

# Quantum States and Spectra of Gases

## ONE WEIGHT

### RECOMMENDED READINGS

- 1) R. Harris: Modern Physics, Ch.4, pp. 122-124; Ch.7 pp. 238-268; Ch.8.4-8.9. Addison Wesley, 2008.
- 2) PASCO Wireless Spectrometer Manual  
[https://www.pasco.com/file\\_downloads/Downloads\\_Manuals/Wireless-Spectrometer-Manual-PS-2600.pdf](https://www.pasco.com/file_downloads/Downloads_Manuals/Wireless-Spectrometer-Manual-PS-2600.pdf).
- 3) PASCO video Guide <https://youtu.be/i5BexMng2WY>.
- 4) Glow Discharge review [https://en.wikipedia.org/wiki/Glow\\_discharge](https://en.wikipedia.org/wiki/Glow_discharge).

### INTRODUCTION

A spectrum is a function  $I(E)$ , where  $I$  is usually called intensity and is proportional to the number  $N(E)$  of photons of energy  $E$ , registered per a second by a spectrometer. The energy may be replaced by a frequency  $f$  or by a wavelength  $\lambda$ , so that the spectrum is either  $I(E)$ , or  $I(f)$ , or  $I(\lambda)$  (Fig. 1).

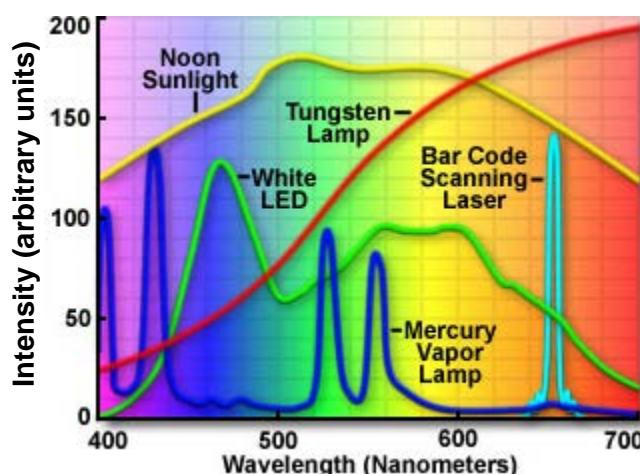


FIG. 1: Intensity of different sources of visible light.

<http://www.olympusmicro.com/primer/lightandcolor/lightsourcesintro.html>

To numerically find this function for all energies been registered, the photons of different energies are spatially separated by a dispersive device: a prism, a diffraction grating (in this experiment) or a crystal. Resolution of the device can be qualitatively treated as an ability to distinguish photons with very close values of energy as different photons.

The spectra of Fig.1 can be referred to either continuous (noon, sunlight, tungsten lamp and white LED) or linear (a laser and a mercury vapor lamp). From the point of view of quantum mechanics, all atomic spectra must be the linear ones, because every photon has a specific energy

$$E_f = hf = hc / \lambda , \quad (1)$$

where  $h = 6.626070040(81) \times 10^{-34}$  J·s, is the Planck's constant; and  $c$  is the speed of light in vacuum. Emission of photons is a result of transition of an atom or a molecule from one quantum state to the other with smaller energy. Absorption of a photon occurs when the atom undergoes a transition to a higher energy level.

In this experiment, you will study the emission spectra produced by a number of gas discharge tubes and absorption spectra of different solutions. Emitted light is incident on a diffraction grating and then passes a CCD array each element of which detects specific incident energy.

A glow discharge is plasma (a mixture of free electrons and ions) produced in a gas by collisions between the accelerated electrons and the atoms of the gas under high voltage and low pressure. Description of regions of the glow discharge in the gas discharge tube is demonstrated in Fig. 2.

