How Math fits in Chemistry

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Overview

This unit is meant to be implemented into a high school chemistry course in order to supplement the unit on light and atomic emission spectra. The unit may be helpful for students in grades 10 – 12 in a regular education or inclusion classroom. The atomic emission spectra for each element are unique and can be used as a means of identification. This unit can be tied into the exchange of energy in a chemical reaction, and reaffirm that every element and compound has its own unique set of properties. This will also show that there is an energy transfer when atoms get excited, and show the wave-particle duality of light. This unit is to show the uncanny application of math in science. The objective is to allow students to calculate the energy of an electron during transformation. Students should be able to prove the mathematical equation for this transformation from experimental data showing how math and science really do coincide.

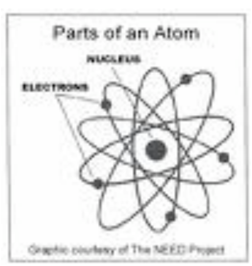
The Learning Research and Development Center at the University of Pittsburgh is currently piloting a chemistry unit on heating and cooling systems. The unit is a design based learning initiative where students learn the design process used by engineers. Students are asked to design a prototype to solve a problem they have related to using a heating or cooling system. The unit allows students to explore both endothermic and exothermic reactions, but does not explore the concepts in depth at a molecular level. In order to supplement the heating and cooling unit, Sally Martin, Rick Rubin, and myself, all of the Pittsburgh Public Schools, created a design based learning unit on special effects for a movie. Some students decided that they were interested in manipulating light or creating lightning. Part of the design based learning initiative is to ask students to be able to describe the chemistry behind the prototype. The special effects unit encourages students to develop a deeper understanding of the change in energy taken on by electrons of various elements. It is a good unit to get the attention of students, and allow them to explore a part of chemistry that they would not have a chance to explore otherwise. The special effects unit requires massive amounts of preparation on the part of the teacher, and impeccable organization, but I was impressed with the results. The unit really focuses on the particulate nature of matter, and how particles interact with each other. It should follow that students would be curious then as a follow-up about how the particles that make up atoms of elements interact with each other. This unit allows students to describe both physically and mathematically what happens inside an atom when there is an energy change.

Rationale

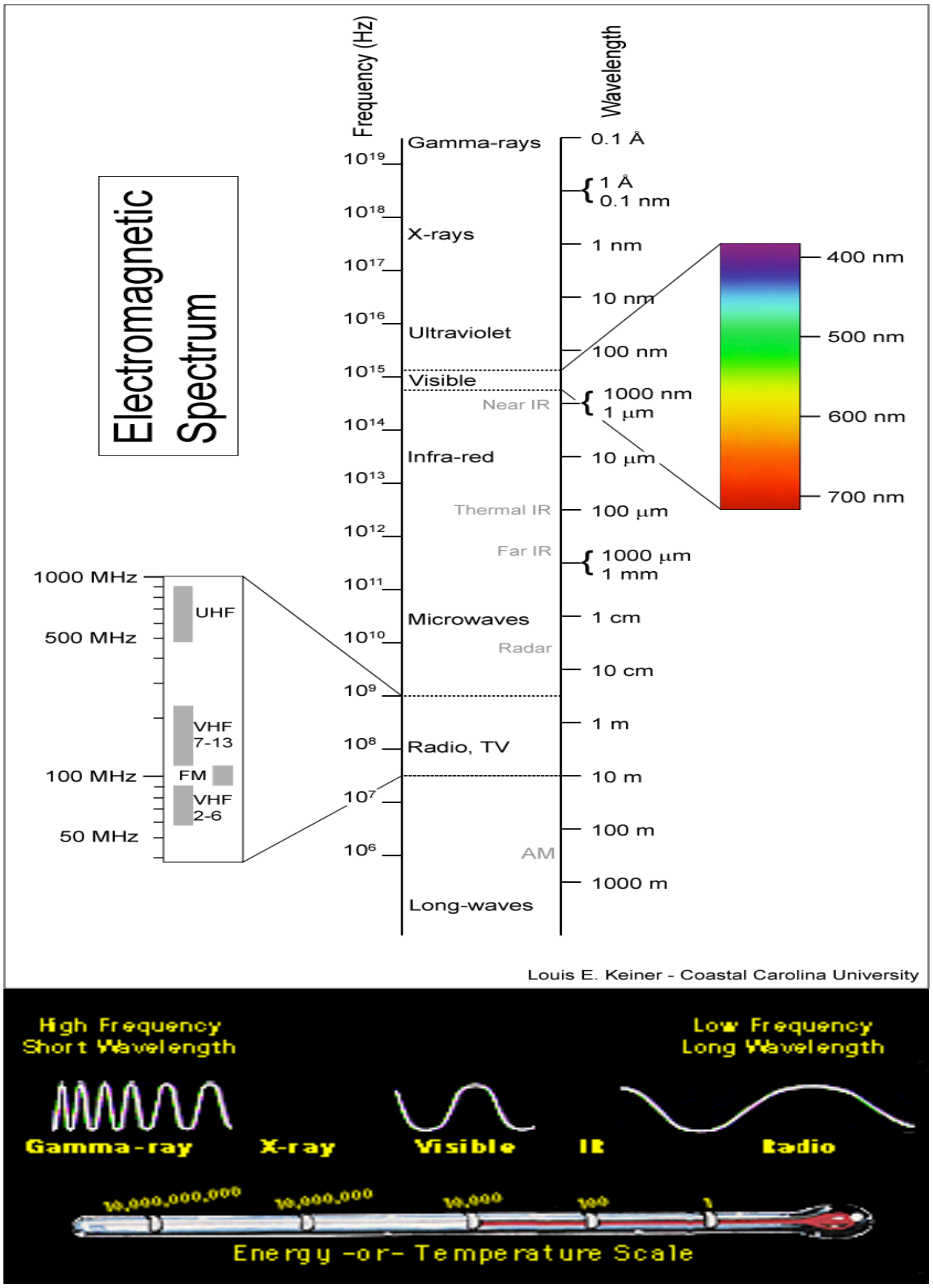
Math and chemistry fit together hand in hand because in order to understand one, we need the other. In some ways it is like the chicken and the egg scenario, which one came first. It can be argued that there is a great creator who used complex math in order to design the universe, or that math was created in order to describe what already was. Regardless of what you personally believe it is inevitable that you come to the conclusion that math and science are so inter-related that they can not be separated. There are courses which use just the theory behind science concepts and skip the math altogether; however, these are usually lower level courses taught to students who do not yet have the mathematical skills necessary for a deep understanding of the content. Students taking chemistry at the Pittsburgh Public Schools have a prerequisite of having passed Algebra I with at least a C or better. For anyone who has taken or taught chemistry it is plain to see that students who do not understand basic algebra will inevitably struggle with the math dependent course. This unit in particular will utilize the algebra skills necessary for the course when students are asked to derive the mathematical equation explaining the experimental data collected in lab from the transition of electrons from an excited state to a ground state.

