2.10.2 Energy levels and photon emission

Electrons are found is discrete (separate) energy levels in atoms. These are often referred to as shells. Inner shells have lower energies than outer shells. Electrons can transition between shells provided they gain or lose precisely the energy required to do this. As we have seen previously, energy can be supplied by a colliding electron. Energy can also be supplied by a photon.

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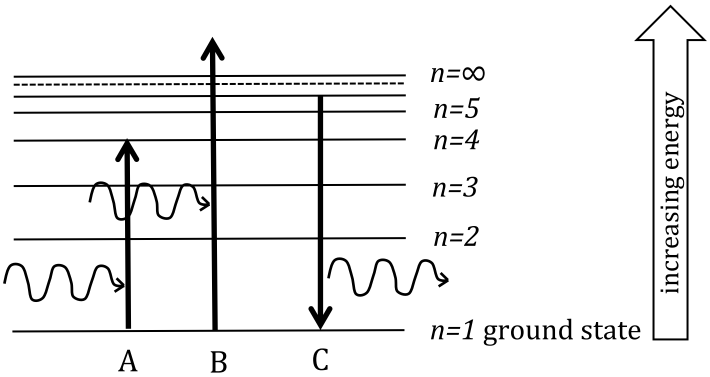
Electrons tend to occupy the lowest energy state, called the ground state. However, if electrons gain energy they can be raised to a higher energy level. This is usually a temporary situation, and the electron will subsequently drop down to the ground state, releasing energy as a photon. This is called de-excitation.



The energy levels are numbered (*n=1,2,3,…*), with *n=1* being the ground state.

Arrow A shows a transition from the ground state to a higher energy level (*n=4*). This could be caused by an electron absorbing a photon. This process is called excitation.

Arrow C shows a transition from *n=6* down to the ground state. The energy released during this transition would produce a photon. This process is called de-excitation.

(1)✐What process does arrow B represent?

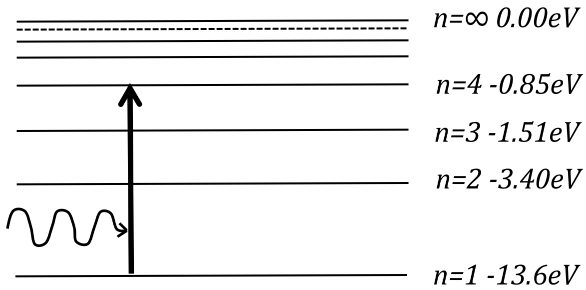
Quantifying excitation and de-excitation

If an electron is excited to a higher energy level by a photon, the energy of the photon ($E=hf$) is exactly equal to the energy required to undertake this transition. The energy change (ΔE) is equal to the difference in energy between the two energy levels. Therefore:

$$hf = \Delta E = E_2 - E_1$$

where h=Planck constant, f=frequency of the photon, E_1= energy of the lower level, E_2= energy of the higher energy level

Consider the energy levels in a hydrogen atom for an example.



Note that energy values are given in electron-volts (Remember *1eV=1.6x10*$^{-19}$*J*). Also note that energy levels are given negative values. This is not a problem in our calculations as we will be looking at the difference between energy levels.

In the diagram, we can see that a photon is absorbed by an electron and that it transitions from the ground state to a higher (less negative) energy level.

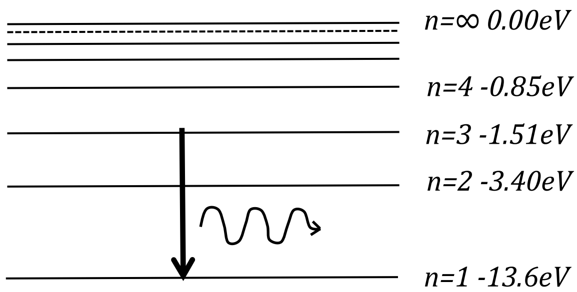
(2)✏Work out the energy difference (ΔE) between these two energy levels (n=1, and n=4) in eV.

(3)✏Convert this value of ΔE to joules using the conversion factor above.

(4)✏Rearrange the formula hf = ΔE, to make frequency the subject.

(5)✏Use this formula to calculate the frequency of photon required to produce this transition.

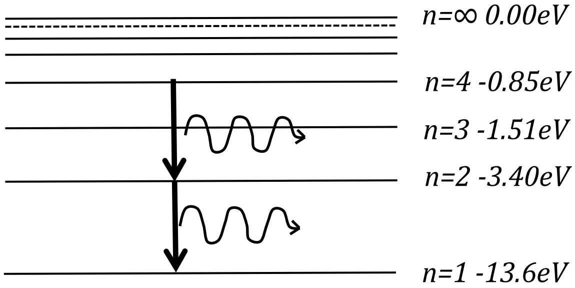
Now consider de-excitation:



This diagram shows that an electron in an excited state drops down to a lower energy level and emits a photon.

(6)✏*Follow the calculations, above, to work out the frequency of photon emitted.*

When an electron falls from a higher energy to the ground state, it may do this in a number of steps, releasing energy as a series of photons, each with a different frequency.



The diagram above shows a two-step de-excitation, with a transition from *n=4* to *n=2*, followed by a transition from *n=2* to the ground state.

(7)✐Which transition will produce a photon with the highest frequency? Why?

Fluorescence

Materials which fluoresce, absorb high energy photons (e.g. from ultraviolet light), causing atoms in the material to be excited. When de-excitation occurs, this happens in a number of steps. Some of the steps produce photons of visible light, so the material appears to glow. On example of this is the special UV marker pens used to mark bank notes. When they are exposed to UV light the marker ink will appear to glow.

(8)✐Find out how a fluorescent tube works.