Experiment 7: Spectrum of the Hydrogen Atom

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Office Hour: Mondays, 5:30-6:30PM

INTRO TO EXPERIMENTAL PHYS-LAB 1493/1494/2699

Introduction

- The physics behind:
 - The spectrum of light
 - The empirical Balmer series for Hydrogen
 - The Bohr model (a taste of Quantum Mechanics)
 - Brief review of diffraction
- The experiment:
 - How to use the spectrometer and read the Vernier scale
 - Part 1: Analysis of the Helium (He) spectrum
 - Finding lattice constant of the grating
 - Part 2: Measuring spectral lines of Hydrogen (H)
 - Determining the initial state of the electron

Light Spectra

 Isaac Newton (1670): shine sunlight through prism and you will observe continuous rainbow of colors.



 John Herschel (1826): shine light from heated gas through spectroscope, and you will observe monochromatic lines of pure color on a dim/dark background.



Discharge lamps and artificial light

- Herschel's discovery of emission spectra from heated gas was studied extensively in the 1800's.
- It was realized that a heated gas emits a unique combination of colors, called emission spectrum, depending on its composition.
- <u>Example</u>: Helium gas in a discharge lamp.
- <u>Main idea:</u> put a large voltage across the gas. It will break down and emit light. The light emitted is composed of discrete colors.



Helium emission spectrum.

Atomic spectra

• This is an example of the lines emitted from different gases



Interlude: Spectra in Astronomy

- In the 19th century astronomers realized that spectroscopy can be used as a powerful tool to study stars
- <u>1802</u>: William Wollaston observed sunlight separated into colors by a prism; he noticed dark lines in the spectrum at definite locations.
- <u>1817</u>: Joseph Fraunhofer observed these spectral lines in other stars as well. Sometimes they are the same, sometimes they are different.



- They correspond to the *absorption* (not emission!) of some wavelengths due to the most external elements of the solar gas
- *Implication:* The stars are made of the same stuff as Earth!

Interlude: How stars emit light

- Emitted spectrum: Very hot stellar center emits continuous (blackbody) radiation
- From the core all the wavelengths are emitted (just like a rainbow)
- Observed spectrum: Before arriving to us the light passes through surface of stars



Visible light

- Gases at surface *absorb* light at characteristic wavelengths
- Therefore, we detect all the wavelengths but the one absorbed. <u>This generates the Fraunhofer's lines</u>

Interlude: How stars emit light

Star's absorption spectrum:



The Balmer formula

- J.J. Balmer discovers (1885) that the quantized (discrete) visible emission spectrum of the <u>hydrogen</u> follows a mathematical formula:
 - He found the following expression for the wavelength of the absorption lines completely <u>empirically</u>. He did not provide any physical explanation for it:

$$\frac{1}{\lambda_{\text{line}}} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right); \text{ with: } R = 1.097 \times 10^7 \text{ m}^{-1}$$

- Different values of n_f correspond to different line series discovered by several scientists before Balmer himself:
 - $n_f = 1$ (Lyman Series)
 - $n_f = 2$ (Balmer Series) which is visible
 - $n_f = 3, 4, 5, \dots$ (Paschen, Brackett, Pfund)

Example: the visible spectrum

- For the <u>**Balmer series</u>** (visible spectrum) we fix $n_f = 2$ </u>
- If we then change *n_i* we obtain:
 - $n_i = 3$ yields $\lambda = 656$ nm (red)
 - $n_i = 4$ yields $\lambda = 486$ nm (green)
 - $n_i = 5$ yields $\lambda = 433$ nm (blue)
 - $n_i = 6$ yields $\lambda = 410$ nm (purple)
- The Balmer's equation describes the observed lines very accurately
- But what is the physics?



<u>NOTE</u>: in the real experiment you will perform:

I. "Green" looks bluish-green

- 2. "Blue" looks purple
- 3. "Purple" looks dark violet. It will be very hard to see...

Bohr's model for H atom

- In 1913 Neils Bohr proposed a physical model to describe the spectrum of the hydrogen atom. It was the birth of Quantum Mechanics!
- He hypothesizes that the angular momentum of an electron in orbit around a proton is quantized (*i.e.* it can only be a discrete multiple of a certain number):

 $L_n = n\hbar;$ with: n = 1, 2, 3, ...; $\hbar = 1.05 \times 10^{-34} \text{ J} \cdot \text{s}$

 Under this simple assumption he managed to compute the energy of the electron around the atom:

$$E_n = -\left(\frac{me^4}{32\pi^2\epsilon_0^2\hbar^2}\right)\frac{1}{n^2}$$

 As a consequence, the energy of the electron around the nucleus is also quantized!

Bohr's model for H atom

- In a real life experiment we can never measure the energy of the electron. We can however measure differences in its energy
- Suppose that an electron makes a transition from a level n_i to a level n_f (with $n_i > n_f$)
- In order to conserve energy it will have to emit a photon with energy exactly $E_v = \Delta E = E_{ni} - E_{nf}$
- Quantum mechanically the energy of a single photon is related to its wavelength as $E_v = hc/\lambda$
- Therefore, the wavelength of the emitted photon is:

$$\frac{1}{\lambda} = \frac{me^4}{64\pi^3 c\epsilon_0^2 \hbar^3} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

• This perfectly describes the spectrum of the hydrogen atom!

Towards Quantum Mechanics

- The Bohr's model was a very first step towards a new paradigm of physics
- It still had unresolved problems like:
 - Could not capture some subtle properties of H atoms
 - Does not apply to other kinds of elements
- Eventually the development of Quantum Mechanics (1920's) lead to a description of matter that is still in agreement with experiments today
- However, the main idea present in the Bohr's model remains: <u>some physical</u> <u>quantities like energy and angular</u> <u>momentum can be discrete. This changes</u> <u>everything!</u>





- How can we measure the energy levels (spectrum) of the hydrogen atom?
 - We **CANNOT** measure individual energy levels!
 - We ARE ABLE TO measure indirectly the transition of an electron from one state to another.
 - Absorption and emission of photons.