One of the key concepts to understanding chemistry is being able to understand atoms and the energy transfers and reorganizations they undergo. One way that scientists can use energy transfers to understand chemical changes is to follow the energy transfers caused by photons. Light has a wave-particle duality. The electromagnetic spectrum shows the various types of radiation which travel as waves. They also act as particles, when we think of particles of light we refer to them as photons. When a photon hits another particle there is a transfer of energy. The energy flows from the photon to the particle that it hits. This energy causes the particle that it hits to become excited, which gives the particle a higher energy. In nature stability matters; the lower the energy the more stable the substance. In order for the substance to get rid of the extra energy acquired the energy has to be transferred again. The law of conservation of energy tells us that just like mass, energy cannot be created or destroyed, but it can be turned into another form. This unit is focusing on how light is emitted when electrons transfer back down to the lowest energy state or the ground state.

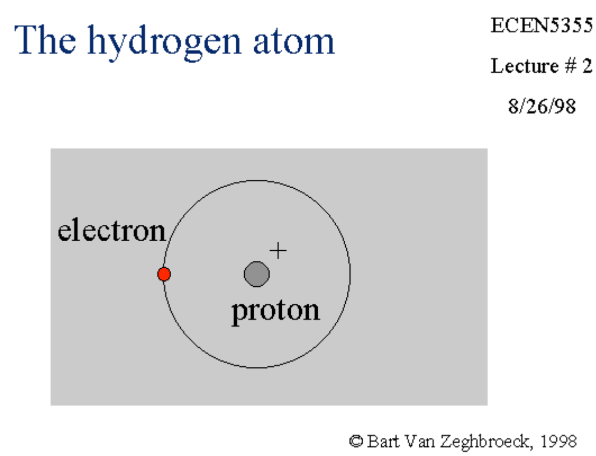
In order to understand the energy transfers around an atom we first need to understand the concept of an atom. Greek philosopher Democritus was the first to say that all matter is composed of tiny indivisible particles called atoms. As time passed and scientists developed better techniques and technologies the theory developed into our current view of atoms. In 1803 John Dalton pictured atoms as tiny, indestructible particles with no internal structure. By 1897 J. J. Thomson discovered the electron using the cathode ray tube. This led to the plum pudding model of the atom, when electrons were embedded in a sphere of positive electrical charges. By 1904 Hantaro Nagaoka suggested that atoms had a central nucleus where electrons moved in orbits like rings around Saturn. Scientists, however, are not satisfied with a theory unless they can back it up, and are sure that they understand the theory completely. It wasn't until 1911 when Ernest Rutherford showed that there was a tiny positively charged nucleus with electrons moving around it using the gold foil experiment. One of the most prominent models of the atom was the Bohr model introduced in 1913. Bohr's model gave electrons a circular orbit which had fixed distances from the nucleus. In 1923 Louis de Broglie proposed that moving particles like electrons have some properties like waves. Within a few years, experimental evidence supported this idea. Just 13 years later Erwin Schrodinger developed the mathematical equations to describe the motion of electrons in atoms. His work led to the electron cloud model. We now know that electrons do not circle the nucleus as if they are on a set track, but that there is an area where the electrons are most likely to be found. By 1932 James Chadwick confirmed the existence of neutrons which have no charge, but do contribute to the mass of an atom. Atomic nuclei contain neutrons and the positively charged protons. The discovery of the neutron was encouraging to scientists because it confirmed the mathematical predictions of what the masses of atoms should have been. The theory of the atom continued to develop further subdividing the particles of the atom into quarks and gluons. However for this unit we do not need to worry about any subatomic particles other than the proton, neutron, and electron.



Spectroscopy is a science that records and analyzes the wavelengths of electromagnetic radiation emitted by samples of materials. Spectroscopists need to use sophisticated machinery and instruments to make their measurements; they have an advanced degree in chemistry with a strong background in mathematics. Spectroscopists observe emission lines from a sample by using detectors and recording its output (PHschool.com, cdb-1053). The data given by spectrometers gives the wavelength and intensity of each emission line. We call the pattern of wavelengths and intensities the emission spectrum of a sample. The more complex an atom is the more complex the spectrum will be.