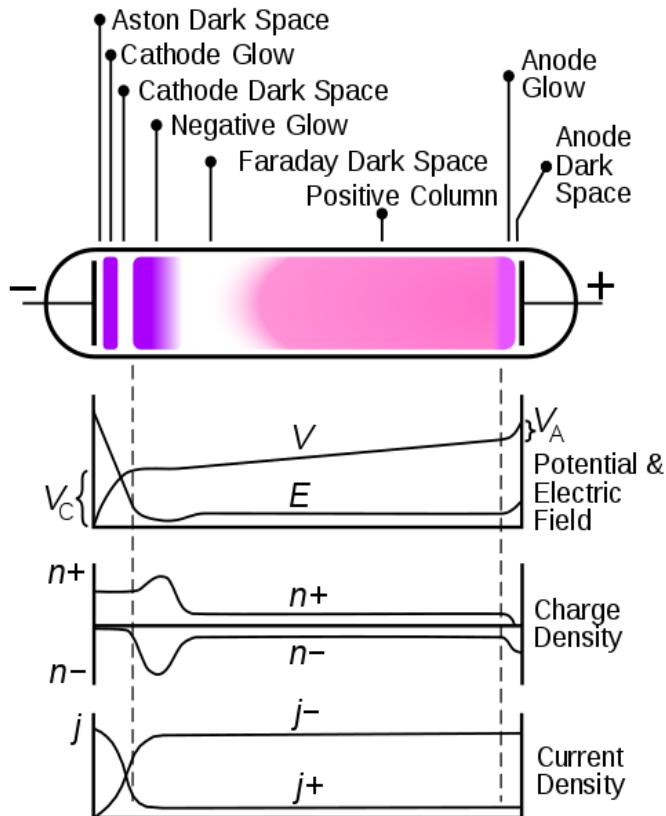


FIG. 2: Glow discharge in a gas discharge tube.

[https://commons.wikimedia.org/wiki/File:Glow\\_discharge\\_structure\\_English.svg](https://commons.wikimedia.org/wiki/File:Glow_discharge_structure_English.svg)

High voltage applied between the anode and the cathode of the tube creates current of electrons (cathode rays) accelerated toward the anode. Collisions with atoms of a pure gas inside the discharge tube entail ionization of the atoms. Positively charged ions are moving to the cathode thus increasing the number of collisions among the particles in partially ionized plasma – a state of matter when electrons and ions coexist in a mixture and the number of electrons captured by the ions is equal to the number of electrons released due to ionization. This number stays constant for a particular pressure and excitation energy (in this experiment – the applied voltage).

When an incident electron or an ion ionizes the neutral atom, the quantum state of the latter changes, the atom transits to the excited state with its total energy increased. A vacancy in either of the inner shell is produced and can be then filled by one of the electrons of the other electron shells of the atom. The electron transition is accompanied by emitting a photon and decreasing the ion energy. Thus, if we know the energy of the two states participating in the transition, we can expect the energy of the emitted photon to equal the difference of the energies of the states. This modern understanding of discrete quantum states of an atom and the nature of photons is based on genius contribution into quantum physics by Max Plank and Niels Bohr. The empirical Rydberg equation [Johannes Rydberg] and the Bohr's theory of atom result in the relationship between the energy of the emitted photon and the particular transition between the two quantum states with

principal quantum numbers  $m$  and  $n$ . For the hydrogen atom and a *hydrogen-like ion* with  $Z$  protons in the nucleus and one orbiting electron, e.g.  $\text{He}^+$ , this gives:

$$hf = R_{EH} Z^2 \left( \frac{1}{m^2} - \frac{1}{n^2} \right) \quad (2a)$$

where  $R_{EH}$  is the Rydberg constant for hydrogen, and  $R_{EH} = 13.605693 \text{ eV} = 2.179872 \times 10^{-18} \text{ J}$ . The above conversion of energy units shows that the **electronvolt** (eV) is a very convenient unit and is widely used in particle, atomic and nuclear physics. The conversion from **electronvolts** (eV) to **joules** (J) is given by

$$1 \text{ eV} = 1.60218 \times 10^{-19} \text{ J.}$$

A formula similar to (2a) was empirically obtained by Johann Balmer in 1885 for a visible spectrum of hydrogen and describes so called Balmer series as:

$$hf = R_{EH} \left( \frac{1}{2^2} - \frac{1}{n^2} \right) \quad (2b)$$

After replacing frequency by wavelength, the Eqs. (2) can be rewritten as:

$$\frac{hc}{\lambda} = R_{EH} Z^2 \left( \frac{1}{m^2} - \frac{1}{n^2} \right) \quad \frac{1}{\lambda} = \frac{R_{EH}}{hc} Z^2 \left( \frac{1}{m^2} - \frac{1}{n^2} \right) = R_H Z^2 \left( \frac{1}{m^2} - \frac{1}{n^2} \right) \quad (3)$$

where  $R_H = 10973731.57 \text{ m}^{-1}$  and is the Rydberg constant for hydrogen in different units. The transition takes place from the excited state with a principal quantum number  $n$  to the ground state with the principal quantum number  $m$ .

