ACTIVITY 1

Exploring LEDs and Lamps

Goal
In this activity, you will explore the effect of changing the energy supplied to incandescent lamps and light emitting diodes, then look for similarities and differences among the different light sources.

Matter emits light through various processes that transform other forms of energy into light. For example, a flame from a candle or fireplace during the burning process emits light. The incandescent lamp — a light bulb — is a standard light source that is recognizable by its characteristic shape and appearance. The light bulb contains a solid tungsten filament that emits light when energy is provided by an external energy source such as a battery or electrical power plant.

? Examine the incandescent lamp that you have been provided. Draw the location of the filament and wires inside the lamp.

Voltage is a measure of energy being supplied to an electrical device like an incandescent lamp. Although household incandescent lamps typically require a high voltage to operate, incandescent lamps such as the one that you have been supplied operate with low voltages from a battery.

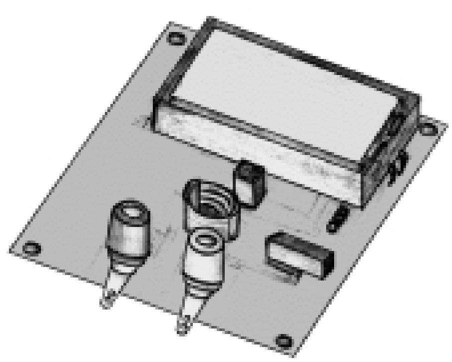
Another modern light source that requires low voltages is the light emitting diode (LED). LEDs are typically used as on/off indicator lights in electrical appliances such as televisions, VCR's, video cameras, computers, and stereos. They are also used to display numbers in some alarm clocks, radios, and microwave ovens. Another use is very large video displays at sporting events and concerts. For example, the music group *U2* during its 1997 *POPMART* tour was using a 56 feet x 170 feet video screen consisting of LEDs. In the 1997 movie, *Batman & Robin*, the Mr. Freeze costume worn by Arnold Schwarzenegger consisted of 3,800 blue LEDs to illuminate his appearance. The low voltage requirements needed to operate LEDs as well as their small size and mass make them an attractive light source to use for these applications.

A diagram of the inside of an LED is shown in Figure 1-1. The chip at the heart of the LED consists of two different solid materials that have been joined together. It is surrounded by a transparent, hard plastic that protects the LED from vibration and shock. The LED is constructed in such a way that the light emitted by the chip is reflected off the base it sits on and is focused through the top of the LED. Thus, the light is brightest at the top of LEDs.

Figure 1-1: Schematic Diagram of an LED

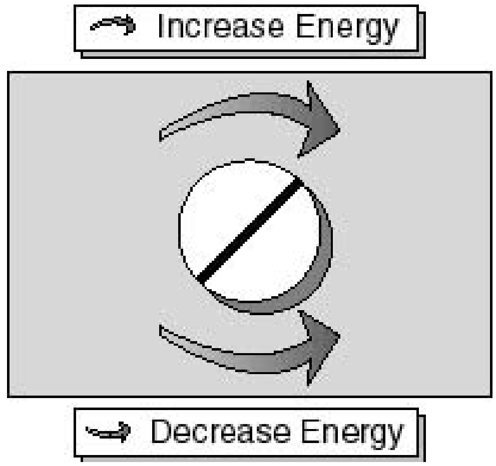
Examine the LEDs that you have been provided. Notice that the two connecting wires have different lengths. These connecting wires are connected to the chip by very thin wires inside the LED.

In this first activity, we will connect lamps and LEDs individually to an electrical energy source and investigate the effects of changing the amount of electrical energy supplied to them. The apparatus uses a small battery as an energy source. The amount of energy reaching the lamp or LED is controlled by a potentiometer — a small rectangular device with a screw on one end. Turning this screw changes the energy going to the lamp or LED in the sockets. The apparatus used for these measurements are shown below.



Check the apparatus by inserting the incandescent lamp in its socket and connect the battery to the battery clips. Adjust the meter in the circuit so that it will have a range of about 0 - 9 Volts, and it will measure voltage. The incandescent lamp should come on and the meter displays the voltage. If not, turn the screw on the potentiometer in the counterclockwise direction until the light comes on. If the lamp still does not emit light, check the connections or ask the instructor to help.

To vary the energy supplied to the incandescent lamp turn the knob of the potentiometer by using the tiny screwdriver. Either end of this screwdriver can be used. However, the end where the blade of the screwdriver is recessed in a cavity is easier to use. The diagram below shows how to increase and decrease the energy.



The voltage measured by the meter is directly related to energy, so we will use it in the observations. As you vary the voltage record your observations about the change in the emitted light. When you have finished, you should be able to answer the following questions.

- ? What is the color of light emitted by the lamp?
- ? How does the color of light change with voltage?
- ? What is the voltage when the lamp barely emits light, called threshold voltage, and the voltage at maximum brightness?
- ? How does the brightness of the lamp change with voltage?

? In the space provided below, record your measurements, observations and answers to the questions.

Remove the incandescent lamp from its socket and place an LED assigned by your instructor in the LED holder.

With the LED in the circuit adjust the potentiometer to vary the energy supplied to the LED. Measure the voltage for the LED. Again when you are finished you should answer the following questions:

- ? What is the color of light emitted by the LED?
- ? If the color of light changes with voltage, how does it change with voltage?

- ? What are the threshold voltage of the LED and the voltage at maximum brightness?
- ? How does the brightness of the LED change with voltage?

In the space provided below, record your measurements and observations. (If the LED does not come on for any voltage, record that information and continue to the next experiment.)

Color of Light Emitted by LED	Threshold Voltage (V)	Maximum Voltage (V)

Remove the LED from the circuit and reverse its connection to the socket. Repeat taking voltage measurements for the LED by focusing on the same four questions as before.

In the space provided record similarities to and differences from the previous investigation.

After everyone has completed the investigation, each group should share its results with the class. The resulting discussion should allow you to make a class table that illustrates the threshold voltage and voltage that results in maximum brightness for the incandescent lamps and LEDs in order of increasing voltage.

Light Source	Threshold Voltage (V)	Maximum Voltage (V)

Based on the results recorded by all groups, answer the following questions:

? What patterns do you see in the observations?

? How is the incandescent lamp similar to the LEDs?

? How is the incandescent lamp different from LEDs?

From our investigations, we found that the LEDs emit individual colors of light with different threshold voltages. Christmas lights are also examples of light sources that emit individual colors of light.

Examine one of the Christmas lights furnished by your instructor. Don't remove the lights from their sockets. The leads have a tendency to break off very easily.

? In the space provided below, describe the physical features of these Christmas lights by concentrating on what you can see inside the lamp.

Connect a Christmas light to the apparatus by using the terminals that are used for the voltmeter. Then, adjust the potentiometer to vary the voltage applied across the Christmas light. When finished answer the following questions:

? What is the color of light emitted by the Christmas light?
? If the color of light changes with voltage, how does it change?
? What are the threshold voltage of the Christmas light and the voltage at maximum brightness?
? How does the brightness of the Christmas light change with voltage?

? In the space provided below, record your measurements, observations and answers to these questions.

? Compare your results with others in the class by recording threshold voltage for each color of light.

? Based on your observations of the Christmas lights, make a prediction on whether the lights are incandescent lamps or LEDs. Explain your reasoning.

A final comparison between the LEDs and Christmas tree lights involve their colors. Summarize the similarities and differences in the color of light emitted by considering the following questions:

? Can you determine the color that will be emitted before you turn on a Christmas tree light and an LED? Why or why not?

? What property of the light source seems to determine the color?

Christmas lights, like LEDs, emit individual colors of light. However, all other physical properties of Christmas lights are similar to incandescent lamps. Incandescent lamps typically emit “white” light but can produce individual colors of light when a colored filter or coating is used. Christmas lights are examples of tiny, incandescent lamps that produce color due to colored coating found on the glass surrounding the tungsten filament. LEDs, unlike Christmas lights, emit various colors of light not as a result of the filament and a colored filter but as a result of a different process that we will learn more about in future activities. In the next activity, we will continue to investigate the physical properties of LEDs and compare these properties with an incandescent lamp and another light source – the gas lamp. This exploration may answer some of your questions and will certainly raise others.

Homework Question:

■ Traffic lights emit their characteristic red (stop), yellow (caution), or green (go) light. Based on your observations and what you have learned, can you determine if traffic lights are incandescent lamps or LEDs? If yes, explain how. If not, describe an experiment that could make the determination (even if you cannot do the experiment).

Name: Class:

SOLIDS & LIGHT

Visual Quantum Mechanics

ACTIVITY 2

Exploring Light Patterns

Goal

We will continue to investigate the properties of LEDs and the incandescent lamp by observing and exploring the light patterns emitted by some devices.

In the previous exploration we saw that incandescent lamps and LEDs have quite different properties when the energy supplied to them is varied. Further, LEDs that look alike on the outside can emit different colors of light, even though they are not painted any color. Thus, the LED, a rather recent invention, acts quite differently from ordinary lamps. Our goal for this series of activities is to understand how these devices work. To accomplish this understanding we need to learn about the emission of light by atoms. Because we cannot see atoms as they emit light, we will need to build a conceptual model of what is happening at the atomic level and use this model to understand LEDs.

The LED is made up of a very small solid consisting of a large number of atoms which are closely packed together and interact with one another in a complex manner. When energy is supplied to the LED, these complex interactions result in the light emitting properties that you have seen. In these solids each atom is very close to its neighbors. Just as with closely spaced people the nature of the interactions can be difficult to understand at first. Thus, we will begin with atoms that are far away from each other; study how they emit light and then work back to a situation where atoms are close together.

Atoms are relatively far apart in a gas. In fact, one of the defining properties of a gas is that the atoms or molecules have only a few interactions with each other. So, we will supply electrical energy to gases confined in a tube. These gas lamps, which are somewhat similar to fluorescent tubes, will emit light. By investigating this light we will be able to build a conceptual model of how gas atoms emit light. We will then extend this model to the closely spaced atoms in a solid and, thus, to LEDs.

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In our investigations we will be particularly interested in the energy of the light emitted by the gas. Two factors — brightness and color — contribute in very different ways to the energy of a light. When we think about the definition of energy, the brightness makes sense. A bright light has more energy in it than a dim light. This conclusion matches the observation from the first activity — as we increased the electrical energy supplied to the lamps, they became brighter.

The color connection is not quite so obvious. Atoms emit light in small packets of energy . These packets are called photons. Each individual photon contains an amount of energy that is related to its color. So, if we wish to discuss the energy of one of these photons, we need to know its color.

For light that we can see the energy ranges from red at the low energy to violet at the high-energy end. Not visible but still a form of light are infrared photons with an energy lower than red and ultraviolet photons which have energies higher than violet. The order of energies for the various colors of photons is shown below.

Low energy photons:	Infrared
	Red
	Orange
	Yellow
	Green
	Blue
	Violet
Higher energy photons:	Ultraviolet

Each time an atom produces light, it emits one photon. Thus, in our investigations we will be primarily interested in the energy of individual photons. As we will see, this energy will tell us something about the atoms of a material. Thus, the color of a light will be an important variable. Each photon of visible light carries a very small amount of energy. This energy ranges from about 2.56×10^{-19} Joules for red light to 4.97×10^{-19} Joules for violet. Using these very small numbers is inconvenient, so we will use different units – the electron volt (eV). In these units, visible light energies range from about 1.6 eV (red) to 3.1 eV (violet) – much easier numbers to deal with.

The brightness of the light is related to the number of photons emitted. A dim light will emit fewer photons than a bright light. Thus, we have two measures of energy — brightness and color. Because color is related to the light from each individual atom, we will concentrate on it.

Most light is composed of several different colors. To separate the colors we use a spectroscope. Inside the spectroscope you will see each of the colors which are present in the light. If you look at white light through a spectroscope, you will see all of the colors of the rainbow. Other light will have fewer colors. This display of color is called a spectrum.

Some spectroscopes provide scales directly in eV. Others show measurements in nanometers (nm) or Ångstroms (Å). If yours uses one of these units, recording your observations on the scales provided on the following pages will enable you to determine the energy value in eV directly.

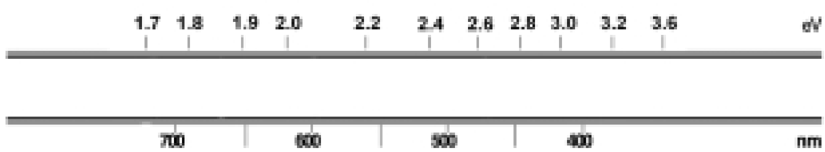
> Caution: (1) Some power supplies for gas tubes have exposed metal contacts. Because the gas lamp is a high voltage light source, do not touch the metal contacts that connect the gas tube to the power supply.
> (2) Never look at the sun or a tanning lamp with a spectroscope. Eye damage may occur from brightness and from high energy ultraviolet photons.

On the following scales, draw the pattern of emitted light observed with the spectroscope for three gas lamps.[Hint] Use colored pencils or markers to indicate the position of color(s). Add a written description to record which colors seem bright

Light Patterns Emitted by Gas Lamps

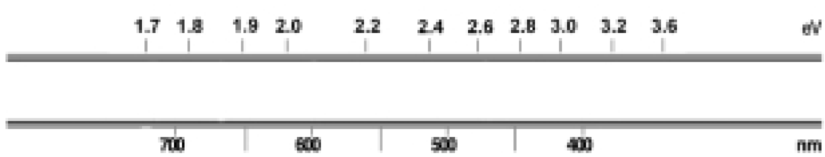
Hydrogen or ________________:

Color of the light without spectroscope__________



Helium or __________________:

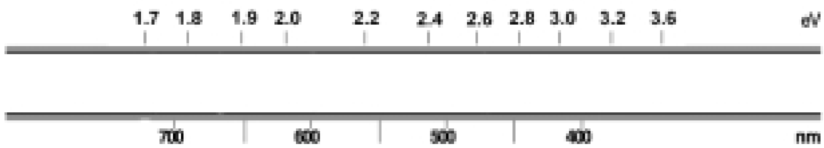
Color of the light without spectroscope__________



[Hint] To ensure that the light patterns are clearly visible, position the vertical slit of the spectrometer (found on the end with a screen) so that it is directly facing the light source and, if possible, hold the spectrometer less than a foot away from the light source. Dim the lights of the room so that the light patterns may be seen. The room, however, should be lighted enough for the energy scale to be seen.

