

Experiment: Spectroscopy

Introduction to light

Light is a form of energy called electromagnetic radiation. A chart of the electromagnetic spectrum is shown below.

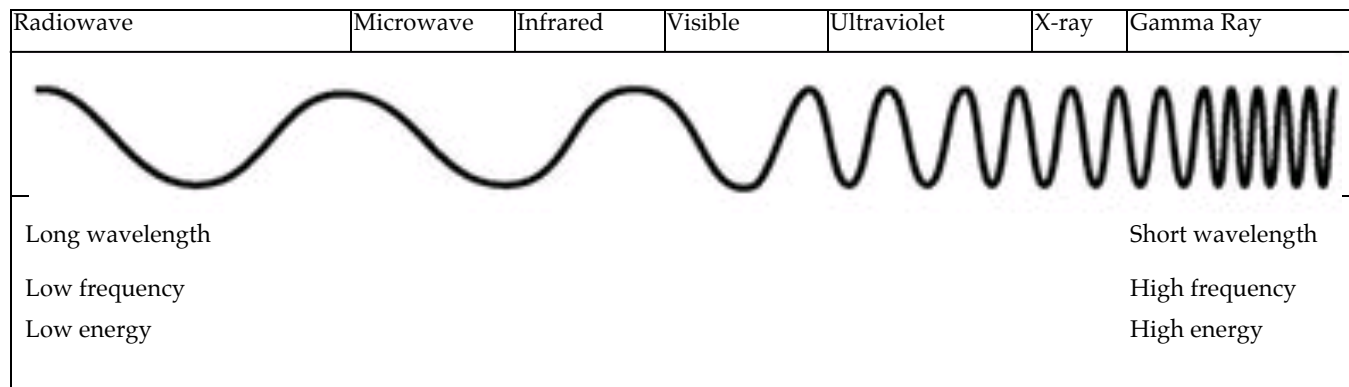


Figure 1. Types of electromagnetic radiation

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

Wavelength is denoted by the symbol lambda (λ) and 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) A micron (a micrometer, or 10^{-6} meters). It's often used in biology since it is the appropriate size of cells and bacteria. You will encounter microns when using most microscopes.

(b) An Angstrom (\AA). One Angstrom is 10^{-10} meters and is used for describing the diameter of atoms.

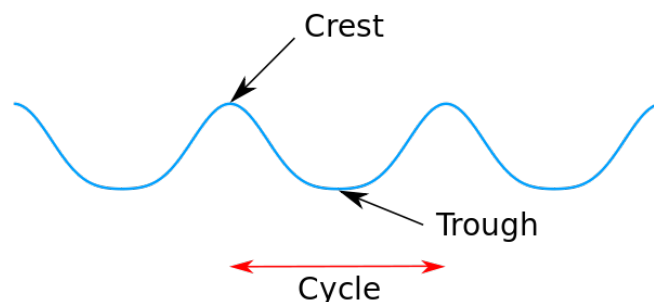


Figure 2. Terminology of a wave.

Source: Wikimedia Commons

The second variable used to describe waves is called **frequency**, which is the number of cycles which pass per unit time (see figure 2 above for "cycle"). 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).

❖ **Shorter \downarrow wavelength = higher \uparrow 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.

❖ **Higher \uparrow frequency = Higher \uparrow energy.**
This is a direct relationship. (See the chart above to confirm this.)

Let's compare radiowaves (a) with ultraviolet radiation (b):

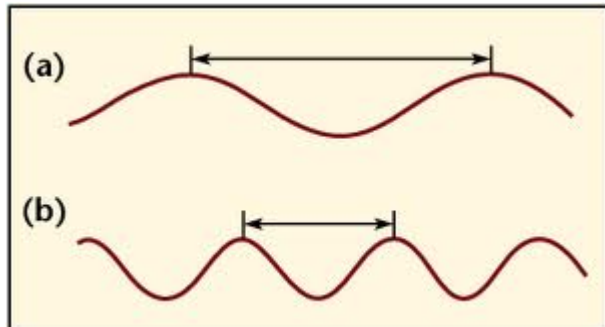


Figure 3. Comparing two waves corresponding to different types of electromagnetic radiation.

Which type of radiation is higher energy?

Looking at **radiowaves**, when you tune your radio you are receiving radio waves with frequencies between 530 and 1700 kHz for AM and 87.5 to 108 MHz for FM. (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. These waves pass through us without harm.

Sunscreens protect your skin from **UV radiation** by absorbing 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. Thus, photons of UV light have a higher energy than photons of radiowaves.

Both radio waves and ultraviolet light are invisible to the human eye. Colors are due to visible light, which have wavelengths between radiowaves and UV, with a range between 400-700 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. The visible spectrum (the "rainbow") 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.

