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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 2
How Much, How Fast and How Far?

Paper 2 Style Practice · Short-answer and extended-response questions

Medium to hard exam-style questions on Reactivity 2.1 – 2.3, with a full markscheme

Standard Level & Higher Level

Timing and marks

Higher Level: answer all five questions. Total **65 marks**. Suggested time **about 105 minutes**.

Standard Level: answer Questions 1–3 and omit Questions 4 and 5 (HL). Total **41 marks**. Suggested time **about 75 minutes**.

These timings use the same rate as the real examination (SL Paper 2: 50 marks in 1 h 30; HL Paper 2: 90 marks in 2 h 30). Reactivity 2 contains a large amount of additional higher level material, so the HL paper is longer.

Inside this booklet

1. Question 1 — Reacting quantities, yield and atom economy (14 marks)
2. Question 2 — Rates of reaction and collision theory (14 marks)
3. Question 3 — Equilibrium and Le Chatelier's principle (13 marks)
4. Question 4 — HL: rate equations, mechanisms and the Arrhenius equation (12 marks)
5. Question 5 — HL: Q, equilibrium calculations and Gibbs energy (12 marks)
6. Markscheme and worked solutions

Instructions

- Work under examination conditions: set a timer, and do not look at the markscheme until you have finished.
- You need a **calculator** and a copy of the **IB Chemistry data booklet** (first assessment 2025). Where a question gives data, use those values. Molar volume of an ideal gas at STP (273 K, 100 kPa) = 22.7 $dm^3\ mol^{-1}$; R = 8.31 $J\ K^{-1}\ mol^{-1}$.
- Write your answers in the spaces provided. Show all working in calculations and give answers with units and an appropriate number of significant figures.
- The number of marks for each part is shown in brackets. As a guide, one mark usually requires one distinct, correct point.

Question 1 Reacting quantities, yield and atom economy (14 marks)

Aspirin is made by reacting salicylic acid with ethanoic anhydride.

$$C_7H_6O_3(s) + C_4H_6O_3(l) \rightarrow C_9H_8O_4(s) + CH_3COOH(l)$$

salicylic acid (M = 138.13) ethanoic anhydride (M = 102.10) aspirin (M = 180.17) ethanoic acid (M = 60.06)

A student heated 2.50 g of salicylic acid with 3.00 cm^3 of ethanoic anhydride (density 1.08 g cm^{-3}) and a few drops of acid catalyst.

(a) Determine the limiting reactant. Show your working. **[3]**

(b) Calculate the theoretical yield of aspirin, in g. **[1]**

(c) After recrystallization the student obtained 2.31 g of dry aspirin. Calculate the percentage yield. **[1]**

(d) Suggest two reasons why the percentage yield is less than 100 %. **[2]**

(e) Calculate the atom economy of this method of making aspirin. **[2]**

(f) Aspirin can also be made using ethanoyl chloride, CH_3COCl (M = 78.50): $C_7H_6O_3 + CH_3COCl \rightarrow C_9H_8O_4 + HCl$. The atom economy of this route is 83.2 %. Suggest two reasons why manufacturers still prefer to use ethanoic anhydride. **[2]**

(g) To check its purity, 0.300 g of the product was dissolved in ethanol and titrated with 0.100 mol dm^{-3} sodium hydroxide solution. 16.20 cm^3 was required. Aspirin reacts with NaOH in a 1 : 1 ratio. Calculate the percentage purity of the product, assuming the impurities do not react with NaOH. **[3]**

Question 2 Rates of reaction and collision theory **(14 marks)**

Hydrogen peroxide decomposes slowly at room temperature. The reaction is catalysed by manganese(IV) oxide, MnO_2.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

(a) State what is meant by the activation energy of a reaction. **[1]**

(b) Sketch, on the axes below, the Maxwell–Boltzmann distribution of molecular kinetic energies for a sample of gas at temperature T_1 and at a higher temperature T_2. Label the axes and mark the activation energy, E_a. Use your sketch to explain why the rate of reaction increases with temperature. **[4]**

(c) Explain, with reference to the Maxwell–Boltzmann distribution, how MnO_2 increases the rate of the decomposition. **[2]**

