**Line Emission Spectra**

In this lab, you will observe and compare the emission spectra of **three samples**: Deuterium, Helium and Water Vapor. You will follow the scientific inquiry process and construct a poster to be used for your presentation and as an aid to the discussion of your findings.

**Learning Objectives:**

* To understand the relationship between the emissions spectral lines and energy changes and electron activity in the atom
* Explain connection between emitted photons and electron configuration
* Apply proper scientific investigation procedures and reporting techniques to share conclusions
* Understand and explain how energy absorbed by atoms is connected to photons

**Background:**

Emission spectra of elements are caused by electron transitions within atoms. They provide information about the arrangements of electrons in the atoms. The **emission spectrum** of a [chemical element](http://en.wikipedia.org/wiki/Chemical_element) or [chemical compound](http://en.wikipedia.org/wiki/Chemical_compound) is the spectrum of [frequencies](http://en.wikipedia.org/wiki/Frequencies) of [electromagnetic radiation](http://en.wikipedia.org/wiki/Electromagnetic_radiation) [emitted](http://en.wikipedia.org/wiki/Emission_(electromagnetic_radiation)) by the element's [atoms](http://en.wikipedia.org/wiki/Atom) or the compound's [molecules](http://en.wikipedia.org/wiki/Molecules) when they are returned to a lower energy state. Each element's emission spectrum is unique. Therefore, [spectroscopy](http://en.wikipedia.org/wiki/Spectroscopy) can be used to identify the elements in matter of unknown composition. Similarly, the emission spectra of molecules can be used in chemical analysis of substances.

Emission is the process by which a higher energy quantum mechanical state (energy level) of a particle becomes converted to a lower one through a [photon](http://en.wikipedia.org/wiki/Photon), resulting in the production of [light](http://en.wikipedia.org/wiki/Light). The frequency of light emitted is a product of the energy of the transition. Since energy must be conserved, the energy difference between the two states equals the energy carried off by the photon. The energy states of the transitions can lead to emissions over a very large range of frequencies, [visible light](http://en.wikipedia.org/wiki/Visible_light) is emitted by the coupling of electronic states in atoms and molecules.

When the [electrons](http://en.wikipedia.org/wiki/Electrons) in the atom are excited, for example by being heated, the additional [energy](http://en.wikipedia.org/wiki/Energy) pushes the electrons to higher energy orbitals. When the electrons fall back down and leave the excited state, energy is re-emitted in the form of a [photon](http://en.wikipedia.org/wiki/Photon). The wavelength of the photon is determined by the difference in energy between the two states. These emitted photons form the element's emission spectrum.

The fact that only certain colors appear in an element's atomic emission spectrum means that only certain frequencies of light are emitted. Each of these frequencies are related to energy by the formula:

*E*photon = *h* ν,

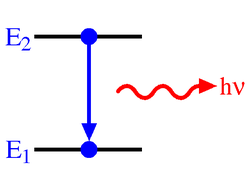
where *E* is the energy of the photon, *ν* is its [frequency](http://en.wikipedia.org/wiki/Frequency), and *h* is [Planck's constant](http://en.wikipedia.org/wiki/Planck%27s_constant). This concludes that only [photons](http://en.wikipedia.org/wiki/Photon) having certain energies are emitted by the atom. The principle of the atomic emission spectrum explains the varied colors in [neon signs](http://en.wikipedia.org/wiki/Neon_sign).

The frequencies of light that an atom can emit are dependent on states the electrons can be in. When excited, an electron moves to a higher energy level/orbital. When the electron falls back to its ground level the light is emitted.

**Emission spectroscopy**

Light consists of electromagnetic radiation of different wavelengths. Therefore, when the elements or their compounds are heated either on a flame or by an electric arc they emit energy in form of light. Analysis of this light, with the help of [spectroscope](http://en.wikipedia.org/wiki/Spectroscope) gives us a discontinuous spectrum. A spectroscope or a spectrometer is an instrument which is used for separating the components of light, which have different wavelengths. The spectrum appears in a series of lines called the line spectrum. This line spectrum is also called the Atomic Spectrum because it originates in the element. Each element has a different atomic spectrum. The production of line spectra, by the atoms of an element, indicates that an atom can radiate only a certain amount of energy. This leads to the conclusion that electrons cannot have just any amount of energy but only a certain amount of energy.

