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 observe visible continuous and discrete spectra of various light sources
2. To measure the wavelength of spectral lines from a hydrogen light source.
3. 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:

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

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).

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[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: Qualitative Spectroscopy

Using a spectroscope observe the following light sources. The spectroscope will allow you observe 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. *Feel free to include any additional observations you make…*

1. an incandescent light

2. a fluorescent light

3. hydrogen gas tube

4. helium gas tube

5. other light source or gas tube: ______________

Part 2: Quantitative Spectroscopy

Using a computer-interfaced spectrometer, the wavelengths associated with the line spectra for a gas tube can be measured. A spectrometer measures the light intensity over the full spectrum of wavelengths and generates a graph of the date. The objective of this section is to identify the wavelengths corresponding to the visible line spectra, observed as peak intensity values for an excited gas tube.

Procedure:

1. Turn on the computer located at your workstation.
2. Obtain a Vernier spectrometer and fiber optic cable from the instructor. Connect it to a USB port on your computer.
3. Start the LoggerPro software then click on *"Experiment"*→ *"Change Units"*→ *"Spectrometer"*. This will set the data collection to collect intensity values for the emitted light over the range of visible wavelengths.
4. Obtain a gas tube, preferable hydrogen although mercury, helium or sodium will work as well, and insert it into a power supply.
5. Click on the *"Collect"* button in the upper right hand corner of the window and observe the onscreen graph of Intensity vs. wavelength. When a well-defined pattern of spectral peaks is observed click on the *"Stop"* button to capture that set of data.
6. Observe the graph. Do the peaks agree with the expected line spectra for the gas tube. Use your observations from ***Part 1*** to answer this question. If you did not previously observe this particular gas tube in ***Part 1***, refer to the reference poster in the lab room.
7. If you are satisfied with the graph, cut-and-paste the graph into Microsoft Word. Print out a copy of the graph (in Word) for each member of your group.
8. In LoggerPro, to identify the wavelengths associated with the intensity peaks on the graph, click on *"Analyze"*→ *"Interpolate"*. A small textbox will appear indicating the values of wavelength and intensity corresponding to the cursor position on the graph.
9. Move the cursor to the top of each intensity spike then record the associated wavelength (in nm) in the table below.
10. From the wavelength in nm, convert the value to meters then calculate the photon energy for each line.
11. Using the wavelength values, calculate the energy associated with a "photon" of that wavelength.

Gas Tube:			
Color	**Wavelength (in nm)**	**Wavelength (in m)**	**Energy (J)**

QUESTIONS:

1. Arrange the following light colors in order of <u>increasing</u> energy: blue, orange, green, violet, and red. *Use your data table above as a guide.*

2. Using your data above predict the color of these spectral 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. In your own words, describe the significance of the energy values you calculated above in terms of structure and function of the atom.

5. The size of an atom, according to the Bohr Model, has a diameter of approximately $1x10^{-10}$ meters. How does the "order of magnitude" for your values of the wavelengths measured for light compare to the size of an atom?

6. What is a physical consequence of your answer to question 5?

7. Neon lights are widely used in signs and come in many colors. However, 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:*** *He is credited with coining the term "photon" (around 1920)*

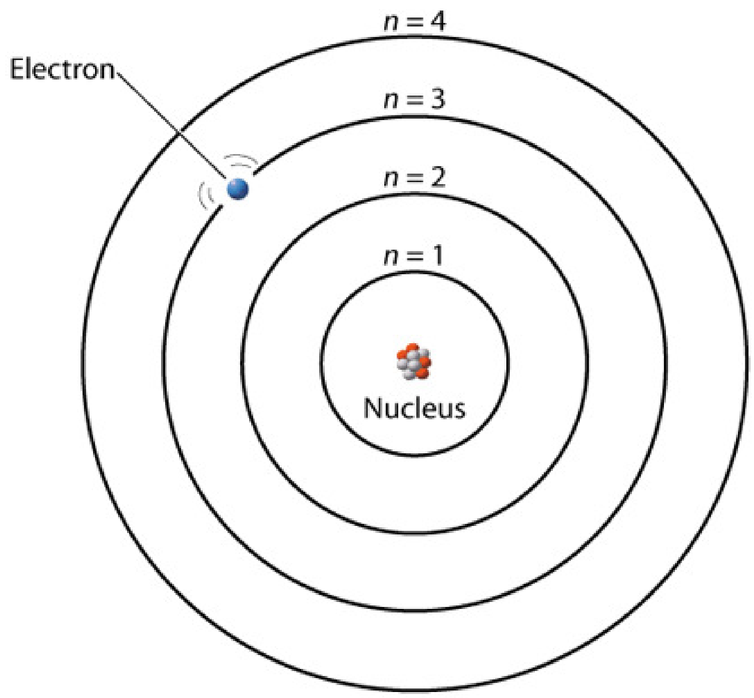
Neils Bohr (around 1915):

- 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