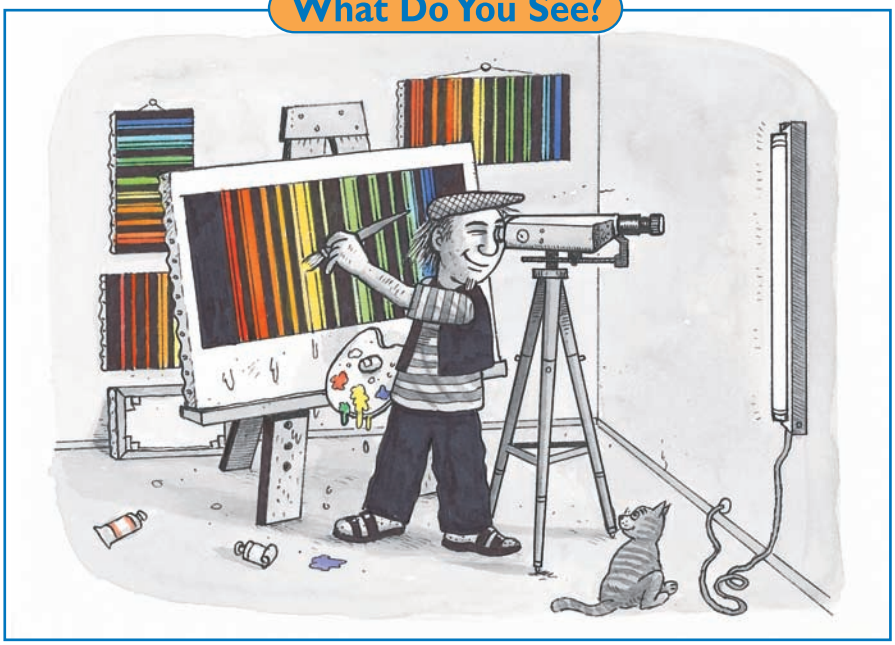
Section 4

Hydrogen Spectra and Bohr's Model of the Hydrogen Atom

What Do You See?



Learning Outcomes

In this section, you will

- **Observe** the spectrum of light emitted by energetic atoms, and describe differences in the spectra of different kinds of atoms.
- **Develop** the Bohr model of a hydrogen atom.
- **Calculate** the energy levels of the electron in the Bohr model of a hydrogen atom.
- **Describe** how the electrons jump from one orbit to another and give off light of a specific wavelength.
- **Cite** evidence for why scientists think the Sun is composed of hydrogen and helium gases.
- **Calculate** the wavelengths of light emitted from transitions of electrons in the Bohr atom.

What Do You Think?

You are very fortunate. With the invention of the electric light, you can engage in the same activities at night as you can during the day. You use fluorescent bulbs, incandescent bulbs, light-emitting diodes, and street lamps to produce light.

- **How is the light from different sources similar and how is it different?**

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

Investigate

You are going to use your observation of colors to find information that will contain clues about the structure of an atom. Your teacher will set up tubes of hydrogen, helium, and neon gases and connect each to a high-voltage power supply. You will identify the gases in each tube by matching the colors you see to a specific set of wavelengths. Finally, you will explore Bohr's model of an atom and how light is emitted by electrons.

Only your teacher will handle the power supply and the tubes of gas. The power supply uses high voltage, which can be dangerous. The tubes of gas are glass and can be broken, leaving sharp edges. Do not look into bright lights (other than the sample of study) or the Sun with the spectrometer.

1. Observe the light of each tube with the naked eye. Then observe the same light with a spectrometer. A spectrometer is a device used to measure a wavelength of light. You teacher will explain how to use the spectrometer.

a) Record your observations in your *Active Physics* log.

b) The spectrometer has a scale and values within for measurement of the wavelengths of light. Record the wavelengths that correspond to each color of light that you are observing in each tube. The wavelengths are measured in nanometers (nm). The prefix *nano* means 10^{-9}. $1\text{ nm} = 10^{-9}\text{ m}$ $(1\text{ nm} = 0.000000001\text{ m})$. The wavelengths of visible light range from about 450 nm (violet) to about 700 nm (red).

