Section 5 Line Spectra and Electron "Jumps"

What Do You See?

Learning Outcomes

In this section you will

- **View** the spectrum of hydrogen.
- **Interpret** changes in electron energies in the hydrogen atom to develop an explanation for where the colored light in the hydrogen spectrum comes from.
- **Use** Bohr's model of the atom to predict parts of the hydrogen-atom spectrum.
- **Compare** the wavelengths, energies, and frequencies of light of different colors.
- **Identify** regions in the electromagnetic spectrum.
- **Explain** the photoelectric effect.

What Do You Think?

Neon signs make up some of the lights of Broadway in New York City. They are also used in advertising signs throughout the world.

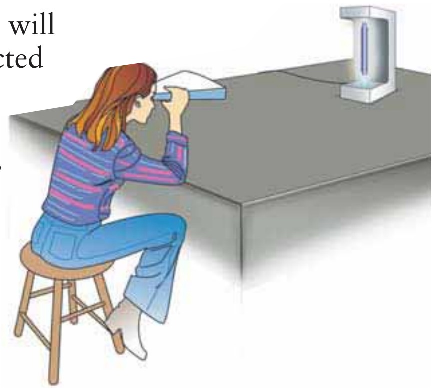
- **How is the color produced in a neon sign?**

Record your ideas about this question in your *Active Chemistry* log. Be prepared to discuss your response with your small group and the class.

Investigate

1. In order to directly observe the behavior of atoms, you can observe the spectrum of visible *light* given off when atoms are excited by a high-voltage, electric-power supply. You probably have already seen this effect in the familiar red color of neon signs.

   For this demonstration, your teacher will set up a tube of hydrogen gas connected to a high-voltage power supply. This light can be viewed through a spectroscope or a diffraction grating, as shown in the diagram. When the slit at the end of the spectroscope is aimed toward the light, the colors of the spectrum appear separately off to the sides of the slit.



a) What colors do you see in the spectrum of light given off by hydrogen gas?

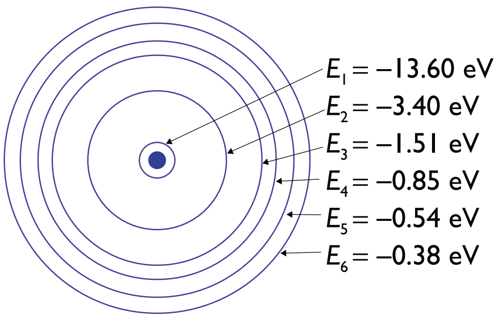
b) Make a colored diagram in your *Active Chemistry* log of what you see inside the spectroscope. Make sure to draw and label the colors in the proper order and spacing between them that you observe.

2. When you observed the spectrum of light given off by hydrogen gas, you probably saw three or four distinct lines, each having a different color. The color of light is a measure of its energy. The colors closer to red in the spectrum have the least energy and the colors closer to violet have the greatest energy.

a) List the colors that you observed from lowest energy to highest energy.

3. In 1913, Niels Bohr, a Danish physicist, tried to explain the line spectrum of hydrogen. He hypothesized that the electron in the hydrogen atom is allowed to have only certain amounts of energy. These energy levels would be *orbits*, or energy levels, in which the electron in hydrogen could circle the nucleus of hydrogen.

The diagram shows a sketch of some of the possible orbits of the hydrogen electron with their corresponding energies.



a) Think of the energy values as being on a number line. Which orbit has the greatest energy?

b) Which orbit has the least energy?

4. Bohr believed that an electron in a higher energy level would give off light when it jumps to a lower energy level. The amount of energy in the light would be the difference in these energy levels. When an electron jumps from E_2 to E_1, the amount of energy in the light would be $E_2 - E_1 = (-3.4) - (-13.6) = +10.2$ units of energy. Note that E_2 has a larger energy than E_1 (that's because $-3.4 > -13.6$).

