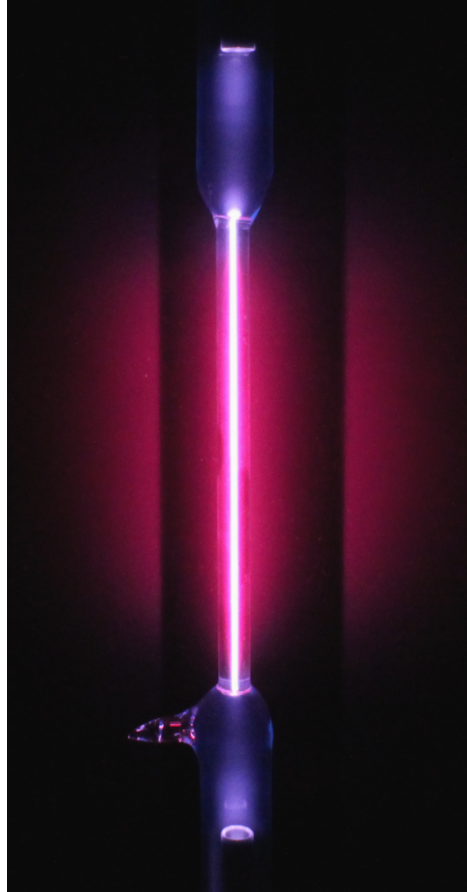


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# PHY143 LAB 4: ATOMIC SPECTRA

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## Introduction

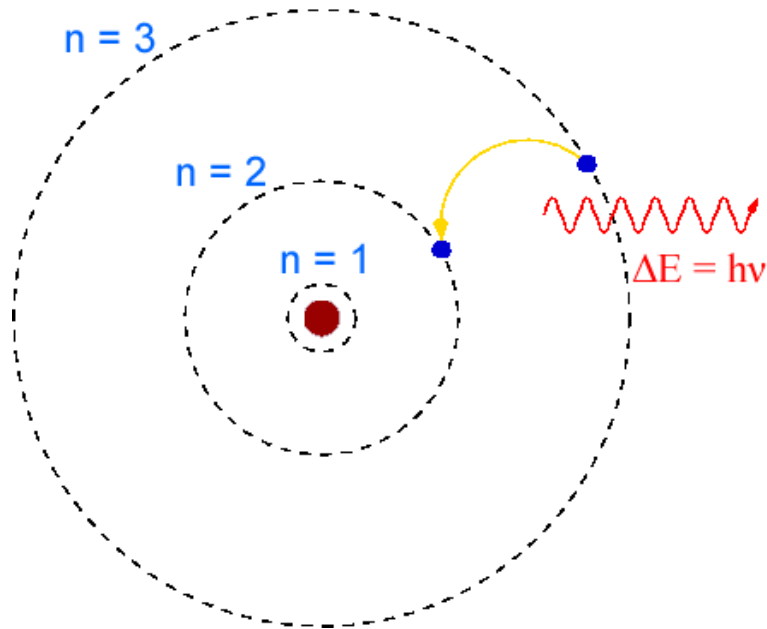
When an atom is excited it eventually falls back to its ground state, releasing the extra energy as photons. Since the energy of these photons directly corresponds to the gap between different energy levels in the atom, we can study the energy structure of the atom by measuring the wavelengths of these photons.

Every atom emits a unique set of energies known as its emission spectrum. Once measured, these spectra allow scientists to identify atoms or molecules based purely on the light they emit: a technique known as spectroscopy. This technique allows us to investigate the material composition of objects ranging from very small samples to distant stars.

In this lab you will use a diffraction based spectrometer to measure the emission spectrum of hydrogen and use the Rydberg formula to match each line in the spectrum with an atomic transition. You will then use the spectrometer to identify three different elements enclosed in electric discharge tubes.

## THEORY

The Bohr model coupled with the photon theory of light accurately describes the spectrum of hydrogen. We only need classical mechanics and Bohr's assumption that angular momentum is quantized according to  $L = n\hbar$ , where  $\hbar = \frac{h}{2\pi}$  is the reduced Planck's constant and  $n$  is an integer.



