

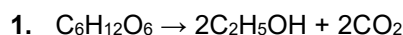
Oxford Resources for IB

Chemistry – 2023 Edition

Answers

Reactivity 2.1 – How much? The amount of chemical change

Practice question (page 462)



Practice questions (page 463)

2. a. 100 molecules of O_2
 b. 0.20 mol of C_4H_{10}
 c. 28.8 g of O_2

Practice questions (page 465)

3. a. $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$
 b. $\text{Cl}_2(\text{g})$, as both gases are consumed in equal volumes.

c.	$\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$		
$V_{\text{initial}} / \text{dm}^3$	4.54	2.27	0
$\Delta V / \text{dm}^3$	-2.27	-2.27	+4.54
$V_{\text{final}} / \text{dm}^3$	2.27	0	4.54

Note that the volumes of gases at STP are proportional to their amounts, so either quantity can be used directly in stoichiometric calculations.

Practice questions (page 467)

4. a. $2\text{KOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
 b. The volumes of initial solutions are equal, so we can choose these volumes arbitrarily, for example, 1.00 dm^3 each. In this case, $n_{\text{initial}}(\text{KOH}) = n_{\text{initial}}(\text{H}_2\text{SO}_4) = 0.100 \text{ mol}$. Therefore:

	$2\text{KOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$			
$n_{\text{initial}} / \text{mol}$	0.100	0.100	0	excess
$\Delta n / \text{mol}$	-0.100	-0.050	+0.050	+0.100
$n_{\text{final}} / \text{mol}$	0	0.050	0.050	excess

Assuming that the volumes of diluted solutions are additive, the volume of the final solution will be $1.00 + 1.00 = 2.00 \text{ dm}^3$. Therefore:

$$\begin{aligned}
 c_{\text{final}}(\text{H}_2\text{SO}_4) &= c_{\text{final}}(\text{K}_2\text{SO}_4) \\
 &= \frac{0.050 \text{ mol}}{2.00 \text{ dm}^3} \\
 &= 0.025 \text{ mol dm}^{-3}
 \end{aligned}$$

Practice question (page 468)

5. $0.455 \text{ mol dm}^{-3}$

Skills questions (page 469)

Responses to this task will vary depending on experimental designs. Guidance:

1. The guidance given in *Tool 1* should be followed. Plans will likely involve reacting an excess of standard acid solution, of known volume and concentration, with a known mass of crushed eggshells. The remaining acid can then be titrated against a standard NaOH solution to determine the amount, in moles, that reacted with the eggshells.
2. Will depend on methodology. The guidance in *Tool 1* should be followed.
3. A rough titration should be performed first and then the titration should be repeated until concordant values are obtained.
4. For example:
 - Determine the amount, in moles, of base that neutralised the acid remaining from the first titration.
 - Determine the amount, in moles, of acid that reacted with the base. Determine the initial amount, in moles, of acid used in the first titration. From these data, determine the amount, in moles, of acid that reacted with the eggshells.
 - Determine the mass of calcium carbonate in the eggshells.
 - Find the percentage by mass of calcium carbonate in the eggshells.
5. Answers will vary depending on the data.
6. Follow the guidance in *Tool 3*.
7. Answers will vary depending on the origin of the eggshells and the literature values found.
8. Answers will vary. For example:
 - Assumption that there are no other carbonates or bases in the eggshells that can also react with the acid.
 - Assumption that the eggshell was completely dry when its initial mass was measured.

Activity (page 470)

Initial mixture:

$$m(\text{CaC}_2) = 16.0 \text{ g}$$

$$n(\text{O}_2) = \frac{19.3 \text{ dm}^3}{22.7 \text{ dm}^3 \text{ mol}^{-1}} \\ \approx 0.850 \text{ mol}$$

$$m(\text{O}_2) = 0.850 \text{ mol} \times 32.00 \text{ g mol}^{-1} = 27.2 \text{ g}$$

$$m(\text{mixture}) = 27.2 \text{ g} + 16.0 \text{ g} = 43.2 \text{ g}$$

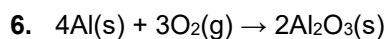
Final mixture:

$$m(\text{CaCO}_3) = 25.0 \text{ g}$$

$$m(\text{CO}_2) = 0.250 \text{ mol} \times 44.01 \text{ g mol}^{-1} \approx 11.0 \text{ g}$$

$$m(\text{O}_2) = 0.225 \text{ mol} \times 32.00 \text{ g mol}^{-1} = 7.2 \text{ g}$$

$$m(\text{mixture}) = 25.0 \text{ g} + 11.0 \text{ g} + 7.2 \text{ g} = 43.2 \text{ g}$$

Practice question (page 471)

$$n(\text{Al}) = \frac{9.443 \text{ g}}{26.98 \text{ g mol}^{-1}}$$

$$= 0.3500 \text{ mol}$$

$$n(\text{O}_2) = \frac{7.945 \text{ dm}^3}{22.7 \text{ dm}^3 \text{ mol}^{-1}}$$

$$= 0.350 \text{ mol}$$

The limiting reactant is Al(s), so

$$n_{\text{theoretical}}(\text{Al}_2\text{O}_3) = \frac{0.3500}{2} = 0.1750 \text{ mol}$$

$$n_{\text{practical}}(\text{Al}_2\text{O}_3) = \frac{17.13 \text{ g}}{101.96 \text{ g mol}^{-1}} \approx 0.1680 \text{ mol}$$

$$\text{percentage yield} = \frac{0.1680 \text{ mol}}{0.1750 \text{ mol}} \times 100\%$$

$$= 96.00\%$$

Data-based questions (page 472)

Answers to questions 1 and 2 are given in the table below.

	Experiment 1	Experiment 2	Experiment 3
amount of salicylic acid / mol	0.0149	0.0186	0.0170
amount of ethanoic anhydride / mol	0.0244	0.0279	0.0305
limiting reactant	salicylic acid	salicylic acid	salicylic acid
amount of aspirin / mol	0.0190	0.0128	0.0104
yield of aspirin / %	128	69	61

- The percentage yield of aspirin in experiment 1 is greater than 100%. This sample of aspirin is likely to be impure, as its melting point (122–128 °C) is much lower than the literature value (136 °C), and it has a very broad range (6 °C instead of typical 1–2 °C).
- Most likely, the purest sample of aspirin was produced in experiment 3, as the melting point of this sample (135–136 °C) is the closest to the literature value (136 °C), and it has the narrowest range (2 °C).

