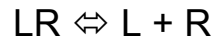


Bioenergetics Practice Questions Answers

1. For the dissociation of a ligand, L, from a membrane receptor, R:



- The equilibrium constant for dissociation is $K_d = 2 \times 10^{-6}$ M. Intuitively, how much dissociation do you expect, a little or a lot?
- If $[L] = 1 \times 10^{-5}$ M, $[R] = 5 \times 10^{-7}$ M, calculate $[LR]$.
- For the dissociation reaction at equilibrium calculate ΔG° at 25 °C (as written).
- Which reaction is favored, dissociation or association?

Since K_d is such a small number, you would expect to get very little dissociation.

$$K_d = [L][R] / [LR] = (1 \times 10^{-5} \text{ M})(5 \times 10^{-7} \text{ M}) / [LR] = 2 \times 10^{-6} \text{ M}$$

$$[LR] = (5 \times 10^{-12} \text{ M}^2) / (2 \times 10^{-6} \text{ M}) = 2.5 \times 10^{-6} \text{ M}$$

The system is at equilibrium so $\Delta G^\circ = -RT \ln K_{eq}$

$$= -(8.315 \times 10^{-3} \text{ kJ / K} \cdot \text{ mol})(298 \text{ K}) \ln(2 \times 10^{-6} \text{ M})$$

$$= + 32.52 \text{ kJ/mol}$$

Since $\Delta G^\circ \gg 0$, the dissociation reaction is not favorable (but we knew that because of the value for K_d !).

2. For the hydrolysis of ATP: $ATP \rightleftharpoons ADP + P_i$

If the **equilibrium** concentrations of ATP = 1×10^{-7} M, ADP = 0.165 M and $P_i = 0.1$ M, what are the equilibrium constant and ΔG° for the hydrolysis of ATP at 37°C?

Use the data to calculate K_{eq}' .

$$K_{eq}' = \frac{[ADP][P_i]}{[ATP]} = \frac{[0.165M][0.1M]}{[1 \times 10^{-7} M]}$$
$$= 1.65 \times 10^5 M$$

Use the value of K_{eq}' to calculate $\Delta G^{\circ'}$.

$$\Delta G^{\circ'} = -RT \ln K_{eq}' = -(8.315 \times 10^{-3} \text{ kJ/K} \cdot \text{mol})(310 \text{ K}) \ln(1.65 \times 10^5)$$

$$\Delta G^{\circ'} = -(2.578)(12) = -31.0 \text{ kJ/mol}$$

3. In a typical cell at 37°C the concentration of ATP = $8 \times 10^{-3} \text{ M}$, ADP = $1 \times 10^{-3} \text{ M}$, and P_i = $8 \times 10^{-3} \text{ M}$. What is the actual free energy change ($\Delta G'$) for ATP hydrolysis under these conditions?

The KEY is to recognize that this system is NOT at equilibrium.

$$\Delta G' = \Delta G^{\circ'} + RT \ln \frac{[ADP][P_i]}{[ATP]}$$

$$\Delta G' = -31.0 \text{ kJ/mol} + (8.315 \times 10^{-3} \text{ kJ/mol} \cdot \text{K})(310 \text{ K}) \ln \frac{[1 \times 10^{-3} \text{ M}][8 \times 10^{-3} \text{ M}]}{[8 \times 10^{-3} \text{ M}]}$$

$$\Delta G' = -31.0 \text{ kJ/mol} + (-17.8 \text{ kJ/mol}) = -49 \text{ kJ/mol}$$

4. The conversion of A to E is coupled by the two reactions below:



Calculate the total $\Delta G^{\circ'}$ for the coupled reactions.

$$\Delta G^{\circ'}_{total} = \Delta G^{\circ'}_1 + \Delta G^{\circ'}_2 = (+15 \text{ kJ/mol}) + (-35 \text{ kJ/mol})$$
$$= -20 \text{ kJ/mol}$$

Thus, coupling these two reactions together permits the thermodynamically favorable conversion of A to E.