12. PARTICLE PROPERTIES OF LIGHT

You will investigate some of the basic ideas of quantum physics and the particle nature of light. You will calibrate a spectrograph using a number of light sources emitting known wavelengths and then use it to study the visible spectrum of hydrogen. This is a classic experiment in the histories of physics and astronomy.

Theory

Light is a Particle and a Wave

It was pointed out in the experiment “Wave Properties of Light” that the nature of light—particle or wave—was disputed for many years. Diffraction and interference were explained very well by assuming that light is a wave phenomena. Indeed, well into this century the wave theory of light was the predominant view, a view shared by Isaac Newton amongst others.

There were, however, a number of experiments involving light whose results could not be explained, at least well, by the wave theory. Among them were those dealing with the colors or wavelengths of light emitted by hot gases like hydrogen and helium. The experiment you are about to do is one of those experiments which is taken to demonstrate the particle nature of light.

Light Sources and Spectrographs

There are many different sources of light. The most common in the laboratory or in the home is probably the ordinary light bulb. A light bulb has a tungsten filament, and when an electric current is passed through the tungsten, the tungsten heats up until it eventually glows white hot. The white light that is emitted consists of all visible wavelengths distributed continuously. Sunlight is of this form.

In contrast, discharge tubes which contain pure gases or vapors of single elements emit light of definite wavelengths and definite colors: red, green, blue, and so on. Astronomical sources emit (and absorb) light of similar colors. Discharge tubes enable us to focus on the study of the electronic structure of elements.

A discharge tube is made by encapsulating a pure gas or vapor in a glass envelope containing embedded electrodes. If a high enough voltage is applied between the electrodes the ions in the gas are attracted to the electrodes and collide with other atoms. According to the modern explanation, large numbers of atoms are excited to higher energy states. When the atoms decay to lower energy states they emit light—the gas glows.

In this experiment you will study this light by means of a prism spectrograph. The light comes from the discharge tube and passes through a glass prism (Figure 1) and out a telescope and eyepiece. Because of dispersion in the glass prism, light of shorter wavelength is refracted through a greater angle than light of longer wavelength. Thus the component colors or wavelengths of the light are separated one from another. You can measure the angle of deviation $D$ of a given color as sketched in Figure 2. Then if you calibrate the angle $D$ in terms of wavelength $\lambda$, you can find the wavelengths of an unknown source of light from the calibration curve. This is the primary objective of this experiment.
Early in the last century experimentalists discovered that elements in the form of gases and vapors in a “state of discharge” emit light of specific characteristic wavelengths or colors. These wavelengths or “lines” form regular mathematical series. (Line is a jargon word for the image of the slit in the spectrograph one actually sees.) In 1885, Johann Jacob Balmer, a Swiss secondary school teacher who studied science as a hobby, deduced that the wavelengths of the visible lines in the emission spectrum of hydrogen that had by then been published can be represented to great accuracy by the empirical formula:

\[
\frac{1}{\lambda_n} = R \left[ \frac{1}{2^2} - \frac{1}{n^2} \right], \quad \ldots [1]
\]

where \( n = 3, 4, 5 \ldots \) and \( R \) is a constant. \( R \) is called the Rydberg constant. The most recent measurement of \( R \) has yielded the value 10,973,731.5 m\(^{-1}\). Though ignorant of the physics giving rise to this formula (as was everyone else for the next 30 years), Balmer was able to predict the existence of other series in hydrogen, series described by eq[1] with the “2” replaced by integers “1”, “3”, “4”, etc. These series were finally explained in 1913 by the model of the atom proposed by Niels Bohr.

The Bohr Theory

The Bohr model of the hydrogen atom was one of the first great successes of the early quantum theory. Bohr assumed in his model that the atom consists of a kind of miniature solar system. At the center of the system is a rather heavy, positively charged nucleus and around this revolves the lighter negative electron. The electron is held in orbit by the Coulomb attraction that exists between the positive nucleus and the negative electron. Unlike classical mechanics Bohr’s new mechanics requires that the angular momentum \( l \) of the electron be quantized, that is,

\[
l = mvr = \frac{nh}{2\pi}, \quad \ldots [2]
\]

where \( n = 1, 2, 3 \ldots \) The factor \( h \) is today called Planck’s constant and has the value 6.6260755 \times 10^{-34} \text{ Js}. The integers \( n \) are called principal quantum numbers. Using eq[2] in a calculation of the total energy of the electron (kinetic plus potential) yields