Mercury or _________________:

Color of the light without spectroscope__________



In the table below record the color of light emitted by each gas lamp that is related to the greatest and least energy per photon.

Gas	Greatest Energy	Least Energy

? How can you tell which particular color of light emitted by each gas lamp results in the greatest number of photons emitted?

In the table below record the color(s) of light for which the greatest numbers of photons are emitted by each gas lamp.

Gas	Greatest Number of Photons

? What are the similarities among the light patterns observed for the various gases?

? What are the differences?

Now use the spectroscope to observe the light pattern emitted by the clear incandescent lamp. Connect the incandescent lamp to the circuit that you used in Activity 1 (See Figure 1-2) but without the use of the voltmeter. We will observe the light emitted by the incandescent lamp with the spectroscope when it is at maximum brightness.

? On the following scale, draw the pattern of emitted light observed with the spectroscope for the incandescent lamp. Use colored pencils or markers to indicate the position of observed colors. Add a written description to indicate any colors that are brighter or dimmer than others do.

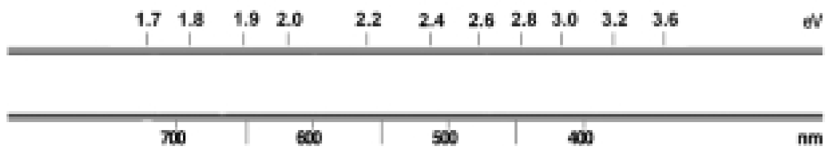
Light Emitted by the Clear Incandescent Lamp

? In terms of the color, intensity, and patterns of light emitted, how is the incandescent lamp similar to the gas lamps?

? How are they different?

Now look at the spectra of one colored incandescent lamp as assigned by your teacher. Record the spectrum below and indicate the portion of the spectrum with the brightest light.

Color of light ______________.



The pattern of light emitted by gas lamps is called a *discrete spectrum*. These light patterns appear as a limited number of bright lines of certain colors. The pattern of light observed for the incandescent lamp is called a *continuous spectrum* for its broad pattern of various colors with no dark regions.

Reduce the brightness of the incandescent lamp by using the potentiometer and the trimmer tool.

? What do you notice about the color of light that is emitted as you reduce the brightness to the point where light is barely visible?

Use the spectroscope to observe the light emitted by the lamp when the brightness is reduced.

? How is the resulting spectrum similar to what you observed before?

? How is the resulting spectrum different from what you observed before?

? Which situation - when the incandescent lamp is bright or dim - results in the greatest number of photons emitted?

? Which situation results in the emission of a larger number of high-energy photons? Explain how you reached your answer.

We will now compare the spectrum emitted by the LED to those emitted by the gas lamps and the incandescent lamp. Carefully remove the incandescent lamp from the circuit and insert the LED that is assigned to you by your instructor. Recall the appropriate manner in which to connect the LED to the circuit so that it will emit light. The best way to observe the light spectrum emitted by the LED is to look at the top of the LED down from above. Recall that the LED focuses light through the top.

Spectrum emitted by a _____________ LED.

After everyone has completed the investigation, each group should share its results with the entire class. The resulting discussion should focus on the following questions.

Compare your observations for LEDs and incandescent lamp with other students who looked at different colors of light sources.

? How are they similar?

? How are they different?

Summarize the results of the class's observations of the LEDs by completing the table below.

LED	Color(s) of Light Observed	Energy (in eV) of Brightest Light Observed

Summarize the results of the class's observations of the incandescent lamp.

Summarize the results of your observations of the light emitted by each source, their spectra and their physical characteristics by completing the table below with the differences and similarities among the three light sources.

Light Source	Gas Lamps	Incandescent Lamps	LEDs
Gas Lamps			
Incandescent Lamps			
LEDs			

Similarities

Differences

As we stated above, gas atoms have fewer interactions than atoms in solids. The spectra for gases show only a few energies while the spectra of solids contain a large number of energies. This observation is a hint that light emission from gases might be less complex than emission from solids. So, we will concentrate on gases in the next activity.

Homework Questions:

Use the spectroscope to observe the light emitted by the fluorescent lights. Notice that the fluorescent lights emit complex spectra that consist of bright lines of several energies and a continuous spectrum. The continuous spectrum is a result of high-energy light interacting with the coating found inside the lights. Based on what you have observed in this activity, what do you think is responsible for the discrete spectrum?

Identify the gas that creates this spectrum.

How do you know?

Spectra can be considered the fingerprints of matter. It allows us to identify not only the material found in artificial light sources but also natural ones like the sun and other stars. We will focus on this procedure in a later activity.

Name: Class:

SOLIDS & LIGHT

Visual Quantum Mechanics

ACTIVITY 3

Introducing Energy Diagrams for Atoms

Goal

Now that we have explored spectral properties of LEDs, incandescent lamps, and gas lamps, we will build a model that can be used to explain these observations. This model will be applied first to explain the spectral properties of gas lamps.

Observations of Light

In the previous activities we observed somewhat different patterns of light emitted by the different types of light sources. We see one difference by just looking at the various light sources. The gas lamps and LEDs are made of clear material yet they emit light of different colors. For the gas lamps the color depends on the type of gas in the tube. For LEDs, the color comes from a process that we will study later. The incandescent lamps, such as Christmas tree lights, are different; the color is contained in the glass surrounding the filament. Understanding these differences is one of our goals.

Another difference appears when we view the spectra. Both the LEDs and incandescent lamps display parts of a continuous spectrum. Their color determines which part of the spectrum is emitted. Gas lamps are different. Each of them emits only certain colors resulting in a spectrum that we call discrete.

In all cases matter inside the lamp emits the light. This material is made of atoms. So we must learn something about atoms to understand the emission of light.

The attraction between the electrons and nucleus means that energy in the form of electrical potential energy is stored in the atom. In addition the electron's motion contributes kinetic energy. So, each electron has a total energy that is equal to its kinetic energy plus its electrical potential energy.

Electrical potential energy occurs for attraction (opposite charges) and repulsion (same charges). To distinguish these two situations we use positive and negative numbers. The positive numbers indicate potential energy associated with repulsion, while negative numbers go with attraction. Because we will work with attraction, we will be using negative potential energies.

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To get the total energy we add kinetic energy (a positive number) and potential energy (a negative number). For an electron in an atom the result for the total energy will always be negative. The idea of a negative energy may seem strange at first. To get an idea of its meaning consider an electron which is *not* attached to an atom, not near any other electrical charges, and is not moving. It is interacting with nothing and not moving, so it has zero potential energy, zero kinetic energy and zero total energy.

If this electron is attached to an atom, its energy becomes negative. The *magnitude* of the energy must be added to the electron to get it back to zero energy — to get it to be no longer attach to an atom and not moving.

For example, suppose we know that an electron has an energy of -13.6 eV. From this information we know that

- the electron is attached to an atom, and
- to get the electron completely free from that atom we must give it 13.6 eV of energy.

Thus, the negative total energy can convey some valuable information about the electrons.

? Which energies below indicate that the electron is attached to an atom?

-1 eV 0 eV 18 eV -8.6 eV

? For each of the energies below indicate how much energy you must add to get the electron free from the atom.

-3.4 eV -54.4 eV -11.5 eV

? An electron has an energy of -4.6 eV. An interaction occurs and it *loses* 5.1 eV of energy. What is its new energy?

? Is it still attached to the atom? Explain your answer.

? An electron is attached to an atom and has a total energy of -8.9 eV. Kevin adds 12.0 eV to this atom. What will be the electron's new energy?

? Will it be moving? Explain your answer.

A useful way to describe the energy of electrons in an atom is to use an energy diagram. The diagram plots the electron's energy on the vertical axis of a graph. We simply draw a line at the energy of the electron. As an example the diagram in Figure 3.1 represents an energy of -3.4 eV.

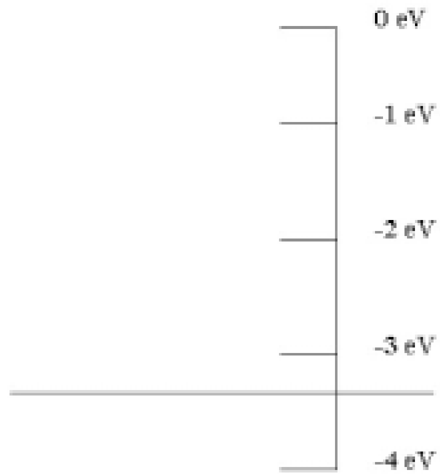


Figure 3.1: An energy diagram for an electron with -3.4 eV of energy.

In this scheme the horizontal axis has no particular meaning. We are only dealing with one variable — the electron's energy. We could just draw dots on the energy axis, but lines are easier to see.

In our studies we will always be interested in electrons that are attached to atoms. So, we place zero energy at the top of the diagram and do not include positive energies.

Changing Energies — Transitions

To emit light an electron must change its energy. This statement reflects conservation of energy.

Electron energy before = Electron energy after + Light (photon) energy

Each time an electron decreases its energy it emits one photon. Thus, by looking at the energy of photons we can learn about what is happening in an atom. From what we can see (light) we infer about what we cannot see (the atom). This process allows us to build models of the atom.

We will use energy diagrams to indicate the changes in the electron's energy. The process is shown in Figure 3.2

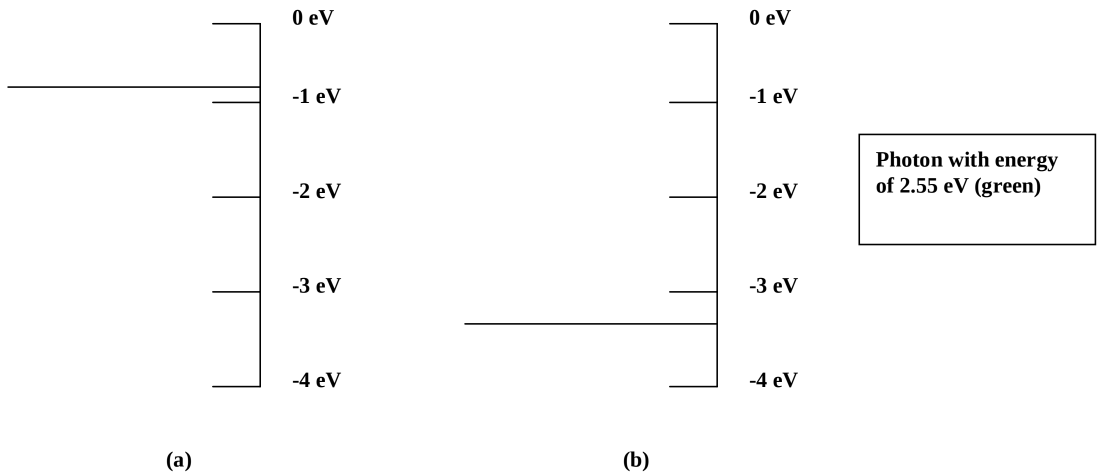


Figure 3.2 (a) Before the emission of light the electron has an energy of -0.85 eV. (b) After the emission of light the electron has an energy of -3.40 eV and a photon of 2.55 eV has been emitted.

The diagrams show the before and after pictures for the electron's energy and indicate that a photon was emitted. To simplify our drawings we generally combine all of the information onto one graph as in Figure 3-3.

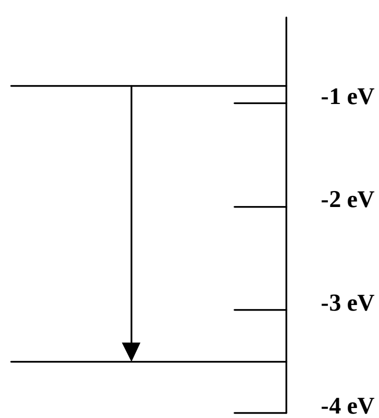
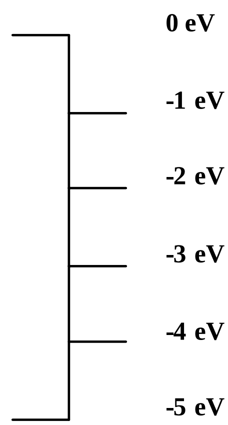


Figure 3-3 The interaction that was shown in the previous figure but combined onto one graph.

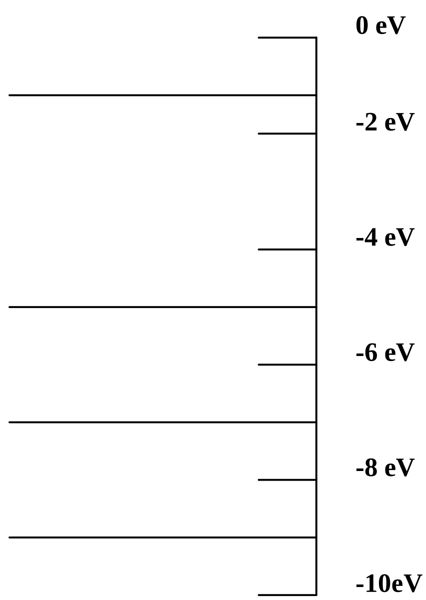
The arrow in Figure 3-3 indicates that electron changed from an energy of -0.85 eV to an energy of -3.40 eV. The sketch above the energy diagram represents what we would see in a spectroscope when the photon is emitted. (One photon is too few to see but it is representative of the energy.)

The process during which an electron changes energy is called *transition*. Thus, Figure 3-3 represents a transition from -0.85 eV to -3.40 eV.

Draw an energy diagram which represents a transition from -2.3 eV to -4.6 eV.

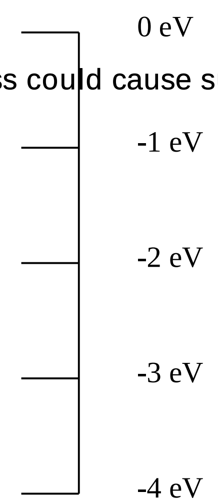


The energy diagram below has four possible energies for the electron. Indicate all transitions that could occur.



? Determine the energies of the photons for each transition.

Another type of transition involves the electron gaining energy rather than losing it. Sketch a diagram which indicates that an electron changed from -3.47 eV to -1.1 eV.



