

Y13 Chemistry Homework Book 1

Physical and Inorganic Chemistry

Name:

Teacher:

Y12-13 Transition

Grade achieved in end of year exam:

My personal target grade:

What am I applying to do at Uni:

What topics I found difficult (and why you think that was the case)

Rates and Kp

Rate equations	
T 1.	Define the terms order of reaction and rate constant
T 2.	Perform calculations using the rate equation
T 3.	Explain the qualitative effect of changes in temperature on the rate constant k
T 4.	Perform calculations using the equation $k = Ae^{-E_a/RT}$
T 5.	Understand that the equation $k = Ae^{-E_a/RT}$ can be rearranged into the form $\ln k = -E_a/RT + \ln A$ and know how to use this rearranged equation with experimental data to plot a straight line graph with slope $-E_a/R$
T 6.	Use concentration–time graphs to deduce the rate of a reaction
T 7.	Use initial concentration–time data to deduce the initial rate of a reaction
T 8.	Use rate–concentration data or graphs to deduce the order (0, 1 or 2) with respect to a reactant
T 9.	Derive the rate equation for a reaction from the orders with respect to each of the reactants
T 10.	Use the orders with respect to reactants to provide information about the rate determining/limiting step of a reaction.
Equilibrium and Kp	
T 1.	Derive partial pressure from mole fraction and total pressure
T 2.	Construct an expression for Kp for a homogeneous system in equilibrium
T 3.	Perform calculations involving Kp
T 4.	Predict the qualitative effects of changes in temperature and pressure on the position of equilibrium
T 5.	Predict the qualitative effects of changes in temperature on the value of Kp
T 6.	Understand that, whilst a catalyst can affect the rate of attainment of an equilibrium, it does not affect the value of the equilibrium constant.

Exam booklet reflection:

WWW

EBI

Checked by teacher

Doddle quiz homework:

Kinetics

Score:

Rate equations

Score:

Rate graphs

Score:

A LEVEL CHEMISTRY

TOPIC 11 – HOW FAR HOW FAST II

TEST

Answer all questions

Max 50 marks

Name		
Mark/50%	Grade

SECTION A

1. The rate of hydrolysis of an ester X ($\text{HCOOCH}_2\text{CH}_2\text{CH}_3$) was studied in alkaline conditions at a given temperature. The rate was found to be first order with respect to the ester and first order with respect to hydroxide ions.

(a) (i) Name ester X.

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(1)

(ii) Using X to represent the ester, write a rate equation for this hydrolysis reaction.

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(1)

(iii) When the initial concentration of X was $0.024 \text{ mol dm}^{-3}$ and the initial concentration of hydroxide ions was $0.035 \text{ mol dm}^{-3}$, the initial rate of the reaction was $8.5 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$.

Calculate a value for the rate constant at this temperature and give its units.

Calculation

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Units

(3)

(iv) In a second experiment at the same temperature, water was added to the original reaction mixture so that the total volume was doubled.

Calculate the initial rate of reaction in this second experiment.

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(1)

(v) In a third experiment at the same temperature, the concentration of X was half that used in the experiment in part (a) (iii) and the concentration of hydroxide ions was three times the original value.

Calculate the initial rate of reaction in this third experiment.

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(1)

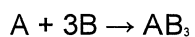
(vi) State the effect, if any, on the value of the rate constant k when the temperature is lowered but all other conditions are kept constant. Explain your answer.

Effect

Explanation

(2)

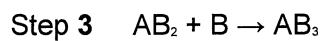
(b) Compound **A** reacts with compound **B** as shown by the overall equation



The rate equation for the reaction is

$$\text{rate} = k[\text{A}][\text{B}]^2$$

A suggested mechanism for the reaction is



Deduce which one of the three steps is the rate-determining step.

Explain your answer.

Rate-determining step

Explanation

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(2)
(Total 11 marks)

2. (a) The initial rate of the reaction between compounds **A** and **B** was measured in a series of experiments at a fixed temperature. The following rate equation was deduced.

$$\text{rate} = k[\text{A}][\text{B}]^2$$

- (i) Complete the table of data below for the reaction between **A** and **B**.

Expt	Initial [A] /mol dm ⁻³	Initial [B] /mol dm ⁻³	Initial rate /mol dm ⁻³ s ⁻¹
1	4.80×10^{-2}	6.60×10^{-2}	10.4×10^{-3}
2	4.80×10^{-2}	3.30×10^{-2}	
3		13.2×10^{-2}	5.20×10^{-3}
4	1.60×10^{-2}		10.4×10^{-3}

- (ii) Using the data for experiment 1, calculate a value for the rate constant, k , and state its units.

Calculation

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Units

(6)

- (b) State how the value of the rate constant, k , would change, if at all, if the concentration of **A** were increased in a series of experiments.

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(1)

(Total 7 marks)

3. (a) The following data were obtained in a series of experiments on the rate of the reaction between compounds **A** and **B** at a constant temperature.

Experiment	Initial concentration of A /mol dm ⁻³	Initial concentration of B /mol dm ⁻³	Initial rate/mol dm ⁻³ s ⁻¹
1	0.12	0.15	0.32×10^{-3}
2	0.36	0.15	2.88×10^{-3}
3	0.72	0.30	11.52×10^{-3}

- (i) Deduce the order of reaction with respect to **A**.

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- (ii) Deduce the order of reaction with respect to **B**.

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(2)

- (b) The following data were obtained in a series of experiments on the rate of the reaction between NO and O₂ at a constant temperature.

Experiment	Initial concentration of NO/mol dm ⁻³	Initial concentration of O ₂ /mol dm ⁻³	Initial rate/mol dm ⁻³ s ⁻¹
4	5.0×10^{-2}	2.0×10^{-2}	6.5×10^{-4}
5	6.5×10^{-2}	3.4×10^{-2}	To be calculated

The rate equation for this reaction is

$$\text{rate} = k[\text{NO}]^2[\text{O}_2]$$

- (i) Use the data from experiment 4 to calculate a value for the rate constant, k , at this temperature, and state its units.

Value of k

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Units of k

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- (ii) Calculate a value for the initial rate in experiment 5.

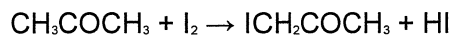
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(4)
(Total 6 marks)

4. Propanone and iodine react in acidic conditions according to the following equation.



A student studied the kinetics of this reaction using hydrochloric acid and a solution containing propanone and iodine. From the results the following rate equation was deduced.

$$\text{rate} = k[\text{CH}_3\text{COCH}_3][\text{H}^+]$$

- (a) Give the overall order for this reaction.

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(1)

- (b) When the initial concentrations of the reactants were as shown in the table below, the initial rate of reaction was found to be $1.24 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$.

	initial concentration / mol dm^{-3}
CH_3COCH_3	4.40
I_2	5.00×10^{-3}
H^+	0.820

Use these data to calculate a value for the rate constant, k , for the reaction and give its units.

