Emission Spectra of Atoms – Neon signs

Question

Neon signs - the same or different?

Skills

1. Construct appropriate tables.
2. Explore and document the pattern of emission wavelengths as a “fingerprint” for various elements.
3. Quantitatively and quantitatively compare emission spectra of two lights to determine if the product is a potential replacement for the competitor’s lights.

Introduction

After recently graduating from GVSU with a degree in chemistry, you have been hired by Flashy Lights Inc., a company that wants to break into the stop-motion strobe photography business. The standard in the industry is a xenon arc flash tube. When an atom of xenon receives energy from the electric current, the energy can cause one or more electrons in that atom to transition between a lower atomic energy level and a higher atomic energy level. This excess energy is released when it is emitted in the form of visible light. Your new employers want to produce a less expensive lamp than their competitors, Blinky Light LLC. Blinky Light, LLC makes a high-quality xenon arc flash tube that retails for $120.95 (Topbulb). Your job is to record the emission spectrum of the lamp made by Flashy Lights Inc. using a spectroscope. Then, use the spectroscope to measure the emission spectrum of Blinky Lights LLC, and determine whether Flashy Lights Inc. was successful at mimicking the competition.

The Interaction of Energy and Matter

Atoms in their lowest energy state can absorb light (photons). In an absorption event, when a photon interacts with the atom, the light provides energy for the atom to transition between a

lower-energy and a higher-energy state. This transition involves one or more electrons being promoted from one or more of the lower quantized energy levels to one or more of the higher quantized energy levels. Atoms can also emit light if they have first had energy added to them to put them into a high energy (excited) state. When emission

occurs, light is given off (emitted) from the atom as a result of one or more electrons transitioning from a high energy level to a lower energy level.

For example, if neon gas is provided with energy in a high-voltage tube, the energy is absorbed by the neon atoms, putting them into a high energy state. When the atoms relax, the atoms emit light. These emissions give the orange-red color characteristic of neon signs. As is true for all atomic gases, the photons emitted do not have many different wavelengths, but only a few; the wavelengths correspond to the differences in energy between the allowed energy levels of the neon atoms. The relationship between the energy (E) of a photon of light and its wavelength (λ) is shown in equation 1, where h is Planck's constant and c is the speed of light.

$$E = \frac{h\,c}{\lambda} \qquad \text{(Equation 1)}$$

Historical Development: The Puzzle Posed by Atomic Line Spectra

Atomic spectra were being studied by the middle of the 1800s. (Robert Bunsen invented the Bunsen burner as a tool for spectroscopy.) By 1900 the wavelengths emitted by many different atoms had been cataloged, but there was no general understanding of why particular wavelengths (colors) were emitted by particular atoms. Work by Johannes Rydberg, Albert Einstein, Niels Bohr, and others led to an explanation of the lines and further development of atomic theory. A new theory, quantum mechanics, was developed between 1925 and 1927. Because the electronic structure of each element's atoms is unique, the spectral lines are also unique. The wavelengths/colors of the spectral lines are determined by the electron energy levels within the atom.

Quantum Mechanical Picture of the Hydrogen Atom

The energy levels for electrons in an atom can be represented graphically by horizontal lines, as shown in Figure 1. Since the hydrogen atom has only one electron, we draw a single arrow to indicate which energy level the electron is in. In Figure 1, the electron is in the lowest quantized energy level, so the hydrogen atom is in its lowest (ground) energy state.

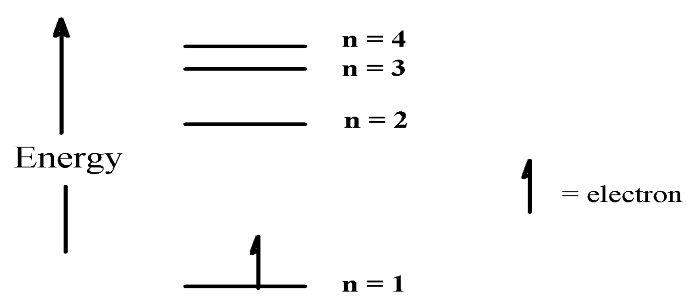

Figure 1. Energy levels for the hydrogen atom.