2. The light emitted by each of the three tubes is comprised of a distinct pattern of colors. These colors correspond to specific wavelengths of light. Each gas has its own distinct set of wavelengths.

a) Write down three wavelengths of light from one gas tube. Pass this list on to someone else in your group. Have that person identify the name of the gas you chose. Were you all successful?

Scientists try to determine characteristic properties of substances. A characteristic property of a substance is a unique attribute that can be used to identify that substance and distinguish it from other substances. Fingerprints or DNA patterns for humans are characteristic properties. No two people have been found who share an identical fingerprint or identical DNA (other than identical twins). The spectrum of light from a gas is a characteristic property of specific atoms in that gas.

3. When the spectrum of light from the Sun was analyzed, a set of observed wavelengths had the following values: 434 nm, 471 nm, 486 nm, 588 nm, 656 nm, and 668 nm.

a) Which gas on Earth emits three of these wavelengths of light?

b) Which gas on Earth emits the other three wavelengths?

c) From these observations, what can you conclude about the gases that comprise the Sun?

4. In 1913 Niels Bohr, as a young physist, constructed a model of the hydrogen atom that could account for the spectral lines of hydrogen. Bohr's model consists of only a few simple assumptions:

- A proton forms a nucleus and the single electron orbits the proton.
- The electron orbits in a circular path.
- The electron is held in orbit about the proton by Coulomb's law (unlike charges attract).

So far, the Bohr model is a replica of a tiny solar system, where the proton is like a little Sun, and the electron a planet. In this model, Coulomb's law plays the same role as Newton's law of universal gravitation. Bohr's model can be described by standard physics. However, Bohr made the following radical assumptions. He hypothesized that:

- the electron could only exist in specific orbits of specific radii. These specific radii corresponded to specific energies, and
- the radiation from the electron (the spectral lines) only occurs when the electron jumps from one orbit to another orbit.

From this model, Bohr derived an equation for the specific energies at different energy levels.

$$E_n = -13.6\left(\frac{1}{n^2}\right)\text{eV}$$

where $n = 1, 2, 3, 4\ldots$

$E_n = E_1, E_2, E_3, E_4\ldots$

eV = electron volt (a unit of energy, the amount of energy given an electron by a 1-V battery)

The energy of the electron's first orbit

$(n = 1)$ is $E_1 = -13.6\left(\frac{1}{1^2}\right)\text{eV} = -13.6\text{ eV}.$

$E_1(n = 1)$ is also called the ground state.

The energy of the electron in the second orbit

$(n = 2)$ is $E_2 = -13.6\left(\frac{1}{2^2}\right)\text{eV} = -3.4\text{ eV}.$

$E_2(n = 2)$ is also called the first excited state.

a) Calculate the energy of the electron in the 3rd, 4th, 5th, and 6th orbits.

5. When a particle is bound to another particle, the system is defined to have "negative" energy. To liberate one particle from the other, energy must be put into the system to raise the energy of an electron from a negative value up to zero. In hydrogen, an electron in orbit about the proton in the $n = 1$ orbit (the "ground state") has an energy of -13.6 eV. An electron in the ground state would have to be given 13.6 eV to free it.

a) Explain why an electron in the $n = 2$ orbit (the "first excited state," the state just above the ground state) would have to be given 3.4 eV to free it.

b) How much energy would have to be given to an electron in the $n = 3$, 4, 5, and 6 states to free it?

6. The energy required to free an electron from a nucleus is called its ionization energy. The ionization energy of an electron in the ground state is 13.6 eV. The ionization energy of an electron in the $n = 2$ state is 3.4 eV.

a) What are the ionization energies of the electron in the $n = 3$, 4, 5, and 6 states?

7. Niels Bohr explained why light was given off by the hydrogen atom. Light is emitted when the electron begins in one orbit and then jumps to a lower orbit. During that jump the electron loses energy. The energy lost by the electron in making the downward jump becomes the energy of light that is emitted. In this process, energy is conserved. If one object loses energy, something else must gain that same amount of energy for the total energy to remain the same.

