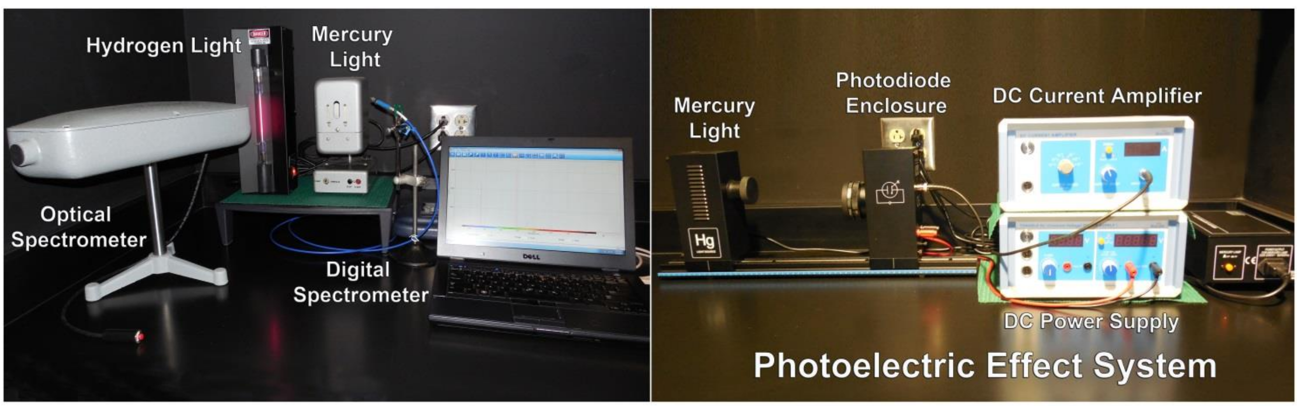
A7a: Atomic Line Spectra and Determining Planck's Constant

Introduction:

The purpose of this experiment is to introduce some of the physics that helped increase our understanding of the atom. Since the atom is of such small size, not visible by the eye, one of the main methods for obtaining information about it is through light. This will be the approach used here. The intent is to demonstrate some of what we have observed and how we explain our observations.

The sequence for this exploration will be as follows. First, observe the atomic line spectra produced by Mercury Gas using both an optical spectrometer and a digital spectrometer. Notice the differences between the two observations. Second, also using Mercury Gas, determine the constant that relates the frequency of light to the energy of light. This constant is called Planck's constant. The method used for determining the constant is called the photoelectric effect; something that Einstein received the Nobel Prize for explaining. Third, observe the spectra produced by Hydrogen Gas using both an optical spectrometer and a digital spectrometer. From the observations determine the energies of the photons producing the spectral lines. Fourth, determine the theoretical energies predicted by Bohr's model of the Hydrogen atom and compare these to the energies you determined from your observations.



Apparatus:

Optical Grating Spectrometer [Sargent-Welch model VP86950-00]
Digital Grating Spectrometer [Pasco Amadeus model SE-7183]
Optical Fiber 300nm-1100nm range [Oceans Optical model P200-2UV-VIS]
2 Spectral Tube Power Supply [Pasco model SE-9460]
Mercury Gas Spectral Tube [Pasco model SE-9466]
Hydrogen Gas Spectral Tube [Pasco model SE-9461]
Photoelectric Effect System [Pasco model SE-6609]
System includes:

Mercury Light Source Enclosure
Track, 60 cm
Photodiode Enclosure
Mercury Light Source Power Supply
DC Current Amplifier
Tunable DC (Constant Voltage) Power Supply
Filter Wheel Dial (365, 405, 436, 546, 577 nm)
Aperture Dial (2 mm, 4 mm, 8 mm diameter)

Discussion:

Our method for exploring the atom is through light. Previously we've observed diffraction of light using a diffraction grating. The experiment used a monochromatic light source from a laser and we determined the wavelength of its light. If we choose a different light source, such as light emanating from a very hot piece of metal like iron or light from the sun or light from a fluorescent gas tube lighting fixture, the observed effect from a diffraction grating will be different. The first difference to notice concerns colors. The laser produced a diffraction pattern entirely of one color. The new light sources produce diffraction patterns that contain many different colored lines. Also the different sources, such as the hot metal verses the fluorescent gas, produced very different color patterns. The question becomes, what is causing the different patterns we observe?

