EXPERIMENT 17: Atomic Emission

PURPOSE:

- To construct an energy level diagram of the hydrogen atom
- To identify an element from its line spectrum.

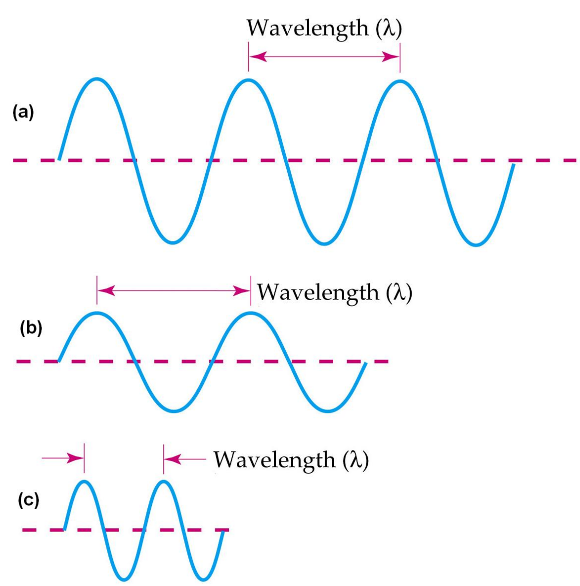
PRINCIPLES:

White light, such as emitted by the sun or an incandescent bulb, is a form of energy. When sunlight passes through raindrops it spreads out in a spectrum of colors called **continuous spectrum**. That white light is indeed composed of many colors can also be shown by placing a triangular piece of glass (called a prism) or a diffraction grating in the beam of white light. As the light passes through the prism, the prism changes the direction of some colors more than others. The result is the spreading out of the colors and it is referred to as **dispersion**.

Although we consider light to be colored, color is only a perception in our minds caused by our brains interpreting the sensation of light energy striking the retina of our eyes. There must be some variable property of visible light that tells our eyes and brains what color to see. This property must be different for the various colors of the visible light spectrum.

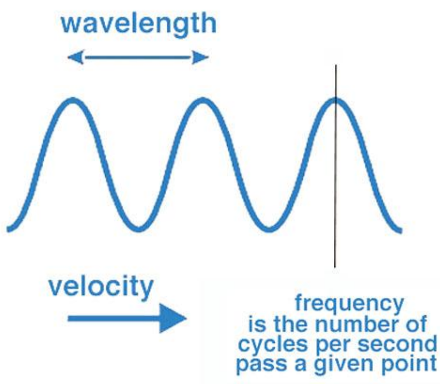
One theory used to describe light considers light as consisting of a series of energy waves that can be transmitted through space. One important characteristic quantity associated with waves is the wavelength, which is the distance from any point on one wave to the corresponding point on the next wave (for example from crest to crest).

The symbol for wavelength is the Greek letter, λ (pronounced "lambda"). The figure below shows diagrams of three waves, each with a different wavelength. Wave A has the longest wave length and wave C has the shortest.



Wave length values are expressed in length units, such as meters (m), centimeters (cm), and nanometers (nm) or Angstroms (A).

Note that 1m = 10^2cm = 10^9nm = 10^{10}Å



Another Important characteristic of waves is the frequency. The frequency of a wave is the number of waves passing through a point in a time interval. The symbol for frequency is the Greek letter ν (pronounced mu).

NOTE:

λ = wavelength = distance from crest to crest
(measured in units of length such as: m, nm, A)

ν = frequency = number of waves passing through a point in a time interval.
(measured in 1/second or Hertz, Hz)

The wavelength multiplied by the frequency is equal to the velocity of the waves.

