Visual Quantum Mechanics Sampler
AOK-2001

A Collection of Example Materials for Teaching Quantum Mechanics Through Interactive Engagement and Visualization

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The Visual Quantum Mechanics (VQM) project is creating a series of teaching/learning units to introduce quantum physics to several different groups of students. Separate materials are being created for each of the following groups:

- high school students and college students with little background in science or mathematics (VQM-original),
- biology and pre-medical students who are interested in how modern physics is applied to medical instrumentation, and
- science and engineering students who have completed a calculus-based physics course (VQM – The Next Generation).

In this sampler you will see parts of several sets of lessons for teaching quantum physics to the first and third groups. Much of the learning is related to simple devices such as the light emitting diode (LED), gas lamps, and light sticks. You will also see parts of units that discuss one really complex device – the scanning tunneling microscope (STM). This sampler does not contain all of the activities which students would do. Instead you will experience some of the highlights.

 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.

Exploring Light Emission

Emission of light is a quantum effect. To see how we could get to quantum physics from something that seems as straight forward as the light emission as a function of variables such as voltage and temperature for common devices, we will explore the spectra of light emitting diodes (LEDs) and light sticks.

Parts A, B and C below can be done in any order.

A. The Spectra of LEDs as a Function of Voltage

This part of the exploration is adapted from the VQM – Original unit *Solids & Light.*

To save time we will perform this experiment as a cooperative venture. Each group will record the spectra of two LEDs as a function of the voltage applied and investigate briefly the light output from an infrared LED.

An important part of this experiment is being able to see the light emitted by the LED. You might wonder how you are going to see the light from the infrared LED. That is the value of the infrared detection card that is available at your station. Before beginning the experiment try the IR detection card with one of the remote control devices. (TV repair people find these devices very valuable when troubleshooting remote controls.) This device is simple and inexpensive but requires knowledge of quantum physics to understand its operation. We will use it now and return to understanding it later.

The apparatus that we will use during the activity is a circuit board which will be used to explore the variation of brightness with voltage for Light Emitting Diodes (LEDs). Because we do not expect students to know about circuits, we use a fabricated circuit board. The LED can be easily replaced with the incandescent lamp, Christmas tree light, or laser diode with some minor adjustments.

To make sure the apparatus is working properly, insert the incandescent lamp and turn it on. If it does not emit light, turn the knob of the blue potentiometer clockwise. If the lamp still does not emit light, check the connections or ask for help.

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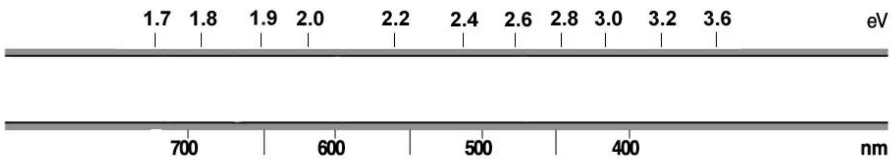
Emission spectra of LEDs

Remove the incandescent lamp and place one of the LEDs in the square, black socket. Use a spectroscope to observe the light emitted by the LEDs. Record your observations for the maximum voltage below.

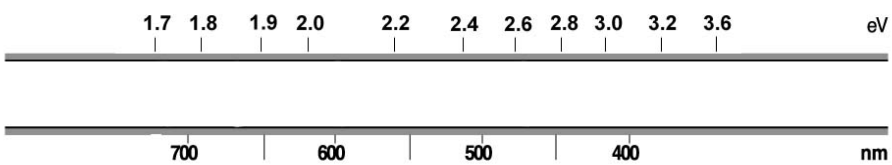
***LED color:*__________________________**



? Adjust the voltage across the LED by turning the knob of the potentiometer and record the changes, if any, in the brightness of the light and in the spectrum with voltage:

? Record the voltage at which the LED begins to emit light: _______________

***LED color:*__________________________**



? Vary the voltage and record the changes, if any, in the brightness of the light and in the spectrum:

? Record the voltage at which the LED begins to emit light: _______________

Infrared LED

We cannot view the spectrum of the infrared LED, but we can see if it is emitting light by using the infrared detector card

? Record the changes, if any, in the infrared emission with voltage:

? Record the voltage at which the infrared LED begins to emit light: ________

An electrical property of LEDs

? Place one of the LEDs in the circuit so that it is emitting the maximum light. Now, reverse the LED in the socket. Describe what happens to the LED.