The emission spectrum can be used to determine the composition of a material, since it is different for each [element](http://en.wikipedia.org/wiki/Chemical_element) of the [periodic table](http://en.wikipedia.org/wiki/Periodic_table).. The emission spectrum characteristics of some elements are plainly visible to the naked eye when these elements are heated These definite characteristics allow elements to be identified by their atomic emission spectrum. Not all lights emitted by the spectrum are viewable to the naked eye, it also includes ultra violet rays and infra red lighting, an emission is formed when an excited gas is viewed directly though a spectroscope.

[](http://en.wikipedia.org/wiki/File:AtomicLineSpEm.png)

[http://bits.wikimedia.org/skins-1.5/common/images/magnify-clip.png](http://en.wikipedia.org/wiki/File:AtomicLineSpEm.png)

Schematic diagram of [spontaneous emission](http://en.wikipedia.org/wiki/Spontaneous_emission)

**Emission spectroscopy** **is a** [**spectroscopic**](http://en.wikipedia.org/wiki/Spectroscopy) **technique which examines the wavelengths of** [**photons**](http://en.wikipedia.org/wiki/Photon) **emitted by atoms or molecules during their transition from an** [**excited state**](http://en.wikipedia.org/wiki/Excited_state) **to a lower energy state. Each element emits a characteristic set of discrete wavelengths according to its** [**electronic structure**](http://en.wikipedia.org/wiki/Electronic_structure)**, by observing these wavelengths the elemental composition of the sample can be determined.**

The energy associated to an electron is that of its orbital. The energy of a configuration is approximated as the sum of the energy of each electron. The configuration that corresponds to the lowest electronic energy is called the [ground state](http://en.wikipedia.org/wiki/Stationary_state). Any other configuration is an [excited state](http://en.wikipedia.org/wiki/Excited_state).

As an example, the ground state electron configuration of the [sodium](http://en.wikipedia.org/wiki/Sodium) atom is 1s22s22p63s. The first excited state is obtained by promoting a 3s electron to the 3p orbital, to obtain the 1s22s22p63p configuration, abbreviated as the 3p level. **Atoms can move from one configuration to another by absorbing or emitting energy.** In a [sodium-vapor lamp](http://en.wikipedia.org/wiki/Sodium-vapor_lamp) for example, sodium atoms are excited to the 3p level by an electrical discharge, and return to the ground state by emitting yellow light of wavelength 589 nm.

Usually **the excitation of** [**valence electrons**](http://en.wikipedia.org/wiki/Valence_electron) **(such as 3s for sodium) involves energies corresponding to** [**photons**](http://en.wikipedia.org/wiki/Photon) **of visible or** [**ultraviolet**](http://en.wikipedia.org/wiki/Ultraviolet) **light**. The excitation of [core electrons](http://en.wikipedia.org/wiki/Core_electron) is possible, but requires much higher energies generally corresponding to [x-ray](http://en.wikipedia.org/wiki/X-ray) photons. This would be the case for example to excite a 2p electron to the 3s level and form the excited 1s22s22p53s2 configuration.

**The Poster Presentation Will Include:**

* A question, claim(hypothesis), evidence(data and results), conclusion and “the next question”
* A discussion of the similarities and differences between the spectra
* Inferences, data interpretations and conclusions.
* A visual representation of energy absorption and the emission spectra.
* An explanation of what causes the color emission.
* An explanation of the connection between the emission spectra, energy absorption and photons.
* A table with the samples electron configurations, observations of colors, and wavelengths of the emission lines.
* **A strong conclusion** which will address **all** of the following questions.
  + Why do certain colors appear in the emission spectra of the samples?
  + What is the connection between the emission spectral lines and the electrons in the samples?
  + What differences would you expect to see in the amount of emission spectral lines as you move down the periodic table?
  + What can examining a light emission spectrum tell you about electron arrangement?
  + What is the connection between emitted photons and electron configuration?

**Be prepared for your presentations which will begin on Friday 1/7/2011!**