Atomic Emission Spectra (Teacher Demonstration)

When a high voltage current is passed through a glass tube that contains hydrogen gas at low pressure, the gas glows with a pinkish-purple colour. If you observe this gas discharge tube through a spectroscope or a diffraction grating, you can see distinct, coloured lines. This type of spectrum is called a line spectrum or an emission spectrum. Chemists first observed spectra like this in the mid-1800s.

Question
Can the Rutherford model of the atom explain the emission spectra of elements?

Predictions
The Rutherford model of the atom described electrons as tiny masses in motion around a nucleus. According to this model, excited atoms should emit a continuous spectrum of light energy as their electrons give off radiation. In fact, according to this model, atoms should continually radiate energy whether they are excited or not. Read through the procedure and predict whether your observations will support or contradict the Rutherford model. Give reasons for your answer.

Safety Precautions

- A very high voltage is required to operate the gas discharge tubes. Do not come into contact with the source while viewing the tubes.
- Do not work with the gas discharge tubes yourself. Your teacher will demonstrate how it works.

Materials
- gas discharge tube apparatus (teacher use only)
- low pressure hydrogen gas tube and tubes for other available elements
- spectroscope or diffraction grating

Procedure

1. Practise using the spectroscope or diffraction grating by observing an incandescent light bulb. Point the slit of the spectroscope toward the bulb and move the spectroscope until you can clearly see the spectrum.

2. Record the appearance of the spectrum from the incandescent bulb.

3. Your teacher will set up the gas discharge tube apparatus to demonstrate the emission spectra of hydrogen and other elements.

4. Observe the hydrogen discharge tube, and note its colour when high-voltage current is applied to it.

5. With the lights dimmed, examine the hydrogen emission spectrum with the aid of a spectroscope or diffraction grating. Make a careful sketch to record your observations. Note in particular the colours you see. Estimate their wavelengths. (Use the figure below to help you estimate the wavelengths.)

6. Follow steps 4 and 5 for the emission spectra produced by other gas discharge tubes that your teacher has available.
Analysis

1. According to Rutherford’s model of the atom and nineteenth-century physics, electrons could emit or absorb any quantity of energy.
   (a) Compare the spectrum you observed for the incandescent bulb to the spectrum that Rutherford’s model predicts for hydrogen.
   (b) Compare the spectrum you observed for hydrogen to the spectrum you observed for the incandescent bulb. Do your observations seem to support or contradict Rutherford’s model? Explain your answer.
   (c) How do you think Rutherford’s model would have to be modified in order to account for your observations?

2. How do you think the lines you observed in the hydrogen emission spectrum relate to the energy of the electrons in a hydrogen atom? Include a diagram in your answer.

3. Compare the hydrogen emission spectrum to the other spectra you observed. Why do you think emission spectra are different for different elements?

4. Why do you think that scientists initially focused their attention on the hydrogen emission spectrum rather than the emission spectrum of another element?

5. Each element has a characteristic emission spectrum. Why is it only possible to observe the emission spectra of selected elements in a high school laboratory?

Conclusion

6. Reread your prediction and state whether it was accurate or not. Explain your answer in detail.

Applications

7. The emission spectrum of each element is characteristic of that element. In other words, it is a “fingerprint” that can be used to identify the element.
   (a) Suggest reasons to explain why each element has a distinct and characteristic emission spectrum.
   (b) Helium was discovered in the Sun before it was discovered on Earth. How do you think this was possible? Use print resources or the Internet to help you answer the question.

8. Examine the spectrum below. It is the absorption spectrum for sodium. Absorption spectra are produced when atoms absorb light of specific wavelengths.
   (a) If you examined the emission spectrum for sodium in the laboratory, compare it to the absorption spectrum for sodium. (If you did not, look for a picture of the emission spectrum for sodium in a book or on the Internet.)
   (b) In general, how do you think emission spectra compare to absorption spectra?
   (c) Instruments based on absorption spectra are used extensively in analytical chemistry. One such instrument is an atomic absorption spectrophotometer. Use print and electronic resources to find out how the instrument works, and for what applications it is used.
The Bohr Model of the Atom

Scientists of the nineteenth century lacked the concepts necessary to explain line spectra. Even in the first decade of the twentieth century, a suitable explanation proved elusive. This changed in 1913 when Niels Bohr, a Danish physicist and student of Rutherford, proposed a new model for the hydrogen atom. This model retained some of the features of Rutherford’s model. More importantly, it was able to explain the line spectrum for hydrogen because it incorporated several new ideas about energy. As you can see in Figure 3.8, Bohr’s atomic model pictures electrons in orbit around a central nucleus. Unlike Rutherford’s model, however, in which electrons may move anywhere within the volume of space around the nucleus, Bohr’s model imposes certain restrictions.

