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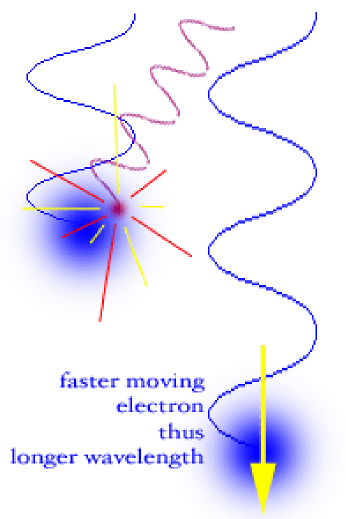
Return

— **Atomic Structure**
—— The Nature of ...
——— ... electrons and energy.

Energy and Electrons

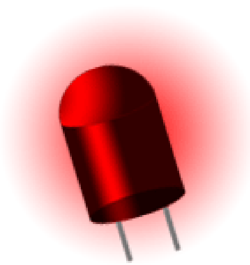


When an electron is hit by a photon of light, it absorbs the quanta of energy the photon was carrying and moves to a higher energy state.

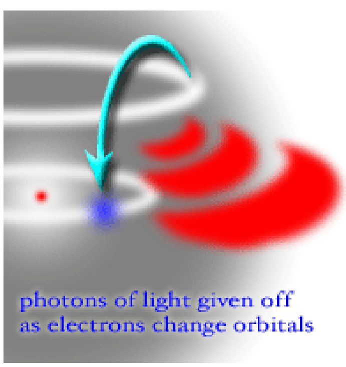
One way of thinking about this higher energy state is to imagine that the electron is now moving faster, (it has just been "hit" by a rapidly moving photon). But if the velocity of the electron is now greater, it's wavelength must also have changed, so it can no long stay in the original orbital where the original wavelength was perfect for that orbital-shape.

So the electron moves to a different orbital where once again its own wavelength is in phase with its self.

Electrons therefore have to jump around within the atom as they either gain or lose energy. This property of electrons, and the energy they absorb or give off, can be put to an every day use.



Almost any electronic device you buy these days comes with one or more Light Emitting Diodes (usually called "**LEDs**"). These are tiny bubbles of epoxy or plastic with two wire connectors. When electricity is passed through the diode it glows with a characteristic color telling you that the device is working, switched on and ready to do it's work.



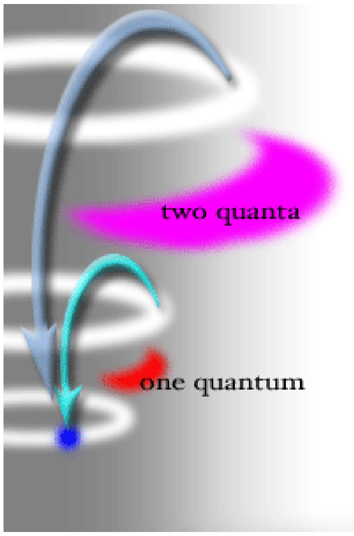
Deep in the semiconductor materials of the LED are "impurities", materials such as aluminum, gallium, indium and phosphide. When properly stimulated, electrons in these materials move from a lower level of energy up to a higher level of energy and occupy a different orbital.

Then, at some point, these higher energy electrons give up their "extra" energy in the form of a photon of light, and fall back down to their original energy level. The light that has suddenly been produced rushes away from the electron, atom and the LED to color our world.

Typically, the light produced by a LED is only one color (red or green being strong favorites). Although they are cheap, easy to make, don't cost a lot to run, LEDs are **not** usually used to light a room, because they cannot normally produce the wide range of different colors needed in "white" light.

This is because of the quantum nature of the atoms being used in the LED and the quantum energies of the electrons within them.

When an excited electron within a LED gives up energy it must do so in those lumps called **quanta**. These are fixed packets of energy that cannot be changed or used in fractions; they must always be transferred in whole amounts.



