Title: Atomic Spectra II (Hydrogen)

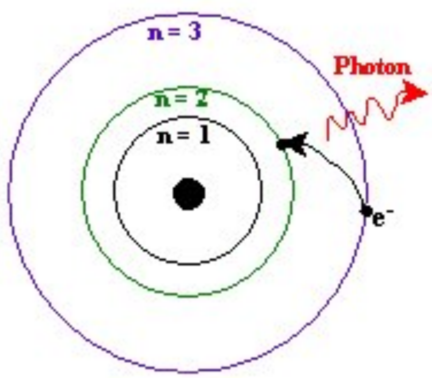
Equipment: USB spectrometer with fiber, Hydrogen light source.

Purpose: This lab demonstrates the quantized nature of the electronic energy levels of an electron constrained by the electrostatic potential to a nucleus. In this lab you will observe the quantized nature of these electronic energy levels by calculating and then measuring the atomic emission spectra of Hydrogen.

Introduction: Electrons constrained to small regions of space demonstrate quantized energy levels. When they are attached to an atom, electrons can only absorb and emit photons of energy consistent with transitions between allowed discrete energy states. Since every atom is different, these energy levels, and subsequent emission and absorption spectra, provide a "fingerprint" for the elements present. In today's lab, you will measure the atomic spectra for various elemental/molecular gases, including hydrogen and helium.

Light is given off by an atom when an excited electron decays from a higher energy orbit to a lower energy orbit. Calculating the energy levels and transitions for the more complex elements (or molecule in the case of water vapor) is outside our scope, however the energy levels of the electron in a **hydrogen atom** are given by:

$$E = -\left(\frac{m_e e^4}{8\varepsilon_o^2 h^2}\right)\left(\frac{1}{n^2}\right) \quad (1)$$



where m_e is the mass of the electron, e is the charge of the electron, ε_o is the permittivity constant, h is Planck's constant, and n is the energy level number (1,2,3,...). Plugging these numbers into Equation (1) gives:

$$E = -(13.6eV)\left(\frac{1}{n^2}\right). \tag{2}$$

Where 13.6eV is the ionization energy of Hydrogen. The energy of the emitted photon, ΔE, is the negative of the loss of energy of the electron and is given by:

$$\Delta E = E_f - E_i = (13.6eV)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right). \tag{3}$$

Where i and f refer to the initial and final energy level of the electron. For the visible photons given off by Hydrogen, the final energy level is nf = 2. The wavelength of the emitted photon is determined by:

$$\lambda = hc/\Delta E \tag{4}$$

Procedure:

1. ***Calculation:*** *Use equations 3 and 4 to calculate the* ***expected wavelengths*** *of the* ***visible wavelength****h photons emitted by Hydrogen atoms. Make a table of these theoretical wavelengths and the associated* n_f *and* n_i.

2. Measure the emission spectra of the Hydrogen source.

Capture the spectrum plot while the fiber is positioned for maximum signal. Make sure you can see all of the visible emitted wavelengths (you may capture more than one spectra if you need to 'zoom in' to see the dimmer wavelengths), and mark the wavelengths on the plot.

Analysis:

3. ***Hydrogen Spectrum, comparison with theory:*** *Compare your measured wavelengths to the theoretical wavelengths you calculated in #1, compute the % error and comment on sources of error.*

4**. Ionization Energy of Hydrogen:** If the values of the constants in equation 1 were not already known (you weren't able to calculate that the ionization energy of Hydrogen was 13.6 eV) you could use your data (your measured wavelengths) and equation 3 to determine the ionization energy.

Using your ***measured*** *wavelength values for Hydrogen and the associated* n_f*s and* n_i*s create a linear plot whose slope is the ionization energy of Hydrogen. Perform a linear fit and compare your empirically determined value of the ionization energy to the theoretical value, calculate % error.*

Hand in:

-Print outs of the Hydrogen spectrum.

-**Calculations** of the theoretical Hydrogen emission wavelengths

-**Measured** Hydrogen emission wavelengths, with %error and possible error sources

-Plot and fit giving **calculated** ionization energy of Hydrogen, with % error