Calculation

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Units

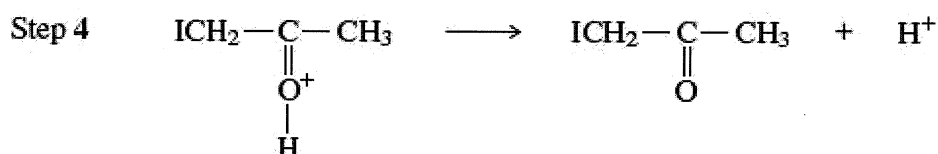
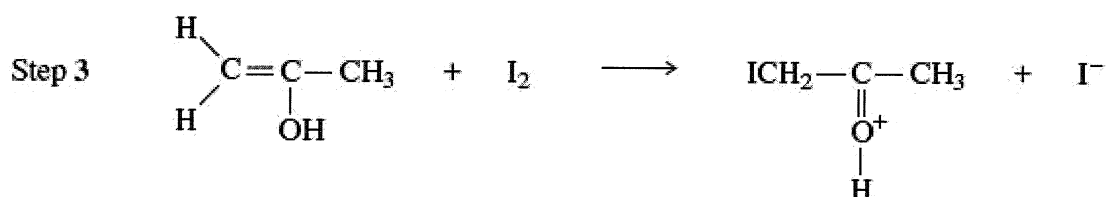
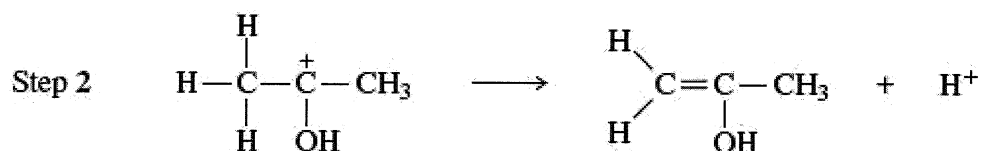
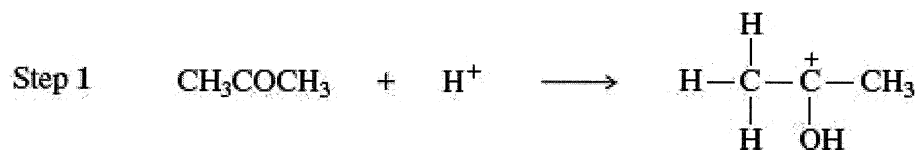
(3)

- (c) Deduce how the initial rate of reaction changes when the concentration of iodine is doubled but the concentrations of propanone and of hydrochloric acid are unchanged.

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(1)

(d) The following mechanism for the overall reaction has been proposed.



Use the rate equation to suggest which of the four steps could be the rate-determining step. Explain your answer.

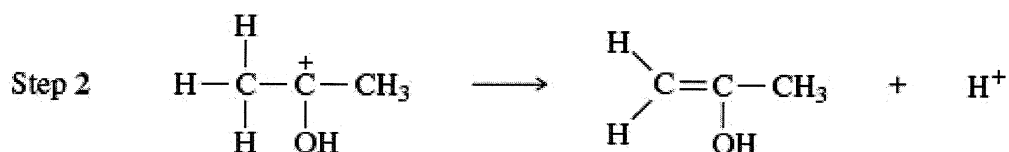
Rate-determining step

Explanation

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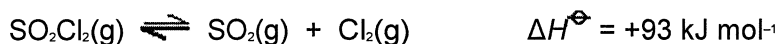
(2)

(e) Use your understanding of reaction mechanisms to predict a mechanism for Step 2 by adding one or more curly arrows as necessary to the structure of the carbocation below.



(1)
(Total 8 marks)

5. At high temperatures, SO_2Cl_2 dissociates according to the following equation.



When 1.00 mol of SO_2Cl_2 dissociates, the equilibrium mixture contains 0.75 mol of Cl_2 at 673 K and a total pressure of 125 kPa.

- (a) Write an expression for the equilibrium constant, K_p , for this reaction.

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(1)

- (b) Calculate the total number of moles of gas present in the equilibrium mixture.

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(2)

- (c) (i) Write a general expression for the partial pressure of a gas in a mixture of gases in terms of the total pressure.

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- (ii) Calculate the partial pressure of SO_2Cl_2 and the partial pressure of Cl_2 in the equilibrium mixture.

Partial pressure of SO_2Cl_2

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Partial pressure of Cl_2

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(5)

- (d) Calculate a value for the equilibrium constant, K_p , for this reaction and give its units.

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(3)

- (e) State the effect, if any, of an increase in temperature on the value of K_p for this reaction. Explain your answer.

Effect on K_p

Explanation

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(2)

- (f) State the effect, if any, of an increase in the total pressure on the value of K_p for this reaction.

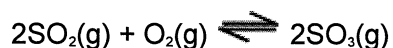
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(1)

(Total 14 marks)

SECTION B

This question relates to the equilibrium gas-phase synthesis of sulphur trioxide:



Thermodynamic data for the components of this equilibrium are:

This equilibrium, at a temperature of 585 K and a total pressure of 540 kPa, occurs in a vessel of volume 1.80 dm³. At equilibrium, the vessel contains 0.0500 mol of SO₂(g), 0.0800 mol of O₂(g) and 0.0700 mol of SO₃(g).

6. The mole fraction of SO₃ in the equilibrium mixture is

A 0.250
B 0.350
C 0.440
D 0.700

(Total 1 mark)

7. With pressures expressed in MPa units, the value of the equilibrium constant, K_p , is

A 4.90
B 6.48
C 9.07
D 16.8

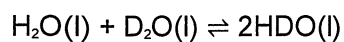
(Total 1 mark)

8. Possible units for the equilibrium constant K_p include

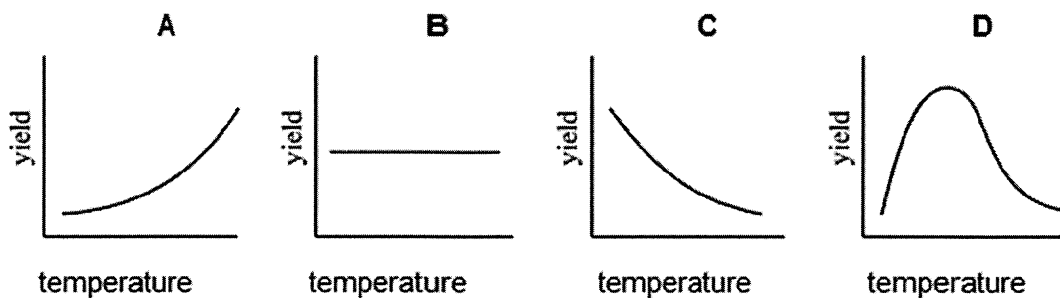
A no units
B kPa
C Mpa⁻¹
D kPa⁻²

(Total 1 mark)

9. Normal water and heavy water react together to form isotopically mixed water according to the equation



The standard enthalpy of formation of $\text{H}_2\text{O}(\text{l})$ is -286 kJ mol^{-1} , that of $\text{D}_2\text{O}(\text{l})$ is -294 kJ mol^{-1} , and that of $\text{HDO}(\text{l})$ is -290 kJ mol^{-1} . Which one of the following best represents the variation with temperature of the yield of HDO at equilibrium?



(Total 1 mark)

Acids

Acids and bases
T 1. Convert concentration of hydrogen ions into pH and vice versa
T 2. Calculate the pH of a solution of a strong acid from its concentration.
T 3. Use K_w to calculate the pH of a strong base from its concentration.
T 4. Construct an expression for K_a
T 5. Perform calculations relating the pH of a weak acid to the concentration of the acid and the dissociation constant, K_a
T 6. Convert K_a into pK_a and vice versa.
T 7. Perform calculations for acid base titrations based on experimental results.
T 8. Sketch and explain the shapes of typical pH curves
T 9. Use pH curves to select an appropriate indicator.
T 10. Explain qualitatively the action of acidic and basic buffers
T 11. Calculate the pH of acidic buffer solutions.

