

Activity 5

The Chemical Behavior of Atoms

**GOALS**

In this activity 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.
- Calculate and compare the wavelengths, energies, and frequencies of light of different colors.
- Identify regions in the electromagnetic spectrum.

What Do You Think?

A neon sign uses electricity and a gas-filled tube to produce a colored light. A fluorescent bulb uses electricity, a gas-filled tube and a phosphor coating to produce white light.

- 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 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 lens, as shown in the diagram on the next page. 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.

What Do You Think?

Neon emits certain colors in its line spectrum when electrons of the gas become excited and then return to lower energy levels. By using colored glass we can absorb certain wavelengths and allow only certain colors to be emitted. A variety of colored glass tubing is used for neon signs in order to create the effect desired. When you demonstrate gas discharge tubes you may want to show students how a colored piece of glass placed in front of the discharge tube will eliminate one of the colors that could originally be seen.

Student Conceptions

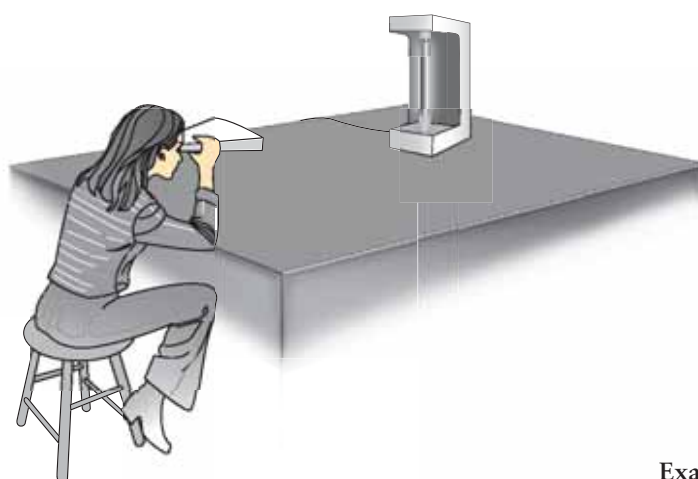
Explaining the origin of the hydrogen spectrum is a conceptual task. Although there is a great deal of evidence to support them, the concepts involved here—particularly the quantization of atomic energy levels—were difficult even for scientists to accept. Nothing in the tangible world exhibits quantization that can be seen. It is difficult, then, for students to see that only certain colors of light can be produced by a hydrogen atom because its electron is restricted to certain energy levels. To view the question from the other direction, one might ask students, “Why can the hydrogen atom not produce orange or yellow light?”

Students tend to cling strongly to the misconception that electrons orbit around the nucleus as planets do around the Sun. This incorrect model of the atom, called the Bohr model, was an important step in the process of understanding the quantization of energy levels. While the Bohr model correctly predicts the hydrogen spectrum, it breaks down for atoms with more than one electron. The model that is accepted today, the quantum mechanical model, is explored in the next two activities. It is much more difficult to visualize because it deals with predictions of the energies of electrons rather than their locations in space. This is why the Bohr model is difficult to overwrite.

The electromagnetic spectrum is another source of misconceptions for students. Students are apt to describe it as extending from red to violet light (ROY.G.BIV). However, this is only the visible part of the electromagnetic spectrum, that is, the part to which human’s eyes have adapted. The full spectrum includes radio waves, infrared radiation, ultraviolet radiation and x-rays as well.



Active Chemistry The Periodic Table



- 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 the colors with the 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 determined by its frequency; the greater the energy in the light, the greater the frequency.

The diagram on the right shows the visible lines of the line spectrum of hydrogen.

The wavelengths are measured in nanometers (nm). The prefix *nano*

means 10^{-9} . Thus, 656.5 nm is 656.5×10^{-9} m. The frequency of each line can be determined using the equation:

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

where,

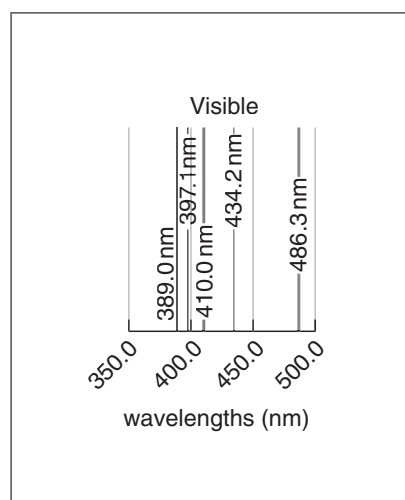
f is the frequency (waves/s or cycles/s or hertz),

c is the speed of light and is a constant (2.998×10^8 m/s), and

λ is the wavelength measured in nanometers.

