

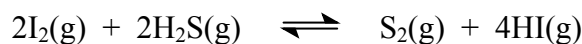
- At 700 °C, hydrogen and iodine react according to the following equation.



Hydrogen also reacts with sulfur at 700 °C:



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



**Marks**  
**5**

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



The 1<sup>st</sup> reaction is doubled so the original equilibrium constant is squared.

The 2<sup>nd</sup> 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:

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

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

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  $\Delta n = +1$  and:**

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

Note that as  $K_c$  is in terms of concentration units of mol L<sup>-1</sup>,  $R = 0.08206$  atm L mol<sup>-1</sup> K<sup>-1</sup> has been used.

As  $\Delta G^\circ = -RT \ln K_p$ :

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

$$\text{Answer: } +51.2 \text{ kJ mol}^{-1}$$

**THIS QUESTION CONTINUES ON THE NEXT PAGE.**