

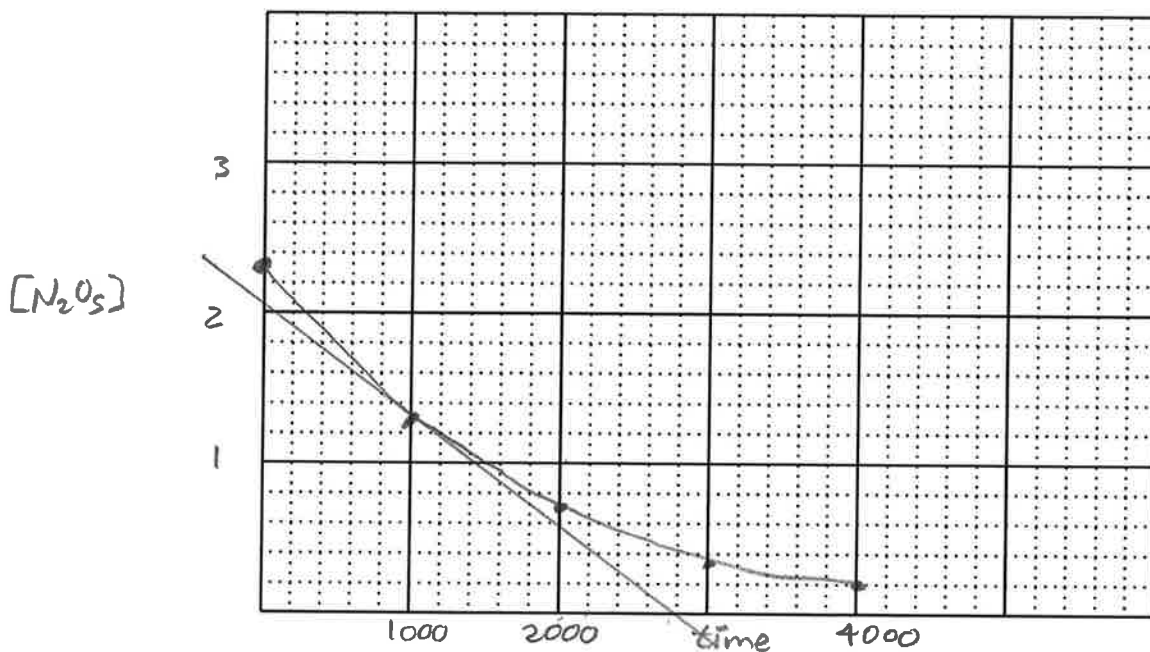
Chapter 13 Number 1 (13.1 – 13.3) (Unit 1) 19 January 2018



| time | [N ₂ O ₅] | [NO ₂] | [O ₂] |
|--------|----------------------------------|--------------------|-------------------|
| 0 s | 2.330 mol/L | 0 | 0 |
| 1000 s | 1.260 | | |
| 2000 s | 0.681 | | |
| 3000 s | 0.369 | | |
| 4000 s | 0.199 | | |

See notes

1. Graph these data with time on the x-axis and [N₂O₅], [NO₂], and [O₂] on the y-axis.



See notes

2. Is this reaction slowing down or speeding up with time? Circle **Slowing down** or **Speeding up**

3. Sketch a tangent line on the graph and estimate the rate at t = 1000 s in terms of rate = $-\Delta[\text{N}_2\text{O}_5]/\Delta t$. Make sure to use the right units.

$$m = \frac{\Delta y}{\Delta x} = \frac{2.0 - 0}{0 - 2800}$$

$$m = -7.1 \times 10^{-4}$$

$$\text{rate} = +7.1 \times 10^{-4} \text{ mol/L}\cdot\text{s}$$

4. Using your value from Question 3, what is the rate in terms of rate = $\Delta[\text{NO}_2]/\Delta t$ and rate = $\Delta[\text{O}_2]/\Delta t$

$$\text{rate} = \frac{7.1 \times 10^{-4} \text{ mol N}_2\text{O}_5 / \text{L}\cdot\text{s}}{2 \text{ mol N}_2\text{O}_5} \times \frac{1 \text{ mol O}_2}{2 \text{ mol N}_2\text{O}_5}$$

$$= 3.6 \times 10^{-4} \text{ mol O}_2 / \text{L}\cdot\text{s}$$

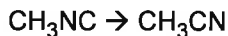
5. Would the rate of the reaction increase or decrease if one started with a larger [N₂O₅]?