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 - Absorption and emission of photons.
- Absorption spectroscopy: Shine a range of various wavelength radiation on a sample and measure the wavelengths that get <u>absorbed</u> (excitation of atoms).
 - Not the kind of spectroscopy we will use
- <u>Emission spectroscopy</u>: Given a sample of excited atoms, measure the radiation that is <u>emitted</u> by the atoms.
 - This will be the principle behind this week's experiment.

The emission spectrum of hydrogen

Energy levels of the hydrogen atom:



The emission spectrum of hydrogen

 Lyman series: It is made of all the de-excitations that end up on the n_f = 1 level



The emission spectrum of hydrogen

Balmer series: All de-excitation emissions to the n_f = 2 state
Infinite number of them: n_i = 3, 4, 5... → n_f = 2

Ε



Review: interference of light

- <u>Question</u>: How can we measure the wavelength of the emitted light?
- **Answer:** We can use interference!
- Recall: For slits with spacing *d*, the condition for <u>constructive</u> <u>interference</u> (bright spots) is:

$d\sin\theta = m\lambda$



Review: interference of light

- One feature of this equation is the it is *wavelength dependent*. $d \sin \theta = m \lambda$
- Consequence: Each wavelength will interfere at a different angle:
 - $\lambda_i = \frac{d}{m} \sin \theta_i$; with: $\lambda_i =$ blue, red, yellow, ...
- A diffraction grating can be used as a wavelength separator!





The Experiment

Main goals

 As usual this experiment is divided in two parts. The first one is preliminary and used for "calibration". The second one is the actual measure of interest

Part 1: Calibration

- Determine the lattice constant ("d" in previous eqns.) of the diffraction grating
- Use He discharge lamp for a line with known wavelength (effectively a calibration)

Part 2: Balmer series

- Four visible de-excitation emission lines (*i.e.* wavelengths)
- Measure the angle of interference
- Calculate initial state of transition
- <u>Tips</u>

Equipment

You can rotate the telescope tube and change the angle θ



Equipment

• Light is emitted from the arc lamp and focused with a lens



Equipment

 When it reaches the diffraction grating, every wavelength is deviated with a different angle. By looking at the Vernier scale you can measure this angle



Reading the Vernier scale

Vernier scale is a very precise way to measure angles.



Reading the Vernier scale

Vernier scale is a very precise way to measure angles.



The zero of the arcminute scale will tell you how many degrees you have rotated the apparatus. 50 degrees plus a bit more than 30 arcminutes. $\theta = 50^{\circ}30' + a$ little bit

Reading the Vernier scale

Vernier scale is a very precise way to measure angles.



Looking at when the fine scale and coarse scale align you can determine the "little bit". In this case, they are aligned on the 13 mark of the fine scale. So:

$$\theta = 50^{\circ}30' + 13' = 50^{\circ}43' = 50.717^{\circ}$$

Part 1: calibration of the grating

• Using He lamp (yellow lamp), you will determine what the spacing between adjacent slits is.

• Three steps:

- Align so that 0th order line is at 0 degrees.
- Turn the telescope to the left, find *first* yellow line (m = 1) and measure angle θ_L .
- Turn the telescope to the right, find *first* yellow line (m = 1) and measure angle θ_R .
- Make sure to record errors for your angle measurements.
- Take <u>average</u> of both angles. Note that the angle you want to use is the <u>relative angle to the central maximum</u>. This means that your angle must be always between 0° and 90°

Part 1: calibration of the grating

Take as a given the wavelength of the yellow line of He:

 $\lambda_{\text{yellow}} = 587.56 \text{ nm}$

• Using the grating equation and setting m = 1, solve for d.

$$d = \frac{\lambda}{\sin \theta}$$

- Don't forget to propagate errors.
- Your final calibration result should be: $\,d\pm\sigma_d$

Part 2: Balmer series

- Replace the helium lamp with the hydrogen one
- Procedure:
 - For the first and second order lines (*m* = 1 and *m* = 2) do:
 - Measure θ_L and θ_R for each spectral line (*i.e.* dark-purple, purple, greenish-blue, and red line)
 - Calculate the average angle θ_{avg}
 - Find the wavelengths λ for each average angle $\theta_{\rm avg}$, taking into account the order m
- Determine the uncertainty σ_{λ} taking into account the error in the angle σ_{θ} and the lattice constant σ_{d} .
- Use a weighted average between m = 1 and m = 2 to find the final value of $\overline{\lambda} \pm \sigma_{\overline{\lambda}}$ for each color

Part 2: initial energy levels

Recall the Balmer's formula:

$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

- Starting from that find the initial level (n_i) as a function of the wavelength and the final level (n_f)
- Assuming n_f = 2 find the initial value of n associated to each observed wavelength
- You final result will be $n_i \pm \sigma_{n_i}$ for each color

Tips

- I feel like this experiment requires to be <u>careful at the following</u> <u>few aspects</u>:
 - 1. The experiment is clearly best performed in a dark environment. Try to screen your apparatus from ambient light in the best way possible (*e.g.* using the black cloth that you will find on the table)
 - 2. As already mentioned, the purple line will most likely look like a dark violet one and it will be quite difficult to see it. If you are having problems with it ask your TA. If that doesn't solve the problem neither, it is ok to skip it
 - 3. Remember to always measure your angle with respect to the central line. There are no negative angles!