Thoughts for the Day
CH301 Fall 2010
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Things that classical mechanics can't explain. Two important examples

Light

Light is an electro-magnetic wave. For our discussion it is only important to think about the electric field part. This is why light interacts with the electrons in the material.

Key ideas about waves

Amplitude	height of the wave
Wavelength	distance from peak to peak
Speed	speed the peaks are traveling
Frequency	How many peaks a second

The speed of light in a vacuum (interacting with nothing) is a constant of the universe. It is by definition

$$c = 2.99792458 \times 10^{8} \text{ m s}^{-1}$$

This is how a meter is defined. It is denoted by a small c.

Wavelength, frequency, and speed are related

$$c = \lambda\nu$$

Where the wavelength is given the symbol λ and the frequency the symbol ν

Since c is fixed, either the wavelength or the frequency is sufficient to describe the light. One can be found easily from the other. Different wavelengths (frequencies) correspond to different colors that we can see as well as light that we cannot see.

See the diagram in the book for the full range of the EM spectrum.

1. Blackbody radiation.

Planck's model could reproduce the experimental data by assuming that electrons (oscillators) in the material could only have particular energies that were proportional to a constant (Planck's Constant) $h = 6.626 \times 10^{-34}$ J s. He got the value from fitting the temperature dependent black body radiation curves

Photoelectric Effect.

Classically we would associate the energy of light with its amplitude. Big wave = lots of energy. Small wave = little energy.

However, with the photoelectric effect this doesn't appear to be the case. In this experiment you shine light on a piece of metal and sometime electrons come flying off. The idea is a simple one. The electric field of light causes the electrons to oscillate. If they are oscillating with enough energy the can escape the metal (if they have more energy than the potential energy holding them to the metal. It this were an isolated atom we would call this amount of energy the ionization energy). However, we discover something odd. The electrons come off only when a certain color or light issued. Or more specifically only when the light has a frequency equal to or greater than a threshold frequency. If the frequency is too low, we get nothing. If the frequency is at or above the threshold we see electrons. More interesting yet, as the frequency gets higher and higher than the threshold we see the electrons come off with more and more kinetic energy. From this we decide the energy of the light is proportional to its frequency. All of the energy in the light seems to be divided up into little bundles, or photons, each with an energy related to its frequency. The energy is directly proportional to the frequency and the proportionality constant is h, Planck's constant. Thus $E = hv$. That is the energy of the light seems to come in little bundles each of which can interact with one electron. The energy of the photon is hv. If this is enough energy to free the electron from the metal then it does. If it is not, then it doesn't even if there are lots and lots of photons. Think about it as one phonon interacts with one electron. Does the amplitude matter. Yes. Lots of light = lots of photons = interactions with lots of electrons. Therefore if the frequency is high enough, then the electrons will come off. More light more electrons. Let's have a quiz

1. You shine green (500 nm) light on a particular metal and see that electrons are emitted from the surface.

What will happen if you use a higher intensity (brighter) green light source?

More electrons will be emitted from the surface

What will happen is you used blue light instead of green?

The electrons will come off with a higher kinetic energy

Line Spectra for atoms

If you look at the light emitted from excited atoms or absorbed by atoms you notice that each element has a unique set of lines and that the colors are distinct wavelengths. This indicates that the electrons in the atoms are changing in energy by discrete amounts (remember we can relate energy and frequency).

So we go in the lab and look at lots of spectra. Very complicated.
We look at H. Why? It is simple and very important. Key in the sun and it has a simple spectrum as it has only one electron.

Take all the wavelengths for the transitions in the visible, UV, and IR and put them together and we see a trend. It looks like there are discrete energy levels in the atom that the transitions all match energies for moving between that fixed set of levels. The energy of level appears to be related to 1 divided the "number" of the level squared.

The Rhydberg formula relates the difference in these levels to the energy of the light of the transition

$$hv = constant\left[\frac{1}{n_2^2} - \frac{1}{n_1^2}\right]$$

for the Balmer lines in the which $n_2 = 2$
this can be written as

$$v = \left[\frac{1}{4} - \frac{1}{n_1^2}\right]3.29x10^{15}s^{-1}$$

Now we need to figure out why there are such levels.

The Bohr model for an atom

Bohr's model attempt to come up with a model for the atom that explains line spectra.

We need to take a great deal of algebra and condense into a few key ideas. I will try to do so here rather than repeating the derivations of formulas from class or the book.

1. First key idea. Electrons in atoms appear to have discrete (not continuous) energy levels as deduced form their line spectra.

2. How to explain this. Look at the H-atom. Simple. One electron. The line spectra can all be explained by a simple formula that implies a series of energy levels. The energy is proportional to $-1/n^2$ where n is an integer that labels the energy level. The lowest energy (most negative) level is n=1. The energy levels increase in energy as n increase until the limit of n=infinity for which the E=0.

3. Where do these levels come from? Enter Bohr's model of the atom. What does it do? Takes classical mechanics and invokes the idea that there are fixed orbits (or fixed angular momenta) for the electron.

Stable orbits implies a relationship between the angular momentum and the distance

We can write the energy in terms of the angular momentum L and the distance r
These can be combined to make on formula that shows that energy is a function of
the radius. Or we can write the energy as a function of the angular momentum.

THIS IS THE KEY NOTION. When we write the energy as a function of L we find that energy is equal to a constant times $1/L^2$. Therefore if the angular momentum comes in integer choices then the energy will have fixed levels that go as $1/n^2$.

Bohr postulated that the angular moment is fixed and has only limited values that are allowed.

$L = nh/2\pi \quad n = 1, 2, 3, 4, 5\ldots.$

From this the energy could be written as a function of fundamental constants and n

$E_n = -2.18 \times 10^{-18}\ \text{J}\ (Z^2/n^2)$ where Z is the charge on the nucleus and n is the level number.

What else can we do. We can see the energy goes as $1/n^2$ but the radius is proportionally to n^2. We can calculate the size of the atom! Big stuff.

What is the bad news. Bohr is wrong as the electron is not classical and therefore the solution to the problem is inherently quantum mechanical. Where is he right? In the idea of quantization of states of the electron.