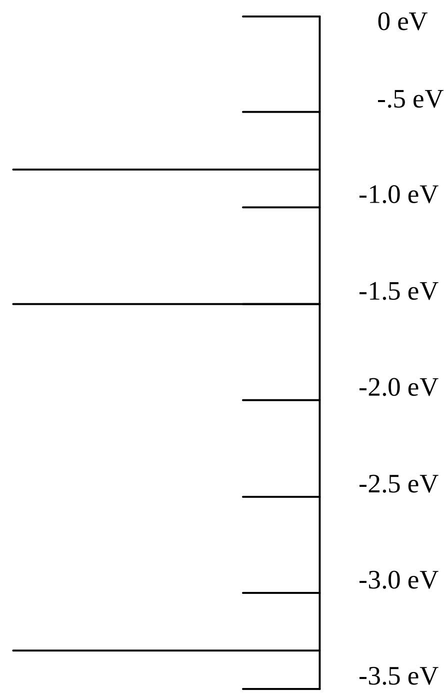
? Speculate about what type of process could cause such a transition. Explain your answer.

An Energy Model for the Atom

The energy diagram provides us with a way to understand some of the processes in the atom. You will use it to understand the various spectra that you have observed. In the process you will learn why, in terms of energy, the spectra of incandescent lamps, LEDs, and gas lamps are different from each other. You will also learn how gas lamps and LEDs can emit certain colors even though they are clear. The first step is to build an energy model of an individual atom. That is the topic of the next activity.

Application Question

The energy diagram below represents a set of energies for an atom.



? What electron transition(s) can exist?

? For the electron transition(s), what are the resulting energies of the emitted light?

? What are the colors of light associated with these electron transitions?

Name: Class:

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Visual Quantum Mechanics

ACTIVITY 4

Understanding the Spectra Emitted by Gas Lamps

> *Goal*
> You will use your observations of gas spectra to build a model of energies in an atom.

In the last activity, we learned that an electron in an atom loses energy equal to the difference between two energy values. The energy lost by the electron appears in the form of light. The energy difference determines the energy and, thus, the color of light emitted by the atom.

We will now use *Spectroscopy Lab Suite* to see how the spectra of light emitted by gases can help us understand more about the energies in an atom.

In *Spectroscopy Lab Suite*, select *Emission* under *Gas Lamps*. Figure 4-1 shows the screen that appears. In this program, we can

- Select a gas tube and drag it to the socket that is just above the lamps. Some of the light in the spectra for that gas will appear at the top of the screen.
- Add energy levels for an electron in a potential energy diagram by using the Add Energy Level button.
- Move the energy levels by selecting them at the left of the vertical energy scale and dragging them to the desired position.
- Create transitions (represented by vertical arrow) by selecting the electron's initial energy on the right of the energy scale. (It turns green.) Drag the transition arrow to the electron's final energy. When you reach the final energy, it will turn green.

- This process will enable you to create an energy level model of the light emitting process in an atom. From the results you will be able to learn about energy levels in atoms. A colored spectral line on the screen above the potential energy diagram will indicate the light emitted by the transition. If the light is not in the visible region of the spectrum, it will not appear on the screen.
- Move any of the energy levels after you have created a transition. Begin with hydrogen. Follow the procedure on the previous page to place the hydrogen gas tube in the socket. Some of the spectral lines for hydrogen will appear in the top spectrum.

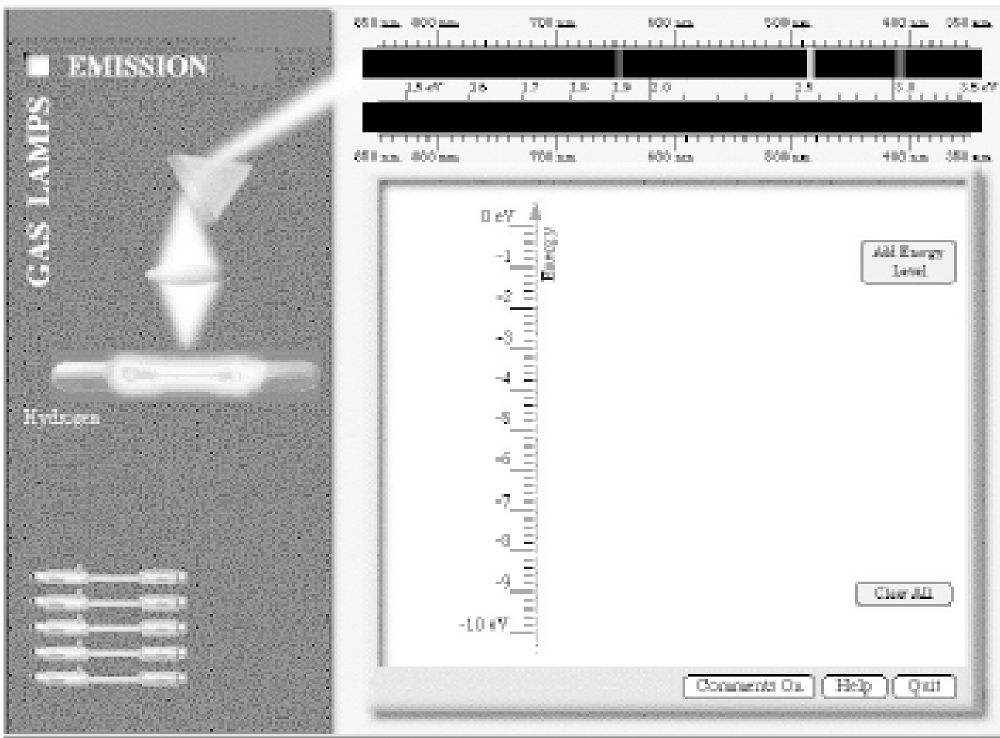


Figure 4-1: *Gas Lamp Spectroscopy* Computer Program

Begin with hydrogen. Follow the procedure on the previous page to place the hydrogen gas tube in the socket. Some of the spectral lines for hydrogen will appear in the top spectrum.

Create energies and a transition that will match one of the spectral lines of hydrogen. The spectrum that you create appears on the lower spectrum.

? How many energy levels are needed to create one spectral line?

? What is the energy of the spectral line as indicated by the eV scale?

? List the values of the energies that you created.

? What is the difference in energy between the electron’s initial total energy and its final total energy?

? How is this energy difference related to the energy of the light emitted by the atom?

Move the energy levels up or down but keep the difference in energy between the electron’s final and initial energy levels constant.

? Why does the spectral line stay at the same energy?

Now create and move energy levels until the bottom spectrum matches the spectrum of hydrogen as shown in the top spectrum.

Sketch the resulting energy level diagram for hydrogen in the space below.

? How many energy levels are needed to create these three spectral lines?

? How many electron transitions are needed to create these three spectral lines?

?	What other, if any, possible electron transitions can take place with the energy levels illustrated on your screen?

Compare your energy diagram with the diagrams created by other students.

?	How are they similar?

?	How are they different?

At this time none of the energy diagrams is more right or wrong than the others. We do not have enough information to distinguish exactly what transitions or initial and final energies occur in nature. Our model is limited by the knowledge that we have. Thus, all sets of energies and transitions that reproduce the spectrum are equally correct. (Scientists have more information to help distinguish the various possibilities, but that is not needed for our purposes.)

We can create energy diagrams that provide all of the spectral lines rather easily. We need only a few energies to have sufficient transitions for all of the visible light. From this construction we conclude that an electron in an atom can have only a few energies. Otherwise we would see light of many more colors. This conclusion is somewhat surprising. When an electron moves in an atom, *it might seem* that the electron could have any one of many energies. But, nature does not behave that way. Instead electrons in atoms are limited to a very few discrete energies. We call these energies the allowable ones.

Repeat the steps to determine the energy levels and transitions necessary to produce the spectral lines emitted by another gas that is assigned by your instructor.

Sketch the resulting energy level diagram for the second gas in the space below.

? How is the energy level diagram for the second gas similar to the diagram for hydrogen?

? How are they different?

Up to this point, we have learned that light is produced when electrons make transitions in atoms. If they have high energy, they naturally lose it in the form of light as they move to a lower energy level. In a normal situation the electrons will be in a low energy level. They must first be given energy to attain high energies so that it can naturally lose that energy. An external energy source, such as electricity must supply that energy. This process is illustrated in Figure 4-2.

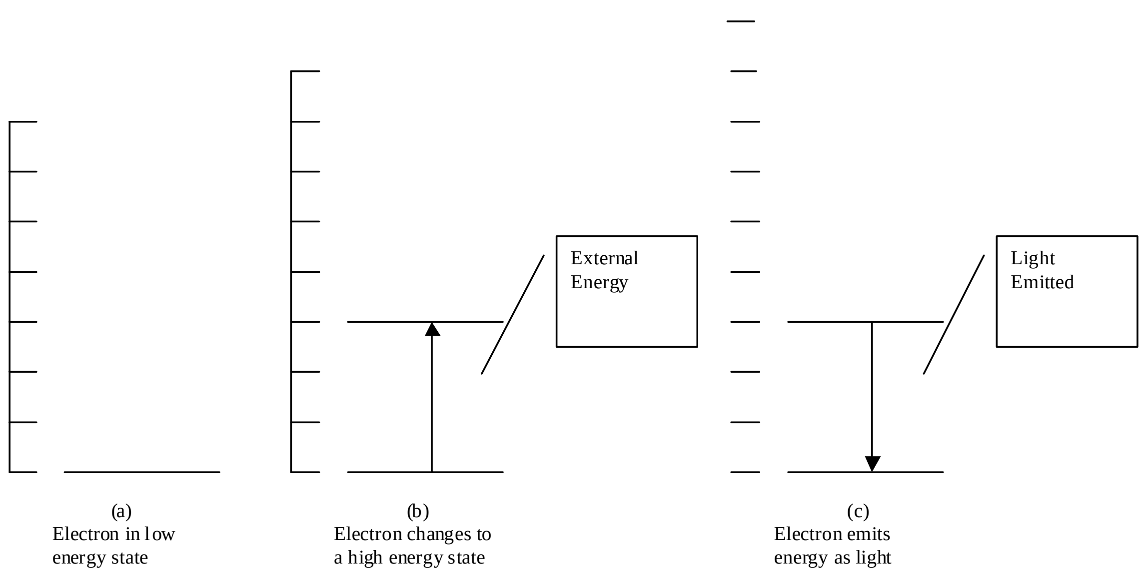


Figure 4-2: Gain and Loss of Energy by Electrons in an Atom

? What external energy source was necessary for the gas lamps to emit light? What form of energy was supplied?

The larger the external energy provided, the greater the number of electrons that will obtain higher energies. For example, suppose the energy difference between two allowed energies is 2.55 eV. See Figure 4-3.

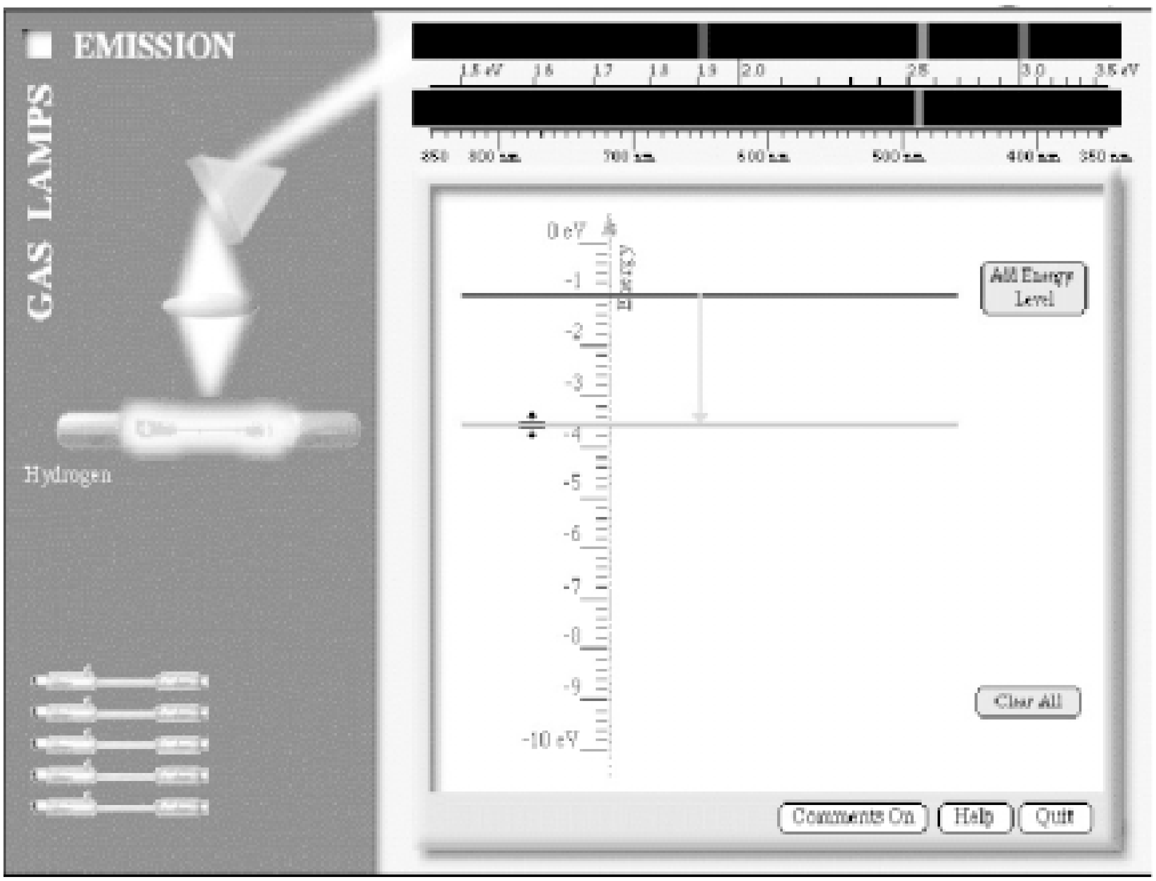


Figure 4-3: Screen capture from the Gas Lamps program.