? Do you obtain similar results with an incandescent lamp?

B. Exploring Photoluminescence

This part of the exploration is adapted from the VQM – Original unit *Luminescence: It's Cool Light*

Now you will explore a different kind of light emission phenomena – **luminescence**. A ***photoluminescent*** object glows as a result of receiving energy from a light source. A ***phosphorescent*** object will glow in the dark even after the lights have been turned off, while a ***fluorescent*** object emits light only in the presence of light. The luminescent objects that you will investigate have been stored in a black plastic bag for a few days.

Place each of the objects on a piece of black paper. Turn off the room lights; turn on a *fluorescent lamp;* and observe the appearance of each object in the light.

? Do the color and appearance of the object change when light strikes it?

? How does the brightness change after the light is turned off?

? How long can you see the light emitted by the object?

Now illuminate each of the objects with a *black light*. In each case answer the following questions.

> *The UV light emitted by some black lights has the potential to cause eye damage. Do not look directly into the black light.*

? Does the color and appearance of the object change when light strikes it?

? How does the brightness change after the light is turned off?

? How long can you see the light emitted by the object?

Based on your observations above, answer the following questions:

? Which of the object(s) exhibits fluorescence and which phosphorescence?

? Describe the characteristics (color, brightness etc.) of light emitted by the *fluorescent* objects.

? What conditions were required for light emission by *fluorescent* objects?

? What effect, if any, did the light source have on the ability of *fluorescent* objects to emit light?

? Describe the characteristics (color, brightness etc.) of light emitted by the *phosphorescent* objects.

? What conditions were required for light emission by *phosphorescent* objects?

? What effect, if any, did the light source have on the ability of *phosphorescent* objects to emit light?

? What are the major similarities and differences between *phosphorescence* and *fluorescence*?

C. Exploring the Temperature Dependence of Light Sticks

This part of the exploration is adapted from the unit *Luminescence: It's Cool Light*

A light stick consists of a sealed plastic tube that contains two solutions. One solution contains a fluorescent dye and a chemical called phenyl oxalate ester. The other solution, dilute hydrogen peroxide, is in a thin glass vial within the plastic tube. A light stick is activated by bending the plastic tube, which causes the glass vial to break so the two solutions can interact. When mixed, the two solutions react and emit visible light. Different groups will investigate light sticks of different colors.

> *The light sticks are non-toxic and are safe to handle. However, the liquid in the sticks can permanently stain clothing and furniture.*

Use three light sticks of the same color – one will be heated on a heating pad, one left at room temperature, and a third cooled. Activate the light sticks and compare the intensities of the light emitted as a function of temperature. Record your results below by answering the following questions.

? How do the intensities of the light sticks vary as a function of temperature?

? How do the spectra of the light sticks vary as a function of temperature?

? Does the time during which the light sticks emit light vary with temperature?

From Observation of Atomic Spectra to an Energy Level Model

Exploring the Spectra of Gases

At the beginning of this part of their study students explore the spectra of several gases. This experiment can be performed by looking at standard gas tubes through spectroscopes or by aiming the spectroscopes at a street lamp. (Even if the students use the gas tubes, we recommend that they look at street lamps and other common light sources as well.) Once they have collected data we ask them to use results similar to theirs to create an energy level diagram of the gases. The following section includes excerpts from two student activities.

For this sampler you may select an introduction to energy level diagrams with minimal mathematics (Version A) or an introduction that leads to Balmer's equation for students with a stronger mathematical background.

Version A
Introducing the Energy Level Diagrams to Non-Science Students

Our transition from viewing spectra to studying atoms uses the following logic: 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.

A useful way to describe the energy of electrons in an atom is to use an energy diagram, which 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 2A.1 represents an energy of -3.4 eV.