Most atoms at room temperature are in their lowest energy states. If energy is supplied, the atom moves to an excited state because one or more electrons moves to higher energy levels. When an electron moves from a higher to a lower energy level, light is emitted. Figure 2 shows an example of an electronic transition for hydrogen that would result in the emission of light.

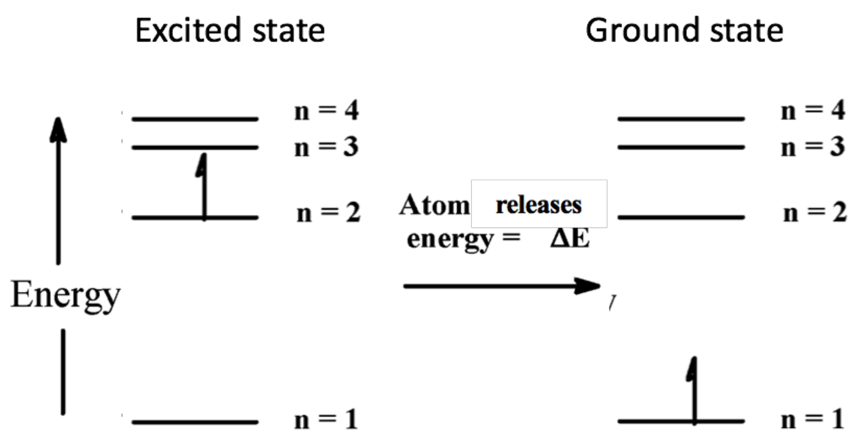

Figure 2. Atomic emission of energy

If the electron in the hydrogen atom were initially excited to the n=2 level, the electron can only fall down to the n = 1 level and only one color of light can be emitted. However, if the electron were excited to a higher level to begin with, that electron could fall down to any of the lower levels. For example, an electron initially excited to n=4 could fall down to n=3, n=2, or n=1. Each of those possible transitions would result in the emission of a different wavelength (color) of light. Only some of the wavelengths (colors) are visible to human eyes.

The photon emitted by the atom must carry away exactly the amount of energy that is the difference between the initial (before-emission) level and the final (after-emission) level. The energy of the light emitted corresponds to the difference or differences between the energy levels of the atom. Electrons in atoms can access different energy levels (such as n = 1, 2, 3, etc. for hydrogen). Every atom has energy levels (n = 1, n = 2, etc.), but the calculated differences in energy between the levels are different for each element.

Light, the Electromagnetic Spectrum, and Spectroscopes

Most light bulbs installed to illuminate rooms appear to our eyes to emit white light, which is a combination of many wavelengths of visible light. A *spectroscope* is a device that separates the light into different colors. If you view white light through a spectroscope, you will see the colors separated into a *spectrum*. A spectrum that contains only a few well separated colors is a *line spectrum*, while a complete rainbow-like band of light is called a *continuum spectrum*. If you are physically in the lab, you and your lab partners will work with the simple spectroscope shown in Figure 3.