(d) (i) A student adds MnO_2 to 25.0 cm^3 of 0.400 mol dm^{-3} hydrogen peroxide. Calculate the maximum volume of oxygen that can form, in cm^3, measured at STP. **[2]**

(ii) 45.0 cm^3 of oxygen was collected in the first 30 s. Calculate the average rate of reaction over this time. **[1]**

(e) The student wants to compare MnO_2 with another catalyst. State two variables that must be controlled. **[2]**

(f) Explain why powdered MnO_2 is a more effective catalyst than the same mass of MnO_2 lumps. **[2]**

Question 3 Equilibrium and Le Chatelier's principle **(13 marks)**

One step in the Contact process for the manufacture of sulfuric acid is the oxidation of sulfur dioxide:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -198 \text{ kJ mol}^{-1}$$

(a) State two characteristics of a system in dynamic equilibrium. **[2]**

(b) Write the equilibrium constant expression, K, for this reaction. **[1]**

(c) At a certain temperature $K = 280$. Calculate K for $SO_3(g) \rightleftharpoons SO_2(g) + \frac{1}{2}O_2(g)$ at the same temperature. **[1]**

(d) Predict and explain the effect on the position of equilibrium, and on the value of *K*, of:

(i) increasing the total pressure at constant temperature **[2]**

(ii) increasing the temperature at constant pressure. **[2]**

(e) The reaction is carried out at about 450 °C and 100–200 kPa, using a vanadium(V) oxide catalyst.

(i) Explain why a temperature of about 450 °C is used. **[2]**

(ii) Suggest why a much higher pressure is not used, even though it would increase the equilibrium yield. **[1]**

(f) State the effect of the vanadium(V) oxide catalyst on the position of equilibrium and on the value of *K*, and explain why it is used. **[2]**

HL Question 4 Rate equations, mechanisms and the Arrhenius equation **(12 marks)**

Nitrogen monoxide reacts with bromine: $2NO(g) + Br_2(g) \rightarrow 2NOBr(g)$. The initial rates below were measured at 298 K.

Experiment	[NO] / mol dm^{-3}	[Br_2] / mol dm^{-3}	Initial rate / mol dm^{-3} s^{-1}
1	0.10	0.10	1.2×10^{-3}
2	0.20	0.10	4.8×10^{-3}
3	0.10	0.20	2.4×10^{-3}

(a) Deduce the order of reaction with respect to NO and with respect to Br_2. **[2]**

(b) Write the rate equation and calculate the rate constant, k, with its units. **[2]**

Two mechanisms are proposed.

Mechanism 1	$2NO + Br_2 \rightarrow 2NOBr$ (one step)
Mechanism 2	step 1: $NO + Br_2 \rightleftharpoons NOBr_2$ (fast) step 2: $NOBr_2 + NO \rightarrow 2NOBr$ (slow)

(c) Identify the intermediate in mechanism 2, and explain why mechanism 2 is consistent with your rate equation. **[2]**

(d) State the molecularity of the rate-determining step in mechanism 2, and suggest why mechanism 2 is more likely than mechanism 1. **[1]**

(e) Sketch a graph of initial rate against [NO], with $[Br_2]$ kept constant. **[1]**

(f) The rate constant is 3.6 $mol^{-2}\ dm^6\ s^{-1}$ at 323 K. Calculate the activation energy, in $kJ\ mol^{-1}$. **[2]**

(g) The Arrhenius equation is $k = Ae^{-E_a/RT}$. State what the Arrhenius factor, A, represents, and explain why k increases with temperature. **[2]**

HL Question 5 *Q*, equilibrium calculations and Gibbs energy (12 marks)

The reverse water–gas shift reaction is used to convert carbon dioxide into carbon monoxide. At 700 K, $K = 0.250$.