Calculate the amount of energy in the light when an electron jumps from the following energy levels.

a) E_3 to E_1

b) E_3 to E_2

c) E_4 to E_2

d) E_5 to E_2

5. Picture the electron in orbit around the nucleus of the hydrogen atom as Bohr did. Allowing the electron to have only certain amounts of energy would mean that the electron could be allowed in orbits of only certain distances from the nucleus. As the electron "jumps" from one energy level to another, it behaves something like a ball falling down a flight of uneven stairs. It is allowed to rest only on one of the steps, but nowhere in between.

a) Every possible jump corresponds to light of a different energy. How many different energies of light can be emitted from hydrogen when the electron jumps down to E_1 from E_2, E_3, E_4, E_5, and E_6?

b) How many different energies of light can be emitted from hydrogen when the electron jumps down to E_2 from E_3, E_4, E_5, and E_6?

c) An electron can also jump from E_5 to E_3 to E_2. Calculate the two energies of light that would be emitted in these two jumps.

6. Each energy of light corresponds to a specific color. The four colors of light that you observed in the hydrogen spectrum corresponded to four specific energies and therefore four specific jumps. Bohr found that the four colors of light had energies that corresponded to the energy differences when electrons jumped from a higher level to the E_2 level. However, Bohr was unable to find light colors corresponding to jumps to the E_1 level or the E_3 level.

a) Why do you think he was unable to find light colors corresponding to the jumps in the E_1 level?

7. The colors corresponding to jumps to the E_1 level have higher energies and are not seen by the eye. They do not correspond to visible light. When detectors were built that could see into the ultraviolet portion of the spectrum, the *wavelengths* of light, or the distance between light wave crests, corresponded to the energy levels that Bohr predicted when the electron jumped down to the E_1 level.

a) Why do you *now* think that Bohr was not able to find light colors corresponding to the jumps to the E_3 level? (Hint: Compare the energies of these jumps to the energies when the electron jumped to the E_2 level.)

Prediction of the energies of the light given off when the electrons jumped from higher energy states to the E_1 level and the E_3 level demonstrated what a powerful model Bohr had invented. He was able to predict accurately some light given off that nobody had ever measured.

