THE ENERGY LEVELS OF ELECTRONS

Learning Objective

1. What happens to the magnitude of V if r is increased?

Model 1: Electron Energy

For an atom, such as hydrogen, with one electron orbiting around a nucleus with charge Z, the energy of the electron is given by the equation below:

$$E_n = -(2.18 \times 10^{-18}\ \text{J})\frac{Z^2}{n^2}$$

where n = 1, 2, 3, 4.... The different values of n correspond to the *allowed* energies that the electron can have. These energies are called "energy levels". The lowest energy level has $n = 1$ and is called the 'ground state'. All other energy levels are called 'excited states'. The average size of the electron's orbit is also controlled by the value of n:

$$r_{\text{average}} = (0.529 \times 10^{-10}\ \text{m})\ \frac{n^2}{Z}$$

Critical thinking questions

1. The hydrogen atom has atomic number $Z = 1$. Using the equations above for the energy and average radius of electron's orbit, complete the table below for hydrogen.

n	E_n (J)	r_{average} (m)
1	-218×10^{-20}	0.529×10^{-10}
2	-54.5×10^{-20}	2.12×10^{-10}
3	-24.2×10^{-20}	
4		
5		
6		
7		

2. Describe *in words* what happens to the energy levels and the average size of the orbit as n increases.

3. What do you predict happens to the energy and orbit of the electron when n becomes very large?

Model 2: Atomic Spectroscopy

When an electron in a hydrogen atom moves from an upper energy level to a lower energy level, it loses energy and a photon is produced. The energy (E) of the photon is equal to the *difference* in energy between the two levels. The wavelength (λ) of the photon can be calculated from this energy using:

$$\lambda = hc / E$$

where c is the speed of light (2.998×10^{17} nm s^{-1}) and h is Planck's constant (6.626×10^{-34} J s). The picture below shows the emission spectrum of the hydrogen atom in the ultraviolet region. The wavelength in nm is plotted on the x axis and the intensity of the light emitted (the number of photons per second) is plotted on the y axis. The spectrum shows a series of bands called the *Lyman series*. They are produced by the electron moving from excited states (high n values) to the ground state ($n = 1$).

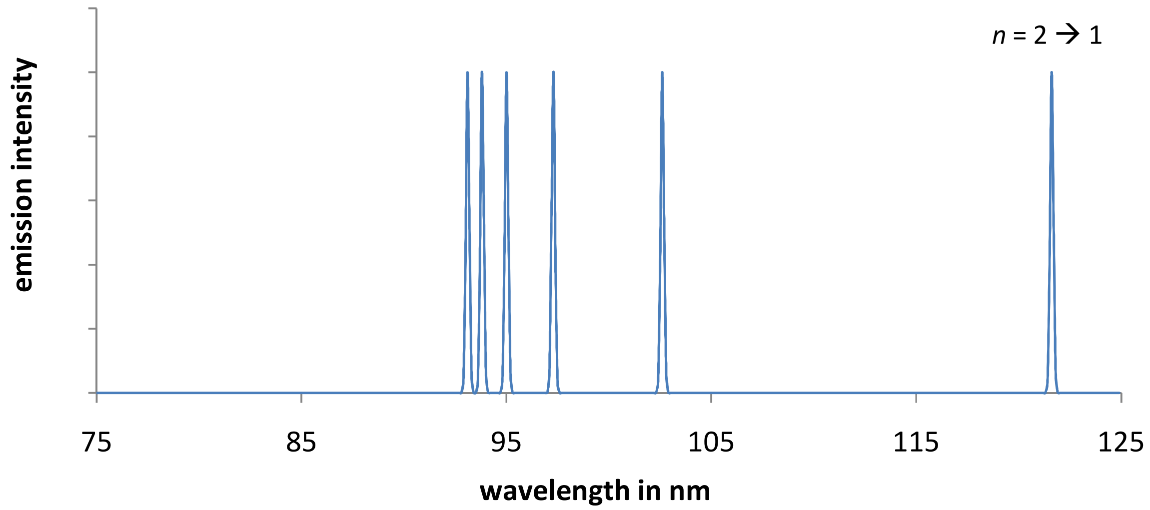
For example, if an electron moves from $n = 2$ to $n = 1$, the energy of the photon produced is equal to the difference in the energies of these two levels. From the table in question 1, this is equal to:

$$\Delta E = E \text{ (upper)} - E \text{ (lower)} = E_{n=2} - E_{n=1} = [(-54.5 \times 10^{-20}) - (-218 \times 10^{-20})] \text{ J} = 1.635 \times 10^{-18} \text{ J}$$

A photon with this energy will have wavelength:

$$\lambda = hc / E = (6.626 \times 10^{-34} \text{ J s}) \times (2.998 \times 10^{17} \text{ nm s}^{-1}) / (1.635 \times 10^{-18} \text{ J}) = 121 \text{ nm}$$

There is a peak at 121 nm in the spectrum below and this has been labelled "$n = 2 \rightarrow 1$" to show it corresponds to the transition between $n = 2$ and $n = 1$.



Critical thinking questions

4. Using your answers to question 1, complete the table below showing the energy and the wavelength of the photon that is emitted when an electron moves from an excited level into the $n = 1$ level.

n (upper)	E (upper)	n (lower)	E (lower)	$\Delta E = E$ (upper) - E (lower) (in J)	λ (in nm)
2	-54.5×10^{-20}	1	-218×10^{-20}	1.635×10^{-18}	121
3		1			
4		1			
5		1			
6		1			
7		1			

5. Above each peak in the spectrum, add labels showing what transitions the peaks correspond to.