Thus, an excited electron has no option but to give off either 1 quanta or 2 quanta of energy, it cannot give up 1.5 quanta, or 2.3 quanta. Also, the electron can only move to very limited orbitals within the atom; it must end up in an orbital where the wavelength is now uses is "in phase" with itself. These two restrictions limit the quality of the quanta of energy being released by the electron, and thus the nature of the photon of light that rushes away from the LED.

Since the energy given off is strongly restricted to quanta, and quanta that allow the electron to move to a suitable place within the atom, the photons of light are similarly restricted to a tiny range of values of wavelength and frequency (a property we see as "color").

Many LEDs have electrons that can only give up quanta of energy that, when converted into photons, produce light with a wavelength of about 700 nm - which we then see as **red** light. These electrons are so restricted in the quanta they can emit that they never shine blue light, or green light, or yellow light, only red light.

Lines in Spectra

Long, long before their were LEDs in our lives, scientists trying to understand electrons in atoms noted a similar phenomenon when light was either shone on certain materials or given off by certain materials.

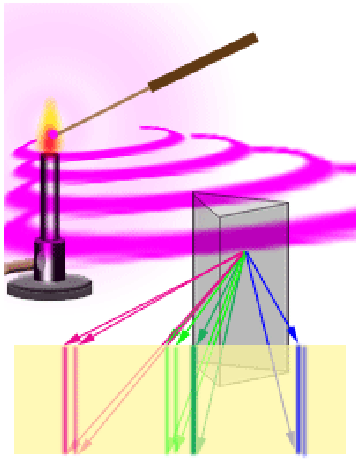
In 1859 the German physicist Gustav Robert Kirchoff, and his older friend Robert Wilhelm Bunsen came up with a clever idea. They used Bunsen's burner to strongly heat tiny pieces of various materials and minerals until they were so hot that they glowed and gave off light.

Sodium, for example, when heated to incandescence, produced a strong yellow light, but no blue, green or red. Potassium glowed with a dim sort of violet light, and mercury with a horrible green light but no red or yellow.

When Kirchoff passed the emitted light through a prism it separated out into its various wavelengths (the same way a rainbow effect is produced when white light is used), and he got a shock. He could only see a few thin lines of light in very specific places and often spread far apart.

Clearly glowing sodium was not producing anywhere near all the different wavelengths of white light, in fact it was only producing a very characteristic band of light in the yellow region of the spectrum - just like a LED!

Kirchoff and Bunsen carefully measured the number and position of all the spectral lines they saw given off by a whole range of materials. These were called ***emission spectra***, and when they had collected enough of them it was clear that each substance produced a very



spectral lines

characteristic line spectrum that was unique. No two substances produced exactly the same series of lines, and if two different materials were combined they collectively gave off **all** the lines produced by both substances.

This, thought Kirchoff and Bunsen, would be a good way of identifying substances in mixtures or in materials that needed to be analyzed. So they did. In 1859 they found a spectrum of lines that they had never seen before, and which did not correspond to any known substance, so, quite rightly, they deduced that they had found a new element, which they called **cesium** from the Latin word meaning "sky blue". (Guess in what part of the spectrum they found the lines!).

Quantum Numbers and Levels of Energy

All the research on atomic structure and the hideously difficult-to-understand properties of electrons come together in the topic of "electron energy".

An atom such as lithium has three electrons in various orbitals surrounding the atomic center. These electrons can be bombarded with energy and if they absorb enough of the quanta of energy being transferred they jump about and in the most extreme case, leave the lithium atom completely. This is called **ionization**.

The amount of energy needed to remove the first electron from a lithium is 124 kilocalories/mole, an amount of energy that is not difficult to supply, so lithium atoms ionize easily.

However, it takes almost 1740 kilocalories/mole of energy to dislodge the second electron from around the lithium ion (it is now an "ion" because it has already lost one electron). It takes a massive 2820 kilocalories/mole to dislodge the third and final electron from around the lithium ion.