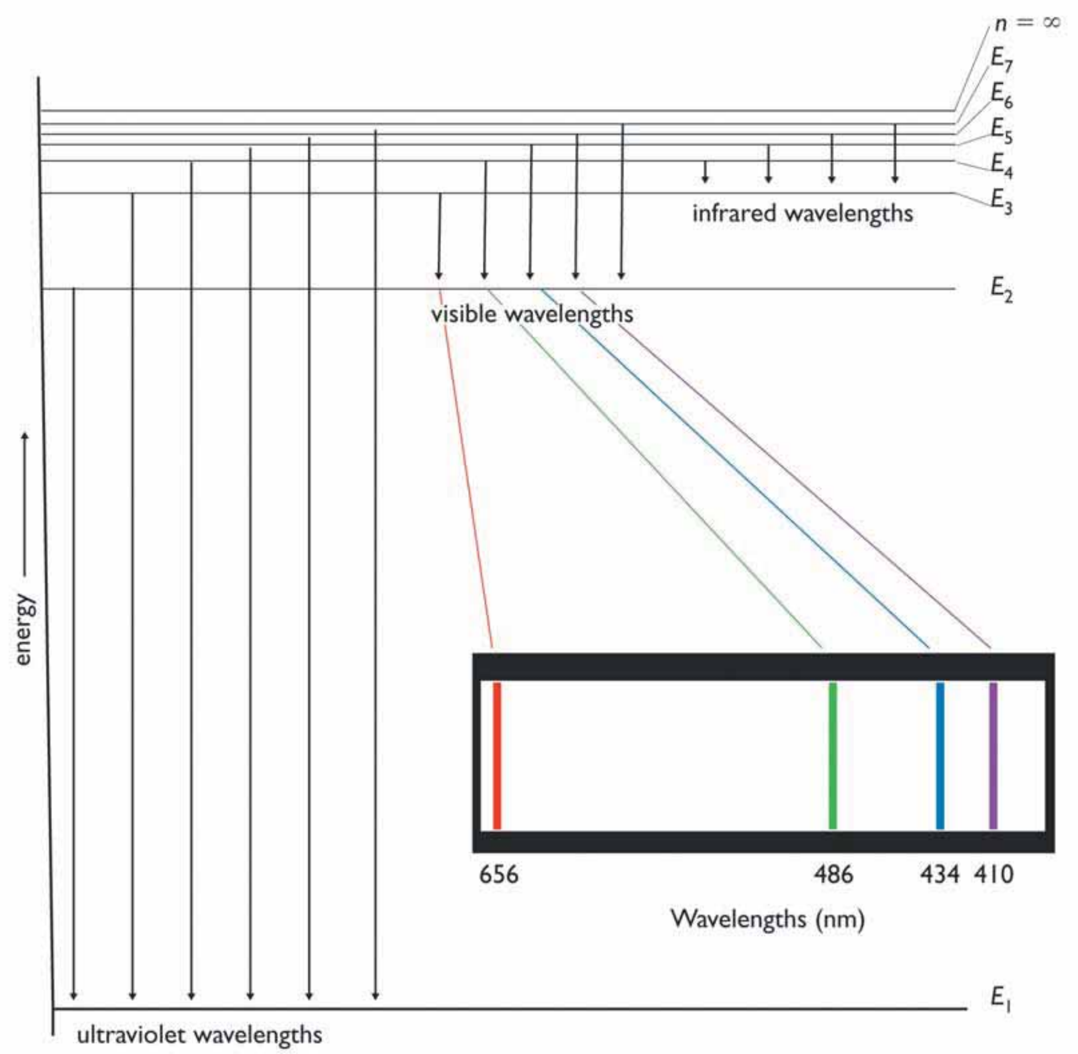
8. The electrons can get "excited" and move from lower energy levels to higher energy levels by absorbing light of just the right energy. An electron in the E_1 orbit can absorb 10.2 units of energy and jump to the E_2 level. Refer back to the energies of the different levels. Provide the energies of light that can be absorbed by an electron to move it through each of the following.

a) E_1 to E_3

b) E_2 to E_3

c) E_2 to E_4

9. You now have grasped the basics of the Bohr model of the atom. The electron of hydrogen is able to orbit the nucleus in specific orbits. Each of these orbits corresponds to a specific energy. When the electron jumps from a higher energy level (further from the nucleus) to a lower energy level (closer to the nucleus), the electron loses energy. Light of exactly this energy is given off. Each jump corresponds to a different energy of light. Some of these light energies can be detected by the human eye and are called visible light. You see these different energies as different colors. Other jumps correspond to energies of light that are not detected by the human eye. However, they can be detected by instruments. These energies correspond to light in the infrared or ultraviolet parts of the spectrum. When electrons jump to lower energy levels, light is given off. When electrons absorb light, they jump to higher energy levels.



Chem Talk

BOHR'S MODEL OF AN ATOM

Electron Orbits

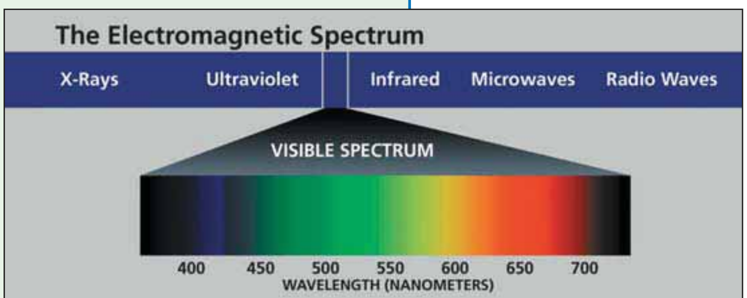
Niels Bohr

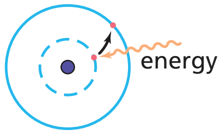
Niels Bohr was a brilliant Danish physicist. He proposed a "planetary" model of the atom. He theorized that electrons travel in nearly circular paths, called **orbits**, around the nucleus. Each electron orbit has a definite amount of energy. The farther away the electron is from the nucleus, the greater is its energy. Bohr suggested the revolutionary idea that electrons "jump" between energy levels (orbits) in a *quantum* fashion. This means that they can never exist in an in-between level. Thus, when an atom absorbs or gives off energy (as in light or heat), the electron jumps to higher or lower orbits. Electrons are the most stable when they are at lower energy levels closer to the nucleus.