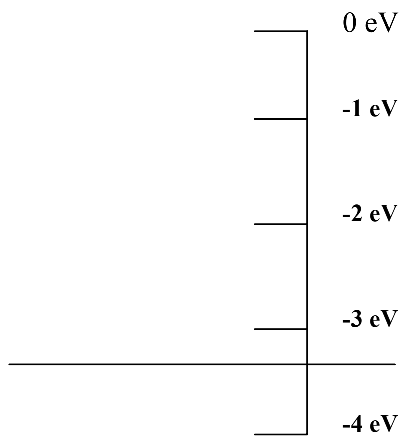


Figure 2A.1: An energy diagram for an electron with -3.4 eV of energy.

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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

Conservation of energy tells us that:

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 the light that they see, the students reach conclusions about the atom. This process allows them to build models of 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. Select *Emission* under *Gas Lamps*. Figure 2A.2 shows the screen that appears. In this program, we can

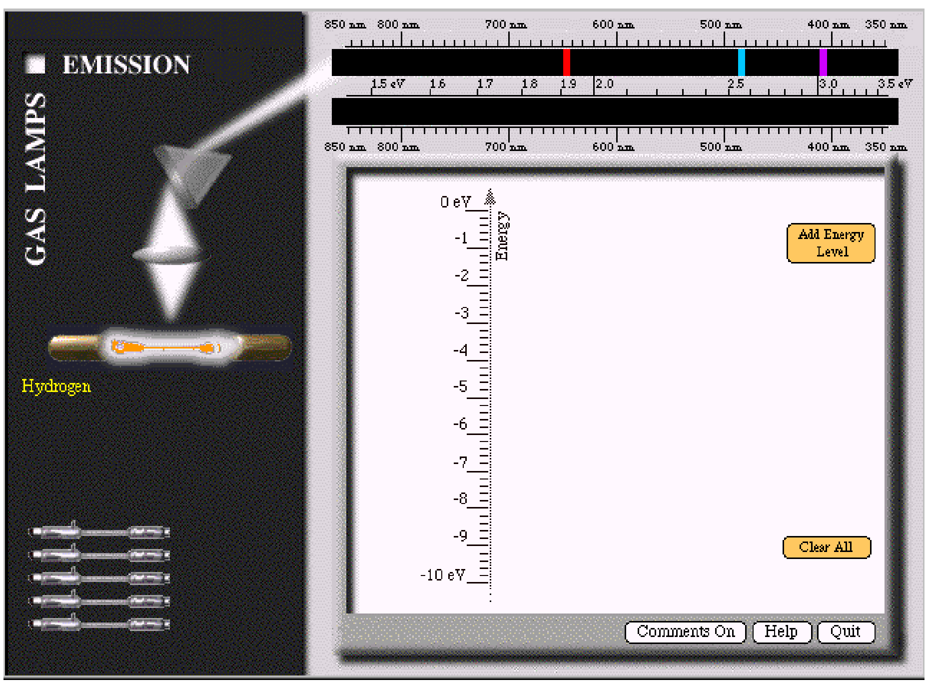

Figure 2A.2: ***Gas Lamp Spectroscopy*** **Computer Program**

- 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 the atom by using the **Add Energy Level** button.
- Move the energy levels added by selecting them to the left of the vertical energy scale and dragging them to the desired position.

Create transitions (represented by vertical arrows) 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.

You may move any of the energy levels after you have created a transition. 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 top black screen 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.

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 help students realize 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 is bound to 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.

Version B
Energy Level Diagrams and an Equation for Spectrum

This is an extract from a self-contained tutorial designed for science and engineering students. This tutorial would be incorporated in a modern physics course to assist student understanding of atomic spectra and atomic structure. We want the students to make connections between their observations of spectra and energy levels.To do this, the students build and test a model of energy levels for the hydrogen atom using visible and UV spectral lines. Once they have identified a working model they determine the relationship between the energy level number (n) and the electron energy (E) and arrive at Balmer's equation.

