AP Chemistry **Name:____________________________**

Experiment #3: Atomic Spectra & Electronic Structure

Background Information:

About 300 years ago, Sir Isaac Newton saw a beam of sunlight through a glass prism. He discovered that light is made up of a spectrum of seven distinct visible colors. This spectrum of colors always appears in the same order. You can see this color spectrum (Red, Orange, Yellow, Green, Blue, Indigo, Violet and all the colors in between) when you look through a diffraction grating. There are two color ranges that are not visible to our eyes in this spectrum: below red is infrared and above violet is ultraviolet. In a rainbow after a rainstorm this same color spectrum appears in the same order. Rainbows are created when sunlight passes through raindrops that act as millions of tiny prisms.

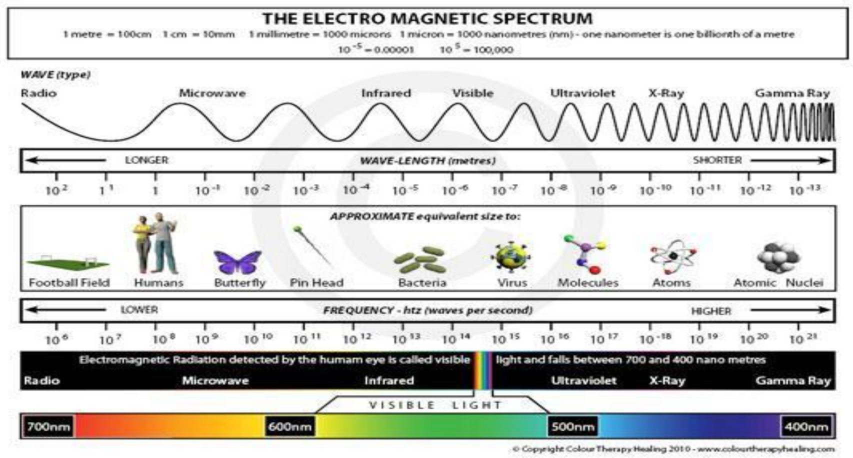
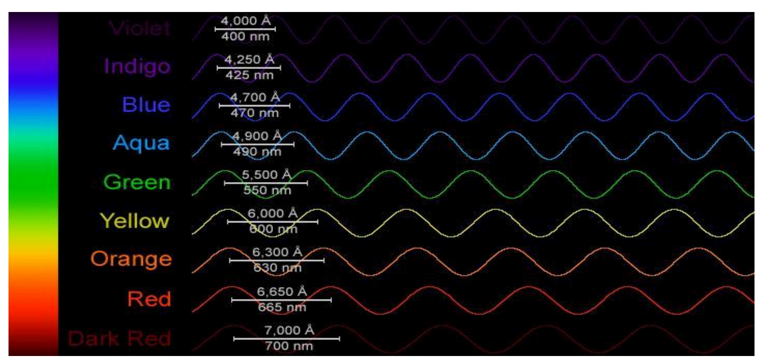


Figure 1. This lab will focus on the visible portion of the spectrum.

The color of a solid object depends on the colors of light that it reflects. A red object looks red because it reflects red light and absorbs all other colors. A blue object looks blue because it reflects blue light and absorbs all other colors. A white object reflects all colors of light equally and appears white. A black object absorbs all colors and reflects no visible light and appears black. Just like when you color with too many colors in one area with crayons or markers, all colors are absorbed, none are reflected and it appears black!

Explanation of visible light at the electronic level:

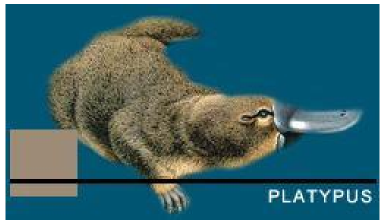
What do fireworks, lasers, and neon signs have in common? In each case, we see the brilliant colors because the atoms and molecules are emitting energy in the form of visible light. The chemistry of an element strongly depends on the arrangement of the electrons. Electrons in an atom are normally found in the lowest energy level called the ground state. However, they can be "excited" to a higher energy level if given the right amount of energy, usually in the form of heat or electricity. Once the electron is excited to a higher energy level, it quickly loses the energy and "relaxes" back to a more stable, lower energy level. If the energy released is the same amount as the energy that makes up visible light, the element produces a color. The visible spectrum, showing the wavelengths corresponding to each color, is shown to the right. **Note: 1 Å = 0.1 nm**



Is light a particle or a wave?

