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?

Reaction: $H_2O_2 \rightarrow \boxed{\text{stuff}}$

one reactant

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: $\ln[A]_t = -Kt + \ln[A]_0$

$[A]_0 = 2.32\ M$

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

$t = 1200\ s$

agreement of units!

$[A]_t = ??$

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

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

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

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

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

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)

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

"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: $\ln[0.232\ M] = -(7.30\times10^{-4}\,s^{-1})t + \ln[2.32\ M]$

NOT significant digit in a logarithm!

$-1.461018 = (-7.30\times10^{-4}\,s^{-1})t + 0.841567$ ← logs are unitless

log with 3 sig. figs ends here

$-2.302585 = (-7.30\times10^{-4}\,s^{-1})t$

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

H_2O_2 will be 90% decomposed after $3.15\times10^{3}\ s$ (52.3 min)