Equation (3) makes it possible to calculate the principal quantum numbers of the states for the measured wavelength of a spectrum line.

Equations (2) permit calculating the value of energy of the state with a particular principal quantum number.

To calculate energy of any stationary state of the hydrogen-like atom, we use two Bohr's postulates as the cornerstone of his model. Basing on the Bohr's postulates, the energy  $E_n$  of the  $n$ -th level of the hydrogen-like atom is given by

$$E_n = -\frac{Z^2 k_e e^2}{2n^2 a_0} \approx -\frac{13.6 Z^2}{n^2} \text{ eV} \quad (4)$$

where  $k_e$  is Coulomb's constant;  $e$  is the elementary charge, and  $a_0$  is the Bohr radius;  $a_0 = 0.0529 \text{ nm}$ . The Bohr's model works for just hydrogen and hydrogen-like ions that have one electron.

The equations (2) - (4) make it clear that any atom has a specific, or characteristic, set of spectral lines that correspond to transitions of this atom between the two specific quantum states. The atomic spectrum is therefore called *characteristic spectrum*. The characteristic spectrum is a tool to determine elemental composition of any substance.

Except the explained way of exciting atoms of a substance in the gaseous state with the technique of the glow discharge, the atoms can be excited by incident photons with energies exceeding the energy of excitation called *ionization potential*. The resulting emission is called *fluorescence*. The *ionization potential*, or *ionization energy*, is actually work required to move an electron from a specific shell and with specific quantum numbers to infinity, and it is usually presented in the units of *electronvolts*. The ionization potential of the hydrogen atom is roughly 13.6 eV.

Intensity of a spectral line depends on the probability of the transition, the lifetime of the excited state, the probability of fluorescent yield and on the absorption of the photons in a medium.

Regardless the way of excitation, there is always many peaks presented simultaneously with different intensities in the emission spectrum. This is due to a nonzero probability for an incident photon to excite the atom at any state with the ionization potential lower than the energy of the photon. With almost infinite number of atoms in the discharge tube, an observer can find all peaks of the characteristic spectrum of the sample, each of the peak with its own intensity.

## EQUIPMENT

The setup includes

- PASCO wireless spectrometer PS-2600, connected to a desktop with the USB cable; may be connected also to a smartphone via Bluetooth.
- Fibre optics cable PS-2601 with a probe directed to the light source; specifications are here [https://www.pasco.com/prodCatalog/PS/PS-2601\\_fiber-optics-cable/index.cfm](https://www.pasco.com/prodCatalog/PS/PS-2601_fiber-optics-cable/index.cfm).
- Set of gas discharge tubes; <https://www.pasco.com/prodGroups/spectral-tubes/index.cfm>. We recommend you to open the internet source and find the safety sheets for the gas tubes you are using in this experiment. **A discharge tube should not have current to run through them for more than 30 minutes continuously!**
- High voltage power supply for the gas discharge tubes.
- Set of cuvettes for water solutions of different substances.
- Software to freely download to home computer and a smartphone <https://www.pasco.com/downloads/spectrometry/index.cfm>.

## EXPERIMENT

**ATTENTION!!** EVERY TIME YOU FINISH THE DATA ACQUISITION, SAVE YOUR FILE WITH THE SPECTRUM IN TWO FORMATS: IN THE FORMAT OF THE SPECTROMETER SOFTWARE AND IN ANY OTHER, LIKE PAINT BRUSH, THAT WILL PERMIT YOU TO WATCH THE DATA FROM ANY COMPUTER WITHOUT INSTALLING THE SOFTWARE!!!

Part I. Calibration of the spectrometer with a mercury (Hg) tube.

