- The lowest four energy levels of the He^+ ion are given. **3**

Principal quantum number (*n*)	Energy (J)
1	-8.720×10^{-18}
2	-2.180×10^{-18}
3	-0.969×10^{-18}
4	-0.545×10^{-18}

An electronic transition is identified by specifying the value of *n* of the initial state and the value of *n* of the final state. Identify the electronic transition responsible for the emission of radiation from He^+ with a wavelength of 121.5 nm?

The wavelength of light is related to its energy through Planck's equation:

$$E = \frac{hc}{\lambda}$$

Substituting the values for Planck's constant (*h*), the speed of light (*c*) and the wavelength gives:

$$E = \frac{(6.626 \times 10^{-34}\ \mathrm{J\ s})(2.998 \times 10^{8}\ \mathrm{m\ s^{-1}})}{(121.5 \times 10^{-9}\ \mathrm{m})} = 1.635 \times 10^{-18}\ \mathrm{J}$$

This corresponds to the energy difference between the *n* = 4 and *n* = 2 levels:

$$\Delta E = E_{n=4} - E_{n=2} = (-0.545 \times 10^{18}\ \mathrm{J}) - (-2.180 \times 10^{-18}\ \mathrm{J}) = 1.640 \times 10^{-18}\ \mathrm{J}$$

For the emission of light, the transition is from *n* = 4 to *n* = 2.