In other words, 2.55 eV must be supplied by the external source for a *single* electron to change from a low energy to a higher one. Supplying a larger amount of external energy does not change the allowed energies. Recall that the allowed energies for an electron bound to an atom depend on the type of gas atoms found in the lamp. Supplying a larger amount of external energy causes a larger number of electrons to possess the highest allowed energies. Thus, more electrons will make transitions from higher allowed energies to lower allowed energies that result in the greater emission of photons and brighter light.

A different situation occurs if the energy supplied to the atom is not equal to the difference between energy levels. For the example illustrated in Figure 4-3 no transition will occur if the atom receives less than 2.55 eV. If the atom obtains 2.40 eV of energy, it cannot use it. No energy level exists for that transition to occur. If it gets more than 2.55 eV but less than 5.10 eV of energy, only one electron can make the transition. When 3.20 eV is available, the electron can use 2.55 eV and the remaining 0.65 eV will end up as some other form of energy.

In hydrogen and helium atoms, unlike the *Gas Lamp Spectroscopy* computer program, the allowed energies for an electron cannot be changed because the type of atom uniquely defines them.The values for the energies are determined by electrical interactions between the nucleus and the electrons.

? What differences are there between hydrogen and helium atoms that might account for their electrons having different energies?

The electrical properties of an atom uniquely determine what energies its electrons are allowed to have. So, even though the *Gas Lamp Spectroscopy* computer program allows you to adjust the energies available to the electrons, these energies are fixed at very specific values by the electrical properties of the atom.

Because the atoms of each of the elements have a unique set of energies, the light given off by a material can be used to determine the type of elements present. This property is used to learn about the composition of distant stars as well as substances on earth.

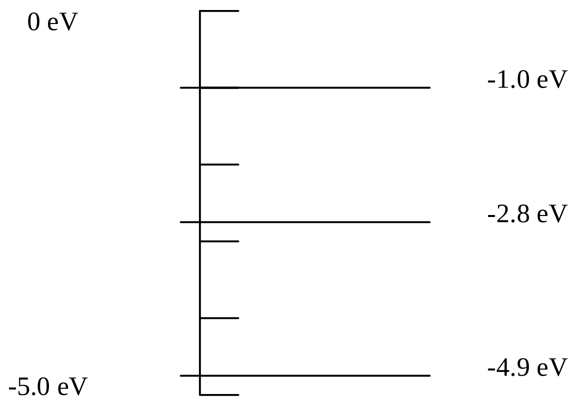
Homework Problem:

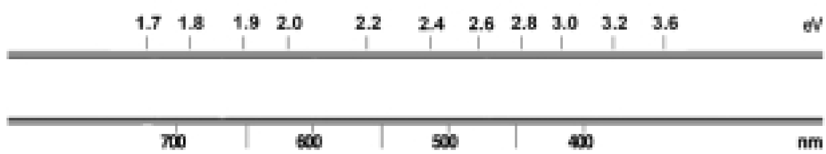
1. Art historians are frequently faced with determining whether paintings are originals or forgeries. Suppose an art dealer who seems a little suspicious has a painting that she claims was painted in 1704. Because of its age she says that it is very expensive. However, one of the colors looks remarkably like pigments from cobalt blue. As an art expert you know that blue paint that used the element cobalt was not used by artists until 1804. How might you use the ideas presented in this activity to determine whether the painting contains cobalt blue?[1]

[1] Adapted from Jacqueline D. Spears and Dean Zollman,The Fascination of Physics, (The Benjamin /Cummings Publishing Company, Inc., Menlo Park, CA 1985). Permission granted by the Authors.

2. The energy level diagram below represents a possible set of energies for an atom.

a) Construct the spectrum of light emitted by a gas represented by the following energy level diagram.





b) Visible light ranges in energy from 1.6 eV (red) to 3.1 eV (violet). How much of the resulting spectrum lies in the visible range?

ACTIVITY 5

Applying Spectra and Energy Diagrams to Learn About Stars

Goal

Now that we can use energy diagrams to describe the emission of light by gases we will apply this knowledge to see how scientists learn about the composition of stars.

The study of spectra in all its forms plays an important role in determining the chemical make-up of matter. Spectra can be considered as the fingerprints of matter. Every atom emits its own characteristic spectrum of light. The study of spectra has been an important tool for scientists to identify the chemical composition of substances isolated in chemical, biological, and astronomical research. For example, the study of spectra from starlight showed that the sun and some of the stars are made of hydrogen and helium. Helium was actually discovered in the analysis of solar spectra before it was discovered on earth. In this activity we will learn how scientists were able to discover helium on the sun even though they had not yet isolated it on earth.

Warning: Never look directly at the sun. To view the sun's spectrum place white paper in sunlight. Look at the light reflected from the paper.

The sun, and other stars, produce the light that we see by a process similar to the one used in incandescent lamps. Material near the surface of the star becomes very hot. Some of the energy is emitted as a light.

During the last century, scientists looked carefully at the spectrum of the sun. They saw something similar to the spectrum that you saw for the incandescent lamp. It was continuous but it had an important difference — dark lines appeared in the sun's continuous spectrum. An example of this type of spectra is represented in Figure 5-1. The gray area represents light coming through while the black lines show some of the places where no light was seen.

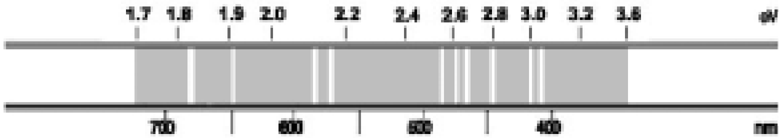


Figure 5-1: Some of the spectrum of a sample star.

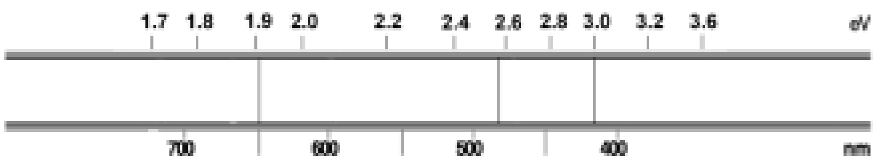
While this spectrum is similar to previous ones we have studied, it also has some unique features. As we understand it we will learn that:

These dark lines can be used to identify elements as well as the bright lines in the emission spectra of gases can.

A process that is similar, but not identical to emission creates the dark regions in this spectrum.

Begin by comparing the dark lines in the solar spectrum with the lines of light from the gases shown in Figure 5-2.

Hydrogen:



Nitrogen:

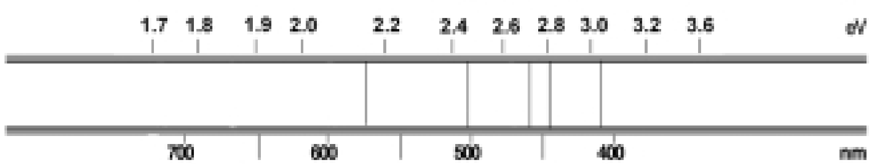


Figure 5-2: The light emission spectrum of hydrogen and nitrogen.

? What similarities do you see between the hydrogen and nitrogen spectra and the spectrum of the star?

? Do you think that you see evidence for the presence of hydrogen or nitrogen on the sun? Explain your answer.

We see a rather close match of the energies in the gas spectra with some of the energies of the dark lines of the star's spectrum. These observations are a hint that the process that produced the dark lines is related to the one that produced light in gases.

A sign which photographers sometimes have on darkroom doors reads "Keep the door closed so the dark will not escape." This joke plays on the idea that dark is the absence of light (or energy), so it cannot escape. We will use the same idea with the star's spectrum. The dark lines are energies that are missing from its spectrum. Each of these energies has been removed while other energies are present.

To understand the processes we need to think about energy in the light. In the previous activity we saw how light is produced by energy changes in atoms. Now think about what needs to happen for one energy of light to be removed from a beam containing all energies of visible light.

? Can you use the energy level diagram to describe how a photon of light can be removed from a beam of light?

A spectrum with dark lines in it is called an *absorption spectrum*. To help understand how absorption spectra occur we will use the *Spectroscopy Lab Suite* computer software.

In *Spectroscopy Lab Suite*, select *Absorption* under *Gas Lamps*.

Drag the figure of the hydrogen gas tube to the gas lamp socket. Figure 5-3 illustrates what should appear on the screen.

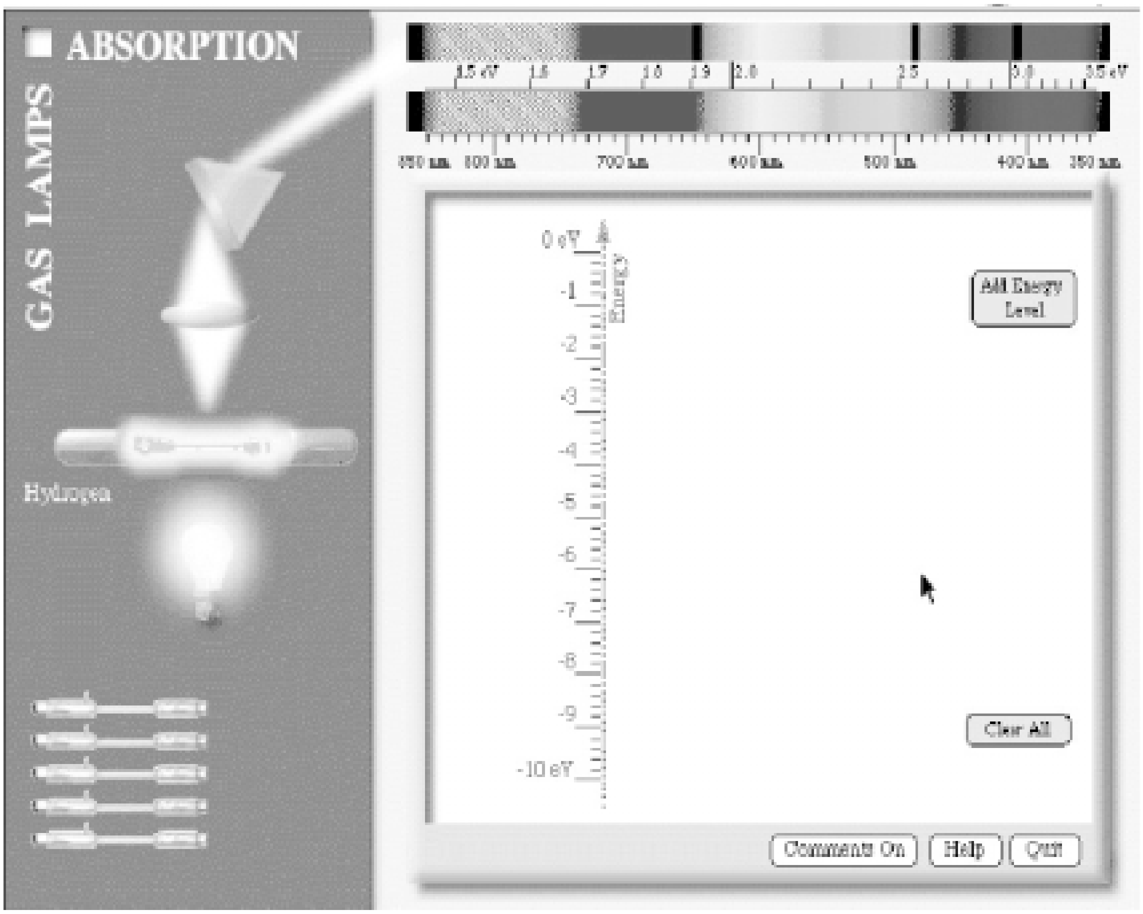


Figure 5-3: A screen from the *Spectroscopy Absorption* Computer Program

The dark (absorption) lines appear in the spectrum at the top of the screen.

To duplicate this spectrum, create and move energy levels and then put in the transitions.

An absorption transition, unlike an emission transition (which we constructed in the previous activity) is created by starting with a lower energy level (largest negative energy value) and ending with a higher energy level (smallest negative energy value). Notice that the arrows that represent the allowed transitions point upward. The arrows point upward to represent that the hydrogen gas absorbs these energies emitted by the light source. As a result, electrons acquire enough energy (equal to the difference of the allowed total energies) to move up to the higher allowed energies.

Sketch below the energy levels and transitions.

Once you have completed your diagram compare it to others in your class.

? How are the diagrams similar?

? How are they different?

Now, compare these energy levels and transitions to those of the emission spectrum for hydrogen that you created in Activity 4.

? How are the emission and absorption spectrum different?

Now, use energy considerations to describe how an atom absorbs a photon from a light source containing all visible energies.

? Suppose that we create a beam of light that has photons of energies only in region of 2.0 to 2.4 eV. That beam passes through a tube with hydrogen. Would you see dark lines in the resulting spectrum? Why or why not?

? Suppose this light (2.0 - 2.4 eV) passes through nitrogen. Would the resulting spectrum have dark lines? Why or why not? (The nitrogen spectrum is shown in Figure 5-2.)

Now that we can use energy levels to explain absorption spectra, we can return to the spectrum of our sample star. In a star a continuous spectrum is created as a result of an extremely hot environment. (More about this process will be discussed in a later activity.) A layer of relatively cool gases surrounds the star. Light emitted by the star passes through these gases. Atoms of the gases absorb photons and move from a lower energy state to a higher one. Then, photons of certain energies are no longer present in the light. Thus, we see dark lines — the absence of certain energies — in the spectrum of our star.

An important conclusion from this analysis is that absorption spectra can be used to identify elements in much the same way as emission spectra. 19th Century scientists used this identification process when they looked at the spectrum of the sun.

? Are both hydrogen and nitrogen present in the gases surrounding the star?

? Do you have any lines that are not accounted for by either hydrogen or nitrogen?

The 19th Century scientists went through this same process. The full solar spectrum had many more absorption lines than shown in Figure 5-1. The scientists were able to identify elements that could have absorbed photons associated with almost all of them. They even identified metals such as iron in the gases surrounding the sun. But, when they got done, they had some lines that they could not associate with any atom. These lines were the same ones that you had left over.

They knew that absorption of certain energies meant the presence of a certain type of atom. But, they had never seen either emission or absorption at these energies. So, an element not detected (in the 19th Century) on Earth must be in the gases around the Sun. They named the element helium after the Greek word helios, which means sun.

