

**Problem 2.** (20 points total, 10 points each)

Consider a hypothetical reaction:  $A \rightleftharpoons B + C$

Molar concentrations of the compounds at the beginning of the reaction that takes place at 25°C were  $[A] = 0.1 \text{ M}$ ,  $[B] = 0$ ,  $[C] = 0$ . Assume we can monitor the course of the reaction by measuring the concentration of B.

(a) Molar concentration of B at the equilibrium was  $[B]_{\text{eq}} = 40 \text{ mM}$ . From the data you have calculate the standard change in free energy,  $\Delta G^{\circ}$ , for the forward reaction.

$$\Delta G^{\circ} = -RT \ln K_{\text{eq}}', \quad K_{\text{eq}}' = \frac{[B]_{\text{eq}} \cdot [C]_{\text{eq}}}{[A]_{\text{eq}}}$$
$$[B]_{\text{eq}} = [C]_{\text{eq}} = 0.04 \text{ M}; \quad [A]_{\text{eq}} = 0.1 \text{ M} - 0.04 \text{ M} = 0.06 \text{ M}$$
$$K_{\text{eq}}' = \frac{0.04 \cdot 0.04}{0.06} \text{ M} = 0.027 \text{ M}$$
$$\Delta G^{\circ} = -(8.314 \cdot 298) \ln(0.027) \text{ J/mol} = 8.95 \text{ kJ/mol}$$

(b) Ten minutes after the reaction started, the concentration of B was  $[B] = 20 \text{ mM}$ . Calculate  $\Delta G$  of the reaction at this point.

$$\Delta G = \Delta G^{\circ} + RT \ln \frac{[B][C]}{[A]}$$
$$[B] = [C] = 0.02 \text{ M}; \quad [A] = 0.1 \text{ M} - 0.02 \text{ M} = 0.08 \text{ M}$$
$$\Delta G = 8.95 \frac{\text{kJ}}{\text{mol}} + 2.478 \frac{\text{kJ}}{\text{mol}} \cdot \ln \frac{0.02 \cdot 0.02}{0.08} =$$
$$= -4.18 \text{ kJ/mol}$$