Now try these problems from the book:
 Section 13.1. (Rates) Problems 1, 2, 39, 52 and 54
 Section 13.2. (Rate Law and Order) Problems 3 and 4
 Section 13.3. (Initial rate method) Problems 5, 6, 7, 8, 40, 46, 60, 66 and 124

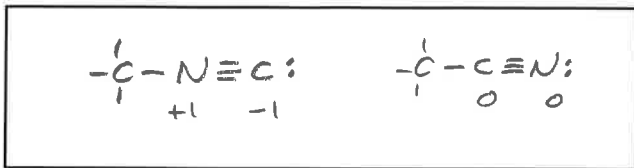
Chapter 13 Number 2 (13.4 – 13.6) (Unit 1)

22 January 2018

1. Consider this reaction in which the nitrogen atom rearranges:



1a. Sketch a Lewis dot structure for both reactant and product. Assign formal charges to each atom.



1b. Use this data to determine the order of the reaction and the value of the rate constant with proper units. Write the rate expression (law).

| | Initial [CH ₃ NC] mol/L | Initial rate = -Δ[CH ₃ NC]/Δt |
|--------|---------------------------------------|---|
| Expt 1 | 0.0439 mol/L | 2.24 × 10 ⁻⁶ mol/Ls |
| Expt 2 | 0.0394 mol/L | 2.01 × 10 ⁻⁶ mol/Ls |
| Expt 3 | 0.0722 mol/L | 3.69 × 10 ⁻⁶ mol/Ls |

Expt 1 & 2

$$\frac{2.24 \times 10^{-6}}{2.01 \times 10^{-6}} = \frac{k(0.0439)^{0.12}}{k(0.0394)^{0.12}}$$

$$1.11 = 1.11^{0.12} \quad \text{1st order}$$

$$3.69 \times 10^{-6} = k(0.0722) \quad k = 5.11 \times 10^{-5} \text{ L/mol s}$$

1c. What is the rate with units if [CH₃NC] = 0.100 M?

$$\text{rate} = k \cdot [\text{CH}_3\text{NC}]^1$$

$$= 5.11 \times 10^{-5} \times 0.100 = 5.11 \times 10^{-6} \text{ mol/L s}$$

1d. Express the rate in terms of rate = +Δ[CH₃CN]/Δt

Same

$$\text{rate} = \frac{\Delta[\text{CH}_3\text{CN}]}{\Delta t} = k[\text{CH}_3\text{NC}]^1$$

1e. Use the appropriate time concentration expression to determine what the concentration will be after 10.00 hr, given [CH₃NC]_{t=0} = 0.200 M

$$\ln \frac{0.200}{[A]_t} = k \cdot 10 \times 3600 \text{ s}$$

$$\ln \frac{0.200}{[A]_t} = 1.84$$

$$\frac{0.200}{[A]_t} = 6.30$$

$$[A]_t = 3.18 \times 10^{-2} \text{ mol/L}$$

2. Consider the reaction of 2 HI(g) → H₂(g) + I₂(g).

Use this data to determine the order of the reaction and the value of the rate constant with proper units. Write the rate expression (law).

| Time (min) | [HI] |
|------------|--------|
| 0 | 0.1000 |
| 100 | 0.0741 |
| 200 | 0.0588 |
| 300 | 0.0488 |
| 400 | 0.0417 |
| 500 | 0.0364 |

Using the spreadsheet at the 205 website, the best trendline is for 2nd order (R² = 1.00) from trendline k = 0.0349

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

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0} \quad \frac{1}{0.488} = k \cdot 300 + \frac{1}{0.1}$$

$$k = 0.0350 \text{ L/mol min}$$

Questions in final exam format:

3. The decomposition of ammonia to nitrogen and hydrogen on a tungsten filament at 800°C is independent of the concentration of ammonia at high pressures of ammonia. What is the order of the reaction with respect to ammonia?