Early in the 1900's scientists determined that light is quantized. This means that light has a smallest size, or smallest amount, or what could be called a smallest quantity. All light is composed of discrete quantities called photons. These individual photons can each have the same energy, or they can have different energies. Your eyes cannot detect an individual photon of light. You are always seeing many photons. In the above example of the diffraction grating, when you see the different colors from the diffraction pattern, you are actually seeing the result of different energies of the photons. Each different color corresponds to different photon energies. Within one color you are also seeing many photons, but all of them have the same energy. Returning back to our different light sources and their different diffraction color patterns, the following becomes apparent. The different color patterns mean that different photon energies are being emitted from the source and therefore different energy characteristics exist for each different source.

To begin with, the source of light must initially be receiving energy in some way. For example, the solid is heated initially from a flame which increases its internal energy. Or in the case of the fluorescent gas tube, an electric potential is applied which increases the internal energy of the gas. We often refer to the increase of energy as causing a high energy state. So with the gas, the atoms in the gas are at a higher energy state. But matter doesn't like to stay at a higher energy state. It always wants to drop back to the lowest energy state possible. The method for returning to a lower energy state requires releasing energy, and this is accomplished by emitting photons of light. It is these photons of light that we are observing. The photons do not all have the same energy because the atoms releasing the energy have many paths back to a lower energy state. It depends on how much energy the atom originally absorbed and the combination of energy packets the atom releases. Each of these paths has a certain probability of occurring. Some occur frequently and others occur more rarely. The ones that occur more frequently will produce brighter lines of color due to the greater number of photons compared to the ones that occur less frequently. The spectral patterns we observe typically have some very bright lines and some very faint ones.

The remarkable thing is that for any one type of atom the probabilities are constant and therefore the spectral line pattern is always the same. But different atoms are composed of different numbers of sub-atomic particles, and this leads to a different probability distribution for each type of atom. Consequently, when you look at the atomic line spectrum pattern for a different source, the observed pattern is different. But since the pattern is the same for all atoms of one type, we can use an observed pattern to help identify the composition of a source.

At the same time that scientists were determining that light is quantized, they also determined that the energy of a photon is proportional to its frequency. So the energy doesn't vary with the intensity of light. Nor is it related to the speed of light. It's the frequency that determines the energy of the photons. Higher frequency has greater energy and lower frequency has lower energy. The constant of proportionality is Planck's constant, named for Max Planck who first published his theory in 1901.

Equation 1

$$E = h * f$$

where: E is the Energy
h is Planck's constant
f is the frequency

Many scientists contributed to our understanding of light and the atom during the early 1900's. Einstein explained the photoelectric effect and was awarded the Nobel Prize in 1921 for his explanation. The photoelectric effect refers to a phenomenon that occurs when light, having sufficient energy, is shined on some types of metal. Energy is given off by the light to atoms in the metal. Some of the electrons in the metal that received the energy are then free to move toward the surface of the metal. Once at the surface, if they have sufficient energy they may be emitted from the surface of the metal. These emitted electrons will have kinetic energy.

Equation 2

$$E = KE_{max} + W_0$$

where: E is the Energy (total Energy received by the electron)
KE_{max} is the maximum kinetic energy of the ejected electron
W_0 is the work function, the energy needed to free the electron from the material

The material being used in this part of the experiment is a photodiode that contains a type of metal that will respond to the light shining on it by emitting electrons. It is also connected to a circuit so that the emitted electrons can be measured as a current by an ammeter. In prior experiments you've worked with some basic circuits measuring electric potential (voltages) and currents (amperage). Electric potential had been defined in terms of energy.

Equation 3

$$V = \frac{E}{q_0}$$

where: V is the Electric Potential
E is the Energy
q_0 is the charge

The photodiode circuit will have an electric potential applied across it that can be varied by a power supply. It will help us control and measure the current coming from the photodiode due to the photoelectric effect. For this experiment, we need to solve *Equation 3* for energy. Also recognizing that the charge refers to the charge on the electron, we substitute the electron charge into the relationship.

