Marks 6

- Ozone in the upper atmosphere absorbs light with wavelengths of 220 to 290 nm. What are the frequency (in Hz) and energy (in J) of the most energetic of these photons?

The energy, E, and frequency, ν, are related to the wavelength, λ, of light by

$E = \frac{hc}{\lambda}$ and $\nu = \frac{c}{\lambda}$ respectively.

As energy is inversely proportional to wavelength, the most energetic of these photons has the shortest wavelength, λ = 220 nm = 220 × 10^{-9} m. Hence,

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\text{ J s})(2.998 \times 10^{8}\text{ m s}^{-1})}{(220 \times 10^{-9}\text{ m})} = 9.0 \times 10^{-19}\text{ J}$$

$$\nu = \frac{c}{\lambda} = \frac{2.998 \times 10^{8}\text{ m s}^{-1}}{220 \times 10^{-9}\text{ m}} = 1.4 \times 10^{15}\text{ Hz}$$

Frequency: **1.4 × 10^{15} Hz**	Energy: **9.0 × 10^{-19} J**

Carbon-carbon bonds form the backbone of nearly every organic and biological molecule. The average bond energy of the C–C bond is 347 kJ mol^{-1}. Calculate the wavelength (in nm) of the least energetic photon that can break this bond.

A bond energy of 347 kJ mol^{-1} corresponds to a bond energy per molecule of

$$E = \frac{347 \times 10^{3}}{6.022 \times 10^{23}} \frac{\text{J mol}^{-1}}{\text{molecules mol}^{-1}} = 5.76 \times 10^{-19}\text{ J molecule}^{-1}$$

As $E = \frac{hc}{\lambda}$, the corresponding wavelength is:

$$\lambda = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34})(2.998 \times 10^{8})}{(5.76 \times 10^{-19})} = 3.45 \times 10^{-7}\text{ m} = 345\text{ nm}$$

Wavelength: **3.45 × 10^{-7} m or 345 nm**

Compare this value to that absorbed by ozone and comment on the ability of the ozone layer to prevent C–C bond disruption.

345 nm is not blocked by ozone, so C–C bond disruption is still possible even with the presence of the ozone layer. Hence one still needs to wear sunblock creams.