The Electromagnetic Spectrum

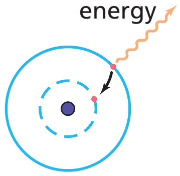
You observed **light** when hydrogen gas was given a large voltage. This visible light is only one part of the **electromagnetic spectrum**. You have probably heard of some of the other parts. These include ultraviolet, infrared, X-rays, gamma rays, microwaves, and radio waves. As you demonstrated in calculations using Bohr's model, the light from some of the transitions is in the ultraviolet region. Infrared light is also emitted as the electron jumps from E_4 to E_3 and E_5 to E_3 and other higher energy levels.

Many people do not think of radio waves as being similar to light waves. However, they are part of the electromagnetic spectrum. Your eyes can see only a very small part of the electromagnetic spectrum. Often, you hear radio announcers say that they are broadcasting at a certain **frequency**. Your FM radio dial may have MHz (megahertz) printed on the side. This tells you that the numbers correspond to frequencies in units of MHz or 10^6 Hz. Frequency tells you the number of cycles or waves that are being produced per second. The unit for frequency is a hertz (Hz). 1 Hz = 1 cycle/s = 1 s^{-1}. Normally, frequency is read as per second and the cycles are dropped from the terminology.





(a) An electron gains a quantum of energy.



(b) An electron loses a quantum of energy.

Chem Words

orbit: in Bohr's model of the hydrogen atom, the circular path of an electron traveling around the nucleus (similar to the planet's orbit around the Sun).

light: electromagnetic radiation that lies within the visible range.

electromagnetic spectrum: the entire range of electromagnetic radiation. At one end of the spectrum are gamma rays, which have the shortest wavelengths and high frequencies. At the other end are radio waves, which have the longest wavelengths and low frequencies. Visible light is near the center of the spectrum.

frequency: the number of waves per second or cycles per second or Hertz (Hz).

Chem Words

wavelength: the distance measured from crest to crest of one complete wave.

photon: a quantum or discrete packet of electromagnetic radiation that has a given amount of energy.

Wavelength is the distance from crest to crest of a wave. The symbol for wavelength is λ, the Greek letter *lambda*. The speed of electromagnetic radiation is constant and it is called the speed of light, c. Its value is 2.998×10^8 m/s. This is often approximated to 3.00×10^8 m/s.

Frequency and wavelength are related through the speed of light:

$$c = \lambda f$$

From this information, you can calculate the frequency of light of a given wavelength. The equation that is used for this is:

$$f = \frac{c}{\lambda}$$

As an example, the wavelength of blue light is 434.2 nm and the frequency can be calculated as follows:

$$\begin{aligned} f &= \frac{2.998 \times 10^8 \text{ m/s}}{434.2 \times 10^{-9} \text{ m}} \\ &= 6.905 \times 10^{14} \text{ cycles/s} \\ &= 6.905 \times 10^{14} \text{ Hz} \end{aligned}$$

As you go across the electromagnetic spectrum from right to left, you should note that the wavelength continues to get smaller as the frequency increases. Also, you should understand that the energy of the spectrum increases as you go from radio waves to X-rays or gamma rays. Max Planck, a German physicist, found that light comes in quantum or discrete packets called **photons**. The energy of a specific wavelength or frequency of light could be calculated.

The equation he developed is

$$E = hf$$

where h is Planck's constant and is 6.63×10^{-34} J • s and f is the frequency.