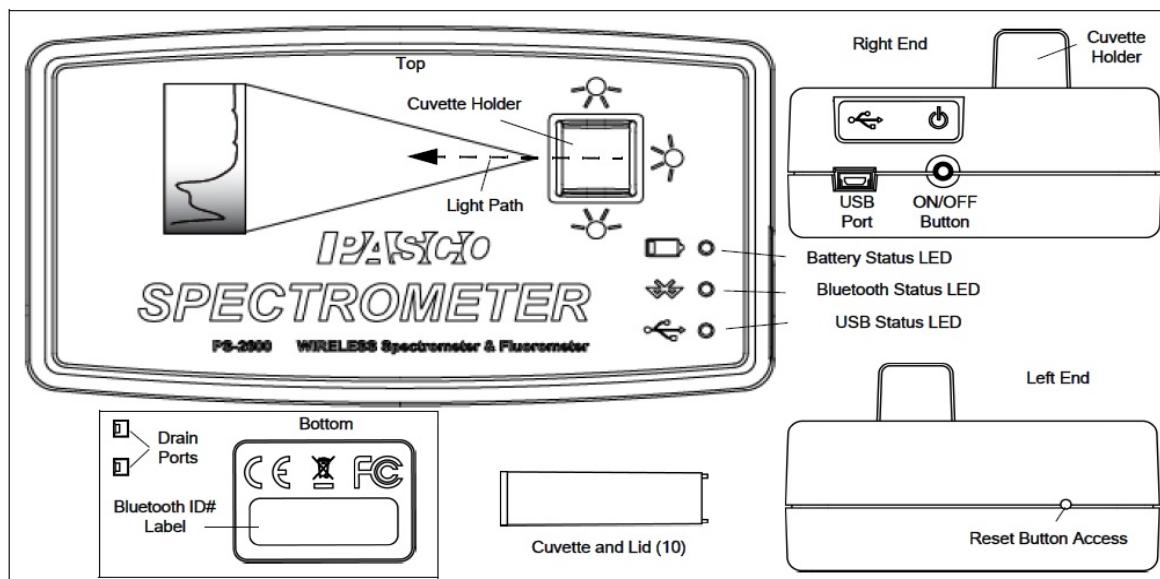


FIG. 3: Spectrometer PS-2600.

1. Begin with reviewing the Wireless Spectrometer Manual [ftp://ftp.pasco.com/support/documents/english/PS/PS-2600/013-14074B%20PS-2600.pdf](http://ftp.pasco.com/support/documents/english/PS/PS-2600/013-14074B%20PS-2600.pdf). Connect the spectrometer to the laboratory desktop with a USB cable. You may connect your tablet or smartphone with Bluetooth in parallel with the wired connection to the desktop. Get familiar with LED's functions and determine whether the spectrometer battery is charged. Charging the battery takes more than one hour, so, don't begin this procedure. Instead, request a replacement battery. Turn the spectrometer off when you stop taking measurements.
2. Turn the spectrometer on by the ON/OFF Button: all three LEDs must glow. Study functions of all icons on the spectrometer display with the APPENDIX 1.
3. The power supply for a discharge tube must be unplugged. Carefully insert a Hg gas tube into the power supply to plug it in on both ends. The connection is loose.
4. Place the rectangular end of the fibre optics cable in the cuvette opening of the spectrometer. The arrow on top of the rectangular end must show the direction toward the built-in detector, which is the direction toward the colored spectrum on the top of the spectrometer box. Fix the probe end (a black cylinder with the rounded end) of the fibre optics cable in the clump vertically above the tube and point the probe to the tube center at about 0.5 cm from the tube. You may improve the spectrum intensity later by slightly adjusting this distance and moving the probe along the tube. Never touch the gas tube with the probe end.
5. Plug the wire of the power supply into the net. Turn on the power supply with a big red switch on its side.

**Safety:**

- Don't touch the spectrum tube or the high-voltage power supply after turning it on!
- Use caution around the high-voltage power supply!

6. Open software by clicking on the icon  on the desktop:
7. Select ANALYZE LIGHT function on the spectrometer panel of application tools (see the list of application tools in APPENDIX 1).
8. Begin a test spectrum recording by pressing the Start Record button. To set up integration time, press the Autoset button under the Integration Time function in the left-hand tools panel. This function automatically selects the best time of data acquisition, and you should press the Autoset at the beginning of every run. Adjust the distance between the probe and the tube, number of scans to average and number of points for smoothing on the left of the screen. In your report, explain how the shape and height of the spectrum line change during your tuning of the index of smoothing. Give your guess on the origin of the change.
9. After about 2 minutes after all adjustments are made, stop taking the data.
10. Turn off the power supply of the gas discharge tube. Unplug the cable of the power supply.
11. The spectrum of mercury will include peaks whose wavelengths are given in Table 1 as the expected values.
12. Press SCALE TO FIT or expand/compress a scale on the axes by a cursor if required.
13. Use the Reference arrows at the bottom-right of the screen to compare your spectrum with the standard one.