Is light composed of waves or of particles? If light is waves, then one can always reduce the amount of light by making the waves weaker, while if light is particles, there is a minimum amount of light you can have - a single ``particle'' of light. In 1905, Einstein found the answer: Light is both! In some situations it behaves like waves, while in others it behaves like particles.

This may seem odd. How can light act like both a wave and a particle at the same time? Consider a duck-billed platypus. It has some duck-like properties and some beaver-like properties, but it is neither. Similarly, light has some wavelike properties and some particle-like properties, but it is neither a pure wave nor a pure particle.



A wave of light has a wavelength, defined as the distance from one crest of the wave to the next, and written using the symbol λ. The wavelengths of visible light are quite small: between 400 mm and 650 nm, where 1 nm = 10^{-9} m is a ``nanometer'' - one billionth of a meter. Red light has long wavelengths, while blue light has short wavelengths.

A particle of light, known as a photon, has an energy *E*. The energy of a single photon of visible light is tiny, barely enough to disturb one atom; we use units of "electron-volts", abbreviated as eV, to measure the energy of photons. Photons of red light have low energies, while photons of blue light have high energies.

The energy *E* of a photon is proportional to the wave frequency *v*,

$$E = h\,v$$

where the constant of proportionality h is the **Planck's Constant**, h = 6.626 x 10^{-34} J s.

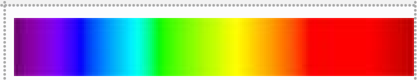
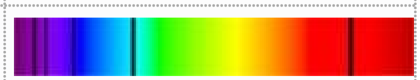
Also, the relationship between frequency and wavelength can be defined as:

$$v = \frac{c}{\lambda}$$

where c is the speed of light (2.998×10^{8} metres per second).

So photons still have a wavelength. A famous result of quantum mechanics is that the wavelength relates to the energy of the photon. The longer the wavelength, the smaller the energy. For instance, ultraviolet photons have shorter wavelengths than visible photons, and thus more energy. This is why they can give you sunburn, while ordinary light cannot.

Spectra can be simplified to one of three basic types. Simple examples in the visible wavebands are shown below:

Type of Spectrum	**Photographic example**
Continuous (or continuum)	
Absorption (dark line on colored background)	
Emission (bright line on a dark background)	

One means by which a continuous spectrum can be produced is by thermal emission from a black body. This is particularly relevant in astronomy. Astronomical spectra can be combination of absorption and emission lines on a continuous background spectrum.

The important thing to know about absorption and emission lines is that every atom of a particular element will have the same pattern of lines all the time. And the spacing of the lines is the same in both absorption and emission, only emission lines are added to the continuum, while absorption lines are subtracted. The spectral lines come from electrons jumping between allowed energy states. This jumping is either electrons absorbing a photon and increasing in energy or emitting a photon and decreasing in energy. Because there are many allowed energy states above the ones that are filled in an element's ground state electron configuration there are multiple spectral lines in a given element.

In **Part One**, wooden splints dipped in solutions of metal salts are heated using a Bunsen burner, producing different colored flames. By comparing the color given off by an unknown with the known metal salts, the identity of the metal salt can be determined. In **Part Two**, electricity is passed through the gas discharge tube resulting in different colors. When the light is passed through a diffraction grating, it is broken into components, producing a line spectrum of the element. Each gas will produce a different line spectrum that will be observed and recorded.

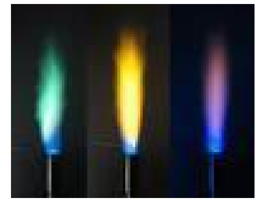
Pre-lab Questions:

1. When scientists first detected the spectrum for hydrogen in the mid-1800s, it contained only four lines in the visible portion of the electromagnetic radiation spectrum. Calculate the frequency and energy associated with each line. (Draw the table in your pre-lab. **Show your work for one row** of calculations.)