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g) \quad \Delta H = +41 \text{ kJ mol}^{-1}$$

(a) A mixture at 700 K contains $[CO_2] = [H_2] = 0.200$ mol dm^{-3} and $[CO] = [H_2O] = 0.150$ mol dm^{-3}. Calculate the reaction quotient, Q, and deduce the direction in which the reaction proceeds. **[2]**

(b) 0.500 mol dm^{-3} each of CO_2 and H_2 are allowed to reach equilibrium at 700 K. Calculate the equilibrium concentrations of all four species. **[3]**

(c) Calculate $\Delta G^{\ominus}$ for the reaction at 700 K, and state what its sign shows about the position of equilibrium. **[2]**

(d) Predict the effect of increasing the temperature on the value of K and on the value of $\Delta G^{\ominus}$. Give a reason. **[2]**

(e) In a car engine at 1500 K, nitrogen and oxygen react: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$, $K = 1.0 \times 10^{-5}$. The initial concentrations are $[N_2] = 0.800$ and $[O_2] = 0.200$ mol dm^{-3}. Calculate the equilibrium concentration of NO, stating and justifying the approximation you make. **[3]**

End of questions

Markscheme and worked solutions

Bold **[1]** shows where each mark is awarded. Accept equivalent wording that conveys the same chemistry. Text in italics gives examiner-style guidance. Allow follow-through (ECF) from an earlier error unless stated.

Question 1 (14 marks)

(a) n(salicylic acid) = 2.50/138.13 = 0.0181 mol **[1]**; mass of anhydride = 3.00 × 1.08 = 3.24 g, so n = 3.24/102.10 = 0.0317 mol **[1]**; the ratio is 1 : 1, so salicylic acid is the limiting reactant **[1]**.

(b) 0.0181 × 180.17 = 3.26 g **[1]**.

(c) 2.31/3.26 × 100 = 70.8 % **[1]** *(ECF from (b))*.

(d) Any two **[2]**: the reaction is reversible (or incomplete) or side reactions occur; product lost on transfer or left in the filtrate during recrystallization (aspirin is slightly soluble); product lost during filtration; not all the crystals dried or collected.

(e) Atom economy = 180.17/(138.13 + 102.10) × 100 **[1]** = 75.0 % **[1]**.

(f) Any two **[2]**: the by-product HCl is a corrosive, toxic gas; ethanoyl chloride reacts violently with water and is hazardous to handle; ethanoyl chloride is more expensive; the by-product ethanoic acid from the anhydride route is useful and can be sold or recycled. *Do not accept "it has a higher atom economy".*

(g) n(NaOH) = 0.100 × 0.01620 = 1.62 × 10^{-3} mol = n(aspirin) **[1]**; mass = 1.62 × 10^{-3} × 180.17 = 0.292 g **[1]**; purity = 0.292/0.300 × 100 = 97.3 % **[1]**.

Question 2 (14 marks)

(a) The minimum energy that colliding particles need for a reaction to occur **[1]**.

(b) Axes labelled: number (fraction) of particles against kinetic energy, with both curves starting at the origin **[1]**; the T_2 curve has a lower peak at higher energy, the same area, and lies above T_1 at high energy **[1]**; E_a marked on the energy axis, to the right of both peaks **[1]**; at T_2 a greater proportion of particles has $E \geq E_a$ (larger area beyond E_a), so a greater frequency of successful collisions **[1]**.

(c) The catalyst provides an alternative pathway with a lower activation energy; the lower E_a is further to the left on the distribution **[1]**; a greater proportion of molecules have energy ≥ the new E_a, so there are more successful collisions per unit time **[1]**.

(d) (i) $n(H_2O_2)$ = 0.400 × 0.0250 = 0.0100 mol, so $n(O_2)$ = 5.00 × 10^{-3} mol **[1]**; V = 5.00 × 10^{-3} × 22.7 × 1000 = 114 cm^3 **[1]**.

(ii) 45.0/30 = 1.50 $cm^3\ s^{-1}$ **[1]**.

(e) Any two **[2]**: volume and concentration of the hydrogen peroxide; mass of catalyst; particle size (surface area) of the catalyst; temperature.

(f) The powder has a larger surface area, so more active sites are available **[1]**; more H_2O_2 molecules can collide with (be adsorbed on) the catalyst surface per unit time **[1]**.

Question 3 (13 marks)

(a) Any two **[2]**: the rates of the forward and backward reactions are equal; the concentrations of reactants and products remain constant; it is established in a closed system; both reactions continue at the molecular level.

