Exp. 8 Pre-Lab ASSIGNMENT **Name: __________________________**

Lab Section____________ Score: _____ / 10

(1) Use the equations [see the discussion on the next page] to calculate the energies of the 6 lowest states for the hydrogen atom, and enter your answers in Table 1.
(2) Calculate the energy differences between each pair of the six lowest levels and place them in the upper halves of the appropriate boxes in Table 2.
(3) Use equation (1) [see the discussion on the next page] to calculate the wavelength corresponding to each energy difference calculated in part (2). Convert your wavelengths to nanometers and place them in the bottom halves of the appropriate boxes in Table 2.

Table 1: Energy Levels of the Hydrogen Atom (from Equation 2) *(Watch your sig. figs.!)*

Quantum number, n	Energy, E, in J	Quantum number, n	Energy, E, in J
1	________________	4	________________
2	________________	5	________________
3	________________	6	________________

Table 2: Use the energy differences and to calculate the wavelengths (nm) of the following transitions.

(report to 4 sig. figs.!)

$6 \rightarrow 1$ _____________	$6 \rightarrow 2$ _____________	$6 \rightarrow 3$ _____________	$6 \rightarrow 4$ _____________
$5 \rightarrow 1$ _____________	$5 \rightarrow 2$ _____________	$5 \rightarrow 3$ _____________	$5 \rightarrow 4$ _____________
$4 \rightarrow 1$ _____________	$4 \rightarrow 2$ _____________	$4 \rightarrow 3$ _____________	$6 \rightarrow 5$ _____________
$3 \rightarrow 1$ _____________	$3 \rightarrow 2$ _____________		
$2 \rightarrow 1$ _____________			

Use the internet to determine what compounds used commonly in fireworks give specific colors, list three compounds and the corresponding colors.

Compound	Color
__________________________	__________________________
__________________________	__________________________
__________________________	__________________________

What are the energies (J) and wavelengths (in nm) for these colors?

Color	Energy	wavelength
_________________	_________________	_________________
_________________	_________________	_________________
_________________	_________________	_________________

ATOMIC SPECTRA

INTRODUCTION

At room temperature essentially all atoms are in their ground (lowest-energy) electronic state. When heated or placed in an electric discharge, electrons in the atoms are excited to higher energy levels. As the electrons relax to lower levels, light is emitted. The frequency of the light emitted is specific to the transitions between the higher and lower energy levels. The photon emitted must have the following energy based on Planck's Law:

$$E_{photon} = E_{upper} - E_{lower} = \frac{hc}{\lambda}$$

Where h (Planck's constant) = 6.626076×10^{-34} Js, c (speed of light in a vacuum) = 2.997925×10^{8} m/s, and λ= the wavelength of the photon.

A few of the energy levels of the hydrogen atom are shown below:

Energy	O (zero)	____________________	n = ∞ (Ionization occurs)
	-2.421×10^{-19} J	____________________	n = 3 (2^{nd} excited state)
	-5.448×10^{-19} J	____________________	n = 2 (1^{st} excited state)
	-2.179×10^{-18} J	____________________	n = 1 (ground state)

The energy of the photon emitted when the electron in the hydrogen atom makes a transition from the first excited state to the ground state is:

$$E_{photon} = -5.448 \times 10^{-19}\text{ J} - (-2.179 \times 10^{-18}\text{ J}) = 1.634 \times 10^{-18}\text{ J}$$

The wavelength of the photon is:

$$\lambda = \frac{hc}{E_{photon}} = \frac{6.626 \times 10^{-34}\text{ J}\cdot\text{s} \times 2.998 \times 10^{8}\text{ m/s}}{1.634 \times 10^{-18}\text{ J}} = 1.216 \times 10^{-7}\text{ m} \times \frac{1 \times 10^{9}\text{ nm}}{1\text{m}} = 121.6\text{ nm}$$

This particular wavelength is in the ultraviolet region of the electromagnetic spectrum and belongs in the Lyman series of electronic transitions of the hydrogen atom.

A hydrogen atom can also emit light at two other wavelengths (λ_{3-1} and λ_{3-2}) as well as light at many more wavelengths involving energy levels not included in the diagram above. The set of all wavelengths ("lines") emitted by an atom is known as its "atomic spectrum." Since each element has a unique set of electronic energy levels, the atomic spectrum for each element is also unique. As a result, elements can readily be identified from their atomic spectra. In the first part of this experiment, you will use a spectroscope to view lines in the visible portions of several atomic spectra.