Name: Class:

SOLIDS & LIGHT

Visual Quantum Mechanics

ACTIVITY 6

Using Spectra to Search for an Earth-like Planet

Goal

Now that we can explain why gas lamps emit their characteristic spectra and how absorption spectra are related to emission, we will apply our knowledge in a science fiction type scenario. The purpose will be to identify the gases present on a mythical planet and determine whether conditions similar to earth exist there.

To continue the study of spectra imagine we are a group of scientists in a star ship. Our job is to look for planets that might be hospitable for humans. These plants are called Class M on Star Trek. Lots of planets exist around other stars so it is a big job. If we were to take the time needed to travel to each one – even at warp speed – and transport down, we might never finish the job. In fact, we would not want to transport down until we knew something about the atmosphere. So we need to investigate the gases in the atmosphere from afar.

Fortunately, we have the perfect tool for learning about the atmosphere without needing to get close to the planet. We can look at the light that passes through the atmosphere of the planet and see what energy photons are absorbed by it. The basic arrangement is shown in Figure 6-1.

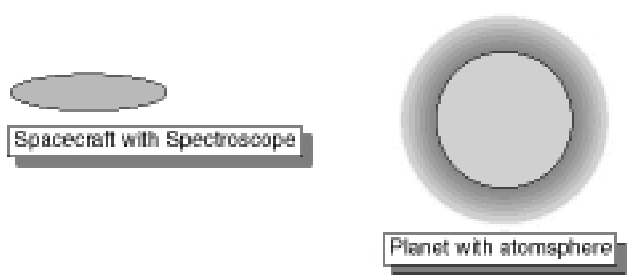


Figure 6-1: The arrangement of the starship, planet and a nearby star.

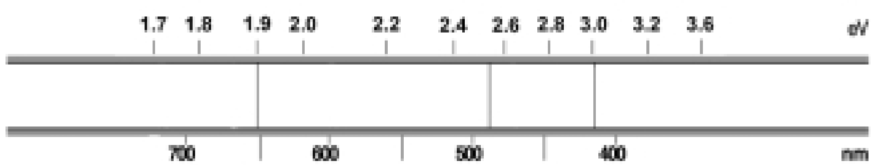
 Visual Quantum Mechanics is supported by the National Science Foundation under grants ESI 945782 and DUE 965288. Opinions expressed are those of the authors and not necessarily of the Foundation.

? Our starship is equipped with very good spectroscopes. How would this arrangement allow us to investigate the elements in the atmosphere of the planet?

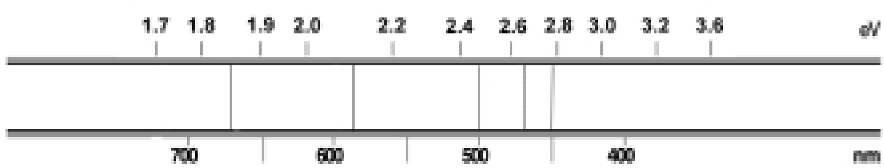
The astronomers on our starship team bring data showing the spectra absorbed by the planet's atmosphere, as observed through a telescope that has a spectroscope attached.

Our task is to use the recorded spectra and compare them with those of some commonly known terrestrial gases. Then we can determine what gases are present in the planet's atmosphere. The spectra of various known gases are illustrated in Figure 6-2.

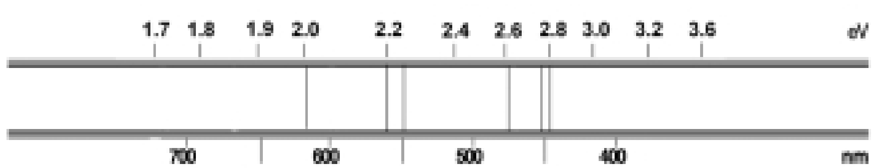
Hydrogen:



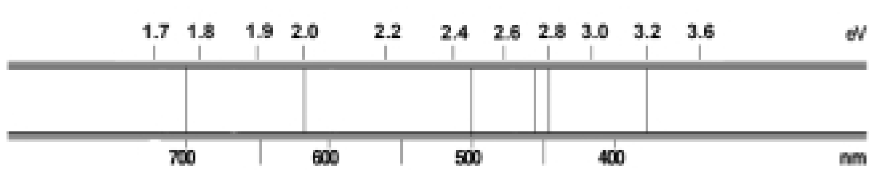
Helium:



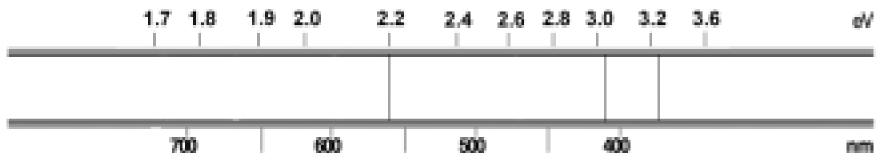
Carbon Dioxide:



Oxygen:



Ozone:



Nitrogen:

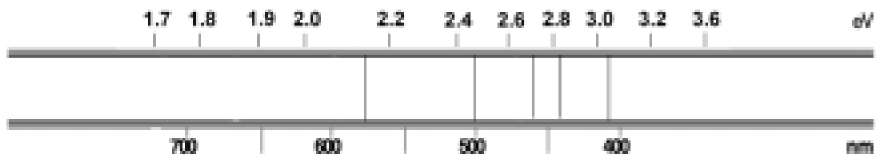


Figure 6-2: The Visible Spectra of Various Gases

? Looking at the list of gases found in Figure 6-2, what other gas(es) or substances, if any, would you add to this list if we were looking for conditions on a planet that could support life? Explain.

? Could any of the gases in our list on the previous page be omitted? Why or why not?

? Your teacher will give you a copy of the planet's spectrum. Compare the spectra of the planet's atmosphere with that of the other gases. Which of these gases are present in the planet's atmosphere? Support your conclusions.

? Are there any gas(es) other than the ones listed above present in the planet's atmosphere? Explain how you reached your conclusion. List the energies of any light that cannot be accounted for in gases listed on the previous page.

? Based on the information and your analysis describe and justify your conclusions about the similarity of this distant planet's atmosphere and earth's.

If other solar systems in the universe do contain Earthlike worlds they must have a planet at a distance from a star so that water can exist in liquid form. If the planet is too far away, it will only contain ice. Too close and all the water will be steam. Planets such as earth are very fortunate to be located where water can exist as gas, liquid and solid. Scientists use the term, "habitable zone" (defined as the planet's distance from its star in which water exists in liquid form rather than as a solid or a gas) in describing the probability of whether conditions could support life. The Galileo space probe launched by NASA passed over Jupiter and found that the atmosphere not only contained hydrogen, neon, helium, and noxious gases such as ammonia and methane but also had higher winds and less lightning and water than scientists had predicted.

The Galileo probe also found that Europa, one of Jupiter's moons, has an icy white surface and shows evidence of water and oxygen. Some scientists speculate that the surface could conceal oceans of water and perhaps some sort of living organisms. Europa was the focus of the science fiction movie, *2010*, in which the humans witnessed the birth of a cosmic event and were given a second chance to reconcile their national differences. NASA is making plans to possibly send three spacecraft to Europa to determine if it does have a water ocean.

Our neighboring planet, Mars, has also been the focus of science fiction movies like *War of the Worlds* and *Total Recall*. The thin atmosphere of Mars (less than 1% of Earth's) is made almost entirely of carbon dioxide with traces of nitrogen and argon. As a result of finding evidence of massive floods, scientists speculate that Mars, now a bone-dry salmon-colored surface, had water 1 to 4 billion years ago. If that is the case, scientists are optimistic in that one-cell microorganisms may still be surviving or at least fossil evidence may exist somewhere under the Martian landscape.

The *Mars Pathfinder* spacecraft[Note], was launched by NASA in December of 1996 and arrived on Mars in July of 1997. The project involved the landing of a spacecraft (called *Pathfinder*) that contains a miniature ground-roving vehicle (called *Sojourner*) with several instruments on board. As *Sojourner* moved across the surface of Mars at a speed of 2 feet per minute, the instruments on board identified and analyzed a wide variety of surface materials from the heavily cratered terrain to the ridged plains to channel deposits. The rover has a life span of about a month while the lander has a lifespan of a year. During these times, data will continue to be collected, analyzed, and reported back to Earth. Up to 10 missions are being planned by NASA by 2005, including a rock-gathering round trip.

One of the instruments is a circular spectrometer that is mounted near the front of *Sojourner* (See Figure 6-3.). By shooting electrically charged particles at objects on Mars it causes the atoms in the materials to emit energy. The spectrometer records the spectrum and sends the data back to earth. The instrument not only acquires spectra from Martian dust, but more importantly, is deployed to distinct rocks where it analyzes the native rock composition for the first time. The spectrometer determines the chemical make-up of surface materials and can find most major elements except hydrogen. The approach used is to expose the Martian rock to high-energy particles with a known energy and to analyze the resulting energy spectra emitted by the sample.

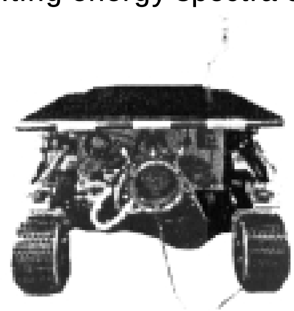


Figure 6-3: Front View of Spectrometer Mounted on the Rover

Note For more information about the Mars Pathfinder Mission, check out the following web page: http://mpfwww. jpl.nasa.gov/default1.html

The study of spectra includes energy that is outside the visible region such as infrared and radio (less than 1.6 eV) and ultraviolet and x-rays (greater than 3.1 eV). The Sojouner uses x-ray spectra in many of its analysis.

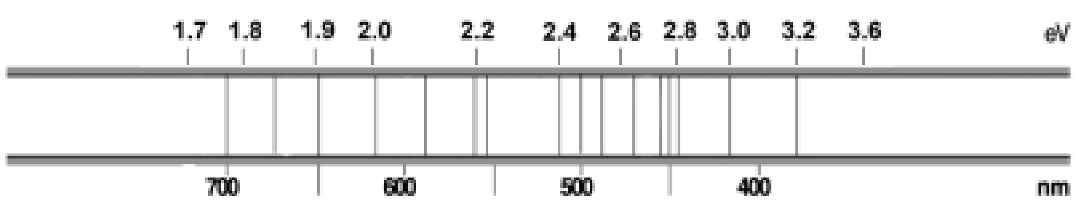
Initial analyses of the spectra from rocks have identified a rock rich in silica — a quartz material found in sand. Such a rock could have been brought to the surface by volcanic activity or by a meteorite impact. A rock rich in silica could only be formed by repeated heating in an active crust of a planet. A second type of rock may be andesite - one of the most common types of lava found on Earth. Since some types of andesite only form in the presence of water, scientists are speculating that Mars may have had water in its interior about the same time life began on Earth about 4 billion years ago.

We have learned that spectra can be used to determine some of the properties of the atoms of a gas. The spectrum from each type of atom is unique, somewhat like the uniqueness of human fingerprints. Just as fingerprints can be used to identify people, spectra are used to identify atoms in materials. For example,

- Astronomers determine the elements in a distant star by analyzing its spectra.
- Crime labs use spectra to match samples of paint on hit-and-run victims and suspected cars.
- Historians and archeologists use spectra to determine the elements used in painting and pottery. By knowing when different cultures started using substances containing certain elements they can determine the age and origin of the object.
- Biologists use spectra to identify the presence and concentration of substances.

These differences in spectra are, in turn, related to the energy levels of electrons in the atoms. These energy levels are related to the interactions between electrical charges. Thus, the spectra and our representation of atoms with the energy diagram are very closely related. Much of our knowledge of the universe, ranging from the enormous scale of galaxies to the small scale of atoms, is built on the study of spectra and our knowledge of the energy transition in atoms.

Planet's Atmosphere Spectrum



ACTIVITY 7

Using Gas Lamps to Understand LEDs

Goal
With the help of computer programs, we have been able to use the energy level diagram to explain the spectra emitted by gas lamps. We will now apply what we have learned to explain the spectra emitted by LEDs.

Our study of gas spectra has led us to the conclusion that only certain energy levels can exist in a gas atom. Now, we wish to extend our investigation to solids so that we understand how LEDs emit light. As a first step we will explore how we might create a spectrum similar to that of an LED.

Open the *Emission* version of the *Gas Lamp Spectroscopy* computer program and place the unknown gas tube in the gas lamp socket.

When the unknown is in the socket, you can create your own spectrum by dragging the lines near the top of the screen. Edit the energy values for the computer-generated spectral lines so that the spectrum is similar to the spectrum of one LED that you observed in Activity 2.

Create an energy level diagram for an atom that could produce this spectrum.

In the space below sketch this energy level diagram.

 Visual Quantum Mechanics is supported by the National Science Foundation under grants ESI 945782 and DUE 965288. Opinions expressed are those of the authors and not necessarily of the Foundation.

Compare your results with other students who observed different LEDs in Activity 2. The resulting discussion should focus on answering the following questions.

? How are the allowed energies in your diagram similar to the others?

? How are they different?

? How are the energy level diagrams for the LED similar to ones for a gas lamp?

? How are they different?

This exploration shows us that we could get a spectrum similar to that of an LED by having several closely spaced energy levels. When we look at the spectrum of an LED, we see a broad spectrum with no dark regions in it. So, atoms in an LED must have many energy states that are extremely close together. No real gas has the energy levels to create this type of spectrum. We can create it only on a computer with our "un-known" gas. So, we must look beyond gas atoms to explain the spectra of LEDs. This conclusion is not surprising because LEDs are made of small bits of solids.

Solids have many atoms that are close together and interact with each other. These interactions create very closely spaced energy levels. In addition to having energy levels which are very close together, a solid has an extremely large number of levels – literally billions and billions. Because of the large number and the close spacing we treat each group as a band of energy level. When you tried to match an LED spectrum with the *Emission Spectroscopy* program, you created something similar to an energy band with just a few levels. A solid may have several bands of energy. However, only two of the bands are involved in light emissions. (So, it works just like the model you created with closely spaced spectral lines.) The band with the highest energy contains electrons that cannot leave the solid but are not firmly attached to any atom. They can move throughout the solid. This freedom of motion allows these electrons to carry (or conduct) energy through the solid. So, we call this band the *conduction band*.