Goal
We will introduce atomic spectra by focusing on the hydrogen atom, which has the simplest atomic structure. By observing the atomic spectra of the hydrogen atom, we will learn about the limitation on the energy levels of the hydrogen atom. Then, we will look at other process involving light.

Introduction
Because we can not see atoms directly, we must learn about them through indirect observations. For example, light is emitted from gases that are electrically exited. Thus the atoms in the gas are involved in the process of emitting light. In this activity, we will look at this light and see what we can learn about the atom from the properties of the light. We will use Planck's hypothesis to connect wavelength and energy.

$$E = \frac{hc}{\lambda} \qquad (1)$$

Open *Spectroscopy II*. Figure 2B.1 shows the screen that appears.You will see a spectrum at the top of the screen. Use the program to create these levels and transitions and match the hydrogen spectrum. The program will allow you to:

- Increase the number of energy levels with the "add energy level" button.
- Change the energy of a level by grabbing and dragging the level on the left side of the energy scale.
- Make a transition between two energy levels by clicking on one energy level on the right of the energy scale and dragging the arrow to another level until it becomes green.
- Change the difference between two energy levels by dragging one of energy levels on the left of the energy scale.

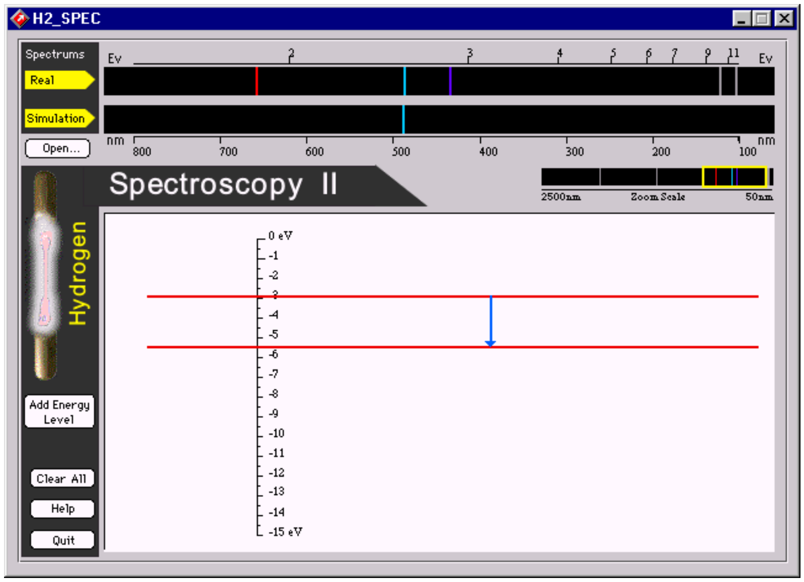

Figure 2B.1: ***Spectroscopy II*** **computer program.**

Start with the visible spectrum. As you create an energy diagram consider the following questions.

- Which transition - energy gain or loss - do we need for emitting the light?
- How many transitions do we need for 3 spectral lines?
- What is the minimum number of energy levels that we need to create these three spectral lines?

We have found a number of different energy level *models* to represent the spectral lines for the hydrogen atom. The common and important feature of these models is the energy differences that correspond to the wavelength of the spectral lines. We have discovered one very important fact-energy states in the atom have only discrete values. We have modeled the energy states of the atom using only our observations of the spectral lines.

Now look at the ultraviolet and infrared lines. If you wish to have a closer view, the yellow window at the top can be used to zoom in on various lines in the spectrum. Just use your mouse to change the size and position of the yellow box.

- How many transitions are needed for 7 spectral lines?
- What is the minimum number of energy levels we need to create these lines?

There should be only one arrangement of energy levels, this is because the hydrogen atom is unique and has specific properties. Decide on an appropriate model for the energy levels of a hydrogen atom using your observations, computer trials and discussions with your neighbors.

Sketch your energy level model for the hydrogen atom:

- What can you say about the energy of each level in your diagram?
- Could this energy be varied?
- What can you say about the difference in energy between two levels?
- Could this difference be varied?

