Chem. 6A: Atomic Spectra and Flame Test **Dr. Mack**

Pre–lab Assignment (*to be submitted at the beginning of class*)

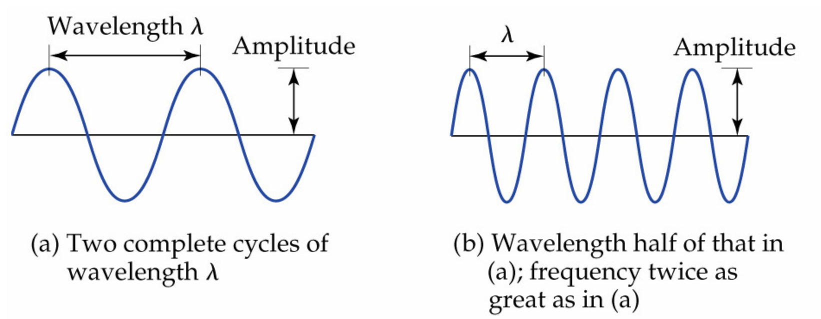
In addition to a brief summary of the experiment, perform an Internet search to find compounds that fireworks and pyrotechnics manufacturers commonly use to produce specific colors. List the color along with the compound and its chemical name. List at least three compounds.

Part 1: Atomic Line Spectra (work in groups of no more than 3)

Introduction:

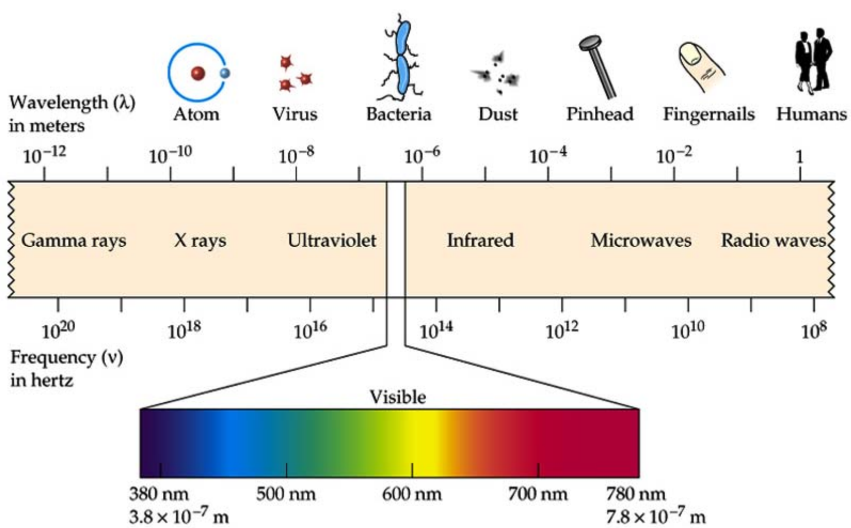
What is light? Visible light is a member of the family of radiant energy Radiant energy travels in waves and can radiate in all directions. The wavelength of a wave is the distance from crest to crest or trough to trough. The shorter the wavelength of a wave, the more energy that wave has. Waves of longer wavelength are of lower energy.

Figure 1:



Radio and television waves, radar waves, microwaves, infrared rays, visible light, x-rays, ultraviolet rays, and gamma rays are all forms of radiant energy. (The family of radiant energy is often called the electromagnetic spectrum due to the electronic and magnetic properties of radiant energy.)

Figure 2:



Visible light includes those forms of light that the human eye is capable of detecting. Notice that visible light is a small component of the electromagnetic spectrum (Figure 2). White light is visible light that comes from the sun and includes all of the colors of light (ROYGBIV). Each color of light is of a specific wavelength. Red light is made of radiant energy with wavelengths around 650 nm while blue–violet light has a much shorter wavelength, about 450 nm. (1nm = 10^{-9}m) When white light passes through a prism or off of a diffraction gating, it is separated into its component colors and a continuous rainbow of is observed. This smear of color is called a continuous spectrum. The sun or an incandescent source of light (regular light bulb) produce such a spectrum. Atomic sources of light however do not.