Skills questions (page 473)

1. $n(\text{CaCl}_2) = (0.025 \text{ cm}^3) \times (0.50 \text{ mol dm}^{-3}) = 0.0125 \text{ mol}$

$$n(\text{Na}_2\text{CO}_3) = (0.015 \text{ cm}^3) \times (1.0 \text{ mol dm}^{-3}) = 0.015 \text{ mol}$$

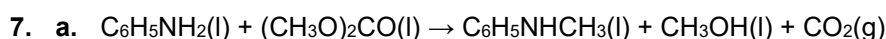
Calcium chloride and sodium carbonate react in a 1 : 1 ratio. Thus, calcium chloride is the limiting reactant and sodium carbonate is in excess.

2. $n(\text{CaCO}_3) = (0.0125 \text{ mol}) \times (100.09 \text{ g mol}^{-1}) = 1.25113 \text{ g} \approx 1.3 \text{ g}$ (3 sf)

3. Answers will vary depending on the experimental yield.

4. Answers will vary depending on class results.

5. Answers will vary but might include ideas such as drying the precipitate more efficiently by breaking up any clumps and weighing to a constant mass, better transfer of suspension from beaker to the filter paper, etc.

Practice questions (page 476)

b. The by-product methanol can be recycled and used for the synthesis of dimethyl carbonate.

End of topic questions (pages 478–479)

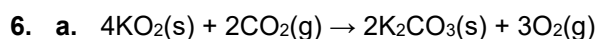
1. Answers will depend on the student, but a possible answer might be: Stoichiometric coefficients in chemical equations show the mole ratio of reactants and products in a reaction. In turn, this ratio can be used to determine the changes in amounts and masses of reactants and products. The mole ratio can also be used to determine the changes in volumes of gaseous species and the molar concentrations of reactants and products in solutions. Finally, the theoretical atom economy of a reaction can also be deduced from the reaction stoichiometry and molar masses of participating species.

2. B

3. C

4. D

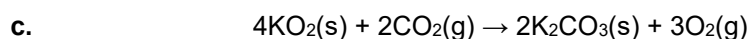
5. C



b. $n(\text{KO}_2) = \frac{28.44 \text{ g}}{71.10 \text{ g mol}^{-1}} \approx 0.400 \text{ mol}$

$$n(\text{CO}_2) = \frac{6.81 \text{ dm}^3}{22.7 \text{ dm}^3 \text{ mol}^{-1}} \approx 0.300 \text{ mol}$$

KO_2 and CO_2 react with each other in 2 : 1 ratio, so 0.300 mol CO_2 could potentially react with 0.600 mol KO_2 . However, there is only 0.400 mol KO_2 , so KO_2 is the limiting reactant.



$n_{\text{initial}} / \text{mol}$	0.400	0.300	0	0
$\Delta n / \text{mol}$	-0.400	-0.200	+0.200	+0.300
$n_{\text{final}} / \text{mol}$	0	0.100	0.200	0.300

$$m(\text{K}_2\text{CO}_3) = 0.200 \text{ mol} \times 138.21 \text{ g mol}^{-1} \approx 27.6 \text{ g}$$

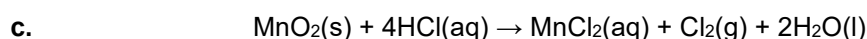
$$V(\text{O}_2) = 0.300 \text{ mol} \times 22.7 \text{ dm}^3 \text{ mol}^{-1} \approx 6.81 \text{ dm}^3$$

7. a. $n(\text{MnO}_2) = \frac{17.4 \text{ g}}{86.94 \text{ g mol}^{-1}} \approx 0.200 \text{ mol}$

$$n(\text{HCl}) = 0.100 \text{ dm}^3 \times 10.0 \text{ mol dm}^{-3} = 1.00 \text{ mol}$$

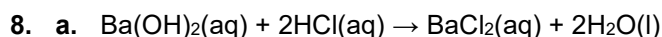
MnO_2 and HCl react with each other in 1 : 4 ratio, so 1.00 mol HCl could potentially react with 0.250 mol MnO_2 . However, there is only 0.200 mol MnO_2 , so MnO_2 is the limiting reactant.

b. $n(\text{Cl}_2) = n(\text{MnO}_2) = 0.200 \text{ mol}$, so $V(\text{Cl}_2) = 0.200 \text{ mol} \times 22.7 \text{ dm}^3 \text{ mol}^{-1} \approx 4.54 \text{ dm}^3$



$n_{\text{initial}} / \text{mol}$	0.200	1.00	0	0	excess
$\Delta n / \text{mol}$	-0.200	-0.800	+0.200	+0.200	+0.400
$n_{\text{final}} / \text{mol}$	0	0.200	0.200	0.200	excess

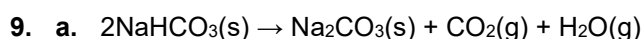
$$c(\text{HCl}) = c(\text{MnCl}_2) = \frac{0.200 \text{ mol}}{0.100 \text{ dm}^3} = 2.00 \text{ mol dm}^{-3}$$



b. $n(\text{HCl}) = 0.100 \text{ mol dm}^{-3} \times 0.0117 \text{ dm}^3 = 0.00117 \text{ mol}$

$$n(\text{BaCl}_2) = 0.5 \times n(\text{HCl}) = 0.000585 \text{ mol}$$

$$c(\text{BaCl}_2) = \frac{0.000585 \text{ mol}}{0.0100 \text{ dm}^3} = 0.0585 \text{ mol dm}^{-3}$$



Note that at high temperature water is likely to be produced as steam rather than liquid water.

b. $n(\text{NaHCO}_3) = \frac{10.0 \text{ g}}{84.01 \text{ g mol}^{-1}} \approx 0.119 \text{ mol}$

$$n(\text{CO}_2)_{\text{theor}} = 0.5 \times n(\text{NaHCO}_3) = 0.0595 \text{ mol}$$

$$m(\text{CO}_2)_{\text{theor}} = 0.0595 \text{ mol} \times 44.01 \text{ g mol}^{-1} \approx 2.62 \text{ g}$$

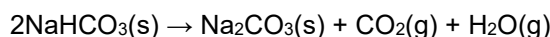
c. $V(\text{CO}_2)_{\text{theor}} = 0.0595 \text{ mol} \times 22.7 \text{ dm}^3 \text{ mol}^{-1} \approx 1.35 \text{ dm}^3$

$$\text{percentage yield} = \frac{1.20 \text{ dm}^3}{1.35 \text{ dm}^3} \times 100\% \approx 88.9\%$$

d. $n(\text{CO}_2)_{\text{pract}} = \frac{1.20 \text{ dm}^3}{22.7 \text{ dm}^3 \text{ mol}^{-1}} \approx 0.0529 \text{ mol}$

According to the equation, $0.0529 \times 2 \approx 0.106 \text{ mol}$ NaHCO_3 has decomposed.