We have found the hydrogen atom can only have certain allowed energies to produce the spectral lines. In our present model of hydrogen the energy levels can be at different values but the differences between the energy levels must be equivalent to the energy of each spectral line.

In a real hydrogen atom the energy levels have specific values that reflect the physical properties of the atom. In this section we will attempt to determine the correct physical energy level model for hydrogen. To distinguish the energy levels, label them in order starting with n= 1 for the lowest energy level.
As we can see, the energy of a level is related to the number n-the higher the n, the higher (closer to zero) the energy.To see if a more analytic relationship exist, complete the following exercises:

- Plot the following four graphs for your energy level diagram: E vs. n, E vs. 1/n, E vs. n^2 and E vs. 1/ n^2.
- Which one has the linear relation? Find the curve fit line.
- What is the slope of your line? Is it negative or positive? What is the intercept?
- Compare your slope and intercept with your neighbor. What is similar? What is different?
- Write down the equation that you have derived to represent the relationship between energy level number and energy. Compare your equation with your neighbors.

In situation such as this one we can set the zero for energy at any value, only the difference between energy levels are important in explaining the light that we see. We have used negative energies to indicate the electron is bound to the atom. So, a convenient choice is that E=0 is the energy at which the electron is no longer bound to the atom. Studies of the electrical energy in atoms show us that an infinite number of energy states are possible. So, n= infinity is an unbound electron and is also E=0.

- Does your equation reflect this assumption?
- Make any changes in your equation so that it reflects $E \to 0$ as $n \to \infty$.
- By adjusting the value (but not the separation) of your energy levels you should obtain an equation which is similar to:

$$E = \frac{-13.6\ eV}{n^2} \qquad (2)$$

where n is 1,2,3..., and E is energy. Because you were working with your experimental values and fitting the curve, your constant may be different from –13.6, which is the approximate presently accepted value.

Emission Series

We will use our observations to develop an equation for the Balmer's series, which applies to the visible spectral lines. Actually Balmer, who was a schoolteacher, discovered that formula only with observing hydrogen spectrum without any knowledge of energy states (Banet 1970). But here we will start with the discrete energy levels we have found and Planck's hypotheses.

The energy of the photon is the same as the energy difference between two energy levels.

$$E_{photon} = \Delta E_{electron} = E_{final} - E_{initial} \qquad (3)$$

- Suppose the electron is in the n^{th} energy state. And n is greater than 2. As we have seen, when the electron moves to the n=2 energy state when it emits visible light. Using equation (2), (3) and Planck's hypothesis, derive the wavelength of the light as a function of n for these transitions.

The visible spectral lines are called the Balmer's series. The relation between energy levels and the wave-length is

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad (4)$$

where n = 3,4,5..., and R = 1.097×10^{7} m^{-1}. Similar equations can be found for other series. In each case a series of the spectral lines arises when electrons make transitions to a particular final energy state. Thus, the general equation is

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right) \qquad (5)$$

where

- m=1 is the Lyman Series and is created by transitions to the ground state,
- m=2 is the Balmer Series and is created by transitions to the second state,
- m=3 is the Paschen Series and is created by transitions to the third state.

While these equations are easy to drive once one knows about the discrete energy levels in atoms, they were much more difficult to determine in Balmer's time, before we had knowledge of how light was emitted in atoms or even firm knowledge that atom exist.

References

Banet, L. (1970). Balmer's manuscripts and the construction of his series. *American Journal of Physics*, **38**(7), 821-828.

Extending the Energy Level Model to LEDs

Now, have the students extend their investigation to solids and begin to understand how LEDs emit light. How do these spectra compare with those of a gas? Then, explore how we might create a spectrum similar to that of an LED.

- Open the *Emission* module of the ***Spectroscopy Lab Suite*** and place the unknown gas tube in the gas lamp socket.
- When the unknown is in the socket, you can create your own spectrum. Click on the **Edit** button and enter energy values which are similar to the spectrum that you observed for an LED.
- Create an energy level diagram for an atom that could produce this spectrum, sketch it below.