Exam booklet reflection:

WWW

EBI

Checked by teacher

Doddle quiz homework:

Acids and bases

Score:

Neutralisation

Score:

Buffers

Score:

A LEVEL CHEMISTRY

TOPIC 12 – ACIDS, BASES AND BUFFERS

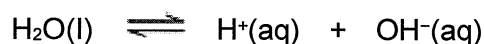
TEST

Answer all questions

Max 50 marks

Name		
Mark/50%	Grade

1. Water dissociates slightly according to the equation:



The ionic product of water, K_w , is given by the expression

$$K_w = [\text{H}^+][\text{OH}^-]$$

K_w varies with temperature as shown in the table.

Temperature / °C	K_w / mol ² dm ⁻⁶
25	1.00×10^{-14}
50	5.48×10^{-14}

- (a) Explain why the expression for K_w does **not** include the concentration of water.

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(2)

- (b) Explain why the value of K_w increases as the temperature increases.

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(2)

- (c) Calculate the pH of pure water at 50 °C.
Give your answer to 2 decimal places.

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(3)

- (d) Calculate the pH of 0.12 mol dm^{-3} aqueous NaOH at 50°C .
Give your answer to 2 decimal places.

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(3)
(Total 10 marks)

2. (a) A sample of hydrochloric acid has a pH of 2.34
Write an expression for pH and calculate the concentration of this acid.

pH

Concentration

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(2)

- (b) A $0.150 \text{ mol dm}^{-3}$ solution of a weak acid, HX, also has a pH of 2.34

- (i) Write an expression for the acid dissociation constant, K_a , for the acid HX.

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- (ii) Calculate the value of K_a for this acid and state its units.

Calculation

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Units

- (iii) Calculate the value of $\text{p}K_a$ for the acid HX. Give your answer to two decimal places.

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(5)

- (c) A 30.0 cm^3 sample of a $0.480 \text{ mol dm}^{-3}$ solution of potassium hydroxide was partially neutralised by the addition of 18.0 cm^3 of a $0.350 \text{ mol dm}^{-3}$ solution of sulphuric acid.

- (i) Calculate the initial number of moles of potassium hydroxide.

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- (ii) Calculate the number of moles of sulphuric acid added.

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- (iii) Calculate the number of moles of potassium hydroxide remaining in excess in the solution formed.

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- (iv) Calculate the concentration of hydroxide ions in the solution formed.

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- (v) Hence calculate the pH of the solution formed. Give your answer to two decimal places.

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(6)
(Total 13 marks)

3. The value of the acid dissociation constant, K_a , for the weak acid HA, at 298 K, is $1.45 \times 10^{-4} \text{ mol dm}^{-3}$.

(a) Write an expression for the term K_a for the weak acid HA.

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(1)

(b) Calculate the pH of a $0.250 \text{ mol dm}^{-3}$ solution of HA at 298 K.

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(4)

(c) A mixture of the acid HA and the sodium salt of this acid, NaA, can be used to prepare a buffer solution.

(i) State and explain the effect on the pH of this buffer solution when a small amount of hydrochloric acid is added.

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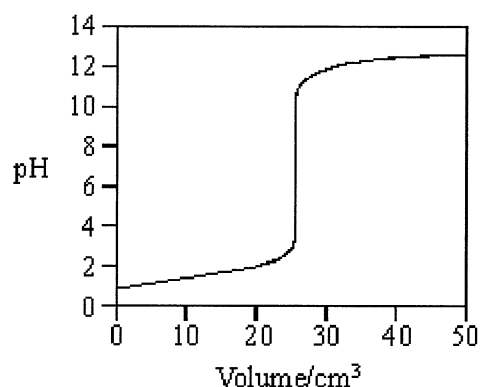
(ii) The concentration of HA in a buffer solution is $0.250 \text{ mol dm}^{-3}$. Calculate the concentration of A^- in this buffer solution when the pH is 3.59

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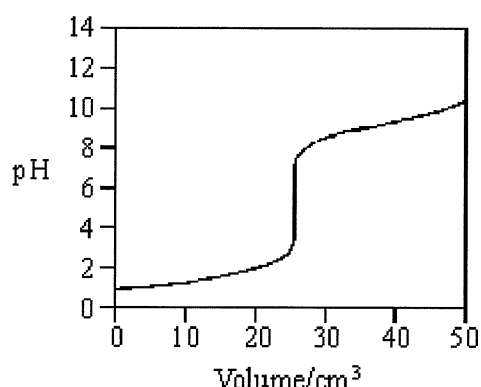
(6)

(Total 11 marks)

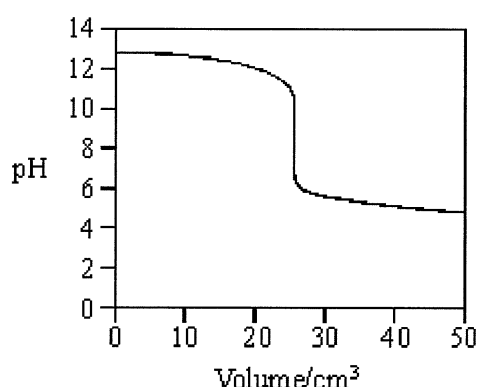
4. (a) Titration curves labelled **A**, **B**, **C** and **D** for combinations of different acids and bases are shown below. All solutions have a concentration of 0.1 mol dm^{-3} .



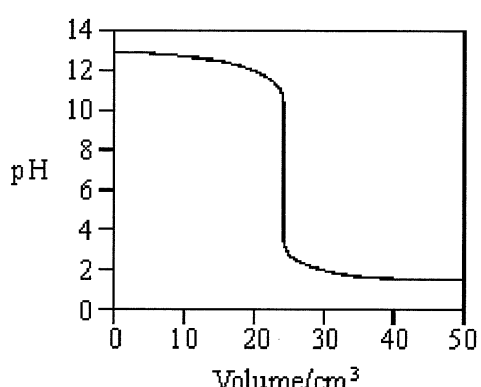
A



B



C



D

- (i) Select from **A**, **B**, **C** and **D** the curve produced by the addition of
- ammonia to 25 cm^3 of hydrochloric acid
- ethanoic acid to 25 cm^3 of sodium hydroxide
- sodium hydroxide to 25 cm^3 of hydrochloric acid
- (ii) A table of acid–base indicators and the pH ranges over which they change colour is shown below.

Indicator	pH range
Thymol blue	1.2 – 2.8
Bromophenol blue	3.0 – 4.6
Methyl red	4.2 – 6.3
Cresolphthalein	8.2 – 9.8
Thymolphthalein	9.3 – 10.5

Select from the table an indicator which could be used in the titration which produces curve **A** but not in the titration which produces curve **B**.

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(4)

- (b) (i) Write an expression for the term pH .

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- (ii) A solution of potassium hydroxide has a pH of 11.90 at 25°C. Calculate the concentration of potassium hydroxide in the solution.

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(4)

- (c) The acid dissociation constant, K_a , for propanoic acid has the value of $1.35 \times 10^{-5} \text{ mol dm}^{-3}$ at 25 °C.

$$K_a = \frac{[H^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$

In each of the calculations below, give your answer to 2 decimal places.

- (i) Calculate the pH of a $0.117 \text{ mol dm}^{-3}$ aqueous solution of propanoic acid.

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- (ii) Calculate the pH of a mixture formed by adding 25 cm³ of a $0.117 \text{ mol dm}^{-3}$ aqueous solution of sodium propanoate to 25 cm³ of a $0.117 \text{ mol dm}^{-3}$ aqueous solution of propanoic acid.

