Lab: Spectra of Atomic Hydrogen

Introduction

The emission and absorption of light energy of particular wavelengths by atoms and molecules is a common phenomenon. The emission and absorptions are characteristic for each element's atoms and arise from transitions of electrons among the various energy levels of the atom. The spectra produced by certain gaseous substances , consists of only a limited number of coloured lines with dark spaces between them. These *discontinuous* spectra are called **atomic or line spectra.** Each element has its own distinctive line spectrum - a kind of atomic fingerprint, this regularity in the line spectra was the key to understanding electronic structure.

The apparatus used to study the wavelengths of light emitted /absorbed by atoms is called a ***spectroscope.*** You will examine the line spectra of hydrogen using a simple spectroscope of the sort indicated in the diagram below:

The spectroscope includes four major features: a slit for admitting a narrow, collimated beam of light, a prism or diffraction grating that spreads the incident light into its component wavelengths, a telescope for viewing the spectrum, and an illuminated reference scale against which the spectrum may be viewed (as an aid in locating the positions of the lines in the spectrum). The scale of the spectroscope is calibrated by viewing a known element that produces especially sharp lines in its spectrum and whose spectrum has been previously characterised (with the emission wavelengths being known with great precision).

The most extensively studied atomic spectra is that of hydrogen. Hydrogen is the simplest of the atoms, consisting of a single proton and a single electron. The emission spectrum of hydrogen is of interest because this spectrum was the first to be completely explained by a theory of atomic structure, by the scientist Niels Bohr.

Atoms absorb and emit radiation as light of fixed, characteristic wavelengths when excited. This absorption and emission of light is now known to correspond to electrons within the atom moving away from the nucleus (energy absorbed) or closer to the nucleus (energy emitted). Atoms emit and absorb energy of only certain wavelengths (bright or dark lines in the spectrum) because electrons do not move randomly away from and toward the nucleus, but may only move between certain fixed, allowed "orbits," each of which is at a definite fixed distance from the nucleus.

When an electron moves from one of the fixed orbits to another orbit, the attractive force of the nucleus changes by a definite amount that corresponds to a specific change in energy. The quantity of energy absorbed or emitted by an electron in moving from one allowed orbit to another is called a **quantum (photon)**, and the energy of a particular quantum is indicated by the wavelength(or frequency) of the light emitted or absorbed by the atom. The energy of a photon is given by the Planck equation:

$$\Delta E = h f \qquad \Delta E = \frac{hc}{\lambda} \qquad c = f\lambda$$

where f is the frequency of light emitted or absorbed and λ is the wavelength (m), corresponding to that frequency (has units of cycles per seconds, since cycles is a dimension less quantity, the unit is written as s^{-1}, reciprocal seconds. In the SI system one cycle per second is a hertz, Hz, 1 Hz = 1 s^{-1}), and h is the Planck's constant, it has the units of energy x time, and has a value of 6.626×10^{-34} Js. The speed of light in a vacuum , $2.9979 \times 10^{8} ms^{-1}$, is one of the fundamental constants of nature and does not vary with the wavelength or any other properties of light.

A quantum is like a package or bundle of something that is available only in specific and separate amounts. A quantum is somewhat like the scoops of ice-cream, you can order one scoop, or two scops, but not 1.5 or 2.35 scoops. The amount of ice-cream that you can get are 'n' times the size of the scoop. Something that is **quantized** is restricted to amounts that are *whole-number multiples* of the basic unit, or quantum, for the particular system. A whole-number multiplier that specifies an amount of energy (or anything else that is quantized) is called a **quantum number. Quantum theory** is a general term for the idea that energy is quantized and the consequences of that idea.. A quantum of radiant energy is called a **photon**, and each photon has energy equal to hf.

Bohr postulated that the energy of an electron when it is in a particular orbit was given by the formula:

$$E_n = -\frac{\text{constant}}{n^2}$$

where n is the number of the orbit as counted out from the nucleus (n = 1 means the first orbit, n = 2 means the second orbit, etc.) and is called the **principal quantum number.**

The proportionality constant in Bohr's theory is called the **Rydberg constant** (given the symbol R_H) and has the value 2.179 x 10 $^{-18}$ J.

The electron energy value becomes negative, with its value lowered, energy has been released. According to Bohr's theory, if an electron were to move from an outer orbit (designated as n_{outer}) to an inner orbit (designated by n_{inner}), a photon of light should be emitted, having energy given by:

$$\Delta E = E_{inner} - E_{outer} = -R_H\left(\frac{1}{n_{inner}^2} - \frac{1}{n_{outer}^2}\right)$$

The wavelength, of this photon would be given by the Planck formula as :

$$\lambda = \frac{hc}{\Delta E}$$

The spectral lines observed in the visible spectrum of hydrogen arise from transitions from upper states back to the n = 2 principal quantum level.

Bohr performed calculations of wavelengths for various values of the principal quantum number, n, and found that the predicted wavelengths from theory agreed exactly with experimental wavelengths measured with a spectroscope. Bohr even went so far as to predict emissions by hydrogen atoms in other regions of the electromagnetic spectrum (ultraviolet, infrared) that had not yet been observed experimentally but that were confirmed almost immediately. Bohr's simple atomic theory of an electron moving between fixed orbits helped greatly to explain observed spectra and formed the basis for the detailed modern atomic theory for more complex atoms with more than one electron.

In this experiment, you will measure the wavelengths of the lines in the emission spectrum of hydrogen with a spectroscope and then determine by calculation to which transition (of the electron between the various orbits) each of these spectral lines corresponds.

Apparatus

Spectroscope with illuminated scale, hydrogen lamp (discharge tube), high-voltage power pack with lamp holder.

Procedure

1. Turn on the illuminated scale of the spectroscope.
2. Look through the eyepiece to make sure that the scale is visible.
3. Position the power supply pack containing the hydrogen vapour lamp so that the lamp is directly in front of the slit opening of the spectroscope.
4. Turn on the power supply switch to illuminate the hydrogen lamp.
5. Look through the eyepiece, and adjust the slit opening of the spectroscope so that the hydrogen spectral lines are as bright and as sharp as possible.

6. Record the colour and the location on the numbered scale of the spectroscope for each line in the visible spectrum of hydrogen. (You should easily observe red, blue-green and violet lines, a second very faint violet line may also be visible if the room is sufficiently dark !)
7. Turn off and unplug the power pack containing the hydrogen lamp.

Data Collection

Description of the spectrum:

Lines observed:

Colour	Location on spectroscope scale
__________	________________________
__________	________________________
__________	________________________
__________	________________________

Discussion

Use the equations provided in the introduction to the lab to perform the following calculations: For the following calculations show one sample calculation then present the complete results in a table.

1. a) Calculate the energy of hydrogen's first six energy levels.
   b) Calculate the predicted wavelengths in nanometres (according to Bohr's theory) from the electronic transitions in the hydrogen atom corresponding to the following :

   $n = 3$ ----------->> $n = 2$
   $n = 4$ ------>> $n = 2$
   $n = 5$ ------>> $n = 2$
   $n = 6$ ------>> $n = 2$
2. How do these predicted wavelengths correspond to those you have measured for hydrogen?
3. What transition corresponds to the point where the spectral lines merge?
4. What happens to the electron if $n = \propto$?
5. Could you use the energy value at this point to calculate a value for the ionisation energy? Explain.

Diagram of the Spectroscope