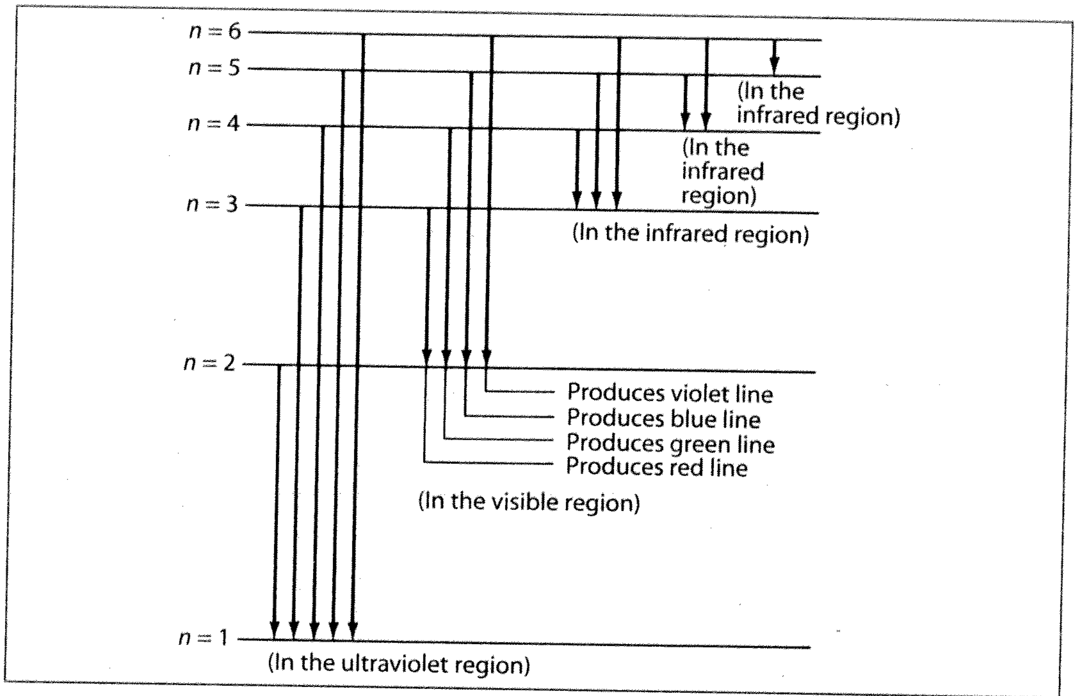


Figure 6-3. The relationship between possible energy level transitions and the observed series of frequencies in the hydrogen spectrum

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

Note to student: The answers to some of the following questions are based on information from the energy-level diagrams for hydrogen and mercury found in the *Reference Tables for Physical Setting/Physics.*

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

16. Hydrogen atoms undergo a transition from the $n = 3$ energy level to the ground state. What is the total number of different photon energies that may be emitted by these atoms?

17. An electron in a mercury atom jumps from level *a* to level *g* by absorbing a single photon. What is the energy of the photon in electronvolts?

18. Which phenomenon provides evidence that the hydrogen atom has discrete energy levels?

19. As an atom absorbs a photon of energy, one of its electrons will (1) exchange energy levels with another of its electrons (2) undergo a transition to a higher energy level (3) undergo a transition to a lower energy level (4) increase its charge

20. Which transition between the energy levels of mercury causes the emission of a photon of highest frequency? (1) *e* to *d* (2) *e* to *c* (3) *c* to *b* (4) *b* to *a*

21. As an atom goes from the ground state to an excited state, the energy of the atom (1) decreases (2) increases (3) remains the same

22. It is possible for an excited hydrogen atom to return to the ground state by the emission of a single photon. Regardless of the initial excited state, this electron transition produces a spectral line in which region of the electromagnetic spectrum? (1) ultraviolet (2) infrared (3) visible light (4) radio waves

23. Determine the frequency of the photon emitted when an excited hydrogen atom changes from energy level $n = 3$ to $n = 2$.

24. An electron in a mercury atom changes from energy level *b* to level *e*. This energy-level change occurs as the atom (1) absorbs a 2.03-eV photon (2) absorbs a 5.74-eV photon (3) emits a 2.03-eV photon (4) emits a 5.74-eV photon

Base your answers to questions 25 through 27 on the information that follows.

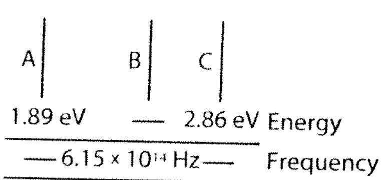
A hydrogen atom emits a 2.55-electronvolt photon as its electron changes from one energy level to another.

25. Determine the energy level change for the electron.

26. Express the energy of the emitted photon in joules.

27. Determine the frequency of the emitted photon.

Base your answers to questions 28 through 30 on the following diagram, which represents three visible lines in the hydrogen spectrum. Either the energy or the frequency for each of these lines is given below the diagram.



28. Which energy level transition produced line A? (1) $n = 2$ to $n = 1$ (2) $n = 3$ to $n = 2$ (3) $n = 4$ to $n = 3$ (4) $n = 5$ to $n = 2$

29. Determine the energy of the photons that produced line B.

30. Express the energy of line C in joules.

Base your answers to questions 31 through 34 on the following information.

A photon with 14.60 electronvolts of energy collides with a mercury atom in its ground state.

31. Express the energy of the incident photon in joules.

32. Determine the frequency of the incident photon.

33. In what region of the electromagnetic spectrum is the frequency of the incident photon? (1) gamma rays (2) infrared (3) visible (4) ultraviolet

34. If the photon collision ionizes the atom, what is the maximum energy that the electron removed from the atom can have? (1) 0.00 eV (2) 4.22 eV (3) 10.38 eV (4) 14.60 eV

The Nucleus

Rutherford's experiments showed that all of the atom's positive charge and nearly all of its mass is contained in the nucleus. The **nucleus** is the core of an atom made up of one or more protons and (except for one of the isotopes of hydrogen) one or more neutrons. The protons and neutrons that make up the nucleus of an atom are called nucleons.

Nuclear Force

The positively charged protons in any nucleus containing more than one proton are separated by a distance of 10^{-15} meter. Consequently, a large repulsive Coulomb force exists between them. The gravitational force of attraction between protons is far too weak to counterbalance this electrostatic force of repulsion. Thus, there must exist a very strong attractive nuclear force to keep the protons concentrated in the nucleus of an atom. It is this **nuclear force,** which is an attractive force between protons and neutrons in an atomic nucleus, that is responsible for the stability of the nucleus.

The nuclear force of attraction between two protons in a nucleus is about 100 times stronger than the electrostatic force of repulsion. At distances greater than a few nucleon diameters, however, the nuclear force diminishes rapidly and becomes much less than the gravitational or electrostatic forces. Although nuclear forces are the strongest forces known to exist, they are effective only over a short distance.