Table 1. Strong spectral lines of mercury. The data for wavelengths  $\lambda$  are taken from  
[https://physics.nist.gov/PhysRefData/Handbook/Tables/mercurytable2\\_a.htm](https://physics.nist.gov/PhysRefData/Handbook/Tables/mercurytable2_a.htm)

Color	Violet	Violet	Blue	Green	Yellow	Yellow
Expected $\lambda$ , nm	404.6565	407.7837	435.8328	546.0735	576.9598	579.0663
Expected energies, eV						
Experiment $\lambda$ , nm						
Experiment energy, eV						

14. Press ADD COORDINATE and drag the square with a small blank window to each maximum to obtain its wavelength and intensity. Since the spectrometer resolution is 2 - 3 nm according to the manufacturer's manual, not all peaks may be completely resolved, i.e. two close lines may look like one peak.
15. You can see that the spectrometer was actually calibrated by the manufacturer, however, the energies found by the software may not coincide with those in the Table 1. Copy the table to your report and fill in the row for your experiment wavelengths. Save the reasonable number of significant figures according to the given resolution of the spectrometer.
16. Using Eq. (1), convert wavelengths into energies of the lines in electronvolts for expected and measured values.
17. Calibrate the spectrometer for true wavelength vs measured wavelength. Find the slope and y-intercept with their uncertainties and the goodness of fit criteria (reduced chi-squared and residuals).
18. Calibrate the spectrometer to find true energy vs measured energy as a linear function. Find the slope and y-intercept with their uncertainties and two goodness of fit criteria.

## Part II. Calculating energy of the quantum states of hydrogen (H).

In your report, all data of this exercise must be organized in a table.

After the mercury tube has cooled down, replace it by the hydrogen tube. Take a spectrum of the hydrogen discharge tube. It will look like the one in Fig. 4. All lines belong to the Balmer series of hydrogen (see Eq. (2b)). Turn off the power supply and unplug the wire.

The other sources can report the other intensity distribution in the spectrum of hydrogen (see <https://www.vernier.com/innovate/measure-gas-discharge-tube-emissions/> ).

➤ **Why may the relative intensities of the lines vary for the same atoms in the discharge tube from experiment to experiment?**

1. Calculate energies of *quantum states* of a hydrogen atom with  $n = 3, 4$  and  $5$ , using Eq. (4).
2. Calculate the expected energy of *spectral lines* of Balmer series with different  $n$  using Eq. (2b).
3. Convert the theoretically obtained energies of the *spectral lines* into wavelengths in nm.
4. Compare the experimental and theoretical values for the spectrum peaks' energy and discuss the difference considering all experimental uncertainties.

