Flame Test Formal Lab Report; SC4- LDC

L1: How can atomic absorption and emission spectroscopy be used to identify various metal ions?

After reading the two articles provided, write a report that addresses the question (L1). Be sure to provide examples from the two texts. In your report you will also need to incorporate the rerun format after conducting the lab. So your entire written report will include a detailed response to the L1 prompt in addition to a rerun conclusion. Be sure to always identify any gaps or unanswered questions from the lab (**L3**).

A bibliography is not required. The “rerun” format is outlined on the last page of this file.

Please note that you will also have a traditional lab report (title, purpose, standard, data table, questions etc.) stapled to your typed response to the above prompt.

Article #1- The Mysterious Mr. Biv.

Lexile: 880L
Publication: Odyssey(Dec2007)
Author: Cox, Mary Beth

Sir Isaac Newton was the first to make Mr. Biv's acquaintance. The two met in the year 1666. Newton's most famous discovery, that gravity causes apples to fall and moons to orbit, was still two decades in the future. In 1666, Newton experimented not with gravity but with light.

Sir Isaac discovered that a crystal prism separates a beam of white sunlight into a band of many colors. The colors fan out in a regular pattern: first red, then orange, yellow, green, blue, indigo, and finally violet. The pattern spells a name: Roy G. Biv.

At first glance, it seemed innocent enough. Mr. Biv was a rainbow, more or less. But hidden from Newton's inquisitive gaze was a dark secret — a secret that would embroil Mr. Biv in a great scientific mystery.

The trouble started in 1814. An employee of the Munich Philosophical Instrument Company, Joseph von Fraunhofer, had his own fateful encounter with Mr. Biv. Fraunhofer specialized in the manufacturing of quality optical glass. In the course of his duties, Herr Fraunhofer discovered something unsettling about Mr. Biv. The company's high-resolution optics didn't spread sunlight into the unbroken spectrum Sir Isaac observed. Rather, dark lines repeatedly interrupted the colors. Fraunhofer had stumbled upon Mr. Biv's secret! Hundreds of slices of color were missing from the Sun's spectrum! Fraunhofer was flabbergasted. He couldn't fathom a reason for those dark lines. Sadly, he never would. Fraunhofer died prematurely of tuberculosis at age 39. He left the mystery of the missing colors unsolved and his questions about Mr. Biv unanswered.

The amazing truth wouldn't be known until 1859. Two chemists, Robert Bunsen and Gustav Kirchhoff finally exposed Mr. Biv's true character. Bunsen and Kirchhoff ran a business selling samples of chemical elements to research scientists. To check the purity of their wares, they used a "flame test." They heated their samples in a flame. The color of the flame helped them identify the elements in the sample. Sodium turned the flame yellow. Potassium turned the flame violet. Lithium turned it red, and so on.

Bunsen improved the business by enhancing the flame test. First, he developed the famous "Bunsen burner," now used in school labs worldwide. The Bunsen burner produced a clean, clear, almost colorless flame perfect for identifying elements. Then Bunsen borrowed an idea from Sir Isaac. He directed the light from flame tests through a prism, just as Newton had done with sunbeams. The results were nothing short of revolutionary.

Bunsen discovered that the flames weren't just one color. Like sunlight, they were actually a blend of colors. The prism separated the light from flame tests into sharp, distinct colored lines. Each element had a unique flame spectrum, much like its own colored line fingerprint. As with any suspect, the fingerprint identified the element beyond question. This method of using light to make identifications became known as spectroscogy.

Gustav Kirchhoff proved what a powerful tool spectroscopy could be. Kirchhoff became an expert at identifying elements by their spectra. And he discovered two previously unknown elements by detecting unusual patterns of colored lines. One element he named rubidium from the Latin word for "red" and the other cesium from the Latin word for "sky blue."

Kirchhoff had heard rumors of the mysterious Mr. Biv. He decided to examine the dark lines in the Sun's spectrum himself. To his amazement, he found that the Sun's dark lines exactly matched the flame spectra of elements. The elements' fingerprints slid into the Sun's dark lines just like pieces of a puzzle. Kirchhoff had found the slices of missing color!

It was then that Gustav Kirchhoff had a stroke of inspired genius. He realized that the mysterious dark lines and the colorful fingerprints were two sides of the same coin. When exposed to a flame, elements release colored light. But, under more extreme conditions like those in the Sun, the very same elements absorb colors. Elements suck their signature fingerprint colors right out of the Sun's spectrum. Mr. Biv's missing colors are evidence that the Sun and Earth are made of the same stuff. Like Sir Isaac's gravity, elements are everywhere in the universe.

