

The Atomic Spectrum of Hydrogen

When atoms are excited by heat or by electric discharge, they tend to give off electromagnetic energy. These photons of energy are emitted only at certain wavelengths that are characteristic of the specific atoms involved in the sample. These wavelengths make up the atomic spectrum of the element and tell us much about the electronic structure of the atom. Quantum theory tells us that atoms can only exist in certain states, each of which has a fixed amount of energy. When an atom changes energy states, it must absorb or emit an amount of energy that is just equal to the difference between the energies of the initial and the final states. When this energy is emitted, it is done so in the form of electromagnetic energy. This energy may fall within the visible region of the spectrum, or it may lie outside the visible in the lower energy infrared or the higher energy ultraviolet. In some cases the energy is even given off in the form of X-rays, but that takes specialized instruments to measure.

The emission spectrum of an element is obtained when excited electrons fall from a high-energy state to a lower state and emit a photon of "light". Since there are many possible transitions that an electron could make within an atom, the spectra of most atoms are extremely complex. Before you begin this lab, be sure that you have viewed the spectra of several elements through a diffraction rating to see the uniqueness of each pattern of colored lines.

Light is absorbed or emitted by atoms in the form of photons, each of which has a specific amount of energy, E . This energy is related to the wavelength of light by the equation

$$E_{\text{photon}} = hc/\lambda$$

where h is Planck's constant, 6.626×10^{-34} joule seconds, c is the speed of light in a vacuum, 2.997925×10^8 meters/second, and λ is the wavelength in meters. The energy, E_{photon} is in joules and is the energy given off by one electron when it jumps from a higher to a lower energy level. Since total energy is conserved, the change in energy of the atom, ΔE_{atom} , must equal the energy of the photon emitted.

$$\Delta E_{\text{atom}} = \Delta E_{\text{photon}}$$

where ΔE_{atom} is equal to the energy in the upper level minus the energy in the lower one. Thus, the energy transition from a higher energy level to a lower one is equal to the energy of the photon released.

$$E_{\text{higher}} - E_{\text{lower}} = hc/\lambda$$

The simplest of all atomic spectra is that of the hydrogen atom. In 1886 Balmer showed that the lines in the spectrum of the hydrogen atom had wavelengths that could be expressed by a rather simple equation. Bohr, in 1913, explained the spectrum on a theoretical basis with his famous model of the hydrogen atom. According to Bohr's theory, the energies allowed to a hydrogen atom are given by the equation

$$E_n = -b/n^2$$

Where b is a constant predicted by the theory and n is the Principal quantum number (1, 2, 3; ...etc.). It has been found that all the lines in the atomic spectrum of hydrogen can be associated with energy levels in the atom that are predicted with great accuracy by Bohr's equation. When we write the equation above in terms of a mole of H atoms we get the following equation:

$$N(E_{\text{higher}} - E_{\text{lower}}) = Nhc/\lambda$$

Multiplying both sides of the equation by N , the number of atoms/mol (6.02205×10^{23}), converting J to kJ , and substituting the numerical value for b , we obtain

$$E_n = -1312.04 \text{ kJ}/n^2 \text{ mol}$$

Atomic Spectrum Of Hydrogen Pre Lab Answers

**Harold R. Hunt, Toby F. Block, George
M. McKelvy**



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