

SPECTRAL ANALYSIS

LAB ELEC

From World of Chemistry, L. Fruen, 1994

PURPOSE

To observe the visible emission spectrum of several elements and simple compounds and relate it to their electronic configuration.

To identify an unknown substance by observing its spectrum.

To understand the relationship between various primary and secondary colors.

INTRODUCTION

This activity will serve as an introduction to qualitative analysis by using the light emitted by substance when its electrons absorb energy and move to higher energy levels. As the electrons return to the ground state they emit light equal to the energy absorbed. Since every substance will have a specific and unique electron configuration they will also have a specific and unique pattern of light emitted.

EQUIPMENT/MATERIALS

Various Spectrum Tubes Color Triangle Chart

Power Supply for Spectrum Tubes Spectroscope

Electromagnetic Spectrum Chart Full Color Spectrum Chart

SAFETY

The power supplies deliver 5000 VAC to the spectrum tubes. ***DO NOT TOUCH THE TERMINALS WITH THE POWER ON!*** There are no waste disposal issues with this activity.

PROCEDURES

1. Observe and record the visible line spectra of several substances by looking through a spectroscope at the spectrum tube in energized power supply.
2. Determine the energy of the photon associated with each line
3. Observe the spectrum of an unknown substance obtained from your instructor.
4. Record the number of the unknown and the color of the visible lines
5. Determine the identity of the substance by comparing it to the full color spectrum chart.
6. Discuss the color triangle chart and how it relates to the colors we see in the spectrum and in general.

7. Be sure to include in your data table the following information:
 - a. name of substance
 - b. color of visible lines from left to right
 - c. wavelength (λ) of each color
 - d. energy (joules) of each photon
 - e. electron configuration of substance
 - f. at bottom of table list the unknown number and your identification of the unknown

QUESTIONS

1. Why are some areas of the spectrum dark (black)?

2. What is the complementary color of Blue?

3. What is meant by “subtractive primary colors”?

4. The color white means that which wavelengths are being absorbed?

5. What is the identity of your unknown?

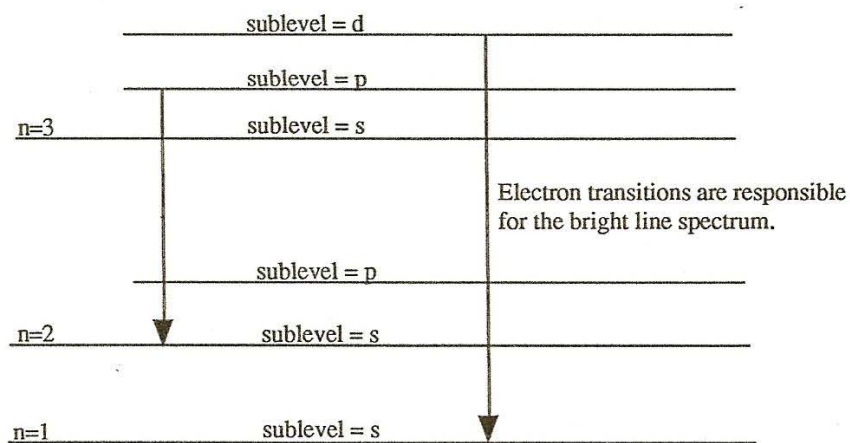
6. Which color (λ) contains the most energy?

EXTENSIONS

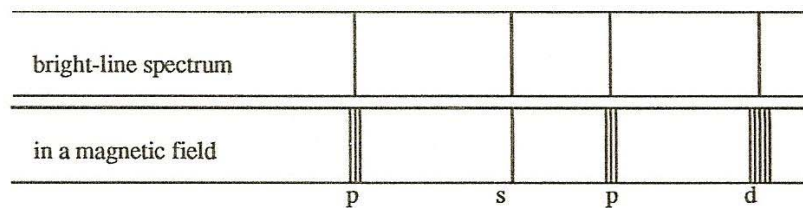
Quantum-Mechanical Model of the Atom

The Bohr model of the atom cannot account for the number of visible spectral lines in the spectra of elements which have more than one electron. For this reason, Bohr's model of the atom was replaced by the Quantum-mechanical Model of the atom.

Atoms contain energy sublevels as well as principal energy levels. The number of sublevels in each principal energy level depends on the number of the principal energy level. Principal energy level $n=1$ contains 1 energy level. Principal energy level $n=2$ contains 2 sublevels, and $n=3$ contains 3 sublevels and so on. These sublevels are designated by the letters s, p, d, and f. The relative energies of these sublevels within a principal energy level increase in order: $s < p < d < f$. The distance between principal energy levels as well as the distance between sub-energy levels varies from atom to atom, accounting for the different spectra.



