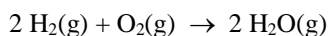


FINAL EXAM U02(M,W,F) Wednesday, December 7, 9:45m to 11:45am, PG5-155
 U03(T,R) Thursday, December 8, 7:15pm to 9:15pm, PG6-116

For problems involving calculations you must show your work for credit. Unless otherwise stated, you may assume $T = 25.0\text{ }^{\circ}\text{C}$.

1) For the irreversible chemical reaction



the following data were found.

time (min)	[H ₂]	[O ₂]	[H ₂ O]
0.0	0.0805	0.0418	0.0000
10.0	0.0781	0.0406	0.0024
20.0	0.0757	0.0394	0.0048
30.0	0.0733	0.0382	0.0071
40.0	0.0711	0.0371	0.0095

Find the average rate of reaction between 20.0 min and 30.0 min. Do this by using the data for H₂, for O₂, and for H₂O, and show that you get the same result for all of these (to within experimental error).

$$\text{For H}_2 \text{ ave rate} = -\frac{1}{2} \frac{\Delta[\text{H}_2]}{\Delta t} = -\frac{1}{2} \frac{(0.0733 - 0.0757) \text{ mol/L}}{(30.0 - 20.0) \text{ min}} = 1.20 \times 10^{-4} \text{ mol/L}\cdot\text{min}$$

$$\text{For O}_2 \text{ ave rate} = -\frac{\Delta[\text{O}_2]}{\Delta t} = -\frac{(0.0382 - 0.0394) \text{ mol/L}}{(30.0 - 20.0) \text{ min}} = 1.20 \times 10^{-4} \text{ mol/L}\cdot\text{min}$$

$$\text{For H}_2\text{O ave rate} = \frac{1}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = \frac{1}{2} \frac{(0.0071 - 0.0048) \text{ mol/L}}{(30.0 - 20.0) \text{ min}} = 1.15 \times 10^{-4} \text{ mol/L}\cdot\text{min}$$

Notice that we get approximately the same answer no matter how we find the average rate, as expected.

2) A particular irreversible chemical reaction obeys the rate law

$$\text{rate} = k [\text{A}]^2 [\text{B}]$$

The reaction is

- a) first order homogeneous
- b) second order homogeneous
- E** c) third order homogeneous
- d) second order heterogeneous
- e) third order heterogeneous

Overall order = 2 + 1 = 3 (third order). Rate depends on more than one concentration, so heterogeneous.

3) For a particular irreversible chemical reaction the initial concentration of A is $[A]_0 = 4.38 \times 10^{-4} \text{ M}$. After 20.0 s, the concentration drops to $[A] = 3.81 \times 10^{-4} \text{ M}$. Assuming the reaction is first order homogeneous ($\text{rate} = k [A]$) find the numerical value for k (including correct units) and the half-life ($t_{1/2}$).

$$[A]_t = [A]_0 e^{-kt} \quad \text{Divide both sides by } [A]_0$$

$$\frac{[A]_t}{[A]_0} = e^{-kt} \quad \text{Take the ln of both sides}$$

$$\ln([A]_t/[A]_0) = -kt \quad \text{Divide both sides by } -t$$

$$k = -\frac{1}{t} \ln([A]_t/[A]_0) = -\frac{1}{20.0 \text{ s}} \ln(3.81/4.38) = 6.97 \times 10^{-3} \text{ s}^{-1}$$

$$\text{Then } t_{1/2} = \frac{\ln(2)}{k} = \frac{\ln(2)}{(6.97 \times 10^{-3} \text{ s}^{-1})} = 99.4 \text{ s}$$

4) A particular chemical reaction obeys the rate law

$$\text{rate} = k [A]^m [B]^n [C]^p$$

The initial rate of reaction is measured for a variety of initial conditions, and the data are given below. Based on this information, find the values for m, n, p, and k (including correct units).

experiment	[A] (mol/L)	[B] (mol/L)	[C] (mol/L)	initial rate (mol/L·s)
1	0.0200	0.0200	0.0100	0.60×10^{-7}
2	0.0200	0.0200	0.0200	1.23×10^{-7}
3	0.0200	0.0400	0.0100	2.37×10^{-7}
4	0.0500	0.0200	0.0200	1.32×10^{-7}

$$\text{Compare 4 and 2} \quad \frac{(0.0500)^m}{(0.0200)^m} = \frac{(1.32 \times 10^{-7})}{(1.23 \times 10^{-7})} \quad \text{so } (2.5)^m = 1.07 \quad m = 0$$

$$\text{Compare 3 and 1} \quad \frac{(0.0400)^n}{(0.0200)^n} = \frac{(2.37 \times 10^{-7})}{(0.60 \times 10^{-7})} \quad \text{so } (2)^n = 3.95 \quad n = 2$$

$$\text{Compare 2 and 1} \quad \frac{(0.0200)^p}{(0.0100)^p} = \frac{(1.23 \times 10^{-7})}{(0.60 \times 10^{-7})} \quad \text{so } (2)^p = 2.05 \quad p = 1$$

So the rate law is $\text{rate} = k [B]^2 [C]$

$$\begin{aligned} \text{And so } k &= \frac{\text{rate}}{[B]^2 [C]} \quad \text{If we use trial 1, } k = \frac{(0.60 \times 10^{-7} \text{ mol/L}\cdot\text{s})}{(0.0200 \text{ mol/L})^2 (0.0100 \text{ mol/L})} \\ &= 1.5 \times 10^{-2} \text{ L}^2/\text{mol}^2\cdot\text{s} \end{aligned}$$