LAB: Spectroscopy

Neon lights are orange. Sodium lamps are yellow. Mercury lights are bluish. Electricity is doing something to the electrons of these elements to produce light of a distinctive color.

PURPOSE:

1. To build a simple spectroscope
2. To observe visible continuous and discrete spectra of various light sources
3. To measure the wavelength of spectral lines from a hydrogen light source.
4. To calculate the energy of a photon of light using the relation: **$E = h \cdot f$**

THEORY:
Elements that exist in the ground state (i.e. unexcited) emit no light. Energy applied to the atoms, in the form of an electric current, may be absorbed by their electrons. As the energy is absorbed, electrons become excited and are bumped up to higher orbits. According to the Bohr model of the atom, only quantum levels of excitation are allowed. Electrons do not remain in the excited state forever. They eventually drop back to the ground state. The energy that made them excited is released as electromagnetic radiation (in other words, light). This is why neon lights glow when plugged in.

Electromagnetic radiation seen by the human eye is called visible light. Differences in visible light energy result in color. At the macro level, color gives us a subjective measure of light energy. In the micro view, color is the result of electrons bouncing up and down between orbits. Different orbits give different colors. This objective measure of energy is expressed by frequency, f, and wavelength, λ (lambda).

Light travels in waves much like those seen on the surface of the ocean before they crash onto the shore. The distance from wave peak to wave peak is called the wavelength (λ)[1]. You've probably seen ocean waves with wavelengths of 3 <u>meters</u> or more. Visible light is commonly expressed in wavelengths of 300-700 <u>nanometers</u>. Wavelength determines color, and color indicates energy. If the wavelength of light is known, then the <u>energy</u> of that wavelength may be calculated via the following equations from <u>laws</u> of physics:

Energy is	h times frequency…………………..	$E = h \cdot f$
	h is Planck's constant……………..	$h = 6.626 \times 10^{-34}$ J*s
	c is the speed of light……………..	$c = 2.998 \times 10^{8}$ m/s
Velocity is	frequency times wavelength…..	$c = f \cdot \lambda$
Frequency is velocity divided by wavelength.		$f = c / \lambda$

Putting it all together:

$$E = \frac{h \times c}{\lambda} = \frac{(6.626 \times 10^{-34} J \cdot s)(2.998 \times 10^{8} m/s)}{\lambda}$$

In summary, we see different colors because light consists of different wavelengths (corresponding to each color). The amount energy (in Joules) that is associated with a given color, actually the energy per photon of that color, can be calculated by a "simple" equation (see above).

[1] (Note: λ is different from <u>peak height</u> which is of interest to surfers and small craft)

Spectroscopy:
A single source of light may contain radiation of many different wavelengths. Evidence of this is the way a rainbow reveals that plain sunlight is actually composed of many colors. Droplets of water in the atmosphere act as thousands of prisms to produce this effect. The simple spectroscope uses a special prism, called a diffraction grating, to do the same thing.

In this experiment, light from an energized source will be viewed through a spectroscope. A spectroscope is an instrument that uses a diffraction grating to split up the light into its component colors. The component colors may then be viewed against the calibrated scale inside the spectroscope.

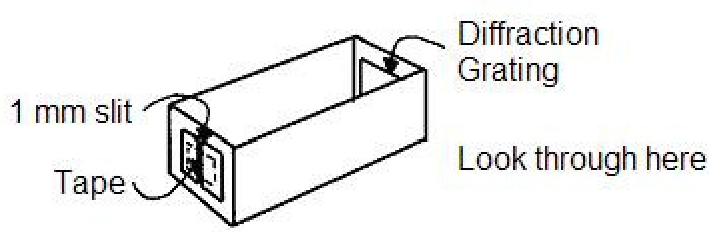
EXPERIMENT
Part 1: Building a Spectroscope (time permitting)

A simple spectroscope can be constructed from the following items:

- a thin box (such as a shoe box or cereal box)
- a diffraction grating
- black electrical tape
- a scalpel (or razor blade).

Procedure:

1. Cut a 2 cm square at each end of the box
2. Cover one hole with two pieces of tape so that you have a slit about 1 mm wide (see diagram below)
3. Cover the other hole with the diffraction grating (be sure grating is aligned with the slit)
4. Hold the box so that the grating is close to your eye and point the other end toward a light source.
5. Congratulations. You have just constructed a simple spectroscope.



Part 2: Qualitative Spectroscopy

In the following steps you will observe several light sources with your spectroscope. The spectroscope will allow you identify the individual colors that make up the spectrum of that light source. Using colored ink pens or crayons (sorry but yah gotta do it…) sketch the spectrum or line spectra for the following light sources: *{feel free to include any additional observations you make…}*

1. an incandescent or fluorescent light

2. hydrogen gas tube

3. helium gas tube

4. other light source or gas tube: ______________

Part 3: Quantitative Spectroscopy

Obtain a spectroscope (with wavelength scale) from the instructor. Set-up and view the line spectra for the hydrogen gas tube. Observe and record the color and wavelength of each line.

