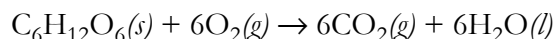


Thermodynamics of Metabolism

Metabolism is a the set of reactions occurring in cells that sustain life. Groups of reactions that accomplish a particular need are called metabolic pathways. One important metabolic pathway is glycolysis, in which the combustion of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, produces water and carbon dioxide:



Characterize the thermodynamic favorability of this reaction with respect to temperature. The standard state free energy of formation, standard state enthalpy of formation and the absolute entropy for glucose are shown here; you will find other necessary thermodynamic values in your textbook:

$$\Delta G_f^\circ = -910.52 \text{ kJ/mol}_{\text{rxn}} \quad \Delta H_f^\circ = -1274.4 \text{ kJ/mol}_{\text{rxn}} \quad S^\circ = 212.1 \text{ J/mol}_{\text{rxn}} \cdot \text{K}$$

Answer. The values of ΔH° and ΔS° for the reaction are

$$\Delta H^\circ = -2806 \text{ kJ/mol}_{\text{rxn}} \quad \Delta S^\circ = 262 \text{ J/mol}_{\text{rxn}} \cdot \text{K}$$

The reaction, therefore, is favorable at all temperatures.

Imagine that a man weighing 60 kg eats a candy bar containing 60 g of glucose. Imagine, as well, that his body is a perfect calorimeter (not a good assumption, of course). Assuming the specific heat of his body is the same as that for water, calculate the change in the man's body temperature? In your experience as a consumer of candy bars, is your answer reasonable? Explain. If your answer isn't reasonable, provide at least one plausible explanation—other than the body is not a perfect calorimeter—for why the calculated ΔT is larger or smaller than expected.

Answer. Knowing that ΔH° is $-2806 \text{ kJ/mol}_{\text{rxn}}$, we first calculate the heat released when converting 60 g of glucose to carbon dioxide and water

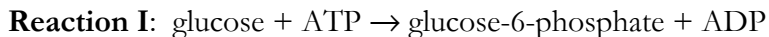
$$q_{\text{rxn}} = \frac{-2806 \text{ kJ}}{\text{mol}_{\text{rxn}}} \times \frac{1 \text{ mol}_{\text{rxn}}}{\text{mol C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g C}_6\text{H}_{12}\text{O}_6} \times 60 \text{ g C}_6\text{H}_{12}\text{O}_6 = -934.3 \text{ kJ}$$

We know that q_{soln} is $-q_{\text{rxn}}$ so the body is absorbing 934.3 kJ of heat. To find the increase in temperature we solve as follows

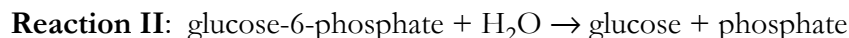
$$q_{\text{body}} = 934.3 \text{ kJ} = m\Delta T = (60,000 \text{ g}) \times (4.184 \times 10^{-3} \text{ kJ/g} \cdot ^\circ\text{C}) \times (\Delta T)$$

giving a ΔT of 3.7°C . Of course, this answer is not reasonable as personal experience attests. The reason that the body's temperature doesn't rise this quickly is that the reaction doesn't take place with a single burst of released energy. Instead, the overall reaction proceeds through many steps with the energy leaking out bit-by-bit. And, of course, much of the released energy is used to accomplish useful chemical work and not to just heat up the body.

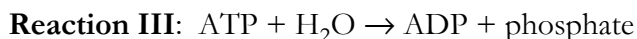
The first step in the glycolytic pathway is the transfer of a phosphate from adenosine-triphosphate (ATP) to glucose:



Under standard state conditions and in the absence of ATP, glucose-6-phosphate is unstable and reacts with water to produce glucose and a free phosphate:



Additionally, for the hydrolysis of ATP under standard state conditions



the free energy change is favorable.

$$\Delta G^\circ_{\text{III}} = -32.48 \text{ kJ/mol}_{\text{rxn}}$$

What prediction can you make concerning the sign of ΔG° for Reaction II under standard state conditions? Explain.

Answer. We are told that glucose-6-phosphate is “unstable and reacts with water to produce glucose and a free phosphate.” From this statement we know that the hydrolysis of glucose-6-phosphate is favorable and that $\Delta G^\circ_{\text{II}} < 0$.

