

Key

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \quad \text{or} \quad \log\left(\frac{[A]_t}{[A]_0}\right) = -\frac{kt}{2.303}$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \log[A]_t = -kt/2.303 + \log[A]_0$$

$$1/[A]_t = kt + 1/[A]_0$$

$$k = Ae^{-E_a/(RT)}$$

$$\ln k = \ln A - (E_a / R)(1 / T)$$

$$R = 8.314 \times 10^{-3} \text{ kJ/(mol}\cdot\text{K)}$$

(10 pt) 1. Compound A reacts to give B. The rate law for the reaction is 1st order in A and 0 order in B.

(a) What is the rate law for the reaction $A \rightarrow B$?

$$\text{rate} = k[A] \quad (1 \text{ pt})$$

(b) Explain how you could use integrated rate laws to confirm the rate law for this reaction.

(1 pt) - measure $[A]$ vs t

(1 pt) - Plot $\ln[A]$ vs t

(2 pt) - Getting a straight line with a negative slope confirms the reaction is 1st order

(c) If the rate constant for this reaction is 1.6 min^{-1} at a particular temperature, how long will it take until only 10% of the original A is left?

$$\text{means } \frac{[A]_t}{[A]_0} = 0.10$$

(1 pt) for correct eqn

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \Rightarrow t = -\frac{1}{k} \ln\left(\frac{[A]_t}{[A]_0}\right)$$
$$t = -\frac{1}{1.6 \text{ min}^{-1}} \ln(0.10) = 1.4 \text{ min} \quad (1 \text{ pt})$$

CONTINUED ON BACK →

(10 pt) 2. Initial rates shown below were measured for the following hypothetical reaction:



Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.050	0.15	4.82×10^{-2}
2	0.025	0.15	2.41×10^{-2}
3	0.050	0.30	0.193

- (a) What is the rate law for this reaction? Show your reasoning. (You can assume that only reactants are in the rate law.)

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

to determine x , compare runs 1 + 2

$$\frac{\text{run 1}}{\text{run 2}} = \frac{0.050^x}{0.025^x} = \frac{4.82 \times 10^{-2}}{2.41 \times 10^{-2}} = 2 \quad 2^x = 2$$

so $x = 1$

to determine y , compare runs 1 + 3

$$\frac{\text{run 3}}{\text{run 1}} = \frac{0.30^y}{0.15^y} = \frac{0.193}{4.82 \times 10^{-2}} = 4 \quad 2^y = 4$$

so $y = 2$

Overall $\boxed{\text{rate} = k[A][B]^2}$ (1 pt)

- (b) What is the value of the rate constant for this reaction? Be sure to include units with your answer.

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

$$k = \frac{\text{rate}}{[A][B]^2} = \frac{4.82 \times 10^{-2} \text{ M/s}}{(0.050 \text{ M})(0.15 \text{ M})^2}$$

give full credit if do OK
but have wrong rate law from (a)

$$= 43 \frac{1}{\text{M}^2 \cdot \text{s}}$$

(1 pt)

KEY

Name _____

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \quad \text{or} \quad \log\left(\frac{[A]_t}{[A]_0}\right) = -\frac{kt}{2.303}$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \log[A]_t = -kt/2.303 + \log[A]_0$$

$$1/[A]_t = kt + 1/[A]_0$$

$$k = Ae^{-E_a/(RT)}$$

$$\ln k = \ln A - (E_a / R)(1 / T)$$

$$R = 8.314 \times 10^{-3} \text{ kJ/(mol}\cdot\text{K)}$$

(10 pt) 1. Compound A reacts to give B. The rate law for the reaction is 2nd order in A and 0 order in B.

(a) What is the rate law for the reaction A → B?

$$\text{rate} = k[A]^2 \quad (1 \text{ pt})$$

(b) Explain how you could use integrated rate laws to confirm the rate law for this reaction.