From the wavelength, calculate the energy of each line. Use the results of your calculations to prepare a table showing the "ranking" of colors (from highest to lowest) in terms of the energy and wavelength.

Data Sheet:

Gas Tube: ___________________

Color	Wavelength (nm)	Wavelength (m)	Energy (J)

QUESTIONS:

1. Arrange in order of increasing energy: blue, orange, green, violet, red

2. Use your data to predict the color of these lines: 444 nm, 500 nm, 650 nm

3. What is the explicit relationship between energy and frequency to wavelength as the latter increases or decreases?

4. How do the values of the wavelength measured for light compare to the size of an atom (roughly $1x10^{-10}$ m)? For discussion, consider an atom to be a sphere, i.e. a Bohr atom.

5. Is the gas in a blue "neon" light actually neon? Explain.

Appendix 1. Historical Development of Spectroscopy

Max Planck (in the 1890's):

- explained thermal spectra by assuming that radiant energy (like light) was emitted in discrete packets. Planck proposed that the energy of each energy packet (E) is:

$$E = h \cdot f \quad \{\text{where f is the frequency \& h} = 6.626\text{x}10^{-34}\ \text{J}\cdot\text{s}\}$$

- did not believe this explanation nor did his contemporaries
- ultimately won the Nobel prize for this work

Albert Einstein (1905):

- explained a seemingly unrelated phenomena, the "photo-electric" effect by adopting Planck's hypothesis and thus demonstrated that radiant energy is emitted as discrete particles
- validated Planck's work and his work was thus supported by Planck

Ernest Rutherford (around 1910):

- performed experiments that led to his proposed "nuclear atom" hypothesis. He received the Nobel Prize for his contributions.
- his experimental approach transformed how experimental physics was conducted and is still implemented to this day
- ***Side Note:*** *credited with coining the term "photon" (around 1920)*

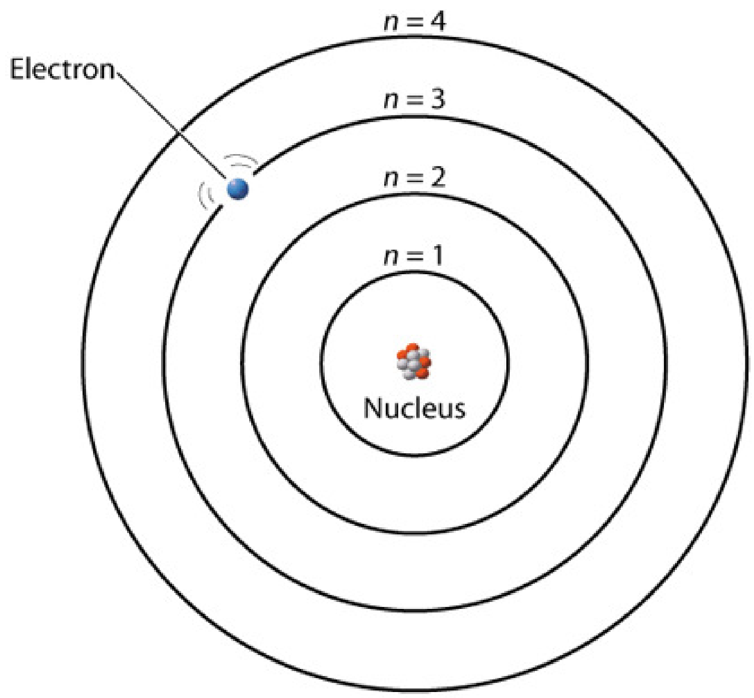
Neils Bohr (around 1915):

- Proposed a "planetary" model of the atom to explain the discrete spectra emitted by excited atoms. This model ultimately failed but led the way for the modern development of the quantum theory of the atom.
- According to this model:
  1. the electrons orbit the nucleus in circular orbits.
  2. the electrons can only exist in specific energy "states" (the orbits) that depend on the particular atom and they reside in the lowest available energy state unless excited
  3. when electrons absorb incoming radiant energy packets (i.e. photons) they jump to higher (excited) energy states, but only if the energy absorbed exactly equals the amount of energy required to be in a higher state. The transition is assumed to be instantaneous.
  4. the electrons then drop back to their lower energy states ("relax") and they release the excess energy in the form of photons, <u>exactly one photon is emitted</u> for each transition to a lower energy level.
  5. the energy of the emitted photon is exactly equal to the difference between the energy levels of the transition as the electron drops to lower energy.

Bottom Line → *by measuring the energy of the lines observed in discrete spectra for excited atoms, we can infer the electron transitions within the atom and determine relative energy associated with the allowed energy levels for those electron transitions.*

Appendix 2. Models of the Atom

The Bohr Model of the Atom



Modern Model of the Atom