Marks

- Calculate the energy (in J) and wavelength (in nm) expected for an emission associated with an electronic transition from n = 4 to 3 in the B^{4+} ion. **3**

For the one electron ion, B^{4+}, the energy levels are given by

$$E_n = \frac{-E_R Z^2}{n^2} \text{ where } E_R = 2.18\times10^{-18} \text{ J}$$

with atomic number Z = 5. The energies of the n = 3 and 4 levels are then:

$$E_3 = \frac{-E_R(5)^2}{(3)^2} = -\frac{25}{9}E_R \text{ and } E_4 = \frac{-E_R(5)^2}{(4)^2} = -\frac{25}{16}E_R$$

The energy separation is $1.215E_R = 1.215 \times (2.18 \times 10^{-18}$ J) = $\underline{2.65 \times 10^{-18} \text{ J}}$

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

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

Substituting the values for Planck's constant (h), the speed of light (c) and the value of E from above gives:

$$\lambda = \frac{(6.626\times10^{-34} \text{ J s})(2.998\times10^{8} \text{ m s}^{-1})}{(2.65\times10^{-18} \text{ J})} = 7.50 \times 10^{-8} \text{ m} = 75.0 \text{ nm}$$

Energy = **2.65×10^{-18} J**	Wavelength = **7.50×10^{-8} m or 75.0 nm**

- Describe how EITHER the *photoelectric effect* OR the *visible spectrum of hydrogen* contributed to the development of quantum mechanics. **2**

Photoelectric effect:
Certain aspects of the photoelectric effect could only be explained by considering light as particulate - a steam of photons. The energy of the photons was proportional to the frequency (not intensity) of the light. This explained the facts that there was a minimum threshold energy and that there was no time lag.

Visible spectrum of hydrogen:
The visible spectrum of hydrogen showed distinct bands at certain wavelengths only. This showed that energy was quantised (ie not continuous) and that only certain energy levels were allowed.