

- Consider the reaction of  $\text{H}_2(\text{g})$  with  $\text{I}_2(\text{g})$  at 298 K to give  $\text{HI}(\text{g})$ .



If partial pressures of 0.20 atm of all three gases are mixed, in which direction will the reaction proceed?

$$\text{The reaction quotient } Q = \frac{(p_{\text{HI}})^2}{(p_{\text{H}_2})(p_{\text{I}_2})} = \frac{(0.20)^2}{(0.20)(0.20)} = 1.0$$

As  $Q < K_p$ , the reaction will proceed towards HI, to increase the partial pressure of HI and decrease the partial pressure of  $\text{H}_2$  and  $\text{I}_2$ , until  $Q = K$ .

Answer: **towards products**

Calculate  $\Delta G^\circ$  for this reaction at 298 K.

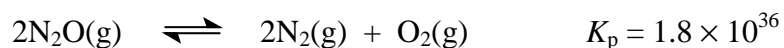
Using  $\Delta G^\circ = -RT \ln K_p$ :

$$\Delta G^\circ = -(8.314) \times (298) \times \ln(2.24) = -2000 \text{ J mol}^{-1} = -2.00 \text{ kJ mol}^{-1}$$

Answer: **-2000 J mol<sup>-1</sup> = -2.00 kJ mol<sup>-1</sup>**

**Marks**  
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- Nitrous oxide decomposes at 25 °C according to the following equation.



What is the value for  $K_p$  at 40 °C?

Using  $\Delta_r G^\circ = -RT \ln K_p$ , at 25 °C (= 298 K):

$$\Delta_r G^\circ = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K}) \times \ln(1.8 \times 10^{36}) = -207 \text{ kJ mol}^{-1}$$

Assuming that  $\Delta_r G^\circ$  does not change over the temperature range, at 40 °C = 313 K:

$$\ln K_p = \frac{\Delta_r G^\circ}{RT} = \frac{(210 \times 10^3 \text{ J mol}^{-1})}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 79$$

$$K_p = 3.3 \times 10^{34}$$

Answer:  $3.3 \times 10^{34}$

Is the reaction endothermic or exothermic? Give a reason for your answer.

**The equilibrium constant decreases when the temperature increases. This is consistent with an exothermic reaction since these become less favourable as the temperature increases (Le Chatelier's principle).**

- The reaction  $2A + B \rightarrow C + 3D$  has reached equilibrium.  
What is the expression for the equilibrium constant,  $K_c$ ?

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$$K_c = \frac{[C][D]^3}{[A]^2[B]}$$

Explain how the equilibrium constant,  $K_c$ , changes when more C is added to the reaction mixture.

**$K_c$  is the equilibrium constant: it does not change.**

**$K_c$  changes only with temperature.**

- The autoionisation of water conforms to the following balanced equation:



Is this an exothermic or endothermic reaction?

**Endothermic (as  $\Delta H$  is positive).**

What will happen to the equilibrium if the temperature is raised?

**The reaction will shift to the right. From le Chatelier's principle, the equilibrium will shift to reduce the effect of the change. As the forward reaction is endothermic, it is able to mitigate the increase in temperature by shifting forwards.**

The equilibrium constant,  $K$ , for this reaction is  $1.8 \times 10^{-16}$  at  $25^\circ\text{C}$ . Calculate  $\Delta G$ .

Using  $\Delta G = -RT \ln K$ ,

$$\Delta G = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K}) \times 1.8 \times 10^{-16} = +89 \text{ kJ mol}^{-1}$$

Answer: **+89 kJ mol<sup>-1</sup>**

Why is  $\Delta G$  not equal to  $\Delta H$  for this reaction?

**By definition,  $\Delta G = \Delta H - T\Delta S$ . As the entropy change for the reaction and the temperature are not zero, so  $\Delta G \neq \Delta H$ .**

The pH of pure water is 6.81 at  $37^\circ\text{C}$ . Is water acidic, basic or neutral at this temperature? Explain.

**Neutral. Pure water is neutral at all temperatures as the chemical equation always gives  $[\text{H}_3\text{O}^+(\text{aq})] = [\text{OH}^-(\text{aq})]$ . A pH value of 7.0 only corresponds to a neutral solution at  $25^\circ\text{C}$ .**