To this day, Mr. Biv's mystery has had lasting consequences. Spectroscopy is an important tool for astronomers, who use it to glean information from the light of stars, planets, galaxies, and nebulae too far away to ever be visited. NASA has four Great Orbiting Observatories that monitor light in space. The Hubble Space Telescope observes visible light — the very same colors that Roy G. Biv represents. The other three observatories watch for light not seen by the human eye. The Spitzer Space Telescope observes the infrared spectrum, which includes the invisible colors before Roy. The Compton Gamma Ray Observatory and the Chandra X-ray Observatory handle the other end of the light spectrum, which includes the invisible colors after Biv.

Spectroscopy is also used by chemists for research, in manufacturing, and to test the environment. But don't think that chemists have completely yielded the skies to astronomers. At the end of this very month, you might witness a few breathtaking flame tests in the night sky. Most people call them fireworks. Excited by the energy of a detonation, the elements in fireworks pour forth their fingerprint colors at high altitude. It's celebratory spectroscopy. To know which element glows what color, you only have to remember one name: Roy (strontium, calcium, sodium) G. (barium) Bi (copper) v (copper + strontium).

Article #2 →

Article #2- Principle of Atomic Absorption /Emission Spectroscopy- Flame Test Lab LDC

5.1 ATOMIC EMISSION-THE FLAME TEST

When a small amount of a solution of a metal ion is placed in the flame of a Bunsen burner, the flame turns a color that is characteristic of the metal ion. A sodium solution gives a yellow color, a potassium solution results in a violet color, a copper solution gives a green color, etc. Such an experiment, called the flame test, has been used in conjunction with other tests in many qualitative analysis schemes for metal ions. Whatever color our eye perceives indicates what metal ion is present. When more than one metal ion is present, viewing the flame through a colored glass filter can help mask any interference. Figure 1 shows this experiment.

The phenomenon just described is an "atomic emission" phenomenon. This statement may seem inappropriate, since it is a solution of metal ions (and not atoms) that is tested. The reason for calling it atomic emission lies in the process occurring in the flame. One of the steps of the process is an atomization step. That is, the flame converts the metal ions into atoms. When a solution of sodium chloride is placed in a flame, for example, the solvent evaporates, leaving behind solid crystalline sodium chloride. This evaporation is then followed by the dissociation of the sodium chloride crystals into individual ground state atoms -a process that is termed atomization. Thus sodium atoms are actually present in the flame at this point rather than sodium ions, and the process of light emission actually involves these atoms rather than the ions.

