

# Atomic Spectra

## Light, Energy, and Electron Structure



### Introduction

When a pure atomic gas such as hydrogen or helium is subjected to a high-voltage electrical discharge, light is produced and the gas glows. When this light is observed through a diffraction grating, a series of bright colored lines is observed. The series of bright lines is called an atomic emission spectrum and is unique to each element. In this activity, a spectroscope will be used to view the “bright line” emission spectra of different elements and determine their wavelengths. A spectroscope contains a diffraction grating that separates light into its component wavelengths.

### Concepts

- Atomic emission spectrum
- Quantization of energy
- Electron energy levels
- Electron transitions

### Materials

Colored pencils or crayons

Gas discharge (spectrum) tubes, such as hydrogen, helium, mercury, and neon

Light sources, such as incandescent and fluorescent lightbulbs, street lamps, neon signs, and novelty “neon” lamps

Power supply

Ring clamps, 2 (optional)

Ring stand (optional)

Spectroscope

### Safety Precautions

*Spectrum tubes operate at very high voltages and can produce a large electric shock. Do not touch the ends of the tube when the power supply is on. Do not touch the contacts on the transformer when the power is on. Make sure the power supply is turned off before inserting or removing the spectrum tube. To extend the life of the tubes, do not leave the tubes on for more than 30–45 sec at a time—cycle the power on and off as needed to complete the observations. Spectrum tubes may get very hot. Never touch a spectrum tube when the power is on. After turning off the power, allow the tube to cool before removing it from the power supply.*

### Procedure

1. Using a spectroscope, observe the continuous “rainbow” spectrum from an incandescent lightbulb.
2. Observe the colors of light in the visible spectrum and the wavelength range for each band. Sketch the spectrum of white light using colored pencils in appropriate wavelength boxes. Note that the units of wavelength on the spectroscope are nanometers ( $1 \text{ nm} = 10^{-9} \text{ m}$ ).
3. (Optional) For optimum viewing of the emission spectra of gas discharge tubes, stabilize the spectroscope on a ring stand. Set up a ring stand in front of the power supply and attach one ring clamp. Place the spectroscope on the ring clamp and adjust the height of the ring clamp so that the eyepiece on the spectroscope is approximately level with the middle of the gas discharge tube. Attach a second ring clamp on top of the spectroscope so that it will be held firmly in position without moving.
4. With the power OFF, ask the instructor to insert the hydrogen spectrum tube between the contacts on the power supply.
5. Move the power supply so that the spectrum tube is about 3–5 cm away from the spectroscope.
6. Turn on the power on the power supply and observe the atomic emission spectrum of hydrogen.

7. Turn OFF the power supply. Record the following information for the emission spectrum of hydrogen: the number of lines, their colors, and their approximate wavelengths.
8. Using colored pencils, sketch the atomic spectrum of hydrogen in appropriate wavelength boxes. Turn the power supply on and off, as necessary, to complete the observations in steps 7 and 8.
9. *Check to make sure the power supply is off*, then remove the hydrogen spectrum tube and insert a mercury spectrum tube.
10. Observe and note the atomic emission spectrum of mercury.
11. Turn OFF the power supply. Record the following information for the emission spectrum of mercury: the number of lines, their colors, and their approximate wavelengths.
12. Using colored pencils, sketch the atomic spectrum of mercury in appropriate wavelength boxes. Cycle the power supply on and off, as necessary, to complete the observations in steps 11 and 12.
13. Repeat steps 9–12 for any other gas spectrum tubes that are available.
14. Using a spectroscope, observe the spectrum of visible light obtained from a fluorescent light. What kind of spectrum is produced? If any bright lines are present, record the number of lines, their colors, and their approximate wavelengths.
15. (*Optional*) Using a spectroscope, observe the emission spectrum of other light sources, such as neon signs, street lights, headlights, novelty lamps, etc. What kind of spectra are produced? If any bright lines are present, record the number of lines, their colors, and their approximate wavelengths.

## Disposal

None required. Carefully store all equipment for future use.

## Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

### ***Unifying Concepts and Processes: Grades K–12***

Evidence, models, and explanation

Constancy, change, and measurement

### ***Content Standards: Grades 9–12***

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, structure of atoms, structure and properties of matter, interactions of energy and matter

Content Standard G: History and Nature of Science, nature of scientific knowledge, historical perspectives

## Tips

- Using a spectroscope is an interesting and fun activity for students. Many teachers are discouraged from doing an “Atomic Spectra” lab because of the expense involved in purchasing power supplies. Many inexpensive alternatives to spectrum tubes are readily available, however, and may be used. Inexpensive “neon” novelty lamps available at many party stores and discount stores are excellent sources of bright line emission spectra of mercury (blue) and neon (red).
- Check with the physics teacher at your school for power supplies, spectrum tubes, and spectroscopes.
- The “Spectrum Analysis Chart” available from Flinn Scientific (Catalog No. AP8676) is a poster-size, full-color chart that shows the bright line emission spectra of ten elements.
- Take advantage of Internet resources to locate full-color reference spectra for many different gases. The following Web site (accessed January 2010) showed a wide selection of atomic spectra.  
<http://astro.u-strasbg.fr/~koppen/discharge>
- The line emission spectrum of an element can be compared to the UPC code found on many packaged items. The UPC code consists of a series of lines—the pattern of lines, not any individual line, is characteristic of the item in question.
- Show students the colors of light reflected from a CD to illustrate what a diffraction grating is and what it does.