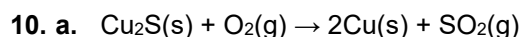
This means that $0.119 - 0.106 = 0.013 \text{ mol}$ NaHCO_3 remained unchanged. The amount of Na_2CO_3 formed is equal to the amount of CO_2 . Therefore:



$n_{\text{initial}} / \text{mol}$	0.119	0	0	0
$\Delta n / \text{mol}$	-0.106	+0.0529	+0.0529	+0.0529
$n_{\text{final}} / \text{mol}$	0.013	0.0529	0.0529	0.0529

$$m(\text{NaHCO}_3)_{\text{final}} = 0.013 \text{ mol} \times 84.01 \text{ g mol}^{-1} \approx 1.09 \text{ g}$$

$$m(\text{Na}_2\text{CO}_3)_{\text{final}} = 0.0529 \text{ mol} \times 105.99 \text{ g mol}^{-1} \approx 5.61 \text{ g}$$



b. $m(\text{Cu}_2\text{S}) = 753 \text{ kg} = 7.53 \times 10^5 \text{ g}$

$$n(\text{Cu}_2\text{S}) = \frac{7.53 \times 10^5 \text{ g}}{159.17 \text{ g mol}^{-1}} \approx 4730 \text{ mol}$$

$$n(\text{Cu})_{\text{theor}} = 2 \times 4730 \text{ mol} = 9460 \text{ mol}$$

$$m(\text{Cu})_{\text{theor}} = 9460 \text{ mol} \times 63.55 \text{ g mol}^{-1} \approx 6.01 \times 10^5 \text{ g} = 601 \text{ kg}$$

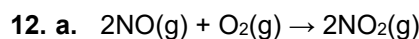
$$\text{percentage yield} = \frac{506 \text{ kg}}{601 \text{ kg}} \times 100\% \approx 84.2\%$$

11. a. $\text{atom economy} = \frac{M_r(\text{C}_2\text{H}_4)}{M_r(\text{CH}_3\text{CH}_2\text{OH})} \times 100\% = \frac{28.06}{46.08} \times 100\% = 60.89\%$

b. $\text{atom economy} = \frac{M_r(\text{C}_6\text{H}_{12}\text{O}_6)}{6 \times M_r(\text{CO}_2) + 6 \times M_r(\text{H}_2\text{O})} \times 100\%$

$$= \frac{180.18}{6 \times 44.01 + 6 \times 18.02} \times 100\% = 48.41\%$$

c. All atoms from the reactant are used to form the product, so the atom economy is 100%.



b. All atoms from the reactants are used to form the product, so the atom economy is 100%.

c. Oxygen is in excess, so nitrogen monoxide is the limiting reactant. One mole of nitrogen monoxide can theoretically produce one mole of nitrogen dioxide, so the theoretical yield of nitrogen dioxide is 1.00 dm^3 .

$$\text{percentage yield} = \frac{0.700 \text{ dm}^3}{1.00 \text{ dm}^3} \times 100\% = 70.0\%$$

d. The theoretical atom economy is 100%, so the actual atom economy is equal to the percentage yield (70.0%).

Reactivity 2.2 – How fast? The rate of chemical change

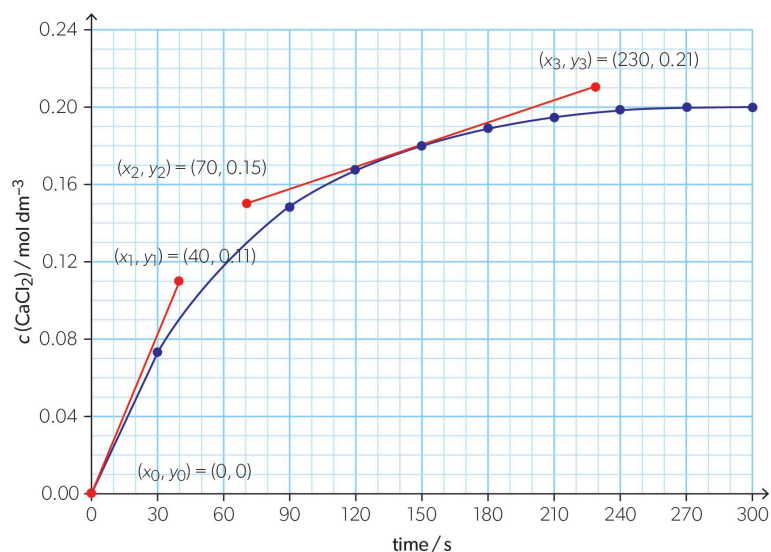
Practice questions (page 482)

1. a. $v(\text{NO}) = 1.2 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$
 $v(\text{O}_2) = 1.5 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$
 b. $v_{\text{overall}} = 3.0 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$

Data-based questions (page 484)

- a. $2.44 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$
 b. $7.41 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$

Activity (page 486)

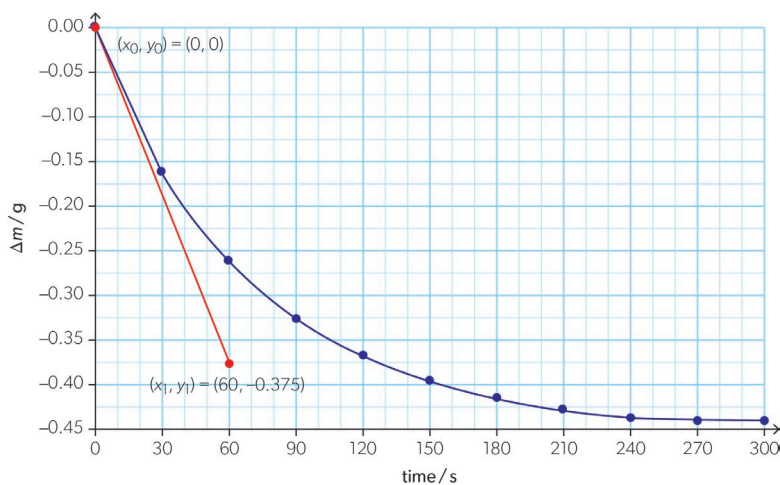


$$v_{\text{init}} = 2.8 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$$

$$v_{\text{inst}}(150) = 3.8 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$$

Practice questions (page 487)

2. a.