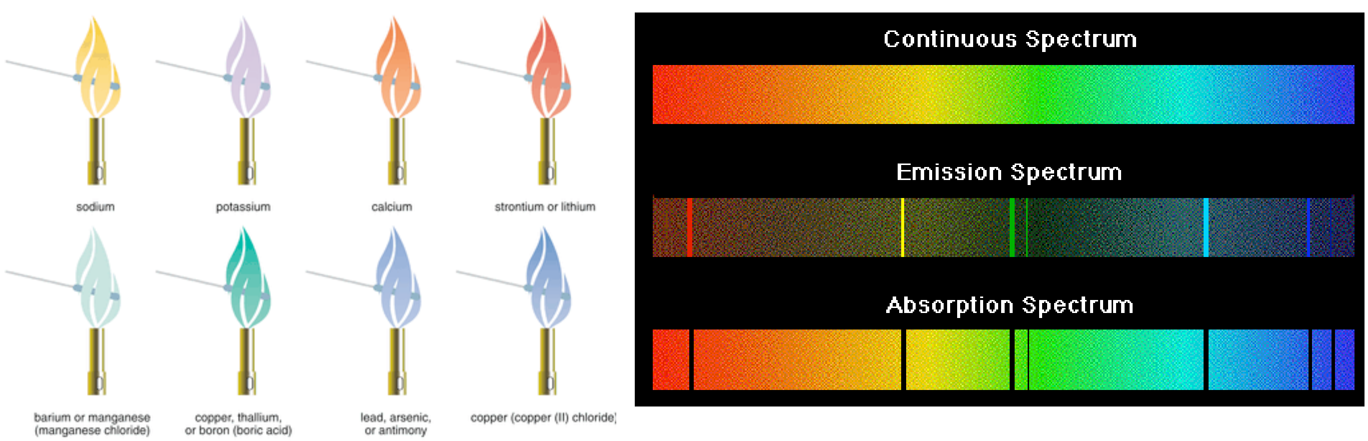


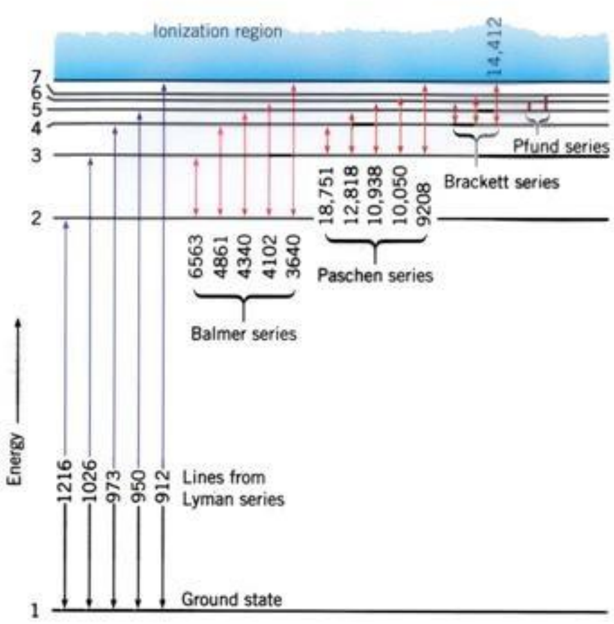
Figure 1. The flame test: observation of the flame color (a) without colored glass and (b) through prism

$$\frac{Na^+}{Cl^-} \xrightarrow[\text{of solvent}]{\text{evaporation}} NaCl \xrightarrow{\text{atomization}} \frac{Na}{Cl} \xrightarrow[\text{of metal atoms}]{\text{excitation}} Na^* \xrightarrow[\text{ground state}]{\text{drop back to}} Na \xrightarrow[\text{emission}]{\text{light}}$$

Figure 2. The sequence of events occurring in the flame test

As with fluorescence, the atomic emission is a result of electrons dropping from an excited state to lower states. The difference is that (1) atoms are involved here, rather than molecules, and (2) light is not absorbed prior to this atomic emission. Following atomization, a small percentage of the atoms absorb sufficient energy from the flame (as opposed to a light beam) so as to be promoted to an excited state. As with molecules in fluorescence, these atoms quickly return to a lower state, and light corresponding to the energy that is lost in the process is generated. It is this light that our eye perceives. The complete sequence of events is depicted in Figures 2 and 3.

The discussion of the facts regarding atomic energy levels and molecular energy levels presented in the previous three chapters is applicable here. Since there are no vibrational levels in atoms, the energy of emission is a discrete amount of energy corresponding to the difference between two electronic levels. Also, since there are usually a number of electronic levels to which an electron in an atom can be promoted, there are a number of possible discrete energy jumps back to the lower energy states. These represent a number of distinct wavelengths of light to be emitted. What is actually emitted by the atoms in a flame is then a line emission spectrum as indicated in Figure 4. (Compare with Figures 10 and 11 in Chapter 12). Figure 5 depicts an explanation of the atomic emission phenomenon. When atoms fall back to lower energy states following the absorption of energy from a flame, a line spectrum is emitted which our eye perceives as a particular color of light. Each kind of atom is different in terms of the separation between energy levels and the line emission spectra are therefore different. Because of this, different elements are found to emit the different colors noted earlier.



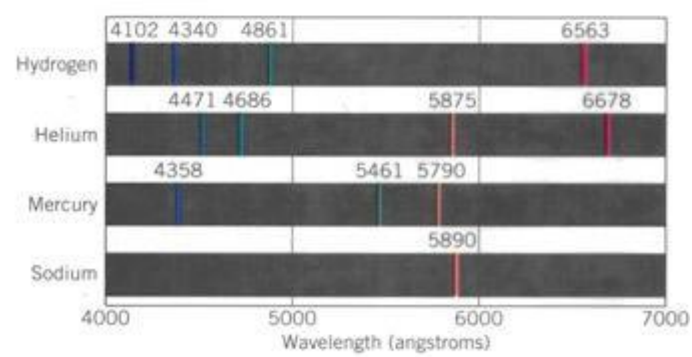


Figure 4. The atomic emission phenomenon: (a) Allowed energy jumps back to lower levels. (b) A line spectrum representing the wavelengths of the emitted light corresponds to the jumps