Example:

Given red light with a wavelength $\lambda = 656.5 \times 10^{-9}$ m and the speed of light $c = 2.998 \times 10^8$ m/s, calculate the corresponding frequency and energy of the light.



Investigate

Teaching Suggestions and Sample Answers

1. a) Red (656 nm); Green (486 nm); Violet (435 nm); Violet (410 nm).
b) Students will produce a scale similar to the values in Part a.

Teaching Notes

The students should understand that the electron falling from a higher energy level down to the 2nd energy level causes the visible emission spectrum that they are seeing. If it falls to the 1st energy level it will be in the ultraviolet region.

If you do not have a special spectrum tube power supply setup, then you must be aware that a common power supply may provide too much current and could shorten the life of the spectrum tube and can also cause it to overheat. A DC rheostat volt control should be used in conjunction with the power supply unit.

First calculate the frequency.

$$\begin{aligned}
 f &= \frac{c}{\lambda} \\
 &= \frac{2.998 \times 10^8 \text{ m/s}}{656.5 \times 10^{-9} \text{ m}} \\
 &= 4.567 \times 10^{14} \text{ cycles/s (or Hz)} \\
 &= 4.567 \times 10^{14} \text{ Hz (Hertz)}
 \end{aligned}$$

To calculate the corresponding energy of a photon of light, you can use an equation called Planck's equation that relates the frequency to the energy. A photon is a tiny, indivisible packet of light. Planck's constant is h and has the value of $6.63 \times 10^{-34} \text{ J}\cdot\text{s}$.

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

- a) Calculate the frequencies and energies for the other visible wavelengths of light in the hydrogen spectra. Complete a table like the one below in your *Active Chemistry* log.
3. In 1913, Niels Bohr, a Danish physicist, tried to explain the line

spectrum of hydrogen by hypothesizing that the electron in the hydrogen atom is allowed to have only certain amounts of energy. Bohr believed that an electron in a higher energy level would give off light when it fell to a lower energy level, and the amount of energy in the light would be the difference in these energy levels. Bohr pictured the atom to be similar to the Solar System, with electrons orbiting around the nucleus in the way that planets orbit around the Sun. If you picture the electron to be in orbit about 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 distance from the nucleus. As the electron "jumps" from one energy level to another, it behaves something like a ball falling down a flight of stairs. It's allowed to rest only on one of the steps, nowhere in between.

The light was the result of electron energy jumps. These jumps should be able to produce the corresponding energies of the light. Bohr calculated the lowest energy level to be $-2.18 \times 10^{-18} \text{ J}$. Each of the other

Wavelength λ (nm)	Frequency f (cycles/s or Hz)	Energy E
410.3		
434.2		
486.3		
656.5	4.567×10^{14}	$3.03 \times 10^{-19} \text{ J}$

2.

Wavelength λ (nm)	Frequency f (cycles/s or Hz)	Energy E (kJ/mol)
410.3	7.307×10^{14}	4.84×10^{-19}
434.2	6.905×10^{14}	4.58×10^{-19}
486.3	6.165×10^{14}	4.09×10^{-19}
656.5	4.567×10^{14}	3.03×10^{-19}



Active Chemistry The Periodic Table



Niels Bohr

energy levels can be computed from this first one with the equation:

$$E_n = \frac{1}{n^2} E_1$$

The second energy level:

$$\begin{aligned} E_2 &= \frac{1}{2^2} E_1 \\ &= \frac{1}{4} (-2.18 \times 10^{-18} \text{ J}) \\ &= -5.45 \times 10^{-19} \text{ J} \end{aligned}$$

The third energy level:

$$\begin{aligned} E_3 &= \frac{1}{3^2} E_1 \\ &= \frac{1}{9} (-2.18 \times 10^{-18} \text{ J}) \\ &= -2.42 \times 10^{-19} \text{ J} \end{aligned}$$

- a) Calculate the energy levels for the fourth, fifth, and sixth levels and record all these values in your log.