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(5)

(Total 13 marks)

6. The pH of 0.001 M NaOH at 25°C is

- A 13
- B 11
- C 9
- D 3

(Total 1 mark)

7. A weak acid HA dissociates in aqueous solution as shown below

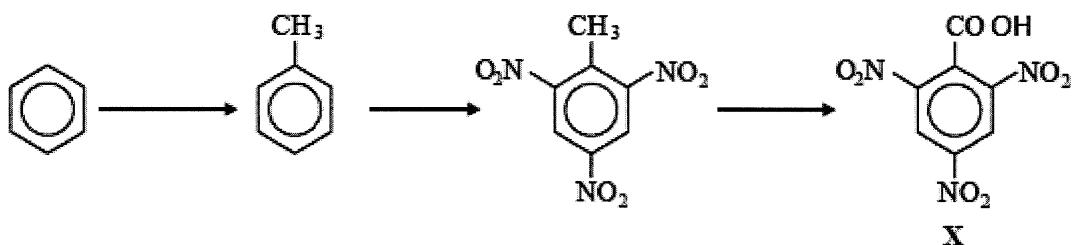


Which one of the following changes will result in a decrease in the pH of an aqueous solution of the acid?

- A addition of a little aqueous sodium hydroxide solution
- B raising the temperature of the solution
- C dissolving a little of the sodium salt, NaA, in the solution
- D adding a platinum catalyst to the solution

(Total 1 mark)

8. This question is based on the reactions and compounds shown in the scheme below.



A 0.100 mol dm⁻³ solution of **X** is found to have a pH of 2.50. The value of K_a in mol dm⁻³ is

- A 3.16×10^{-2}
- B 3.16×10^{-3}
- C 1.00×10^{-4}
- D 1.00×10^{-5}

(Total 1 mark)

Thermodynamics

Thermodynamics
T 1. Define each of the terms enthalpy of formation, ionisation energy, enthalpy of atomisation, bond enthalpy, electron affinity and lattice enthalpy
T 2. Construct Born–Haber cycles to calculate lattice enthalpies using these enthalpy changes
T 3. Construct Born–Haber cycles to calculate one of the other enthalpy changes
T 4. Compare lattice enthalpies from Born–Haber cycles with those from calculations based on a perfect ionic model to provide evidence for covalent character in ionic compounds.
T 5. Define the term enthalpy of hydration
T 6. Perform calculations of an enthalpy change using these cycles.
T 7. Calculate entropy changes from absolute entropy values
T 8. Use the relationship $\Delta G = \Delta H - T\Delta S$ to determine how ΔG varies with temperature
T 9. Use the relationship $\Delta G = \Delta H - T\Delta S$ to determine the temperature at which a reaction becomes feasible.

Exam booklet reflection:

WWW

EBI

Checked by teacher

Doddle quiz homework:

Enthalpy change

Score:

Lattice enthalpy

Score:

Enthalpy and entropy

Score:

A LEVEL CHEMISTRY

TOPIC 10 –THERMODYNAMICS

TEST

Answer all questions

Max 50 marks

Name

Mark/50% Grade

1. Calcium fluoride occurs naturally as the mineral fluorite, a very hard crystalline solid that is almost insoluble in water and is used as a gemstone.

Tables 1 and 2 contain thermodynamic data.

Table 1

Process	$\Delta H^\circ / \text{kJ mol}^{-1}$
$\text{Ca(s)} \rightarrow \text{Ca(g)}$	+193
$\text{Ca(g)} \rightarrow \text{Ca}^+(\text{g}) + \text{e}^-$	+590
$\text{Ca}^+(\text{g}) \rightarrow \text{Ca}^{2+}(\text{g}) + \text{e}^-$	+1150
$\text{F}_2(\text{g}) \rightarrow 2\text{F(g)}$	+158
$\text{F(g)} + \text{e}^- \rightarrow \text{F}^-(\text{g})$	-348

Table 2

Name of enthalpy change	$\Delta H^\circ / \text{kJ mol}^{-1}$
Enthalpy of lattice dissociation for calcium fluoride	+2602
Enthalpy of lattice dissociation for calcium chloride	+2237
Enthalpy of hydration for F^- ions	-506
Enthalpy of hydration for Cl^- ions	-364
Enthalpy of hydration for Ca^{2+} ions	-1650

- (a) Write an equation, including state symbols, for the process that occurs when the calcium fluoride lattice dissociates and for which the enthalpy change is equal to the lattice enthalpy.

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(1)

- (b) (i) Define the term *standard enthalpy of formation*.

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(3)

- (ii) Write an equation, including state symbols, for the process that has an enthalpy change equal to the standard enthalpy of formation of calcium fluoride.

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(1)

- (iii) Use data from the **Tables 1** and **2** to calculate the standard enthalpy of formation for calcium fluoride.

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(3)

- (c) Explain why the enthalpy of lattice dissociation for calcium fluoride is greater than that for calcium chloride.

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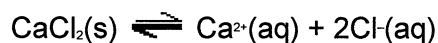
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(2)

- (d) Calcium chloride dissolves in water. After a certain amount has dissolved, a saturated solution is formed and the following equilibrium is established.



- (i) Using data from **Table 2**, calculate the enthalpy change for this reaction.

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(2)

- (ii) Predict whether raising the temperature will increase, decrease or have no effect on the amount of solid calcium chloride that can dissolve in a fixed mass of water.
Explain your prediction.
(If you have been unable to obtain an answer to part (d) (i), you may assume that the enthalpy change = -60 kJ mol^{-1} . This is **not** the correct answer.)

Effect on amount of solid that can dissolve

Explanation

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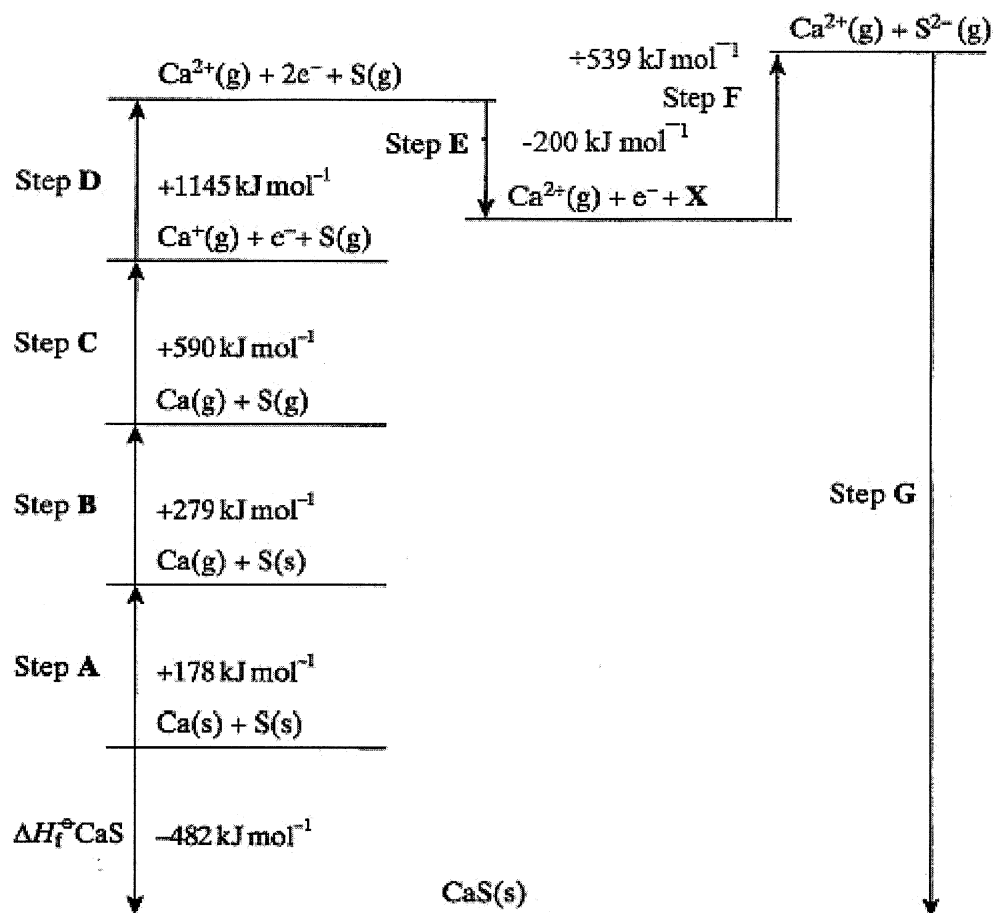
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(3)

(Total 15 marks)

2. (a) A Born–Haber cycle for the formation of calcium sulphide is shown below. The cycle includes enthalpy changes for all steps except step **G**. (The cycle is not drawn to scale.)



