<u>Abstract</u>

The purpose of this lab was to further our understanding of atomic structure and its relation to the production of light. To do this we used different spectrometers to look at the color spectrum that hydrogen, helium, mercury, and nitrogen produced. We observed and recorded the different wavelengths of the spectral lines emitted by each gas. We used this data to compare the intensity of the energy emitted from each gas. We realized that both hydrogen and helium had similar spectral lines but that hydrogen emitted more energy because of the smaller wavelengths. Observing mercury and nitrogen, we noted that the color of the glowing gas tube to the unaided eye hinted what the spectrum would look like. Mercury's gas tube color was very blue, and therefore had a spectrum that lacked red, meaning it produced greater energy than the rest of the gases we observed. In contrast, nitrogen, which glows a yellowish color, had a spectrum that lacked violet. Nitrogen therefore emits less energy than the rest of the gases we observed. We later used our measurements of wavelengths for hydrogen's spectral lines to calculate the energy difference between the orbits in the hydrogen atom. This proved that the longer the wavelength, the smaller the energy emitted.

Introduction

Light is the result of an event that took place within an atom. In an atom, electrons with negative charge orbit the nucleus which is made up of a combination of protons (positively charged) and neutrons (no charge). In an ordinary atom there are as many positive protons as there are negative electrons so the atom has no net electric charge. The electric forces attracting the positively charged protons and negatively charged electrons keep the atom from coming apart. Trying to understand the structure of hydrogen (the simplest and lightest of the elements), Bohr studied the connection between atomic spectra and atomic structure. Because of him we learned that an electron can jump from one Bohr orbit (as he named them) to another by having the atom gain or lose specific amount of energy. The atom must absorb energy for the electron to go from an inner to an outer orbit; the atom must release energy for the electron to go from an outer to an inner orbit. When the electron returns to the ground state after orbiting a higher orbit, the energy emitted is now emitted in the form of light. The color or wavelength of light emitted depends on the energy involved. Emitted light is like a wave where the distance between the crests is the wavelength. For larger jumps which require greater amounts of energy the light emitted is of a shorter wavelength. We know that the relation between the energy and the wavelength of the emitted light is: E (electric volts eV) = $12,400 \lambda$ (Angstroms). In order to actually see the difference in wavelengths that different gases emitted, our lab used a spectrometer. A spectrometer is a device that separates light by wavelength in order to see the component colors the gasses emit.

On October 2, 2007 we started out by looking through the spectrometer to observe the hydrogen spectrum and measure the wavelength of the four different colored lines visible. Next,

we observed the spectral lines on Helium's spectrum and again measured and recorded their wavelengths. We repeated this process with both Mercury and Nitrogen, using the spectrometer to observe the difference in spectrums.

Data and Analysis

Hydrogen Spectrum:

As we observed, hydrogen emits only four wavelengths in the visible region of the spectrum. As expected, the wavelengths decrease from red to violet, meaning that there is more energy released the smaller the wavelength gets.

Helium Spectrum:

Unlike what we expected, we observed that the spacing and location of the spectral lines was similar to hydrogen. The difference from each distinguishable spectral line is actually exactly the same as the difference between hydrogen's spectral lines. However, helium's spectrum began and ended at a higher wavelength than that of hydrogen's. Because violet's wavelength is larger than hydrogen's wavelength, hydrogen has the potential to emit greater energy than helium.

Mercury Spectrum:

The mixed color spectrum of the glowing gas tube light was very blue. The data explains that this is because the mercury spectrum lacks red spectral lines.

Nitrogen Spectrum:

Because the spectrum was not as clear, we were not able to get exact measurements of the spectral line's wavelengths. However, we could see from the spectrum that it was mostly red and that violet was almost not visible at all. The spectrum starts at the green and ends around the red, meaning that nitrogen emits less energy than the other gases observed.

Hydrogen Model:

The relation between the energy and the wavelength of the emitted light is:

E (electric volts eV) = $12,400 \lambda$ (Angstroms).