Example:

Calculate the energy of light emitted by a hydrogen atom when the electron jumps from the $n = 3$ to the $n = 2$ state.

The energy of electron in $n = 3$ state = -1.51 eV

The energy of electron in $n = 2$ state = -3.40 eV

An electron jumping from $n = 3$ to $n = 2$ would have a change of energy.

$$\begin{aligned}\Delta E &= E_{\text{final}} - E_{\text{initial}} \\ &= E_2 - E_3 \\ &= -3.40 \text{ eV} - (-1.51 \text{ eV}) \\ &= -3.40 \text{ eV} + 1.51 \text{ eV} \\ &= -1.89 \text{ eV}\end{aligned}$$

The electron lost 1.89 eV of energy. Light was created with exactly this 1.89 eV of energy. Light of this energy leaves the atom, and can be observed by the spectrometer as red light. The wavelength of this light is a measure of its 1.89 eV of energy.

a) Calculate the change of energy ΔE when an electron jumps from E_4 to E_2.

b) Calculate the change of energy ΔE when an electron jumps from E_5 to E_2.

c) Calculate the change of energy ΔE when an electron jumps from E_6 to E_2.

Each of these energies corresponds to a specific wavelength of light. These were the colors of light that you observed in the spectrometer for hydrogen.

8. The success of Bohr's model was not limited to a way to calculate the light emitted from hydrogen that was observed with a spectrometer. Bohr also predicted that there would be light emitted from hydrogen that had never been observed.

a) The energy of this light would be due to electron jumps from $n = 2$ to $n = 1$, $n = 3$ to $n = 1$, $n = 4$ to $n = 1$.

Calculate the energies corresponding to these three transitions when the electron jumps from one energy level to another.

b) Other light emitted would be due to electron jumps to the $n = 3$ level. List three transitions to the $n = 3$ level and calculate the energies corresponding to these transitions.

c) Compare the energies of light emitted when electrons jump from higher levels to the $n = 2$ level, to the $n = 1$ level, and to the $n = 3$ level.

Physics Talk

BOHR'S MODEL OF AN ATOM

In the *Investigate*, you observed the **spectral lines** of several gases. These lines of different wavelengths tell you something about the structure of hydrogen, helium, neon, and the other elements as well. The lines are a means by which nature reveals its secrets.

Each element gives off only certain colors. The colors of each element are unique to that element. Niels Bohr began with the Rutherford model of the atom, which had a tiny nucleus in the center. From the Rutherford model he created a model of the hydrogen atom that was able to account for the specific colors of light given off by hydrogen. Hydrogen has only one proton and one electron.

In the Bohr model of an atom,

- the proton is the nucleus and the single electron orbits (revolves) around the nucleus in the same way that the planets orbit the Sun.
- the electron is able to move in a circle about the proton because of the attractive Coulomb force between the positively charged proton and the negatively charged electron.
- the electron can only be found in specific orbits with specific radii. Each orbit (path of an electron) has a specific energy associated with it. The energy levels of the orbits follow a simple pattern:

$$E_n = -13.6\left(\frac{1}{n^2}\right)\text{eV}$$

where n = 1,2,3,4...

$$E_n = E_1, E_2, E_3, E_4 \ldots$$

eV = electron volt (a unit of energy)

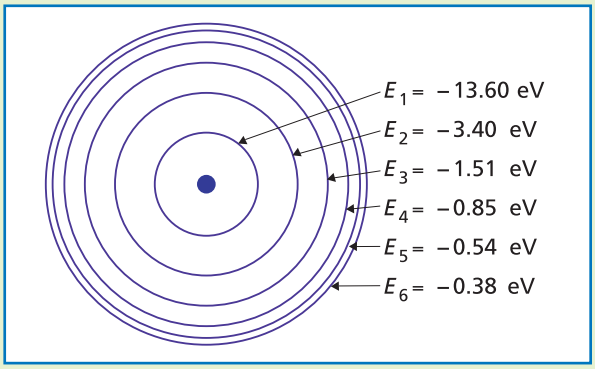
The negative sign in the equation indicates that the energy of the electron is negative. When a particle is bound to another particle, the system is said to have "negative" energy, because to free one particle from the other, and therefore raise its energy up to zero, energy must be put into the system. In hydrogen, an electron in orbit about the proton in the n=1 orbit (the "ground state") has an energy of -13.6 eV. An electron in the ground state would have to be given 13.6 eV to free it. The energy required to free an electron from a nucleus is called its ionization energy. The **ionization energy** of an electron in the ground state is 13.6 eV.

