Experiment: Spectroscopy and Element identification

Introduction to light

Light is a form of energy called electromagnetic radiation. A chart of the electromagnetic spectrum is shown below. (Your text may have a nicer diagram, but this simple chart may help you with the following discussion.)

Radiowave	Microwave	Infrared	Visible	Ultraviolet	X-ray	Gamma Ray
Long wavelength						Short wavelength
Low frequency						High frequency
Low energy						High energy

One model used to describe electromagnetic radiation is called the **wave model.** This model uses several variables to describe this radiation: <u>wavelength</u>, <u>frequency</u> and <u>energy.</u>

Wavelength, lambda (□) is the distance between the crests of a wave. Wavelength can be measured in any length unit. For example, radio waves are often measured in meters. Gamma rays are often measured in nanometers.

Some other units for wavelength include a millimicron. A milimicron is an older term used for a nanometer. Another unit you may run across is called an Angstrom (A). One Angstrom is 10^{-10} meters. Notice that all of the following represent the <u>same</u> wavelength:

400 millimicrons = 400 nanometers = 4000 Angstroms = 400 X 10^{-9} meters

The second variable used to describe waves is called frequency. How often the crest of a wave passes a particular point is described by its frequency. The unit of frequency is the hertz (Hz). Once per second is 1 hertz, twice per second is 2 hertz, etc. One thousand hertz is expressed as a kilohertz (kHz). One million hertz is a megahertz (MHz).

Notice that the radiation with the shorter wavelength has a higher frequency. This is an inverse relationship. (See the chart above to confirm this.)

The energy carried by light is related to both wavelength and frequency. X-rays and ultraviolet rays are high energy radiation and can be damaging to your cells. On the other hand, radio waves are lower in energy and you need not use "blocking" agents to protect yourself!

A second model is often used to describe the energy carried by light. This model describes light as having particle characteristics. These "particles" are bundles of energy called photons. Each photon carries a small bundle of energy. A single photon associated with ultraviolet light has greater energy than a single photon associated with radio waves. Using this model, we can think of light as a stream of photons.

Notice that the energy associated with light increases with increasing frequency. This is a direct relationship. (See the chart above to confirm this.)

Let's compare ultraviolet radiation with radiowaves:

Sunscreens protect your skin from UV radiation with wavelengths between 290 and 400 nanometers. This is a short wavelength compared to radio waves. The frequency of UV light compared to radio waves is extremely high. UV light have frequencies in excess of 700 trillion (7×10^{14}) hertz.
Looking at radiowaves, when you tune your radio you are receiving radio waves with frequencies between 530 and 1700 kHz. (This corresponds to the "numbers" on your dial.) A radio wave corresponding to 1700 kilohertz on your AM radio has a wavelength of approximately 200 meters, approximately twice the length of a football field.

Both radio waves and ultraviolet light are invisible to the human eye, although some insects, like bumblebees, can see ultraviolet light. Visible light has a wavelength longer than that of UV light but shorter than that of radio waves. The wavelength of visible light is between 400 and 700 nanometers. Each color has a corresponding wavelength. For example, red light has a wavelength in the 700 nm range, while violet light has a wavelength near 400 nanometers.

Examine the electromagnetic spectrum in your textbook. The visible spectrum can be remembered with a mnemonic ROY G BIV (Red, Orange, Yellow, Green, Blue, Indigo, Violet). An interesting bit of trivia--the British use the following mnemonic--Richard Of York Gave Battle In Vain.

Which color of visible light has the lowest energy?______________

The longest wavelength?_________________

The lowest frequency? ________________

Hint: To help you remember the visible electromagnetic spectrum, notice that UV light is close to the violet color in the visible spectrum. Infrared light is near the red color in the visible spectrum.

Label the following diagrams of waves (□1, □2, □3).

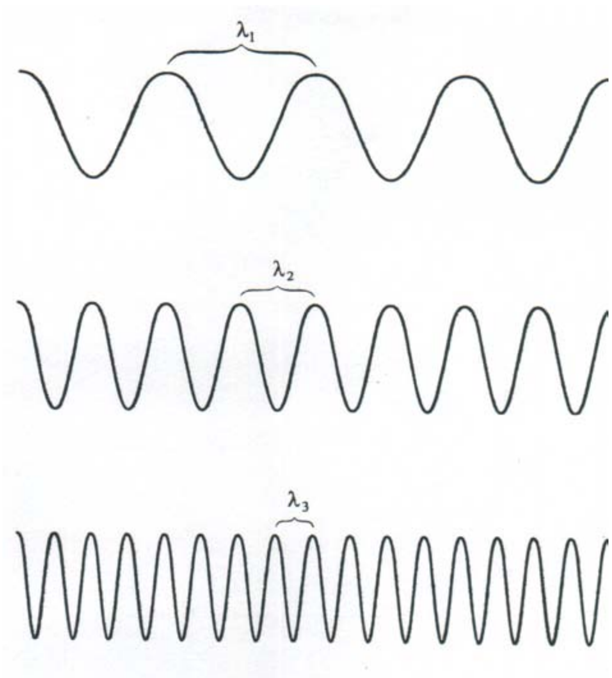
Which wave has the longest wavelength? The shortest wavelength?