Origin of Spectral Lines

The origin of the spectral lines can be explained using the planetary model of the atom, in which the nucleus is in the center and surrounded by orbits (n) on which electrons travel (note – this is not completely accurate – this is an oversimplified model). Atoms are usually in a low energy state, which means their electrons are in the lowest energy orbits such as level 1 ($n=1$). 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 (in figure 4a, the electron jumps from level 1 to the second level, $n=2$). This energy can be added to atoms by an

electric discharge or by heat, an incoming photon. This added energy is emitted when the atom gives off a photon and the excited electron returns to the ground state (see figure 4b).



Figure 4a shows an atom in its ground state, with the electron (green dot) moving from level 1 to level 2 by absorbing energy from an incoming photon. Figure 4b shows an atom in an excited state emitting a photon as its electron moves to a lower energy level. Only two levels ($n=1$, $n=2$) are shown

The light emitted by atoms have a definite wavelength and if visible, a 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.

Objectives

In this experiment you will

- ✓ Learn about light and the relationship between energy, wavelength and frequency.
- ✓ Observe colors from flame emission of various elements and identify an unknown based on the flame color.
- ✓ Observe continuous vs. line spectra resulting from the emission from gas discharge tubes and identify the unknown element in fluorescent light bulbs.

Hazards

☢ Use caution with the Bunsen burners and wooden splints. Wear goggles and tie hair/loose clothing back to prevent fires.

☠ 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.

Procedure

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

1. Light a bunsen burner. Adjust the air vent to obtain a blue flame.
2. Obtain two small beakers filled about 1/3 full with tap water. One will be used to moisten a wooden splint, and the other to extinguish your wooden splint.
3. Obtain a vial containing a known sample (note the element symbol on the vial) and a piece of weigh paper. Using a spatula, take a scoop of solid from the vial onto the weigh paper. Dip a wooden splint it into one of the beakers of water to moisten it. Gently roll the splint onto the weigh paper with the solid so the tip of the splint is covered with sample.
4. Place the splint into the clear blue flame and observe/record the color of the flame for that element. Dip the splint into the other beaker of water to douse the flame.
5. Repeat this procedure for each known.

Part 2: Identification of an unknown solid

Repeat the above procedure for two unknowns. Check your results with your instructor before you proceed.

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. **(See Hazards note.)**
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 H₂ (and others).
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.

Report

Spectroscopy

Name _____ Section _____

Lab Partner _____

Data and Questions**Part 1: Flame tests for known elements**

<u>Metal Ion</u>	<u>Color of flame</u>
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 wavelengths)

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. Based on your results, which element is most likely present in fluorescent light bulbs? (Note: Other elements not listed here may also be present in the bulbs.) Explain your reasoning.

Prelab

Spectroscopy

Name _____ Section _____

Complete the following questions **BEFORE** class.

1. a. List the colors in the visible spectrum, from highest to the lowest frequency.

b. Which color of visible light has the lowest energy? _____

The longest wavelength? _____

The lowest frequency? _____

2. The following diagrams of waves are labeled λ_1 , λ_2 , and λ_3 . Fill in each blank with the best choice.

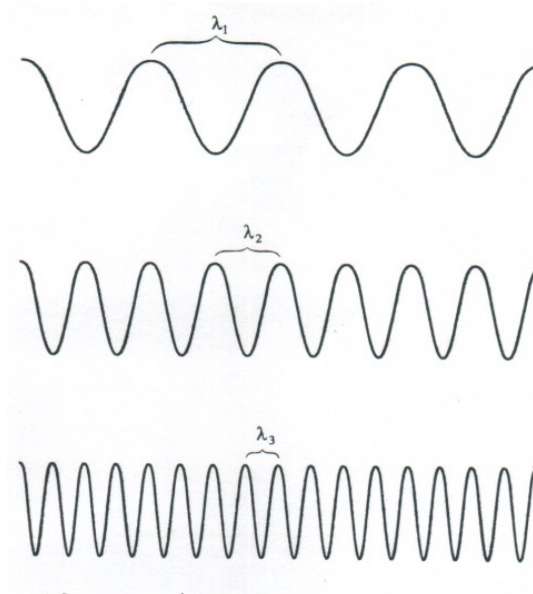
a. Which wave has the longest wavelength? _____ The shortest wavelength? _____

b. Which wave has the greatest frequency? _____ The smallest frequency? _____

c. Which wave would be a good model for UV light? _____ Radio waves? _____
Visible light? _____

d. Which wave has photons with the greatest energy? _____

The smallest energy? _____



3. What is the difference between a continuous spectrum versus a line spectrum?