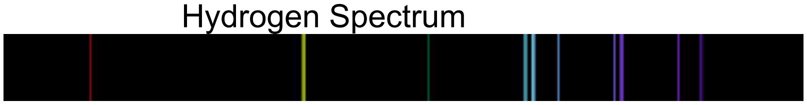
I think it is important to show this, because the only portion of the electromagnetic spectrum students tend to acknowledge is the visible portion. This makes sense because for us it is the easiest to detect, but they should be aware of the other types of radiation. If you take time to discuss the various portions of the electromagnetic spectrum students realize that they have at least heard of uses for or have been exposed to the other parts of the electromagnetic spectrum. Obviously they have all used television and radio waves, but many just have never thought of how they work. If they have ever had a broken bone or gone to the dentist then they have most likely been exposed to X-ray radiation. For anyone who has ever had sunburn they should at least have some understanding about how ultraviolet radiation can affect their lives. Microwave radiation is also a popular form of radiation that they have had experience with. Students tend to be more likely to remember that there are more types of radiation than just visible light which we see when they have to come up with examples of or uses for the radiation on their own. Even after discussing the various types of radiation students still seem to have difficulty with the concept of electromagnetic radiation, and that through simple calculations you can determine the type of radiation when given either the frequency or wavelength.



Since hydrogen is the simplest element on the periodic table it is easiest to work with when it comes to understanding the changes in energy which occur when electrons are excited. Since hydrogen has only one electron, there is no question as to which electron is being excited. In my current chemistry courses I have students use a simple spectrometer to view the atomic emission spectra of various elemental gases. This shows them that each element has a unique spectrum, but does not describe or show the excitement of electrons. Since light travels in waves the spectrometer shows colored bands above the wavelengths represented by the various colors in each spectrum. In addition to using the spectrometers I also have students do a flame test lab, which does show the light given off by photons as they are returning to their ground state. This allows them to calculate the energy change for different metals. Even though the students are capable of completing the calculations by rote they tend not to understand the meaning behind the calculation. It is necessary to discuss several times what the calculations are showing and why they are pertinent.

Niels Bohr, a Danish physicist, made huge contributions to the understanding of atomic structure and quantum mechanics. His model of the hydrogen atom explained why the emission spectrum consists of specific frequencies of light and the values of those frequencies; which were confirmed through experimentation. Bohr showed that the Rutherford atomic model using the application of Planck's quantum theory that light or radiation consists of quanta of energy explain the spectra of elements and also their position in the periodic table (Partington, page 361-362). A simple model of the atom was proposed in 1916 by G. N. Lewis who supposed that the outer layer of electrons is 8 in the atoms of inert gases, and tends to become 8 in other atoms either by gaining or losing outer shell electrons. When atoms gain or lose electrons they become ions. Metal atoms lose outer shell electrons to become cations, nonmetals gain electrons to become anions. Atoms may also share outer shell electrons when forming covalent bonds, depending on how electronegative the atoms are. When atoms get 8 electrons in their outer shell we refer to this as having an octet. In Bohr's theory of the atom electrons are regarded as point-charges revolving around the central positive nucleus of the atom in circular orbits. The energy state of the electron is represented by the principle quantum number, n (Partington page 362-363).

Hydrogen Spectrum

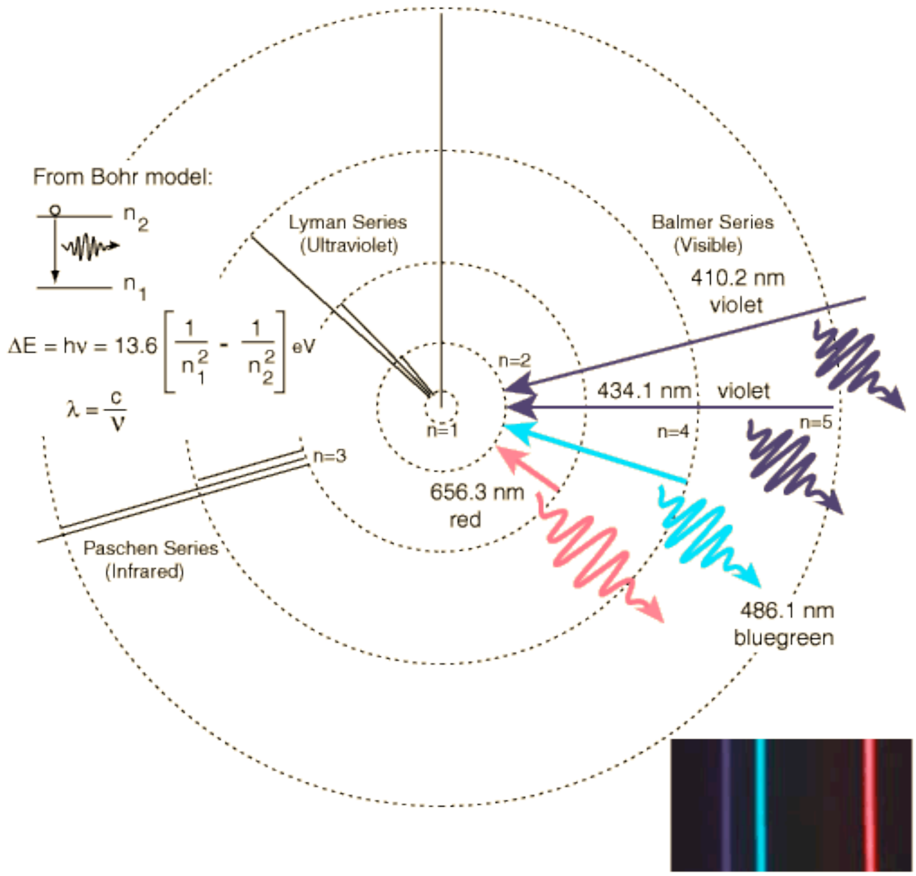


Helium Spectrum

Bohr's model explained the emission spectrum of hydrogen, but not elements with more than one electron. His theory was later replaced by the quantum mechanical model. In Bohr's model the electron in hydrogen can only have certain specific energies, when it is in its ground state, it has the lowest energy. When energy is absorbed by the electron the energy rises to an excited state. We associate the state of the electron with the principal quantum number, which determines the shell or energy level the electron is in. When hydrogen is in its ground state n = 1, but can excite to any of the other energy levels. Once the electron has been excited it naturally wants to fall back down to its ground state. I tell my students that here atoms act like people, they want to use the least amount of energy possible, so they always want to be in the ground state. For hydrogen the electron falls back to its ground state in one step, but for other elements the transition can take multiple steps to get back to its ground state. Bohr already knew from prior work that there is a relationship between the energy of the transition and the frequency of the emitted light. There is a limit to the frequency of emitted light for each set of lines because an electron with enough energy can completely escape an atom. The light that is given off during the fall of the electron is referred to as a photon; this is where the particle nature of light comes into play.