- (i) Give the full electronic configuration of the ion S^{2-} .

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- (ii) Suggest why step **F** is an endothermic process.

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- (iii) Name the enthalpy changes in steps **B** and **D**.

Step **B**

Step **D**

- (iv) Explain why the enthalpy change for step **D** is larger than that for step **C**.

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- (v) Use the data shown in the cycle to calculate a value for the enthalpy change for step **G**.

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(9)

- (b) Using a Born–Haber cycle, a value of -905 kJ mol^{-1} was determined for the lattice enthalpy of silver chloride. A value for the lattice enthalpy of silver chloride using the ionic model was -833 kJ mol^{-1} .

Explain what a scientist would be able to deduce from a comparison of these values.

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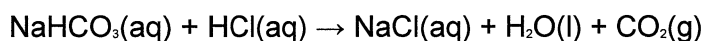
(3)

- (c) Some endothermic reactions occur spontaneously at room temperature. Some exothermic reactions do not occur if the reactants are heated together to a very high temperature.

In order to explain the following observations, another factor, the entropy change, ΔS , must be considered. The equation which relates ΔS to ΔH is given below.

$$\Delta G = \Delta H - T\Delta S$$

- (i) Explain why the following reaction occurs at room temperature even though the reaction is endothermic.



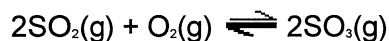
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- (ii) Explain why the following reaction does not occur at very high temperatures even though the reaction is exothermic.



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(6)
(Total 18 marks)

3. Methanol can be regarded as a carbon-neutral fuel because it can be synthesised from carbon dioxide as shown in the equation below.



Standard enthalpy of formation and standard entropy data for the starting materials and products are shown in the following table.

	$\text{CO}_2(\text{g})$	$\text{H}_2(\text{g})$	$\text{CH}_3\text{OH}(\text{g})$	$\text{H}_2\text{O}(\text{g})$
$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	-394	0	-201	-242
$S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$	214	131	238	189

- (a) Calculate the standard enthalpy change for this reaction.

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(3)

- (b) Calculate the standard entropy change for this reaction.

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(3)

- (c) Use your answers to parts (a) and (b) to explain why this reaction is **not** feasible at high temperatures.

Calculate the temperature at which the reaction becomes feasible.

Suggest why the industrial process is carried out at a higher temperature than you have calculated.

(If you have been unable to calculate values for ΔH and ΔS you may assume that they are -61 kJ mol^{-1} and $-205 \text{ J K}^{-1} \text{ mol}^{-1}$ respectively. These are **not** the correct values.)

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(6)

- (d) Write an equation for the complete combustion of methanol. Use your equation to explain why the combustion reaction in the gas phase is feasible at all temperatures.

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(4)

(Total 16 marks)

4. Which one of the following statements is **not** correct?
- A The first ionisation energy of iron is greater than its second ionisation energy.
 - B The magnitude of the lattice enthalpy of magnesium oxide is greater than that of barium oxide.
 - C The oxidation state of iron in $[\text{Fe}(\text{CN})_6]^{3-}$ is greater than the oxidation state of copper in $[\text{CuCl}_2]^-$.
 - D The boiling point of C_3H_8 is lower than that of $\text{CH}_3\text{CH}_2\text{OH}$

(Total 1 mark)

Electrochemistry

Electrode potentials and electrochemical cells
T 1. Use E^\ominus values to predict the direction of simple redox reactions
T 2. Calculate the EMF of a cell
T 3. Write and apply the conventional representation of a cell.
T 4. Use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells
T 5. Deduce the EMF of a cell
T 6. Explain how the electrode reactions can be used to generate an electric current.

Exam booklet reflection:

WWW

EBI

Doddle quiz homework:

Electrochemical cells

Score:

Electrode potentials

Score:

Redox titrations

Score:

Storage and fuel cells

Score:

A LEVEL CHEMISTRY

TOPIC 13 – ELECTROCHEMISTRY

TEST

Answer all questions

Max 50 marks

Name		
Mark/50%	Grade

1. The table contains some standard electrode potential data.

Electrode half-equation	E^\ominus / V
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Au^+ + e^- \longrightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Ag^+ + e^- \longrightarrow Ag$	+0.80
$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$Fe^{2+} + 2e^- \longrightarrow Fe$	-0.44

- (a) In terms of electrons, explain the meaning of the term **oxidising agent**.

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(1)

- (b) Identify the weakest oxidising agent in the table.
 Explain your choice.

Weakest oxidising agent

Explanation

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(2)

- (c) Write the conventional representation of the cell used to measure the standard electrode potential for the Ag^+ / Ag electrode.

State the conditions necessary when measuring this value.

Conventional representation

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Conditions

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(4)

- (d) Use data from the table to explain, in terms of redox, what happens when a soluble gold(I) compound containing Au^+ ions is added to water.

State what you would observe.

Write an equation for the reaction that occurs.

Explanation

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Observation

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Equation

(4)

- (e) A cell is made by connecting $\text{Fe}^{2+} / \text{Fe}$ and Ag^+ / Ag electrodes with a salt bridge.

- (i) Calculate the e.m.f. of this cell.

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Answer

(1)

- (ii) Suggest why potassium chloride would **not** be suitable for use in the salt bridge of this cell.

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(1)

- (f) Use data from the table to explain what happens when a solution of iron(II) chloride is exposed to the air.

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(2)
(Total 15 marks)

2. Use the data below, where appropriate, to answer the questions which follow.

Standard electrode potentials	E^\ominus/V
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(aq) + 6H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions.

- (a) State the hydrogen ion concentration and the hydrogen gas pressure when, at 298 K, the potential of the hydrogen electrode is 0.00 V.

