- Calculate the wavelength of light (in nm) emitted when an electron moves from the $n = 4$ to $n = 2$ energy levels in a hydrogen atom. **3**

The energy of an orbital in an 1-electron atom or ion is given by

$$E_n = -Z^2 E_R (1/n^2)$$

The energy difference between two levels is therefore:

$$\Delta E = E_{n1} - E_{n2} = [-Z^2 E_R (1/n_1^2)] - [-Z^2 E_R (1/n_2^2)] = Z^2 E_R (1/n_2^2 - 1/n_1^2)$$

The energy emitted when an electron moves from $n = 4$ to $n = 2$ for H with $Z = 1$ is therefore:

$$\Delta E = (1)^2 E_R (1/2^2 - 1/4^2)$$
$$= E_R \times (3/16) = (3/16) \times 2.18 \times 10^{-18} \text{ J} = 4.09 \times 10^{-19} \text{ J}$$

Using $E = hc / \lambda$, this corresponds to a wavelength of:

$$\lambda = hc / E = (6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^{8} \text{ m s}^{-1}) / (4.09 \times 10^{-19} \text{ J})$$
$$= 4.86 \times 10^{-7} \text{ m}$$
$$= 486 \text{ nm}$$

Answer: **486 nm**

What is the energy of this radiation (in kJ mol^{-1})?

The energy of each photon is **4.09×10^{-19} J. Therefore, per mol:**

energy of radiation = $(6.022 \times 10^{23} \text{ mol}^{-1}) \times (4.09 \times 10^{-19} \text{ J}) = 246 \text{ kJ mol}^{-1}$

Answer: **246 kJ mol^{-1}**