In the Bohr model of the hydrogen atom, the electron orbits the proton in a fixed, perfectly circular orbit. We start with Coulomb's law, which gives the magnitude of the force of the proton on the electron:

$$F = \frac{kq_1q_2}{r^2} = \frac{ke^2}{r^2}$$

where  $k = \frac{1}{4\pi\epsilon_0}$  is Coulomb's constant,  $r$  is the distance of the electron from the proton, and  $e$  is the charge on the electron. In order for the electron to maintain a circular orbit, as assumed, the proton must exert a force of magnitude

$$\frac{m_e v^2}{r}$$

where  $m_e$  is the mass of the electron and  $v$  is the velocity of the electron. Equating these forces yields

$$F = \frac{ke^2}{r^2} = \frac{m_e v^2}{r} \Leftrightarrow \frac{ke^2}{r} = m_e v^2$$

Now we can substitute in Planck's quantization assumption. Note that

$$L = p_e r = m_e v r = n \hbar \Leftrightarrow (m_e v)^2 = \frac{(n \hbar)^2}{r^2} \Leftrightarrow m_e v^2 = \frac{1}{m_e} \frac{n^2 \hbar^2}{r^2}$$

Substituting to eliminate  $v$ ,

$$k e^2 = \frac{n^2 \hbar^2}{m_e r} \Leftrightarrow r = \frac{1}{k} \frac{n^2 \hbar^2}{m_e e^2}$$

Now we can look at the total energy of the electron, given by the difference of the kinetic and potential energy:

$$\begin{aligned} E_n &= \frac{p_e^2}{2m_e} - \frac{k e^2}{r} = \frac{1}{2m_e} \left( \frac{n \hbar}{r} \right)^2 - \frac{k e^2}{r} = \frac{1}{2m_e} (n \hbar)^2 \left( \frac{k m_e e^2}{n^2 \hbar^2} \right)^2 - k e^2 \frac{k m_e e^2}{n^2 \hbar^2} \\ &= \frac{-k^2 m_e e^4}{2 n^2 \hbar^2} \end{aligned}$$

Note that all energies are conveniently proportional to the ground state energy  $E_1$ . Solving for this energy, we find

$$E_1 = \frac{-k^2 m_e e^4}{2 \hbar^2} = -13.6 \text{ eV}$$

This energy is known as the Rydberg energy  $R_e$ . Then we can express any energy level of the electron as

$$E_n = \frac{1}{n^2} \frac{-k^2 m_e e^4}{2 \hbar^2} = \frac{1}{n^2} E_1 = -\frac{13.6 \text{ eV}}{n^2}$$

When the electron jumps to a lower energy level, it emits a photon that carries the released energy. This energy is the difference of the initial and final energies.

$$E_\gamma = E_f - E_i = R_e \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

Where  $n_i$  is the electron's initial energy level and  $n_f$  is the electron's final energy level. According to the photon model, a photon has energy  $E = hc/\lambda$ , where  $c$  is the speed of light and  $\lambda$  is the wavelength of the photon. Since the wavelength is observable in the atomic spectra, we can use the wavelength to observe hydrogen transitions.

$$E_\gamma = \frac{hc}{\lambda} = R_e \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \Leftrightarrow \frac{1}{\lambda} = R_c \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

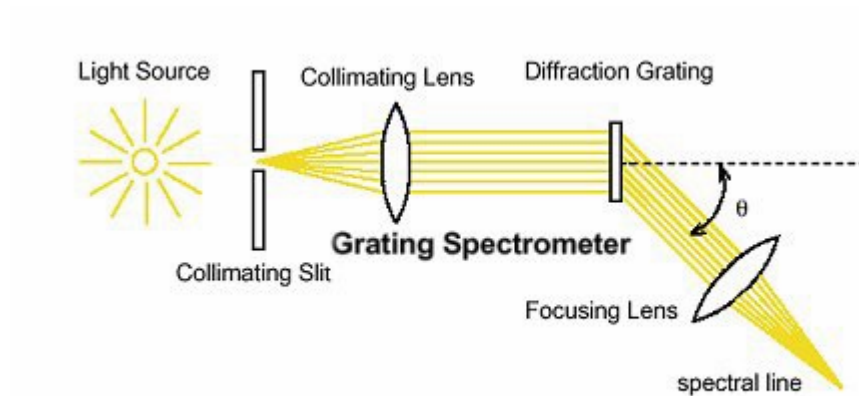
Where  $R_H = -\frac{13.6 \text{ eV}}{hc}$  is the Rydberg constant for hydrogen.