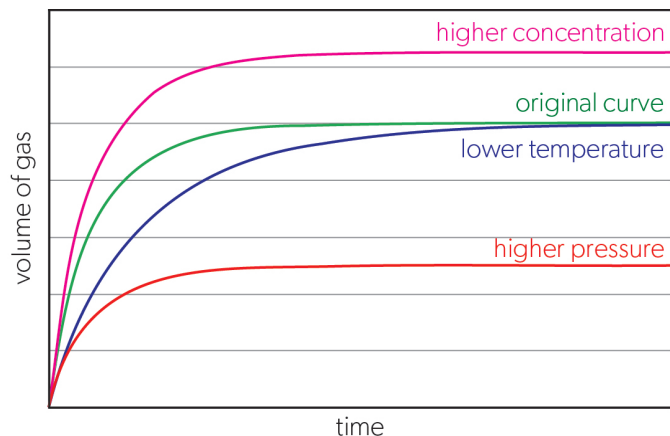


b. $v_{\text{init}} = 6.25 \times 10^{-3} \text{ g s}^{-1}$

c. $v_{\text{init}}(\text{overall}) = v_{\text{init}}(\text{CO}_2) = \frac{6.25 \times 10^{-3} \text{ g s}^{-1}}{44.01 \text{ g mol}^{-1} \times 0.050 \text{ dm}^3}$
 $\approx 2.8 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$

Notice that this answer matches the value of v_{init} obtained in the activity on page 486.

Activity (page 490)



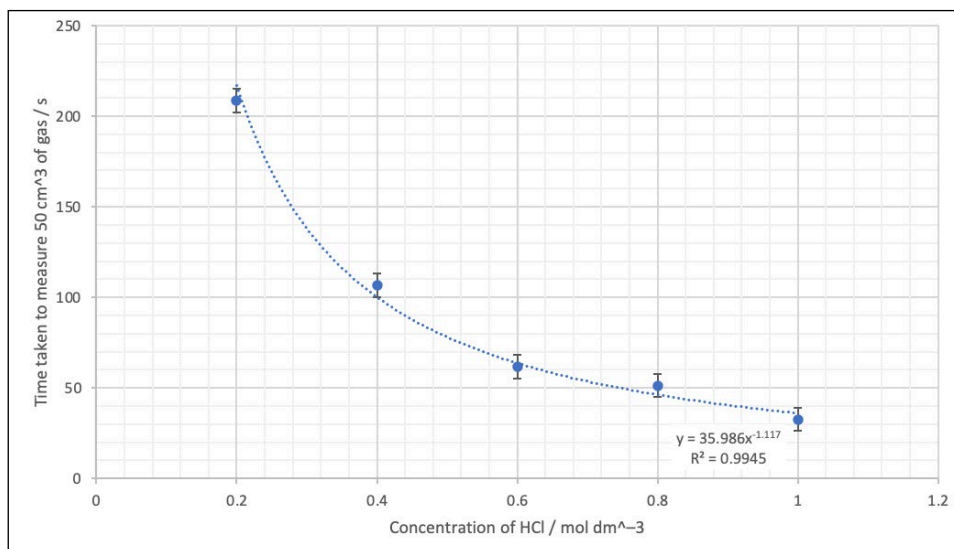
Skills question (page 491)

1. Independent: concentration of HCl; dependent: time taken to collect 50 cm³ of gas.
2. Answers might include: temperature of the water bath, temperature of the reaction mixture, volume of acid, etc.
3. The second trial performed with 0.2 mol dm⁻³ is an outlier as it is much larger than the other two trials.
- 4.

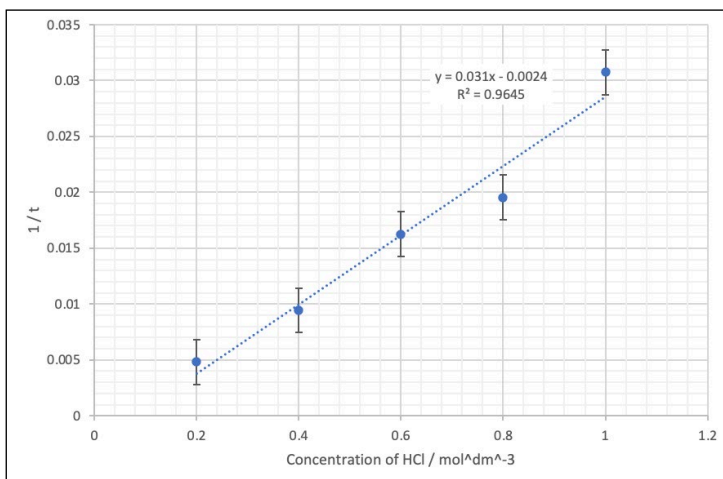
	A	B	C	D	G	H	I	J
	Time taken to measure 50 cm ³ of CO ₂ / s			1/t				
Concentration of HCl / mol dm ⁻³	1	2	3	1	2	3	mean	
0.2	202		215	0.0049505		0.00465116	0.00480083	
0.4	118	104	98	0.00847458	0.00961538	0.01020408	0.00943135	
0.6	62	58	65	0.01612903	0.01724138	0.01538462	0.01625168	
0.8	54	48	52	0.01851852	0.02083333	0.01923077	0.01952754	
1	30	35	33	0.03333333	0.02857143	0.03030303	0.03073593	

5 and 6.