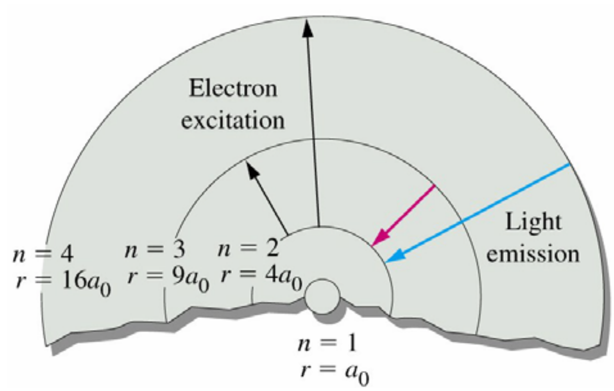
Atomic Emission:

Let's consider the modern model of an atom as proposed by Bohr. Electrons reside in discrete energy levels called "orbitals". The electrons cannot exist between these orbitals. When an electron jumps to a higher level, it must fall back down. When this occurs, light is emitted.

Figure 3:

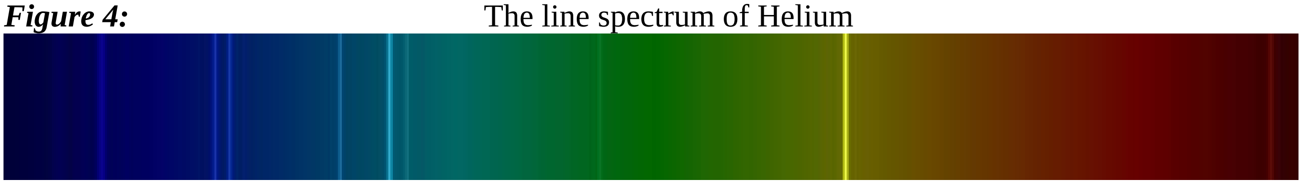
Each drop from a higher level results in the emission of light. The grater the drop, the higher the energy of the light emitted and therefore the shorter the wavelength. (more towards the blue side of the spectrum)

Excitation can be attained by an electrical discharge, a flame source or the absorption of light



When an atomic source of light passes through a prism or is viewed through a ***spectroscope***, the light appears as lines rather than a continuous spectrum.

Figure 4: The line spectrum of Helium



When one views the light emitted from an atomic source with the naked eye, the light appears as an overall color since our eyes are incapable of discerning the individual wavelengths that correspond to the lines. As a result, a discharge lamp containing neon appears red to the human eye.

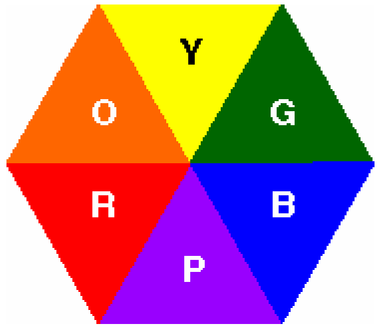
Fireworks and pyrotechnic manufacturers use the different emission characteristic of elements to produce the brilliant panorama of colors one sees in their shows. By adding certain amounts of different elements to the explosive shell, they produce the purple, green and orange hues that light up the night sky.

Part 2: Flame Test

In part 2 of this experiment, your group (*no more than 3 people*) will observe the characteristic emission colors of several elements. When a solution of a metal cation evaporates in a flame, the cation quickly converts back to a neutral atom because of the electron rich environment of the flame. The heat in the flame is sufficient to excite the electrons in the atoms, which results in the emission of the distinct lines that are specific to the element. Your eye observes the combination of these lines as the characteristic color of the element.

When colors mix, new colors form:

Figure 4:



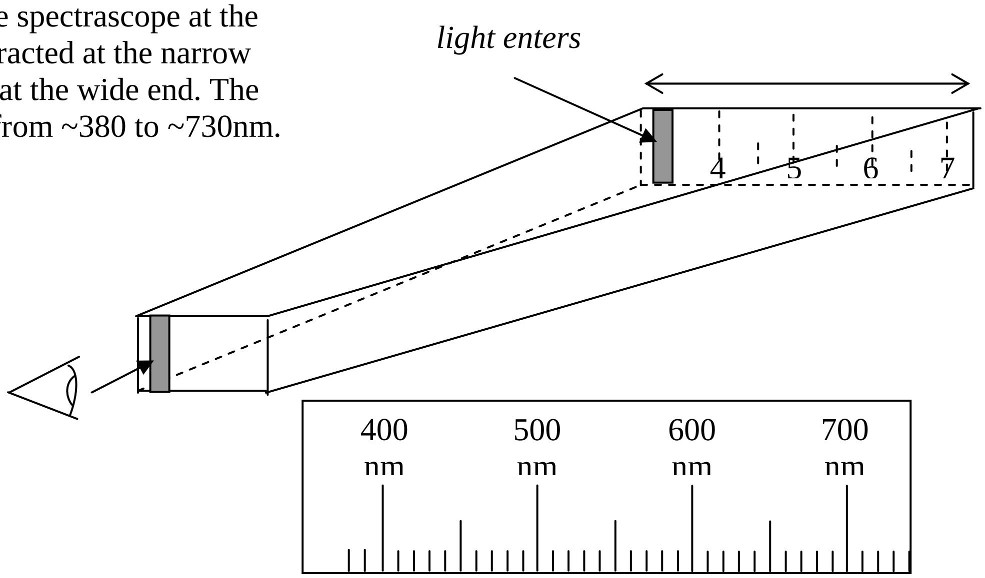
Red and yellow make orange.
Yellow and blue make green.
Blue and red make purple and so on…

By noting the color emitted by each element, one can identify an unknown element if needed.

Procedure: **Part 1: Atomic Spectra** (*You need not wear goggles for this part*.)

Obtain a spectroscope form your instructor; be careful not to touch the viewing aperture at the smaller of the two ends. You will record your observations in the provided date sheets.

The light enters the spectrascope at the wide end. It is diffracted at the narrow end onto the scale at the wide end. The scale covers light from ~380 to ~730nm. (4, 5, 6, 7)



Observation of various sources of light.

(A) Look out the window with the spectroscope and record your observations. If it is a cloudy day, you may need to quickly go out side and look towards the sun, but not directly at it. What do you observe?

(B) Look at the fluorescent lights in the room with the spectrometer. Record your observations in the space provided on the data sheet.

(C) Look at the discharge lamps in the room (H_2, He, Ne and Hg). Sketch the lines that appear brightest. (For some sources there may be too many, just sketch the strongest.)

SAFETY: DO NOT touch the leads of the discharge lamps, as you will receive quite a jolt! Do not look directly into the sun as it may damage your eyes.

Procedure: **Part 2: Flame test** (*You **MUST** wear goggles for this part*.)

1. Place approximately 0.5 mL (10 drops) of the following solutions into your smallest test tubes: Label them first with a wax pencil.

   LiCl, $CuCl_2$, KCl, BaCl, NaCl, $SrCl_2$, $CaCl_2$

2. Fill a 250ml beaker half–way with deionized water.
3. Place ~ 25 mL of 6M HCl into a small beaker.
4. Obtain a ni-chrome wire with a cork holder.
5. Carefully light the Bunsen burner.
6. Dip the metal end of your ni-chrome wire first into in the 6M HCl solution, then into the DI – water and heat the wire in the hottest part of the flame. (blue part) Repeat twice. This will clean off your wire so that any metal contaminates present on the wire are removed. You will need to repeat this step before examining each different solution.
7. You are now ready to perform the flame tests on the known solutions. PERFORM THE FLAME TEST FOR THE SODIUM SOLUTION (NaCI) LAST. It will contaminate all of your results if you don't. Dip the wire into your first solution and place it in the hottest part of the Bunsen burner flame. Note the color of the flame and record your observations on your data sheet. You should perform the test a few times for each solution.
8. Clean your wire with the 6M HCl and DI–water as described in step 6. Repeat the tests on the remaining solutions as described in step 7. Record your observations on the data sheet. Make sure to clean the wire between solutions.
9. Dispose of all waste in the jugs provided in the hood. If the container is full, ***let your instructor know immediately***!
10. Make sure that you turn the gas to Bunsen burner off when you are finished.

Safety: The 6M HCl you will work with in today's experiment is concentrated enough to cause serious skin /eye burns. You should be wearing safety goggles at **ALL** times and be careful not to spill the HCL or any of the solutions on your skin or clothing. If an accident occurs and you spill or get solution on you skin or in your eyes, immediately rinse with plenty of water and notify your lab instructor. Be careful lighting and working around the open flame of the Bunsen burner.