A Level Chemistry Summer Independent Learning

Y12 - Y13 2025

- Key Topics Consolidation & Review
    - Amount of Substance
    - Energetics
    - Bonding
    - Equilibria
    - Redox
    - Intro to Organic
- Required Practicals (1-6)
- Graph Practice
- Past Paper Question and Exam Practice
- New Content: Kp

In September you will be tested on certain areas with written tasks and Socrative for other topics.

You should complete this work in full and bring it to your first Chemistry lesson of Year 13

→ Key Topics Consolidation & Review

For each of the key topics you need to create a Knowledge Organiser using the guidance below and learn the 10 retrieval questions for each topic. You will be tested on a random selection of these as part of the Initial Assessment in Y13.

- **Amount of Substance**
- **Energetics**
- **Bonding**
- **Equilibria**
- **Redox**
- **Intro to Organic**

Instructions for creating a Knowledge Organiser

Step-by-Step Guide to Creating Your Knowledge Organiser

- ***Knowledge only; ensure the information is condensed***
- ***Organised Layout with clear sections***
- ***Include diagram, key terms and formula***

Retrieval Question Answers

<table>
<tr><th colspan="2">Title : Amount of Substance – AQA Chemistry (3.1.2)</th></tr>
<tr><td><i>Key Definitions:</i> <i>Include definitions of the following:</i>

• Mole
• Avogadro constant
• Relative atomic mass (Ar)
• Relative molecular mass (Mr)
• Empirical formula
• Molecular formula
• Concentration (mol/dm^3)
• Percentage yield
• Atom economy</td><td><i>Equations to Memorise and Use:</i> <i>Create a section titled</i> <i>"Key Equations"</i> <i>with these formulas:</i>

$n = m / M$ → Moles = mass / molar mass

$n = C \times V / 1000$ → Moles = concentration × volume (in cm^3)

$pV = nRT$ → Ideal gas equation

Empirical formula → Divide % or mass by Ar, simplify ratio

% Yield → (Actual / Theoretical) × 100

Atom Economy → (Mr of useful product / total Mr of products) × 100</td></tr>
<tr><td colspan="2"><i>Tips and Common Mistakes</i>
Make a list of 5–7 tips such as:

• Always convert cm^3 to dm^3 (÷1000)
• Check units in ideal gas equation
• Use significant figures correctly
• Balance equations carefully</td></tr>
</table>

Amount of Substance Retrieval Practice:

1. What is a mole?
2. Define the Avogadro constant.
3. What are the units for pressure, volume, and temperature in the ideal gas equation?
4. Rearrange $pV=nRT$ to calculate temperature.
5. Calculate the number of moles in 20 g of NaOH (Mr = 40).
6. Convert 250 cm^3 into dm^3.
7. A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen. Find the empirical formula.
8. Define atom economy and state its importance.
9. What is the concentration of a solution that contains 0.25 mol in 500 cm^3?
10. What is the purpose of performing a titration?

Title: *Bonding – AQA Chemistry (3.1.3)*

Key Definitions

Create a list of key definitions, such as:

- Ionic bond
- Covalent bond
- Dative covalent bond
- Metallic bond
- Electronegativity
- Polar molecule
- Hydrogen bond
- Van der Waals forces

Types of Bonding

Create comparison tables or diagrams for:

- Ionic bonding (e.g., NaCl)
- Covalent bonding (e.g., Cl_2, CH_4, NH_3)
- Metallic bonding (e.g., Mg, Al)

Explain:

- Particle types involved
- Electron behaviour
- Properties (melting point, conductivity, solubility)

Intermolecular Forces

Explain:

- Induced dipole–dipole interactions (van der Waals)
- Permanent dipole–dipole interactions
- Hydrogen bonding (e.g., H_2O, NH_3)

Include a strength comparison and how these affect boiling/melting points.

Shapes of Molecules and Bond Angles

Include a VSEPR table showing:

- Shape names (linear, trigonal planar, tetrahedral, etc.)
- Number of bonding and lone pairs
- Example molecules
- Bond angles

Structure and Physical Properties

Link structure type (giant ionic, giant covalent, simple molecular, metallic) to:

- Melting/boiling point
- Electrical conductivity
- Solubility
- Examples (e.g., NaCl, diamond, graphite, I_2)

Diagrams

Draw:

- Energy profile diagrams for exothermic and endothermic reactions
- A labelled calorimeter setup
- A Hess's Law enthalpy cycle

Bond Polarity and Molecular Polarity

Include:

- Electronegativity definition
- How bond and molecular polarity ariseh
- Examples of polar and non-polar molecules

Retrieval Practice:

1. What is an ionic bond?
2. Define a dative covalent bond.
3. What causes a molecule to be polar?
4. Why does graphite conduct electricity but diamond does not?
5. Draw the shape and give the bond angle of methane (CH_4).
6. Which type of bonding exists in sodium metal?
7. Name three types of intermolecular forces.
8. Why does water have a higher boiling point than hydrogen sulfide (H_2S)?
9. What is the shape and bond angle of ammonia (NH_3)?
10. What is the difference between simple molecular and giant covalent structures?