- a. Note error bars have been estimated using the half-range method (see *Tool 3*).



- b. Note error bars have been estimated using the half-range method (see *Tool 3*).



7. Answers will vary, depending on what you find. Guidance: plot the data with and without the outlier and see what the effect is on the graphs. Justify the removal (or not) of the outlier based on what might have caused it and whether you think it is due to external factors.
8. The points do not go through all the error bars in the second graph indicating a degree of random error. One would expect the line in the second graph to go through the origin because the rate would be 0 when the concentration of acid is 0. However, the equation of the line shows a non-zero y -intercept, which suggests the presence of a systematic error.
9. Answers may include:
- Carbon dioxide is soluble in water and therefore the volume measured in the measuring cylinder does not account for all the CO_2 formed in the reaction.
 - Uneven concentration of acid throughout the mixture.
 - Difficulty reading the volume from the side of the measuring cylinder.
 - Loss of gas, particularly if the bung was removed and then replaced in order to add the second reactant. This would have also impacted the volume measurements, particularly those taken at higher concentrations.
 - The reaction's own exo- or endothermicity would have affected the rate.

Etc.

10. Answers may include:
- Use of a gas syringe or a mass loss method to determine the rate.
 - Maintaining a constant water temperature to approximate the amount of dissolved carbon dioxide. This value could then be used to correct the data collected.
 - Constant stirring of the reaction mixture with a magnetic stir bar and stirrer hotplate.
 - Adding the acid through a thistle funnel.

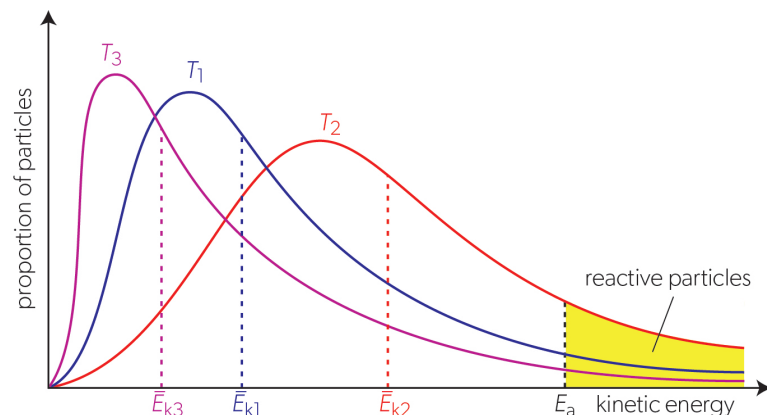
Etc.

11. Answers will vary greatly. For example, the reaction can be carried out at different temperatures, the order of reaction with respect to HCl could be investigated, data from different volume-measurement methods could be compared to establish the systematic error introduced by each, etc.
12. a. Answers will vary but could include measuring the volume of hydrogen released, the pH changes in the solution, etc.
- b. Answers will vary but could include measuring the time taken to obscure a mark made on a paper below the reaction flask, the changes in pH, turbidity of the reaction mixture, etc.
- c. Answers will vary but could involve measuring the volume or mass of carbon dioxide produced, the colour change of the solution (through spectrophotometry), etc.

13. Answers will vary widely.

Practice questions (page 493)

3. a. Your sketch should be very similar to that shown in figure 13 (left). The height of the E_a barrier should be approximately twice as large as the drop in enthalpy between the reactants and products.
- b. $E_a' = E_a - \Delta H_r = 30 \text{ kJ mol}^{-1}$

Activity (page 493)

The maximum for the curve at T_3 will be higher and lie to the left of that for the curve at T_1 . The average kinetic energy of the particles at T_3 will be lower than that at T_1 . At the right-hand side, the curve at T_3 will flatten below the curve at T_1 . Therefore, fewer particles at T_3 will have $E_k > E_a$, so the proportion of reactive particles at T_3 will be lower than that at T_1 .

Practice questions (page 495)

4. a. $\text{NO}_2(\text{g})$ is a catalyst while $\text{NO}(\text{g})$ is an intermediate. $\text{NO}(\text{g})$ is formed in the first step of the reaction, but it is almost immediately oxidized back to $\text{NO}_2(\text{g})$ in the second step. This two-step process has a lower E_a than the uncatalysed reaction, so the reaction rate increases in the presence of $\text{NO}_2(\text{g})$.
- b. Your sketch should be very similar to that shown in figure 15. The label “reactants” can be replaced with “ $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$ ”, the label “intermediate” with “ $\text{NO}(\text{g})$ ”, and the label “products” with “ $2\text{SO}_3(\text{g})$ ”.

Activity (page 496)

Factor	Frequency of collisions	Average E_k of reacting species	Activation energy	Rate of forward reaction
decrease in a reactant concentration	decreases	no change	no change	decreases
increase in pressure of a gaseous reactant	increases	no change	no change	increases
decrease in pressure of a gaseous product	decreases*	no change	no change	no change
increase in volume of a gaseous reaction mixture	decreases	no change	no change	decreases
decrease in temperature	decreases	decreases	no change	decreases
increase in surface area of a solid reactant	increases	no change	no change	increases
addition of a catalyst	no change	no change	decreases	increases

* The overall frequency of collisions decreases, but the frequency of collisions between reactant particles does not change.

Practice questions (page 503)

5. a. First order with respect to each of A(g) and B(g)
b. Rate equation: $v = k[A][B]$
Overall reaction order: second
c. $k \approx 0.67 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$

Practice questions (page 506)

6. a. Second order
b. *Advantages:* more intuitive method; experimental data are used directly; the knowledge (or derivation) of linear equations for the 1st and 2nd orders is not required; final result can be obtained in fewer steps.
Disadvantages: limited precision, as the positions and slopes of tangent lines depend on personal judgement; the final answer is less obvious and often requires rounding.

