

## Activity 5

# The Electronic Behavior of Atoms

### What Do You See?



### 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.
- 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.

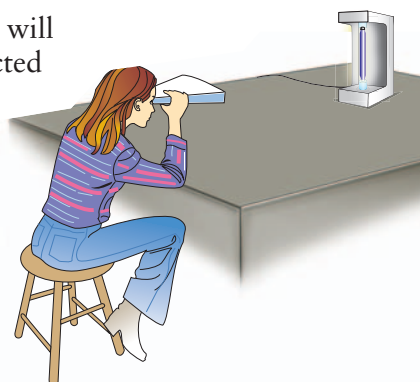
- **How is the color produced in a neon sign?**

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 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.

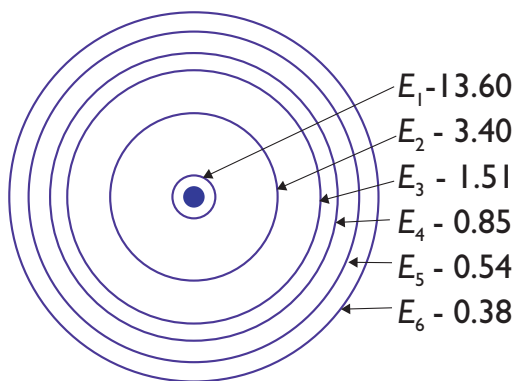
For this demonstration, your teacher will set up a tube of hydrogen gas connected to a high-voltage power supply. This light can be viewed through a spectroscope or a diffraction grating, as shown in the diagram. When the slit at the end of the spectroscope is aimed towards the light, the colors of the 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 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  $E_1$       c)  $E_4$  to  $E_2$
  - b)  $E_3$  to  $E_2$       d)  $E_5$  to  $E_2$
5. Picture the electron 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 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, 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  $E_6$ ?