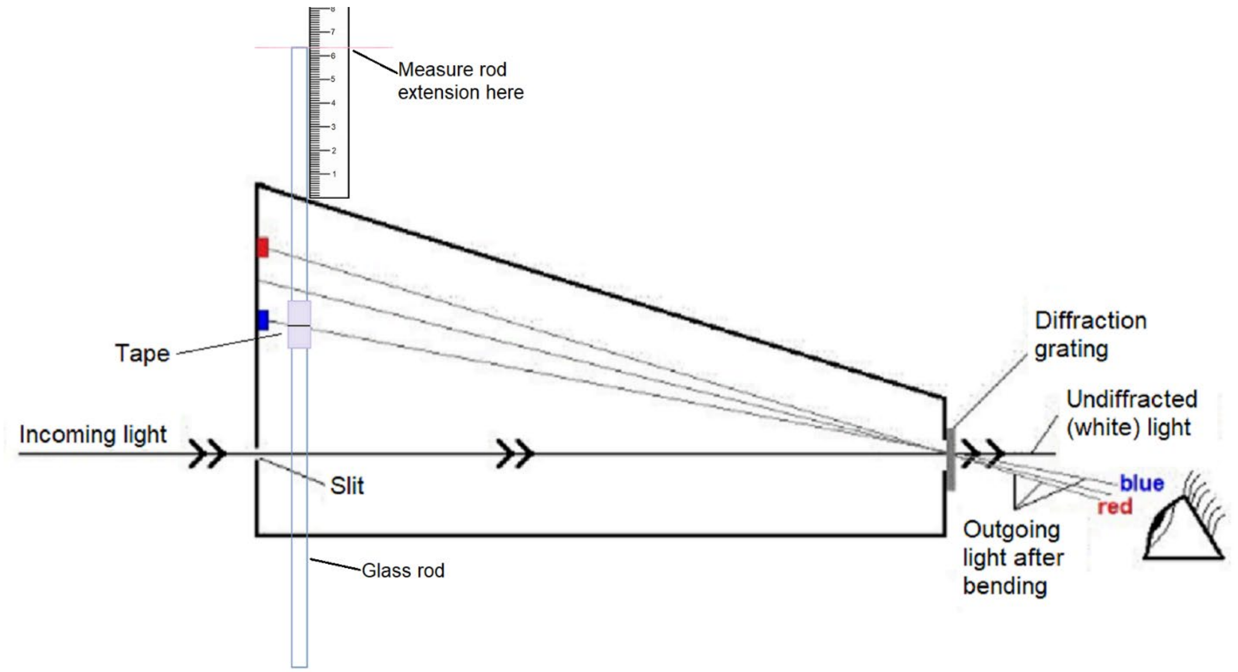

Figure 3. Observation of light with a box spectroscope, ruler shows where to measure.

This spectroscope uses a piece of plastic with hundreds of closely spaced parallel scratches, a diffraction grating, to separate the light of different wavelengths. The diffraction grating acts like a prism. When light goes through a diffraction grating (or prism), its path is bent through some angle; the longer the wavelength, the greater the bending.

If you point your spectroscope at a fluorescent bulb, you see both the continuous (rainbowlike) spectrum emitted by the phosphor (coating inside the bulb), and, when the slit is narrow, the line spectrum emitted by the gaseous mercury atoms (fluorescent bulbs are filled with gaseous Hg atoms). You will record the location of the five mercury emission lines visible with this box spectroscope as part of the lab. For calibration, three sets of measurements are needed. You and your labs partners will each measure the five visible lines in the mercury spectra using the same spectroscope. The measurements will be used to create a calibration curve and to determine error ranges in the calibration measurements.

Claim, Evidence, and Reasoning Arguments

The development of scientific argumentation skills happens in different ways. One common method is the use of the Claim, Evidence, Reasoning (CER) framework. In this model, the *claim* is a statement that answers the original question. In this lab, the question involves whether or not the lamp made by your new employer, Flashy Lights Inc., is the same as the light made by their competitor, Blinky Light LLC using the measured emission spectra. The *evidence* is qualitative (observations) and quantitative (measurements) data that support your claim. Evidence comes from experiments, measurements, observations, research, readings, data sets, etc. Finally, reasoning is an explanation of the logic connecting the claim and evidence. The *reasoning* also includes scientific principles that support the claim and evidence. In this experiment, information about emission spectra and the structure of atoms will be important to include in your reasoning.

Connecting Wavelength, Frequency, and Energy

Physically, you can think of electromagnetic radiation in two different ways. First, you can think of it as a spread-out wave of electric and magnetic fields. Figure 4 shows how the electric field would be arranged in space if you could take a "snapshot" of such a wave; at some places the electric field points up, at some there is no electric field, and at some the field points down.

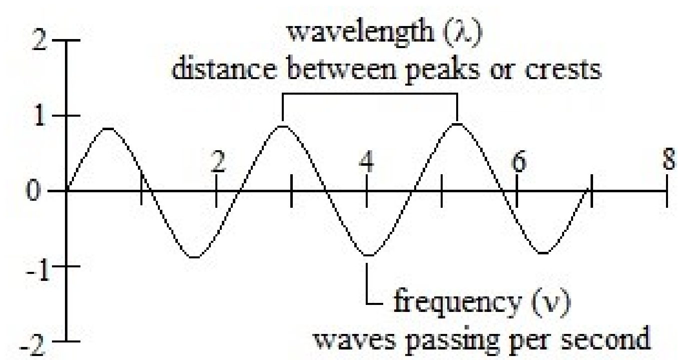

Figure 4. Wavelength and frequency

If you took several such snapshots in a row, you would find that the whole picture moves to the right very rapidly (at the speed of light). Therefore, if you stood in one place and watched the electric field at that spot, you would see it go up and down, just like a cork bobbing on the ocean as waves go by.

