

## Practice: pages 20–25

- 1 Answer is B. From experiments 1 and 3 the order with respect to  $\text{Cl}_2$  is 1 (the rate quadruples as the concentration doubles. Therefore, in an experiment using  $0.80 \text{ mol dm}^{-3}$  of CO and  $0.20 \text{ mol dm}^{-3}$  of  $\text{Cl}_2$ , 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.  $\text{I}_2$  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  $\text{N}_2\text{O}_5$  that reacts 4 mol of  $\text{NO}_2$  is formed. Therefore, the rate of formation of  $\text{NO}_2$  is twice as fast as the loss of  $\text{N}_2\text{O}_5$ , 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  $\text{N}_2\text{O}_5$ ,  
i.e.  $1.75 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$  [1]



# Reaction Energy And Rate Answers

**DP Hallahan**





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