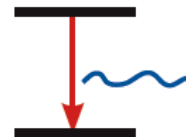


Atoms First

Recent interest in an “Atoms First” general chemistry organization has created a vacant spot early in the general chemistry lab curriculum. Instructors are discussing light, atomic spectra, electron structure, and bonding. Students aren’t yet able to deal with the traditional lab topics of reactions and wet chemistry.



Likewise, in a traditional course organization, students have little corresponding visual “hands-on” experience with these lecture topics.

Consider using MicroLab’s inexpensive, visual, and interactive **Energy of Light** and photographic **Atomic Spectra** experiments to help fill this gap.

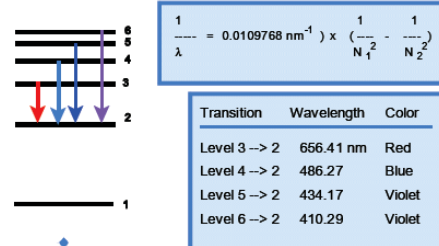
The relationship between color and energy of light is basic to understanding of atomic emission and molecular absorption spectra, and to our understanding of the behavior of electrons.



The Balmer (visible) series of atomic hydrogen (pictured here with a Hg reference spectrum) provided early support for the Bohr electron model. Sharp spectral lines were thought to be due to sharply defined electron energy levels. Broadening of this red emission line is photographic flare.

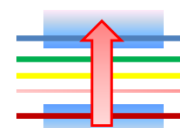
The Energy of Light – Planck’s Constant

Students use **MicroLab’s Model 214 Energy of Light Module** to visualize, quantify, and understand the relationship between energy, color, wavelength, and frequency of light. They do this by easily measuring the voltage required to initiate emission of light in a series of visible and IR Light-Emitting Diodes.



This relationship between energy and color/wavelength/frequency of light can be demonstrated at four different conceptual levels, depending on the level of the course.

1. The energy of light *increases* as color changes from infra-red to violet. UV has enough energy to break chemical bonds and to cause sunburn.
2. The energy of light is *inversely proportional* to the wavelength of the light wave.
3. The energy of light is *directly proportional* to the frequency of the light wave.
4. The mathematical proportionality between energy (E) and frequency (ν) of light is known as Planck’s Constant (h): $E = h\nu$.



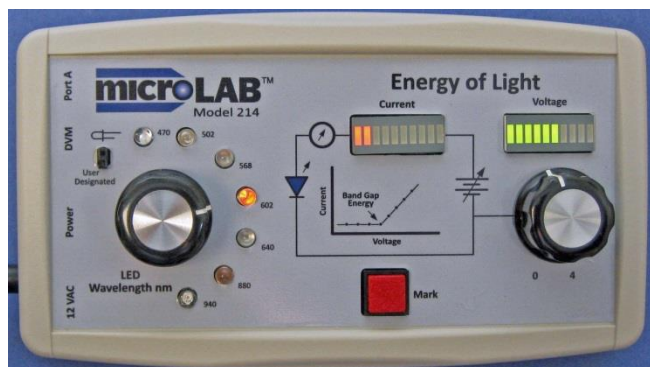
Planck’s constant and the measured wavelengths of the atomic emission spectra of excited hydrogen atoms were historically used to support the Bohr electron model. It and spectral measurements are commonly used today to calculate the electron energy level changes within an atom.

Students can use **MicroLab’s Model 214** to determine Planck’s constant by measuring the band-gap voltage of a series of inexpensive LED’s.

LED’s emit light when sufficient voltage is applied to force electrons to jump the energy gap between the N (electron excess) and P (electron-deficient) semi-conductor layers in the LED.



Atoms in molecules can vibrate and rotate, broadening electron energy levels and producing broad absorption and emission bands.



When the arriving “N” electrons recombine with holes on the “P” side, the energy imparted to the electron is given off as light. The color of light is determined by the energy difference across the energy band-gap in the LED semiconductor. Band-gap is determined by the chemistry of the semiconductor, which is selected to create LED’s of different colors.

Students can determine the band gap energy of a selected LED, and therefore the relationship between energy and the color of light it produces, by simply rotating the voltage control until light just starts to be produced by the LED. They press the “Mark” button when they see light production begin, recording the band gap energy in software.

The applied voltage and the current through the LED are reported both digitally and graphically on the MicroLab display, as well as live with colored bar graphs on the Energy of Light module (see photo below). The left (red) LED bar graph shows current through the experimental LED, and the right (green) LED bar graph reports applied voltage.