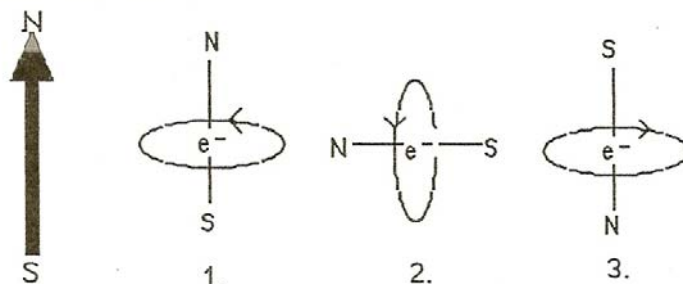
When exposed to a strong magnetic field, some spectral lines split into more lines. This is called the Zeeman Effect. Electron transitions from s sublevels show no splitting in a magnetic field, electron transitions from p sublevels split into 3 spectral lines, electron transitions from d sublevels split into 5 spectral lines, and electron transitions from f sublevels split into 7 spectral lines.



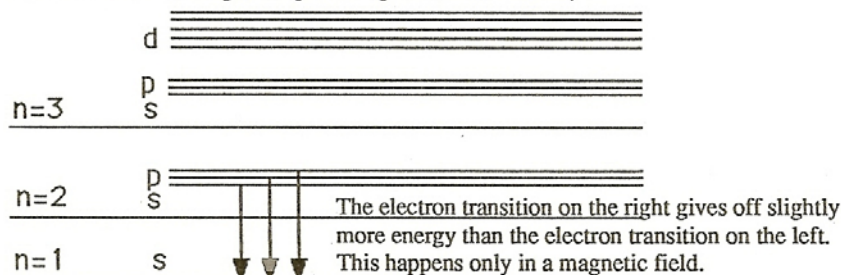
This splitting shows that there are sub-sublevels (called orbitals) within the sublevels. The number of orbitals per sub-energy level depends on the principal energy level and can be calculated by n^2 where n equals the principal energy level number. For example: The second principal energy level, $n=2$, contains $n^2 = 2^2 = 4$ orbitals.

The reason that the s sublevels do not split in a magnetic field is that s sublevels are spherical and so are not affected by a magnetic field, whereas the p, d, f sublevels are not spherical. Spinning electrons are like little magnets. So when confronted with another magnet, some electrons in p, d, and f orbitals have to adjust in order to stay in the orbital. This accounts for the slight difference in energy of the spectral lines in a magnetic field. In the diagram on the next page, electron number 1 has less potential energy in the magnetic field (represented by the bold arrow) because less work is required to rotate the electron away from

its normal position. Electron number 2 will not be affected by this magnet. Electron number 3 has more potential energy in the magnet field because work must be done to rotate it from its normal position (*College Physics Franklin Miller*).



Suppose each one of these electrons is in a different p orbital. The three p orbitals are oriented three-dimensionally on X, Y, Z axes respectively. When a magnet is brought up to the gas discharge tube, the energy of each electron depends on which way the electron is aligned with the magnet. One will have a bit less energy, one will not be affected, and one will have a bit more energy. The diagram shows the possible transitions these electrons could make from the $n=2$ p orbitals into the $n=1$ s orbital. Each transition would give off slightly different energy, which accounts for the splitting of the spectral lines. The orbitals still exist even when the magnetic field is taken away, but the energy given off by electrons making transitions from like orbitals is the same, and so the spectral lines show no splitting. (The diagram below is somewhat misleading because three electrons cannot be pulled into the $n=1$ s orbital at the same time. When a magnet is held up to a gas discharge tube, the excited electrons make all possible transitions, and the splitting of the spectrum can be seen. This diagram represents possible transitions.)



The **Pauli Exclusion Principal** says that only two electrons spinning in opposite directions can occupy the same orbital at a time. This is because there is a very slight attraction between oppositely spinning electrons because the spinning creates a slight magnetic field, but keep in mind the electrons still repel each other because of their negative charges. Each orbitals can hold a **maximum** of 2 electrons.

