

Part I (4 points each)

1. For the reaction: $3\text{BrO}^- (\text{aq}) \rightarrow \text{BrO}_3^- (\text{aq}) + 2\text{Br}^- (\text{aq})$, what is the rate of reaction, if $\Delta[\text{BrO}^-]/\Delta t = -0.12 \text{ M/min}$?

- a. -0.12 M/min
- b. 0.12 M/min
- c. 0.06 M/min
- d. 0.04 M/min
- e. 0.36 M/min

$$\text{rate of reaction} = -\frac{1}{3} \frac{\Delta[\text{BrO}^-]}{\Delta t} = -\frac{1}{3} \frac{-0.12 \text{ M}}{\text{min}} = 0.04 \text{ M/min}$$

2. Suppose for the above reaction, under different conditions, the initial concentration of BrO^- is 0.51 M . Twenty-one seconds later its concentration is 0.15 M . What is the average rate of change of Br^- concentration?

- a. -0.017 M/s
- b. 0.017 M/s
- c. 0.026 M/s
- d. 0.011 M/s
- e. 0.034 M/s

$$\frac{\Delta[\text{BrO}^-]}{\Delta t} = \frac{0.15 - 0.51 \text{ M}}{21 \text{ s}} = -0.0171$$

$$\text{rate of reaction} = -\frac{1}{3} \frac{\Delta[\text{BrO}^-]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{Br}^-]}{\Delta t}; \quad \frac{\Delta[\text{Br}^-]}{\Delta t} = -\frac{2}{3} \frac{\Delta[\text{BrO}^-]}{\Delta t} = -\frac{2}{3} (-0.0171) = 0.0114 \text{ M/s}$$

3. For the reaction $\text{A(g)} + 2\text{B(g)} \rightarrow \text{C(g)} + \text{D(g)}$, the following experimental data were obtained for the rate. What is the overall order of reaction?

- a. 0 From runs 1 and 2, doubling A at constant B, quadruples the rate.
- b. 1 So the rate of reaction with respect to A is 2nd order. From runs
- c. 2.5 1 and 3, changing B at constant A has no effect. So, the rate of
- d. 2 reaction is 0 order with respect to B. Over all order = $2+0=2$.
- e. 3

Run	[A] (M)	[B] (M)	Rate (M/min)
1	0.025	0.20	2.2
2	0.050	0.20	8.8
3	0.025	0.40	2.2

4. For the above reaction, what is the rate constant, k ?

- a. $3,500 \text{ M}^{-2} \text{ min}^{-1}$
- b. 180 min^{-1}
- c. $180 \text{ M}^{-2} \text{ min}^{-1}$
- d. $180 \text{ M}^{-2} \text{ min}^{-1}$
- e. $2.8 \times 10^{-4} \text{ min}^{-1}$

$$\text{rate of reaction} = k[\text{A}]^2; \quad k = \frac{\text{rate}}{[\text{A}]^2} = \frac{8.8 \text{ M/min}}{(0.05 \text{ M})^2} = 3520 \rightarrow 3,500 \text{ M}^{-2} \text{ min}^{-1}$$

5. Suppose the rate of reaction is second order for $\text{A} \rightarrow \text{B} + \text{C}$. What is the approximate concentration of A (in M) after 55 s if $[\text{A}]_0 = 0.80 \text{ M}$ and $k = 0.035 \text{ M}^{-1} \text{ s}^{-1}$?

- a. 0.1 M
- b. 0.2 M
- c. 0.3 M
- d. 0.4 M
- e. 0.6 M

$$\frac{1}{[\text{A}]} = \frac{1}{[\text{A}]_0} + kt = \frac{1}{0.8 \text{ M}} + (0.035 \text{ M}^{-1} \text{ s}^{-1})(55 \text{ s}) = \frac{3.175}{\text{M}}; \quad [\text{A}] = \frac{\text{M}}{3.175} = 0.31 \text{ M}$$

6. For a certain reaction at 220°C , when the natural log of the reactant concentration is plotted against the time a straight line is obtained whose slope is -0.0016 s^{-1} . The initial reactant concentration is 0.80 M . What is the concentration after 7.5 min?

- a. 0.2 M
- b. 0.4 M
- c. 0.3 M
- d. 0.7 M
- e. 0.8 M

The reaction is first order. So, $\ln[\text{A}] = \ln[\text{A}]_0 - kt$ and $k = -\text{slope}$

$$\ln[\text{A}] = \ln(0.8) - \frac{0.0016}{\text{s}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{7.5 \text{ min}}{1} = -0.943; \quad [\text{A}] = e^{-0.943} = 0.39 \text{ M}$$

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