Chapter 13 Notes: Photons and Atomic Spectra

How are the e-m waves emitted by atoms related to energy levels of electrons?

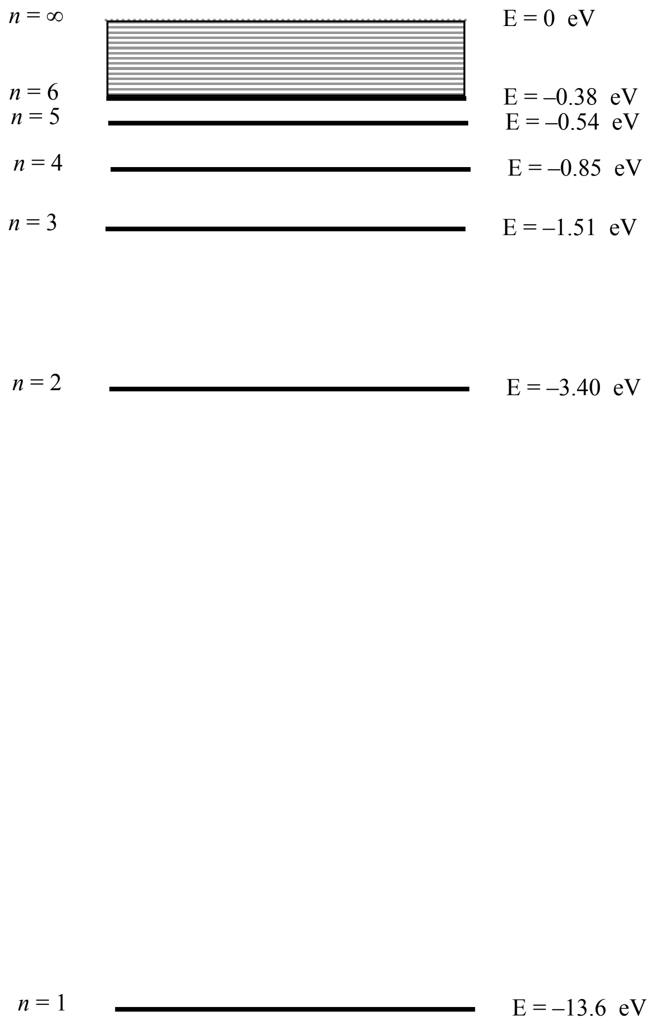
Around 1900 it began to become clear that the principles of electromagnetism as understood at that time were not adequate to explain all observable phenomena. In particular, in order to obtain consistent explanations of the spectrum of e-m waves emitted by hot substances, some very unexpected assumptions were required. According to Maxwell's theory, e-m waves could have any desired energy, high or low (at least in principle). You could simply reduce the magnitude of the electric and magnetic fields as low as necessary in order to obtain e-m waves with energies as low as you desired. To do this, you could for instance produce very small amplitude oscillations of the charges that emitted the waves.

However, in 1900 German physicist Max Planck had proposed that the energies of these oscillators could ***not*** be as small as desired. Instead, they behaved as if there were a lower limit to their energies, and the value of this lower limit depended on the ***frequency*** of the oscillation according to this equation: ***E = hf***, where f was the frequency of the oscillator and h was a constant that later became known as "Planck's constant." ($h = 6.63 \times 10^{-34}$ J-s.) This appeared to be the "minimum package size" for the energy of the oscillating charges. In 1905, Einstein proposed that the e-m waves themselves could only be produced with energies of this "minimum package size." For instance, the lowest possible energy with which one could observe an e-m wave of 1000 Hz was $E = hf = (6.63 \times 10^{-34}$ J-s) $\times$ (1000 Hz) $= 6.63 \times 10^{-31}$ J. This is a very tiny amount of energy, but it was still very shocking that there might be ***any*** sort of a minimum energy value. This "minimum package" of e-m waves came to be called a ***"photon."*** So, the energy of a photon of e-m waves of a particular frequency could be found from the equation $E = hf$.

Around this same time, there was an intensive effort underway to understand the ***spectra*** of e-m waves emitted by gas atoms that had been struck by fast-moving electrons. It had been found that if one filled a glass tube with a low-pressure gas of any type, and then caused high-velocity electrons to race through the tube (by imposing a strong electric field on the tube), the gases would emit very distinctive "spectra" of e-m waves with various frequencies. These could be observed by a "spectroscope," which is a fancy version of a prism that turns white light into a rainbow (as discussed in Chapter 12). When the light emitted by a gas is passed through a spectroscope, a series of brightly colored lines is observed, with a different pattern of lines – different numbers of lines of different colors – characteristic of each different type of gas atom. These "emission spectra" are so distinctive they are a sort of fingerprint, which can be used to positively identify practically any substance (in gaseous form) by the spectrum of emitted e-m waves. Before 1913, there was essentially no understanding of how these spectra came to be produced.

At this time, the Danish physicist Niels Bohr proposed a model of atomic structure which provided some important clues as to how this process might work. Over the next 12 years his model underwent radical revisions by many physicists, but some of the basic mechanisms he proposed are still considered to be reasonably accurate. Bohr proposed that the electrons in any given atom could not, as previously believed, have ***any*** desired amount of energy. In a manner consistent with the ideas of Planck and Einstein, Bohr suggested that in each atom the electrons could only exist in a very limited and specific set of different "energy levels." This simply meant that the electrons could have those values of energy, and no others. Why this should be the case was not understood at that time. (Later on, explanations for this phenomenon would be found that produced very consistent results.)

Bohr considered the hydrogen atom in detail; with only one electron, hydrogen has the simplest structure of all atoms. By building on the observations and theories of other workers, Bohr came to propose that the electron in hydrogen could only take on a specific set of different energy values, of which the first six may be represented by the "energy level diagram" shown on the next page.