Equation 4

$$E = e * V$$

where: e represents the charge of an electron

So now we have three equations related to energy: *Equation 1*, *Equation 2* and *Equation 4*. These will be combined to derive an expression that can be used to solve for Planck's constant in the experiment.

First set *Equation 1 and Equation 2* equal to each other.

$$h * f = KE_{max} + W_0$$

Next solve for kinetic energy.

$$KE_{max} = h * f - W_0$$

Now if we control the electric potential applied to the photodiode during the experiment, we can apply just enough to stop the current from flowing. In essence the electrons are still emitted from the metal but they don't have any kinetic energy once free and therefore no current. So in the equation we set the energy applied from the power supply *Equation 4* equal to the kinetic energy.

$$e * V = h * f - W_0$$

Finally isolating the electric potential, which in this case is commonly referred to as the stopping potential, we have the expression that can be used in the experiment to solve for the Planck's constant.

Equation 5

$$V = \frac{h}{e} * f - \frac{W_0}{e}$$

where: V is the stopping potential

The photoelectric effect apparatus will allow you to collect stopping potentials for five of the wavelengths in the spectrum of Mercury. This will be graphed as a function of the frequency. The slope of the line from this graph is equivalent to (*h/e*).

Earlier in the term you conducted an experiment involving mechanical waves. The relationship between the speed of the wave, the wavelength and the frequency was examined. The same relationship also applies here for light.

Equation 6

$$c = \lambda * f$$

where: c is the speed of light
λ is the wavelength
f is the frequency

With the wavelengths you'll measure during the experiment, *Equation 6* can be applied to solve for the frequencies.

Niels Bohr presented a model of the Hydrogen atom around 1913 that incorporated several of the ideals that had begun to emerge in physics at the time. Although his model is not accurately correct by today's knowledge, it is still a good introductory approximation within its limits. In brief, the model looked at the Hydrogen atom as a positively charged nucleus with a negatively charged electron orbiting the nucleus in a circular orbit. The electrostatic force between the nucleus and the electron supplied the centripetal force necessary to keep the atom from flying apart. Also the velocity that the electron orbited the nucleus was perfectly matched to keep it from either spiraling into the nucleus or escaping from it. Another assumption of the model was that the electron could not assume an arbitrary orbit, but was limited to specifically defined, discrete orbits. The electron could jump to a higher orbit only by receiving energy. Thus a photon striking the atom could give exactly the right amount of energy sufficient to jump the electron to a higher orbit, but it must correspond to an allowed orbit. Each of these allowed orbital levels correlate to specific hydrogen energy levels. These energy levels for the hydrogen atom can be calculated using the following expression.

Equation 7

$$E_\eta = \frac{-13.6}{\eta^2}(eV)$$

where: η is an integer = 1,2,3,…
(eV) means electron Volts

In order for the atom to return to a lower energy level, it must release energy. Since only specific orbits, and therefore energy levels, are allowed, the energy released is also defined in specific amounts. During your experiment, use *Equation 7* to calculate the first seven energy levels for the hydrogen atom. The differences between these energy levels are the transition energies and would correspond to the energies of the photons being released by the hydrogen.

Please read the relevant material in your textbook for this experiment.

Photoelectric Effect	Cutnell & Johnson. **Physics**, *Chapter* **29** *section* **2, 3**
Atomic Line Spectra	Cutnell & Johnson. **Physics**, *Chapter* **30** *section* **2, 3**

Procedure & Analyses:

The spectrometers, gas emission tubes and photoelectric effect system should already be assembled. Please see one of the lab instructors for guidance on proper use of the equipment and safety precautions. Please make sure all of the equipment is turned off when you finish the experimental portion of this lab.