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 levels. 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. For the LED, 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*.

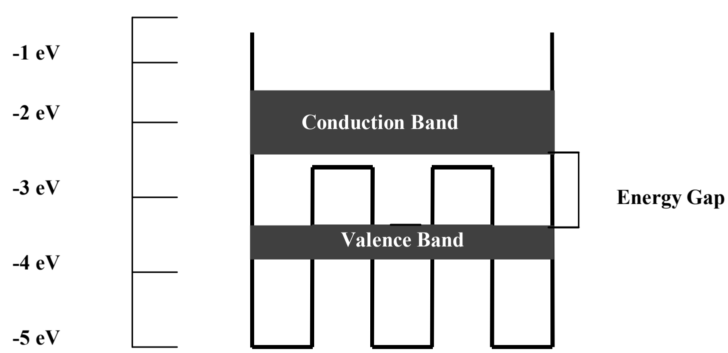

Figure 3-1: Energy diagram with a Very Large Number of Solid Atoms

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.*

No electron energies are allowed between the conduction and valence bands. This range of energies is called the *energy gap*.

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**. Use this program to see how the spectrum of an LED depends on the energies of the bands and the gap. Then, create a set of bands that reproduce the spectrum of an LED that you observed.

Application of Energy Bands to LEDs

This activity is adapted from the unit *Solids & Light*

This section represents a rather lengthy discussion. In class this material can be presented as part of a discussion which frequently refers back to the experiments performed by the students. In addition, interactive use of the LED Constructor *computer program is required for this section. Here we limit our discussion to the LED.*

The energies of the conduction and valence band electrons and the size of the energy gap depend upon the types of the atoms in the solid. For example, we can find solids which have energy diagrams similar to the ones shown in Figure 4-1.

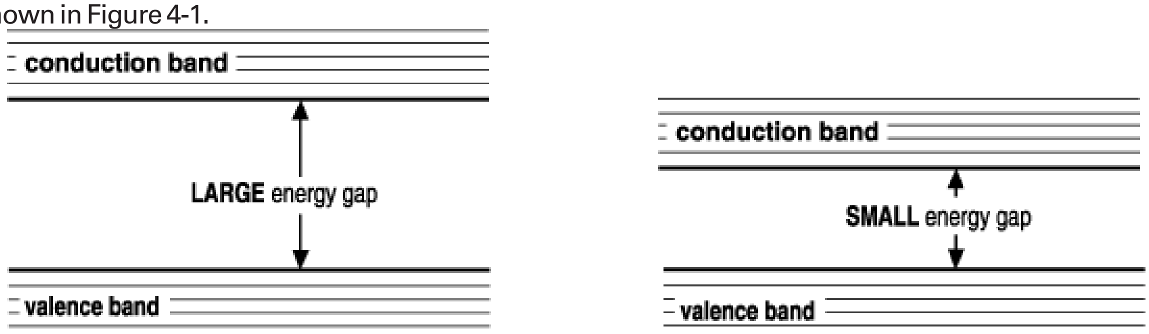

Figure 4-1 : Energy band diagram of two different materials

Modern technology allows manufacturers to control very precisely the types of atoms and thus the energies of the conduction and valence bands.

An LED consists of two types of semiconductor materials adjacent to one another. One of these materials possesses more electrons than the other. We can use the *LED Constructor* computer program to demonstrate how the potential energy looks for such a device. The energy band diagram of these two materials as shown by the *LED Constructor* is seen in Figure 4-2.

Open the *LED Constructor* program and drag one of the LEDs into the circuit. The two blocks at the top right represent the semiconductor materials of the LED. Add acceptor atoms, with fewer electrons, and donor atoms, with more electrons by clicking on the **Add Impurities** button. Now bring these two materials together with the **Merge** button and observe the energy level diagram for the LED.You should see the energy and diagram on the screen, similar to Figure 4-2 without the text.