Practice question (page 507)

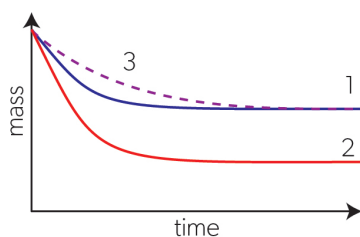
$$7. \ln \frac{1}{8.2} = \frac{E_a}{8.31 \text{ J K}^{-1} \text{ mol}^{-1}} \left(\frac{1}{318 \text{ K}} - \frac{1}{298 \text{ K}} \right)$$
$$E_a \approx 83000 \text{ J mol}^{-1} = 83 \text{ kJ mol}^{-1}$$

Skills question (page 509)

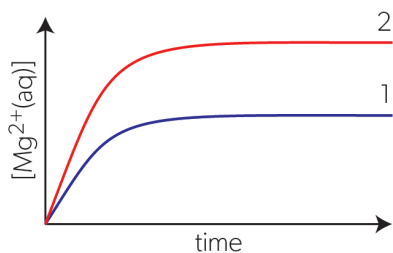
Answers will vary widely depending on the reaction and variables selected. The guidance in the *Tools* and *Inquiry cycle* chapters should be followed.

End of topic questions (page 510–511)

1. Answers will depend on the student, but a possible answer might be: The rate of a reaction depends on pressure (for gaseous species), concentration (for species in solution and gas phase), surface area (for heterogeneous reactions), temperature and the presence of a catalyst. The first four factors affect the frequency of collisions between reactant particles. In addition, temperature affects the average energy of these collisions and thus the proportion of successful collisions that lead to chemical changes. Catalysts increase the rate of reaction by providing an alternative reaction pathway with a lower activation energy. In a multi-step reaction, the rate is largely controlled by the rate-determining step, which is the step that has the highest activation energy. The study of a reaction mechanism allows us to identify this step and thus gain a better understanding of how the reaction rate can be controlled.
2. D
3. B
4. a. As both reactants are consumed, the concentration of HCl(aq) in the solution decreases. This reduces the frequency of collisions between magnesium metal and $\text{H}^+(\text{aq})$ ions. The frequency of successful collisions decreases proportionally to the overall frequency of collisions. As a result, the mass of $\text{H}_2(\text{g})$ released per unit time decreases, so the curve becomes more and more gradual. Once the HCl(aq) is consumed completely, the release of $\text{H}_2(\text{g})$ stops, so the curve flattens out.
b. The reaction rate will be lower, but the final mass of $\text{H}_2(\text{g})$ released will be the same as that in the original experiment. Therefore, curve 3 will be more gradual than curve 1, but it will begin and end at the same levels as curve 1.



c.



In both experiments, HCl(aq) is the limiting reactant, so the final $[\text{Mg}^{2+}(\text{aq})]$ in the second experiment will be higher than that in the first experiment. Therefore, curve 2 will flatten at a higher level than curve 1.

5. a. $pV = nRT$

$$c = \frac{n}{V}$$

$$c = \frac{p}{RT}$$

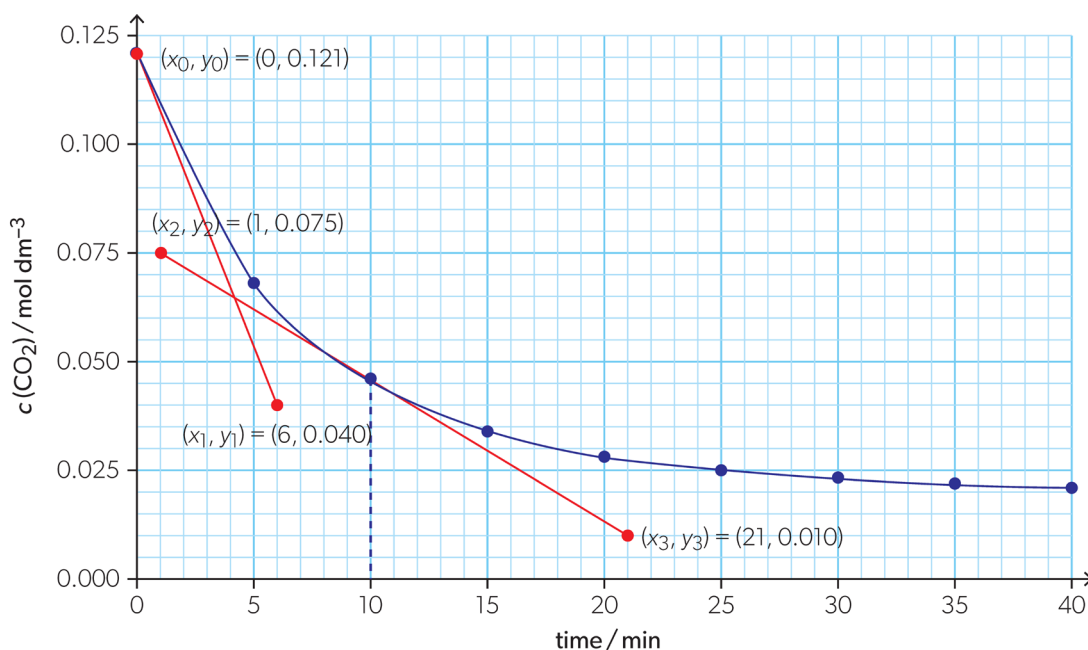
Using the last expression and the data from figure 28, we can first determine the pressures and then concentrations of $\text{CO}_2(\text{g})$:

t / min	0	5	10	15	20	25	30	35	40
p / kPa	300	169	113	85	70	62	57	54	52
$c(\text{CO}_2) / \text{mol dm}^{-3}$	0.121	0.068	0.046	0.034	0.028	0.025	0.023	0.022	0.021

b. $\Delta c(\text{CO}_2) = 0.028 - 0.121 = -0.093 \text{ mol dm}^{-3}$

$$v_{\text{avr}} = \frac{|-0.093 \text{ mol dm}^{-3}|}{20 \text{ min}} \approx 0.0047 \text{ mol dm}^{-3} \text{ min}^{-1}$$

c.

d. $v_{\text{init}} = -\frac{0.040 - 0.121}{6 - 0} \approx 0.014 \text{ mol dm}^{-3} \text{ min}^{-1}$

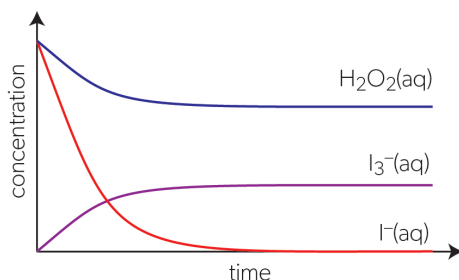
$$v_{10} = -\frac{0.010 - 0.075}{21 - 1} \approx 0.0033 \text{ mol dm}^{-3} \text{ min}^{-1}$$

