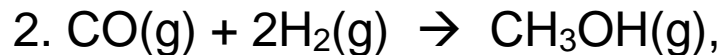


Solution to question 1



$$\begin{aligned}\Delta S^\circ &= \{(3 \times S^\circ \text{CO}_2) + (4 \times S^\circ \text{H}_2\text{O})\} - \{(S^\circ \text{C}_3\text{H}_8) + (5 \times S^\circ \text{O}_2)\} \\ &= \{(3 \times 214) + (4 \times 189)\} \text{ J/mol}\cdot\text{K} - \{270 + (5 \times 205)\} \text{ J/mol}\cdot\text{K} \\ &= (642 + 756) \text{ J/mol}\cdot\text{K} - (270 + 1025) \text{ J/mol}\cdot\text{K} \\ &= 103 \text{ J/mol}\cdot\text{K}\end{aligned}$$



$$\begin{aligned}\Delta S^\circ &= (S^\circ \text{CH}_3\text{OH}) - \{(S^\circ \text{CO}) + (2 \times S^\circ \text{H}_2)\} \\ &= 240 \text{ J/mol}\cdot\text{K} - \{198 \text{ J/mol}\cdot\text{K} + (2 \times 131 \text{ J/mol}\cdot\text{K})\} \\ &= 240 \text{ J/mol}\cdot\text{K} - 460 \text{ J/mol}\cdot\text{K} = -220 \text{ J/mol}\cdot\text{K}\end{aligned}$$

Solution to Question 2

$$\Delta G^\circ \text{C}_3\text{H}_8 = -23.56 \text{ kJ/mol}$$

$$\Delta G^\circ \text{CO}_2 = -394.4 \text{ kJ/mol}$$

$$\Delta G^\circ \text{H}_2\text{O} = -237.2 \text{ kJ/mol}$$

$$\Delta G^\circ \text{O}_2 = 0 \text{ kJ/mol}$$

$$\begin{aligned} \Delta G^\circ &= 3*(-394.4) + 4*(-237.2) - 5*0 - (-23.56) \\ &= -2108,4 \text{ kJ/mol} \end{aligned}$$

The reaction is spontaneous at 298 K.

$$\begin{aligned} \ln K &= -(-2108,4 * 10^3 \text{ J/mol}) / (298 \text{ K} * 8.314 \text{ J/K.mol}) \\ &= 851 \end{aligned}$$

$$K = e^{851} \text{ (reaction goes to completion)}$$

Solution to Question 3

$$\Delta H^\circ = 206 \text{ kJ/mol}$$

$$\Delta S^\circ = 216 \text{ J/mol}\cdot\text{K} = 0.216 \text{ kJ/mol}\cdot\text{K}$$

At 298°C, $T\Delta S^\circ = 298 \text{ K} \times (0.216 \text{ kJ/mol}\cdot\text{K}) = 64.4 \text{ kJ/mol}$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 206 \text{ kJ/mol} - 64.4 \text{ kJ/mol} = 142 \text{ kJ/mol};$$

→ reaction is *nonspontaneous* at 298K.

At 1200 K, $T\Delta S^\circ = 1200 \text{ K} \times (0.216 \text{ kJ/mol}\cdot\text{K}) = 259 \text{ kJ/mol};$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 206 \text{ kJ/mol} - 259 \text{ kJ/mol} = -53 \text{ kJ/mol};$$

→ reaction is spontaneous at 1200 K.

$$\Delta G^\circ = \Delta H^\circ - T_r\Delta S^\circ = 0,$$

$$T_r = 206 \text{ kJ/mol} / (0.216 \text{ kJ/mol}\cdot\text{K}) = 954 \text{ K} = 681^\circ\text{C}$$

Under standard pressure (1 atm), this reaction is not spontaneous below 681°C, but becomes spontaneous above this temperature.



Solution to Question 4

The ratio will decrease.

