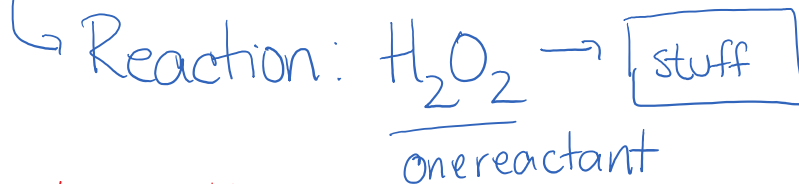


A 2.32 M solution of hydrogen peroxide is allowed to decompose. After 1200 s, what will $[H_2O_2]$ be, if $k = 7.30 \times 10^{-4} s^{-1}$ for the decomposition?

rate law:

$$\text{rate} = k [H_2O_2]^m$$

units for k : (s^{-1}) matches with
1st order reaction



I.R.L. for 1st order: $\ln [A]_t = -kt + \ln [A]_0$

$$[A]_0 = 2.32 \text{ M}$$

$$k = 7.30 \times 10^{-4} s^{-1}$$

$$t = 1200 \text{ s}$$

agreement of units!

$$[A]_t = ??$$

$$\ln [A]_t = - (7.30 \times 10^{-4} s^{-1}) (1200 s) + \ln (2.32 \text{ M})$$

$$\ln [A]_{1200} = -0.876 + 0.84157$$

$$\ln [A]_{1200} = -0.0344$$

$$[A]_{1200} = e^{-0.0344}$$

$$[A]_{1200} = 0.966 \text{ M}$$

For the decomposition of peroxide, $k = 7.30 \times 10^{-4} s^{-1}$ for a certain set of conditions. If the peroxide concentration is initially 2.32 M, how long will it take for 90% of the peroxide to decompose?

- This reaction is 1st order (from units of k)
- Looking for t when $[H_2O_2]$ reaches a certain value
 → use integrated rate law (1st order)

$$\ln \underbrace{[A]_0}_{2.32 \text{ M}} = - \underbrace{k}_{7.30 \times 10^{-4} s^{-1}} t + \ln [A]_t$$

"90% decomposed"

- Looking for time when H_2O_2 is 90% decomposed

This means 10% of the peroxide is left!

10% of 2.32 M is 0.232 M = $[H_2O_2]_t$

solving:

$$\ln [0.232 \text{ M}] = - (7.30 \times 10^{-4} \text{ s}^{-1}) t + \ln [2.32 \text{ M}]$$

$$\underbrace{-1.461018}_{\text{log with 3 sig. figs ends here}} = \underbrace{(-7.30 \times 10^{-4} \text{ s}^{-1}) t}_{\text{not significant digit in a logarithm!}} + \underbrace{0.841567}_{\text{logs are unitless}}$$

log with 3 sig. figs ends here

$$-2.302585 = (-7.30 \times 10^{-4} \text{ s}^{-1}) t$$

$$3154.23 \text{ s} = t \quad (3154.23 \text{ s} \times \frac{1 \text{ min}}{60 \text{ s}} = 52.5704 \text{ min})$$

H_2O_2 will be 90% decomposed
after $3.15 \times 10^3 \text{ s}$ (52.3 min)