6. a. i. $v(I^-) = \frac{(0.050 - 0.020) \text{ mol dm}^{-3}}{2.0 \times 60 \text{ s}} = 2.50 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$

$$v(\text{H}_2\text{O}_2) = \frac{1}{3} \times 2.50 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1} \approx 8.33 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$$

ii. The stoichiometric coefficient before $\text{H}_2\text{O}_2(\text{aq})$ is 1, so the overall average reaction rate will be the same as the average reaction rate with respect to $\text{H}_2\text{O}_2(\text{aq})$.

b. $\text{H}_2\text{O}_2(\text{aq})$ is in excess, so $\text{I}^-(\text{aq})$ will eventually be consumed completely. The stoichiometric coefficient before $\text{I}^-(\text{aq})$ is three times larger than those before $\text{H}_2\text{O}_2(\text{aq})$ and $\text{I}_3^-(\text{aq})$, so at any given time the curve for $\text{I}^-(\text{aq})$ will be three times steeper than the other two curves.



c. By colorimetry or spectrophotometry, as aqueous solutions of iodine are coloured.

d. i. The rate is likely to increase, as a higher concentration of hydrogen peroxide will lead to more frequent collisions between reacting species.

ii. The rate will not be affected noticeably, as liquids are almost incompressible.

iii. The rate will decrease, as a lower temperature will decrease both the frequency of collisions between reacting species and the average kinetic energy of these species (so fewer species will have $E_{\text{kin}} > E_a$).

iv. The rate will not be affected, as the composition of the resulting solution will not change.

7. a. See figure 13 (left) on page 492.

b. See figure 15 on page 494.

c. See figure 16 on page 494, but without the vertical lines or text.

d. As shown in figure 16, a catalyst lowers the activation energy, so more particles have enough energy to react with one another. Although the overall frequency of collisions remains unchanged, the frequency of successful collisions increases and so does the reaction rate.

8. a. Doubling $[\text{Cl}_2]$ doubles the rate, so the order with respect to chlorine is 1. Doubling $[\text{NO}]$ increases the rate by a factor of four, so the order with respect to nitrogen monoxide is 2.

b. $v = k[\text{Cl}_2][\text{NO}]^2$

overall order = $1 + 2 = 3$

c. Using the first row of the table gives the following:

$$5.30 \times 10^{-3} \text{ mol dm}^{-3} \text{ min}^{-1} = k \times (1.50 \times 10^{-2} \text{ mol dm}^{-3}) \times (1.50 \times 10^{-2} \text{ mol dm}^{-3})^2$$

Solving for k gives $k \approx 1.57 \times 10^3 \text{ dm}^6 \text{ mol}^{-2} \text{ min}^{-1}$

d. If step 1 were slow, the rate would be independent of $[\text{Cl}_2]$. The rate equation shows first order with respect to chlorine, so step 2 is the RDS.

9. a. $2\text{N}_2\text{O}(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + \text{O}_2(\text{g})$

b. $\text{N}_2\text{O}(\text{g})$: reactant; $\text{O}(\text{g})$: intermediate; $\text{N}_2(\text{g})$ and $\text{O}_2(\text{g})$: products.

c. 1 (monomolecular)

d. $v = k[\text{N}_2\text{O}]$

overall order = 1

e. See figure 13 (left) on page 492.

10. a. Second order, as shown by the units of the rate constant.

b. Use the following expression:

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\ln \frac{0.183}{5.45} = \frac{E_a}{8.31 \text{ J K}^{-1} \text{ mol}^{-1}} \left(\frac{1}{333 \text{ K}} - \frac{1}{303 \text{ K}} \right)$$

Solving for E_a gives $E_a \approx 94856 \text{ J mol}^{-1} = 94.9 \text{ kJ mol}^{-1}$

c. Using the same expression and substituting the values of k_1 , T_1 , T_3 and E_a gives the following:

$$\ln \frac{0.183}{k_3} = \frac{94856 \text{ J mol}^{-1}}{8.31 \text{ J K}^{-1} \text{ mol}^{-1}} \left(\frac{1}{318 \text{ K}} - \frac{1}{303 \text{ K}} \right)$$

Solving for k_3 gives $k_3 \approx 1.08 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$

11. a. In the linear form of the Arrhenius equation, the line slope gives $-\frac{E_a}{R}$, in kelvin. We can find the slope from any two points on the plot, for example, (0.00285, 15.68) and (0.00385, 14.10):

$$\text{slope} = \frac{15.68 - 14.10}{0.00285 - 0.00385}$$

$$\approx -1580 \text{ K}$$

$$-\frac{E_a}{8.31 \text{ J K}^{-1} \text{ mol}^{-1}} = -1580 \text{ K}$$

$E_a \approx 13130 \text{ J mol}^{-1} = 13.1 \text{ kJ mol}^{-1}$

b. From the first point on the plot, $\ln k = 15.68$ and $1/T = 0.00285 \text{ K}^{-1}$. Substituting these values, together with the E_a value, in J, into the logarithmic form of the Arrhenius equation gives the following:

$$15.68 = -1580 \text{ K} \times 0.00285 \text{ K}^{-1} + \ln A$$

Solving for A gives $A \approx 5.8 \times 10^8$

The units for A are the same as those for k . Since the reaction is second order, the units for k are $\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$, so $A \approx 5.8 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$.

Reactivity 2.3 – How far? The extent of chemical change

Practice questions (page 517)

1. a. $K = \frac{[\text{N}_2][\text{H}_2]^3}{[\text{NH}_3]^2}$

b. $K = \frac{0.50 \times 1.50^3}{1.00^2} \approx 1.69$

Practice questions (page 520)

2. a. Equilibrium position shifts to the right; K is not affected.

b. Equilibrium position shifts to the right; K is not affected.

Practice questions (page 521)

3. a. Equilibrium position shifts to the left; K is not affected.

b. Neither equilibrium position nor K is affected.

c. Equilibrium position shifts to the right; K is not affected.

d. Equilibrium position shifts to the right; K is not affected.

Practice question (page 526)

$$4. \quad Q = \frac{0.1}{0.025^2} = 160$$

$Q > K$, so the reverse reaction is favoured.