4. In Bohr's model, the electrons jump from E_3 to E_2 and give off light with that energy.

$$\begin{aligned} E_3 - E_2 &= (-2.42 \times 10^{-19} \text{ J}) - \\ &\quad (-5.45 \times 10^{-19} \text{ J}) \\ &= 3.03 \times 10^{-19} \text{ J} \end{aligned}$$

This is equivalent to the energy of the red light. The model is successful, so far.

- a) Calculate the energy of light when the electron of hydrogen jumps from:

E_4 to E_2

E_5 to E_2

E_6 to E_2

- b) Do these energies correspond to the energies of the light that is emitted from hydrogen that you calculated earlier?

- c) If the model is a truly good model, it should be able to predict something unknown. Bohr's model does just this. Calculate the energy of light when the electron of hydrogen jumps from E_2 to E_1 .

The light corresponding to this transition was not observed because it is not in the visible range of light. It is in the ultraviolet spectrum. When an ultraviolet detector is used, the wavelength of the ultraviolet light is just as predicted by Bohr's model. The light emitted from transitions of E_3 to E_1 and E_4 to E_1 are also in the ultraviolet.

3. a) The First energy level: $E_1 = \frac{1}{1^2}(-2.18 \times 10^{-18}) = -2.18 \times 10^{-18} \text{ J}$
The Second energy level: $E_2 = \frac{1}{2^2}(-2.18 \times 10^{-18}) = -5.45 \times 10^{-19} \text{ J}$
The Third energy level: $E_3 = \frac{1}{3^2}(-2.18 \times 10^{-18}) = -2.42 \times 10^{-19} \text{ J}$
The Fourth energy level: $E_4 = \frac{1}{4^2}(-2.18 \times 10^{-18}) = -1.36 \times 10^{-19} \text{ J}$
The Fifth energy level: $E_5 = \frac{1}{5^2}(-2.18 \times 10^{-18}) = -8.72 \times 10^{-20} \text{ J}$
The Sixth energy level: $E_6 = \frac{1}{6^2}(-2.18 \times 10^{-18}) = -6.06 \times 10^{-20} \text{ J}$
4. a) $E_3 \text{ to } E_2 = (-2.42 \times 10^{-19}) - (-5.45 \times 10^{-19}) = 3.08 \times 10^{-19} \text{ J}$
 $E_4 \text{ to } E_2 = (-1.36 \times 10^{-19}) - (-5.45 \times 10^{-19}) = 4.09 \times 10^{-19} \text{ J}$
 $E_5 \text{ to } E_2 = (-8.72 \times 10^{-20}) - (-5.45 \times 10^{-19}) = 4.58 \times 10^{-19} \text{ J}$
 $E_6 \text{ to } E_2 = (-6.06 \times 10^{-20}) - (-5.45 \times 10^{-19}) = 4.84 \times 10^{-19} \text{ J}$
- b) The same energies.
- c) $E_2 \text{ to } E_1 = (-5.45 \times 10^{-19}) - (-2.18 \times 10^{-18}) = 1.64 \times 10^{-18} \text{ J}$
 $E_3 \text{ to } E_1 = (-2.42 \times 10^{-19}) - (-2.18 \times 10^{-18}) = 1.94 \times 10^{-18} \text{ J}$
 $E_4 \text{ to } E_1 = (-1.36 \times 10^{-19}) - (-2.18 \times 10^{-18}) = 2.04 \times 10^{-18} \text{ J}$
 $E_5 \text{ to } E_1 = (-8.72 \times 10^{-20}) - (-2.18 \times 10^{-18}) = 2.09 \times 10^{-18} \text{ J}$
 $E_6 \text{ to } E_1 = (-6.06 \times 10^{-20}) - (-2.18 \times 10^{-18}) = 2.12 \times 10^{-18} \text{ J}$



When the electron jumps from a higher energy level to a lower energy level, light is emitted. The light emitted from a neon sign is based on this phenomenon.

The opposite can also occur. If the electron absorbs light of just the right energy, the electron can jump up to a higher energy level. For instance, an electron in the E_2 energy level can absorb light of 3.02×10^{-19} J and jump up to the E_3 energy level. If enough energy

is applied to the hydrogen atom, the electron can be totally removed. This is commonly called the ionization energy; for one electron of hydrogen, the ionization energy is 2.18×10^{-19} J. If the electron in the E_1 level absorbs light of energy 2.18×10^{-19} J, then its new energy will be 0 J. This means that it is no longer bound to the hydrogen nucleus. The atom is ionized (the electron has been removed).