Electrons that have energies in the next lower band are bound to their respective atoms more strongly and are unable to break free from the atoms. This lower energy band is called the *valence band.* Electrons with these low energies have large negative values because they require more energy to escape their respective atoms.

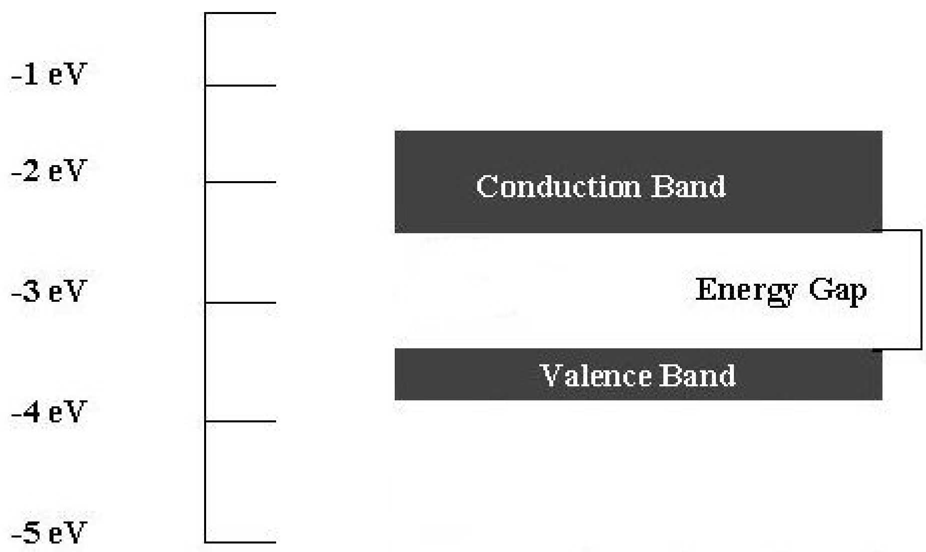


Figure 7-1: Energy diagram with a Very Large Number of Solid Atoms

The space between the conduction band and valence band has no allowed electron energies. This region is called the *energy gap*.

Now let's look at how energy bands are related to the spectra of LEDs.

At the beginning of this activity, we used *Gas Lamp Spectroscopy* to get an idea about how the energy level diagram must look to explain the spectrum emitted by an LED. At that time we created a pseudo-band by putting several energy levels close together. Now, we will look specifically at the energy bands in LEDs.

In the *Spectroscopy Lab Suite* software package, select LEDs from the main menu. Figure 7-2 illustrates how the screen should appear and provides basic instruction for using the program.

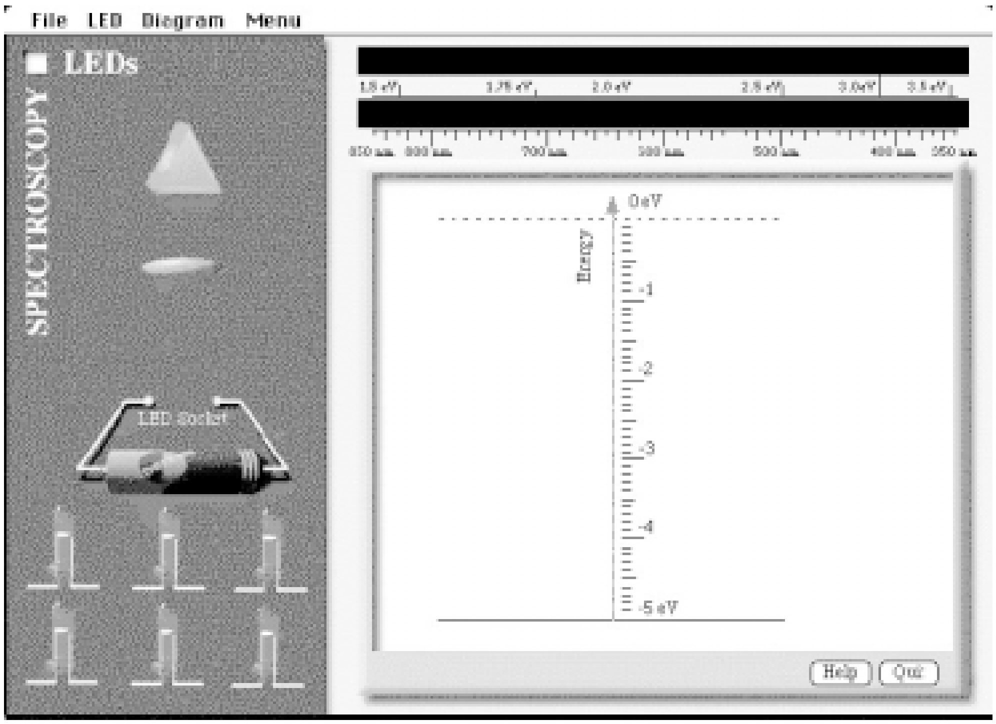


Figure 7-2: *LED Spectroscopy* Computer Program

Drag the LED assigned to you to the LED socket found on the left of the screen. The computer-generated spectrum emitted by the LED will appear on the top screen.

Click the Add Conduction Band button. A red rectangle that represents the conduction band for the LED should appear near the top of the energy scale.

Click on the Add Valence Band button. A faded-red rectangle that represents the valence band for the LED should now appear near the bottom of the energy scale.

The broad, orange vertical arrow represents the allowed transitions for electrons as they move from any energy in the conduction band to any energy level in the valence band. As these electrons make transitions, they emit energy in the form of light.

Place the cursor in the center of one of the bands. The band turns green and a hand symbol appears. You can now change the energy of the band by dragging it up and down.

Place the cursor on the top or bottom edges of one of the bands. The band turns green and up-down arrows appear. You can change the range of energies allowed in the band.

As you change the location of or range in a band, you will see the spectrum. Now manipulate the location and range of both energy bands until the spectra described by the energy level matches.

In the space below sketch the resulting energy band diagram and indicate the range of energy values (in eV) for each band and the resulting energy gap.

After all groups have sketched the energy band diagram for his/her LED, they should share their results with the class. The resulting discussion should focus on answering the following questions.

? How is the size of the energy gap related to the color of light emitted by the LED?

? How are the ranges of energies for the conduction and valence bands important?

We are now able to apply energy diagrams to explain the spectra of LEDs. However, to be successful we needed to extend the concept of individual allowed energies to allowed energy bands. The spectra emitted by LEDs is the result of electrons making transitions from a number of energy levels in the conduction band to a number of energy levels in the valence band. The electron transitions that are allowed can range from the highest energy level of the conduction band to anywhere between the lowest and highest energy levels of the valence band (See A on the left side of Figure 7-3.). Other electronic transitions that are allowed can range from the lowest energy level of the conduction band to anywhere between the highest and lowest energy levels of the valence band (See B of Figure 7-3.). These electronic transitions result in the emission of a broad continuous spectrum that is concentrated on a particular color (and thus energy) of light. The size of the energy gap in solids inside an LED determines the color of light emitted by the LED. Thus, if we know the spectra of light emitted by an LED we can predict its energy gap.

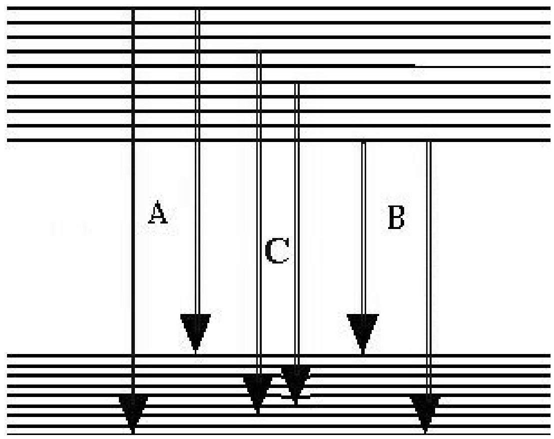


Figure 7-3: Ranges of Allowed Transitions for a Solid that Makes Up an LED

Homework Problem:

An LED is created with a semiconductor that has the energy bands and gaps

a) What is the energy gap of the semiconductor?

b) Determine, using the diagram above, the range of energies emitted by electrons in this solid.

c) Describe the spectrum of visible light emitted by this LED.[Hint]

d) Use the *LED Spectroscopy* computer program to check your answers.

Hint The range of visible light ranges from 1.6 eV (red) to 3.1 eV (violet).

Appendix: Conductors and Insulators

If you have studied the movement of either thermal or electrical energy through solids, you may be interested in how the energy bands are related to conduction and insulation.

The electric current that flows through a solid consists of electrons with energies associated with the conduction band. Solids that conduct heat or electric current (called *conductors*) will have many electrons with energies in the conduction band. Solids that do not readily conduct heat or electric current (called *insulators*) will have few electrons in the conduction band. Thus, the electrons in insulators are found mostly in the valence band. Solids that are good conductors will have many electrons in the valence band because these materials have enough electrons to fill all available energies in the valence band and many left over in the conduction band. Further, these two bands overlap in conductors.

In addition to the difference in the number of electrons found in their respective conduction bands, conductors and insulators also differ in the size of their respective energy gaps. For conductors, the two bands either overlap or are so close together that electrons with energies in the valence band can easily acquire enough energy to have energies associated with the conduction band. Insulators, on the other hand, have energy gaps that are so large that thermal or light energy is not sufficient for electrons to move to the conduction band from the valence band.

The solids that make up LEDs are called *semiconductors*. As the name implies, a semiconductor is a solid whose physical properties place it somewhere between conductors and insulators. In semiconductors the energy gap is small enough that when sufficient energy (i.e. thermal, electrical, or light) is supplied, electrons gain enough energy to become the conduction band electrons.

The electrons in the conduction band carry the energy. Materials that have electrons in this band can transmit easily both thermal and electrical energy. The good electrical conductors are also good thermal conductors.

Name: Class:

SOLIDS & LIGHT

Visual Quantum Mechanics

ACTIVITY 8

Applying Energy Bands to Incandescent Lamps

Goal
We will apply what we have learned about energy bands and gaps to explain the physical properties of the incandescent lamp which we observed earlier.

In the previous activity, we developed the energy diagram representation of a solid (like the one inside an LED) starting from our observation of the LED spectrum and the energy level diagrams of gases. We concluded that energy bands must be present to obtain the characteristic broad spectra emitted by LEDs. We also found that the size of energy gap and the energy range of the two bands determine the color of light emitted by the LED.

The energy level model has successfully described the spectra emitted by gas lamps and LEDs. So, we will apply our model to explain the light emitted by incandescent lamps.

Open the *LED Spectroscopy* computer program and place any one of the LEDs in the socket.

Do not try to match the LED spectrum. Instead, use the program to construct an energy band diagram that would produce the spectrum emitted by an incandescent lamp when it is at maximum brightness.

? What is the range of energies (and colors) for the spectrum emitted by the incandescent lamp?

? How does the spectrum emitted by the incandescent lamp compare with the spectrum emitted by an LED?

In the space below sketch the resulting energy band diagram for your incandescent lamp.

? How does energy band diagram for the light emitted by the incandescent lamp compare with the energy band diagram for an LED?

The energy band diagram that we just constructed represents the diagram for a “white” (mixture of all colors) light LED. Although incandescent lamps emit “white” light, typical LEDs only emit light of a single color. Incandescent lamps are also different from LEDs in that when the electrical energy supplied to the lamp is increased, the color (and energies) of light emitted by the tungsten filament (a solid) changes from the red region of the spectrum to “white” light as the filament heats up.

The energy diagram in Figure 8-1 represents the bands and gaps for tungsten, the material in the filament of most incandescent lamps.

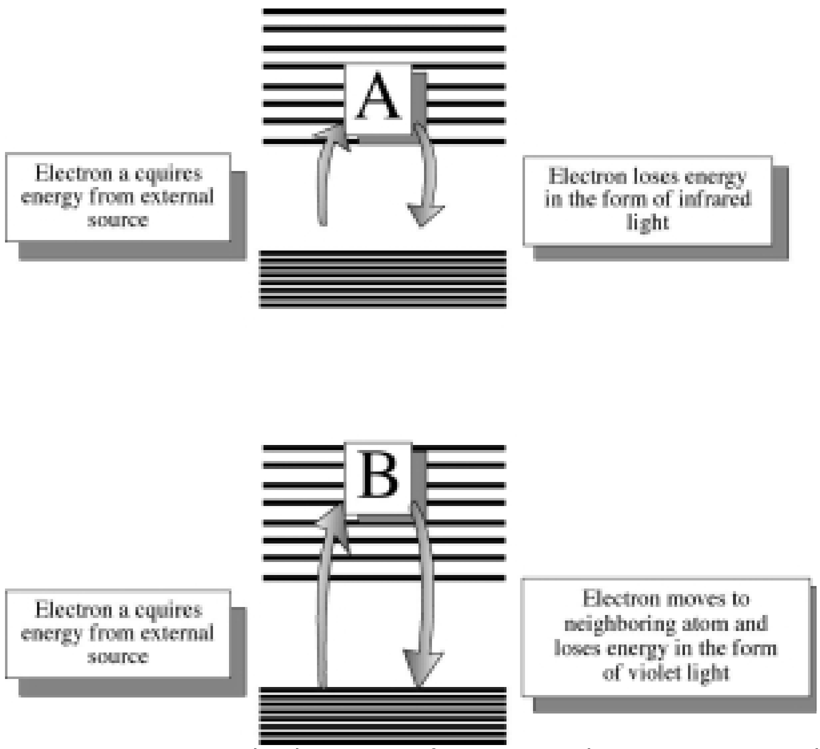


Figure 8-1: Energy Band Diagram of an Incandescent Lamp Filament

Since electrons seek lower energy levels, the electrons of the tungsten filament have energies associated with the valence energy band. When the battery provides sufficient electrical energy, electrons in the valence band make the transition to an energy level of the conduction band. (See Figure 8-1.) These electrons will then lose the energy they recently acquired when photons of light are emitted. As a result, the electrons make a transition back to the valence band of the filament. The light energy emitted can range from infrared to ultraviolet.