The frequency of red light is 4.567×10^{14} Hz which is lower than blue light. The energy of red light can be calculated as follows:

$$\begin{aligned} E &= hf \\ &= (6.63 \times 10^{-34} \text{ J} \bullet \text{s})\ (4.567 \times 10^{14} \text{ Hz}) \\ &= 3.03 \times 10^{-19} \text{ J} \end{aligned}$$

So, the next time that you are standing around a campfire, you can inform your fellow campers that red light has less energy than blue light. You can also tell them how to calculate these values.

Example:

In hydrogen the energy change of an electron jumping from E_3 to E_2 is 3.03×10^{-19} J. Conservation of energy insists that this is equal to the energy of the light that is emitted.

You can find the frequency of the light.

$$\begin{aligned} E &= hf \\ 3.03 \times 10^{-19} \text{ J} &= (6.63 \times 10^{-34} \text{ J} \bullet \text{s})\ (f) \\ f &= 4.57 \times 10^{14} \text{ Hz} \end{aligned}$$

Knowing the frequency of light, you can find the wavelength.

$$c = \lambda f$$
$$3 \times 10^8 \text{ m/s} = (4.57 \times 10^{14} \text{ Hz})\ \lambda$$
$$\lambda = 6.56 \times 10^{-7} \text{ m}$$

Referring to a chart that shows wavelength and color, you find that this is red light. It is the red light that you observed when looking at the hydrogen spectrum.

Calculations and Units

When you make chemistry calculations, it is very important to pay close attention to the units in your answer. Look closely at the units in the example given. Do they make sense?

When you calculated the frequency of the light, you used the unit for energy, the joule (J). You also used the unit joule seconds (J·s) for Planck's constant. The resulting unit in the calculation was the hertz (Hz). How is this possible? Look only at the units in the calculation.

$$f = \frac{\text{J}}{\text{J}\cdot\text{s}}$$
$$= \frac{\cancel{\text{J}}}{\cancel{\text{J}}\cdot\text{s}}$$
$$= \frac{1}{\text{s}} \text{ or } 1\ \text{s}^{-1}$$

Notice how the units J cancel out, leaving s^{-1}. You read earlier that $1\ s^{-1} = 1$ Hz. Therefore, the resulting unit is the hertz, which is the unit for frequency.

Now look at the units in the calculation for wavelength.

$$\lambda = \frac{\frac{\text{m}}{\text{s}}}{\text{Hz}}$$
$$= \frac{\text{m}\cdot\cancel{\text{s}}^{-1}}{\cancel{\text{s}}^{-1}}$$
$$= \text{m}$$

The unit Hz is equal to s^{-1}. The units s^{-1} cancel out, leaving m. The m unit is a measure of length, in this case the wavelength.

Analyzing the units in a calculation is called dimensional analysis, and it is a very important tool that chemists use. Checking to see if the units make sense ensures that their calculations make sense and that they have not made an error. Use this tool as you work through all your calculations in chemistry.

The Photoelectric Effect

Light has wave-like properties like wavelength and frequency. You observed these wave-like properties in your investigation of the hydrogen spectrum. Light also has particle-like properties like momentum and can be seen in collisions with matter. The photoelectric effect is a collision between

Checking Up

1. How are visible light, ultraviolet light, infrared light, X-rays, gamma rays, microwaves, and radio waves related?
2. Explain the meaning of wavelength.
3. Why is "planetary" model an appropriate name for Bohr's model of the atom?
4. How do the energy levels of different electron orbits compare?
5. Why do elements produce certain color light when heated?
6. What two particles collide in the photoelectric effect?
7. To prove that blue light is higher energy radiation than red light, calculate its energy.

a particle of light (a photon) and an electron on the surface of a metal. The energy of this photon is calculated using the equation $E=hf$. The energy of light is *quantized*. That means that it comes in fixed amounts. In Einstein's explanation of the photoelectric effect, he showed that light must behave like a particle with fixed energies that depend on the frequency. For this he was awarded the Nobel Prize in 1921.