(1 pt) - measure [A] vs t

(1 pt) - plot $1/[A]$ vs t

(2 pt) - getting a straight line with a positive slope confirms the reaction is second order

(c) If the rate constant for this reaction is $2.5 \text{ M}^{-1} \text{ min}^{-1}$ at a particular temperature and the initial concentration is 0.50 M , how long will it take until only 0.010 M A remains?

$$\begin{aligned} \frac{1}{[A]_t} &= kt + \frac{1}{[A]_0} \\ (2 \text{ pt}) \quad t &= \frac{1}{k} \left(\frac{1}{[A]_t} - \frac{1}{[A]_0} \right) = \frac{1}{2.5 \text{ M}^{-1} \text{ min}^{-1}} \left(\frac{1}{0.010 \text{ M}} - \frac{1}{0.50 \text{ M}} \right) \\ &= \boxed{39 \text{ min}} \end{aligned}$$

(1 pt) for showing correct eqn

(1 pt) (1 pt)

CONTINUED ON BACK →

(10 pt) 2. Initial rates shown below were measured for the following hypothetical reaction:



Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.60	0.15	1.31×10^{-3}
2	0.15	0.30	6.58×10^{-4}
3	0.60	0.30	2.63×10^{-3}

(a) What is the rate law for this reaction? Show your reasoning. (You can assume that only reactants are in the rate law.)

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

to determine x compare runs 2 & 3

$$\frac{\text{run 3}}{\text{run 2}} = \frac{0.60^x}{0.15^x} = \frac{2.63 \times 10^{-3}}{6.58 \times 10^{-4}} = 4 \quad 4^x = 4 \quad \text{so } x = 1$$

to determine y compare runs 1 & 3

$$\frac{\text{run 3}}{\text{run 1}} = \frac{0.30^y}{0.15^y} = \frac{2.63 \times 10^{-3}}{1.31 \times 10^{-3}} = 2 \quad 2^y = 2 \quad \text{so } y = 1$$

$$\boxed{\text{rate} = k[A][B]} \quad (1 \text{ pt})$$

(b) What is the value of the rate constant for this reaction? Be sure to include units with your answer.

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

$$k = \frac{\text{rate}}{[A][B]} = \frac{1.31 \times 10^{-3} \text{ M/s}}{(0.60 \text{ M})(0.15 \text{ M})} = 1.5 \times 10^{-2} \frac{1}{\text{s} \cdot \text{M}}$$

(1 pt) (1 pt)

give full credit if do OK
but have wrong rate law
from part (a)

Key

Name _____

Give partial credit

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \quad \text{or} \quad \log\left(\frac{[A]_t}{[A]_0}\right) = -\frac{kt}{2.303}$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \log[A]_t = -kt/2.303 + \log[A]_0$$

$$1/[A]_t = kt + 1/[A]_0$$

$$k = Ae^{-E_a/(RT)}$$

$$\ln k = \ln A - (E_a / R)(1/T)$$

$$R = 8.314 \times 10^{-3} \text{ kJ/(mol}\cdot\text{K)}$$

(10 pt) 1. Compound A reacts to give B. The rate law for the reaction is 1st order in A and 0 order in B.

(a) What is the rate law for the reaction $A \rightarrow B$?

$$\text{rate} = k[A] \quad (1 \text{ pt})$$

(b) Explain how you could use integrated rate laws to confirm the rate law for this reaction.

(1 pt) - measure $[A]$ vs t

(1 pt) - plot $\ln[A]$ vs t

(2 pt) - getting a straight line with a negative slope confirms the reaction is first order

(c) If the rate constant for this reaction is 0.055 min^{-1} at a particular temperature, how long will it take for 25% of A to react?

↳ means 75% remains and
 $\ln\left(\frac{[A]_t}{[A]_0}\right) = 0.75$ (1 pt) for correct eqn

$$\left\{ \begin{aligned} \ln\left(\frac{[A]_t}{[A]_0}\right) &= -kt \rightarrow t = -\frac{1}{k} \ln\left(\frac{[A]_t}{[A]_0}\right) \\ t &= -\frac{1}{0.055 \text{ min}^{-1}} \ln(0.75) = \boxed{5.2 \text{ min}} \end{aligned} \right. \quad (1 \text{ pt}) \quad (1 \text{ pt})$$

CONTINUED ON BACK ⇒

- (10 pt) 2. Initial rates shown below were measured for the following hypothetical reaction:



Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.050	0.10	4.82×10^{-2}
2	0.050	0.20	4.82×10^{-2}
3	0.025	0.20	1.21×10^{-2}

- (a) What is the rate law for this reaction? Show your reasoning. (You can assume that only reactants are in the rate law.)