- This activity is best performed in a partially darkened room to avoid interference due to stray light. Dim the lights or close the shades if possible. Do NOT work in a totally darkened environment, however! It is not safe and students will not be able to read the wavelength scales on their spectrosopes.

## Discussion

The phenomenon of atomic spectra has been known since the mid-1800s. Their cause, however, remained unexplained until the structure of the atom and, in particular, its electronic structure, was solved. Rutherford's discovery of the nucleus of the atom in 1911 answered many questions concerning the structure of the atom. It also raised many new questions, including, where are the electrons?

In 1913 Niels Bohr proposed a model of electron structure that would explain the phenomenon of atomic spectra. According to Bohr's model, an electron is restricted to certain specific orbits around the nucleus of the atom. These orbits differ in their distance from the nucleus and in their energy levels. Electrons that are closer to the nucleus are lower in energy than electrons that are farther away from the nucleus. This idea is called the *quantization of energy*—electrons can only occupy specific energy levels, they may not have intermediate energy levels between these allowed states. An electron may be “excited” or promoted from a lower energy level to a higher energy level by absorbing energy of the appropriate wavelength. Conversely, an electron may be “relaxed” down to a lower energy level from a higher energy level by emitting energy of the appropriate wavelength, in the form of a photon.

Bohr's theory successfully predicted the atomic spectrum of hydrogen. When electrical energy is supplied to hydrogen atoms in a gas discharge tube, also called a spectrum tube, the atoms absorb energy and the electrons are promoted to excited energy levels. Once excited, however, the electrons have a natural tendency to drop back down to a lower energy level by emitting light of the appropriate wavelength and energy. The emitted light for a given transition is observed through a diffraction grating as a bright line in the emission spectrum of hydrogen.

The relationship between the energy of light and its wavelength is shown in Equation 1.

$$\Delta E = \frac{hc}{\lambda} \quad \text{Equation 1}$$

$\Delta E$  is the difference in energy between the two energy levels in joules,  $h$  is Planck's constant ( $h = 6.626 \times 10^{-34}$  J·sec),  $c$  is the speed of light ( $c = 2.998 \times 10^8$  m/sec), and  $\lambda$  (lambda) is the wavelength of light in meters.

When Bohr calculated the allowed energy levels for the electron in the hydrogen atom, he found that the results correctly predicted the wavelengths of visible light observed in the emission spectrum of hydrogen (Figure 1).

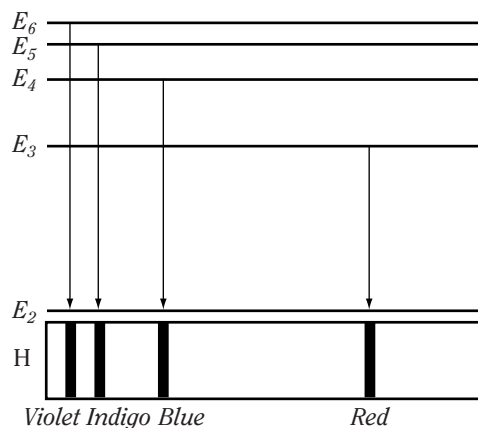


Figure 1.

The Bohr model of electron structure was found to be inadequate for atoms containing more than one electron. However, the idea that only certain stable electron energy levels are allowed has endured. The development of quantum mechanics in the 1920s built on the idea of quantized energy levels and introduced the idea of the wave nature of matter to describe the properties of electrons. According to quantum mechanics, the location of an electron is not restricted to specific orbits but can only be defined in terms of the probability of finding an electron. A system of atomic orbitals was introduced to account for the arrangement of electrons around the nucleus of an atom. An atomic orbital is a region in space where an electron may be found.

## Atomic Spectra *continued*

Atomic orbitals differ in their size, shape, and orientation in space, and also in their energy. The characteristic atomic emission spectrum of an element can be interpreted based on the unique arrangement of atomic orbital energy levels for its atoms.

### Sample Data

Light Source	Spectrum (Number of Lines)	Colors (Wavelength, nm)
Incandescent Light	Continuous spectrum	Red (680–600), orange (600–570), yellow (570–550), green (550–490), blue (490–460), violet (450–400)
Hydrogen Spectrum Tube	Line spectrum (4 lines)	Red (649), blue (476), indigo (425), violet (400)
Mercury Spectrum Tube	Line spectrum (3 lines)	Yellow/orange (570), green (539), blue-violet (430)
Fluorescent Light	Continuous and line spectrum	Continuous “rainbow” visible spectrum with the following bright lines superimposed red (605), green (536), blue (426)
“Blue neon” Novelty Lamp (optional)	Line spectrum (4 lines)	Yellow (572 and 570, two closely spaced lines), green (539), blue (427)
Street Light (optional)	Line spectrum (7 lines)	Red, orange (2 closely spaced lines), yellow (bright), green, blue, and blue-violet (wavelengths could not be measured in the dark!)

### Materials for Atomic Spectra are available from Flinn Scientific, Inc.

Catalog No.	Description
AP1327	Spectrum Tube Power Supply
AP1334	Hydrogen Gas Spectrum Tube
AP1337	Mercury Vapor Spectrum Tube
AP6040	Project Star Student Spectrometer

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.