By studying the hydrogen model, we can see that the electron's jumps from one orbit to another produce the light visible on the spectrum. However, when the electron in a hydrogen atom jumps from orbit 2 to orbit 1, the light emitted has a wavelength of 1216 Angstroms, which is in the ultraviolet portion of the spectrum and therefore is not visible. We can use the equation stated earlier to find the energy difference between the orbits that the electrons jumped using each wavelength measured in the earlier part of the lab.

RED: E = **12,400 # Angstroms** = # eV when electron jumps from orbit # to orbit #.

YELLOW: E = **12,400 # Angstroms** = # eV when electron jumps from orbit # to orbit #.

BLUE-GREEN: E = 12,400 **#** Angstroms = # eV when electron jumps from orbit # to orbit #.

VIOLET: E = **12,400 # Angstroms** = # eV when electron jumps from orbit # to orbit #.

ULTRAVIOLET: E = 12,400 **#** Angstroms = # eV when electron jumps from orbit # to orbit #.

We can see from these calculations that the greater the wavelength, or more distance traveled by the electron from one orbit to another, the smaller the energy emitted. Because the wavelength of the light emitted when an electron jumps from orbits # to # is the smallest, it produces the greatest energy (# eV).

Because the separation of the orbits is proportional to the energy difference, we can construct a scale drawing illustrating the jumps made by the electron in producing the four wavelengths in the visible region of the spectrum. A distance of 1 cm will equal an energy difference of 1 eV.

Discussion

This lab presented us four examples of an emission light spectrum. We observed hydrogen, helium, mercury and nitrogen's spectral lines in order to measure their wavelength. This experiment shows the relationship between wavelength, spectral lines, and energy emitted. As stated earlier, the greater the wavelength, the more red appeared on the spectrum. Also, using the formula to find the total electric volts of the energy emitted in the form of light we see the inverse relationship between the wavelength and the total energy emitted. The shorter the wavelength, the greater the energy emitted by the element's atom.

In 1859, physicist Gustav Kirchhoff and chemist Robert Bunsen found that each chemical element produces its own unique pattern of spectral lines. This gave birth to the technique of spectral analysis, which is the identification of chemical substances by their unique patterns of spectral lines. Kirchhoff later summarized his findings into three important statements today called Kirchhoff's laws. His first law states that a hot opaque body or hot, dense gas produces a continuous spectrum which is a complete rainbow of colors without any spectral lines. His second law stated that a hot, transparent gas (like those observed in the lab) produces an emission line spectrum which includes a series of bring spectral lines against a dark background. His third law states a cool, transparent gas in front of a source of a continuous spectrum produces an absorption line spectrum that has a series of dark spectral lines among the colors of the continuous spectrum. The dark lines in this absorption spectrum occur at exactly the same wavelengths as the bright spectral lines in the emission spectrum of that same gas. The relative temperatures of the gas and its background also affects whether an emission or absorption spectrum is observed. Absorption lines are seen if the background is hotter than the gas, and emission lines are seen if the background is cooler. Therefore, the gases we observed in our experiment had to be warmer than their background, which they were.

Spectroscopy is truly important because it can determine chemical composition at any distance. For example, by studying the absorption spectrum of the Sun, we can identify the spectral lines and therefore determine the chemical composition of the Sun's atmosphere. Spectroscopy can also help us analyze gas clouds in space. In addition to informing us about temperature and chemical composition, the spectrum of a planet, star, or galaxy can also reveal something about that object's motion. In 1842, Doppler concluded that the wavelength of light is affected by motion between the light source and the observer. When a light source is moving, it emits wavelengths from various positions. The wavelengths are shortened when the light source moves toward the observer; this is called blueshift. Redshift occurs when the light source moves away from the observer and the wavelengths get longer. These findings are called the Doppler effect. Doppler also proved that the wavelength shift ($\Delta\lambda$) follows the equation:

 $\Delta\lambda\lambda 0 = VC$ where $\lambda 0$ = wavelength is source not moving, v = velocity of source measured, and c = speed of light. The Doppler effect is important in astronomy because it reveals information about the motions of planets, stars, and galaxies.