## Diffraction Grating

The spectrophotometer uses a diffraction grating to separate the component wavelengths of incident light. As we saw in the diffraction lab, a diffraction grating with line separation  $d$  will yield a pattern with maxima at angles  $\theta$  from the normal for

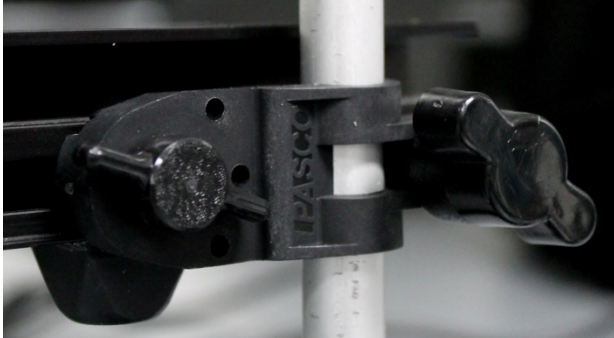
$$d \sin\theta = m\lambda, m = (0,1,2,3 \dots)$$

Measuring the peak positions of a known spectrum allows the separation  $d$  to be calculated precisely. This relation can then be used to find the wavelengths in unknown spectra.



## Setup

- 1) Place the spectrophotometer on the optics track.
- 2) Mount the track on two stands.



- 3) Place the collimating slit holder at one end of the track. Place the discharge lamp behind it, propping it up the wooden platform if necessary to align the middle of the lamp or the hole with the slits. Cover the lamp with the black cloth to block the extra light.



- 4) Lift the track up until the center of the collimating slits is level with the middle of the discharge lamp. Fix the track in place at this height.

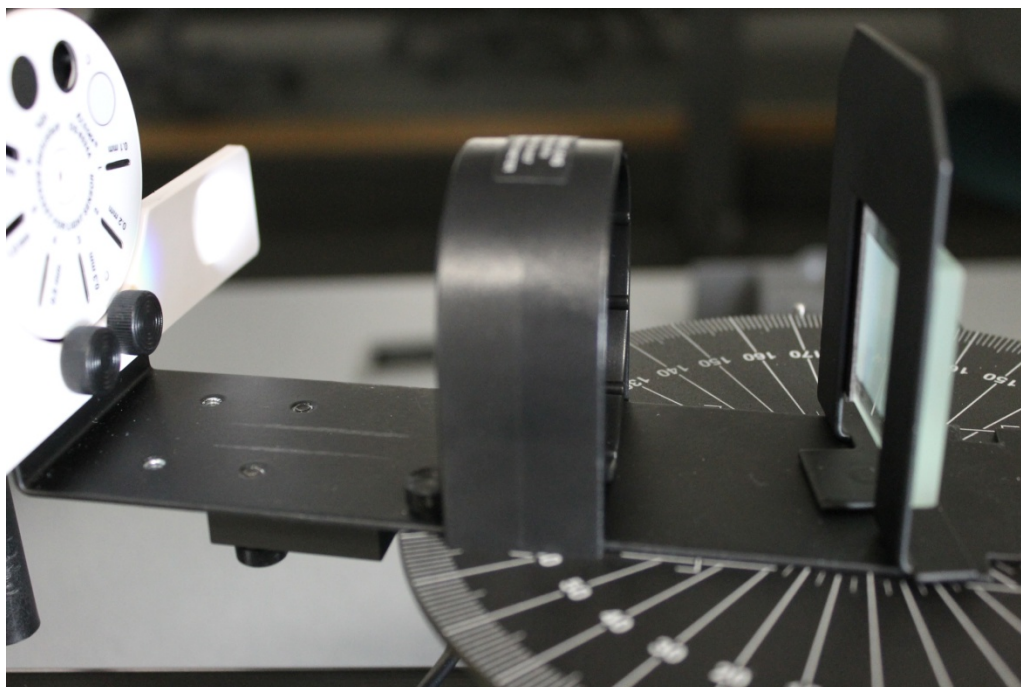
5) Attach the diffraction grating (which is magnetic) to the mount in the center of the spectrophotometer. Avoid touching the surface of the grating!