Hydrogen ion concentration

Hydrogen gas pressure

(2)

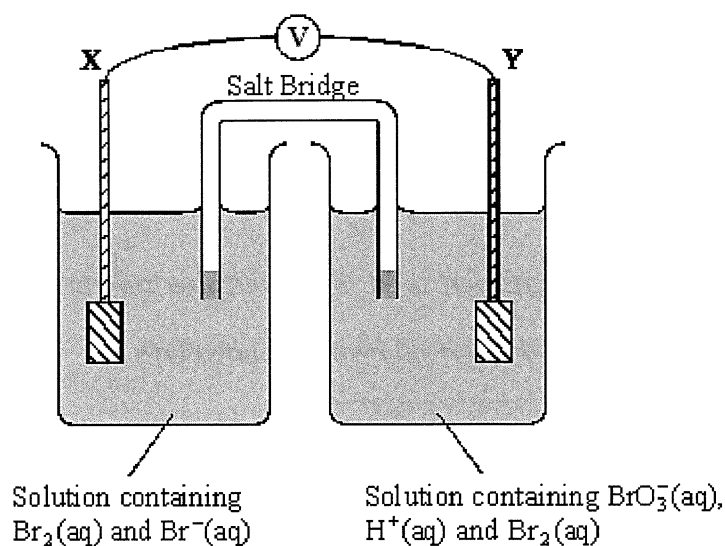
- (b) The electrode potential of a hydrogen electrode changes when the hydrogen ion concentration is reduced. Explain, using Le Chatelier's principle, why this change occurs and state how the electrode potential of the hydrogen electrode changes.

Explanation of change

Change in electrode potential

(3)

- (c) A diagram of a cell using platinum electrodes **X** and **Y** is shown below.



- (i) Use the data above to calculate the e.m.f. of the above cell under standard conditions.

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- (ii) Write a half-equation for the reaction occurring at electrode **X** and an overall equation for the cell reaction which occurs when electrodes **X** and **Y** are connected.

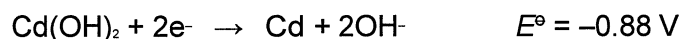
Half-equation

Overall equation

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(4)
(Total 9 marks)

3. Nickel–cadmium cells are used to power electrical equipment such as drills and shavers. The electrode reactions are shown below.



- (a) Calculate the e.m.f. of a nickel–cadmium cell.

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(1)

- (b) Deduce an overall equation for the reaction that occurs in the cell when it is used.

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(2)

- (c) Identify the oxidising agent in the overall cell reaction and give the oxidation state of the metal in this oxidising agent.

Oxidising agent

Oxidation state

(2)
(Total 5 marks)

4. Redox reactions occur in the discharge of all electrochemical cells. Some of these cells are of commercial value.

The table below shows some redox half-equations and standard electrode potentials.

Half-equation	E^\ominus / V
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76
$\text{Ag}_2\text{O}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Ag}(\text{s}) + \text{H}_2\text{O}(\text{l})$	+0.34
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	+1.23
$\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$	+2.87

- (a) In terms of electrons, state what happens to a reducing agent in a redox reaction.

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(1)

- (b) Use the table above to identify the strongest reducing agent from the species in the table.

Explain how you deduced your answer.

Strongest reducing agent

Explanation

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(2)

- (c) Use data from the table to explain why fluorine reacts with water.

Write an equation for the reaction that occurs.

Explanation

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Equation

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(3)

- (d) An electrochemical cell can be constructed using a zinc electrode and an electrode in which silver is in contact with silver oxide. This cell can be used to power electronic devices.

(i) Give the conventional representation for this cell.

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(2)

(ii) Calculate the e.m.f. of the cell.

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(1)

(iii) Suggest **one** reason why the cell cannot be electrically recharged.

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(1)

- (e) The electrode half-equations in a lead–acid cell are shown in the table below.

Half-equation	E° / V
$\text{PbO}_2(\text{s}) + 3\text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq}) + 2\text{e}^- \longrightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.69
$\text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Pb}(\text{s}) + \text{HSO}_4^-(\text{aq})$	to be calculated

- (i) The $\text{PbO}_2/\text{PbSO}_4$ electrode is the positive terminal of the cell and the e.m.f. of the cell is 2.15 V.

Use this information to calculate the missing electrode potential for the half-equation shown in the table.

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(1)

- (ii) A lead–acid cell can be recharged.
Write an equation for the overall reaction that occurs when the cell is being recharged.

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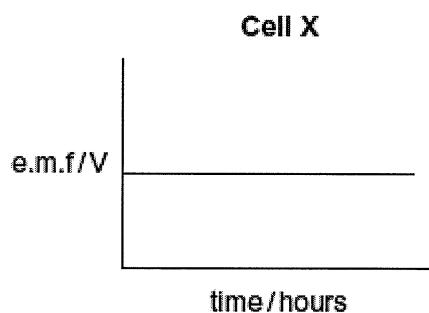
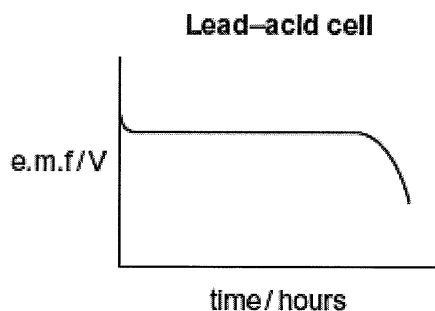
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(2)

- (f) The diagrams below show how the e.m.f. of each of two cells changes with time when each cell is used to provide an electric current.



- (i) Give **one** reason why the e.m.f. of the **lead–acid cell** changes after several hours.

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(1)

- (ii) Identify the type of cell that behaves like **cell X**.

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(1)

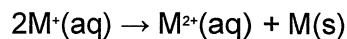
- (iii) Explain why the voltage remains constant in **cell X**.

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(2)

(Total 17 marks)

5. A disproportionation reaction occurs when a species M^+ spontaneously undergoes simultaneous oxidation and reduction.



The table below contains E^\ominus data for copper and mercury species.

	E^\ominus / V
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	+ 0.15
$Cu^+(aq) + e^- \rightarrow Cu(s)$	+ 0.52
$Hg^{2+}(aq) + e^- \rightarrow Hg^+(aq)$	+ 0.91
$Hg^+(aq) + e^- \rightarrow Hg(l)$	+ 0.80

Using these data, which one of the following can be predicted?

- A Both Cu(I) and Hg(I) undergo disproportionation.
 B Only Cu(I) undergoes disproportionation.
 C Only Hg(I) undergoes disproportionation.
 D Neither Cu(I) nor Hg(I) undergoes disproportionation.

(Total 1 mark)

6. Use the data in the table below to answer this question.

	E^\ominus / V
$MnO_4^- (aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.52
$Cr_2O_7^{2-} (aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+ 1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+ 0.77
$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	- 0.41
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	- 0.76

The most powerful oxidising agent in the table is

- A $Mn^{2+}(aq)$
 B $Zn(s)$
 C $MnO_4^-(aq)$
 D $Zn^{2+}(aq)$

(Total 1 mark)

7. In this question consider the data below.

	E^{\ominus} / V
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13

The e.m.f. of the cell $\text{Pt}(\text{s}) \mid \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}) \parallel \text{Ag}^+(\text{aq}) \mid \text{Ag}(\text{s})$ would be increased by

- A increasing the concentration of $\text{H}^+(\text{aq})$.
- B increasing the surface area of the Pt electrode.
- C increasing the concentration of $\text{Ag}^+(\text{aq})$.
- D decreasing the pressure of $\text{H}_2(\text{g})$.

(Total 1 mark)

8. In this question consider the data below.

	E^{\ominus} / V
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13

The e.m.f. of the cell $\text{Ag}(\text{s}) \mid \text{Ag}^+(\text{aq}) \parallel \text{Pb}^{2+}(\text{aq}) \mid \text{Pb}(\text{s})$ is

- A 0.93 V
- B 0.67 V
- C -0.67 V
- D -0.93 V

(Total 1 mark)

Transition metals and P3 oxides

Properties of period 3 elements and their oxides

T 1. Explain the trend in the melting point of the oxides of the elements Na–S in terms of their structure and bonding

T 2. Explain the trends in the reactions of the oxides with water in terms of the type of bonding present in each oxide

T 3. Write equations for the reactions that occur between the oxides of the elements Na–S and given acids and bases.

