Practice Problems For Unit III

A bottle of hydrogen peroxide, H_2O_2, slowly decomposes to produce water and oxygen

$$H_2O_2(aq) \rightarrow H_2O(aq) + \tfrac{1}{2}O_2(g)$$

The following data were recorded in an experimental study of the kinetics of this decomposition reaction (a small amount of I^- was added as a catalyst to make the reaction go faster).

Time (s)	$[H_2O_2]$ (M)	Time (s)	$[H_2O_2]$ (M)	Time (s)	$[H_2O_2]$ (M)
0	0.882	240	0.372	480	0.152
60	0.697	300	0.298	540	0.120
120	0.566	360	0.236	600	0.094
180	0.458	420	0.188	660	???

A graph of $[H_2O_2]$ vs. time is shown on the back of this page.

1. What is the average rate for the period in which the reaction is monitored?

   The average rate is $-\Delta[H_2O_2]/\Delta t$ for the time period in question; thus, the average rate in this case is $-(0.094 - 0.882)/(600 - 0) = 1.31\times10^{-3}$ M/s.

2. Estimate the instantaneous rate at $t = 60$ s?

   The instantaneous rate is the slope of the tangent to the curve of $[H_2O_2]$ vs. time at the time $t = 60$ s. Using the data in the figure gives the rate as 2.8×10^{-3} M/s.

3. What is the rate law for this reaction, including the value of the rate constant?

   Plotting $\ln[H_2O_2]$ vs. time gives a straight line with a slope of 0.00370 s^{-1}; thus, the reaction is first-order. You also could estimate half-lives to arrive at the same result.

4. The table shows ??? as the concentration of H_2O_2 at 660 s. What is the missing value?

   Here you can use the integrated rate law; thus

   $$\ln[H_2O_2]_{660} = \ln(0.882) - (0.00370\ s^{-1})(660\ s)$$

   Solving gives the concentration of H_2O_2 as 7.67×10^{-2} M.

5. Suppose you have a solution of 3.6 M H_2O_2. How long will it take for the concentration to decrease to 0.25 M?

   Again, use the integrated rate law; thus

   $$\ln(0.25) = \ln(3.6) - (0.00370\ s^{-1})t$$

   and solving to get t as 720 s.

6. The concentrated H_2O_2 we purchase is 3.6 M and comes with a warning that it needs to be kept refrigerated. Why do you think that warning is placed on the bottle?

   Reactions go more slowly at lower temperatures. Cooling the peroxide ensures that it won't accidently react too quickly; producing lots of O_2 in a closed container leads to explosive results.