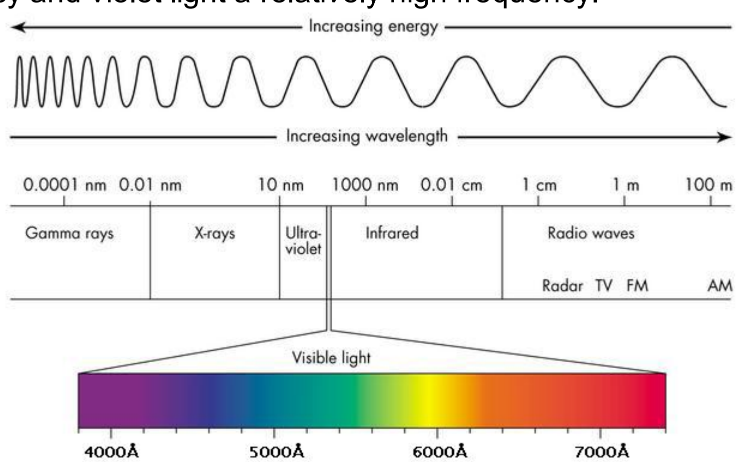
WAVELENGTH X FREQUENCY = VELOCITY

This relationship between frequency and wavelength holds true for all types of wave motion. In the case of light traveling through a vacuum, velocity has the specific value of 3.00×10^{10} cm/sec. This value is an important physical constant and is given the symbol, c. Thus, for all light in a vacuum:

$$c = \lambda \times \nu$$

The various colors of the visible spectrum correspond to various wavelengths of light. Violet light has a relatively short wavelength and red light has a relatively long wave length. Because "c" is a constant for all colors, it is apparent that red light must have a relatively low frequency and violet light a relatively high frequency.

In the spectrum of white light there is a continuous gradation of color; violets gradually merge into blues, blues blend into greens, etc. Because the visible spectrum is continuous, it consists of all wavelengths from approximately 400 nm (4000 Å) to 700 nm (7000 Å).



You should realize now that you cannot describe light precisely by just indicating its color. For example, if you wished to specify an orange light, the information in the figure above indicates that the light could have a wavelength around 620 nm (590 nm would be a yellowish orange, while 640 nm would be a reddish orange).

Word descriptions of colors are not exact and vary from person to person. Although catchy phrases such as *scarlet red* and *aqua blue* may serve the world of advertising, scientists need a much more precise system to specify color. This is done by specifying the wavelength (or frequency) range of light that is associated with a particular color.

It is sometimes useful to consider light as consisting of particles rather than waves. This model stresses the concept of light energy and how this energy is converted into other forms. It originated with the physicist, Max Planck, who suggested in 1900 that substances absorb or emit light in discrete amounts called quanta. Albert Einstein extended Planck's theory by proposing that light is composed of particles of energy called photons. Each photon has a definite quantity of energy. When light interacts with a substance, the photon transfers its energy to the substance as a packet. For anyone frequency of light, all the photons have the same energy. The two different models of light, particle and wave, are related quantitatively by the Planck Equation:

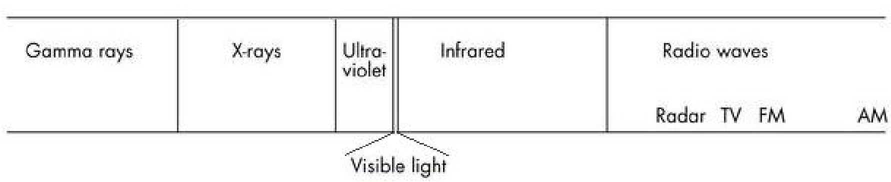
$$E = h \times \nu$$

NOTE:

E = the energy of the photon
h = a proportionality constant called Planck's constant.
($h = 6.63 \times 10^{-34}$ Joule x second)

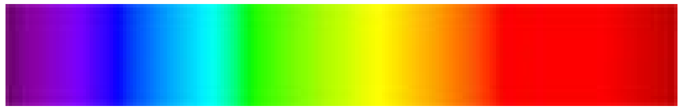
All the photons of light of a particular frequency have the same energy. Moreover, as the frequency of light increases, the energy also increases. Therefore, for example photons of blue light have higher energies than photons of red light.

Actually, what we call "light" is only a part of a very broad spectrum of waves called the electromagnetic spectrum. Our eyes see only the part called visible spectrum. Radio and television waves, Infrared rays (IR), Ultra-violet rays (UV), X-rays, γ-rays and cosmic rays are part of the same electromagnetic spectrum, but they cannot be detected by our eyes.



