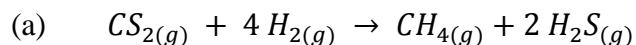


Answers to Exercise 7.3

Calculating Free Energy for Reactions Under Nonstandard Conditions

1.



(b) $\Delta_r G^\circ = \sum \Delta_f G^\circ(\text{products}) - \sum \Delta_f G^\circ(\text{reactants})$

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

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

$$\Delta_r G^\circ = -184.9 \frac{\text{kJ}}{\text{mol}}$$

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

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

$$T = 25 \text{ }^\circ\text{C} = 298.15 \text{ K}$$

Step 2: Determine activities for all reactants and products

$$a_{CS_{2(g)}} = \frac{0.15 \text{ bar}}{1 \text{ bar}} = 0.15$$

$$a_{CH_{4(g)}} = \frac{1.25 \text{ bar}}{1 \text{ bar}} = 1.25$$

$$a_{H_{2(g)}} = \frac{1.65 \text{ bar}}{1 \text{ bar}} = 1.65$$

$$a_{H_2S_{(g)}} = \frac{0.35 \text{ bar}}{1 \text{ bar}} = 0.35$$

Step 3: Calculate reaction quotient

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

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

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

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

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

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

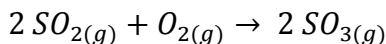
Step 5: Check your work

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.



Step 2: Convert temperature to Kelvin (if necessary)

$$T = 25 \text{ }^\circ\text{C} = 298.15 \text{ K}$$

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

$$\Delta_r G^\circ = \sum \Delta_f G^\circ(\text{products}) - \sum \Delta_f G^\circ(\text{reactants})$$

$$\Delta_r G^\circ = 2 \Delta_f G^\circ(\text{SO}_{3(g)}) - [2 \Delta_f G^\circ(\text{SO}_{2(g)}) + \Delta_f G^\circ(\text{O}_{2(g)})]$$

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

$$\Delta_r G^\circ = -141.8 \frac{\text{kJ}}{\text{mol}}$$

Step 4: Determine activities for all reactants and products

$$a_{\text{SO}_{2(g)}} = \frac{0.48 \text{ bar}}{1 \text{ bar}} = 0.48$$

$$a_{\text{SO}_{3(g)}} = \frac{0.72 \text{ bar}}{1 \text{ bar}} = 0.72$$

$$a_{\text{O}_{2(g)}} = \frac{0.18 \text{ bar}}{1 \text{ bar}} = 0.18$$

Step 5: Calculate reaction quotient

$$Q = \frac{(a_{\text{SO}_{3(g)}})^2}{(a_{\text{SO}_{2(g)}})^2 (a_{\text{O}_{2(g)}})} = \frac{(0.72)^2}{(0.48)^2 (0.18)} = 12.5$$

but only 2 sig. fig.

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

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

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

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

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

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.