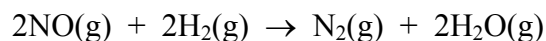


- Consider the results of the following set of experiments studying the rate of the reaction of nitric oxide with hydrogen at 1280 °C.



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.

Between experiments 1 and 2, [H₂] is kept constant. Doubling [NO] (from 5.0×10^{-3} to 1.0×10^{-2} M) leads to the rate quadrupling. The reaction is second order with respect to [NO].

Between experiments 2 and 3, [NO] is kept constant. Doubling [H₂] (from 2.0×10^{-3} to 4.0×10^{-3} M) leads to the rate doubling. The reaction is first order with respect to [H₂]. Thus,

$$\text{rate} \propto [\text{NO}]^2[\text{H}_2] = k[\text{NO}]^2[\text{H}_2]$$

$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

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

Using experiment 1 and rate = $k[\text{NO}]^2[\text{H}_2]$:

$$(1.3 \times 10^{-5} \text{ M s}^{-1}) = k \times (5.0 \times 10^{-3} \text{ M})^2 \times (2.0 \times 10^{-3} \text{ M}) \quad \text{so } k = 260 \text{ M}^{-2} \text{ s}^{-1}$$

$$(\text{M s}^{-1}) = (\text{units of } k) \times (\text{M})^2 \times (\text{M}) \quad \text{so the units of } k \text{ are } \text{M}^{-2} \text{ s}^{-1}$$

$$k = 260 \text{ M}^{-2} \text{ s}^{-1}$$

What is the rate of the reaction when [NO] is 1.2×10^{-2} M and [H₂] is 6.0×10^{-3} M?

Using rate = $(260 \text{ M}^{-2} \text{ s}^{-1})[\text{NO}]^2[\text{H}_2]$:

$$\text{rate} = (260 \text{ M}^{-2} \text{ s}^{-1}) \times (1.2 \times 10^{-2} \text{ M})^2 \times (6.0 \times 10^{-3} \text{ M}) = 2.2 \times 10^{-4} \text{ M s}^{-1}$$

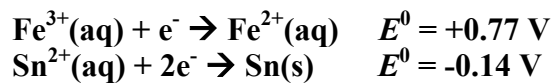
$$\text{Rate} = 2.2 \times 10^{-4} \text{ M s}^{-1}$$

Marks
4

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



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



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

$$E^0 = ((+0.77) - (-0.14)) \text{ V} = +0.91 \text{ V}$$

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

$$\ln K = E^0 \times \frac{nF}{RT} = (+0.91 \text{ V}) \times \left(\frac{2 \times 96485 \text{ C mol}^{-1}}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}} \right) = 70.9$$

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

Answer: 6.05×10^{30}