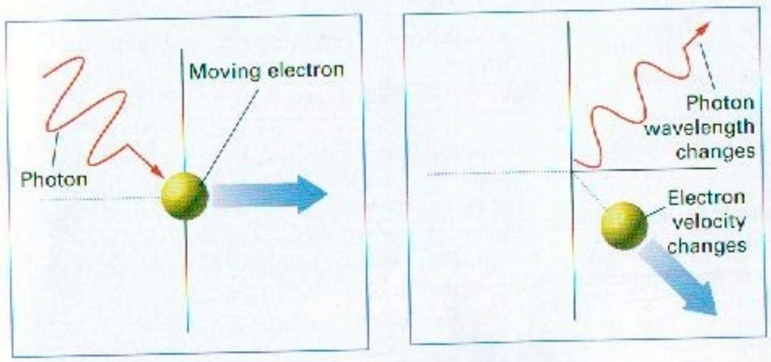
When atoms absorb energy, electrons move into higher energy levels, which we refer to as exciting electrons. These electrons then lose energy by emitting light when they return to lower energy levels. Light emitted by atoms consists of a mixture of only specific frequencies from the visible light emitted. Each specific frequency corresponds to a specific color, so when light passes through a substance the frequencies of light emitted by an element separates into discrete lines to give the atomic emission spectrum of each element.

Robert Wilhelm Bunsen along with Kirchhoff discovered spectrum analysis in 1859. As a teacher Bunsen was very successful; his inclinations were practical, and he took no part in the discussions on theory which raged during the first half of the nineteenth century (Partington 1989, page 236-238). Atomic spectra are part of the chapter on electrons in atoms. The Prentice Hall Chemistry text describes the emission spectrum of each element like fingerprints, just as fingerprints are never duplicated, neither are emission spectra. If you go to the following website you can choose almost any element on the periodic table and see what the emission spectra is: http://jersey.uoregon.edu/vlab/elements/Elements.html. Much of what we know about the universe comes from the study of atomic spectra of stars. In order to use the spectrometers you need a power supply and bulbs with the gas of a specific element inside. You can order both of these from Carolina, each bulb is approximately $30.



We focus on hydrogen because it is the simplest element. The three groups of lines in the hydrogen spectrum correspond to the transition of electrons from the higher energy levels to the lower energy levels. The Lyman series corresponds to the transition to the n = 1 energy level. The Balmer series corresponds to the transitions to the n = 2 energy level. The Paschen series corresponds to the transitions to the n = 3 energy level. These different parts of the spectra are in different portions of the electromagnetic spectrum. The Lyman series is in the ultraviolet portion, the Balmer series is in the visible range, and the Paschen series is part of the infrared section

German physicist Werner Heisenberg, who was a student of Bohr, examined quantum mechanics; which is not addressed in classical mechanics. He came up with the uncertainty principle which states that it is impossible to know exactly both the velocity and the position of a particle at the same time. His findings are particularly important when dealing with small particles such as electrons; but do not apply to macroscopic objects. In order to see an electron you would need to strike it with a photon; however, since electrons are so tiny and have such little mass doing so changes its motion so that it can not be predicted precisely.



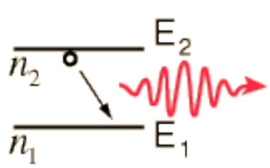
Before collision
A photon strikes an electron during an attempt to observe the electron's position.

After collision
The impact changes the electron's velocity, making it uncertain.

Figure 5.16 The Heisenberg uncertainty principle states that it is impossible to know exactly both the velocity and the position of a particle at the same time.

Section 5.3 Physics and the Quantum Mechanical Model **145**

After the discovery of matter waves the Schrodinger equation was discovered, which describes electrons in atoms. Schrodinger's theory led to the concept of electron orbitals and configurations; including the wavelike motion of matter and the uncertainty principle. This is a great advancement for chemistry, but a little too advanced for high school students. What they can do however is calculate the energy of the transition involving the emission of a photon. If students know what element they are observing, then they can figure out what energy level the photon is dropping to. We know that energy is proportional to the frequency of the radiation. We use the equation E = hυ, where E is the change in energy, h is 6.626 x 10-34 J·s, and υ is the frequency.

