Answer on Question #56897, Physics / Molecular Physics | Thermodynamics

A 28.45 g metal specimen of gold and 25.75 g metal specimen of iron, each at a temperature of 100.0 °C were dropped into 570.0 mL of water at a temperature of 17.70 °C. The molar heat capacity of iron and gold are 25.19 J mol^{-1} $°C^{-1}$ and 25.41 J mol^{-1} $°C^{-1}$ respectively. Assume the density of water to be 1.00g/mL, what is the final temperature of the water and pieces of metals?

Solution:

This problem can be summarized thusly:

$$q_{lost\ by\ gold} + q_{lost\ by\ iron} = q_{gained\ by\ water}$$

$$c_{gold} m_{gold}(T_1 - T_x) + c_{iron} m_{iron}(T_1 - T_x) = c_{water} m_{water}(T_x - T_2)$$

The specific heat of water is = 4.187 joule/gram.

The molar mass of gold M_{Au} = 196.97 g/mol

The molar mass of iron M_{Fe} = 55.8450 g/mol

Thus,

$$c_{gold} = \frac{25.19}{196.97} = 0.128 \frac{J}{mol\ ℃}$$

$$c_{iron} = \frac{25.41}{55.8450} = 0.455 \frac{J}{mol\ ℃}$$

Here are what the several of the terms mean:

(25.75 g / 55.845 g/mol) ---> moles of Fe

(28.45g / 196.97 g/mol) ---> moles of Au

(100.0 - x) ---> temp change of Fe and Au (they each start at 100 °C and go down to the final temperature, symbolized by 'x')

(x - 17.70) ---> temp change of water

Solve for x, which is the final temperature.

Therefore:

$$0.128 \cdot 28.45 \cdot (100 - x) + 0.455 \cdot 25.75 \cdot (100 - x) = 570 \cdot 4.187 \cdot (x - 17.70)$$

$$15.3579 \cdot (100 - x) = 2386.59 \cdot (x - 17.7)$$

$$43778.4 - 2401.95\, x = 0$$

$$x = 18.2262 \approx 18.2℃$$

Answer: 18.2℃