If the supplied energy is great enough, electrons from the valence band can make a transition to the highest energy level of the conduction band. These electrons make the transition back to the valence band by losing energy as photons of violet light (3.1 eV). Since enough electrical energy is being supplied to move electrons from the valence band to the highest energy level of the conduction band, the energy is more than sufficient for electrons to make a transition from the valence band to any energy level found between the lowest and highest energy levels of the conduction band. As a result, photons of light ranging from red (1.6 eV) to violet (3.1 eV) are also emitted.

When the energy supplied to the incandescent lamp is low, the color of the light has a large red component. At low voltages many of the electrons will receive just enough energy to reach lower energies in the conduction band. So, they can only emit light in the low energy end of the spectrum. Infrared light, which has a lower energy than visible light, is also emitted. This energy causes the lamp to be hot.

As the energy increases, the color of the light becomes “white.” The number of transitions that result in the release of photons in the middle or high energy end of the visible spectrum increases. Thus, the color of the light shifts from reddish to white.

Use this discussion to draw energy level diagrams that explain the color shift in an incandescent lamp. Draw one diagram for low voltages and another for higher voltages. Then describe the differences.

Unlike the incandescent lamp, the colors (and energies) of light emitted by the LED do not change with voltage. This difference is a reflection of the range of energies available in the conduction band of the materials that compose the devices. Incandescent lamps are made of electrical conductors that have a wide range of energies available in the conduction band. Semiconductors, which make up LEDs, have only a narrow range of energies in this band. This difference is displayed in the behaviors as the voltage (input energy) is changed.

An additional difference is the method by which the light-emitting electrons gain energy. In an LED electrical interactions directly transfer energy to these electrons. The incandescent process works differently. The energy to electrons comes from heating the material. In the incandescent lamp electricity causes the wire to become hot. The thermal energy is transferred to the electrons, and then they emit energy in the form of light. Incandescence is the process by which thermal energy is converted to light.

Homework

1. Electric cooking stoves emit heat and light from their burners. The heat that we feel is infrared radiation. The light that we see is always in the red end of the spectrum. With this information create an approximate energy level diagram for the material in the burners of an electric stove. Explain how you reached your answer.

2. Another example of incandescence is a glowing piece of metal in a fire. Use energy diagrams to explain why very hot metal emits light.

ACTIVITY 9

Creating an Energy Level for Model LEDs

Goal

The model we developed with energy bands and gaps explained why we see a partial continuous spectrum for LEDs. They even allowed us to understand why we needed some minimal energy (voltage) to turn on each LED and why that energy depended on color. The model does not, however, explain how electrons get into the conduction band so they can emit light. To understand this process we must expand our model of the LED energy levels.

We have used energy bands and gaps to explain the spectral properties of an incandescent lamp and LEDs. Recall that the conduction and valence energy bands and the resulting energy gap result from interactions of numerous atoms that are very close together - a situation that occurs in solids. As we have seen, the energy gap of the solid materials (called semiconductors) that make up the LED determine the energies (and thus, colors) of light emitted by the LEDs.

A solid can be a pure material in which all atoms are the same element. As a result, each nucleus of the atom contained in this solid has the same electrical charge. Thus, each atom in this solid has identical properties. The interactions among these atoms create the energy bands and gaps that we have studied. Modern technology can create materials that are very close to being all identical atoms. These pure materials have light emitting properties much like we have studied in previous activities.

For today's technology pure materials are not the most valuable. Instead, a wide range of devices — from LEDs to computer chips — use almost pure materials into which a small quantity of a different element — an *impurity* has been introduced. Then materials with different impurities are joined.

Suppose we start with a pure material and add atoms of a different element. These different elements will have a different number of electrons than the atoms of the original material. We place the impurities into two groups:

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- *Donors* have more electrons than the material's pure elements. They donate electrons to the solid.
- *Acceptors* have fewer electrons than the material's pure element. They accept electrons from the solid.

Both the donors and acceptors have zero electrical charge. They have more or less charge in the nucleus to balance the more or fewer electrons.

The LED chip consists of two solids – a material that has been supplied with donor atoms and the same material that has been supplied acceptor atoms. We will now use the *LED Constructor* computer program to understand how these two materials are combined to construct a simulation of an LED. Begin the process by opening the *LED Constructor* computer program.

Notice that on the left of the screen is an illustration of 6 different LEDs and a version of the circuit apparatus that we used in the first activity when we explored the properties of LEDs. On the right of the screen are the two blocks of material that make up the LED chip with an energy scale and output spectrum illustrated below the blocks.

In this program, we can:

- Select an LED and drag it to the power supply (LED socket).
- Create an energy band diagram for each of the two semiconductor blocks that make up each LED.
- Add donors or acceptors to each semiconductor block with a click of the mouse inside each block and thus create the LED chip.
- Control the voltage applied across the LED by moving the slider.

When the voltage is appropriate, transitions (represented by a vertical arrow) will occur and light will be emitted. A spectrum will appear below the energy scale.

Drag the LED assigned to you by the instructor to the LED socket. The energy bands that appear on both sides of the energy scale represent the bands and gap for materials associated with these LEDs.

Notice that the valence bands of both semiconductor materials are shaded darker than the conduction band. This shading indicates that the majority of electrons have energies associated with the valence bands. Electrons are naturally found in these bands because they seek the lowest possible energies. Also notice that energy of the energy gaps of both blocks are the same to represent that the LED is constructed of two blocks of the same semiconductor.

? Record the color of the LED you were assigned and the energy (in eV) of the energy gap.

Now, click on the Add Impurities button. This places acceptors in the left block and donors in the right block.

? How does adding acceptors — atoms with fewer electrons — affect the energy bands of the left block?

? What is the effect on the energy bands of adding donor atoms to the right block?

? Do the sizes of the energy gaps for either semiconductor block change as donor or acceptor atoms are added?

? A material that has donor atoms has more electrons than one with acceptors. When these two materials are joined together, the electrons can move from one material to the other. Which way would you expect the electrons to move? Why?

To see the effect of this process on the energy bands, click the Merge button to bring the two semiconductor blocks together and, thus, create a “chip” that makes up the LED. Figure 9-1 illustrates the screen that should appear.

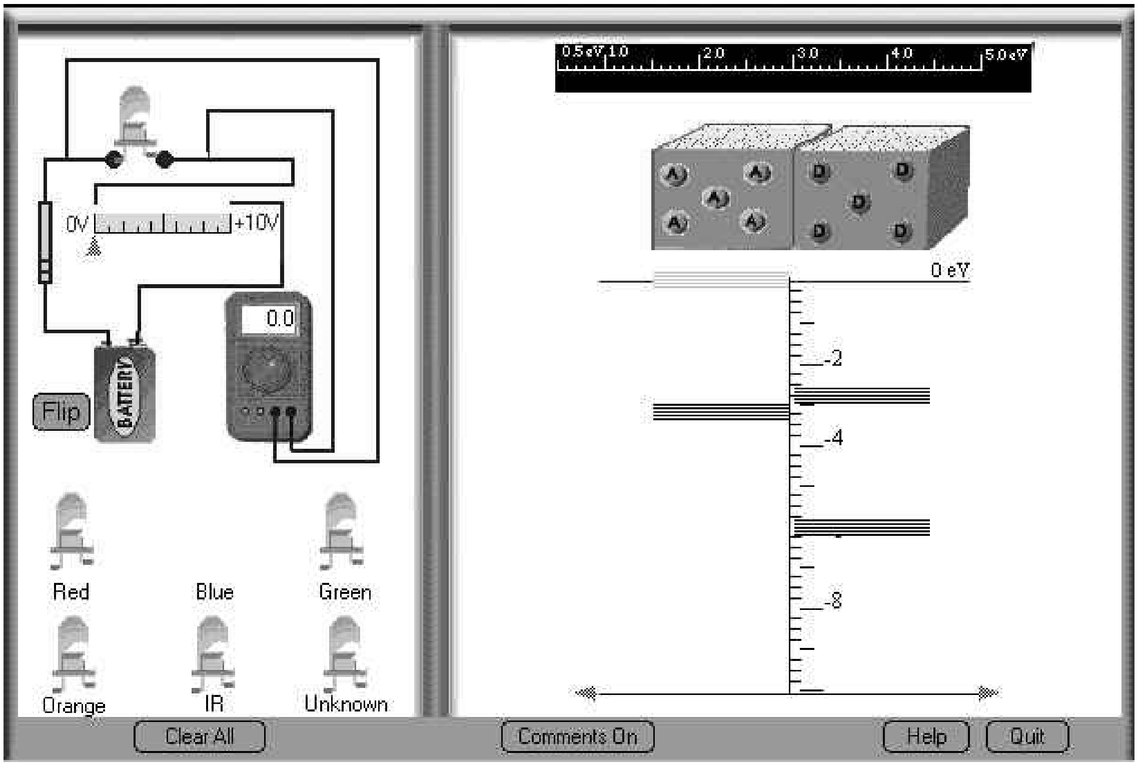

Figure 9-1: The screen of the *LED Constructor* after merging two semiconductor blocks.

As you see, the bands and gaps shift from their usual energies. This shift is a result of the movement of electrons from one material to the other.

Compare the energy diagram for your LED with the diagrams of other students who used different LEDs.

? How are they similar?

? How are they different?

The situation that is now represented on the screen is equivalent to an LED that is not connected to an energy source. The energy diagram for the merged material is similar to Figure 9-2.

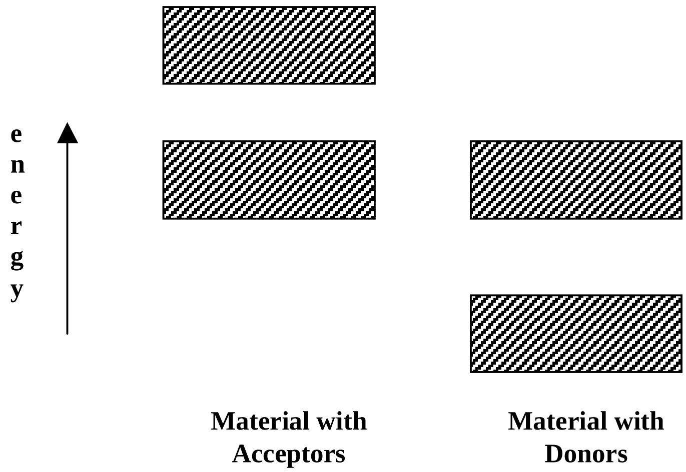

Figure 9-2: Energy level diagram for a chip made of a material with acceptors and a material with donors.

In this situation extra electrons are still available on the right (donor) side. However, they cannot change energies from the conduction band to valence band in the donor side. All of the valence energies are already filled with electrons. We must move the electrons to the other material where the change in energy is easy. In the next activity we will use a battery for this purpose.

Name: Class:

SOLIDS & LIGHT

Visual Quantum Mechanics

ACTIVITY 10

Applying Energy Level Models to LEDs

Goal

The energy levels of two joined materials showed some additional properties that we had not seen previously. By looking at the response of these levels to an applied voltage we will see how light is emitted. As a bonus we will learn why an LED only emits light when the battery is connected in one way.

We will begin where we ended in Activity 9. Start the *LED Constructor*. Then, add acceptors and donors, and merge the materials to create an LED.

Now, we wish to apply a voltage to the LED and see what happens to

- the energy levels in both materials, and
- the light emitted.

Use the slider control (upper left) to vary the voltage (input energy) applied the LED. The energy diagrams show how the energy level responds to the voltage.

? What is the effect on the energy bands as the applied voltage changes from 0.0 Volts to +1.0 Volts?

To apply a negative voltage we need to flip the LED and place it in the circuit with the wires reversed (long lead on the right). Use the flip button to reverse the LED.

? What is the effect on the energy bands as the applied voltage changes from +1.0 Volts to -1.0 Volts by flipping the LED?

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? For which of the two applied voltages (-1.0 Volts or +1.0 Volts) would it be more likely for the electrons to move from the conduction band of the right block to the conduction band of the left block? Explain.

Reverse the LED again by flipping it so that the long lead is on the left.

Move the slider to the minimum voltage necessary for the LED to begin emitting light. Recall we defined this voltage as the *threshold voltage*.

? What happens to the energy bands of the two semiconductor blocks when the threshold voltage is applied to the LED?

Record the threshold voltage displayed on the voltmeter. Indicate the LED that you were assigned.

Record the range of color(s) and energies (in eV) of the output spectrum of your LED.

Compare your results with others in the class.

? How are all results similar?

LED	Threshold Voltage (V)	Energy Gap (eV)	Output Spectral of Range of Energies (eV)

How are they different?

Since the energy bands on the material with acceptor atoms are higher than the bands of the material with donor atoms, the excess, free electrons found in the conduction band of the right block must be supplied additional energy to move to the conduction band of the left block which has a deficiency of electrons. Electrons cannot flow from the left block to the right block because there are not many free electrons on the left block.

When a battery is properly connected across an LED and when the electrical energy (voltage) supplied by the battery is increased, the difference in energy between the right and left energy bands decreases. As the difference in energy decreases, the likelihood of the free electrons moving from the right block to the junction between the blocks increases. When the appropriate voltage (threshold voltage) is applied, the right and left energy bands reach the same energy level, and the free, energetic electrons of the conduction band of the donor material move toward the acceptor material which has fewer electrons. Once there, the energetic free electrons combine with atoms and lose their energy in the form of light emitted by the LED while making the transition to the valence band as shown in Figure 10-1.