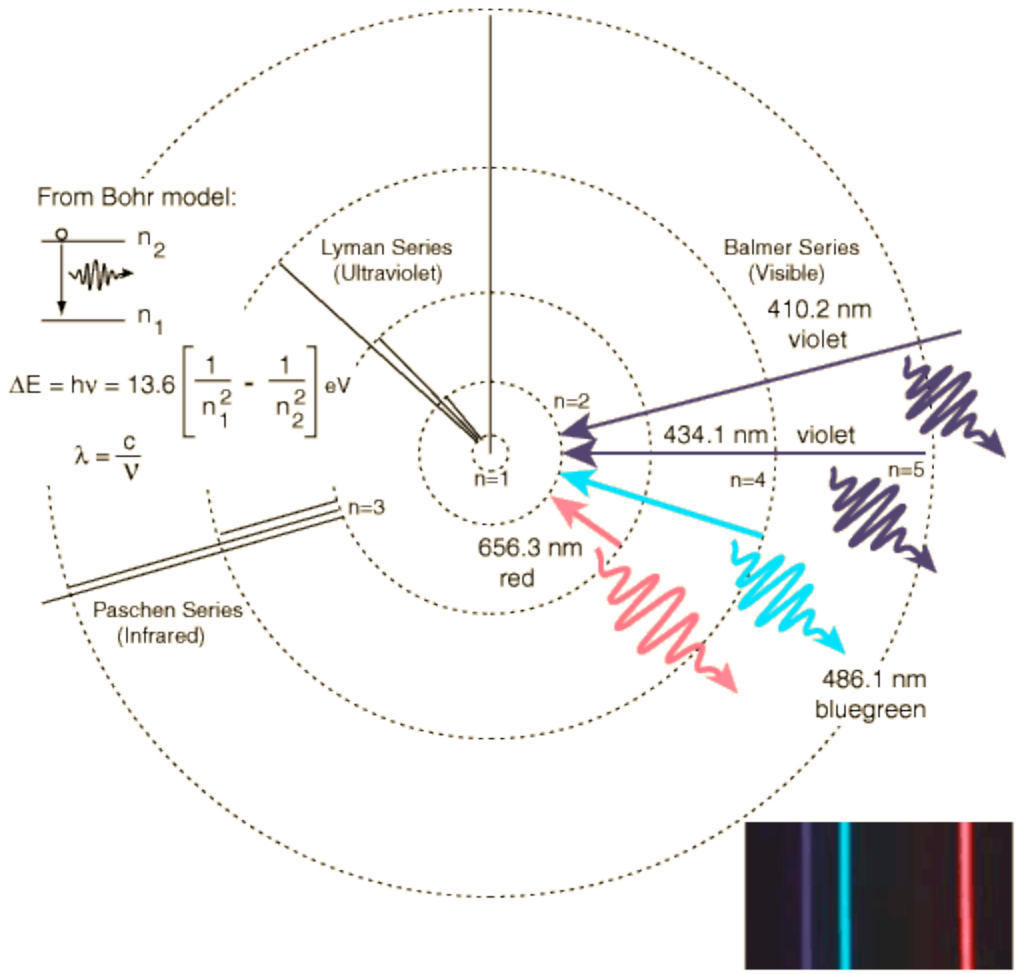


A downward transition involves emission of a photon of energy:

$$E_{photon} = h\upsilon = E_2 - E_1$$

Given the expression for the energies of the hydrogen electron states:

$$h\upsilon = \frac{2\pi^2 me^4}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = -13.6\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \text{eV}$$



The Bohr model picture of the orbits is useful for visualization as long as we remember that this is just the most probably range of values. We can use the above equations to determine which energy levels the electrons come from and go to based on the change in energy. Students can determine this based on the emission spectra they see when looking through the spectrometer.

Measured Hydrogen Spectrum

The measured lines of the Balmer series of hydrogen in the nominal visible region are:

Wavelength (nm)	Relative Intensity	Transition	Color
383.5384	5	9 -> 2	Violet
388.9049	6	8 -> 2	Violet

397.0072	8	7 -> 2	Violet
410.174	15	6 -> 2	Violet
434.047	30	5 -> 2	Violet
486.133	80	4 -> 2	Bluegreen (cyan)
656.272	120	3 -> 2	Red
656.2852	180	3 -> 2	Red

Objectives:

This unit is meant to relate mathematics and chemistry. Students need to realize how they are connected to one another and are used together in nature. Students can understand the differences between solids, liquids, and gases; by understanding phase changes, hopefully this unit will be able to add another level of understanding when it comes to the energy needed for these changes to occur. They can understand and visualize how the molecules within the phases are energized when struck and how the molecules rearrange themselves and then are attracted to other particles as a result of the initial interaction. Students will also be able to see how for different effects different amounts of energy will be needed. By using the math behind the concept they have something concrete to hang onto, something to show the various effects. By using the labs they have the opportunity to see the effect first and then understand it by calculating the energy changed in order for the reaction to happen. Again this takes us back to the chicken and the egg scenario, did we create the math to fit the science or vice versa. To me this proves that it really doesn't matter which one came first, the point is that you cannot truly understand one without the other.

The changes in curriculum over the past decade have headed towards student centered learning, and hands on learning. We are also trying to incorporate more contemporary tropics into the curriculum to help get more students interested in studying science as a post-secondary option. The design unit implemented by the University of Pittsburgh has students design their own experiments and heating and cooling system products. This allows students to take ownership of their learning and apply the background knowledge they get from class. Since it is important for students to develop critical thinking and be able to express their thought processes for standardized tests this unit incorporates these skills as well. Students should further develop their understanding of the energy transfer processes and take that knowledge to the next level by incorporating the scientific theories behind it. I think that once they understand what happens at an atomic level the more macroscopic scientific world makes more sense. They will be able to see more easily how certain chemical reactions require more or less energy, by understanding that it takes energy to break chemical bonds and that energy is given off when new bonds form. The idea of exciting electrons via photons is one which will allow them to understand how these interactions can take place initially.

Strategies:

In order for students to get the most out of this unit they need to take ownership of their learning. If you use an inquiry approach they are more likely to actually remember what they learned. Students should be introduced to the wave-particle duality of light prior to the labs associated with this unit. Allow students to explore the range of spectra using the spectrometers in lab. Then show them examples of the math associated with what they saw. They can then complete the flame test lab and determine the representative wavelengths based on the changes in energy.

Classroom Activities:

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Light: An Historical Perspective

At some time or other, you may have noticed one or more rainbows projected onto a table or wall from the rim of a glass or jelly jar or crystal that was placed in the sun. Where did these colors come from? Why were they always in the same order with red on one side and blue on the other? Only a few hundred years ago, these simple observations were the source of great interest for many scientists. At a time when it was commonly believed that light was made up of tiny particles called corpuscles, Christian Huygens (1629-1695), who had been studying the behavior of water waves, made a conceptual leap and transformed the scientific community by using models that described the behavior of water waves to explain the reflection and refraction of light. Light was not necessarily a particle after all, it behaved like waves of water!