Students can see the initiation of current with the LED current bar graph at the same time they see visible light beginning to be produced by the LED. This display also enables them to observe electron movement with the two infra-red LED’s, which produce light invisible to the eye.

The graphs presented on this page show student data illustrating the four key concepts listed on the previous page.

The experiment takes only about 20 minutes –more than one lab group can share an Energy of Light Module.

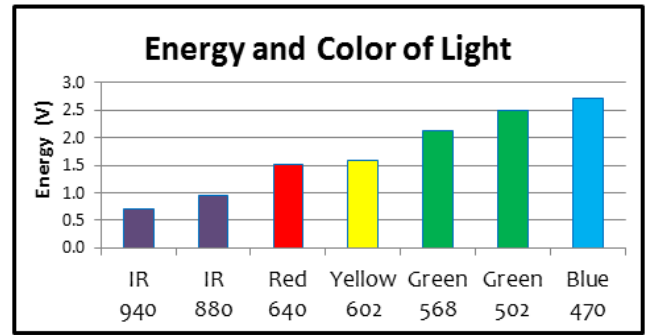
Seven LED’s are provided, with wavelengths from 940 nm to 470 nm. An eighth switch position and a socket permit use of a user-designated LED.

Stand-alone Experiments and Demonstrations

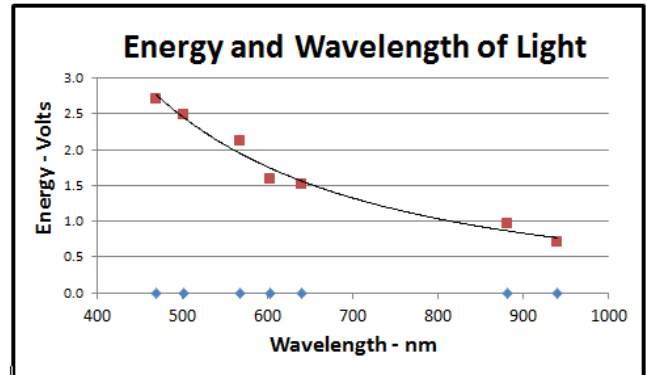
MicroLab’s Model 214 Energy of Light apparatus can be used “stand-alone” without a MicroLab FS-522. In this operating mode, an inexpensive digital voltmeter (DVM) is used as the readout device. Power is supplied by a small plug-in wall power pack to eliminate the need for batteries.

The 214 and DVM can be used together as a demonstration with a video camera and a projector. Students can watch the instructor rotate the voltage control. They can observe the increase in voltage on the voltage LED bar graph, the increase in current on the current LED bar graph, and watch the LED come on.

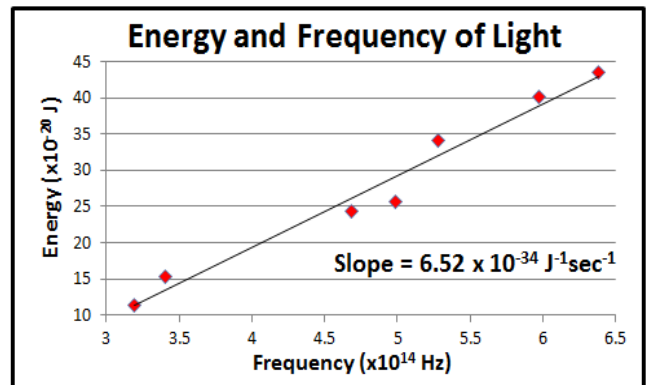
Students can also read the applied voltage on the DVM display, and participate in graphical analysis of data on a whiteboard or computer display using the provided Excel analysis template. Graphs presented on this page were produced with the Excel analysis template.



Concept 1: The energy of light increases as color changes from infra-red to violet.



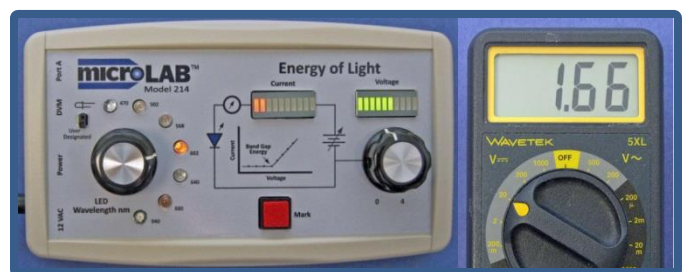
Concept 2: The energy of light is inversely proportional to the wavelength (λ) of light.



Concept 3: The energy of light is directly proportional to the frequency of the light wave.

Concept 4: The mathematical proportionality between energy (E) and frequency (ν) of light is known as Planck’s Constant (h): $E = h\nu$.