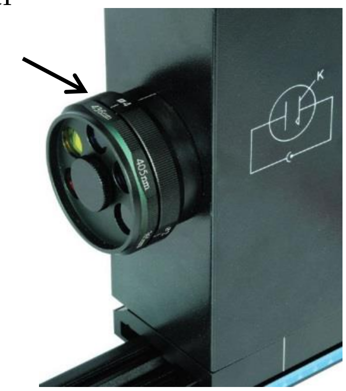
Mercury

1. Turn on the Mercury light source and observe its line spectra using the optical spectrometer. Fill in Table 1 by drawing in each of the lines at the appropriate wavelength and including the line color.
2. Observe the Mercury light source using the digital spectrometer. Have one of the lab assistants demonstrate the software and the method for making measurements. Measure the wavelengths and intensities, using the computer software tool, for all observable lines between 350nm and 800nm. Fill in Table 2 with the wavelengths and relative intensities. **Turn off the Mercury light source once measurements are completed**.
3. Calculate the frequency for each observed wavelength in Table 2.

Photoelectric Effect System

4. Check that the caps are on both the Mercury Light Source Enclosure and the Photodiode Enclosure. Turn on the Mercury Lamp Power Supply. This apparatus must warm-up for 10 minutes before beginning to collect data.
5. Turn on the Tunable DC Power Supply and the DC Current Amplifier.
6. Wait 10 minutes for warm-up before continuing.
7. On the Tunable DC Power Supply set the Voltage Range button switch to (-4.5V to -0V) range; note the push button will be out.
8. On the DC Current Amplifier, set the Current Ranges Switch to ($10^{-13}A$) range. Push the Signal button to the in position for Calibration. Adjust the Current Range dial until the ammeter shows the current is zero. Push the Signal button to the out position for Measure.

9. On the Photodiode Enclosure, rotate the aperture wheel so that the 4mm diameter aperture is aligned with the white line on the enclosure. Rotate the filter wheel until the 365nm filter is also aligned with the white line. Uncover the photodiode by removing the cap covering the enclosure's window.



WARNING the Mercury Light Source Enclosure is very **HOT**, please don't touch it!

10. Carefully uncover the window of the Mercury Light Source being cautious as to not burn your hand.
11. On the Tunable DC Power Supply, carefully adjust the Voltage Dial until the ammeter on the DC Current Amplifier shows the current is exactly zero.
12. Record the voltage measurement from the DC Power Supply voltmeter in Table 3.
13. Rotate the filter wheel on the Photodiode Enclosure until the 405nm filter is aligned with the white line on the enclosure.
14. Carefully adjust the Voltage Dial on the DC Power Supply until the ammeter on the DC Current Amplifier shows the current is exactly zero.
15. Record the voltage measurement from the DC Power Supply voltmeter in Table 3.
16. Repeat steps 13, 14 & 15 for each of the other filters on the Photodiode Enclosure filter wheel (436nm, 546nm, 577nm).
17. **Turn off the Mercury Lamp Power Supply. Also turn off the DC Power Supply and the DC Current Amplifier**. Replace the covers over the Mercury Light Source Enclosure window and the Photodiode Enclosure window.
18. For each of the wavelengths on Table 3, **copy the corresponding frequencies from Table 2** according to the closest matching wavelength.
19. Using Excel, plot a graph of the Stopping Potential as a function of the Frequency. Determine the statistical slope of the best fit line.
20. Use the slope to calculate Planck's constant. Compare your experimental value to the accepted value for Planck's constant and calculate the percent error.

Hydrogen

21. Turn on the Hydrogen light source and observe its line spectra using the optical spectrometer. Fill in Table 4 by drawing in each of the lines at the appropriate wavelength and including the line color.
22. Observe the Hydrogen light source using the digital spectrometer. Measure the wavelengths and intensities, using the computer software tool, for all observable lines between 375nm and 700nm. Fill in Table 5 with the wavelengths and relative intensities. **Turn off the Hydrogen light source once measurements are completed.**
23. Calculate the frequency for each observed wavelength in Table 5.
24. Calculate the energy for each wavelength in Table 5 **using the Planck's constant that you calculated on step 20**. First calculate in units of joules and then convert into units of electron-Volts. Note these are the energies of the observed photons.

Hydrogen Theoretical

25. Calculate the first seven theoretical energy levels for the Hydrogen atom filling in Table 6.
26. Calculate the difference in the energy levels for the transitions listed in Table 7.
27. Compare the theoretical transition energies from Table 7 to the experimental photon energies from Table 5. Complete the comparison in Table 8.
28. Identify the appropriate Hydrogen Series by name and complete an Energy Level diagram for the series.

