

A-Level Chemistry – Physical: Equilibria, Acids and Redox

A-Level Chemistry (Edexcel/AQA/OCR) · Total marks: 40 · Answers on the last page

This worksheet covers chemical equilibria (K_c and K_p), acids, bases and buffers, and redox/electrochemical cells. A calculator is allowed; use $K_w = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 298 K.

Section A – Chemical Equilibria

- For the equilibrium $\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)$, the equilibrium concentrations in a 1 dm^3 container are $[\text{N}_2] = 0.20 \text{ mol dm}^{-3}$, $[\text{H}_2] = 0.60 \text{ mol dm}^{-3}$ and $[\text{NH}_3] = 0.50 \text{ mol dm}^{-3}$.
 - Write the expression for K_c .
 - Calculate the value of K_c , including units. [4]
- The forward reaction $\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)$ is exothermic.
 - Predict and explain the effect of increasing the pressure on the position of equilibrium.
 - Predict and explain the effect of increasing the temperature on the value of K_c . [4]
- For the equilibrium $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$, an equilibrium mixture at total pressure 200 kPa contains 0.40 mol SO_2 , 0.20 mol O_2 and 1.40 mol SO_3 (total 2.00 mol).
 - Calculate the partial pressure of each gas.
 - Write the expression for K_p and calculate its value, including units. [5]
- Explain why adding a catalyst to an equilibrium reaction changes the rate at which equilibrium is reached but does not change the value of K_c or the position of equilibrium. [2]

Section B – Acids, Bases and Buffers

- Sulfuric acid, H_2SO_4 , is a strong diprotic acid that can be assumed to fully dissociate to give two moles of H^+ per mole of acid. Calculate the pH of a $0.020 \text{ mol dm}^{-3}$ solution of sulfuric acid. [3]
- A weak acid HA has $K_a = 1.7 \times 10^{-4} \text{ mol dm}^{-3}$. Using the approximation $[\text{H}^+] \approx \sqrt{K_a c}$, calculate the pH of a $0.050 \text{ mol dm}^{-3}$ solution of HA. [3]
- A buffer solution is made from 0.25 mol dm^{-3} propanoic acid ($K_a = 1.3 \times 10^{-5} \text{ mol dm}^{-3}$) and 0.15 mol dm^{-3} sodium propanoate. Use $\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$ to calculate the pH of the buffer. [3]
- Explain how the propanoic acid/sodium propanoate buffer resists a change in pH when:
 - a small amount of acid is added,
 - a small amount of alkali is added. [3]

Section C – Redox and Electrochemical Cells

- A cell is made from a silver half-cell ($\text{Ag}^+ + e^- \rightleftharpoons \text{Ag}$, $E^\ominus = +0.80 \text{ V}$) and a zinc half-cell ($\text{Zn}^{2+} + 2e^- \rightleftharpoons \text{Zn}$, $E^\ominus = -0.76 \text{ V}$).
 - State which half-cell is the positive electrode.
 - Calculate the standard EMF of the cell.
 - Write the overall (doubled) ionic equation for the spontaneous cell reaction. [5]
- Using $E^\ominus(\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$ and $E^\ominus(\text{Fe}^{3+}/\text{Fe}^{2+}) = +0.77 \text{ V}$, predict, with a calculation,

whether Fe^{3+} ions can oxidise copper metal to Cu^{2+} under standard conditions. [3]

11. Chlorine reacts with cold dilute sodium hydroxide: $\text{Cl}_2 + 2\text{NaOH} \rightarrow \text{NaCl} + \text{NaClO} + \text{H}_2\text{O}$.
(a) Give the oxidation state of chlorine in Cl_2 , in NaCl , and in NaClO .
(b) State the type of redox reaction this represents, where the same element is both oxidised and reduced. [3]
12. A student uses standard electrode potentials to predict that a reaction with a large positive $E_{\text{cell}}^\ominus$ will occur rapidly at room temperature. Explain why this prediction may not always be correct. [2]

Answers

Check your working against these. If an answer does not match, try the question again before looking for the method.