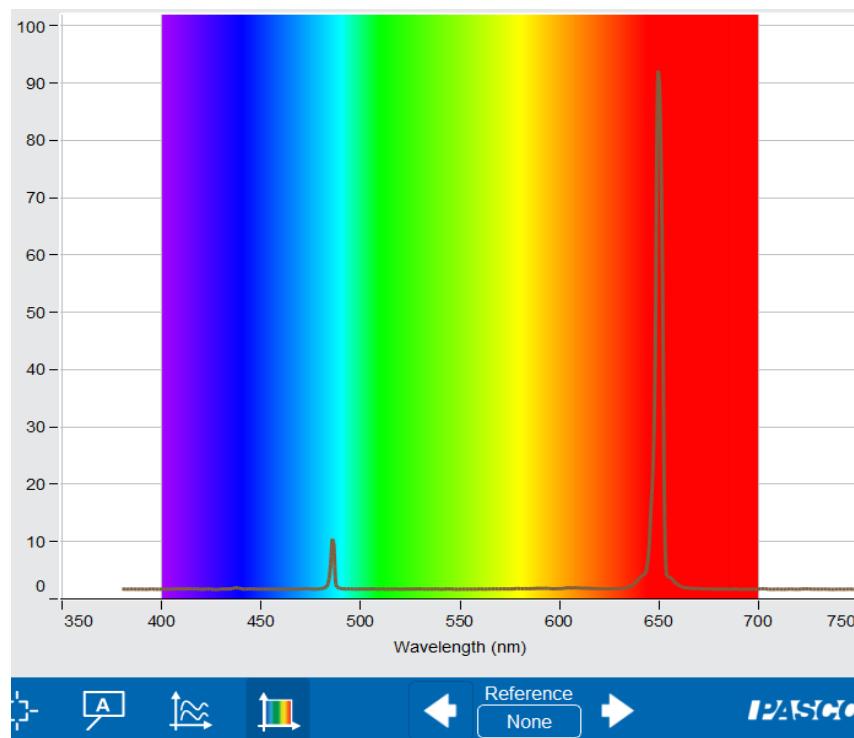


FIG. 4: Spectrum of atomic hydrogen.

### Part III. Calculating energy of quantum states and transitions of helium.

Knowing the quantum states of the He atom was very important for designing the He-Ne laser. Replace the hydrogen tube by the helium tube.

Perform all steps in setting up and taking spectrum of helium as for the previous two gases. Turn off the power supply by the switch. Unplug the wire of the power supply. After the gas discharge tube cools down, put it back to its box.

The transition responsible for emission of the strongest peak with  $\lambda \approx 588$  nm takes place between the quantum states with  $m = 2$  and  $n = 3$ . Using Eq. (4) for hydrogen-like ions, calculate energies of these states for the ion  $\text{He}^+$ . Calculate the expected energy for the ion transition between the two states. Convert the energy into the wavelength.

➤ **Is your measured wavelength equal to the calculated one? Explain why it is or it is not.**

Each quantum state is characterized with a so-called electron configuration that represents the principal quantum number  $n$  and the orbital quantum number  $l$  of each of two electrons in the He atom. The quantum number  $l$  relates to the orbital angular momentum of the electron  $\vec{L}$  as

$$|\vec{L}| = \hbar\sqrt{l(l+1)}, \text{ where } \hbar = h/2\pi \text{ is a reduced Planck constant (h-bar).}$$

The quantum number  $l$  can take specific values depending on the value of  $n$  as follows:  $l = n - 1, n - 2, \dots, 0$ . The quantum state with  $l = 0$  is denoted as “s-state”, with  $l = 1$  as “p-state”, with  $l = 2$  as d-state, and so on. The ground state of the He atom is  $1s^2$ , which means that  $n = 1, l = 0$ , and there are two (superscript) electrons in this state. The first excited state is  $1s2s$ , which means that one of the electrons has  $n = 1, l = 0$ , and the other one has  $n = 2, l = 0$ . The state  $1s2p$  means that one of the electrons has  $n = 1, l = 0$ , and the other one has  $n = 2, l = 1$ .

Besides quantum numbers  $n$  and  $l$ , the quantum state of the electron can vary due to the different direction of spin. A combination of the spin quantum number and the orbital quantum number may be different for same magnitude of spin and angular momentum, because both of them are vectors. Their sum is the quantum number  $J$ , the total angular momentum, and this quantity also obeys some selection rules, i.e. the rules that permit or not permit specific quantum transitions.

Table 2 contains data of the electron configuration of different quantum states of He atom and transitions permitted by the selection rules of quantum mechanics. The quantum number  $J$  is in brackets.

Table 2. Electron configuration and permitted transitions in the He atom. The reference wavelengths and intensity are taken from  
[https://physics.nist.gov/PhysRefData/ASD/lines\\_form.html](https://physics.nist.gov/PhysRefData/ASD/lines_form.html)

Electron configuration of the initial state (upper level)	Electron configuration of the final state (lower level)	Reference intensity in arbitrary units	Reference wavelength of the emission line, $\lambda_r$ , nm	Measured wavelength of the emission line $\lambda_{exp}$ +/- uncertainty, nm	Energy of the quantum state $E_n$ , eV	Energy of the quantum state $E_m$ , eV	Selection rules		
							$\Delta n$	$\Delta l$	$\Delta J$
1s2p (1)	1s <sup>2</sup> (0)	1000	58.43339			-24.57			
1s3s (1)	1s2p (2)	200	706.5190			-3.60			
1s3p (1)	1s2s (1)	500	388.8648			-4.75			
1s3d (3)	1s2p (2)	500	587.5621			-3.60			
1s3d (2)	1s2p (1)	100	667.8151			-3.35			
1s3p (1)	1s2s (0)	100	501.56783			-3.95			
1s4d (1)	1s2p (2)	200	447.14802			-3.60			