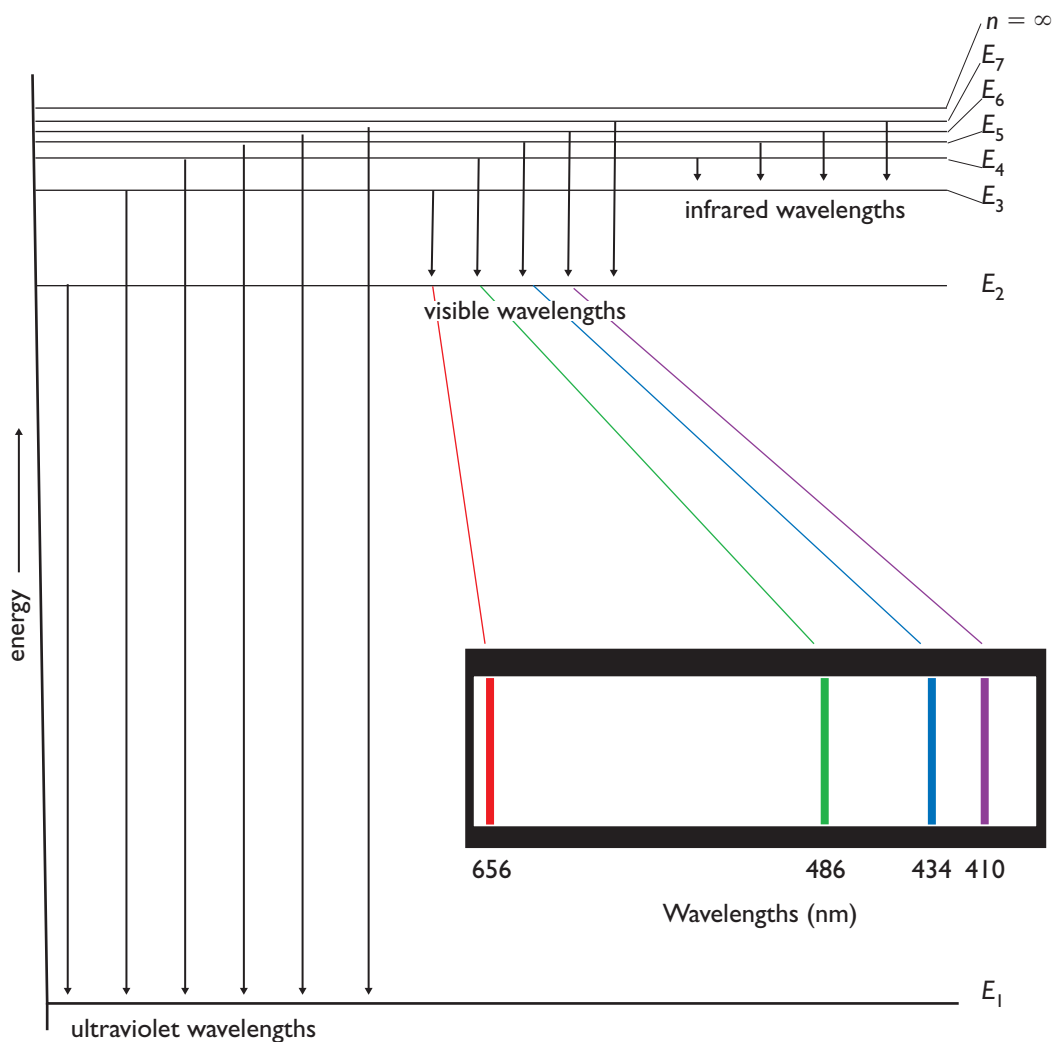


- c) An electron can also jump from  $E_5$  to  $E_3$  to  $E_2$ . Calculate the two energies of light that would be emitted in these two jumps.
6. Each energy of light corresponds to a specific color. The four colors of light that you observed in the hydrogen spectrum corresponded to four specific energies 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 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.





## ChemTalk

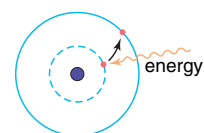
### BOHR'S MODEL OF AN ATOM

#### Electron Orbits

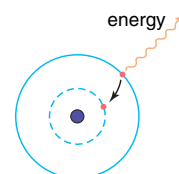
Niels Bohr was a brilliant Danish physicist. He proposed a “planetary” model of the atom. He theorized that electrons travel in nearly circular paths, called **orbits**, around the nucleus. Each electron orbit has a definite amount of energy. The farther away the electron is from the nucleus, the greater 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.



Niels Bohr



(a) An electron gains a quantum of energy.

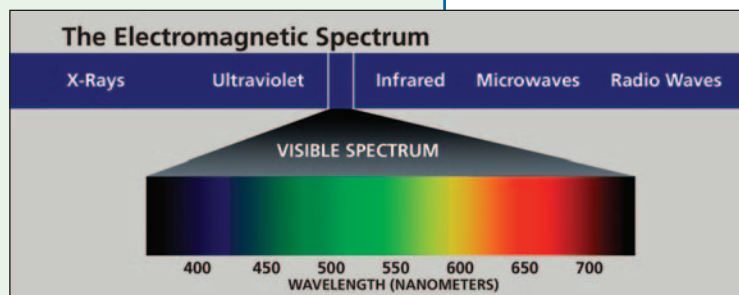


(b) An electron loses a quantum of energy.

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





### 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).

**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.

**photon:** a fixed 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  $\lambda$ , the Greek letter *lambda*. The speed of electromagnetic radiation is constant and it is called the speed of light,  $c$ . Its value is  $2.998 \times 10^8$  m/s. This is often approximated to  $3.00 \times 10^8$  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, if the wavelength is 434.2 nm, then the frequency is:

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

As you go across the electromagnetic spectrum 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 fixed 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  $6.63 \times 10^{-34}$  J·s and  $f$  is the frequency.

The corresponding energy of the red light above would be:

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

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  $3.03 \times 10^{-19}$  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.

$$E = hf$$

$$3.03 \times 10^{-19} \text{ J} = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) (f)$$

$$f = 4.57 \times 10^{14} \text{ Hz}$$

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

$$c = f\lambda$$

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

$$\lambda = 6.56 \times 10^{-7} \text{ m}$$

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.

**The Photoelectric Effect**

Light has wave-like properties like wavelength and frequency. These wave-like properties were seen in the activity as you observed the spectrum of hydrogen. Light also has particle-like properties like momentum and can be seen in collisions with matter. The photoelectric effect is a collision between 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.

**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?



## What Do You Think Now?

At the beginning of the activity you were asked:

- 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
<i>Describe what you saw through the spectroscope or the diffraction grating when the tube of hydrogen gas was subjected to a large voltage.</i>	<i>Explain how the light is created in the hydrogen atom. Why are only certain colors emitted?</i>	<i>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 measure the different electron energy levels?</i>

### 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 Activity and the Challenge

In this activity 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 activity 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  $7.707 \times 10^{14}$  Hz. Show how this frequency value was calculated.
5. Show that the wavelength of 389.0 nm has an energy of  $5.11 \times 10^{-19}$  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  $\lambda$  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  $3.03 \times 10^{-19}$  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).

# Chapter mini Challenge



Your challenge for this chapter is to create a game that will both entertain and teach people about the periodic table. Since you have completed five activities, this is a good time to give your game a *first try*.

Your *mini-challenge* is to create or adapt a game that has chemistry content as part of its structure. You will only be providing the class with a two-minute explanation of the game that you are developing. Within your two minutes, you should explain some of the rules of the game and some of the chemistry content that a player will get to learn from the game.

## Chemistry Content

Your team should first review the chemistry content that may be included in this periodic-table game challenge.

**Activity 1:** You learned about organization. The periodic table is a way in which your knowledge about the properties of matter can be organized.

**Activity 2:** You investigated the properties of some elements. These properties included the ability to conduct electricity, whether the element reacted with HCl, and whether the material was a metal or nonmetal. You also attempted to organize the elements in terms of these properties.