Physics Words

spectral lines: the lines of different colors that tell something about the structure of an element.

ionization energy: the energy required to free an electron from its energy level.

Using the equation, $E_n = -13.6\left(1/n^2\right)\text{eV}$ with $n = 2,\ 3,\ 4,\ 5,$ and 6 yields the corresponding energies for excited states of the electron as shown in the sketch of the possible orbits for the hydrogen electron.



According to Bohr's model, these are the only allowable orbits for the electron. When the electron orbits the nucleus, it is restricted to these orbits and these energies. It cannot orbit at any distance from the nucleus, but only in these specific orbits.

The puzzle that Bohr was seeking to solve was finding the relationship between these energy levels and the wavelengths of light emitted in the hydrogen spectra. Bohr proposed that light is emitted when the electron jumps from a higher orbit to a lower orbit. For example, when the electron jumps from the $n = 3$ orbit to the $n = 2$ orbit, red light is emitted. The energy of the emitted red light is exactly equal to the change in energy of the electron. When the electron jumps from the $n = 4$ orbit to the $n = 2$ orbit, green light is emitted. The energy of the emitted green light is exactly equal to the change in energy of the electron.

Bohr's model and calculations correctly predicted the energy, wavelengths, and color of the light you observed from the hydrogen spectra with the transitions $E_3 \rightarrow E_2$, $E_4 \rightarrow E_2$, $E_5 \rightarrow E_2$, and $E_6 \rightarrow E_2$.

Bohr's model did more than this. Bohr also predicted that light would be emitted when the electron jumps from

$E_2 \rightarrow E_1$, $E_3 \rightarrow E_1$, $E_4 \rightarrow E_1$, $E_5 \rightarrow E_1$ and $E_6 \rightarrow E_1$.

His calculations indicated that these energies of light are not visible to the human eye. Visible light is only one part of the electromagnetic spectrum. If an ultraviolet detector is used, these exact wavelengths are observed. The visible light rays of the electromagnetic spectrum of hydrogen are called the **Balmer series**. The ultraviolet light of the electromagnetic spectrum rays are called the **Lyman series.**