Planck’s Constant is the slope of the energy-frequency graph. Its accepted value is $6.63 \times 10^{-34} \text{ J}^1\text{sec}^{-1}$, within about 1.6% of this measurement.



Measuring Atomic Spectra

MicroLab's Model 141 Visual Spectrometer measures both atomic emission spectra and molecular absorption spectra. A large 500 line/mm diffraction grating and an adjustable 3-position slit work with the focusing lens to produce sharp slit images and bright spectral lines.



The Visual Spectrometer has two light inputs, both entering from the large end of the spectrometer block. The bottom input comes via a flexible fiber-optic cable and presents a Hg reference spectrum. The upper light source is from an atomic emission discharge tube or a white light source for solution absorption spectra (the sample vial sits in the block). Light from the two light sources pass through the same slit as they travel to the diffraction grating. Spectra can be viewed simultaneously in wavelength alignment.

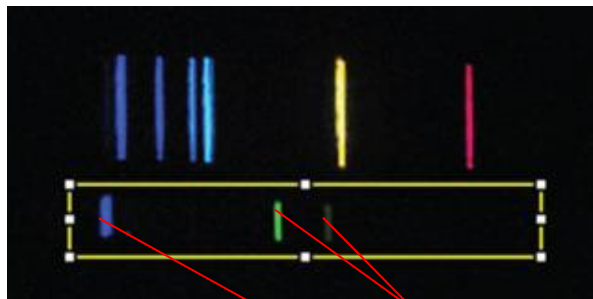
The Spectrometer's sensor is an inexpensive digital point-and-shoot or cell phone camera. Point-and-shoot or SLR cameras work better because their sensors are larger, but cell phones give acceptable results. Even a 3 MB camera has a sensor that is about 2000 pixels across. This is about 20 times the resolution of a \$500 educational diode array spectrophotometer.

Because the camera's CCD sensor is designed to take high quality colored pictures, it is sensitive across the whole visible spectrum. Inexpensive diode array spectrophotometers have poor sensitivity in the blue. Pre-recorded spectra are available on our web site.

Data analysis is provided by powerful image analysis software. *Image J* is a photographic pixel density analysis program developed for life-sciences researchers by the National Institutes of Health. It is available without charge on their web site (<http://rsbweb.nih.gov/ij/>). *Image J* is extremely easy to use – in this application the student just draws a box around the spectrum of interest, clicks the mouse, and the intensity profile of the spectrum is calculated and displayed.

Measuring Atomic Spectra

Visualize and Measure



This is a point-and-shoot camera photo of the spectrum of an unknown gas with a fiber-optic mercury spectrum in wavelength alignment below it. The student draws the yellow analysis box with a mouse, wide enough to cover both the reference and unknown spectra.

Visualize & Measure

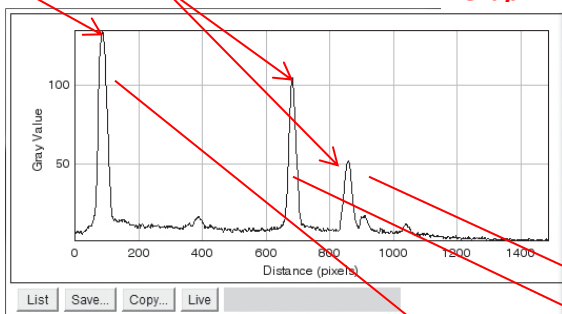
Graph

Create a Math Model

Predict

Test

Graph



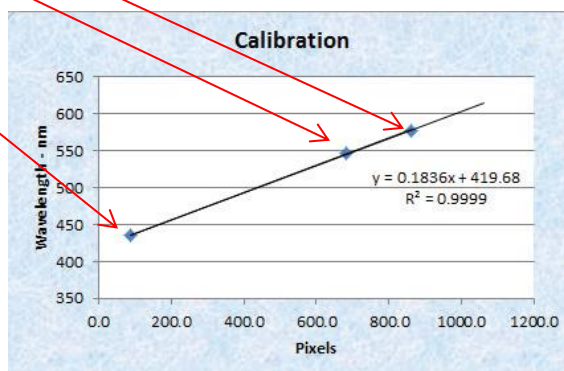
Clicking "Analysis" and "Plot Profile" produces an *Image J* plot for this mercury reference spectrum. The plot shows three distinct peaks at known spectral lines of 436 nm (violet), 546 nm (green), and 578 nm (yellow). The list view to the left shows the 546 nm line peaking at 683 pixels.