6) Attach the High Sensitivity Light Sensor and the Aperture Bracket to the arm of the spectrophotometer using the black rod. Plug the High Sensitivity Light Sensor into Analog Channel A on the ScienceWorkshop interface. Turn on the discharge lamp.

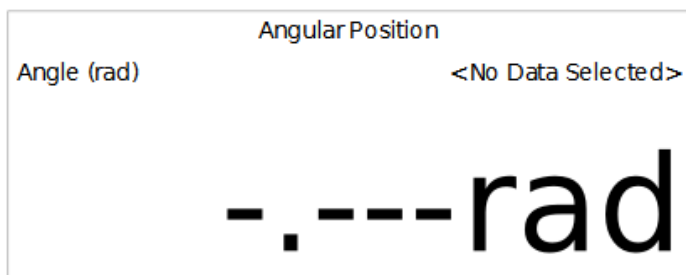
7) Place the collimating lens about 12 cm from the collimating slits. Have someone with 20/20 vision (corrected with glasses is ok) look through the collimating lens at the slits. Adjust the collimating lens until the slits are in sharp focus. The collimating lens should be about 10 cm from the collimating slits. Move the spectrophotometer close to the collimating lens.

8) Place the focusing lens on the spectrophotometer arm in between the light sensor and the grating. You should then adjust the focusing lens so that the light that is shown on the Aperture Bracket is focused.



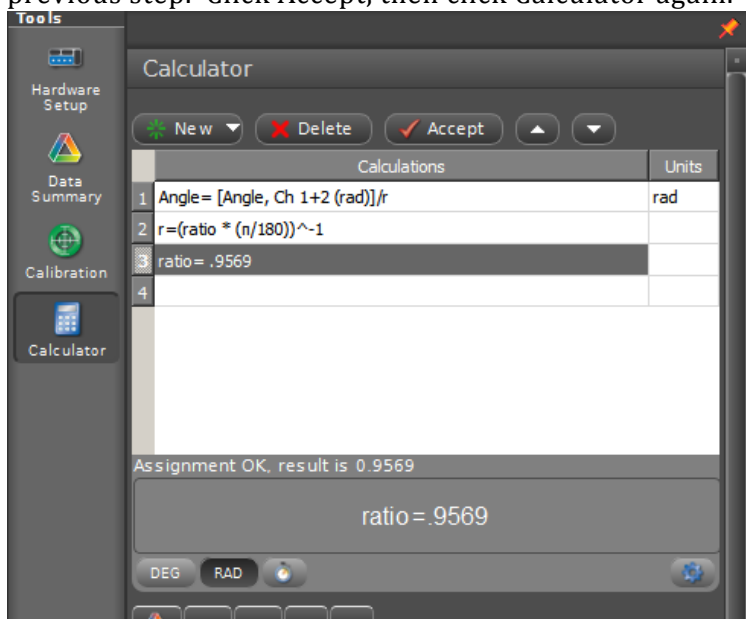
## Computer Setup

- 1) Open the Rotary Sensor calibration Capstone file. The purpose of this program is to determine the relationship between the rotation of the spectrometer arm and the rotation recorded by the rotary sensor.
- 2) Click "Record", then rotate the spectrometer arm between two degree marks. If the reading goes negative, reverse the rotary sensor's connection to the science workshop interface.
- 3) Write down the number of radians the rotary motion sensor rotates (shown on the screen) for your given rotation.