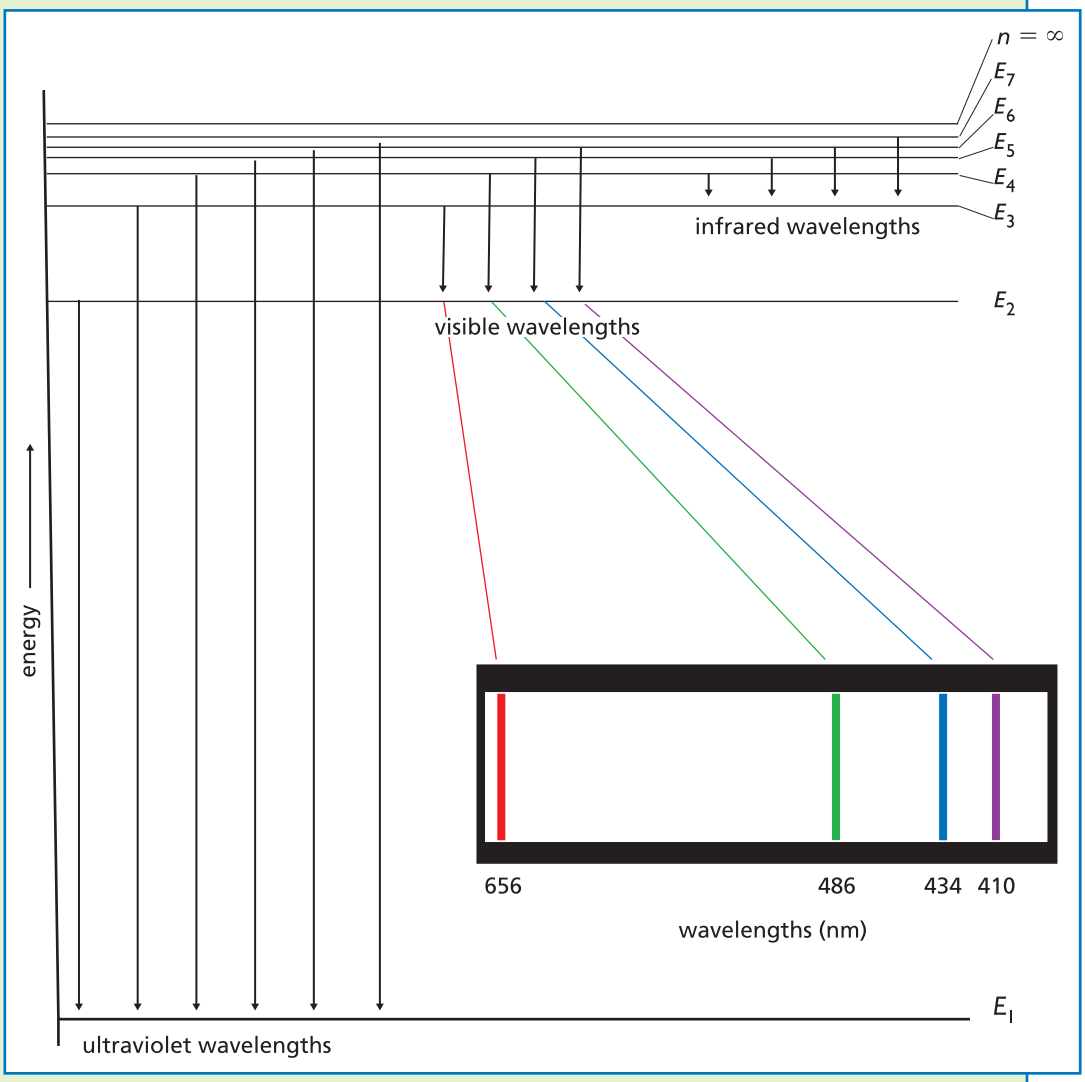
Light should also be emitted when the electron drops from $E_4 \rightarrow E_3$, $E_5 \rightarrow E_3$, and $E_6 \rightarrow E_3$. These energies can also be calculated. The light with these energies is also not visible to the human eye. If an infrared detector is used, these additional wavelengths are observed. These infrared light rays are called the **Paschen series.**

Physics Words

Balmer series: the visible light rays of the hydrogen electromagnetic spectrum.

Lyman series: the ultraviolet light rays of the hydrogen electromagnetic spectrum.

Paschen series: the infrared light rays of the hydrogen electromagnetic spectrum.



A good new theory should be able to explain whatever the old theory could explain. The new theory should also be able to explain something that the old theory could not explain. Finally, the new theory should be able to make a prediction of something that nobody had thought of previously. If that prediction turns out to be true, then you can sense the power of the theory. The Balmer and Paschen series had been observed before Bohr's theory. The Lyman series had never been observed. Bohr predicted the wavelengths of the Lyman series and then they were observed with the predicted wavelengths. Other spectra corresponding to transitions to the $n = 4$ state were also found later.

Energy can be provided to free an electron by having the electron absorb light. When the electron absorbs light, it jumps up to a higher energy state, the reverse of Bohr's downward transitions that emit light.

Unless the electron is completely removed from the atom, the only light that can be absorbed (or emitted) by the electron is the light that has just the right energy and just the right wavelength to get the electron to another fixed orbit. If the light has a little less energy or a bit more energy, it has no effect. If the electron absorbs enough light, it can be freed from the attractive force of the nucleus. So a neutral atom loses its electron to become a **positive ion**. This transition process is called **ionization.**

Physics Words

positive ion: an ion created when a neutral atom loses its electron.

ionization: the process in which a neutral atom becomes an ion.