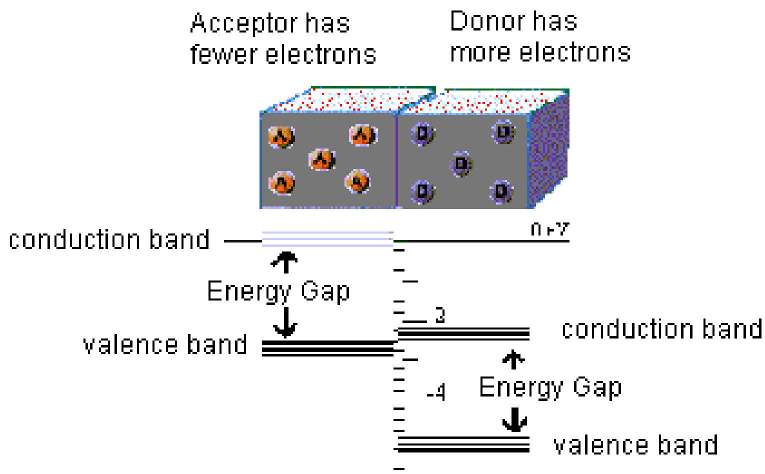

Figure 4-2 : Energy band diagram of an LED with no voltage applied

In keeping with potential energy considerations we expect more electrons to reside in the material on the right because its energy level diagram is lower than the one on the left. For an LED to emit light, electrons must flow across it (we will soon explain why this is necessary). When a battery is connected across the LED, it provides the necessary impetus for the electrons to flow from one side to another. There are two possible ways in which a battery can be connected as seen in Figure 4-3 and Figure 4-4.

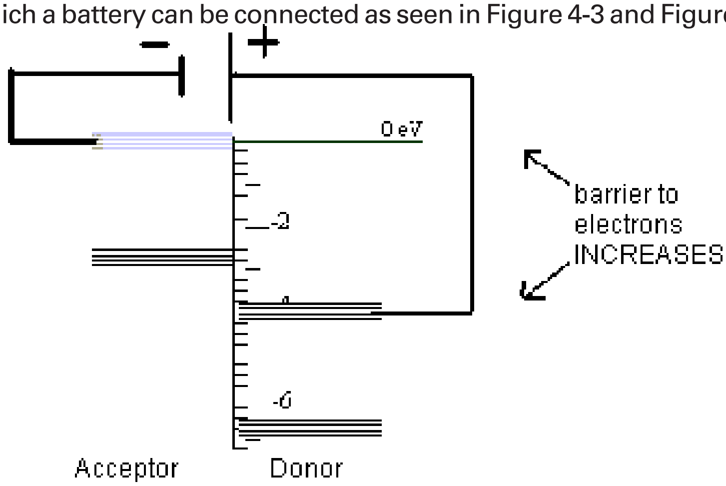

Figure 4-3 : When the *positive* terminal of battery is connected to the side with more electrons, the "barrier" experienced by the electrons *increases*.

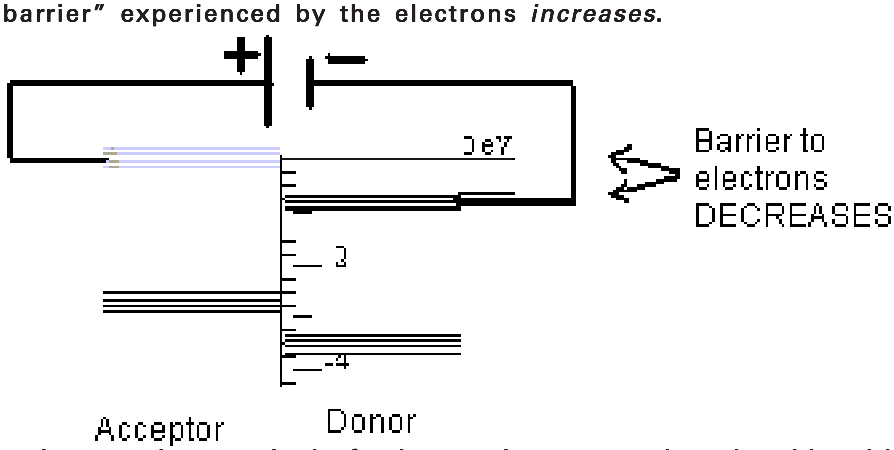

Figure 4-4: When the negative terminal of a battery is connected to the side with more electrons, the "barrier" experienced by electrons *decreases*.

