- Calculate the heat input required (in J) for the conversion of 9.0 g of water from ice at 273 K to steam at 373 K. **2**

  Data: C_p $H_2O(l)$ = 75 J K^{-1} mol^{-1}

  ΔH_{vap} $H_2O(l)$ = 41 kJ mol^{-1} ΔH_{fus} $H_2O(s)$ = 6.0 kJ mol^{-1}

The molar mass of water is (2 × 1.008 (H)) + 16.00 (O) = 18.016 g mol^{-1}

Therefore, 9.0 g corresponds to:

$$\text{number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}} = \frac{9.0\text{ g}}{18.016\text{ g mol}^{-1}} = 0.50\text{ mol}$$

The heat required can be broken down into 3 contributions.

(i) Heat required to melt ice (q_1)

6.0 kJ is required to melt 1 mole so:

$$q_1 = (0.50\text{ mol}) \times (6.0\text{ kJ}) = 3.0\text{ kJ} = 3.0 \times 10^3\text{ kJ} = 3000\text{ J}$$

(ii) Heat required to warm water from 273 K to 373 K (q_2):

Using $q = n \times C_p \times \Delta T$,

$$q_2 = (0.50\text{ mol}) \times (75\text{ J K}^{-1}\text{ mol}^{-1}) \times (373 - 273\text{ K}) = 3700\text{ J}$$

(iii) Heat required to vaporise water (q_3):

41 kJ is required to vapourize 1 mole so:

$$q_1 = (0.50\text{ mol}) \times 41\text{ (kJ mol}^{-1}\text{)} = 20\text{ kJ} = 20 \times 10^3\text{ kJ} = 20000\text{ J}$$

The total heat required is therefore:

$$q_{total} = q_1 + q_2 + q_3 = (20000\text{ J}) + (3700\text{ J}) + (3000\text{ J}) = 27000\text{ J}$$

As the question gives the mass and heats of fusion and vaporization to 2 significant figures, the answer is also quoted to this level of accuracy.