Flame Test Formal Lab; SC4 **(LDC Module #1)**

Chemists began studying colored flames in the 18th century and soon used "flame tests" to distinguish between some elements. Various elements burn with different colored flames. The purpose of this lab is to investigate atomic absorption and emission spectroscopy through the use of flame tests. Although some of the flames you will be seeing will appear similar in color, their light can be resolved (separated) with a prism into distinctly different bands of colors on the electromagnetic spectrum (ROYGBIV). These bands of colors are called **atomic line spectra**, and they are UNIQUE to each element. Niels Bohr studied the line spectrum for hydrogen, and wondered what the specific line spectrum had to do with the structure of the atom. He postulated that an electron can have only specific energy values in an atom, which are called **energy levels**. Bohr believed that the energy levels for electrons were **quantized**, meaning that only certain, specific energy levels were possible. How does an electron move between energy levels? By gaining the right amount of energy, an electron can move, or undergo a transition, from one energy level to the next. We can explain the emission of the light by atoms to give the line spectrum like this:

1. An electron in a high energy level (excited state) undergoes a transition to a low energy level (ground state).
2. In this process, the electron loses energy, which is emitted as a photon (a particle which behaves like a wave)
3. The energy difference between the high energy level and the low energy level is related to the frequency (color) of the emitted light.

Pre-lab questions:
1. Bohr's important discovery was that energy levels of electrons are quantized (only existing in certain, specific levels). In what year was this discovery made? _____________

2. What happens to an electron when energy is added?

3. What is released when an electron loses energy?

4. What determines the frequency (color) of photons?

5. Why do you think the frequencies (color) for a specific element is always the same?

Procedure: In this lab, you will be observing the colors of the flames for 9 different elements: **lithium, zinc, cobalt, sodium, potassium, calcium, strontium, barium,** and **copper.** Each element is dissolved in a solution of its chloride salt. There is a different solution at each lab station. You will go around to all 9, perform the flame test, and make **CAREFUL** observations of the colors. You will then be given an unknown solution, for which you will have to use your notes below to determine which unknown you were given.

Data Table #1:

Metal Name (salt formula)	Description of Crystals	Observed Flame Color
Calcium ($CaCl_2$)		
Sodium (NaCl)		
Cobalt(II) ($CoCl_2$)		
Lithium (LiCl)		
Copper ($CuCl_2$)		
Potassium (KCl)		
Strontium ($SrCl_2$)		
Zinc ($ZnCl_2$)		
Barium (if available)		
Unknown #1		
Unknown #2		

Based on your observations, what are the identities of your 2 unknowns?

Unknown # _____ is _________ **Unknown # _____ is __________**

Post- Lab Questions:

1. If you had 2 colors that seemed identical, how could you tell them apart more accurately?

2. Albert Einstein determined this equation:
energy (in joules) of a photon is equal to **Planck's constant** times the **frequency** of the light:
E = h • υ • Frequency (υ) has units of 1/sec (which is a Hertz, or Hz)
• Planck's constant (**h**) = 6.63 x 10-34 J·sec

a) If the frequency of a **red** spectrum line is at 1.60 x 1014 Hz, how much energy does each photon of this light have?

b) If the frequency of a **violet** spectrum line is at 2.50 x 1014 Hz, how much energy does each photon of this light have?

c) On the far ends of the visible spectrum of light, there exists ultraviolet (**UV**) radiation and infrared (**IR**) radiation.
- **UV radiation is dangerous.** UV radiation is located just past violet on the spectrum.

- **IR radiation is harmless.** It is located just past red on the spectrum.
- Based on what you calculated in parts a & b, explain -why- UV is more dangerous than IR:

Use the ideas below to help you create an RERUN outline to draft your conclusion before you type it in paragraph format as part of this formal lab report. **This is in addition and incorporated into the L1 writing prompt discussing the two articles.**

Recall what your group did during this lab (think about the procedures and calculations).

Explain why you did this lab and what you were trying to find out (refer to the purpose).

Reflect on the lab's meaning and your results (what did you determine) and did it match your hypothesis or was your hypothesis wrong? Why?

Uncertainty (errors that were in the lab that you could not control or just any errors that you came across during this lab that you could fix for the next time you perform the lab). This should be very specific and related to the procedures. For this lab, reflect on the answer for analysis question #1-5. Think about differences between group's data, included human error, among other reasons, that would cause our average atomic mass for Candium to be different.

New questions or new discoveries (normally you put in at least two) generated from this lab.