Marks
4

- A 150.0 g block of iron metal is cooled by placing it in an insulated container with a 50.0 g block of ice at 0.0 °C. The ice melts, and when the system comes to equilibrium the temperature of the water is 78.0 °C. What was the original temperature (in °C) of the iron?

  Data: The specific heat capacity of liquid water is 4.184 J K^{-1} g^{-1}.
  The specific heat capacity of solid iron is 0.450 J K^{-1} g^{-1}.
  The molar enthalpy of fusion of ice (water) is 6.007 kJ mol^{-1}.

The heat from the iron is used to melt the ice and to warm the water from 0.0 °C to 78.0 °C.

The molar mass of H_2O is (2 × 1.008 (H) + 16.00 (O)) g mol^{-1} = 18.02 g mol^{-1}. Hence 50.0 g of ice corresponds to:

number of moles = mass / molar mass = (50.0 g) / (18.02 g mol^{-1}) = 2.775 mol.

Hence the heat used to melt ice is:

q_1 **= 6.007 kJ mol^{-1} × 2.775 mol = 16.67 kJ = 16670 J**

The heat used to warm 50.0 g water by 78.0 °C is:

$q_2 = m \times C \times \Delta T$ **= (50.0 g) × (4.184 J K^{-1} g^{-1}) × (78.0 K) = 16320 J**

Overall, the heat transferred from the iron is:

$q = q_1 + q_2$ **= 16670 J + 16320 J = 32990 J**

This heat is lost from 150.0 g of iron leading to it cooling by ΔT:

$q = m \times C \times \Delta T$ **= (150.0 g) × (0.450 J K^{-1} g^{-1}) × ΔT = 32990 J**

ΔT **= 489 K = 489 °C**

As the final temperature of the iron is 78.0 °C, its original temperature was (78.0 + 489) °C = 567 °C.

Answer: **567 °C**