Only a small part of the electromagnetic spectrum is the visible region as shown above. The general regions of the spectrum, such as infrared and ultraviolet, are also indicated. However, there are no sharp boundaries between the parts of the electromagnetic spectrum, just as there are no sharp boundaries within the visible spectrum.

Most substances will emit light energy if heated to a high enough temperature. If the light from an ordinary electric bulb (in which a thin tungsten wire is heated white-hot by an electric current) is passed through a prism, or a diffraction grating a **continuous spectrum**, containing all the wavelengths of visible light is obtained.



Modern street lights produce a different type of light. They consist of high intensity mercury or sodium vapor lamps and are referred to as gas discharge tubes since the element inside the lamp is in its gaseous state and is made to emit light by passing an electric discharge through it. When the spectrum of this light is observed through a prism, only a few bright lines, corresponding to specific wavelengths will be seen.

Each chemical element gives rise to a characteristic **bright line spectrum** (also called atomic emission spectrum) and in a way can be considered as the fingerprint of the element, since no two elements have identical emission spectra. Consequently, the appearance of a line spectrum and the interpretation of the specific spectral lines and their specific wavelengths can be used to identify the element.

The origin of the spectral lines baffled scientists for many years. The explanation of line spectra is based on an intimate knowledge of the atomic structure of the element involved and the energy changes that occur within the atom when energy is absorbed (heat or electrical energy) or released (light energy).

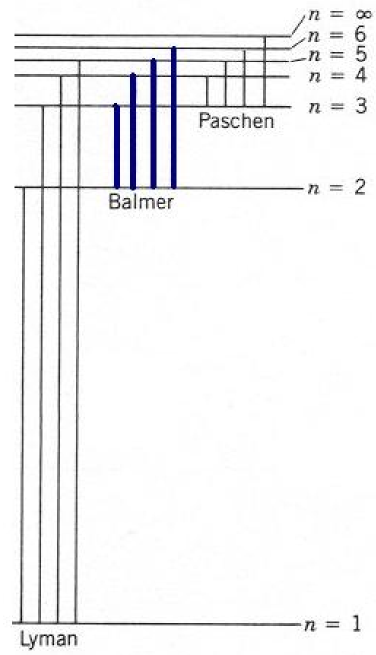
The observation that atoms of a given element which have absorbed energy emit light energy at only fixed wavelengths indicates that atoms can absorb or release energy only in fixed, definite amounts. This in turn implies that an electron in an atom can possess only certain, specific, definite amounts of energy.

This can be summed up by saying that the electron is restricted to specific energy levels in the atoms, usually designated by an integer "n" whose values can be n = 1, 2, 3, 4, 5… ∞. When n = 1, the electron is in a position closest to the nucleus, and it is in the lowest energy level.

Electrical or heat energy is absorbed when an electron jumps from a lower to a higher energy level (lower to higher value of "n"). The energy thus absorbed is equal to $\Delta E = E_{final} - E_{initial}$

Energy is emitted in the form light energy (brightly colored lines) when the electron "falls" from a higher to a lower energy level (higher to lower value of "n”). The energy thus emitted is also equal to $\Delta E = E_{final} - E_{initial}$. Summing up the two situations, gives:

ΔE = ENERGY CHANGE (absorbed or emitted) = E_{final} - $E_{initial}$



Niels Bohr, a Danish physicist, studied and explained the origin of the bright lines in the emission spectrum of hydrogen. Hydrogen is the simplest atom (it has only one electron) and hence, it has the simplest atomic spectrum. Four characteristic bright lines can be seen in the visible region of the atomic spectrum of Hydrogen. These lies are called the Balmer lines. The lines result from tile following electron transitions: 3→2, 4→2, 5→2, 6→2 (the last one is very difficult to see since it corresponds to a wavelength which is close to the ultra-violet region of the spectrum).

Similar series of lines, which occur in the UV (Lyman series) and IR (Paschen series), have been later identified in the atomic spectrum of hydrogen.