Apply a voltage of about 1 volt to the LED by dragging the slider above the battery. The LED is originally connected with the negative side of the battery connected to the donor side of the LED, shown in Figure 4-4. Now click on the **Flip** button on the left side of the screen. This reverses the LED in its socket so that the donor side of the LED is now connected to the positive side of the battery.

When the side of the LED with more electrons is connected to the positive terminal of the battery, the negative electrons get attracted to the positive terminal of the battery. Consequently, they would tend *not* to move across the LED, or the "barrier" they experience in moving to the other side of the LED increases. When the side of the LED with more electrons is connected to the negative terminal of the battery, however, the negative terminal of the battery repels the negative electrons. They would hence tend to move towards the other side of the LED i.e. the "barrier" they experience in doing so would decrease.

Flip the LED again and increase the voltage from the battery until your LED lights up.

? How do the spectrum shown and the threshold voltage for this LED compare with the real LED that you observed in the first activity?

The battery with its negative terminal connected to the right end of the LED in Figure 4-4 supplies electrons on the right side with sufficient energy so that they can flow over to the left. When these electrons reach the other side they collide with atoms, knocking out their electrons. The electrons that are knocked out of their atoms are free to wander about (i.e. they are in the conduction band). They eventually lose their additional energy (supplied by the collisions with the electrons that moved over) and make a transition to the valence band emitting light as shown in Figure 4-5. Thus, for the LED to emit light electrons must flow from the right (where they are more abundant) to the left (where they are less abundant). Only the proper polarity of the battery (as shown in Figure 4-4) can facilitate that movement.

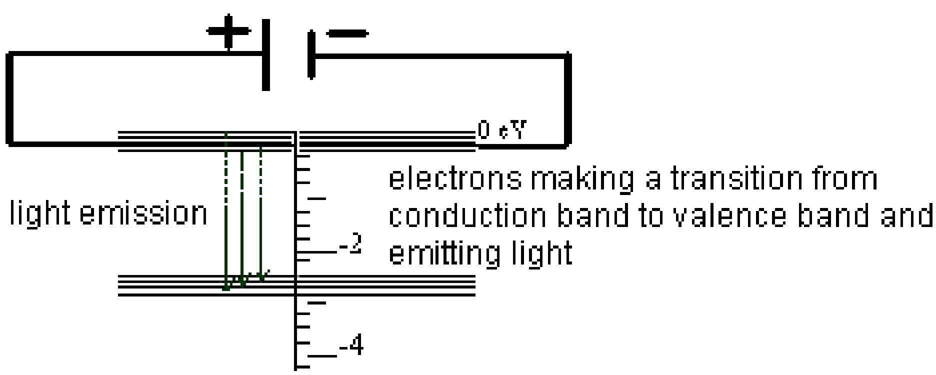

Figure 4-5: Energy band diagram of an LED emitting light

Putting It All Together

This activity is adapted from the unit *Solids & Light*

In this section we bring the various pieces of information together and apply the quantum model of a solid to the observations that we have made.

In our explorations of solid state devices that emit light we have noticed three important characteristics:

1. The color of light emitted by the LED is related to the voltage at which this abrupt change occurs.
2. The electrons in the LED flow preferentially in only one direction.
3. The spectrum of light of an LED is continuous but covers only a fraction of the visible spectrum.

These observations apparently contradict classical physics as applied to electrons moving in the device but are consistent with the model of discrete electron energies in an atom and a quantum physics model that leads to energy bands when many atoms are in close proximity.

? Use these ideas and the simulations to describe how the quantum model of solids explains each of the observations listed above.

In this sampler you have explored light sources in addition to LEDs. Activities similar to the ones we have completed here help students build energy band and gap models for other light emitting processes. You may get a flavor for these activities by working with other components of *Spectroscopy Lab Suite*.

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