1. Copy the Table 2 to your report.
  2. Use the data for  $\lambda_r$ , convert them to the energy of the spectral line in electronvolts, calculate the energy of the initial state and enter it to the column for the energy,  $E_n$ , in eV.
  3. Calculate the change of quantum numbers  $\Delta n$ ,  $\Delta l$ , and  $\Delta J$ , in the observed transitions.
  4. Enter the wavelengths of the emission lines you obtained experimentally with corresponding uncertainty and conclude on the difference between the  $\lambda_r$  and  $\lambda_{exp}$ .
- How many spectral lines correspond to the transition of the atom of He from  $n = 3$  to  $m = 2$  in your experiment? Why could not you observe same result for a hydrogen atom?
  - Explain the main difference between the hydrogen-like ion and the He atom that makes Eq. (2a) not applicable to calculating energies of spectral lines of the neutral He atom.
  - Write the selection rules in terms of all possible  $\Delta n$ ,  $\Delta l$  and  $\Delta J$  for transitions corresponding to the spectral lines registered in the experiment.

#### Part IV. Determine the unknown gas.

Replace the He tube by the tube with unknown gas. Using your calibration, some standard spectra stored in the spectrometer and known wavelengths of different gases from APPENDIX 2, determine the unknown gas. In your report, give the list of experimental wavelengths for all spectral lines you have observed and their relative intensity compared to the strongest line in the spectrum. Organize all data in a table.

#### Part V. Transmittance and absorbance measurements.

This part of the experiment does not require a discharge tube. You will now work with the built-in source of light and a sample of liquid with a dissolved dye. Due to interaction between an atom of

the medium and a photon, emitted by the source, the photon can be absorbed or scattered. In both cases, the photon will not reach a detector behind the sample with the solution. Absorption, scattering and transmission are described in terms of probability of the process. The probability of transmission is calculated as the ratio of the number of the photons not absorbed in the sample to the number of the photons incident on the sample surface. In terms of intensity, the probability for a beam of light of safely passing through the sample is equal to the intensity  $I$  of the exit beam of light divided by the incident intensity  $I_0$ . This probability depends on the energy (or wavelength) of the photon and therefore may be presented as a function of the wavelength, like any spectrum.

The quantity  $T = I / I_0$  is called *transmittance*.

The quantity  $A = -\log_{10}T$  is absorbance. Then,  $T = 10^{-A}$ . (5a)

If transmittance is expressed in percent, absorbance is calculated as  $A = -\log_{10}(T/100)$ . (5b)

The next measurements of  $T$  and  $A$  demonstrate how the visible color of a dye corresponds to transmittance and absorbance of photons with different wavelength inside the beam of *white light*.

1. Remove a fiber optics probe from the spectrometer. In a software window, press ANALYZE SOLUTION on the upper panel of the screen.
2. Press CALIBRATE DARK at the bottom of the screen. Wait for a check mark.
3. Insert a cuvette containing distilled water (a blank solvent) into the CUVETTE HOLDER with the gloss side of the cuvette facing the built-in source of the *white light*.
4. Allow the light source inside the spectrometer to warm up for about 1 minute.
5. Press CALIBRATE REFERENCE.
6. Replace the cuvette with the distilled water by a cuvette with a solution of a blue dye.
7. Press the option ABSORBANCE / TRANSMITTANCE to analyze these properties of the solution with white light. Start the data acquisition. In about 2 – 3 minutes stop recording.
8. Switch from ABSORBANCE to TRANSMITTANCE. By choosing the option  “Show Dual Y-Axes”, superimpose the two diagrams on one display and find the values of  $T$  and  $A$  at the points of interest, e.g. minima and maxima of the functions, or the points of drastic growth.
9. Calculate  $A$  using the experiment values of  $T$ . Compare your experiment results to the expected values calculated with Eq. (5). Enter all results into a table.
10. Repeat measurements of  $T$  and  $A$  for the red and green dyes. Turn the spectrometer off.
11. Basing on the results of pp.9 - 11, explain the colors of the solutions in white light.