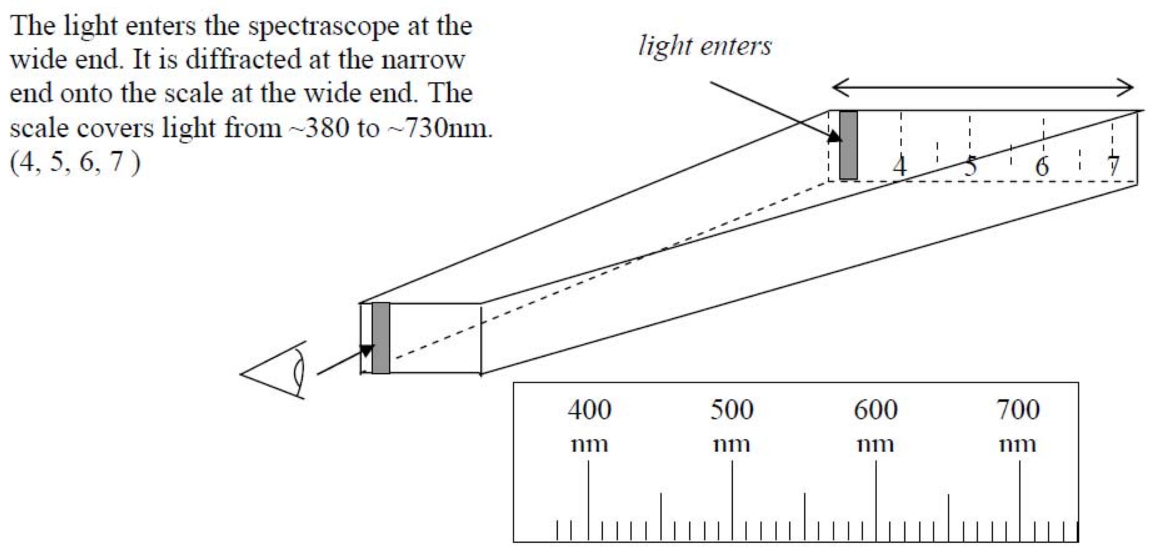
The hydrogen atom has a particularly simple spectrum. According to the Bohr Theory, the energy levels of the hydrogen atom are given by the equation

$$E_n = -\frac{Rhc}{n^2}$$

Where R = 1.097×10^{7} m^{-1}, h is Planck's constant (6.626×10^{-34} J·s), c is the speed of light (2.998×10^{8} m/s) and n = 1, 2, 3 (any positive integer). The lines in the atomic spectrum of hydrogen are grouped into "series," where each member of a series involves a transition to the same lower energy level. In the pre-lab for this experiment you will calculate the energy for the first six levels of the hydrogen atom, and use these energies to calculate the wavelengths of some lines in several different spectral series.

EXPERIMENTAL: Part 1 Line spectra of elements

Working in groups of three or four, view the emission spectra from the discharge tubes set up in the room containing hydrogen and an additional element using a spectroscope. Draw a sketch of the lines you see on the data sheet provided. Indicate the colors observed and label each line in your sketch with its wavelength in nm. You will assign quantum numbers to the lines observed in the H-atom spectra later in the post lab assignment.



EXPERIMENTAL: Part 2 Flame Test Identification

Working in pairs you will observe the emission colors of several aqueous solutions containing certain metal cations.

Procedure:

- Place approximately 0.5 mL of the following solutions into small labeled test tubes: $LiCl$, $CuCl_2$, KCl, $BaCl_2$, $NaCl$, $SrCl_2$, $CaCl_2$
- Obtain a nichrome wire with a cork holder from the front of the room.
- Carefully ignite the Bunsen burner.
- Dip the metal end of your nichrome wire in the 6M HCl solution then into a clean small beaker containing deionized water. Heat the wire in the hottest part of the flame (blue tip). Repeat twice. This will clean off your wire so that any metal contaminates present on the wire will be removed. You will need to repeat this step before examining each different solution. If the HCl solution turns yellow, you will need to replace it with a fresh solution.
- Perform the flame tests on the known solutions: Dip the wire into your first solution and place it in the hottest part of the Bunsen burner flame. Note the color of the flame and record your observations on your data sheet. You should perform the test a few times for each solution. It is important to test the solutions of sodium chloride and lithium chloride last, as these solutions can sometimes give false positives if used first.
- Clean your wire with the 6M HCl as described. Repeat the tests on the remaining solutions. Record your observations on the data sheet. Make sure to clean the wire between solutions.
- Once your data sheet is complete, please show it to your lab instructor to obtain your unknown samples.
- Each student will obtain two unknowns from the lab instructor and individually identify the metal cations. Perform the flame tests on each unknown, record your observations on the data sheet. ***Don't forget to record your unknown numbers on the data sheet.***

DATA AND CALCULATIONS **Name: __________________________**

This Experiment is due at the end of lab. **Lab Section___________** **Score: _______ / 25**

Submit this and the following pages:

Part 1:

Sketches of Line Spectra (Complete the sketch for hydrogen and any one of the three other elements.) Label the color for each line spectrum.

(1) Hydrogen

400nm	500nm	600nm	700nm

(2) Element 2:

400nm	500nm	600nm	700nm

1. How many kilojoules of energy are required to ionize a mole of hydrogen atoms from their ground states?

ans.: ______________________

2. In 1886, Balmer showed that the lines he observed in the visible and near ultraviolet region of the hydrogen atom spectrum had wavelengths which could be expressed by the following equation: Where C is a constant.

$$\frac{1}{\lambda} = C\left(\frac{1}{n_{lower}^2} - \frac{1}{n_{upper}^2}\right)$$

(a) Based on your results in Table 2, what is nlower for this group of lines, now known as the Balmer series?

ans.: ______________________

(b) If wavelength is expressed in nanometers, what is the numerical value of the constant C in Balmer's equation?

ans.: ______________________

(c) What is the longest possible wavelength for a line in the Balmer series?

ans.: ______________________

(d) Calculate the shortest wavelength that a line in Balmer series could have.

ans.: ____________________

(e) The lines you sketched in part A for hydrogen belong to the Balmer series. Assign n_{upper} for each line in your sketch (on the previous page). (If you observed a yellow line, ignore it. It arises from sodium atoms in the glass.)

Part 2: Flame Test

Solution	Observations
Li^{+}	
Cu^{2+}	
K^{+}	
Ba^{2+}	
Na^{+}	
Sr^{2+}	
Ca^{2+}	
Unknown #	
Unknown #	

Based on your observations, what is the identity of the cation in each of your unknowns?

Unknown #_______ Identity_________________________________

Unknown #_______ Identity__________________________________