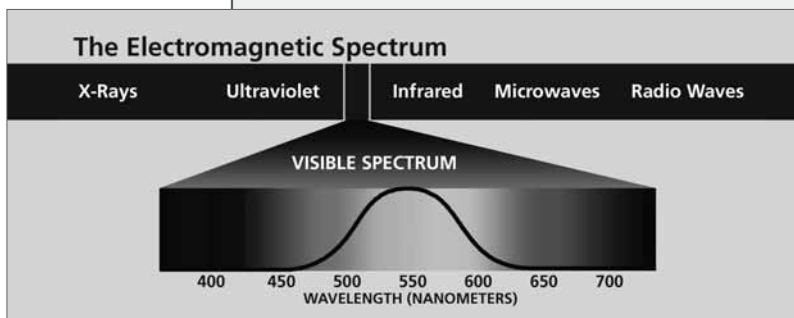
Active Chemistry The Periodic Table

ChemTalk

BOHR'S MODEL OF AN ATOM

The Electromagnetic Spectrum

Visible light is only one part of the **electromagnetic spectrum**. You've probably heard of some of the other parts including 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. 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 light waves, but they are part of the electromagnetic spectrum. The problem is that you can only see a

Chem Words

electromagnetic spectrum: the complete spectrum of electromagnetic radiation, such as radio waves, microwaves, infrared, visible, ultraviolet, x-rays, and gamma rays

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

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

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). $1 \text{ Hz} = 1 \text{ cycle/s} = 1 \text{ s}^{-1}$. Normally, frequency is read as per second and the cycles are dropped from the terminology.

Wavelength (λ), where λ is the Greek letter *lambda*, is the distance from crest to crest of a wave. All light waves travel at the same speed. The speed of electromagnetic radiation is constant and it is called the speed of light (c), $c = 2.998 \times 10^8 \text{ m/s}$ or $3.00 \times 10^8 \text{ m/s}$. 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, if the wavelength is 434.2 nm, then the frequency is:

$$f = \frac{2.998 \times 10^8 \text{ m/s}}{434.2 \times 10^{-9} \text{ m}}$$

$$= 6.905 \times 10^{14} \text{ cycles/s or } 6.905 \times 10^{14} \text{ Hz.}$$

As you go across the electromagnetic spectrum 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 the energy of a wavelength could be calculated. The equation that he developed was based on measuring the change in energy from one level to another level like you did in the **Investigate** section. The equation he developed is

$$E = hf,$$

where h is Planck's constant and is $6.63 \times 10^{-34} \text{ J} \cdot \text{s}$ and f is the frequency.

The corresponding energy of the red light above would be:

$$E = hf$$

$$= (6.63 \times 10^{-34} \text{ J} \cdot \text{s}) (4.567 \times 10^{14} \text{ Hz})$$

$$= 3.03 \times 10^{-19} \text{ J.}$$





Active Chemistry The Periodic Table

Chem Words

orbit: the path of the electron in its motion around the nucleus of Bohr's hydrogen atom

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?

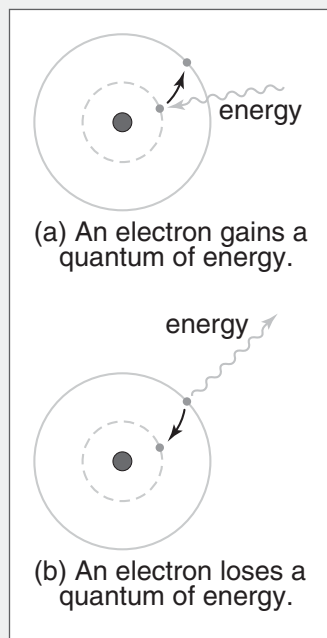
So the next time that you are standing around a campfire, you can inform your fellow campers that red light has less energy than blue light and you can also tell them how to calculate these values.

Bohr's Atomic Theory

Niels Bohr, a brilliant Danish physicist, was aware that his theory of electron jumps had incredible success but also raised some problems.

Bohr 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, and 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. That is, they can never exist in an in-between state. 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. Bohr's theory could only account for the spectrum of hydrogen and not for the spectra of any other element. Bohr's theory could not explain why only certain orbits were allowed, nor 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.



(a) An electron gains a quantum of energy.

(b) An electron loses a quantum of energy.