The wavelength of all spectral lines can be calculated by using the equation below:

$$\frac{1}{\lambda} = 109{,}678\, cm^{-1} \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

NOTE:

n_1 = final energy level
n_2 = initial energy level
$n_2 > n_1$ for emission spectra

Since other elements have a more complex electronic structure, their line spectra are also more complex. However, the pattern of colored lines is characteristic for every element and can be used to identify it.

PROCEDURE:

- First, a simplified spectroscope (an instrument that measures light emission) will be calibrated for wavelength measurements by viewing the emission spectrum of helium.
- In the second part of the experiment, the energy level diagram of the hydrogen atom will be determined.
- In the third part of the experiment, the line spectrum of an unknown element will be used to identify the element from a list of possible choices.

PART I: CALIBRATION OF THE SPECTROSCOPE

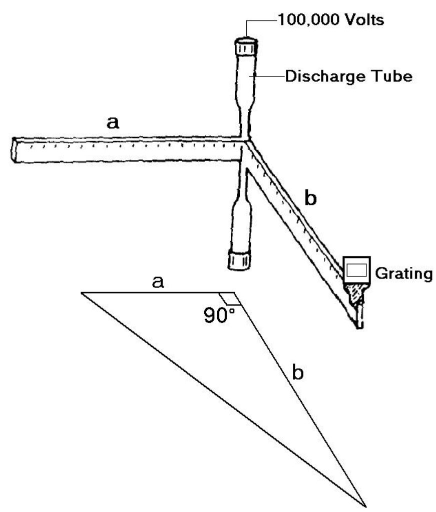
The spectroscope consists of:

- Two meter sticks crossing each other at a right angle
- Diffraction grating

- Gas discharge tube connected to a high voltage power supply.

The diffraction grating is a flat piece of plastic with a series of closely spaced lines etched into its surface and it is used instead of a prism to produce the dispersion of light.

The diffraction grating is placed on meter stick “b" in a fixed position exactly 1.00 m away from the point of intersection of the two meter sticks. Meter stick "a" will be used as the scale of the spectroscope and the units marked on meter stick "a” will be used as the arbitrary divisions of the spectroscope.



Calibration is accomplished by viewing the emission spectrum of helium because the emission wave lengths for helium are precisely known.

You will work in pairs and record the positions of the helium spectral lines on the spectroscope scale (meter stick "a"). You will use these data to prepare a "calibration curve" by plotting the measured positions against the known wavelengths of the lines, which are:

Violet ....447 nm
Blue ....471 nm
Blue-green ....492 nm (not always visible)
Light green...........................502 nm
Yellow..................................588 nm
Red......................................668 nm
Red......................................707 nm Å (not always visible)

Ask your instructor to position the power supply and helium lamp directly behind the intersection point of the two meter sticks. Check to be certain that the meter sticks are at right angles.

CAUTION

1. The power supply develops several thousand volts. DO NOT TOUCH any portion of the power supply, wire leads, or lamps unless the power supply is unplugged from the wall outlet.
2. In addition to visible light, the lamps may emit ULTRAVIOLET RADIATION. Ultraviolet radiation is damaging to your eyes. Use your safety glasses or sunglasses, since they will absorb some of the ultraviolet radiation. DO NOT LOOK DIRECTLY AT ANY OF THE LAMPS WHILE THEY ARE ILLUMINATED FOR ANY EXTENDED PERIOD OF TIME.

CAUTION

1. With you instructor's permission, turn on the power supply to illuminate the helium lamp.
2. One student will look through the diffraction grating (at eyelash distance) and direct their eye of vision to the left to locate the position of a series of colored lines.
3. The second student should move a marker along meter stick "a" until the position of the maker matches the position of the chosen line.
4. Record the distance from the point of Intersection of the meter stick to the marker.
5. Repeat these steps until you have measured and recorded the distances "a" on the data sheet for all visible lines observed for the helium spectrum.
6. Turn off the power supply to the helium lamp.
7. On the graph attached to your Report Form plot the measured distances versus the known wavelength for each line.
8. Connect the experimental points with a straight line.
    a. The ideal curve for this relationship is a smooth curve; however for the purposes of this experiment the small deviation from a straight line variation can be considered negligible.
9. Attach the graph to your Report Form.