Discovery of Helium

When viewing the Sun, the following set of wavelengths can be observed: 410 nm, 434 nm, 471 nm, 486 nm, 588 nm, 656 nm, and 668 nm. From your observations in the *Investigate*, you concluded that these values match the wavelengths emitted by hydrogen and helium. This conclusion led you to the conclusion that the Sun is comprised of hydrogen and helium.

The history of these values, however, is a bit more interesting. The set of values corresponding to hydrogen (410, 434, 486, and 656 nm) were known from the lab. Nobody had ever observed other wavelengths (471, 588, 668 nm) in a lab. In 1868, Pierre Janssen observed one new yellow line from the Sun during a total eclipse in India. J. Norman Lockyear interpreted this yellow line as being evidence of a new element. This set of wavelengths did not correspond to any gas on Earth and so this unknown gas of the Sun was named "helium" after Helios, the Greek god of the Sun. Years later, in 1895, helium gas was discovered on Earth by William Ramsey of Scotland and independently by Per Cleve of Sweden.

Checking Up

1. Using Bohr's model of an atom, explain why an electron remains in its orbit.
2. Why are different wavelengths of light found in a hydrogen spectrum?
3. If the energy of an electron in the ground state is -13.6 eV, would it able to jump to a higher orbit at -3.4 eV? Explain.

Active Physics Plus

+Math	+Depth	+Concepts	+Exploration
◆◆		◆	

Calculations Involving the Hydrogen Atom

The Bohr model has a single electron of hydrogen orbiting a single proton nucleus of hydrogen. The force that holds an electron in orbit is the Coulomb electrostatic force between two unlike charged objects.

a) Using Coulomb's equation:

$$F = k\frac{q_1 q_2}{d^2}$$

calculate the force between the proton and electron (each has a charge of $1.6\times10^{-19}\,\text{C}$. The distance between them is $5\times10^{-10}\,\text{m}$.

As you will see in the next section, in 1905 Albert Einstein proposed the relationship between the energy of light and its frequency in the following equation:

$$E = hf$$

where E is the energy of light,
h is Planck's constant, $\left(6.63\times10^{-34}\ \text{J}\bullet\text{s}\right)$, and
f is the frequency of light.

To know the frequency is to know the wavelength. Wavelengths of light (λ) can be found from the wave equation

$$c = f\lambda$$

where c is the speed of light $(3.0\times10^{8}\ \text{m/s})$.

Einstein's proposed equation for the energy of light $E = hf$ can be combined with Bohr's calculation of the energy given to the light when the electron jumps from E_3 to E_2. You can use this combination to predict the spectrum of emitted light. You must also convert from the energy unit of electron volts to joules if you wish the wavelength to be in meters. To calculate the wavelength using the energy lost by the electron during the jump in electron volts requires using Planck's constant (h) in electron volts, or $h = 4.1\times10^{-15}\,\text{eV}\bullet\text{s}$.

Combining the two equations:
$\Delta E = hf$ and $c = f\lambda$ gives

$$\lambda = \frac{h\bullet c}{|\Delta E|}$$

where $h\bullet c$ is

$(4.1 \times 10^{-15}\ \text{eV}\bullet\text{s})(3 \times 10^{8}\ \text{m/s})$

Therefore,

$$\lambda = \frac{1.24 \times 10^{-6}\ (\text{m})(\text{eV})}{|\Delta E|}$$

ΔE is the energy change in electron volts (eV).

Example:

When the electron jumps from E_3 to E_2, the change in energy is 1.89 eV.

a) Calculate the corresponding wavelength of light.

$$\lambda = \frac{hc}{|\Delta E|}$$

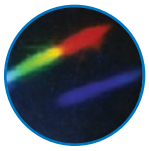
$$\lambda = \frac{1.24 \times 10^{-6}(\text{m})(\text{eV})}{|\Delta E|}$$

$$\lambda = \frac{1.24 \times 10^{-6}(\text{m})(\text{eV})}{|1.89\text{eV}|}$$

$$\lambda = 654\ \text{nm}$$

The wavelength determined by the above calculations equals the measured wavelength of the red line of hydrogen.

