

Practice: pages 20–25

- 1 Answer is B. From experiments 1 and 3 the order with respect to Cl_2 is 1 (the rate quadruples as the concentration doubles. Therefore, in an experiment using 0.80 mol dm^{-3} of CO and 0.20 mol dm^{-3} of Cl_2 , the rate will be halved from the $6.4 \times 10^{-24} \text{ mol dm}^{-3} \text{ min}^{-1}$ obtained in experiment 1. (There is no need to work out the order with respect to CO.) [1]
- 2 Answer is B. The concentration halves every 120 s. So, the sequence of half-lives is $2.4 \rightarrow 1.2 \rightarrow 0.6 \rightarrow 0.3 \rightarrow 0.15$; that is four half-lives which is $4 \times 120 \text{ s}$. [1]
- 3 Answer is C. For a zero order reaction, the rate of decomposition is a constant. [1]
- 4 Answer is A. the decomposition is $\text{Rate} = k[\text{X}]^2$ [1]
From the data given, $0.016 = k[1]^2$ [1]
Therefore, $k = 0.016 \text{ s}^{-1}$. If $[\text{X}] = 0.25 \text{ mol dm}^{-3}$, $\text{rate} = 0.016[0.25]^2$
 $= 0.001 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ [1]
- 5 Answer is B. The overall order is 2, so the units are $\frac{1}{\text{concentration} \times \text{s}}$. [1]
- 6 Answer is D. I_2 has an order of zero and is not involved in the rate-determining step. Therefore, it can be ignored. In D, $[0.1][2]$ has a value that is greater than $[0.1][1]$. [1]
- 7 Answer is C. The slow step + the fast step must give the overall equation. I must, therefore, be removed in the fast step in a reaction with G, and some J and K must be formed. [1]
- 8 Answer is B. The rate equation is $\text{Rate} = k[\text{halide}]$ so statements 1 and 2 are correct. Raising the temperature will increase the rate of reaction by allowing more molecules to exceed the activation energy, but the activation energy is fixed and will not change. [1]
- 9 Answer is D. Increasing the amount of ozone present does not change the order with respect to ozone. A large excess is not necessary to ensure enough ozone is present. The purpose of a large excess is to make sure that during the reaction its concentration is virtually constant. Therefore, only statement 3 is correct. [1]
- 10 Answer is A. Experiments 2 and 3 give the order with respect to X as zero. Experiments 1 and 2 give the order with respect to Y as 2. So statements 1 and 2 are correct. The rate constant is calculated from a set of results, for example, experiment 1 gives $0.08 = k[0.10]^2$.
So $k = 8.0 \text{ dm}^3 \text{ mmol}^{-1} \text{ s}^{-1}$ and statement 3 is, therefore, correct. [1]
- 11 a) $2\text{N}_2\text{O}_5(\text{g}) \rightarrow 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ [1]
b) i) For every 2 mol of N_2O_5 that reacts 4 mol of NO_2 is formed. Therefore, the rate of formation of NO_2 is twice as fast as the loss of N_2O_5 , i.e. $7.0 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$. [1]
ii) For oxygen, the rate of formation will be $\frac{1}{2}$ the rate of loss of N_2O_5 ,
i.e. $1.75 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$ [1]

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Table of Contents Reaction Energy And Rate Answers

1. Understanding the eBook Reaction Energy And Rate Answers
 - The Rise of Digital Reading Reaction Energy And Rate Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Reaction Energy And Rate Answers
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in a Reaction Energy And Rate Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Reaction Energy And Rate Answers
 - Personalized Recommendations
 - Reaction Energy And Rate Answers User Reviews and Ratings

- Reaction Energy And Rate Answers and Bestseller Lists
- 5. Accessing Reaction Energy And Rate Answers Free and Paid eBooks
 - Reaction Energy And Rate Answers Public Domain eBooks
 - Reaction Energy And Rate Answers eBook Subscription Services
 - Reaction Energy And Rate Answers Budget-Friendly Options
- 6. Navigating Reaction Energy And Rate Answers eBook Formats
 - ePub, PDF, MOBI, and More
 - Reaction Energy And Rate Answers Compatibility with Devices
 - Reaction Energy And Rate Answers Enhanced eBook Features
- 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Reaction Energy And Rate Answers
 - Highlighting and Note-Taking Reaction Energy And Rate Answers
 - Interactive Elements Reaction Energy And Rate Answers
- 8. Staying Engaged with Reaction Energy And Rate Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Reaction Energy And Rate Answers
- 9. Balancing eBooks and Physical Books Reaction Energy And Rate Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Reaction Energy And Rate Answers
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Reaction Energy And Rate Answers
 - Setting Reading Goals Reaction Energy And Rate Answers
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Reaction Energy And Rate Answers
 - Fact-Checking eBook Content of Reaction Energy And Rate Answers
 - Distinguishing Credible Sources

13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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