When proceeding from one part of the experiment to the next:

1. Do not change the position of the meter sticks
2. Do not change the position of the diffraction grating on meter stick "b"

PART II: THE ENERGY LEVEL DIAGRAM OF HYDROGEN

1. Replace the helium lamp with a hydrogen lamp.
2. Repeat the procedure used in PART I to determine the exact positions of the three visible spectral lines of hydrogen.
    a. You should easily observe the red, blue-green, and violet lines. A fourth line (faint violet) may also be visible, but its position is difficult to determine.
3. Record in your laboratory notebook, the color and the location of the hydrogen lines.

4. Use the calibration curve, that you have previously constructed, to determine the wavelengths of the spectral lines of hydrogen and enter these data in Table II.

Calculations:

1. Use the following equation to calculate the wavelengths in meters for the four hydrogen lines that appear in the visible region of the spectrum.

$$\frac{1}{\lambda} = 109{,}678\,cm^{-1}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

   Remember that $n_1 = 2$ and that $n_2 = 3, 4, 5$, and 6.

2. Enter these theoretical (calculated) values in Table III.
3. Convert the experimentally determined wavelengths in meters, and enter these values in Table III.
4. Calculate the % error for each wavelength.
5. Using the same equation, calculate the theoretical values of the wavelengths corresponding to the following transitions: $\infty \rightarrow 1$, $2 \rightarrow 1$. Enter: these values in Table IV.
6. Complete Table IV with the experimental values of the wavelengths corresponding to the following transitions: $3 \rightarrow 2$, $4 \rightarrow 2$, $5 \rightarrow 2$. (Transfer these values from Table III.)
7. Using $c = \lambda \times \nu$, calculate the frequency corresponding to each wavelength.
8. Using $E = h \times \nu$, calculate the energy change in Joules for each frequency.
9. Convert the energy change values obtained in Joules (J) into Electron volts (eV). Note that 1 eV = 1.60×10^{-19} J)
10. Using the energy change values expressed in eV (ΔE), calculate the energy values for each energy level (E_1, E_2, E_3, E_4 and E_5)
11. Enter the calculated values in Table V.

   HINT: First determine E_1 by keeping in mind that the transition form $n = \infty$ to $n = 1$ corresponds to $\Delta E = E_\infty - E_1$ and that $E_\infty = 0$.

12. Construct an energy level diagram of the hydrogen atom on the page provided for this purpose.
13. Before you start constructing the diagram, place the energy values you have calculated above on the upper right-hand side corner of the page.

   NOTE: Construct the diagram by simply drawing horizontal lines to indicate the energy levels; there is no need to indicate the electron transitions with arrows.

PART III: THE IDENTIFICATION OF AN UNKNOWN BY SPECTRAL ANALYSIS

1. Repeat the procedure described in PART I, but use as light source a discharge tube containing an unknown element identified only with a number.
2. Record the colors and the positions of the spectral lines on meter stick "a" in your laboratory notebook (Table VI).
3. Determine the wavelengths of the spectral lines of the unknown element by reading them from the calibration curve and enter these values in your laboratory notebook (Table VI).
4. Plot the spectral lines of the unknown element by drawing colored straight vertical lines below the graphical representation of the emission spectra of several elements whose identity is known. (See graph)
5. Color all the spectral lines of the elements whose ATOMIC EMISSION SPECTRUM is given. (Refer to the information on the graph which relates the wavelength range to color).
6. By matching the position and the color of the spectral lines of your unknown with the position and the color of the spectral lines of the elements given as a possible choice, identify your unknown.

NOTE: Do not expect a perfect match in neither the position nor the number of the spectral lines. Use your judgment to interpret possible errors in wavelength values and missing spectral lines. However, a careful comparison between the color of the spectral lines observed for the unknown and the color of the spectral line of the reference spectra is usually very helpful.