- $K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$
 - $K_c = \frac{0.50^2}{0.20 \times 0.60^3} \approx 5.79 \text{ mol}^{-2} \text{ dm}^6$
- The position of equilibrium shifts to the right (towards NH_3), because there are fewer gas moles on the product side (2 mol vs 4 mol), so increasing pressure favours the side with fewer moles of gas.
 - K_c decreases, because the reaction is exothermic, so by Le Chatelier's principle increasing temperature shifts equilibrium in the endothermic (reverse) direction, reducing the proportion of products at equilibrium.
- $p_{\text{SO}_2} = 200 \times \frac{0.40}{2.00} = 40 \text{ kPa}$; $p_{\text{O}_2} = 200 \times \frac{0.20}{2.00} = 20 \text{ kPa}$; $p_{\text{SO}_3} = 200 \times \frac{1.40}{2.00} = 140 \text{ kPa}$
 - $K_p = \frac{p_{\text{SO}_3}^2}{p_{\text{SO}_2}^2 p_{\text{O}_2}} = \frac{140^2}{40^2 \times 20} \approx 0.613 \text{ kPa}^{-1}$
- A catalyst provides an alternative pathway with lower activation energy, speeding up both the forward and reverse reactions equally, so equilibrium is reached faster; because both rates increase by the same factor, the equilibrium concentrations (and therefore K_c) are unchanged.
- $[\text{H}^+] = 2 \times 0.020 = 0.040 \text{ mol dm}^{-3}$
 $\text{pH} = -\log(0.040) \approx 1.40$
- $[\text{H}^+] = \sqrt{(1.7 \times 10^{-4})(0.050)} \approx 2.92 \times 10^{-3} \text{ mol dm}^{-3}$
 $\text{pH} = -\log(2.92 \times 10^{-3}) \approx 2.54$
- $\text{p}K_a = -\log(1.3 \times 10^{-5}) \approx 4.89$
 $\text{pH} = 4.89 + \log\left(\frac{0.15}{0.25}\right) = 4.89 - 0.22 \approx 4.66$
- The propanoate ions ($\text{CH}_3\text{CH}_2\text{COO}^-$) react with the added H^+ to reform propanoic acid, removing most of the extra H^+ and keeping the pH almost unchanged.
 - The propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$) reacts with the added OH^- , donating a proton to neutralise it and forming water and propanoate ions, so the pH changes only slightly.
- The silver half-cell (more positive $E^\ominus$, so silver is reduced) is the positive electrode.
 - $E_{\text{cell}}^\ominus = E_{\text{cathode}}^\ominus - E_{\text{anode}}^\ominus = 0.80 - (-0.76) = 1.56 \text{ V}$
 - $\text{Zn} + 2\text{Ag}^+ \rightarrow \text{Zn}^{2+} + 2\text{Ag}$
- $E_{\text{cell}}^\ominus = E_{\text{reduced}}^\ominus - E_{\text{oxidised}}^\ominus = 0.77 - 0.34 = +0.43 \text{ V}$ (positive), so the reaction $\text{Cu} + 2\text{Fe}^{3+} \rightarrow \text{Cu}^{2+} + 2\text{Fe}^{2+}$ is feasible under standard conditions.
- Cl_2 : 0. NaCl : -1. NaClO : +1.
 - This is a disproportionation reaction: chlorine (oxidation state 0) is simultaneously reduced to -1 (in Cl^-) and oxidised to +1 (in ClO^-).
- $E_{\text{cell}}^\ominus$ values give only thermodynamic information (whether a reaction is energetically feasible), not kinetic information; a reaction with a large positive $E_{\text{cell}}^\ominus$ may still have a high activation energy and therefore proceed very slowly (or not be observed) at room temperature.

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