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.

Between experiments 1 and 2, $[H_2]$ 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 $[H_2]$ (from 2.0×10^{-3} to 4.0×10^{-3} M) leads to the rate doubling. The reaction is first order with respect to $[H_2]$. Thus,

rate α $[NO]^2[H_2] = k[NO]^2[H_2]$

Rate = $\boldsymbol{k[NO]^2[H_2]}$

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

Using experiment 1 and rate = $k[NO]^2[H_2]$:

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

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

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

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

Using rate = $(260\ M^{-2}\ s^{1})[NO]^2[H_2]$:

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

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

- 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:

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

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

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

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}**