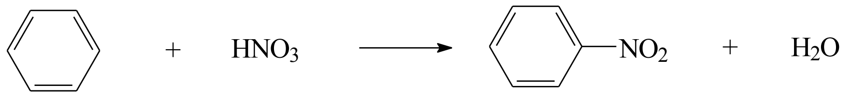
Marks 5

- The nitration of benzene to form nitrobenzene may be written with the following stoichiometry.



The reaction was performed in the presence of excess concentrated sulfuric acid and the following data were obtained.

Experiment number	initial [benzene] (M)	initial [nitric acid] (M)	[nitrobenzene] (M) after 100 s
1	0.010	1.0	1.2×10^{-4}
2	0.020	1.0	2.4×10^{-4}
3	0.020	0.50	1.2×10^{-4}

Determine the rate of the reaction for Experiment 1.

As rate = $\dfrac{\textbf{change in concentration}}{\textbf{change in time}} = \dfrac{\Delta[\textbf{nitrobenzene}]}{\Delta t}$**, for experiment 1:**

$$\textbf{rate} = \frac{[\text{nitrobenzene}]_{t_2} - [\text{nitrobenzene}]_{t_1}}{(t_2 - t_1)} = \frac{(1.2\times10^{-4} - 0)\,\text{M}}{(100-0)\,\text{s}} = 1.2\times10^{-6}\ \text{M s}^{-1}$$

Answer: **1.2×10^{-6} M s^{-1}**

What is the rate equation for this reaction?

Between experiments 1 and 2, [nitric acid] is kept constant. [Benzene] is doubled and this leads to a doubling in the [nitrobenzene] produced after 100 s. Between experiments 2 and 3, [benzene] is kept constant. [Nitric acid] is halved and this leads to a halving in the [nitrobenzene] produced after 100 s. Thus,

rate α [benzene][nitric acid] = *k*[benzene][nitric acid]

Rate = ***k*[benzene][nitric acid]**

What is the value of the rate constant?

As rate = *k*[benzene][nitric acid], for experiment 1 so *k* and its units are:

1.2×10^{-6} M s^{-1} = $k \times (0.010\ \text{M}) \times (1.0\ \text{M})$ so k = 1.2×10^{-4} M s^{-1}

(M s^{-1}) = (units of *k*) × (M) × (M) so the units of *k* are M^{-1} s^{-1}

k = **1.2×10^{-4} M^{-1} s^{-1}**