Principal Energy Level (n)	Sublevel	No. of Orbitals= (n^2)	Maximum no. of e ⁻ /Energy Level = $(2n^2)$
1	s	1	2
2	s	1	2
	ppp	3	6
3	s	1	2
	ppp	3	6
	dddd	5	10
4	s	1	2
	ppp	3	6
	dddd	5	10
	ffff	7	14

There are interesting applications of electrons transitions. Highway flares and sodium lights produce light by electron transitions. The red color of a highway flare is due to electrons transitions in strontium atoms. The energy-saving street lights that give off yellowish light are sodium lamps. The electrons in the sodium atoms are excited by electricity and drop back to lower energy levels in the sodium atom giving off visible light. If you look at highway flare or a sodium lamp using a diffraction grating, you will see the spectrum produced by electron transitions (Bert Rohrer, *Minneapolis Star Tribune*).

Black lights cause posters which have been painted with special chemicals to fluoresce. The chemicals in the paint absorb ultraviolet radiation from the black lights and then radiate back visible light (*The Flying Circus of Physics With Answers* 5.113). Some chemicals have energy levels with "long lifetimes," where electrons linger longer than they would at another level. Normal time for an electron to stay excited is 1×10^{-3} seconds. This "slow" return of electrons in these chemicals results in phosphorescence (*College Physics* by Franklin Miller).

Fluorescent lights use the principle of fluorescence to produce visible light. A common fluorescent bulb contains a very small amount of mercury, which absorbs electricity and gives off ultraviolet light with a wavelength of 365-366 nm. The UV light strikes a thin coating of fluorescent paint on the inside of the bulb, and the paint fluoresces. Overall, about 30% of the electricity is converted to visible light. This compares to 5% for incandescent bulbs, so fluorescent lights save energy and money (*Foundations of Chemistry* 636). A new compact fluorescent light bulb that is now being marketed to replace the incandescent bulb lasts ten times longer and does not get nearly so hot. It is more than 4.5 times more efficient and so saves more energy. These bulbs have been painted with a newly developed fluorescent paint which gives a softer light than a normal fluorescent bulb.

Pigment dyes are made from chemicals that absorb certain wavelengths of white light. The pigment color is due to wavelength(s) of white light that is (are) not absorbed by the dyes. Many of the transition metals have d orbitals into which electrons can make transitions by absorbing certain wavelengths of visible light. Therefore, many transition metal ion compounds are colored while non-transition metals and nonmetals are not. Cobalt, copper, nickel, chromium, and iron all make highly colored compounds that are used for pigment dyes. You may have painted with cobalt blue or chromium yellow.

Pigment dyes absorb wavelengths of visible light and emit infrared. For example, black absorbs all colors of the visible spectrum and radiates the energy as infrared (heat). A recent Harvard study in the Middle East found that the surface of an outfit made with black cloth is about 10°F hotter than one of white cloth because the black cloth absorbs two and one-half times more energy from the sun than white. This explains why white clothes are popular in hot, sunny weather.

Self Test

1. How many s orbitals are there in each principal energy level?
a) 1 b) 2 c) 3 d) 4 e) It depends on the atom
2. How many orbitals are there in the n=6 principal energy level?
a) 2 b) 6 c) 12 d) 36 e) 72
3. How many electrons can the n=6 principal energy level hold maximum?
a) 2 b) 8 c) 18 d) 36 e) 72
4. How many p orbitals are there in each principal energy level except n=1?
a) 1 b) 2 c) 3 d) 4 e) 5 f) 7
5. How many orbitals are there in the n=2 principal energy level?
a) 2 b) 4 c) 9 d) 16 e) 25
6. How many electrons can the n=3 principal energy level hold maximum?
a) 2 b) 8 c) 18 d) 32 e) 72

7. If a principal energy level contains 1s, 3p, 5d, and 7f orbitals, which energy level is it?
a) $n=1$ b) $n=2$ c) $n=3$ d) $n=4$ e) $n=5$ f) $n=6$
8. Which of the following emits a visible spectrum caused by electron transitions
a) incandescent bulbs b) sodium lamps c) florescent bulbs d) highway flares
9. Red pigment
a) absorbs red light.
b) transmits red light.
c) absorbs all white light but red.
d) transmits all white light but red.
10. Black paint is used on solar collectors because
a) black paint absorbs all white light and transmits infrared.
b) black paint transmits all white light.
c) black paint absorbs black light.
d) black paint transmits black light.
11. Fluorescence is caused by
a) electrons which absorb infrared energy and give off visible light.
b) electrons which absorb visible light and give off infrared radiation.
c) electrons which are excited by ultraviolet radiation and give off even higher energy radiation.
d) electrons which absorb ultraviolet radiation and give off visible light.