Practice question (page 527)

$$5. \quad [\text{NO}_2]_{\text{eq}} \approx 0.060 \text{ mol dm}^{-3}$$

$$[\text{NO}_2]_{\text{initial}} \approx 0.142 \text{ mol dm}^{-3}$$

Skills questions (pages 528–529)

7. The formula is used to calculate the reaction quotient, Q :

$$Q = \frac{n(\text{NH}_3)^2}{n(\text{N}_2) \times n(\text{H}_2)^3} \times V^2$$

Since Q is the quotient of the concentrations, each amount in moles must be divided by the volume of the container, V . If you divide each amount in moles by V , you will see that two of the values of V cancel out leaving two more, which is the reason for the V^2 .

8 and 9. Answers will vary, depending on the nature of the exploration carried out in 5 and 6, and your prior experience working with spreadsheets.

10. Answers will vary, but you should consider the new value of K , the stoichiometry of the reaction, volume of the container, and initial concentrations of reactants and products.

Practice question (page 530)

$$6. \quad [\text{OH}^-] \approx 8.61 \times 10^{-6} \text{ mol dm}^{-3}$$

Activity (page 531)

$$-3900 \text{ J mol}^{-1} = -8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K} \times \ln K$$

$$K \approx 4.83$$

End of topic questions (pages 534–535)

1. The equilibrium law can be used to quantify the composition of a reaction mixture at equilibrium. This composition can be influenced by controlling the temperature, pressure and concentrations of reacting species. The effects of specific changes in these conditions on the extent of a reversible reaction can be predicted using Le Châtelier's principle. In many industrial processes, reversible reactions are drawn almost to completion by a constant supply of reactants and removal of products from the reaction mixture. Although catalysts do not affect the position of equilibrium, they increase the rates of all reactions and thus accelerate the changes predicted by Le Châtelier's principle.

2. D

3. B

4. C

5. C

6. A

7. D

8. a. $K = \frac{[\text{O}_3]^2}{[\text{O}_2]^3}$

b. $K = \frac{[\text{NO}]^2[\text{O}_2]}{[\text{NO}_2]^2}$

c. $K = \frac{[\text{NO}][\text{O}_2]^{0.5}}{[\text{NO}_2]}$

d. $K = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$

e. $K = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$

Notice that $\text{H}_2\text{O}(\text{l})$ is the solvent, so it is not included in the K expression.

f. $K = [\text{CO}_2]$

Notice that $\text{CaCO}_3(\text{s})$ and $\text{CaO}(\text{s})$ are solids, so they are not included in the K expression.

9. According to table 1, $K' = \sqrt{K} = \sqrt{0.81} = 0.90$

10. a. $K = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$

$$K = \frac{0.091^2}{0.012 \times 0.015} \approx 46$$

b. When the temperature decreases, the K value increases, so the forward reaction is exothermic. Therefore, for the forward reaction $\Delta H^\ominus < 0$.

11. a. Equilibrium position shifts to the left; K value is not affected.

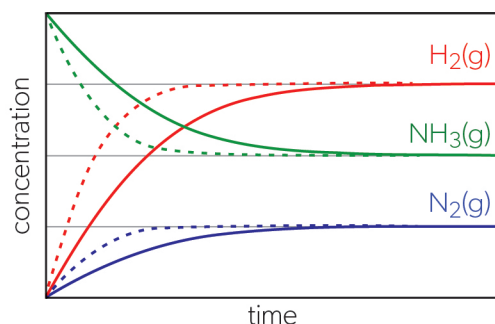
b. Equilibrium position shifts to the left; K value decreases.

c. Equilibrium position shifts to the right; K value is not affected.

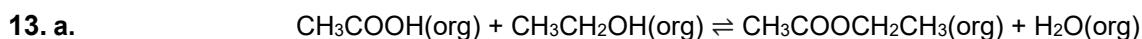
d. Equilibrium position shifts to the left; K value is not affected.

e. Neither equilibrium position nor K value is affected.

12.



The final concentrations of all species will not be affected, but the state of equilibrium will be reached faster.



$n_{\text{initial}} / \text{mol}$	1.00	2.00	0	0
$\Delta n / \text{mol}$	-0.60	-0.60	+0.60	+0.60
$n_{\text{final}} / \text{mol}$	0.40	1.40	0.60	0.60

The solution volume is 1.00 dm^3 , so the concentrations of all species are numerically equal to their amounts.

$$K = \frac{[\text{CH}_3\text{COOCH}_2\text{CH}_3][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{CH}_3\text{CH}_2\text{OH}]}$$

$$= \frac{0.60 \times 0.60}{0.40 \times 1.40}$$

$$\approx 0.64$$

Notice that water in this reaction is not the solvent, so its concentration must be included in the equilibrium constant expression.

b. If all four concentrations are identical, then $Q = 1$. Therefore, $Q > K$, so the reverse reaction is favoured.

14. a. The solution is similar to that given in worked example 5:

$$[\text{C}_6\text{H}_5\text{COO}^-] = [\text{H}^+]$$

$$= \sqrt{0.0200 \times 6.31 \times 10^{-5}}$$

$$\approx 0.00112 \text{ mol dm}^{-3}$$

b. $\Delta G^\ominus = -RT \ln K = -8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K} \times \ln(6.31 \times 10^{-5}) \approx 23\,949 \text{ J mol}^{-1} \approx 23.9 \text{ kJ mol}^{-1}$

Notice that the positive value of ΔG^\ominus suggests that the acid is weak (as the dissociation of that acid is not favoured).

15. a. $K = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$

The plot shows that the equilibrium concentrations of $\text{NO}_2(\text{g})$ and $\text{N}_2\text{O}_4(\text{g})$ are $0.800 \text{ mol dm}^{-3}$ and $0.400 \text{ mol dm}^{-3}$, respectively. Therefore:

$$K = \frac{0.800^2}{0.400} = 1.60$$

b. $\Delta G^\ominus = -RT \ln K = -8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 320 \text{ K} \times \ln 1.60 \approx -1250 \text{ J mol}^{-1} = -1.25 \text{ kJ mol}^{-1}$

c. $Q = \frac{0.10^2}{0.50} = 0.020$

$Q < K$, so the forward reaction is spontaneous.