Partly this difference in the amount of energy needed to dislodge different electrons away from the lithium atomic center is due to the fact that the center of the lithium atom is carrying the positive charges of three protons. Moving a negatively charged electron away from a positively charged atomic center needs more and more energy as the amount of un-neutralized charge increases, thus;

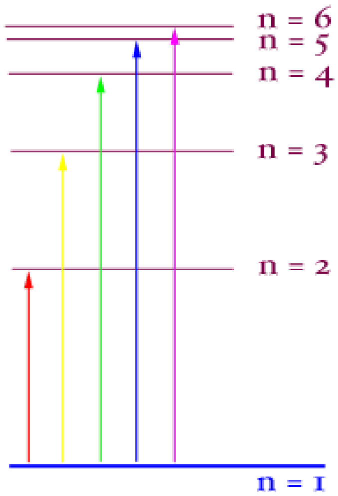
$$Li \dashrightarrow Li^{+} + e^{-}$$

$$Li^{+} \dashrightarrow Li^{++} + e^{-}$$

$$Li^{++} \dashrightarrow Li^{+++} + e^{-}$$

However, the amount of energy needed to remove the first electron is a good measure of what it takes to stimulate an electron to leave its atom, and how tightly it is held there in the first place.

Within the atom, as Bohr pointed out, there are different possible positions for electrons to be found as defined by the ***principal quantum number***, usually written as "**n**".



Bohr defined the energy of electrons located at these different locations of quantum state by the formula:

$$E_n = - E_o/n^2$$

In this formula E_o is a whole collection of physical constants, which for an atom such as hydrogen has a value of 313 kilocalories/mole. Using this formula it is possible to calculate how much energy an electron has at each of the other, different, quantum states (n = 2, n = 3, n = 4, etc.). This is usually presented in the form of a diagram (see left).

For an electron at the ground state (n = 1) to be moved up to the next level (n = 2) it must absorb a quantum of energy that is the perfect amount to make this move. If the quantum is too small the electron could not reach the next level, so it doesn't try. If the quantum is too large the electrons would overshoot the next level, so again, it does not try. Only quanta of exactly the right size will be absorbed and used.

Similarly, if an electron is already at the second level (n = 2), and there is a space for the electron at the lower level (n = 1), it can release a quantum of energy and drop down to the lower level. But the amount of energy given off will be a whole number quantum. If this energy is given off as light (such as happens with emission spectra) then the photons rushing away from the falling electron will be of only one size and quality (color). Hence glowing sodium, or LEDs, only give off very discrete bands of light with distinct colors or bands within their spectrum.

All this implies that if white light (with all the possible wavelengths, colors and possible quanta of energy) is shone on certain materials or substances only certain wavelengths (and their quanta of energy) will be absorbed by the electrons in that substance. Only a narrow band of light will have just the right quanta to move an electron to the next level, or the level above that, and so on.

That wavelength will be taken out of the spectrum of light and leave a dark band of no-light behind. Absorption spectroscopy, therefore, is the equal and opposite of emission spectroscopy. However, in both kinds, it is the absorption of quanta to move electrons, or the emission of quanta to move electrons around in the atom that is the reason why only certain wavelengths of light are affected.

The Quantum Atom - - a Summary

Although Bohr's original picture of a quantum atom has been modified in the years since he first proposed the concept, never the less, the main principles still stand:

1. Electrons are to be found occupying certain volumes of space around an atomic center

("nucleus") - these volumes of space are called **orbitals**

2. An electron in an orbital has a defined wavelength. The actual wavelength can be determined using the de Broglie formula "wavelength = Plank constant / momentum.

3. The shape and location of the orbital is determined by the fact that the only stable shapes and locations are those where the electrons (acting as waves) can have a number of waves that are whole numbers (technically these are called "**standing waves**"). Standing-wave orbitals are the only ones in which the occupying electrons do not either radiate energy, or collapse.

4. The energy carried by electrons has to be a whole number of quanta of energy as given by the formula $\mathbf{E_n = - E_o/n^2}$

   where "n" is the principal quantum number. The energy of an electron, and the atom that carries it, is therefore restricted, or ***quantized,*** to a limited number of values.

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