\[
E_n = -\frac{k_e^2 2\pi^2 e^4 m_e}{h^3 n^2}, \quad \ldots [3]
\]
where \( k_e = 8.99 \times 10^9 \) \( \text{Nm}^2\text{C}^{-2} \) is called the electric constant and \( m_e = 9.109389 \times 10^{-31} \) \( \text{kg} \) is the electron mass. Thus unlike the classical theory, the quantum theory predicts the electron can exist in only a finite number of specific energy states defined by the integers \( n \). Without explaining how (because he didn’t know at the time), Bohr assumed an electron can remain in one of these states indefinitely.

These states can be represented as the highly-schematic “levels” in Figures 3 and 4. Bohr interpreted eq[3] to mean that light of a specific wavelength (a line) was emitted when an electron in the atom decays from a higher energy level \( i \) to a lower one \( f \). The difference in energy

\[
h\nu = E_{ni} - E_{nf} \quad \ldots[4]
\]

between the two levels is emitted as a quantum, or photon, of light. The quantity \( \nu \) is the frequency of the light; and \( n_i \) and \( n_f \) are, respectively, the quantum numbers of the initial and final states.

\[
\frac{1}{\lambda} = R \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \quad \ldots[5a]
\]

where

\[
R = \frac{k_e^2 e^4 2\pi^2 m_e}{c h^3} \quad \ldots[5b]
\]

The simplicity of eq[5a] and its similarity to eq[1]

Figure 3. Orbit representation of four states in the Bohr model of hydrogen.

Figure 4. Energy level representation of the Bohr model of hydrogen.
should be apparent to you. Certainly it was tremendously exciting to Bohr! For each value of \( n_f \) a series of transitions (or lines) exists for all possible values of \( n_i > n_f \). Only the series for \( n_i = 2 \) lies in the visible region of the spectrum. This is the Balmer Series. The Lyman Series lies in the ultraviolet region and results from transitions to the \( n_f = 1 \) state. The Paschen Series lies in the infrared region and results from transitions to the \( n_f = 3 \) state. (These series are named after the relatively obscure experimentalists who first discovered them and measured their wavelengths.) Although a simple prism spectrograph may be used to observe the Balmer Series, as in this experiment, more sophisticated research-grade equipment is required to observe the Lyman and Paschen Series, being as they are, invisible to the unaided eye.

### The Experiment

#### Exercise 0. Preparation

**Orientation**

Carry out the following cold start checks to ensure you operate the spectrometer correctly:

- If ON, turn the apparatus OFF.
- To protect you from electrical shock, the four discharge tubes (He, Hg, Ne, and H\(_2\)) are housed within the large brown insulated cylinder. To change from one tube to another, turn the voltage control to minimum; then rotate the cylinder one “space” by hand.

**IMPORTANT**

Before switching power ON, and before changing tubes, always turn the high voltage control to minimum.

- Switch power ON and turn the voltage control about halfway up. Allow the tube (any one) a few minutes to warm up. Fine tune the voltage control so that the light just stops flickering.
- Confirm that the entrance slit of the spectrograph is aligned against the entrance of the cylinder so as to receive maximum intensity of light. Adjust the width of the slit to resolve sharp, bright lines through the eyepiece. The eyepiece at the near end of the telescope can be moved into or out of the tube so as to focus both the scale and spectrum lines. Make the illuminated scale just bright enough to avoid eyestrain.

### Exercise 1. Calibration

Before any spectrograph is used it must be calibrated. Calibrate your instrument as follows. Measure the positions of the brightest lines in the spectra of the three calibration sources relative to the arbitrary scale in the spectrograph. Identify these lines by color and wavelength from Table 1. Make one plot of wavelength (in nm) vs. scale position for all the lines together on a sample of linear graph paper. (Why should you plot the wavelength along the vertical axis?) The result represents the calibration curve of your particular spectrograph.
Table 1. Wavelengths in nm ($10^{-9}$ m) of a few of the brightest lines of three calibration sources (rounded off to one decimal place). Adapted from information in the CRC Handbook, and R. A. Sawyer, Experimental Spectroscopy (Dover Publication, 1963), page 218.