In the diagram above we have represented some allowed energy levels of the electron in a hydrogen atom. Ordinarily, the electron is in the "ground state" (n = 1) and has energy of –13.6 eV. (1 eV = 1.6×10^{-19} J.) This means that you would have to ***add*** 13.6 eV of energy to the electron in order to break it away from the hydrogen atom. (If the electron just barely escaped from the atom – and had no energy left over to go anywhere – we would say that it has "zero" total energy, or E = 0 eV.) If an electron in the ground state (n = 1) ***gains*** energy its energy level would rise, for instance to n = 2, n =3, or some higher value of n. The ***larger*** the value of n, the ***greater*** is the energy of the electron. The ***maximum*** possible energy for an electron in the hydrogen atom (E ≈ 0 eV) would correspond to $n = \infty$.

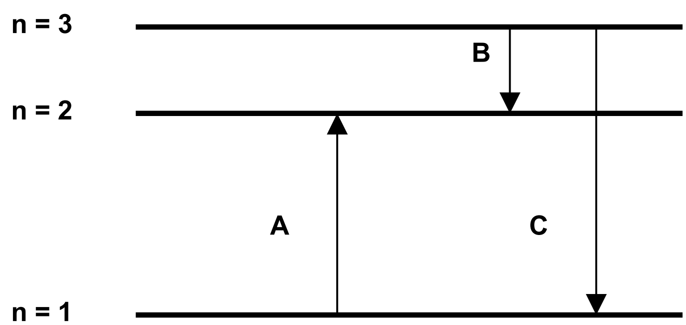
With this model of the energy levels in hydrogen, the origin of the bright lines in the emission spectrum of hydrogen gas could now be explained. This is how it works: when fast-moving electrons are sent flowing through a tube of hydrogen gas, some of them collide with the hydrogen atoms and increase the energy of the electrons in the atoms. In this way, the hydrogen electrons may briefly acquire some of the higher energy levels shown in this diagram. Usually, after a brief moment, these electrons will abruptly lose at least some of their extra energy. When they do this, we say they "drop" in energy level. In this process, they can emit ***one single*** electromagnetic photon – sort of a "particle" composed of e-m waves. ***The energy of the emitted photon must exactly equal the amount of energy that is lost by the electron as it changes its energy level.***

When observing hydrogen gas in an electrical discharge tube (in which high-energy electrons are passed through the gas), the hydrogen glows with a colorful light. When the light is examined through a specially designed filter (a spectroscopic "grating"), it is possible to identify the wavelengths of the individual electromagnetic waves that make up that glowing light. Included among these is an e-m wave with a wavelength of 656.3 nm (1 nm = 1×10^{-9} m). We can identify the electronic transition in the hydrogen atom that corresponds to this e-m photon. *(Here we'll need to recall that the energy of a photon of frequency* f *is given by* E = hf; *Planck's constant* h *has the value* h = 6.63×10^{-34} J-s = 4.14×10^{-15} eV-s.)

We'll use the notation ΔE_{12} to represent the *absolute value* of ($E_{n=2} - E_{n=1}$); this is the energy ***change*** of an electron that makes a transition between level 1 and level 2. Then, according to Planck's formula, $\Delta E_{12} \div h$ = [frequency of e-m wave emitted when electron drops from $n = 2$ to $n = 1$]. The question is, how does the energy level n of the electron have to change in order to emit light with a wavelength of 656.3 nm?

Now, the frequency of light with a wavelength of 656.3 nm can be found from the relationship $c = f\lambda$. We get: $f = c/\lambda = (3 \times 10^8 \text{ m/s}) \div (6.56 \times 10^{-7} \text{ m}) = 4.57 \times 10^{14}$ Hz. Now, the energy of this photon can be found from the equation E = h*f*: $E = hf = (4.14 \times 10^{-15} \text{ eV-s}) \times (4.57 \times 10^{8} \text{ Hz}) = 1.89$ eV. This must be equal to the ***change*** in energy level of the electron that emits this wave. By looking over the energy level diagram, we can see that ΔE_{23} = 1.89 eV; this tells us that an electron dropping down ***from*** the $n = 3$ level ***to*** the $n = 2$ level emits an e-m wave with a wavelength of 656.3 nm.

Here is a schematic energy level diagram of some unknown atom.



As in the diagram of the hydrogen atom, the spacing between the lines is proportional to the ***differences*** in energy levels represented by the lines. So, the spacing between $n = 1$ and $n = 2$ is larger than that between $n = 2$ and $n = 3$, which means that $\Delta E_{12} > \Delta E_{23}$. Transition "A" shows an electron ***increasing*** its energy level; in order to do this, it must have absorbed some energy – either from an incoming e-m wave, or from being struck by a moving electron. If it absorbs an e-m photon, the incoming photon must have an energy ***exactly*** equal to ΔE_{12} . If it does not, the transition cannot take place. Similarly, transition "B" represents the ***emission*** of an e-m wave. ***Question:*** *Is the frequency of the e-m wave emitted in transition B* **greater than, less, than,** *or* **equal to** *the frequency of the e-m wave absorbed in transition A?* ***Answer:*** *The frequency of the "B" photon is* ***less*** *than that of the "A" photon, because its energy is less.* ***Question:*** *How does the wavelength of the photon emitted in transition C compare to that emitted in transition B?* ***Answer:*** *The wavelength of the C photon is* ***smaller*** *than that of the B photon; that is because the* ***frequency*** *(and the energy) of the C photon is greater, and a greater frequency corresponds to a smaller wavelength.*