On what basis could Bohr make these claims? In 1900, a physicist named Max Planck proposed an idea that was so revolutionary that he himself was unwilling to accept its implications. Planck suggested that matter, at the atomic level, can absorb or emit only discrete quantities of energy. Each of these specific quantities is called a quantum of energy. In other words, Planck said that the energy of an atom is quantized. Something that is quantized can exist only in certain discrete amounts. It is not continuous. In this sense, the rungs of a ladder are quantized, while the smooth slope of a ramp is not. Unlike ladders and ramps, however, a quantum is an extremely small “packet” of energy.

Although Planck said that the energy of matter is quantized, he continued to describe energy as travelling in the form of waves. He was, in fact, unwilling to consider that energy might have particle-like properties. In 1905, however, Albert Einstein was prepared to make just such an assertion. Light, according to Einstein, is also quantized. It occurs as quanta of electromagnetic energy that have particle-like properties. These particle-like “packets” of energy were later called photons. In Einstein’s view, light (and, by extension, all electromagnetic energy) travels in the form of photons of energy. Light is emitted as photons of energy, and light is absorbed as photons of energy.

How Bohr’s Atomic Model Explains the Spectrum for Atomic Hydrogen

If you re-read the key points of Bohr’s atomic model above, you can see that he is applying the new ideas of quanta in his model. Bohr proposed that the energy that is emitted and absorbed by an atom must have specific values. The change in energy when an electron moves to higher or lower energy levels is not continuous. It is, rather, quantized.
Look again at the line spectrum for hydrogen shown in Figure 3.7. The energy that is associated with the coloured lines in this spectrum corresponds to the change in energy of an electron as it moves to higher or lower energy levels. For example, when a hydrogen atom is exposed to electrical current, or to another form of electromagnetic energy, its electron absorbs photons of energy. The atom is now said to be in an excited state. Another common way for atoms to become excited is through atomic collisions. For example, when two hydrogen atoms collide, some of the kinetic energy from one atom is transferred to the other atom. The electron of the second hydrogen atom absorbs this energy and is excited to a higher energy level.

When the electron of a hydrogen atom that has been excited to the third energy level falls to the second energy level, it emits light of certain energy. Specifically, an electron that makes a transition from the third energy level to the second energy level emits a photon of red light with a wavelength of 656 nm. Because line spectra result when atoms in an excited state emit photons as they fall to a lower energy level, these spectra are also called emission spectra.

Figure 3.9 shows the energy transitions that are responsible for the coloured lines in hydrogen’s emission spectrum. Notice the use of the symbol \( n \) to designate the allowed energy levels for the hydrogen atom: \( n = 1, n = 2, \) and so on. This symbol, \( n \), represents a positive integer (such as 1, 2, or 3), and is called a quantum number. You will learn more about the significance of quantum numbers in section 3.2.

\[\begin{align*}
\text{Note:} & \text{ Orbits are not drawn to scale.} \\
\text{Note:} & \text{ In its unexcited state, hydrogen’s electron is in the energy level closest to the nucleus: } n = 1. \text{ This is the lowest-possible energy level, representing a state of greatest stability for the hydrogen atom.}
\end{align*}\]

**Figure 3.9** How the Bohr model explains the coloured lines in hydrogen’s emission spectrum. When an excited electron falls from a higher energy level to a lower energy level (shown by the downward-pointing arrows), it emits a photon with a specific wavelength that corresponds to one of the coloured lines in the spectrum.
Another kind of spectrum that is related to an atom’s emission spectrum is an absorption spectrum. This kind of spectrum results when electrons of atoms absorb photons of certain wavelengths, and so are excited from lower energy levels to higher energy levels. Figure 3.10 shows the absorption spectrum for hydrogen. As you can see, discrete, dark lines appear in the precise locations (wavelengths) as those of the coloured lines in hydrogen’s emission spectrum.

Figure 3.10 When light passes through gaseous hydrogen atoms, the electrons in the hydrogen atoms absorb photons of red, green, blue, and purple light at specific wavelengths. When the electrons return to lower energy levels, they re-emit photons with these same wavelengths. However, most of the photons are emitted in a different direction from the one in which the electrons absorbed the light. As a result, dark lines appear in the spectrum where photons of those particular wavelengths are absent.

The Successes and Limitations of the Bohr Atomic Model

Bohr’s realization that the atom’s energy is quantized—that electrons are restricted to specific energy levels (orbits)—was an astounding achievement. As you have seen, this model successfully predicted the coloured lines in the visible-light portion of hydrogen’s emission spectrum. It also successfully predicted other lines, shown in Figure 3.11, that earlier chemists had discovered in the ultraviolet and infrared portions of hydrogen’s emission spectrum.

Figure 3.11 The ultraviolet series of lines appear as a result of electrons falling from higher energy levels to \( n = 1 \). The infrared lines appear as a result of electrons falling from higher energy levels to \( n = 3 \).