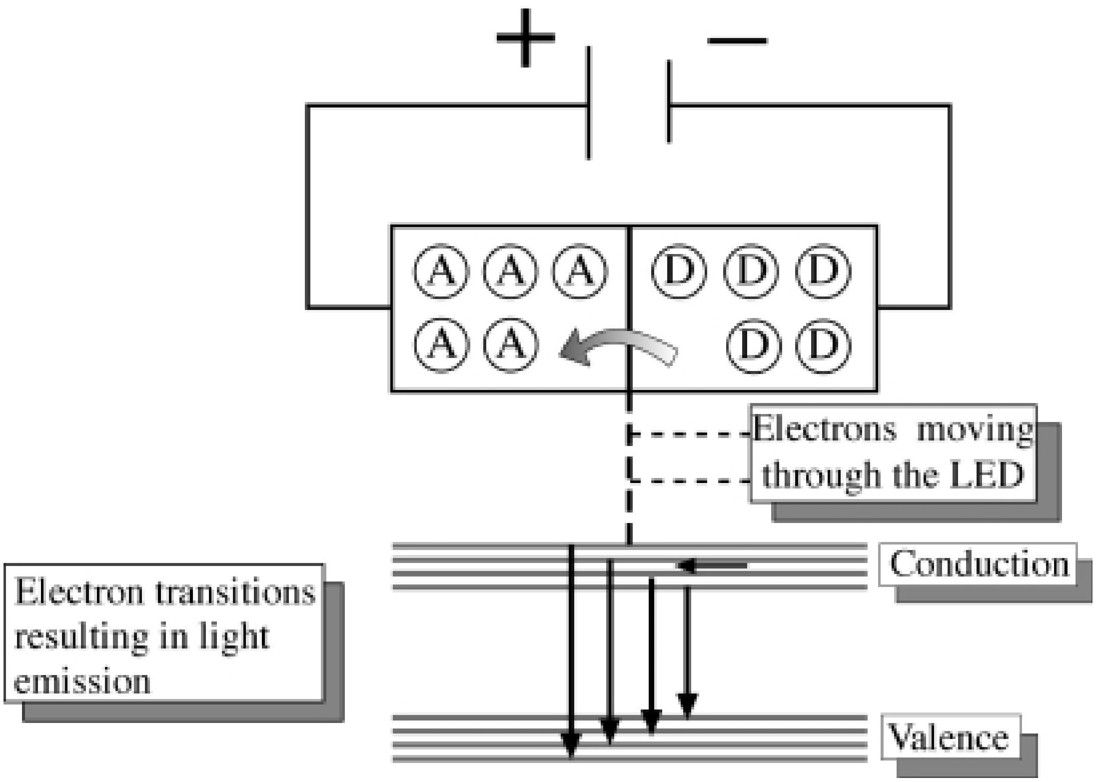

Figure 10-1: Energy Band Diagram of an LED Emitting Light at the Threshold Voltage

The free and excess electrons (which are negatively charged) of the material with donor atoms must be pushed toward the material with acceptor atoms. When a battery is properly connected to an LED, it supplies these excess electrons with sufficient electrical energy to go toward the acceptor side of the chip. Because the negative terminal of the battery repels electrons, the donor side must be connected to the negative terminal of the battery so that the excess electrons are pushed toward the acceptor side.

The short lead of the LEDs is connected to the side of the LED chip that contains the semiconductor with the donor atoms. The long lead connects to the side of the LED chip that contains the semiconductor with acceptor atoms.

When the positive terminal of the battery is connected to the donor side atoms, the free electrons are attracted to the positive terminal of the battery instead of being pushed into the acceptor side. This situation is shown in Figure 10-2. They have insufficient energy to reach the acceptor side. They can only emit light in the junction between the acceptor and donor sides, so no light is emitted.

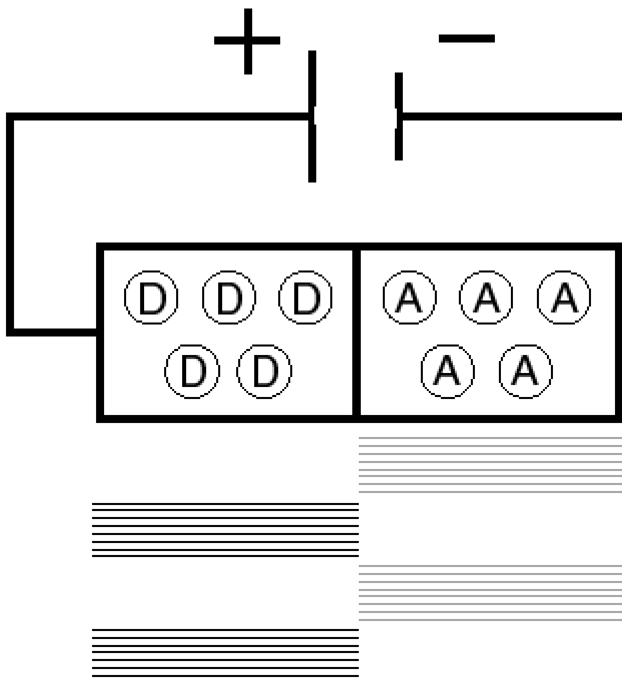

Figure 10-2: Placing the LED in backwards results in the available electrons on the donor side having insufficient energy to move to the acceptor side. No light is emitted.

In our energy level model we have assumed that electrons in a solid acquire and lose energy. Our energy level model of how an LED works, like all models, has limitations. Our energy level model is not perfect and should not be considered a description of all processes that occur in an LED. The model, however, does enable us to explain the general nature of the phenomena that we observed and gives us a general idea of how an LED works at the atomic scale.

We now have gone full circle from exploring the properties of LEDs to constructing an energy level model that explains these properties. The development of this comprehensive model began with atoms of gases to explain the spectral properties of gas lamps and ended with atoms of solids to explain the physical properties of incandescent lamps and LEDs. Along the way we have also learned about gas lamps and the limited number of energies available to an electron in an atom. Knowledge of the energies has been important in almost every area of modern science and technology. One example, the LED and its applications, are important to our daily lives and will play a bigger role as alternate light sources in our future. We could say that the LED and the quantum model that explains how this tiny device works provide us with an “illuminating” experience that allow us to see everyday phenomena in a new “light”.

Homework:

One active research area involves using LEDs for plant growth, especially for space applications where low weight and low power consumption are critical. LEDs are a promising light source for plant-growth experiments in space due to their high intensities and electrical efficiencies, small size and mass, long life, and excellent record of safety and reliability. Several crops – lettuce, spinach, wheat, and potatoes have been successfully grown when illuminated by an array of LEDs supplemented with fluorescent light.

Assume we are members of a group of NASA scientists working on a project that focuses on using LEDs to grow plants in a space station currently orbiting Earth. We must make a decision on what types of LEDs should be used and what the electrical requirements for the batteries to power these LEDs should be. A NASA botanist on our team has collected the following information on the light spectra that would be most conducive for photosynthesis (the process by which plants use light energy to manufacture food). Plants reflect green light but absorb energy in the blue and red regions.

1. What would be the energy gap(s) of the material(s) you would choose to make the two LEDs that would produce the required spectrum? Explain.

2. NASA has to provide batteries to supply the minimal electrical power to these LEDs. If the batteries are too large they will weigh more than necessary and waste valuable fuel. NASA, like any government agency, must be very specific in their request so that bids could be taken on the contract to provide these batteries. What would be the minimum voltage the battery must supply in order to make both LEDs emit light simultaneously? Assume that the LEDs will be connected in such a way that the battery would have to provide a voltage equal to the sum of the threshold voltages for each LED. Explain.

3. Based on our observations using these LEDs, what types of problems can occur when using this set-up?

Name: Class:

SOLIDS & LIGHT

ACTIVITY 11

Can Ohm's Law Explain Your Observations?

Goal

In the first activity, we investigated the electrical and light emitting properties of LEDs and the incandescent lamp. In this activity, we will attempt to explain the electrical properties of the incandescent lamp and the LED.

Because the incandescent lamp and the LEDs are electrical devices, we might assume that these devices obey Ohm's law. For simple electrical components, such as resistors, Ohm's law states that the current flowing through a device is directly proportional to the voltage applied to it. This law is represented by the equation,

Voltage = Current x Resistance

where for many devices the resistance (R) is a constant. We will determine whether Ohm's law with a constant resistance can explain the electrical properties exhibited by the incandescent lamp and the various LEDs investigated in the previous activity.

The best way to determine if an electrical device obeys Ohm's law with constant resistance is by graphing the variation of current (I) with the variation of voltage (V) across the device. A device that obeys Ohm's law exhibits a very characteristic I-V graph. For example, a graph of current (I) versus voltage (V) for a resistor with a given resistance exhibits typical Ohmic behavior as shown in Figure 11-1.

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Note that the current-voltage graph is linear - that is, the slope (change in I divided by a change in V) of the graph is constant. A constant slope indicates that the electrical resistance is constant. Since the slope of the current-voltage graph is equal to $1/R$, the resistance can be determined by calculating the reciprocal of the slope.

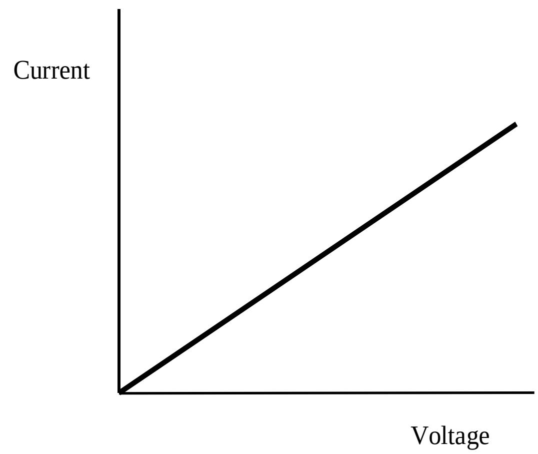

Figure 11-1: I-V Graph of a Resistor

? If you made measurements of current and voltage of a 33 Ω resistor and made a current-voltage graph, is the resulting graph consistent with the general shape of Figure 11-1?

Figure 11-1 can be explained in terms of the electrons in a material. When a voltage is applied to the devices, electrons are attracted to the positive and repelled by the negative side of the battery. They move through the material, and we measure that motion as current.

Because current is directly related to the number of electrons passing a certain location in the circuit in one second, it is related to both the number of electrons and their speeds. The resistance is a property of the material and is related to the availability of electrons and the ease with which they can move.

To determine if incandescent lamps and LEDs behave as simple resistors do, measure the current as a function of voltage of one LED and a Christmas tree light of the same color. Record and plot your data below.

Based on the current-voltage graph for the incandescent lamp, state whether this device appears to obey Ohm's law with constant resistance. Explain your answer.

? How about the LEDs? Do they appear to obey Ohm's law with constant resistance? Explain your answer.

As a starting point in understanding the current-voltage characteristics of the LED, we compare the relation between current and voltage with that of a device that obeys Ohm's law. We find that the LED's and incandescent lamp's variation in current with a change in voltage is quite different from the devices that obey Ohm's law with constant resistance. Our task for the remaining activities in the unit is to understand these differences.

Name: Class:

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ACTIVITY 12

Using LEDs to Measure Planck's Constant

Goal

In this activity, we will use the electrical and spectral data that we have collected from LEDs and apply the concepts we have learned to determine a fundamental constant.

In the previous activities, we learned that the energy gap of the semiconductor that makes up the LED is related to the threshold voltage of the LED and to the energy of the brightest color of light emitted by the LED. An equation expressing this result is:

$$E_{gap} = E_{light} = eV_{threshold} \qquad (12\text{-}1)$$

where E_{gap} is the energy (in eV) of the energy gap of the semiconductor, E_{light} is the energy (in eV) of the brightest color of light emitted by the LED, *e* is the charge of an electron (1.6×10^{-19} Coulombs), and $V_{threshold}$ is the threshold voltage (in Volts) of the LED. Although these energies are actually approximately equal to one another, we will assume they are equal.

Recall that visible light consists of photons which are small packets of energy. The energy of these photons is related to the color of light emitted. The energy of a photon (E_{light}) associated with brightest color of light emitted by the LED is related to its frequency by the Planck-Einstein equation:

$$E_{light} = hf_{light} \qquad (12\text{-}2)$$

where h is Planck's constant and f is the frequency (in Hz or 1/s) of the photon of light. We will not go into the details here, but frequency is a property that allows us to understand the nature of light and is also related to visible light's color.

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Another property that is important in understanding the nature of light is its wavelength. The frequency of the light emitted by the LED is related to its wavelength by the equation:

$$f_{light} = \frac{c}{\lambda_{light}} \qquad (12\text{-}3)$$

where c is the speed of light (3.0×10^{8} m/s), and λ is the wavelength (in meters). When equation (3) is substituted into equation (2), the following equation results:

$$E_{light} = \frac{hc}{\lambda_{light}} \qquad (12\text{-}4)$$

Recall from our earlier investigations in recording spectra that the bottom of the spectral scales that were used had units of nanometers (nm). One nanometer equals 1×10^{-9} meters. This unit is frequently used to describe the wavelengths of visible light. The wavelength of light is just another way to indicate the energy (and color) of a photon that is emitted.

In our investigations with LEDs, we also measured the $V_{threshold}$ for each LED in addition to E_{light}. Since equations (12-1) and (12-2) are related by E_{light}, the relationship between the frequency of the brightest color of light emitted by the LED and the threshold voltage is represented by the equation:

$$eV_{threshold} = hf_{light} = \frac{hc}{\lambda_{light}}$$

Solving for Planck's constant (h), the following equation results:

$$h = \frac{eV_{threshold}}{f} \qquad (12\text{-}5)$$

where e is 1.6×10^{-19} Coulombs, $V_{threshold}$ is in Volts, and f is in 1/seconds. When these units are used, Planck's constant will be in units of Joules x seconds (Js).

Using the data that we collected in the first two activities and the equations found above, record and determine the values for each LED indicated in the following table provided.

LED	E_{light} (eV)	λ_{light}(x 10^{-9} m)	f_{light} (1/s)	$V_{threshold}$ (Volts)

Now use equation (12-4) and our measurements to determine Planck's constant. Construct a graph of the measured threshold voltage versus the peak emission frequency for each LED. Determine the slope of the resulting graph ($\Delta V/\Delta F$). Planck's constant can be calculated by multiplying the resulting slope times the electronic charge, *e*, (1.6 x 10^{-19}).

? What is your measured value for Planck's constant?

? The accepted value for the Planck's constant is 6.626×10^{-34} Js. How does your measured value of Planck's constant compare with the accepted value? Use the following equation to compare these values,

$$\%Error = \frac{|h_{accepted} - h_{measured}|}{h_{accepted}} x100\%$$

? What can explain the difference found between the measured and accept values of h?

Planck's constant is a very important number that helps us determine the quantum nature of both light and matter. German physicist Max Planck who used it to develop a model that explained the photons of light (visible and invisible) emitted first introduced the constant in 1900 by a warm (or hot) body like the sun or the filament of an incandescent lamp. Albert Einstein in 1905 used the constant to develop a theory that explained the phenomenon of electrons being ejected from a metal with kinetic energies determine by the energy of photons of light shined on the metal. Today, Planck's constant is fundamental to any equation of quantum science.