Marks
5

- At 700 °C, hydrogen and iodine react according to the following equation.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K_c = 49.0$$

Hydrogen also reacts with sulfur at 700 °C:

$$2H_2(g) + S_2(g) \rightleftharpoons 2H_2S(g) \qquad K_c = 1.075 \times 10^8$$

Determine K_c for the following overall equilibrium reaction at 700 °C.

$$2I_2(g) + 2H_2S(g) \rightleftharpoons S_2(g) + 4HI(g)$$

The overall reaction corresponds to the twice the first reaction combined with the reverse of the second reaction:

$\mathbf{2H_2(g) + 2I_2(g) \rightleftharpoons 4HI(g)}$	$\mathbf{K_c(1) = (49.0)^2}$
$\mathbf{2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)}$	$\mathbf{K_c(2) = 1/(1.075 \times 10^8)}$
$\mathbf{2I_2(g) + 2H_2S(g) \rightleftharpoons S_2(g) + 4HI(g)}$	$\mathbf{K_c(3) = K_c(1) \times K_c(2)}$

The 1st reaction is doubled so the original equilibrium constant is squared.

The 2nd reaction is reversed so the reciprocal of the equilibrium constant is used.

The two reactions are then combined and the overall equilibrium constant is then the product:

$$\mathbf{K_c(3) = K_c(1) \times K_c(2) = (49.0)^2 \times (1/(1.075 \times 10^8)) = 2.23 \times 10^{-5}}$$

K_c = $\mathbf{2.23 \times 10^{-5}}$

What is the standard free energy change at 700 °C for this overall equilibrium reaction?

The equilibrium constant in terms of pressures is first converted into the equilibrium constant in terms of pressures using $K_p = K_c(RT)^{\Delta n}$. The reaction involves the conversion of 4 mol of gas to 5 mol of gas so Δn = +1 and:

$$\mathbf{K_p = K_c(RT)^{\Delta n} = (2.23 \times 10^{-5}) \times (0.08206 \times 973)^1 = 0.00178}$$

Note that as K_c is in terms of concentration units of mol L^{-1}, R = 0.08206 atm L mol^{-1} K^{-1} has been used.

As $\Delta G° = -RT\ln K_p$:

$$\mathbf{\Delta G° = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) ° (973 \text{ K}) \times \ln(0.00178) = +51.2 \text{ kJ mol}^{-1}}$$

Answer: **+51.2 kJ mol^{-1}**

THIS QUESTION CONTINUES ON THE NEXT PAGE.