Which wave has the greatest frequency? The smallest frequency?

Which wave would be a good model for UV light? Radio waves?

Visible light?

Which wave has photons with the greatest energy? The smallest energy?



Origin of Spectral Lines

The origin of the spectral lines can be explained using a quantum mechanical model for the atom. Atoms are usually in a low energy state. This is called the ground state for an atom. If energy is added to an atom, the electrons can absorb energy and move to a higher energy state, called an excited state. This energy can be added to atoms by an electric discharge or by heat. This added energy is emitted when the atom gives off a photon and the excited electron returns to the ground state.

The light emitted by atoms have definite wavelength and color that depends on the amount of energy originally absorbed. Each excited electron will emit one photon of light as the electron moves to a lower energy state. Billions of excitations and emissions are possible because a typical sample is made up of billions of atoms. There are also many possible energy jumps for each electron. Each of these jumps matches a different color of light. These colors constitute the emission spectrum of the excited atoms. A large jump corresponds to a color associated with high energy photons. A small jump corresponds to a color associated with a low energy photon.

Each element has a unique set of spectral lines. Since each element has a unique set of spectral lines, spectral studies can help us identify an element. The line spectrum of elements is much like a "fingerprint" for the element.

Procedure:

Part 1 Flame tests (All of the samples must be finely ground powders!)

1. Light a bunsen burner.

2. Obtain a vial containing a known sample. Shake the vial. Quickly open the vial and hold the mouth near the air vent of the bunsen burner.

3. Observe and record the color of the flame. Repeat this procedure for each known.

Part 2 Identification of an unknown solid

A. Obtain two unknown solids. Repeat the procedure described above and observe the color of the flame for each unknown.
B. Record the identity of the unknowns on your data sheet.

Part 3 Continuous and line spectra

Using a spectroscope, examine the spectrum given off by an incandescent light bulb and a fluorescent light bulb. Sketch the spectrum for each labeling your diagram with colors and wavelengths.

Part 4 Observing line spectra with the spectroscope

1. Your instructor has set up several gas discharge tubes in power supplies. Use your spectroscope to examine the spectrum for each of the tubes. **Do not touch the connections when the power supply is plugged into an outlet. A serious electric shock can result. Also, these bulbs get very hot. Allow them to cool before touching. Turn off the power supply when you are not using it.**
2. Sketch the results on your data sheet, noting the color and wavelength of the observed lines. Look at tubes containing He, Ne, Hg and H.
3. Use your spectroscope to re-examine the fluorescent light bulbs in the lab room. Which element is used to make fluorescent light bulbs. Justify your answer.

Experiment - Spectroscopy

Report Sheet Name ______________________________

Chem 121 Lab Partner _________________________

Part 1 Flame tests for known elements

Metal Ion	Color of flame
Lithium	___________________
Calcium	___________________
Sodium	___________________
Strontium	___________________
Copper	___________________
Potassium	___________________
Iron	___________________
Barium	___________________

Part 2 Identification of Solid Unknowns

Solid unknown # ___ Color of flame ________ Element _________

Solid unknown # ___ Color of flame ________ Element _________

Part 3--Continuous and line spectra (Draw what you see through the spectroscope. Include colors and wavelenghts)

Incandescent light bulb:

Fluorescent light bulb:

Question 1: How are these spectra different?

Question 2: Which is a line spectrum?

Part 4 Identification of element in fluorescent light bulbs

Draw what you see through the spectroscope when you examine each of the following gas discharge tubes. Include colors and wavelengths. Also, describe the color of the glowing discharge tube.

Helium

Spectrum

Color of the glowing discharge tube: ____________________

Neon

Spectrum

Color of the glowing discharge tube: ____________________

Mercury

Spectrum

Color of the glowing discharge tube: ____________________

Hydrogen

Spectrum

Color of the glowing discharge tube: ____________________

Compare the emission spectrum of these elements with the emission spectrum of the fluorescent light bulb. Which element is present in fluorescent light bulbs? Explain your reasoning.