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Practice: pages 20-25

- 1 Answer is B. From experiments 1 and 3 the order with respect to Cl₂ is 1 (the rate quadruples as the concentration doubles. Therefore, in an experiment using 0.80 mol dm⁻¹ of CO and 0.20 mol dm⁻² of Cl₂, the rate will be halved from the 6.4 × 10⁻⁰⁴ mol dm⁻³ min⁻¹ obtained in experiment 1. (There is no need to work out the order with respect to CO.)
- 2 Answer is B. The concentration halves every 120 s. So, the sequence of half-lives is 2.4 → 1.2 → 0.6 → 0.3 → 0.15; that is four half-lives which is 4 × 120 s. [1]
- Answer is C. For a zero order reaction, the rate of decomposition is a constant. [1]
- 4 Answer is A. the decomposition is Rate = k[X]² [1]
 - From the data given, $0.016 = k[1]^2$ [1]
 - Therefore, $k = 0.016 \text{ s}^{-1}$. If $\{X\} = 0.25 \text{ mol dim}^{-1}$, rate = 0.016 $[0.25]^2$ = 0.001 dm³ mol⁻² s⁻²
- 6 Answer is D. I₂ 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).
- 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 Rate = k[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.
- 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) $2N_2O_3(g) \rightarrow 4NO_2(g) + O_2(g)$ [1]
 - b) i) For every 2 mol of N₂O₃ that reacts 4 mol of NO₂ is formed. Therefore, the rate of formation of NO₃ is twice as fast as the loss of N₂O₃, i.e. 7.0 × 10⁻⁴ mol dm⁻² s⁻¹. [1]
 - For oxygen, the rate of formation will be $\frac{1}{2}$ the rate of loss of N₂O₃. i.e. 1.75×10^{-6} mol dm⁻⁸ s⁻⁶

Reaction Rate And Energy Reinforcement Answers Key

Sunil Kumar, Kavita Rathore, Debjyoti Banerjee

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