As years passed, other great scientists such as Augustin Fresnel (the Fresnel lens) and Thomas Young (1773-1829) contributed experimental evidence to support the wave theory of light. Young's experiments, particularly those dealing with the diffraction of light conclusively demonstrated that light must travel in waves.

This conceptual model of light worked rather well until it began to fray at the edges at the turn of the century when scientists tried to explain spectra made by burning chemical salts in a flame. You see, these spectra were a little different; they did not stretch continuously from red to blue as the spectra of ordinary daylight does, but rather, had several bright lines and many large dark spaces in between.

The tide would soon turn on the wave theory of light and it would again be viewed as a particle when, in 1921, A.H. Compton demonstrated that light possessed momentum, a very particle like quality.

Today, our current theory of light now embraces both its wave and particle aspects. Light is an electromagnetic wave, that travels in small particle like packets called photons. Each photon travels at the same speed: 3×10^8 m/sec, the speed of light. The energy of a photon is determined by its frequency (color). The higher the frequency, the more energy it contains. The higher the frequency, the bluer the light, the lower the frequency the redder the light. We are all familiar with ROY G. BIV. The letters stand for the colors in the rainbow with Red, Orange, Yellow, Green, Blue, Indigo, and Violet. These colors are listed in order of increasing energy (decreasing wavelength) and comprise our visible spectrum. But the electromagnetic spectrum doesn't stop there! It continues beyond the visible into higher energies with ultra violet, x-rays, and gamma rays. It also extends below red into lower energies with infra red, and radio waves.

Where Does Light Come From?

Electromagnetic radiation (in our specific case, light) is created when an electron moves from a higher energy level (electron orbit) to a lower energy level. The photon of light that is emitted has an energy that corresponds *exactly* to the difference in energy between the two orbits.

Diffraction Gratings:

The heart of the spectroscope is the diffraction grating. This is a thin film of plastic with thousands of very closely spaced lines etched in its surface, in our particular case, 5000 lines per inch. This grating uses a combination of diffraction to bend the light waves as they pass by the lines etched on its surface, and interference to constructively add colors at certain locations and destructively eliminate colors at other locations. What you end up with after a beam of light is passed through a diffraction grating is a separation of the colors (wavelengths) of which the incoming light beam was composed.

Using the Spectroscope:

Hold the spectroscope so the small end with the square hole is toward you. The wider, curved end has a narrow slit (which lets light into the spectroscope) and a wide window with a numbered scale. Look through the eyepiece of your spectroscope (the square hole in the small end of the spectroscope which holds the diffraction grating) and point the slit end at an incandescent light bulb. The numbered scale should be to the left of the slit as you look through the eyepiece. You should find that a continuous spectrum is located on the left side of the bright "white" slit. Be sure the slit is pointed directly at the light source for the best and brightest spectrum. Each number on the scale indicates the wavelength of light in angstroms when multiplied by 1000 (that is, a reading of 5 on the scale is equivalent to 5000 angstroms, 1 angstrom equals 1×10^{-10} meters). Most solids that are heated to glowing (like the filament of a light bulb) will produce a smooth distribution of colors called a continuous spectrum.

Types of Spectra

As you've already seen there are different types of spectra in summary, these are:

Continuous spectra whose source is usually a luminous solid or liquid, such as the glowing filament of a lamp.

Emission spectra or bright line spectra whose source is a glowing gas. This gas emits photons with very specific energies (frequencies, wavelengths) that are characteristic of the chemical element of which the gas is composed.

Absorption Spectra or dark line spectra which is generated by having a continuous spectra (white light) pass through a cooler gas located between the source of the continuous spectrum and the observer. The cooler gas absorbs those wavelengths that it would normally emit if it were the glowing source. The dark line spectra has the same spectral fingerprint that the cooler gas would if it were emitting a bright line spectra. You can think of the absorption spectra as the photographic negative of the bright line spectra.

Chemical fingerprints

Because each element has a different atomic structure, the electrons that orbit the nucleus emit photons of light at frequencies that are very characteristic of that particular atom. One could say that the bright line *emission* or dark line *absorption* spectra is a chemical fingerprint for the particular element in question.

So, given a particular spectra, we should be able to identify each wavelength present and pair them up to match the "fingerprints" of known elements to determine which elements are present in the material producing the spectra! Just like detectives fingerprinting the crime scene to determine who the criminal was!

Try looking at a flourescent light bulb with your spectroscope. Now, instead of a smooth continuous spectrum you see several bright lines. One is violet, one is cyan (light blue-green), one is green, one is yellow, one is orange, and a couple of red lines. This spectra is primarily produced by mercury vapor. Try looking at other types of lamps, such as neon, or sodium vapor lamps. How do these compare? Compare other light sources if you can, such as halogen lamps, gas lanterns, automobile head lamps, LED's etc. Can you tell which lamps use solid filaments by looking at their spectra?

Other bright line spectra can be produced by burning salts of various chemicals in a gas (bunsen burner) flame. Some common elements are:

Copper	5300	Green
Sodium	5890	Yellow
Strontium	6060	Orange
Lithium	6708	Red

Note that sodium is very common and very bright and may be present in many flame spectra as a trace of contamination. These salts can be introduced into a flame by dipping a hot wire into the chemical salt then bringing the salt coated wire to the base of the bunsen burner flame. The flame will begin to burn with a color characteristic of the salt used.