**Activity 3:** You learned how to compare and contrast elements and compounds. You were also able to determine the relative masses of elements. The mass of elements is another property by which elements can be organized.

**Activity 4:** The game of battleship helped you learn about the structure of the atom. You learned that the nucleus takes up a very, very small part of the atom. That tiny space holds all of the positive charge and almost all of the mass of the atom. You also learned about the discovery of the electron and its place in the atom.

**Activity 5:** You observed the specific colors of the light given off by hydrogen gas. By investigating these colors (and energies) of light, you learned how Bohr developed a model of the atom. You learned that electrons jumping from a higher energy orbit to a lower energy orbit emitted light of a specific energy and color.

## Criteria

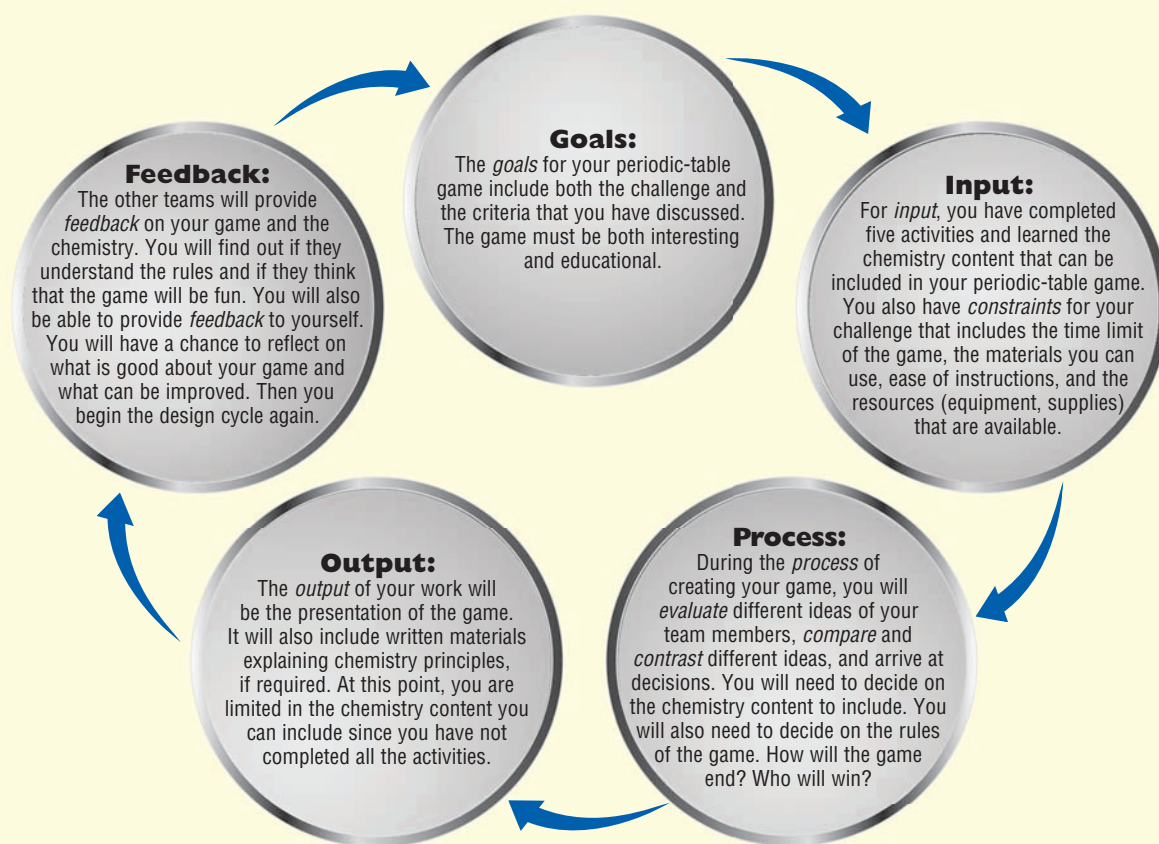
Your team should also review the criteria that you discussed on the first day of this chapter. Your periodic-table game should be creative and fun. It should demonstrate an understanding of chemistry concepts and help teach these concepts.

## Preparing for Your **mini** Challenge

Your team should begin the process of creating a game by brainstorming ideas about games that you enjoy that may be adapted for this purpose. You should then discuss how you could include chemistry concepts from at least two of the activities in the game. You can then decide on the rules of the game.

# Engineering/Technology Design

The engineering-design process involves a number of distinct steps. In creating a game, you are involved in a design process. Note how your efforts follow the design model described.



**Remember:** This is a *first try*. The feedback that you get from this brief presentation of your game will be extremely valuable when you prepare for the full periodic-table game that you will develop at the end of the chapter. At that point, you will begin the entire design cycle again. That's why it is called a cycle. You keep going through it over and over to improve your design or, in this case, the periodic-table game.