The Problem with Bohr's Atomic Theory

Niels Bohr was aware that his theory of electron jumps had incredible success but also raised some problems. Bohr's theory could only account for the spectrum of hydrogen. It could not account for the spectra of any other element. Bohr's theory could not explain why only certain orbits were allowed. It could also not explain how the electron could jump from one orbit to another. Other scientists improved on Bohr's model as they discovered more about the atom and quantum mechanics.

What Do You Think Now?

At the beginning of the section you were asked the following:

- **How is the color produced in a neon sign?**

You now know why light of certain wavelength (and color) is given off when energy is applied to neon gas. How would you explain to a friend how a neon light works?

Chem

Essential Questions

What does it mean?

Chemistry explains a *macroscopic* phenomenon (what you observe) with a description of what happens at the *nanoscopic* level (atoms and molecules) using *symbolic* structures as a way to communicate. Complete the chart below in your *Active Chemistry* log.

MACRO	NANO	SYMBOLIC
Describe what you saw through the spectroscope or the diffraction grating when the tube of hydrogen gas was subjected to a large voltage.	*Explain how the light is created in the hydrogen atom. Why are only certain colors emitted?*	*Make a sketch to represent what is occurring at the atomic level when the line spectrum for hydrogen is produced. What equation is used to calculate the different electron energy levels?*

How do you know?

The excitation of hydrogen atoms produces a visible line spectrum that allows you to measure the wavelength and then determine the energy of that specific wavelength of light. How well did the calculations of the energy correspond to the energy (or wavelength) measured in the lab?

Why do you believe?

The Sun is primarily made of hydrogen. What colors (or wavelengths) of light would you expect to see?

Why should you care?

Understanding what causes the line spectrum of hydrogen atoms will help to understand the behavior of hydrogen's electron. This in turn will help you to understand many of hydrogen's properties, and will be an excellent addition to your game.

Reflecting on the Section and the Challenge

In this section you learned that the electrons in an atom are responsible for the colors of light emitted. Bohr explained the spectrum of light given off by excited hydrogen atoms. He hypothesized that each hydrogen atom's electron was allowed in only certain energy levels. Mendeleev knew nothing about electrons or energy levels when he *first* developed the periodic table. Yet, the periodic table today is seen as a reflection of the number of electrons in an atom of each element and the energy levels those electrons occupy. Continue to think about how the electronic structure of an atom could be included in the game you are designing.

Chem to Go

1. In this section you were told that light with greater frequency has greater energy.
   a) Which color in the visible spectrum of hydrogen has the greatest energy? The least energy?
   b) Which color in the visible spectrum of hydrogen has the highest frequency? The lowest frequency?
2. If an electron were to fall down to the E_1 level from the E_3 level, how would its energy compare to one that only fell to the E_2 level? Explain.
3. What is the difference measured in energy when an electron falls from E_3 to E_1? How many times greater is this value as compared to the difference of the electron falling from E_3 to E_2?
4. A wavelength of light is 389.0 nm and its frequency is 7.707×10^{14} Hz. Show how this frequency value was calculated.
5. Show that the wavelength of 389.0 nm has an energy of 5.11×10^{-19} J.
6. a) Microwave radiation is absorbed by the water in food. As the water absorbs the heat it causes the food to get hot. If the λ of a microwave is 10 cm, calculate the frequency of the microwave and the energy of each photon.
   b) The red light you observed in the hydrogen spectrum had a $\lambda = 656.5$ nm. The energy of the red light was 3.03×10^{-19} J. How many times greater is this value when compared to the energy value that you found for the microwave energy?
7. Describe how the Bohr model explains the spectra emitted by hydrogen.
8. Describe all possible energy transitions that can be emitted from an electron in the E_4 level as it moves to the E_1 level.
9. What are two limitations of the Bohr model for the atom?
10. Draw the set of energy levels of hydrogen to scale. Which transition produces the highest energy light?
11. What do you think happens when an electron in a hydrogen atom moves from a higher energy level to the lowest E_1 level? This series is known as the *Lyman series* (ultraviolet light).