(b) $K = [SO_3]^2/[SO_2]^2[O_2]$ **[1]**.

(c) $K' = (1/280)^{½}$ = 0.0598 **[1]**.

(d) (i) Position shifts to the right, towards fewer gas molecules (3 − 2), which reduces the pressure **[1]**; K is unchanged because temperature is constant **[1]**.

(ii) Position shifts to the left, because the backward reaction is endothermic and absorbs the added heat **[1]**; K decreases **[1]**.

(e) (i) A lower temperature would give a higher equilibrium yield, because the forward reaction is exothermic **[1]**; but the rate would be too slow (and the catalyst would be inactive), so 450 °C is a compromise between rate and yield **[1]**.

(ii) The equilibrium already lies far to the right (high conversion) at low pressure, so the small gain in yield does not justify the cost of high-pressure equipment and energy (or the safety risk) **[1]**.

(f) No effect on the position of equilibrium or on K **[1]**; it increases the rates of the forward and backward reactions equally, so equilibrium is reached faster at a lower temperature, saving time and energy **[1]**.

HL Question 4 (12 marks)

(a) Experiments 1 → 2: [NO] doubles and the rate increases by a factor of 4, so second order in NO **[1]**; experiments 1 → 3: [Br_2] doubles and the rate doubles, so first order in Br_2 **[1]**.

(b) rate = $k[NO]^2[Br_2]$ **[1]**; $k = 1.2 \times 10^{-3}/(0.10^2 \times 0.10) = 1.2$ mol^{-2} dm^6 s^{-1} **[1]**.

(c) $NOBr_2$ **[1]**; the slow step involves $NOBr_2$ and NO, and $NOBr_2$ is formed from one NO and one Br_2, so the rate depends on two NO and one Br_2, giving rate = $k[NO]^2[Br_2]$ **[1]**.

(d) Bimolecular; a single step needs three molecules to collide at the same time with the correct orientation (termolecular), which is very unlikely **[1]**.

(e) A curve through the origin that gets steeper (a parabola, rate $\propto [NO]^2$) **[1]**.

(f) $\ln(3.6/1.2) = (E_a/8.31)(1/298 - 1/323)$ **[1]**; $E_a = 3.51 \times 10^4$ J mol^{-1} = 35.1 kJ mol^{-1} **[1]**.

(g) A represents the frequency of collisions with the correct orientation (the steric factor) **[1]**; as T increases, $e^{-E_a/RT}$ (the fraction of collisions with $E \geq E_a$) increases sharply, so k increases **[1]**.

HL Question 5 (12 marks)

(a) $Q = (0.150 \times 0.150)/(0.200 \times 0.200) = 0.563$ **[1]**; $Q > K$, so net reaction to the left, forming more CO_2 and H_2 **[1]**.

(b) ICE: changes $-x$, $-x$, $+x$, $+x$; $K = x^2/(0.500 - x)^2 = 0.250$ **[1]**; square root: $x/(0.500 - x) = 0.500$, so $x = 0.167$ **[1]**; [CO] = [H_2O] = 0.167 mol dm^{-3} and [CO_2] = [H_2] = 0.333 mol dm^{-3} **[1]**.

(c) $\Delta G^\ominus = -8.31 \times 700 \times \ln 0.250 = +8.06 \times 10^3$ J mol^{-1} = +8.06 kJ mol^{-1} **[1]**; positive, so $K < 1$ and the equilibrium lies to the left (reactants favoured) **[1]**.

(d) The forward reaction is endothermic, so increasing T shifts the equilibrium to the right and K increases **[1]**; since $\Delta G^\ominus = -RT \ln K$, $\Delta G^\ominus$ becomes less positive (decreases) **[1]**.

(e) K is very small, so [N_2] ≈ 0.800 and [O_2] ≈ 0.200 at equilibrium **[1]**; $(2x)^2/(0.800 \times 0.200) = 1.0 \times 10^{-5}$, so $x = 6.3 \times 10^{-4}$ and [NO] = $2x = 1.3 \times 10^{-3}$ mol dm^{-3} **[1]**; x is only 0.3 % of 0.200, so the approximation is justified **[1]**.

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