Color	Wavelength (nm)	Wavelength (m)	Frequency (Hz)	Energy (J)
red	656 nm			
blue-green	486 nm			
blue	434 nm			
violet	410 nm			

2. A photon of light has a wavelength of 400 nm.
    a. Calculate the energy of this photon, in J.
    b. Your answer from part (a) represents the energy of a single photon. Convert this energy to kJ/mol.
    c. The ionization energy of sodium is 496 kJ/mol. Would light with a wavelength of 400 nm be sufficient to cause sodium to lose its electrons? Explain why.
    d. Assuming you said “no” in part (c), calculate the minimum wavelength of light (in units of nm) that would be needed to cause a sodium atom to lose an electron.

Part One: Flame Tests



When solutions of metals are heated in a Bunsen burner flame, they give off characteristic colors. For example, sodium makes the flame turn bright yellow – this is the same yellow color made by sodium street lamps and many fireworks.

Materials:

- Bunsen burner
- Wooden splints
- Solutions of the following metal salts
  - lithium chloride
  - barium chloride
  - strontium chloride
  - calcium chloride
  - copper(II)chloride
  - sodium chloride
  - potassium chloride
  - sodium chloride/potassium chloride mixture
  - an unknown metal chloride solution

Procedure:

1. Fill a 250 mL beaker with 150 mL of tap water to dispose of matches and wooden splints.

2. Light the Bunsen burner and open the air vent to obtain a non-luminous flame with two blue cones. Be sure to avoid a yellow flame. (Why might this be important?)

3. Obtain a wooden splint that has been soaking in the metal salt solutions (Think: Why is soaking the splints important?).

4. Carefully place the end of the wooden splint that was soaked in the metal salt solution at the top of the inner blue cone. Record the color and intensity (bright/faint) of the flame in the data table. The color given off by the salt is the initial color observed, not the yellow-orange color produced by the burning wood. (To avoid burning the wood, wave the splint through the flame rather than holding it right in the flame).

5. Repeat with the other salts and the unknown. Be sure to record the colors as precisely as possible. For the sodium potassium mixture, observe the colors as before.

6. If more observations are needed, dip the clean end of the wooden splint in the solutions for a few minutes and repeat. Otherwise, discard the wooden splints at the end of the experiment.

7. Clean up your lab table.

Data Table:

(Copy and complete in your notebook. Don't forget to title & label each data table!)

Metal found in the salt	Metallic Ion	Flame Color and Intensity
Lithium	Example: Li^{+}	
Barium		
Strontium		
Calcium		
Copper		
Sodium		
Potassium		
Sodium and Potassium		
Unknown		

Part Two: Atomic Spectra

During this experiment you will make use of a device known as a spectrometer or a spectroscope. A diffraction grating splits a single beam of light into the many colors which it contains. The colors should be ordered like a rainbow, from blue on one side to red on the other. The brightness of the image in each color shows you how much light the object is emitting of that color.

We will be using these diffraction gratings to look at "spectral tubes". These are tubes which contain gases composed of different elements. By plugging these tubes into the wall, we can send electricity through them, which adds energy to the gas. This causes the atoms to become "excited", which is how we describe atoms whose electrons have been raised into high energy levels. After a short amount of time the electrons drop back into lower energy states, releasing a photon to carry off the extra energy. Depending on the number of transitions in each atom and the energy levels in it, photons of different wavelengths and thus different colors are released from each gas.

Each element has its own unique line spectrum and is thus referred to as the "fingerprint" for a particular element. The spectra for each element are unique because each element contains differing numbers of electrons and thus different energy levels.

Think about the following questions as you look at the spectral lines corresponding to each gas:

- Are the lines closely packed, or spread out over many different colors?
- Are there many lines you can see, or only a few?
- How do the colors of the lines from each tube relate to the color you see from each tube when you don't look through the gratings?