Transition metals

T 1. Transition metal characteristics of elements Ti–Cu arise from an incomplete d sub-level in atoms or ions.

T 2. The characteristic properties include: complex formation, formation of coloured ions, variable oxidation state and catalytic activity.

T 3. A ligand is a molecule or ion that forms a co-ordinate bond with a transition metal by donating a pair of electrons.

T 4. A complex is a central metal atom or ion surrounded by ligands.

T 5. Co-ordination number is number of co-ordinate bonds to the central metal atom or ion.

T 6. H₂O, NH₃ and Cl[–] can act as monodentate ligands. The ligands NH₃ and H₂O are similar in size and are uncharged.

T 7. Exchange of the ligands NH₃ and H₂O occurs without change of co-ordination number (eg Co²⁺ and Cu²⁺).

T 8. Substitution may be incomplete (eg the formation of [Cu(NH₃)₄(H₂O)₂]²⁺).

T 9. The Cl[–] ligand is larger than the uncharged ligands NH₃ and H₂O

T 10. Exchange of the ligand H₂O by Cl[–] can involve a change of co-ordination number (eg Co²⁺, Cu²⁺ and Fe³⁺).

T 11. Ligands can be bidentate (eg H₂NCH₂CH₂NH₂ and C₂O₄^{2–}). Ligands can be multidentate (eg EDTA^{4–}).

T 12. Oxygen forms a co-ordinate bond to Fe(II) in haemoglobin, enabling oxygen to be transported in the blood. Carbon monoxide is toxic because it replaces oxygen co-ordinately bonded to Fe(II) in haemoglobin.

T 13. Bidentate and multidentate ligands replace monodentate ligands from complexes. This is called the chelate effect.

T 14. Explain the chelate effect, in terms of the balance between the entropy and enthalpy change in these reactions

T 15. Octahedral complexes can display cis–trans isomerism (a special case of E–Z isomerism) with monodentate ligands and optical isomerism with bidentate ligands.

T 16. Transition metal ions commonly form tetrahedral complexes with larger ligands (eg Cl[–]).

T 17. Square planar complexes are also formed and can display cis–trans isomerism.

T 18. The energy difference between the ground state and the excited state of the d electrons is given by $\Delta E = h\nu = hc/\lambda$. This makes transition metal ions coloured

T 19. Changes in oxidation state, co-ordination number and ligand alter ΔE and this leads to a change in colour.

T 20. A simple colorimeter can be used to determine the concentration of coloured ions in solution.

T 21. Vanadium species in oxidation states IV, III and II are formed by the reduction of vanadate(V) ions by zinc in acidic solution.

T 22. The redox potential for a transition metal ion changing from a higher to a lower oxidation state is influenced by pH and by the ligand.

T 23. Perform calculations for redox titrations

Catalysts

T 1. Explain the importance of variable oxidation states in catalysis

T 2. Explain, with the aid of equations, how V₂O₅ acts as a catalyst in the Contact process

T 3. explain, with the aid of equations, how Fe²⁺ ions catalyse the reaction between I[–] and S₂O₈^{2–}

T 4. Explain, with the aid of equations, how Mn²⁺ ions autocatalyse the reaction between C₂O₄^{2–} and MnO₄[–]

Reactions in aqueous solution

T 1. Explain, in terms of the charge/size ratio of the metal ion, why the acidity of [M(H₂O)₆]³⁺ is greater than that of [M(H₂O)₆]²⁺

T 2. Describe and explain the simple test-tube reactions of: M²⁺(aq) ions, limited to M = Fe and Cu, and of M³⁺(aq) ions, limited to M = Al and Fe, with the bases OH[–], NH₃ and CO₃^{2–}

Transition metals and P3 oxides

Exam booklet reflection:

WWW

EBI

Checked by teacher

Doddle quiz homework:

TM Complexes

Score:

TM Properties

Score:

TM Elements

Score:

Catalysis

Score:

Ligand substitution

Score:

Metal Aqua Ions

Score:

A LEVEL CHEMISTRY

TOPIC 15 – TRANSITION METALS AND COMPLEX IONS

TEST

Answer all questions

Max 50 marks

Name		
Mark/50%	Grade

1. Transition elements form complex ions with a range of colours and shapes.

- (a) By considering its electron arrangement, state how an element can be classified as a transition element.

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(1)

- (b) Explain the meaning of the term *complex ion*.

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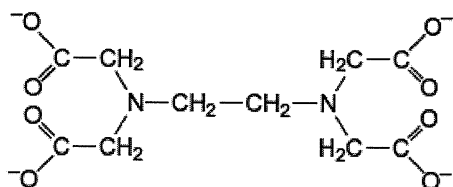
(2)

- (c) In terms of electrons, explain why an aqueous solution of cobalt(II) sulfate has a red colour.

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(3)

- (d) The ligand EDTA^{4-} is shown below.



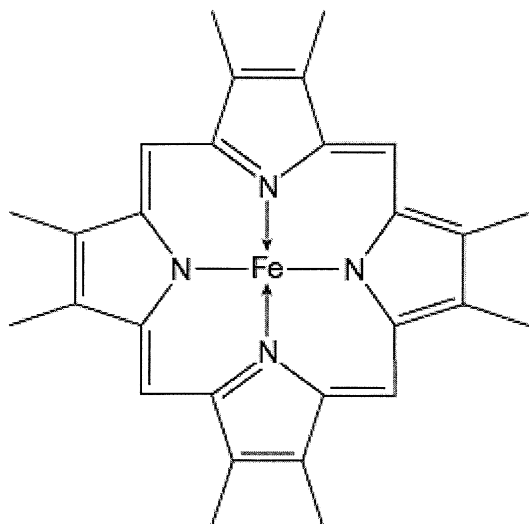
- (i) Draw circles around the atoms of **two** different elements that link to a transition metal ion by a co-ordinate bond when EDTA^{4-} behaves as a ligand.
- (ii) Write an equation for the reaction between EDTA^{4-} and a $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ ion. Use the abbreviation EDTA^{4-} in your equation.
- (iii) Explain why the complex ion, formed as a product of the reaction in part (d) (ii), is more stable than the $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ ion.

(2)

(1)

(2)

- (e) The diagram below shows part of the structure of haemoglobin.



Haemoglobin contains an iron(II) ion bonded to five nitrogen atoms and one other ligand. The fifth nitrogen atom and the additional ligand are not shown in this diagram.

- (i) In this diagram, bonds between nitrogen and iron are shown as $N \rightarrow Fe$ and as $N - Fe$.

State the meaning of each of these symbols.

Meaning of \rightarrow

Meaning of $-$

(2)

- (ii) State the function of haemoglobin in the blood.

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(1)

- (iii) With reference to haemoglobin, explain why carbon monoxide is toxic.

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(2)

(Total 16 marks)

1. (a) Explain the meaning of the terms *ligand* and *bidentate* as applied to transition metal complexes.

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(2)

- (b) Aqueous cobalt(II) ions react separately with an excess of chloride ions and with an excess of ammonia.

For each reaction, draw a diagram to illustrate the structure of, the shape of and the charge on the complex ion formed.

In each case, name the shape and indicate, on the diagram, a value for the ligand-metal-ligand bond angle.

(6)

- (c) The complex ion formed in aqueous solution between cobalt(II) ions and chloride ions is a different colour from the $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ ion.

Explain why these complex ions have different colours.