Some elements are contained in low pressure gas tubes called spectrum tubes that are similar to flourescent lights. Higher resolution spectral lines for some of these elements are given below:

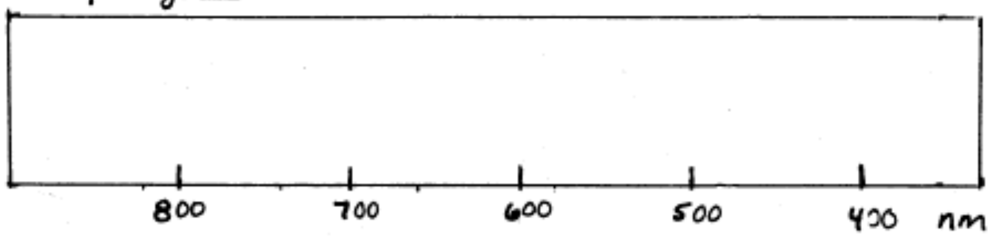
Hydrogen	6562, 4861, 4340
Mercury	4358, 5460, 5769,5790
Helium	4471, 4713, 4921, 5015, 5875, 6678
Sodium	5889, 5895

Note that in many cases closely spaced spectral lines may not be resolvable with this spectroscope and wavelength data is given for reference only.

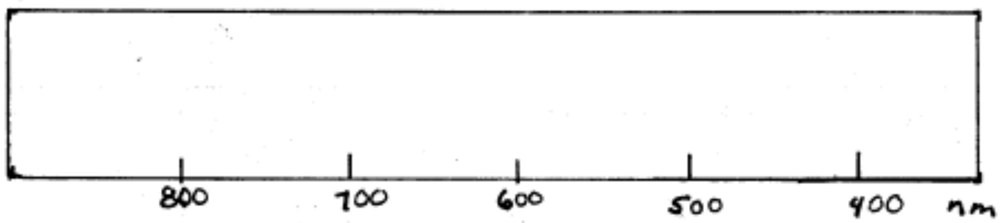
Name ____________

Date ______ Period ____

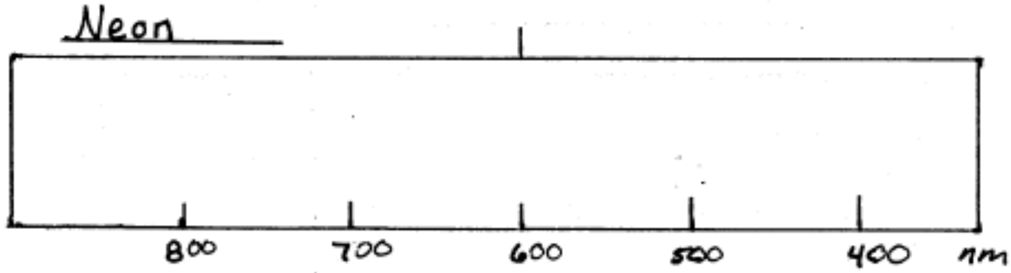
Hydrogen



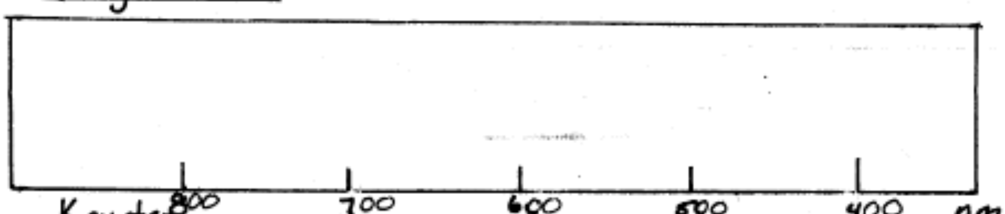
Helium



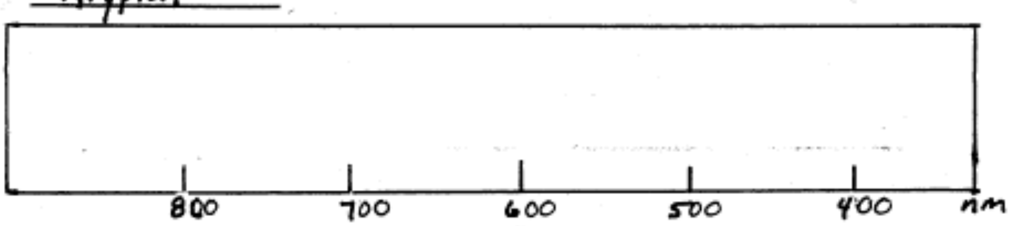
Neon



Argon



Krypton



The Flame Test Student Worksheet

Part I

Procedure

1. Put on lab apron and safety goggles.
2. Add 15 drops of each 0.5M solution to a different clean test tube.
3. To clean the wire, dip it into the test tube of 1M of HCl and heat the wire in the hottest part of the flame until no color shows.
4. When the platinum wire is clean, dip the wire in the test tube containing a 0.5M solution and hold it in the hottest part of the flame. Record your observation of the color of the flame on the data table.
5. Repeat the process of cleaning the platinum wire. Now get ready to test another solution.
6. Test all of the solutions and make sure that you record the color of the flame for each element on the Data Table.
7. Check your flame colors to known results.
8. Fill one clean test tube with 15 drops of one of the 0.5M solutions. The teacher keeps track of what element solution is in this "mystery tube." Repeat the flame test, without telling the students what solution it is. Students must use the information gained from the first part of the experiment to identify the mystery solution.
9. Use the diffraction grating to observe the color of the flame for the following elements: Sodium, Barium, Copper, and Lithium. The students should be able to see the individual lines making up the light from the flame. This can be tricky! In order for it to work, the room will have to be completely dark (in order to block out other light sources) and the students will have to be close to the flame, holding the diffraction grating up to their eyes. It may be necessary to rotate the diffraction grating in order to see the emission lines. Be patient!
10. Record the colors of the elements' emission lines in column three of the Data Table.
11. Before leaving the laboratory, wash your hands thoroughly with soap and water.