The distance between two crests of the wave is called the wavelength and is given the symbol λ (the Greek letter lambda). For ordinary light the wavelength is small, about 5×10^{-7} m, and is usually reported in nanometers. (1 nm = 10^{-9} m). Our eyes can see only radiation with wavelengths between about 400 nm (violet) and about 700 nm (very deep red).

If you stand in one place and watch the electric field go up and down, the number of up-and-down cycles you see per second is called the frequency of the wave and is given the symbol ν (the Greek letter nu, not to be confused with the ordinary Roman letter v). The frequency has units of "number per second", or Hertz (1 Hz = 1 s^{-1}). Since ν crests go by one spot per second, and the distance between crests is λ the overall speed of the wave is $\nu\times\lambda$. This product has the same value for all the different kinds of electromagnetic radiation, and is called c, the speed of light. We write

$$c = \lambda \nu \quad \text{(Equation 2)}$$

where the constant value c = 2.998 x 10^8 m/s. You can see from this formula that as the wavelength of light increases, its frequency must decrease. Red light has a longer wavelength and a lower frequency than blue light. From equation 2 you can find that visible light with wavelength 500 nm has a frequency of about 6×10^{14} Hz.

The second way to think about electromagnetic radiation is as a collection of little packets of energy, called "photons", that move along at the speed of light. Photons have no mass, but each photon carries a certain amount of energy. The energy carried by a photon is proportional to the frequency of the corresponding wave, as expressed in the Einstein relation,

$$E = h\nu \quad \text{(Equation 3)}$$

where E is the energy of one photon in Joules (J), h is Planck's constant, with value h = 6.626×10^{-34} Js, and ν is the frequency in s^{-1} as before.

The photon energy is proportional to the frequency, so it is inversely proportional to the wavelength through equation (2). Photons of light with longer wavelengths therefore carry less energy. Photons of red light carry about half as much energy as photons of blue light. Each photon of 500 nm visible light carries an energy of about 4 x 10^{-19} J.

The energy of light can be calculated from equation (3). This energy represents the energy emitted (or absorbed) when an electron transitions between energy levels in an atom.

Neon Signs the Same or Different?

You are now ready to get to work on your first project at Flashy Lights Inc. Your employer wants to manufacture a less expensive arc flash bulb than the xenon one sold by your competitor. Your manager has asked you to determine the feasibility of making an arc flash bulb by (1) combining (1) a noble gas (helium, neon, argon, or krypton) with (2) another element (carbon, nitrogen, carbon, zinc, selenium, or strontium) to make a bulb with a similar spectrum to that of xenon.

Experiment

Equipment and Materials

Spectroscope	Flashlight	Power Supply
Light bulbs	Ruler	Colored Pencils

Procedures

Safety **Warning:** The light bulbs are connected to a high voltage supply. Ask your instructor to change them if needed. Do not change them yourself, as touching the tubes or metal may result in electric shock.

Part A – Assemble and calibrate the spectroscope with the mercury lights (fluorescent lights) and make a calibration curve using *Microsoft Excel*® - PC or Mac. (If your computer does not have Excel, you may use *Excel*® on a computer available in the lab. The Spectroscope lab has the directions for getting the GVSU site license version of *Excel*® for no charge. Or you may use Google Sheets but realize that program does not have all the functions available. You may not use the Web version of *Excel*®.) The tasks for this part will be similar to the calibration tasks from the spectroscope lab.

Part B – Verify the spectroscope calibration by collecting data for the hydrogen lamp to practice your technique. Then check the results of your measurements against known wavelengths from the hydrogen emission spectrum

Part C – Design and carry out an experiment to collect data from both the *Flashy Lights Inc.* lamp and the *Blinky Lights LLC* lamp to determine if they are the same. Make a claim, identify evidence, and provide an explanation that is supported by your results.

References

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Acknowledgement

Development of this laboratory was made possible by Dr. Julie Henderleiter, Dr. Brittland DeKorver, and Mary Jo Smith, GVSU 2019 with revisions by Dr. Deborah Herrington and Mary Jo Smith, GVSU, 2021.