X	Y
675	72.3716
676	77.4809
677	82.0656
678	89.5137
679	92.4918
680	97.4809
681	102.0765
682	104.6831
683	105.1093
684	104.9290
685	104.8907
686	103.0984
687	100.4754
688	96.6011
689	92.6503
690	88.9891
691	83.3333
692	77.9016

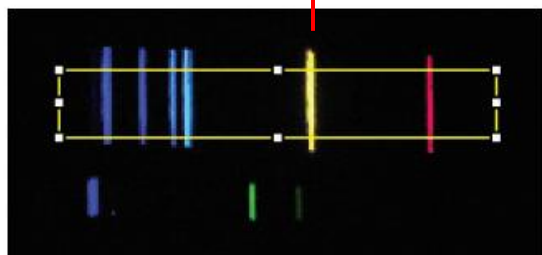
A	B
Wavelength - nm	Pixel Position
436	89
546	683
578	862

This Excel calibration chart correlates pixel position with wavelength. The regression equation shows that each pixel represents 0.1836 nm, with 419.68 nm as the y-intercept. Note the quality of data – four nine's in the correlation coefficient.

Math Model

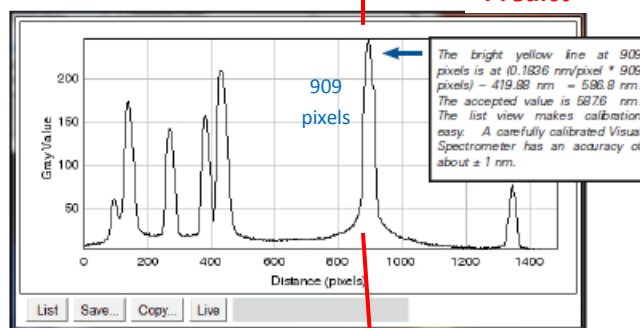


Visualize, Measure



The Image J analysis window was moved up into the unknown spectrum in this photo. The result of this analysis is shown to the right. Six spectral lines are clearly visible. Note high sensitivity in the blue.

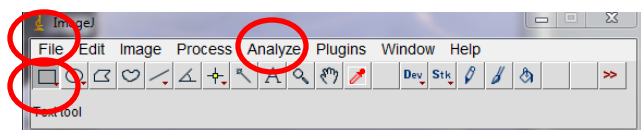
Predict



The calibration equation is used to convert the pixel position of each line into wavelength. The yellow line is at 909 pixels. Note that this spectrum has 1400 channels.

Using Image J

1. Load Image J on your computer. You will see this toolbar. Key controls are circled.



2. Click on **File**. Choose **Open**. Select the your spectrum photo from the camera memory card inserted in your computer, or import it via USB cable.
3. Click on the rectangle below "**File**" in the Image J Tool bar. Using your mouse, draw a rectangle around the reference spectra, taking care that the rectangle is long enough to subsequently hold all of the unknown spectral lines.
4. Click "**Analyze**" on the Image J tool bar, and chose "**Plot Profile**". Image J will draw a pixel intensity profile across your spectrum.
5. Determine the pixel position of each of the reference spectrum Lines. You can identify the peak of each spectral line by placing your mouse on top of the peak, or by clicking the "**List**" button at the lower left of the spectrum profile display. You can scroll up or down to locate the exact position of the center of each spectral line.
6. Plot these reference spectra points in Excel, pixel position (x) against wavelength (y). Run a regression line through your points to determine the equation for the calibration line.
7. Carefully click in the yellow box surrounding your reference spectrum, and use the **UP arrow key** to move the box straight up so that it encloses the unknown spectra. It is not essential to enclose the entire height of the unknown spectrum, but it is essential to move the box straight up so that it remains in wavelength alignment with the reference spectrum. You cannot move the box straight up with the mouse.
8. Measure the pixel positions of each of the unknown spectral lines, and use the calibration equation to calculate their wavelengths.

$$\text{Wavelength} = 0.1836 \text{ nm/pixel} * \text{pixel position} + 419.68 \text{ nm}$$

$$\text{Predict: Wavelength} = (0.1836 \text{ nm/pixel} * 909 \text{ pixels}) + 419.68 \text{ nm}$$

$$\text{Wavelength} = 586.6 \text{ nm}$$

Test

True Helium Wavelengths

Violet	443.7 nm
Violet	447.1 nm
Blue	471.3 nm
Blue	492.2 nm
Cyan	501.6 nm
Yellow	587.6 nm
Red	667.78 nm

$$\text{Error} = (587.6 \text{ nm} - 586.6 \text{ nm}) / 587.6 \text{ nm} = 0.170\%$$

