Answers to Exercise 7.3
Calculating Free Energy for Reactions Under Nonstandard Conditions

1.

(a) $CS_{2(g)} + 4\ H_{2(g)} \rightarrow CH_{4(g)} + 2\ H_2S_{(g)}$

(b) $\Delta_r G° = \sum \Delta_f G°(products) - \sum \Delta_f G°(reactants)$

$\Delta_r G° = \left[\Delta_f G°\left(CH_{4(g)}\right) + 2\ \Delta_f G°\left(H_2S_{(g)}\right)\right] - \left[\Delta_f G°\left(CS_{2(g)}\right) + 4\ \Delta_f G°\left(H_{2(g)}\right)\right]$

$\Delta_r G° = \left[\left(-50.72\frac{kJ}{mol}\right) + 2\left(-33.56\frac{kJ}{mol}\right)\right] - \left[\left(67.1\frac{kJ}{mol}\right) + 4\left(0\frac{kJ}{mol}\right)\right]$

$\Delta_r G° = -184.9\frac{kJ}{mol}$

The answer is significant to one decimal place, but use all digits $\left(-184.94\frac{kJ}{mol}\right)$ *when you take this value to part (c).*

(c) **Step 1: Convert temperature to Kelvin**

$T = 25\ ℃ = 298.15\ K$

Step 2: Determine activities for all reactants and products

$a_{CS_{2(g)}} = \frac{0.15\ bar}{1\ bar} = 0.15$ $\qquad$ $a_{CH_{4(g)}} = \frac{1.25\ bar}{1\ bar} = 1.25$

$a_{H_{2(g)}} = \frac{1.65\ bar}{1\ bar} = 1.65$ $\qquad$ $a_{H_2S_{(g)}} = \frac{0.35\ bar}{1\ bar} = 0.35$

Step 3: Calculate reaction quotient

$$Q = \frac{\left(a_{CH_{4(g)}}\right)\left(a_{H_2S_{(g)}}\right)^2}{\left(a_{CS_{2(g)}}\right)\left(a_{H_{2(g)}}\right)^4} = \frac{(1.25)(0.35)^2}{(0.15)(1.65)^4} = 0.14$$

Step 4: Calculate free energy change for reaction under nonstandard conditions

$\Delta_r G = \Delta_r G° + RT lnQ$

$\Delta_r G = \left(-184.9\frac{kJ}{mol}\right) + \left(8.314\ 462\frac{J}{mol{\cdot}K}\right)\left(\frac{1\ kJ}{1000\ J}\right)(298.15\ K)ln(0.14)$

$\Delta_r G = \left(-184.9\frac{kJ}{mol}\right) + \left(-4.91\frac{kJ}{mol}\right)$

$\Delta_r G = -189.9\frac{kJ}{mol}$

Step 5: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

2.

(a) **Step 1: Write a balanced chemical equation for the reaction**

This was provided in the question.

$2\ SO_{2(g)} + O_{2(g)} \rightarrow\ 2\ SO_{3(g)}$

Step 2: Convert temperature to Kelvin (if necessary)

$T = 25\ ℃ = 298.15\ K$

Step 3: Calculate the standard free energy change for the reaction

$\Delta_r G° = \sum \Delta_f G°(products) - \sum \Delta_f G°(reactants)$

$\Delta_r G° = 2\ \Delta_f G°(SO_{3(g)}) - [2\ \Delta_f G°(SO_{2(g)}) + \Delta_f G°(O_{2(g)})]$

$\Delta_r G° = 2\left(-371.1\frac{kJ}{mol}\right) - \left[2\left(-300.2\frac{kJ}{mol}\right) + \left(0\frac{kJ}{mol}\right)\right]$

$\Delta_r G° = -141.8\frac{kJ}{mol}$

Step 4: Determine activities for all reactants and products

$a_{SO_{2(g)}} = \frac{0.48\ bar}{1\ bar} = 0.48$

$a_{SO_{3(g)}} = \frac{0.72\ bar}{1\ bar} = 0.72$

$a_{O_{2(g)}} = \frac{0.18\ bar}{1\ bar} = 0.18$

Step 5: Calculate reaction quotient

$Q = \frac{\left(a_{SO_{3(g)}}\right)^2}{\left(a_{SO_{2(g)}}\right)^2\left(a_{O_{2(g)}}\right)} = \frac{(0.72)^2}{(0.48)^2(0.18)} = 12.5$ *but only 2 sig. fig.*

Step 6: Calculate free energy change for reaction under nonstandard conditions

$\Delta_r G = \Delta_r G° + RT ln Q$

$\Delta_r G = \left(-141.8\frac{kJ}{mol}\right) + \left(8.314\ 462\frac{J}{mol{\cdot}K}\right)\left(\frac{1\ kJ}{1000\ J}\right)(298.15\ K) ln(12.5)$

$\Delta_r G = \left(-141.8\frac{kJ}{mol}\right) + \left(6.26\frac{kJ}{mol}\right)$

$\Delta_r G = -135.5\frac{kJ}{mol}$

Step 7: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

(b) A catalyst has no effect on the position of equilibrium. The catalyst's role is to speed up both the forward and reverse reactions. It may, therefore help the system reach a state of equilibrium more quickly.