Get some extra practice…

…working with the Gibbs free energy.

1. The first step in the metabolism of glucose yields glucose 6-phosphate in a reaction with an equilibrium constant, *K*, of 4.4×10^{-3} under standard conditions at 298 K. Does this reaction happen spontaneously?

2. The free energy change for the hydrolysis of ATP at 37 °C is −31 kJ mol^{-1}. Does this reaction lie to the left or the right at equilibrium?

3. The free energy change for the formation of a peptide bond is +17 kJ mol^{-1} at 37 °C. What is the equilibrium constant for this reaction?

4. The enzyme carbonic anhydrase catalyses the following reaction:

   $CO_2 + H_2O \rightarrow HCO_3^- + H^+$

   The enthalpy change for this reaction is -303.29 kJ mol^{-1}.

   The entropy change is −96.3 J K^{-1} mol^{-1}

   What is the change in free energy at 25 °C?

Scroll to the following pages to check your answers.

Answers

1. The first step in the metabolism of glucose yields glucose 6-phosphate in a reaction with an equilibrium constant, *K*, of 4.4×10^{-3} under standard conditions at 298 K. Does this reaction happen spontaneously?

The first step in answering this question is to calculate the value of the Gibbs free energy change, using the relationship:

$\Delta G° = -RT\ln K$

$= -((8.31\ \mathrm{J\ K^{-1}\ mol^{-1}}) \times (298\ \mathrm{K}) \times \ln(4.4 \times 10^{-3}))$

$= -(2476.38\ \mathrm{J\ mol^{-1}} \times -5.43)$

$= 13436\ \mathrm{J\ mol^{-1}}$

$= +13.4\ \mathrm{kJ\ mol^{-1}}$

As the value of the Gibbs free energy change is positive, we can deduce that the reaction is **not spontaneous**.

2. The Gibbs free energy change for the hydrolysis of ATP at 37 °C is −31 kJ $\mathrm{mol^{-1}}$. Does this reaction lie to the left or the right at equilibrium?

There are two ways that we could approach the response to this question. First, the value of the Gibbs free energy change is negative, which tells us that the reaction happens spontaneously. From this, we can deduce that the reaction lies to the right at equilibrium.

We can check this assertion by calculating the value of the equilibrium constant for the reaction. To do this, we rearrange the van t'Hoff isotherm to isolate our unknown variable, *K*:

$\Delta G° = -RTln\ K$

$\frac{\Delta G°}{-RT} = \ln K$

$e^{\frac{\Delta G°}{-RT}} = K$

We can now enter our known values into this equation, remembering that temperature, *T*, is cited in kelvin. So we convert 37 °C to 310 K.

$$e^{\frac{\Delta G°}{-RT}} = K$$

$$\mathrm{e}^{\left(\frac{-31 \times 10^3}{-(8.31 \times 310)\ \mathrm{J\ mol^{-1}}}\right)} = K$$

$$K = e^{12.03}$$

$$= 16.77 \times 10^4$$

(Notice that we've also expressed the value of ΔG in J (by multiplying by a factor of 10^3) to take account of the fact that the gas constant is expressed in J K^{-1} mol^{-1}, not kJ K^{-1} mol^{-1}.)

This large, positive value for *K* is consistent with the reaction lying to the right.

3. The free energy change for the formation of a peptide bond is +17 kJ mol^{-1} at 37 °C. What is the equilibrium constant for this reaction?

To calculate *K*, we use the same rearranged form of the van t'Hoff isotherm that we use in question 2:

$$e^{\frac{\Delta G°}{-RT}} = K$$

$$\mathrm{e}^{\left(\frac{17 \times 10^3}{-(8.31 \times 310)\ \mathrm{J\ mol^{-1}}}\right)} = K$$

$$K = e^{-6.60}$$

$$= 1.36 \times 10^{-3}$$

This value is consistent with a non-spontaneous reaction, which requires an input of energy to drive it forward.

4. The enzyme carbonic anhydrase catalyses the following reaction:

   $CO_2 + H_2O \rightarrow HCO_3^- + H^+$

   The enthalpy change for this reaction is -303.29 kJ mol^{-1}.

   The entropy change is −96.3 J K^{-1} mol^{-1}

   What is the change in free energy at 25 °C?

To answer this question, we need to remember the following relationship:

$$\Delta G° = \Delta H° - T\Delta S°$$

We can then feed our known values into this equation. (Again, we need to ensure there is consistency in terms of the use of J versus kJ; entropy, S, uses joules, so we multiply our value of H by a factor of 10^3 so that it is expressed in joules too.)

$$\Delta G° = (-303.29 \times 10^3 \text{ J mol}^{-1}) - (298 \text{ K } \times \ -96.3 \text{ J K}^{-1} \text{ mol}^{-1})$$

$$= -3.0329 \times 10^3 - (-28697.4)$$

$$= 31730.3 \text{ J mol}^{-1} = \mathbf{31.73\ kJ\ mol^{-1}}$$