Stations	**Observed Flame Color**	**Color of Emission Lines**	**l (m)**	**n (Hz)**	**E (J)**
Calcium (0.5M CaCl)					

Sodium (0.5M NaCl)					
Barium (0.5M BaCl)					
Lithium (0.5M LiCl)					
Copper (0.5M CuCl)					
Cesium (0.5MCsCl)					

Think About

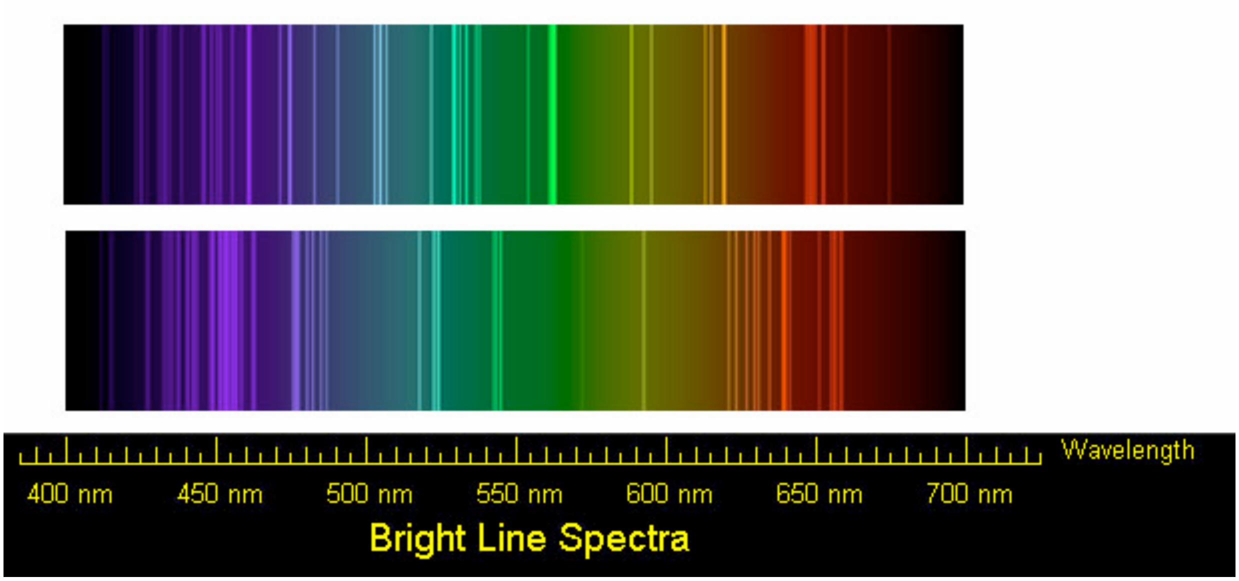
Discuss the following questions in lab groups. Remember you are trying to determine what is taking place during the Flame Test whereby various colors of light are being emitted. One person in your group will have the responsibility of writing the group answers down. After discussing these questions in the group, another person will be responsible for sharing your thoughts with the whole class. You may refer to background material.

- What particles are found in the chemicals that may be responsible for the production of colored light?
- Why do different chemicals emit different colors of light?
- Why do you think the chemicals have to be heated in the flame first before the colored light is emitted?
- Colorful light emissions are applicable to everyday life. Where else have you observed colorful light emissions. Are these light emission applications related? Explain.
- What is the characteristic flame color for Sodium, Lithium, Barium, Copper, Cesium, and Calcium? Explain why.
- When the diffraction grating was used to view the flame, explain why different colorful emission lines

were observed for the elements.

Part II

Use the image below to view the spectra of calcium (top) and sodium (bottom). Solve for frequency and energy of the two brightest emission lines for each element. Use the brightest lines. Show your work and record your answers on the Data Table.



http://imagine.gsfc.nasa.gov/docs/teachers/lessons/xray_spectra/student-worksheet-flame.html

Bibliography/ Resources

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"Electron Transitions" <http://hyperphysics.phy-astr.gsu.edu/hbase/hyde.html#c4>

Standards

Pittsburgh Public Schools Standards

CO2 All students read and use a variety of methods to make sense of various kinds of complex texts

CO4 All students write for a variety of purposes, including to narrate, inform, and persuade in all subject areas

CO6 All students exchange information orally, including understanding and giving spoken instructions, asking and answering questions appropriately, and promoting effective group communications

MA1 All students use numbers, number systems, and equivalent forms to represent theoretical and practical situations

MA2 All students compute, measure, and estimate to solve theoretical and practical problems, using appropriate tools including modern technology such a calculators and computers

MA3 All student apply the concepts of patterns, functions, and relations to solve theoretical and practical problems

MA5 All students understand and apply basic concepts of algebra, geometry, probability and statistics to solve theoretical and practical problems

MA6 All students evaluate, infer, and draw appropriate conclusions from charts, tables, and graphs showing the relationships between data and real-world situations

CI5 All students develop and defend a position on current issues confronting the United States and other nations conducting research analyzing alternatives, organizing evidence and arguments, and making oral presentations

CI8 All students demonstrate that they can work effectively with others

ST1 All students explain how scientific principles of chemical, physical and biological phenomena have developed and relate them to real-world situations

ST2 All students demonstrate knowledge of basic concepts and principles of physical, chemical, biological, and earth sciences

ST4 All students explain the relationships among science, technology, and society

ST5 All students construct and evaluate scientific and technological systems using models to explain or predict results

ST6 All students develop and apply skills of observation, data collection, analysis, pattern recognition, prediction and scientific reasoning, designing and conducting experiments, and solving technological problems.

ST7 All students evaluate advantages, disadvantages, and ethical implications associated with the impact of science and technology on current and future life

ST9 All students demonstrate basic computer literacy, including word processing, software applications, and the ability to access the global infrastructure, using current technology

CW3 All students understand and demonstrate the importance of relating their academic and vocational skills for example interviewing, critical thinking, decision making, problem solving, understanding, and giving written and oral instructions – to their ability to seek, obtain, maintain, and change jobs