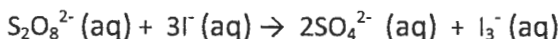


1. The kinetic data shown below was collected for the reaction:



Experiment	$[\text{S}_2\text{O}_8^{2-}]$	$[\text{I}^-]$	Initial rate (M/min)
1	0.0200	0.0155	$1.15 \times 10^{-4}$
2	0.0250	0.0200	$1.85 \times 10^{-4}$
3	0.0300	0.0200	$2.22 \times 10^{-4}$
4	0.0300	0.0275	$3.06 \times 10^{-4}$

a. (8 Pts) Determine the order of the reaction with respect to  $[\text{S}_2\text{O}_8^{2-}]$  and  $[\text{I}^-]$ . Be sure to show setups for each.

The general rate law is:  $\text{rate} = k [\text{S}_2\text{O}_8^{2-}]^x [\text{I}^-]^y$

for  $\text{S}_2\text{O}_8^{2-}$  use Exp 3  $\div$  Exp 2:  $\frac{2.22 \times 10^{-4}}{1.85 \times 10^{-4}} = \frac{[0.0300]^x [0.0200]^y}{[0.0250]^x [0.0200]^y}$

$1.2 = (1.2)^x$  so  $x = 1$

for  $\text{I}^-$  use Exp 4  $\div$  Exp 3:

$\frac{3.06 \times 10^{-4}}{2.22 \times 10^{-4}} = \frac{[0.0300]^x [0.0275]^y}{[0.0300]^x [0.0200]^y}$

$1.38 = (1.38)^y$  so  $y = 1$

and:  $\text{rate} = k [\text{S}_2\text{O}_8^{2-}] [\text{I}^-]$

b. (4 Pts) Calculate the value of the rate constant and determine its units.

$k = \frac{\text{rate}}{[\text{S}_2\text{O}_8^{2-}] [\text{I}^-]} = \frac{1.15 \times 10^{-4} \text{ M/min}}{0.0200 \text{ M} / 0.0155 \text{ M}} = 0.37 \text{ min}^{-1} \text{ M}^{-1}$

2. Given the reaction:  $6\text{I}^-(\text{aq}) + \text{BrO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 3\text{I}_2(\text{aq}) + \text{Br}^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$ .

a. (4 Pts) Write a general rate law.

$\text{rate} = k [\text{I}^-]^x [\text{BrO}_3^-]^y [\text{H}^+]^z$

b. (3 Pts) What can be said about the order of the reaction with respect to  $[\text{BrO}_3^-]$ ?

Order can only be experimentally determined

3. (3 Pts) Determine the units of the rate constant for the rate law:  $\text{rate} = k[\text{A}][\text{B}]^2$  if rate is measured in M/s.

$k = \frac{\text{rate}}{[\text{A}][\text{B}]^2} = \frac{\text{M}}{\text{s} \cdot \text{M} \cdot \text{M}^2} = \text{s}^{-1} \cdot \text{M}^{-2}$

4. (3 Pts) Given the rate law:  $\text{rate} = k[\text{A}][\text{B}]^2$

What would be the effect on the observed rate if the concentration of A is tripled and the concentration of B is doubled?

$\text{rate} = k [3] [2]^2 = 12$

a 12-fold increase in rate