<table>
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<th>Color (approx.)</th>
<th>Source</th>
<th>Hg</th>
<th>Ne</th>
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<td></td>
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Exercise 2. Measurement and Manual Interpolation

Now measure the positions of the brightest lines in the spectrum of hydrogen. With reference to the calibration curve you plotted in Exercise 1, interpolate their corresponding wavelengths. How well do your measurements compare with the wavelengths listed in Figure 4? Note your experimental errors carefully.
The plot you did by hand in Exercise 1 should resemble Figure 5. In Figure 5 the pairs (scale position, wavelength) for the calibration sources (He, Hg and Ne) are all plotted on the same graph and a smooth curve drawn through the points. The wavelengths corresponding to the scale positions of the hydrogen lines were found graphically as should be clear from the figure.

![Calibration Curve](image)

**Figure 5.** What the calibration curve of a typical spectrometer might look like if drawn by hand. Interpolation is done by drawing vertical and horizontal lines meeting on the curve.

Clearly, this method has two obvious problems. Firstly, you must draw a smooth curve through the data points by hand. Secondly, the physical size of the graph paper will determine the accuracy of the value you are able to interpolate. A superior method is clearly desirable.

You can interpolate more accurately using a computer and the programs proFit and Excel. Try the following procedure:

1. Input the data in the form \((S, \Delta S, \lambda, \Delta \lambda)\), where \(S\) is the scale position (in arbitrary units) and \(\lambda\) is the wavelength (in nm) into pro Fit. Fit a function of whatever order you think appropriate.

To illustrate, Figure 5 below, which was plotted by Dr. Quick using proFit, represents a fit (his choice) to the following fourth-order polynomial

\[
\lambda = \text{const} + a_1 S + a_2 S^2 + a_3 S^3 + a_4 S^4
\]

with these values for the coefficients:

\[
\begin{align*}
\text{const} &= 0.80381633 \times 10^3 \\
a_1 &= -0.10646273 \times 10^3 \\
a_2 &= 0.14078903 \times 10^2 \\
a_3 &= -0.10406786 \times 10^1 \\
a_4 &= 0.31014955 \times 10^{-1}
\end{align*}
\]

Your results will, of course, differ from these.

You can enter this information along with the scale positions you measured for the hydrogen lines into an Excel WorkSheet. Try the following:
2. Double click Excel so that a new worksheet is presented.
3. Enter the scale positions of your hydrogen lines in column A of the worksheet as shown in Figure 6.
4. Click on cell B1 and enter the formula you obtained from profIt as described in step 1. Note that the value of S for B1, is found in cell A1. When you are finished entering the formula as shown in Figure 6 press the<ENTER> key, not the<RETURN>key; this will leave B1 selected.
5. Now while holding down the <SHIFT> key, click on cell B4 so that the entire column B is selected as shown in Figure 6. Select FILL DOWN from the edit menu. Your interpolated values of wavelength should now appear in Column B. You may now save this sheet to your account space or print it out in the usual way.

Figure 6. An Excel worksheet created to perform interpolations. A FILL DOWN command has just been executed. The menu bar shows the formula entered in cell B1. Values remaining in column B are the interpolated wavelengths in nm. These compare very well with the expected values of 656.2, 486.1, 434.0 and 410.5 nm, respectively.

Exercise 4. Other Topics

The Rydberg Constant
An experiment like this one enables you to determine the Rydberg constant among other things. First calculate a theoretical value of R by substituting the required physical constants into eq[5b]. (You may wish to confirm these constants from the reference in endnote 1.) How well does this calculated result agree with the “accepted” value quoted in the theory section? Now calculate your own experimental R using eq[5a] and your measurements of λ (one R for each λ). Compare your average R with the accepted value given earlier. Do the values agree to within your experimental error? Using eq[5a] calculate the limiting wavelength; this is the wavelength beyond which no lines occur.
Particle Properties of Light

Videos and Physics Demonstrations on LaserDisc

The episode “Doubt” from the series *The Ring of Truth*, with Philip Morrison, Tape #28
The episode “Electromagnetic Radiation” from the series *Project Universe*, Tape #1,

from Chapter 67 *Atomic Physics*
   Demo 25-01 *Emission Spectra*
   Demo 25-02 *Spectral Absorption by Sodium Vapor*

Activities Using Maple

*E12 Particle Properties of Light*
Bohr's assumptions in his model of the atom are reviewed, and the radius of an electron orbit, velocity and energy are calculated. You have the option of inputting your data into this worksheet to fit a function and to interpolate wavelengths.

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EndNotes for Particle Properties of Light