The Hydrogen Spectrum

PHYS 1301 F99
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Introduction

In a previous laboratory experiment on diffraction, you should have noticed that the light from the mercury discharge tube was composed of only three colors, or three distinct wavelengths of light. Indeed, each element of the periodic table emits its own characteristic wavelengths of light. The collection of the different distinct wavelengths emitted by an atom is called the **emission spectrum** of the atom. Spectra composed from white light but with distinct wavelengths absorbed or removed are called **absorption spectra**. Each element's unique emission spectra can be thought of as a kind of "fingerprint" for the element. The element helium was first discovered in this manner through the spectroscopic analysis of light from the Sun and was only later discovered in natural gas deposits on Earth.

But why are *distinct* wavelengths observed? And why are they different for particular elements? There is nothing distinct about the light from an incandescent source. In an empirical study of the spectrum of hydrogen, it was discovered that the precise frequencies and wavelengths of the light produced could be described by an equation involving a constant and an integer. This equation was then expanded to describe the entire spectrum of hydrogen, including the ultra-violet and the infrared spectral lines. This equation is called the **Rydberg equation**:

$$\frac{1}{l} = R\,(- \ ),$$

where R is the Rydberg constant, and n_1 and n_2 are integers. That such a simple formula describes the emission spectrum of hydrogen is nothing short of amazing. You don't need to know the speed of the electron or the proton, and you do not need to know how rapidly the atom is moving. All you need to know are the constant R and two integers. The presence of integers in this equation created a real problem for physicists until the development of the quantum theory of the atom by Neils Bohr. Bohr's theory suggested that the electron orbiting the nucleus could only have certain *quantized angular momenta*. Ignoring the precise technical meaning of this quantity, the implication of it is that the electron can orbit only at discrete radii and speeds around the nucleus and subsequently can only possess certain discrete energies. The discrete radii can be labeled by integers and the integers are the ones found in the Rydberg equation. The integer n_1 is the **quantum number** of the *initial state* or energy level, and n_2 is the quantum number of the *final state*.

Transitions of an electron from one orbit to another of smaller size produce a burst of light. It is this light which you eventually see with your own eye. Since the energy of an electron in a given orbit is discrete, this implies that the energy of the photon emitted during the electron transition is also discrete. This is why we see discrete colors from the gas in the discharge tube and not a continuous rainbow spectrum. In a future lab, we shall see that the allowed orbit sizes and energies of an electron depend on the number of protons in the nucleus. It is because of this that each element has its own characteristic emission spectrum.

In this experiment, we will be measuring the various wavelengths of the spectral lines of hydrogen, correlating them with their proper quantum numbers, and experimentally determining the Rydberg constant.

Procedure

Set up the same apparatus as was used for measuring the mercury spectrum, except use a mercury discharge lamp rather than hydrogen one. You will notice four lines in the so-called Balmer series. (The Balmer series is just the Rydberg equation when $n_2=2$.) These are as follows:

Red	656.28 nm
Blue-Green	486.13 nm
Blue	434.05 nm
Violet	410.17 nm

Measure the wavelengths of these four spectral lines for yourself using the method from the mercury lab, recording their color and wavelength. You only need to measure the wavelength for the first order diffraction of each spectral line. Note that for today's grating, d = (1/7500) cm = 1333 nm.

Analysis

The integer numbers in the Rydberg equation label the orbit of the electron. For emissions in the visible range, the final state (n_2) is level 2. Substituting this into the Rydberg equation gives us the equation for the so-called Balmer series of spectral lines.

$$= \text{R} (-) \qquad\qquad n = 3, 4, 5 \ldots$$

where the quantum number n is equal to 3, 4, 5... with each larger integer corresponding to a more energetic transition and a <u>shorter</u> wavelength for the emitted light. You will have to associate which value of n goes with each particular spectral line. The observed emission lines should be in order (red=3, blue-green = 4, etc...) but a certain line may be faint and hard to detect. In addition, since R is a constant, your choices for the integers should yield values of R close to one another.

Substitute the proper measured wavelength and the quantum number to get experimental values for the Rydberg constant. Take caution to get the unit right for R.

You should have four values in all, one for each spectral line. Average all the experimental values for R together.

Compute the percent error for R. Assume that R = 109,677.58 1/cm.

Questions

Question 1. Summarize your results for this experiment, reporting your experimental value for R with the percent error.

Question 2. How was the hydrogen spectrum different from the mercury spectrum? Compare the number of lines that were seen and their actual colors.

Question 3. What do you think the *absorption* spectrum of hydrogen would look like? Use a figure to help your explanation.

Error Analysis

What are your primary sources of error? **Explain their significance**. Would a little oxygen or nitrogen contamination in the tube affect the spectrum of hydrogen? Be specific. Would you see more lines or fewer lines? Would these lines have the same color as the hydrogen less or would they be different. A simple sketch might help your grade.

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Name ______________________________________ **Section:**_________

Abstract

Data

Color	N	$2x$	x	y	θ	λ

Calculations (Use back if necessary. Show units!)

Calculations

Calculate the Rydberg constant from each of the wavelengths.

Wavelength λ	initial state *n*	Rydberg constant R
		R_{ave}

= R (-)

Conclusions

Error Analysis