Spectrum of Atomic Hydrogen

Adapted from: ***Advanced Physics with Vernier – Beyond Mechanics ©Vernier Software & Technology*** **21 - 1**

BACKGROUND INFORMATION

Qualitative classification of light sources

From your classroom studies you are aware that we perceive as light is electromagnetic radiation with wavelengths between (roughly) 400 to 700 nm. This light can come from many different kinds of sources. Thermal radiation (aka blackbody radiation) is broadband with light of every color present with no gaps. Atomic emission spectra (aka line spectra) is referred to as the ‘fingerprint of the atom’ in texts for ‘physics without math’. Certain lines of specific wavelength are present with no radiation in the region between the lines. We define narrowband continuum as radiation over a portion of the visible region, but all wavelengths are present within that region. We can also have a ‘composite’ light source that is the superposition of a line plus a continuum where the energy of some of the line emission is used to create a continuum that is superposed over the line spectrum.

Quantitative description of the hydrogen spectrum

You have no doubt been exposed many times to the Bohr model of the atom. You may have even learned of the connection between this model and bright line spectra emitted by excited gases. In this experiment, you will take a closer look at the relationship between the observed wavelengths in the hydrogen spectrum and the energies involved when electrons undergo transitions between energy levels.

The appearance of bright line spectra for excited gases created a seemingly insoluble problem for physicists in the late 19th century. They could not find a simple way to explain why only certain wavelengths were emitted by excited gases. In 1885, Johann Balmer, a Swiss high school mathematics teacher, found an empirical equation

$$\lambda = 364.56\text{nm}\left(\frac{n^2}{n^2 - 2^2}\right)$$

where *n* was an integer greater than 2, that related the wavelengths of the lines in the visible spectrum of hydrogen. Three years later, Johannes Rydberg, a master of spectroscopy, rearranged Balmer’s equation and expressed it in a more general form

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where $1/\lambda$ is the wavenumber (the reciprocal of the wavelength expressed in m), R_H is Rydberg's constant and n_1 and n_2 are integers such that $n_1 < n_2$. This equation led to the discovery of similar sets of spectral lines in the ultraviolet (Lyman series) and infrared (Paschen series) in the early part of the 20th century.

Nevertheless, it was not until Niels Bohr proposed his model of the hydrogen atom in 1911 that a *causal explanation* for the existence of the bright line spectra emerged. Bohr assumed that the electron circled the nucleus in certain well-defined orbits corresponding to specific energy states (see Figure 1 at right). In his model of the hydrogen atom, the electron can exist only in one of these energy states. Ordinarily, the electron exists in its lowest energy condition (called the ground state). So long as the electron is in a particular energy state, the atom does not emit light energy. However, when a hydrogen atom is given enough energy (via an inelastic collision or photon absorption), the electron is bumped up from its ground state ($n = 1$) to an excited state ($n > 1$). When the electron drops back to a lower energy state, a photon is emitted. The lines in the hydrogen spectrum represent various transitions made by electrons from higher to lower energy states.

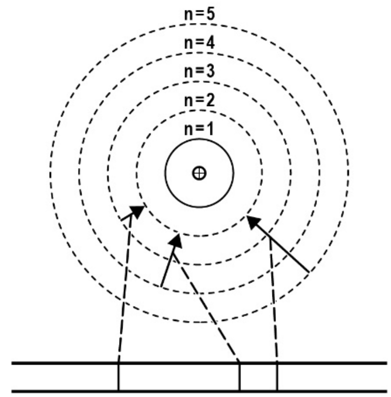

Figure 1

In this experiment you will analyze the visible bright line spectrum for hydrogen and use a variation of the Rydberg equation to relate the energy of the photons associated with each bright line to the energy levels in the Bohr model of the atom.