Materials:

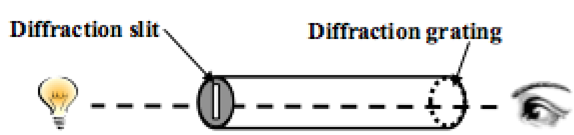
- gas discharge tube
- gas filled tubes (hydrogen, helium, neon, argon, mercury, oxygen and chlorine)
- spectroscope

Procedure:

1. Go to one of the lab stations where a gas discharge tube is set up.

2. Turn on the electricity (**for no more than 30 seconds**) and observe the color given off by the gas. Record your observations in the data table below. If more time for observation is needed, turn off the tube for ~30 seconds and then turn on for another 30 seconds. Do not leave the tube continuously turned on; otherwise, it will burn out the tube. Just remember: "30 on 30 off".

3. Hold the spectroscope so the diffraction slit is towards the light source and the diffraction grating (window) is towards you. (See diagram below).



4. Look through the diffraction grating and observe and record the colored lines (spectrum) produced by hydrogen gas. Complete the data table below and then draw the color lines in the spectrum boxes that follow.

5. Go to the rest of the lab stations and repeat for each element.

Data Table:
(copy into your lab notebook and complete during lab)

Line spectra of elements:

Element/ Compound	Color of the discharge tube	Number of different colored lines in the spectrum

Post Lab Questions:

Part 1:

The electromagnetic spectrum is shown at the beginning of this document. Recall that energy is proportional to frequency, while frequency is inversely proportional to wavelength. Use this information to answer questions 1-4 below.

1. List the colors observed in this lab from the highest energy to the lowest energy.
2. List the colors observed in this lab from the highest frequency to the lowest frequency.
3. List the colors observed in this lab from the shortest wavelength to the longest wavelength.
4. What is the relationship between energy, frequency, and wavelength?
5. Based on the results of your experiment, what metal(s) is found in the unknown? Explain.
6. What inaccuracies may be involved in using flame tests for identification purposes?
7. How are electrons "excited" in this part of the experiment? What does it mean the electrons are "excited"?
8. What particles are found in the chemicals that may be responsible for the production of colored light?
9. Why do different chemicals emit different colors of light?
10. Why do you think the chemicals have to be heated in the flame first before the colored light is emitted?
11. Colorful light emissions are applicable to everyday life. Where else have you observed colorful light emissions? Are these light emission applications related? Explain.

Part 2:

1. Each line in the emission spectrum of the hydrogen corresponds to an electromagnetic radiation with a specific wavelength. Match the 4 observed colors (red, green, blue, and purple) with the following wavelengths: 410 nm, 434 nm, 486 nm, and 656 nm. (Use the electromagnetic spectrum to help if needed.)
2. How are electrons "excited" in this part of the lab? What happens when the electrons relax?

3. What do the different colors in a line spectrum represent? Why are the spectra for each element unique?

4. It has been calculated that the observed colors in a hydrogen atom correspond to the relaxation of the electron from a higher energy level to the second energy level. Which color corresponds to 6→2 transition? 5→ 2 transition? 4→2 transition? 3→2 transition? Explain your reasoning.

5. Each element has its own unique line emission spectrum, just like fingerprints. Explain how this technique can be used to determine the elemental composition of stars.

6. What factors do you think might have contributed to any differences in observations? (ie What is a possible source of error for this part of the lab?)

Detailed view of the emission spectra observed below:
(This can be completed and stapled to the back of your lab report)

Sample Name: ___________________________

Sample Name: ___________________________

Sample Name: ___________________________

Sample Name: ___________________________

Sample Name: ___________________________

Sample Name: ___________________________

Sample Name: ___________________________

Experiment 3 Rubric:
Flame Tests & Atomic Spectra

Name: ____ Period:_______________

Item	Points/Out of
Formatting	/2
Name & Partner	
Title of Experiment	
Date Experiment Started	
Signature and Date turned in	
Use of a Single Line to Correct Errors (-1)	
Pre-Lab	/3
Appropriate Purpose	
Clear & Complete Procedure (Part 1 and Part 2)	
Data and Observations	/11
Complete Part 1 Data Table (2)	
Complete Part 2 Data Table (2)	
Detailed view of Emission Spectra (just attach the lab page) (7)	
Post Lab Questions	/17
Part 1 (11)	
Part 2 (6)	
Conclusion	/2
TOTAL	/35

No need for a separate Data Analysis section because it is already included with post lab questions on this lab.