Marks 4

- A key step in the metabolism of glucose for energy is the isomerism of glucose-6-phosphate (G6P) to fructose-6-phosphate (F6P);

$$G6P \rightleftharpoons F6P$$

At 298 K, the equilibrium constant for the isomerisation is 0.510. Calculate $\Delta G°$ at 298 K.

Using $\Delta G° = -RT\ln K$:

$$\Delta G° = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K}) \times \ln(0.510)$$
$$= +1670 \text{ J mol}^{-1} = +1.6 \text{ kJ mol}^{-1}$$

Answer: **+1.6 kJ mol^{-1}**

Calculate ΔG at 298 K when the [F6P] / [G6P] ratio = 10.

Using $\Delta G = \Delta G^{o} + RT\ln Q$, when the reaction quotient $Q = \frac{[F6P]}{[G6P]} = 10$:

$$\Delta G = (+1670 \text{ J mol}^{-1}) + (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K}) \times \ln(10)$$
$$= +7400 \text{ J mol}^{-1} = +7.4 \text{ kJ mol}^{-1}$$

Answer: **+7.4 kJ mol^{-1}**

In which direction will the reaction shift in order to establish equilibrium? Why?

As $Q > K$, the reaction will shift to decrease Q. It will do this by reducing the amount of product and increasing the amount of reactant: it will shift to the left.

Equivalently, as $\Delta G = +7.4$ kJ mol^{-1}, the forward process is non-spontaneous and the backward reaction is spontaneous.

THE ANSWER CONTINUES ON THE NEXT PAGE

- The specific heat capacity of water is 4.18 J g^{-1} K^{-1} and the specific heat capacity of copper is 0.39 J g^{-1} K^{-1}. If the same amount of energy were applied to a 1.0 mol sample of each substance, both initially at 25 °C, which substance would get hotter? Show all working. **2**

As $q = C \times m \times \Delta T$, the temperature increase is given by $\Delta T = \dfrac{q}{C \times m}$.

As H_2O has a molar mass of (2 × 1.008 (H) + 16.00 (O)) g mol^{1} = 18.016 g mol^{1}, 1.0 mol has a mass of 18 g. The temperature increase is therefore:

$$\Delta \mathrm{T} = \frac{q}{C_{\mathrm{H_2O}} \times m_{\mathrm{H_2O}}} = \frac{q}{(4.18\,\mathrm{J\,g^{-1}\,K^{-1}}) \times (18\,\mathrm{g})} = \frac{q}{(75\,\mathrm{J\,K^{-1}})}$$

As Cu has an atomic mass of 63.55 g mol^{1}, 1.0 mol has a mass of 64 g. The temperature increase is therefore:

$$\Delta T = \frac{q}{C_{\mathrm{Cu}} \times m_{\mathrm{Cu}}} = \frac{q}{(0.39\,\mathrm{J\,g^{-1}\,mol^{-1}}) \times (64\,\mathrm{g})} = \frac{q}{(25\,\mathrm{J\,K^{-1}})}$$

As the same amount of energy is supplied to both, q is the same for both. The temperature increase of the copper is therefore higher.

Answer: **copper**