ChemTalk

Checking Up

1. All of these waves are called electromagnetic waves. Our eyes can only detect a small part of the entire spectrum. All light waves travel at the same speed.
2. Wavelength is the distance between crest to crest of a wave. Actually, the measurement can be taken from any point. For example, distance from the trough to troughs of a wave is also the wavelength.
3. Bohr pictured the electron of hydrogen moving around the nucleus like the planets move around the Sun.
4. Electrons in different orbits have higher energy values than the electron in its ground state.
5. Each element when placed in an excited state will produce a certain line spectrum when the electrons release energy upon return to the ground state. This is sometimes called the fingerprint of the element.

Reflecting on the Activity and the Challenge

In this activity, you've recognized that it is the electrons in an atom that are responsible for the colors of light emitted. Bohr explained the spectrum of light given off by excited hydrogen atoms by hypothesizing that each hydrogen atom's electron was allowed in only certain energy levels. Although Mendeleev knew nothing about electrons or energy levels when he first developed the periodic table (though he did live long enough to know of the electron's discovery by Thomson), 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. Many versions of the periodic table today list the energy levels occupied by each electron in an atom of a given element. You will learn more about this in **Activity 6**. The periodic table is a way of organizing information and knowledge about the atom. Electron structure is a crucial part of that knowledge. Continue to think about how this information could be included in the game you are designing.

Chemistry to Go

1. In this activity you were told that light with greater energy has greater frequency.
 - 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 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 heat is absorbed by the water 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.

Chemistry to Go

- Purple has the greatest energy and red has the least energy.
 - Purple has the highest frequency and red has the lowest frequency.
- An electron would fall a greater distance from E_3 to E_1 and would emit a greater amount of energy than one that only falls to the E_2 level. This energy is in the ultraviolet region.
- E_3 to $E_1 = 1.94 \times 10^{-18}$ J and E_3 to $E_2 = 3.08 \times 10^{-19}$ J.
 Then $(E_3 \text{ to } E_1)/(E_3 \text{ to } E_2) = 1.94 \times 10^{-18} \text{ J} / 3.08 \times 10^{-19} \text{ J} = 6.30$.
- $f = c/\lambda = 2.998 \times 10^8 \text{ m/s} / 389.0 \times 10^{-9} \text{ m} = 7.707 \times 10^{14} \text{ cycles/s}$
 or $7.707 \times 10^{14} \text{ s}^{-1}$
- $E = hf = (6.63 \times 10^{-34} \text{ J}\cdot\text{s})[(2.998 \times 10^8 \text{ m/s}) / (389.0 \times 10^{-9} \text{ m})] = 5.11 \times 10^{-19} \text{ J}$
- $f = c/\lambda = (2.998 \times 10^8 \text{ m/s}) / (1 \times 10^{-1} \text{ m}) = 3 \times 10^9 \text{ s}^{-1}$.
 then, $E = hf = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) \times (3 \times 10^9 \text{ s}^{-1}) = 2 \times 10^{-24} \text{ J}$
 - $(3.02 \times 10^{-19} \text{ J of red light}) / (2 \times 10^{-24} \text{ J of microwave light}) = 2 \times 10^5$.



Active Chemistry The Periodic Table

- b) The red light you observed in the hydrogen spectrum produced a $\lambda = 656.5$ nm. The energy of the red light was 3.02×10^{-19} J. How many times greater is this value when compared to the energy value that you found for the microwave energy?



Inquiring Further

Balmer, Lyman, and Paschen series

In this activity, you learned that the visible light emitted by the hydrogen lamp resulted when an electron moved from a higher energy level to the second lowest energy level in Bohr's scheme. This series of visible spectral lines is said to comprise the **Balmer series**.

What do you think happens when an electron in a hydrogen atom moves

from a higher energy level to the lowest level? This series is known as the **Lyman series** (ultraviolet light).

What if a transition from a higher energy level leads to an electron in the third lowest energy level in Bohr's scheme? How do these energy levels compare with visible light? This series is known as the **Paschen series** (infrared light).

Inquiring Further

Balmer, Lyman, and Paschen series

In the Lyman series, when an electron in a hydrogen atom moves from a higher energy level to the lowest level (E_1), it will give off a greater amount of energy than in the Balmer series' visible light. Ultraviolet light will be emitted. In the Paschen series, when an electron moves from a higher energy level to E_3 , the light given off will be of less energy than in the Balmer series. Infrared light will be emitted.