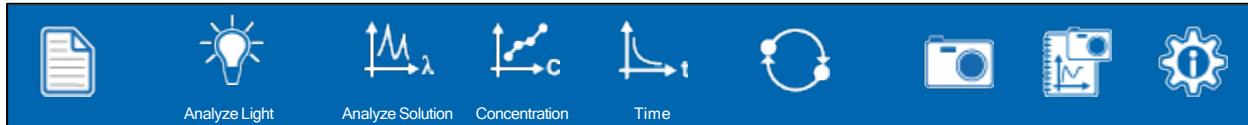
#### Part VI. Fluorescence.

1. Insert a cuvette with the yellow dye with its gloss sides to the sources of violet and blue light. Take spectra at 405-nm excitation and 500-nm excitation. Find the absorption edge, which is the wavelength at which absorption jumps significantly.
2. Use the wavelength of the absorption edge to calculate the band gap energy of the dye in electronvolt as  $E(\text{eV}) = 1240 / \lambda(\text{nm})$ .
3. The dyes with semiconductor properties are used in photovoltaics. Basing on the value of the band gap obtained for the yellow dye, conclude on whether it possesses required properties.

## APPENDIX 1

### I. Application Tools for Analyzing a Spectrum of a Gas Discharge Tube and a solution.

#### 1.1 Upper panel icons with description



Opens the file structure of the software to Open, Create and Save the \*.sp files.



Analyze Light in terms of Intensity vs Wavelength:

**Yellow** - “Currently in use”;

White – Not in use, but available for use;

Grey – disabled (e.g. the spectrometer is not turned on).



Analyze a gas or a solution spectrum in terms of Absorbance, Transmittance and/or Fluorescence vs wavelength.



Analyze a Solution in terms of Absorbance, Transmittance and/or Fluorescence vs solution concentration. Not in use in PHY293Lab.



Analyze a Solution in terms of Absorbance, Transmittance, Fluorescence and/or concentration vs reaction time. Not in use in PHY293Lab.



Connection Status:

Error

OR



Connected



Take Journal Snapshot



Show Journal Snapshots



Software information and settings



Export Snapshots To HTML:

Sharing Options (Tablet only)

Open in another app: save the \*.sp file into Google Drive, DropBox, etc. that are installed on the device.

## 1.2 Lower panel icons with description



Start recording data.



Stop recording data.



Scale to fit the graph (for very strong or very weak peaks).



Add coordinates tool to select points on the graph.



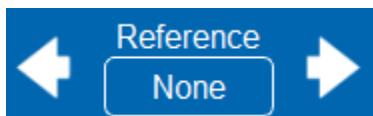
Add text annotation.



Compare spectra of multiple light sources.



Show (or remove) color spectrum background.



Add reference lines. The lines include Hydrogen, Helium, Argon, Xenon, Mercury, Sodium and Neon from NIST Atomic Spectra Database and are given for orientation only.

## 1.3 Left-hand side panel icons with description



With a higher integration time, the spectrometer is more sensitive to less intense light. The “Auto Set” button automatically adjusts the integration time to maximize the spectrum.



The higher number of scans results in a better signal-to-noise ratio. Should be adjusted manually for each spectrum.



Average groups of adjacent data points. May change the shape of the spectrum drastically. Not always recommended to exceed 1.

## II. Application Tools for Analyzing an Unknown Solution.

### 2.1 Left-hand side panel icons with description

#### Absorbance/Transmittance

Analyze absorbance and transmittance of a white light source through the sample. Select “Absorbance” on the graph to switch from Absorbance to Transmittance.

#### Fluorescence (405 nm)

Analyze fluorescence of the sample with 405-nm excitation.

#### Fluorescence (500 nm)

Analyze fluorescence of the sample with 500-nm excitation.

### 2.2 Lower panel additional icons for analyzing a solution (with description)



Calibrate Dark.



Calibrate Reference (e.g. distilled water in a cuvette if it is a solvent of the unknown solution).



Show dual Y-axes of absorbance and transmittance data only.

## APPENDIX 2

Wavelengths of the strongest lines in emission spectra of gases in nm.

Gas	1	2	3	4	5	6	7	8	9
Neon	540	585	622	640	660	693	703	717	725
Argon	427	435	440	459	473	476	502	696	
Krypton	427	432	557	587	646	759	769	810	826
Xenon	484	529	534	542	597	605	610	681	699
Sodium	588.995	589.592							
Nitrogen	404	444	463	501	568	575	594	648	661
Oxygen	412	423	533	559	605	616	626	637	645