ΔS is negative (unfavorable) yet the reaction is spontaneous (ΔG° is negative). Thus, ΔH must be negative (exothermic, favorable). Thus, as the temperature is increased, the reaction proceeds to the left, decreasing the ratio of partial pressure of PCl_5 to the partial pressure of PCl_3 .

Solution to Question 5

We have $\text{Fe}_2\text{O}_3(\text{s}) \rightarrow 2\text{Fe}(\text{s}) + 3/2\text{O}_2(\text{g}); \Delta G^\circ = 740 \text{ kJ}$
 and $2\text{Al}(\text{s}) + 3/2\text{O}_2(\text{g}) \rightarrow \text{Al}_2\text{O}_3(\text{s}); \Delta G^\circ = -1582 \text{ kJ}$

For $\text{Fe}_2\text{O}_3(\text{s}) + 2\text{Al}(\text{s}) \rightarrow 2\text{Fe}(\text{s}) + \text{Al}_2\text{O}_3(\text{s})$
 $\Delta G^\circ = 740 + (-1582) = -842 \text{ KJ.}$

Solution to Question 6

1. E_a is the factor the question asks to be solved. Therefore it is much simpler to use

$$\ln k = -E_a/RT + \ln A$$

$$\text{at } 600 \text{ K } \ln 3.4 = -E_a/8.314 \cdot 600 + \ln A \quad (1)$$

$$\text{At } 750 \text{ K } \ln 31.0 = -E_a/8.314 \cdot 750 + \ln A \quad (2)$$

$$(2)-(1) \quad 2.21 = E_a \cdot 0.0000409 = 55127,1 \text{ J/mol} = 55,127 \text{ kJ/mol}$$

2. Use the equation $\ln(k_1/k_2) = -E_a/R(1/T_1 - 1/T_2)$

$$\ln(15/7) = -[(600 \times 1000)/8.314](1/T_1 - 1/389)$$

$$T_1 = 390.6 \text{ K}$$

Solution to Question 7

$$K = \frac{[\text{N}_2\text{O}_4]^1}{[\text{NO}_2]^2}$$

$$K = \frac{[0.0869]^1}{[0.025]^2} = \frac{[0.0869]}{[0.000625]} = 138$$

$$K = 138 = \frac{[0.12]^1}{[x]^2}$$

$$[x]^2 = \frac{[0.12]}{138}$$

$$[x] = \sqrt{\frac{[0.12]}{138}} = \sqrt{0.000869} = 0.029 \text{ mole / liter}$$

Solution to Question 8

$$1 \quad K_p = \frac{P_{NH_3}^2}{P_{H_2}^3 P_{N_2}}$$

2. Under standard condition,

$$P_{N_2} = P_{H_2} = P_{NH_3} = 1 \text{ atm}, K_p = 1;$$

$$\ln K_p = 0, \text{ and } \Delta G = \Delta G^\circ = -33 \text{ kJ/mol}$$

$$3. K_p = 5^2 / (5 \times 15^3) = 1.5 \times 10^{-3}$$

$$\ln K_p = \ln(1.5 \times 10^{-3}) = -6.5$$

$$\Delta G = \Delta G^\circ + RT \ln K_p;$$

$$= 12 \text{ kJ/mol} + (0.008314 \text{ kJ/mol.K} \times 523 \text{ K} \times (-6.5))$$

$$= 12 \text{ kJ/mol} - 28 \text{ kJ/mol} = -16 \text{ kJ/mol}$$

→ spontaneous reaction



Solution to Question 9

1. Adding products creates a net change in the reverse direction, toward reactants. The opposite occurs when adding more reactants.
2. When the volume of a mixture is reduced, a net change occurs in the direction that produces fewer moles of gas towards products. When volume is increased the change occurs in the direction that produces more moles of gas towards reactants.
3. Adding an inert gas into a gas-phase equilibrium at constant volume does not result in a shift. This is because the addition of a non-reactive gas does not change the partial pressures of the other gases in the container. While the total pressure of the system increases, the total pressure does not have any effect on the equilibrium constant.

Solution to Question 10

Because the reaction rate is extremely low at this temperature.