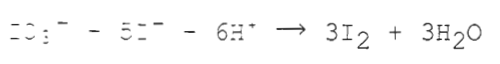


1. 4 Pts For the reaction



the rate of appearance of I_2 at a particular time and concentration is $3.0 \times 10^{-3} \text{ M s}^{-1}$.

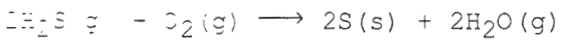
What is the rate of disappearance of I^- in M s^{-1} ?

$$\text{rate} = \frac{-\Delta[\text{IO}_3^-]}{\Delta t} = \frac{-\Delta[\text{I}^-]}{5\Delta t} = \frac{-\Delta[\text{H}^+]}{6\Delta t} = \frac{+\Delta[\text{I}_2]}{3\Delta t} = \frac{+\Delta[\text{H}_2\text{O}]}{\Delta t} = k[\text{IO}_3^-]^x[\text{I}^-]^y[\text{H}^+]^z$$

Since: $\frac{\Delta[\text{I}_2]}{3\Delta t} = \frac{-\Delta[\text{I}^-]}{5\Delta t}$ the $\frac{\Delta[\text{I}^-]}{\Delta t} = \frac{5}{3} \frac{\Delta[\text{I}_2]}{\Delta t}$

So: $\frac{5}{3} (3.0 \times 10^{-3}) = \boxed{5 \times 10^{-3} \text{ M s}^{-1}}$

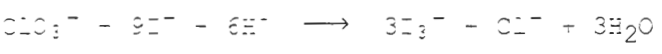
2. 3 Pts From a consideration of the following reaction system



we can conclude

- a. the reaction is second order in H_2S and first order in O_2 .
- b. the reaction is first order in H_2S and second order in O_2 .
- c. rate = $k[\text{H}_2\text{S}]^2[\text{O}_2]$.
- d. rate = $k[\text{H}_2\text{S}][\text{O}_2]$.
- e. none of these conclusions are justified.

3. 6 Pts The rate law for the oxidation of the iodide ion by the chlorate ion in aqueous acidic solution,



has been found to be

$$\text{Rate} = k[\text{ClO}_3^-][\text{I}^-][\text{H}^+]^2$$

What is the overall order for the reaction? 4.

If the time is measured in seconds and the concentration is in M, what are the units for k? units

$$\frac{\text{M}}{\text{s}} = \boxed{\frac{\text{M}^3}{\text{s}}} \text{ M} \cdot \text{M} \cdot \text{M}^2$$

↑
k or $\text{M}^{-3} \text{s}^{-1}$

4. 3 pts For the reaction $aA + bB + cC \rightarrow dD + eE$, the following data were obtained:

Initial Concentrations (mol/L)			Initial Rate (mol/(L · min))
[A]	[B]	[C]	
0.40	0.40	0.20	160
0.20	0.40	0.40	80.
0.60	0.10	0.20	15
0.20	0.10	0.20	5.0
0.20	0.20	0.40	20.
0.20	0.20	0.20	20.

Determine the orders for each reactant, A, B, and C. Then determine the value of the rate constant with proper units.

Use for A: $\frac{\text{rate}}{\text{rate}} = \left(\frac{[A]}{[A]}\right)^x \Rightarrow \frac{15}{5} = \left(\frac{0.60}{0.20}\right)^x \Rightarrow 3 = 3^x \quad x = 1$

Use for C: $\frac{\text{rate}}{\text{rate}} = \left(\frac{[C]}{[C]}\right)^y \Rightarrow \frac{20}{20} = \left(\frac{0.40}{0.20}\right)^y \Rightarrow 1 = (2)^y \quad y = 0$

Use for B: (since C = 0 order): $\frac{\text{rate}}{\text{rate}} = \left(\frac{[B]}{[B]}\right)^z \Rightarrow \frac{20}{5} = \left(\frac{0.20}{0.10}\right)^z \Rightarrow 4 = (2)^z \quad z = 2$

SO rate Law: $\text{rate} = k[A]^1[B]^2[C]^0$

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

$k = \frac{\text{rate}}{[A][B]^2} = \frac{160}{[0.40][0.40]^2} = 2500 \text{ M}^{-2} \text{ min}^{-1}$

4 pts) For the reaction below, which equation represents the rate of reaction?
 $6\text{CH}_2\text{O} + 4\text{NH}_3 \rightarrow (\text{CH}_2)_6\text{N}_4 + 6\text{H}_2\text{O}$

- a) $\text{rate} = (1/6)[\Delta\text{CH}_2\text{O}] / \Delta t$ must be negative
- b) $\text{rate} = -4\Delta[\text{NH}_3] / \Delta t$ must be $\frac{1}{4}$
- c) $\text{rate} = -\Delta[(\text{CH}_2)_6\text{N}_4] / \Delta t$ must be (+)
- d) $\text{rate} = 6\Delta[\text{H}_2\text{O}] / \Delta t$ must be $\frac{1}{6}$
- e) $\text{rate} = (1/6)\Delta[\text{H}_2\text{O}] / \Delta t$