Marks
4

- Consider the results of the following set of experiments studying the rate of the reaction of nitric oxide with hydrogen at 1280 °C.

$$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

Experiment #	[NO] / M	[H_2] / M	Initial Rate / M s^{-1}
1	5.0×10^{-3}	2.0×10^{-3}	1.3×10^{-5}
2	1.0×10^{-2}	2.0×10^{-3}	5.2×10^{-5}
3	1.0×10^{-2}	4.0×10^{-3}	1.0×10^{-4}

Write the rate law expression.

Rate =

Calculate the rate constant, k. Include units in your answer.

k =

What is the rate of the reaction when [NO] is 1.2×10^{-2} M and [H_2] is 6.0×10^{-3} M?

Rate =

- What is the value of the equilibrium constant for the following reaction at 298 K? **2**

$$2Fe^{3+}(aq) + Sn(s) \rightleftharpoons Sn^{2+}(aq) + 2Fe^{2+}(aq)$$

The reduction half cell reactions and E^0 values are:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $E^0 = +0.77$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $E^0 = -0.14$ V

In the reaction, Sn is being oxidized and so the overall cell potential is:

$E^0 = (+0.77) - (-0.14) = +0.91$ V

The reaction involves 2 electrons so, using $E^0 = \frac{RT}{nF}\ln K$:

$$\ln K = E^0 \times \frac{nF}{RT} = (+0.91) \times \left(\frac{2 \times 96485}{8.314 \times 298}\right) = 70.9$$

$K = e^{70.9} = 6.05 \times 10^{30}$

Answer: **6.05×10^{30}**