Experiment : A7a: Atomic Spectra & Planck's Constant

Student Name____

*Lab Partner Name*____

*Lab Partner Name*____

Physics Course____

Physics Professor____

Experiment Start Date__________________________________

Lab Assistant Name	*Date*	*Time In*	*Time Out*

Experiment Stamped Completed

Data Sheet 1: A7a: Atomic Spectra & Planck’s Constant

NAME: ____________________________ **DATE:** ___________________

Table 1: Mercury (observed with the optical spectrometer)

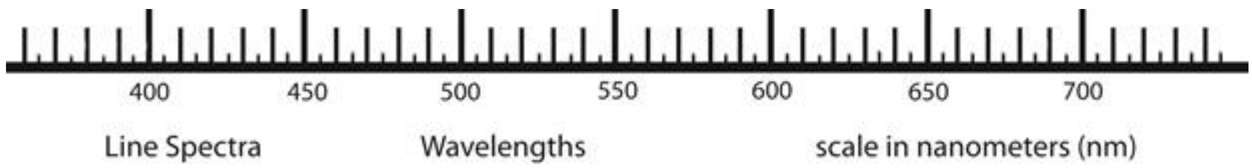


Table 2: Mercury

Wavelength	**Intensity**	**Frequency**
nanometer (nm)	photon count	hertz (Hz)

Date Modified 08/05/20

Data Sheet 2: A7a: Atomic Spectra & Planck's Constant

NAME: ______________________________ **DATE:** ____________________

Table 3: Mercury (photoelectric effect/Planck's constant apparatus)

Filter Wavelengths	Frequency	Stopping Potential
nanometer (nm)	hertz (Hz)	volt (V)
365		
405		
436		
546		
577		

Slope from graph of Stopping Potential as a function of Frequency:

Calculated value for Planck's constant:

Theoretical standard value for Planck's constant:

Percent Error:

Data Sheet 3: A7a: Atomic Spectra & Planck's Constant

NAME: ____________________________ **DATE:** ___________________

Table 4: Hydrogen (observed with the optical spectrometer)

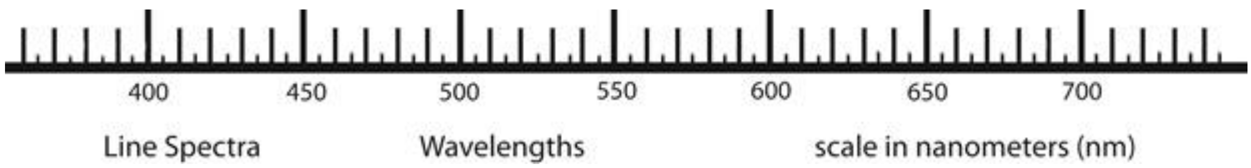


Table 5: Hydrogen (observed with the digital spectrometer)

Wavelength	**Intensity**	**Frequency**	**Energy**	**Energy**
nanometer (nm)	photon count	hertz (Hz)	joule (J)	electronVolt (eV)

Data Sheet 4: A7a: Atomic Spectra & Planck's Constant

NAME: ____________________________ **DATE:** ___________________

Table 6: Theoretical Energies for the Different Energy levels in Hydrogen

η	E_η (eV)
1	
2	
3	
4	
5	
6	
7	

Table 7: Differences between the Theoretical Energies levels in Hydrogen

Transition	ΔE (eV)
$E_2 - E_1$	
$E_3 - E_2$	
$E_4 - E_3$	
$E_4 - E_2$	
$E_5 - E_2$	
$E_6 - E_2$	
$E_7 - E_2$	

Data Sheet 5: A7a: Atomic Spectra & Planck's Constant

NAME: ____________________________ **DATE:** ___________________

Table 8: Comparative Energies between Theoretical and Experimental Values

Theoretical Energy	**Transition**	**Experimental Energy**
(eV)		(eV)
	$E_3 - E_2$	
	$E_4 - E_2$	
	$E_5 - E_2$	
	$E_6 - E_2$	