Marks
4

- Consider the results of the following set of experiments studying the rate of the chemical reaction: $2A + B \rightarrow 3C + D$

Experiment #	initial [A] / M	initial [B] / M	Rate / M hr^{-1}
1	0.240	0.120	2.00
2	0.120	0.120	0.500
3	0.240	0.060	1.00

Write the rate law expression.

Between experiment 1 and 2, [B] is kept constant. [A] is halved which causes the rate to be reduced by a factor of four. The rate is second order with respect to [A].
Between experiment 1 and 3, [A] is kept constant. [B] is halved which causes the rate to halve. The rate is first order with respect to [B]. Thus,

rate α $[A]^2[B] = k[A]^2[B]$

Rate = $k[A]^2[B]$

Calculate the rate constant, k, with units.

Using experiment 1 and rate = $k[A]^2[B]$**:**

$(2.00 \text{ M hr}^{-1}) = k \times (0.240 \text{ M})^2 \times (0.120 \text{ M})$ **so** $k = 289 \text{ M}^{-2} \text{ hr}^{-1}$

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

$k = 289 \text{ M}^{-2} \text{ hr}^{-1}$

What is the rate of the reaction when [A] is 0.0140 M and [B] is 1.35 M?

Using rate = $(289 \text{ M}^{-2} \text{ hr}^{-1})[A]^2[B]$**, the rate is:**

$\text{rate} = (289 \text{ M}^{-2} \text{ hr}^{-1}) \times (0.0140 \text{ M})^2 \times (1.35 \text{ M})$
$= 0.0766 \text{ M hr}^{-1} = 7.66 \times 10^{-3} \text{ M hr}^{1}$

Rate = $7.66 \times 10^{-3} \text{ M hr}^{-1}$