- A. zero B. first C. second D. third

4. Using the method of initial rates for the reaction A → B, if the initial concentration of A is doubled and the rate of reaction quadruples, what is the order of reaction with respect to A?

- A. zero B. first C. second D. fourth

5. The first-order decomposition of hydrogen peroxide occurs according to the equation



Which plot will produce a straight line?

- A. [H₂O₂] versus time B. [H₂O₂]₂ versus time
C. 1/[H₂O₂] versus time D. ln[H₂O₂] versus time

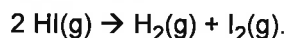
Now try these problems from the book:

- Section 13.4. (Zero order) Problems 9, 10,
Section 13.5. (First order) Problems 11, 12, 13, 14, 15, 16, 42, 68 and 70.
Section 13.6. (Second order) Problems 17, 18, 44, 72, 74, 76 and 80

Chapter 13 Number 3 (13.7 – 13.8) (Unit 1)

24 January 2018

1. The following data were presented in the previous worksheet and led to the conclusion that the best fit was for second order.



Use this data to determine the value of the rate constant with proper units. Write the rate expression (law).

| Time (min) | [HI] |
|------------|--------|
| 0 | 0.1000 |
| 100 | 0.0741 |
| 200 | 0.0588 |
| 300 | 0.0488 |
| 400 | 0.0417 |
| 500 | 0.0364 |

$$k = \frac{1}{0.0488} = k_{300} + \frac{1}{0.10}$$

$$20.49 = k_{300} + 10$$

$$k = 0.0350 \text{ L/mol min}$$

1a. How long does it take for [HI] = 0.0600 M?

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

$$\frac{1}{0.06} = 0.0350t + 10$$

$$t = 190 \text{ min.}$$

1b. What is [HI] after 600 min?

$$\frac{1}{[A]_t} = 0.0350 \times 600 + 10$$

$$[A]_t = 0.0323 \text{ mol/L}$$

2. A first-order reaction has a half-life of 37.0 min. How long does it take for the initial concentration of reactant, A, to decrease by 80% from its original concentration?

$$t_{1/2} \rightarrow k \rightarrow \text{time-conc eqn}$$

$$37.0 \text{ min} = \frac{0.693}{k} \quad k = 0.0187 \text{ min}^{-1}$$

$$\ln\left(\frac{100}{80}\right) = 0.0187 \times t$$

$$t = 11.9 \text{ min.}$$

3. A first-order reaction decreases from 0.0500 M to 0.0400 M in 24.4 minutes. What is its half life?

$$\ln\left(\frac{0.05}{0.04}\right) = k \cdot 24.4 \text{ min}$$

$$k = 0.009145 \text{ min}^{-1}$$

$$t_{1/2} = \frac{0.693}{k} = 75.8 \text{ min}$$

4. Consider the reaction:



The rate constant increases with temperature as summarized here. Use the Arrhenius equation to calculate the energy of activation for this reaction.

| Temperature | k (L/mol s) |
|---------------------|-------------|
| 330 °C K | 0.77 ← |
| 354 °C K | 1.8 |
| 383 °C K | 4.7 ← |

$$\ln(k_2/k_1) = (-E_a/R)(1/T_1 - 1/T_2)$$

$$\ln\left(\frac{4.7}{0.77}\right) = \frac{E_a}{R} \left(\frac{1}{383} - \frac{1}{330}\right)$$

$$E_a = 3.6 \times 10^4 \text{ J/mol}$$

$$= 36 \text{ kJ/mol}$$

Questions in final exam format:

5. What is the minimum energy barrier that must be overcome for a chemical reaction to occur?

- (A) activation energy B. net energy
C. potential energy D. rate limiting energy

6. If a plot of [A] vs time gives a linear plot...

- A. the reaction is first order and the slope is the rate.
B. the reaction is zero order and the slope is the rate constant.
C. the reaction is first order and the slope is the half-life.
(D) the reaction is zero order and the slope is the negative rate constant.

Now try these problems from the book:

Section 13.7. (Temperature) Problems 19, 20, 82, 90 and 118

Section 13.8. (Arrhenius equation) 21, 22, 84, 86 and 88