Laboratory Report
Chemistry 111 Laboratory

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Emission Spectra

Introduction

Emission spectra are the lines of light which indicate the frequencies of photons that are emitted by excited electrons falling to lower energy states. The frequencies of light emitted by excited electrons can be used to identify specific elements and compounds, and therefore emission spectra are very useful in Chemistry. Infrared light can reveal the actual components of a molecule, and in some cases it can show how those elements are arranged within the molecule. Aside from the infrared region of the spectrum telling a chemist what elements are present in a compound and how they are arranged, the intensity of the light emitted can also indicate the concentration of each element within the compound.

The Bohr model is the simplest model for an atom, it consists of a nucleus and an electron shell. The shell has several layers, and electrons can move between the layers by gaining or losing energy. Atoms gain energy when they are exposed to heat or electricity, and then electrons take this energy and move to a higher energy level, called an excited state. When excited electrons lose energy, they travel to a layer of the shell that is closer to the nucleus and emit a photon. The color of the photon is dependent on how many energy levels the electron fell. The energy of the photon which is emitted is simply the absolute value of the energy lost by the falling electron:

$$E_{photon} = h\nu_{photon} = |\Delta E_{electron}| = \left|-2.18\times10^{-18} J\left(\frac{1}{n_f^2}-\frac{1}{n_i^2}\right)\right|$$

Not only do the photons emitted by electrons falling to their ground states have specific energy levels, they have specific wavelengths and frequencies of light as well. The visible range includes every wavelength from about 400nm to about 700nm. The color of the light tells us the wavelength, and also the frequency of the light, which we can use to measure the amount of energy lost by the electron as it fell to its ground state. Each element has its own unique line spectra because every element has different amounts of space between the energy levels of its electrons, which produce different colors of light due to the differences in energy lost by the electron moving down one or more energy levels. We can find the energy of the electron using the first equation, and the frequency of the light can be found using the second:

$$E = \frac{hc}{\lambda} \qquad \text{and} \qquad c = \lambda\nu$$

The objective of this lab is to familiarize students with the visible spectrum of light, and what produces light and causes it to be different colors.