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(3)

- (d) In aqueous ammonia, cobalt(II) ions are oxidised to cobalt(III) ions by hydrogen peroxide. The H_2O_2 is reduced to hydroxide ions.

Calculate the minimum volume of $5.00 \text{ mol dm}^{-3} \text{ H}_2\text{O}_2$ solution required to oxidise the Co^{2+} ions in 9.87 g of $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$

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(5)
(Total 16 marks)

3. When iodine molecules are dissolved in aqueous solutions containing iodide ions, they react to form triiodide ions (I_3^-).



The reaction above between I^- ions and $\text{S}_2\text{O}_8^{2-}$ ions has a high activation energy and $\text{S}_2\text{O}_8^{2-}$ ions are only reduced slowly to SO_4^{2-} ions.
The reaction is catalysed by Fe^{2+} ions.

- (a) Explain why the reaction between I^- ions and $\text{S}_2\text{O}_8^{2-}$ ions is slow.

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(1)

- (b) Other than having variable oxidation states, explain why Fe^{2+} ions are good catalysts for this reaction.

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(1)

- (c) Write a half-equation for the reduction of $\text{S}_2\text{O}_8^{2-}$ ions to SO_4^{2-} ions.

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(1)

- (d) Construct an overall equation for the reaction between $\text{S}_2\text{O}_8^{2-}$ ions and I^- ions.

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(1)

(Total 4 marks)

4. A green solution, **X**, is thought to contain $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ions.

- (a) The presence of these ions can be confirmed by reacting separate samples of solution **X** with aqueous ammonia and with aqueous sodium carbonate.

Write equations for each of these reactions and describe what you would observe.

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(4)

- (b) A 50.0 cm³ sample of solution **X** was added to 50 cm³ of dilute sulfuric acid and made up to 250 cm³ of solution in a volumetric flask.

A 25.0 cm³ sample of this solution from the volumetric flask was titrated with a 0.0205 mol dm⁻³ solution of KMnO₄

At the end point of the reaction, the volume of KMnO₄ solution added was 18.70 cm³.

- (i) State the colour change that occurs at the end point of this titration and give a reason for the colour change.

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(2)

- (ii) Write an equation for the reaction between iron(II) ions and manganate(VII) ions.

Use this equation and the information given to calculate the concentration of iron(II) ions in the original solution **X**.

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(5)

(Total 11 marks)

5. The percentage of iron in a sample of impure iron(II) sulphate crystals can be determined by titrating solutions, made from separate weighed samples acidified with dilute sulphuric acid, against a standard solution of potassium manganate(VII).

Which one of the following would lead to an inaccurate result?

- A transferring the weighed sample of iron(II) sulphate into a wet conical flask
- B failing to measure accurately the volume of water used to dissolve each weighed sample of iron(II) sulphate
- C transferring the standard solution of potassium manganate(VII) from its original container to the burette using a wet beaker
- D failing to measure accurately the volume of dilute sulphuric acid added to the mixture before titration

(Total 1 mark)

6. Which one of the following contains the metal with the lowest oxidation state?

- A CrO_2F_2
- B $[\text{Cr}_2\text{O}_7]^{2-}$
- C $[\text{MnCl}_6]^{2-}$
- D $[\text{Mn}(\text{CN})_6]^{3-}$

(Total 1 mark)

7. Which one of the following could **not** act as a ligand?

- A F^-
- B CH_3CH_3
- C NH_2NH_2
- D CH_3OCH_3

(Total 1 mark)

A LEVEL CHEMISTRY

TOPIC 14 – REACTIONS OF PERIOD 3 ELEMENTS AND THEIR OXIDES

TEST

Answer all questions

Max 50 marks

Name
Mark	<div style="display: flex; justify-content: space-between; align-items: center;">/50% Grade </div>

1. (a) Write an equation for the reaction that occurs when magnesium is heated in steam. Describe what you would observe when this reaction occurs.

Equation

Observations

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(3)

- (b) Write an equation for the reaction that occurs when sodium is heated in oxygen. Describe what you would observe when this reaction occurs.

Equation

Observations

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(3)

(Total 6 marks)

2. Some melting points of Period 3 oxides are given in this table.

	Na ₂ O	SiO ₂	SO ₂	SO ₃
Melting point / K	1548	1883	200	290

- (a) Explain, in terms of structure and bonding, why sodium oxide has a high melting point.

.....

(2)

- (b) Explain, in terms of structure and bonding, why sulfur trioxide has a higher melting point than sulfur dioxide.

.....

(2)

- (c) Some Period 3 oxides have basic properties.

State the type of bonding in these basic oxides.

Explain why this type of bonding causes these oxides to have basic properties.

Type of bonding

Explanation

.....

(3)

- (d) Sulfur dioxide reacts with water to form a weakly acidic solution.

- (i) Ions are formed when sulfur dioxide reacts with water.

Write an equation for this reaction.

.....

(1)

- (ii) With reference to your equation from part (d)(i), suggest why sulfur dioxide forms a weakly acidic solution.

.....

(1)

- (e) Suggest why silicon dioxide is described as an acidic oxide even though it is insoluble in water.

.....

(1)

(Total 10 marks)

3. (a) The melting points of some of the oxides formed by Period 3 elements are given in a random order below.

Oxide	A	B	C	D	E
$T_m/^\circ\text{C}$	2852	-73	1610	1275	300

- (i) Using the letters **A** to **E**, give **two** oxides which have simple molecular structures.

Explain your answer.

Oxide 1

Oxide 2

Explanation

.....

- (ii) Give a simple chemical test which could be used to show which of the oxides in the table is sodium oxide. State the observation you would make.

Chemical test

.....

Observation

(6)

- (b) The base calcium oxide can be used to remove sulfur dioxide from flue-gases produced when fossil fuels are burnt in coal-fired power stations. Calcium oxide is produced when calcium carbonate, is decomposed by heat.

- (i) Write an equation for the action of heat on calcium carbonate.

.....

- (ii) Identify the product formed when sulfur dioxide reacts with calcium oxide.

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- (iii) Despite the additional cost, operators of power stations are encouraged to remove the sulfur dioxide from flue-gases. Explain why this may not be environmentally beneficial.

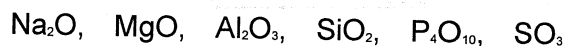
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(4)

(Total 10 marks)

4. Consider the following oxides.



(a) Identify one of the oxides from the above which

(i) can form a solution with a pH less than 3

(ii) can form a solution with a pH greater than 12

(2)

(b) Write an equation for the reaction between

(i) MgO and HNO_3

.....

(ii) SiO_2 and NaOH

.....

(iii) Na_2O and H_3PO_4

.....

(3)

(c) Explain, in terms of their type of structure and bonding, why P_4O_{10} can be vaporised by gentle heat but SiO_2 cannot.

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(4)

(Total 9 marks)

5. (a) State and explain the trend in electronegativities across Period 3 from sodium to sulfur.

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(4)

- (b) Explain why the oxides of the Period 3 elements sodium and phosphorus have different melting points. In your answer you should discuss the structure of and bonding in these oxides, and the link between electronegativity and the type of bonding.

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(6)

- (c) A chemical company has a waste tank of volume 25 000 dm³. The tank is full of phosphoric acid (H₃PO₄) solution formed by adding some unwanted phosphorus(V) oxide to water in the tank.

A 25.0 cm³ sample of this solution required 21.2 cm³ of 0.500 mol dm⁻³ sodium hydroxide solution for complete reaction.

Calculate the mass, in kg, of phosphorus(V) oxide that must have been added to the water in the waste tank.

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(5)

(Total 15 marks)