Under standard state conditions, the sign of ΔG° for Reaction I is unclear given the information provided above. Using Hess’s law, what can you say about the value of $\Delta G^\circ_{\text{I}}$? State your answer as $\Delta G^\circ_{\text{I}}$ must be greater than..., or $\Delta G^\circ_{\text{I}}$ must be less than..., or $\Delta G^\circ_{\text{I}}$ is equal to... Hint: note that Reaction III is the sum of Reactions I and II.

Answer. We know that the $\Delta G^\circ_{\text{III}}$ for the conversion of ATP to ADP is $-32.48 \text{ kJ/mol}_{\text{rxn}}$ and that $\Delta G^\circ_{\text{II}}$ for the hydrolysis of glucose is < 0 . Using Hess’s law, we know that

$$\text{reaction I} + \text{reaction II} = \text{reaction III}$$

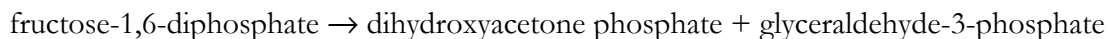
$$\Delta G_{\text{I}} + \Delta G_{\text{II}} = \Delta G_{\text{III}}$$

$$\Delta G_{\text{I}} + \Delta G_{\text{II}} = -32.48 \text{ kJ/mol}_{\text{rxn}}$$

$$\Delta G_{\text{I}} = -32.48 \text{ kJ/mol}_{\text{rxn}} - \Delta G_{\text{II}}$$

From this we can see that ΔG_{I} may be either favorable or unfavorable depending upon the value of ΔG_{II} . For example, if ΔG_{II} is $< -32.48 \text{ kJ/mol}_{\text{rxn}}$, then ΔG_{I} is unfavorable.

Another step in the glycolytic pathway is



for which ΔG° is $+23.8 \text{ kJ/mol}_{\text{rxn}}$ at a temperature of 298 K. Although this reaction is unfavorable under standard state conditions, we know it routinely occurs in cells; thus, the concentrations of reactants and products in the cells must not be at their standard state values. What is the largest reaction quotient, Q , that will produce a favorable free energy change? Recall that Q is a measure of the relative amounts of products to reactants.

Answer. If the reaction occurs then ΔG must be less than zero. To find the relative ratio of products to reactants (Q), we recall that

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$0 = +23.8 \text{ kJ/mol}_{\text{rxn}} + (8.314 \times 10^{-3} \text{ J/mol}_{\text{rxn}} \cdot \text{K})(298 \text{ K}) \ln Q$$

$$Q = 6.73 \times 10^{-5}$$

Thus, the value of Q must be less than 6.73×10^{-5} for the reaction to be favorable.

The value of Q is approximately 10^{-4} in the cells of rats. What is the value of ΔG for this reaction under these conditions and what does this tell you about the favorability of this glycolytic pathway in rats? If you believe that the reaction is unfavorable, why might it still occur?

Answer. The typical ratio found in rats of approximately 10^{-4} suggests that the reaction is still not favorable as it is larger than the value calculated above. The value of ΔG under these conditions, however

$$\Delta G = +23.8 \text{ kJ/mol}_{\text{rxn}} + (8.314 \times 10^{-3} \text{ J/mol}_{\text{rxn}} \cdot \text{K})(298 \text{ K}) \ln(10^{-4}) = 0.98 \text{ kJ/mol}_{\text{rxn}}$$

is only slightly unfavorable. A rat's body temperature, of course, is not 25°C , but is closer to 38°C , which might explain why the reaction is favorable. For example, at 311 K and assuming no change in ΔG° , we find that ΔG is slightly favorable

$$\Delta G = +23.8 \text{ kJ/mol}_{\text{rxn}} + (8.314 \times 10^{-3} \text{ J/mol}_{\text{rxn}} \cdot \text{K})(311 \text{ K}) \ln(10^{-4}) = -0.015 \text{ kJ/mol}_{\text{rxn}}$$

Of course, ΔG° will change, but since we can guess that ΔH° is positive and ΔS° is positive for this reaction, an increase in temperature will make ΔG° less positive and ΔG becomes even more favorable.