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

to determine x , compare runs 2 + 3

$$\frac{\text{run 2}}{\text{run 3}} = \frac{0.050^x}{0.025^x} = \frac{4.82 \times 10^{-2}}{1.21 \times 10^{-2}} = 4 \quad 2^x = 4$$

$$= 2^x \quad \underline{x = 2}$$

to determine y , compare runs 1 + 2

$$\frac{\text{run 2}}{\text{run 1}} = \frac{0.20^y}{0.10^y} = \frac{4.82 \times 10^{-2}}{4.82 \times 10^{-2}} = 1 \quad 2^y = 1$$

$$= 2^y \quad \underline{y = 0}$$

overall
$$\text{rate} = k[A]^2[B]^0 = k[A]^2$$

- (b) What is the value of the rate constant for this reaction? Be sure to include units with your answer.

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

$$k = \frac{\text{rate}}{[A]^2} = \frac{4.82 \times 10^{-2} \text{ M/s}}{(0.050 \text{ M})^2} = 19. \frac{1}{\text{M} \cdot \text{s}}$$

give full credit
 if do OK but
 have wrong rate law
 from (a)

KEY

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \quad \text{or} \quad \log\left(\frac{[A]_t}{[A]_0}\right) = -\frac{kt}{2.303}$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \log[A]_t = -kt/2.303 + \log[A]_0$$

$$1/[A]_t = kt + 1/[A]_0$$

$$k = Ae^{-E_a/(RT)}$$

$$\ln k = \ln A - (E_a / R)(1 / T)$$

$$R = 8.314 \times 10^{-3} \text{ kJ}/(\text{mol}\cdot\text{K})$$

(10 pt) 1. Compound A reacts to give B. The rate law for the reaction is 2nd order in A and 0 order in B.

(a) What is the rate law for the reaction A → B?

$$\text{rate} = k[A]^2 \quad (1 \text{ pt})$$

(b) Explain how you could use integrated rate laws to confirm the rate law for this reaction.

(1 pt) - measure [A] vs t

(1 pt) - plot $\frac{1}{[A]}$ vs t

(2 pt) - getting a straight line with a positive slope confirms the rxn is second order

(c) If the rate constant for this reaction is $0.28 \text{ M}^{-1} \text{ hr}^{-1}$ at a particular temperature and the initial concentration is 0.10 M , what will be the concentration of A after 48 hrs?

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0} = (0.28 \frac{1}{\text{M}\cdot\text{hr}})(48 \text{ hr}) + \frac{1}{0.10 \text{ M}}$$

(1 pt) for correct eqn

$$\frac{1}{[A]_t} = 23.44 \frac{1}{\text{M}}$$
$$[A]_t = \frac{1}{23.44 \frac{1}{\text{M}}} = 0.043 \text{ M}$$

(2 pt) { (1 pt) (1 pt)

(10 pt) 2. Initial rates shown below were measured for the following hypothetical reaction:



Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.30	0.20	6.52×10^{-2}
2	0.10	0.050	1.81×10^{-3}
3	0.10	0.20	7.24×10^{-3}

(a) What is the rate law for this reaction? Show your reasoning. (You can assume that only reactants are in the rate law.)

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

(3 pt) to determine x compare runs 1 & 3

$$\frac{\text{run 1}}{\text{run 3}} = \frac{0.30^x}{0.10^x} = 3^x = \frac{6.52 \times 10^{-2}}{7.24 \times 10^{-3}} = 9 \quad \begin{matrix} 3^x = 9 \\ x = 2 \end{matrix}$$

(3 pt) to determine y compare runs 2 & 3

$$\frac{\text{run 3}}{\text{run 2}} = \frac{0.20^y}{0.050^y} = 4^y = \frac{7.24 \times 10^{-3}}{1.81 \times 10^{-3}} = 4 \quad \begin{matrix} 4^y = 4 \\ y = 1 \end{matrix}$$

$$\boxed{\text{rate} = k [A]^2 [B]}$$
 (1 pt)

(b) What is the value of the rate constant for this reaction? Be sure to include units with your answer.

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

$$k = \frac{\text{rate}}{[A]^2 [B]} = \frac{6.52 \times 10^{-2} \text{ M/s}}{(0.30 \text{ M})^2 (0.20 \text{ M})}$$
 (1 pt)

give full credit if do OK but have wrong rate law from (a)

$$= \boxed{3.6 \frac{1}{\text{M}^2 \text{s}}}$$
 (1 pt)