1. Calculate the wavelengths of light when the electron jumps from E_4 to E_2, E_5 to E_2, and E_6 to E_2.
2. How do these values compare with the ones you found in your observations of the hydrogen spectra?



What Do You Think Now?

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

- **How is the light from different sources similar and how is it different?**

You observed the difference in colors of light emitted by different gases. How does the wavelength of those colors relate to the structure of an atom?

Physics

Essential Questions

What does it mean?

The spectra of light for a given element (hydrogen, helium, neon) is said to be a fingerprint for that element. What is the spectra of hydrogen and why can the spectra be used as evidence that a gas is hydrogen, not neon?

How do you know?

What evidence do you have from the *Investigate* that the spectra of different gases are different?

Why do you believe?

Connects with Other Physics Content	Fits with Big Ideas in Science	Meets Physics Requirements
Optics	Symmetry – laws of physics are the same everywhere	* Good, clear, explanation, no more complex than necessary

* Physics tries to use the same explanation to explain observations anywhere on Earth and anywhere in the universe. You observe certain spectral lines from the Sun that are identical to the lines that are emitted from hydrogen and helium gas on Earth. Why do you believe that the Sun is composed of hydrogen and helium?

Why should you care?

The light from different gases has different colors and wavelengths. This light reveals something about the structure of the atom. How can you use this revelation to both generate interest in museum visitors and help them understand about Bohr's model of the atom?

Reflecting on the Section and the Challenge

The Bohr atom has electrons orbiting in "special" orbits surrounding the nucleus. Light is emitted when electrons jump from a higher-energy orbit to a lower orbit. An electron that absorbs energy can jump from a lower orbit to a higher-energy orbit. The wavelengths of light can be calculated, observed, and measured. The values from Bohr's theory and your observations from hydrogen are almost exactly equal. In your museum exhibit, you may try to show the Bohr model of the atom and the emission of light as electrons jump from one energy level to another.

You may also wish to show how an atom becomes ionized when the electron absorbs enough energy to free it. Finally, you may choose to show that invisible light in the ultraviolet and infrared regions is also emitted. Electron jumps and emitted light can be an interactive museum display. Will your display create an immediate interest? Provide a means by which the museum visitor will want to stop and see what is going on with hydrogen.

Physics to Go

1. Light of greater energy has a higher frequency. In the hydrogen spectrum, which visible line has the greatest energy? Which transition does this line correspond to?
2. Compare the energy of the light emitted from the electron jump from $n = 3$ to $n = 2$ to the light emitted from $n = 5$ to $n = 2$.
3. Given that the speed of light equals 3.0×10^8 m/s and the wavelength of light is 389 nm $\left(389 \times 10^{-9}\text{ m}\right)$, calculate the frequency of the light.
4. Calculate the energies of each Bohr orbit using the equation $E = -13.6\text{ eV}\left(1/n^2\right)$ for $n = 1, 2, 3, 4,$ and 5.
5. Make a scale diagram showing the energies of each Bohr orbit as a vertical number line which goes from -13.6 eV to 0 eV.
6. The hydrogen spectrum is said to be a "fingerprint" for hydrogen. Explain why this is a useful metaphor (comparison).
7. Why can't light of 500 nm be given off from hydrogen?
8. Electrons can jump from the $n = 4$ state directly to the $n = 3$ or $n = 2$ or $n = 1$ states. Similarly, there are multiple jumps from $n = 3$. How many different wavelengths of light can be given off from electrons that begin in the $n = 4$ orbit?
9. Suppose you could build an atom with the following energy levels for its electron:
   a) $E = -(10\text{eV})(1/n)$ where $n = 1, 2, 3,$ and 4. Draw a diagram of the electron's energy levels. Describe the spectrum, that is, what are the energies of the spectral lines?
   b) Repeat 9.*a*) if $E = -(10\text{ eV})(1/n)^2$.
10. A candle flame is mostly yellow; the flame from a welder's torch is bright blue. Why are these colors different?
11. ***Preparing for the Chapter Challenge***

    How could the electron transitions be creatively displayed in a museum exhibit? Describe a way in which the display could be interactive.