Answer on the question #63805, Chemistry / General Chemistry

Question:

How many photons are required to heat 200 mL of coffee from 25 ◦C to 65 ◦C?

Solution:

Let's express the quantity of energy required for the process. In assumption that there is no energy losses, the heating of coffee can be predicted by the following equation:

$$Q = cm\Delta T,$$

where c is the heat capacity of coffee, m is the mass of the substance and ΔT is the temperature change.

Let's take the heat capacity of coffee equal to the heat capacity of water, 4.181 J/ gK, at this temperature range.

The mass of our coffee will be the product of density (we take 1g/mL, as for pure water) and volume:

$$m = dV = 1\left(\frac{g}{mL}\right) \cdot 200\ (mL) = 200\ (g).$$

The change in temperature is:

$$\Delta T = (65 + 273.15)(K) - (25 + 273.15)(K) = \ 40\ (K).$$

Actually we see that when you take a difference for the temperature, it will be the same in Kelvin and Celsius scales. Then we can calculate the energy required for the heating of coffee:

$$Q = 4.181\left(\frac{J}{gK}\right) \cdot 200\ (g) \cdot (40)(K) = 33\,448\ (J)$$

Now, we should calculate how much photons we need to get the energy of 33 448 J.

The energy of one photon is:

$$E = h\nu,$$

where h is the Planck constant $6.626 \cdot 10^{-34}\ (Js)$ and ν is the frequency of the light.

As the frequency is not given, we will assume the wavelength from a microwave oven $\lambda = 11.2\ cm$. Then,

$$\nu = \frac{c}{\lambda}, E = \frac{hc}{\lambda}, n = \frac{Q}{E} = \frac{Q\lambda}{hc}$$

$$n = \frac{33\,448\ (J) \cdot 11.2 \cdot 10^{-2}\ (m)}{6.626 \cdot 10^{-34}\ (Js) \cdot 3 \cdot 10^{8}(m\ s^{-1})} = 18\,845 \cdot 10^{24} = 1.88 \cdot 10^{28}$$

Answer : you will need 1.88· 10^{28} photons