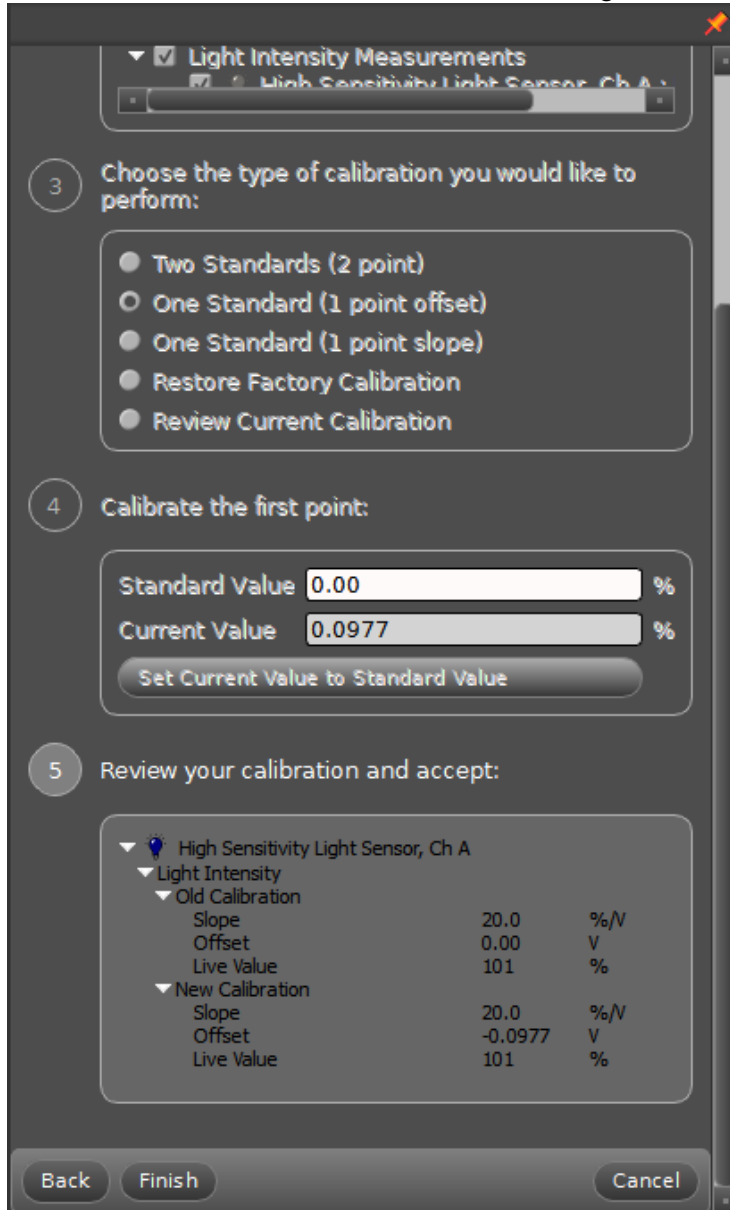


- Should you sweep through a small or large angle to make the procedure more accurate?
- How will ambient light affect your measurements? What are the sources of ambient light around your experiment, and how can you minimize them?

- 4) Take the number of degrees that you rotated the spectrophotometer arm and divide it by the number of radians that you got. The number you should get should be around 0.95-0.96.
- 5) Open the atomic spectra Capstone file. Click on Calculator found on the left side of the screen. On line 3, replace the number .9569 with the number that you got in the previous step. Click Accept, then click Calculator again.



- 6) Calibrate the High Sensitivity Light Sensor. Click Calibration (the green circle on the left side of the screen). From the menu, select Light Intensity and click next. Select the dot that says One Standard (1 point offset) and click next. Cover the Light Sensor then click Set Current Value to Standard Value. Click Finish. Click Calibration again to close the calibration menu.





## Experiment

Sweeping the detector arm through will now record a spectrum of the light from the discharge lamp. Try different aperture and slit sizes, and adjusting the lens locations, to record as many of the spectral peaks as possible. Dim spectral peaks will require careful tuning of the apertures to observe.

1) One of the unlabeled tubes is Hydrogen. The visible part of the hydrogen is known as the Balmer series. Use the Rydberg formula to determine the final and initial energy states ( $n_f$  and  $n_i$ ) for each transition line you observed:

$$\frac{1}{\lambda_{n_i \rightarrow n_f}} = R \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right), \text{ where } R = 1.097 \times 10^7 \text{ m}^{-1}$$

2) Use the spectrophotometer to measure the spectrum of each unlabeled discharge tube. Determine the wavelength of each peak and then compare your data with a table of spectra data to identify each sample. A good table of spectra lines can be found here: <http://hyperphysics.phy-astr.gsu.edu/hbase/quantum/atspect.html>

### THINGS TO THINK ABOUT

- How should you decide what slit apertures and sensor gain to use?
- Since this spectrometer uses a diffraction grating, you will observe copies of the spectrum on either side of the center position (the 1st order diffraction peaks). You will also observe 2nd order diffraction peaks, which may overlap with the 1st order. How can you tell them apart?
- How could you calibrate your readings, relating the measured angles to wavelengths?