Title: *Energetics – AQA Chemistry (3.1.4)*

Key Definitions Section

Include:

- Enthalpy change (ΔH)
- Exothermic and endothermic reactions
- Standard enthalpy of combustion (ΔH_c°)
- Standard enthalpy of formation (ΔH_f°)
- Activation energy

Essential Equations

List and annotate:

- $q=mc\Delta T$
- $\Delta H=q/n$
- Hess's Law: construct a Hess's law cycle.
- Mean bond enthalpy: A bond enthalpy calculation

Common Mistakes

Bullet point 5–6 such as:

- Forgetting to convert mass or volume units
- Using ΔH instead of q for calculations
- Not using the correct number of significant figures

Retrieval Questions

1. Define enthalpy change.
2. What is the sign of ΔH for an exothermic reaction?
3. State the equation used to calculate heat change in a calorimeter.
4. In the equation $q=mc\Delta T$ what does each symbol mean?
5. What is the specific heat capacity of water?
6. Calculate q if 50 g of water increases by 10°C (c = 4.18 J/g·K).
7. State Hess's Law
8. Why are bond enthalpy calculations only approximate?
9. How is enthalpy of formation (ΔH_f°) defined?
10. Why do we use a polystyrene cup in calorimetry?

Title: *Equilibria – AQA Chemistry (3.1.6)*

Key Definitions

Create a section of core terms:

- Reversible reaction
- Dynamic equilibrium
- Le Chatelier’s Principle
- Equilibrium constant (Kc)
- Homogeneous system
- Catalyst

Le Chatelier’s Principle

Explain:

- The effect of changing:
    - Concentration
    - Pressure (for gases)
    - Temperature
    - Use of a catalyst
- Use diagrams where appropriate.

Equilibrium Constant (Kc)

Include:

- General form:

  Kc=[products]/[reactants]

- Example calculation using concentrations in mol/dm^3
- Explanation of how temperature affects Kc (but not concentration or catalyst)

Worked Examples

Show one example of:

- Predicting the shift in equilibrium using Le Chatelier’s Principle
- Calculating Kc from given concentrations
- Using ICE (Initial-Change-Equilibrium) tables

Common Misconceptions

List key misunderstandings to avoid:

- Confusing “rate” with “position” of equilibrium
- Believing a catalyst shifts the position of equilibrium (it doesn’t)
- Thinking that changing concentration affects the value of Kc (it doesn’t)

Retrieval Practice:

1. What is a dynamic equilibrium?
2. State Le Chatelier’s Principle.
3. What happens to the equilibrium position if temperature is increased in an exothermic reaction?
4. What effect does a catalyst have on equilibrium position?
5. Write the expression for Kc for the reaction: $2A + B \rightarrow C + D$

6. If the concentration of a reactant is increased, what happens to the position of equilibrium?
7. What are the units of Kc for the reaction: $A + B \rightarrow C$

8. Does changing pressure affect all equilibria?
9. Why doesn't adding a catalyst affect the value of Kc?
10. A reaction has Kc = 1. What does this indicate?

<table>
<tr><th colspan="2">Title: <i>Redox – AQA Chemistry (3.1.7)</i></th></tr>
<tr><td>Key Definitions
Create a table or list with:
• Oxidation
• Reduction
• Oxidising agent
• Reducing agent
• Oxidation state
• Half-equation
• Redox equation</td><td>Rules for Assigning Oxidation States
Include rules such as:
• Uncombined elements = 0
• Oxygen = –2 (except peroxides or when bonded to fluorine)
• Hydrogen = +1 (except in metal hydrides)
• Sum of oxidation states = charge on species</td></tr>
<tr><td>Worked Examples
Provide one or more examples of:
• Assigning oxidation numbers in compounds and ions
• Writing half-equations (oxidation/reduction)
• Combining half-equations into a balanced redox equation</td><td>Common Mistakes
List mistakes to avoid:
• Incorrect signs on oxidation states
• Forgetting to balance charge and atoms in half-equations
• Confusing oxidising and reducing agents</td></tr>
</table>

Retrieval Practice –

1. Define oxidation and reduction in terms of electrons.
2. What is an oxidising agent?
3. Assign the oxidation state of sulfur in H_2SO_4.
4. What is the oxidation number of oxygen in H_2O_2?
5. Write the half-equation for the reduction of Cl_2 to Cl^-.
6. Write the half-equation for the oxidation of Fe^{2+} to Fe^{3+}.
7. Combine the following half-equations into a redox equation:
   $Fe^{2+} \rightarrow Fe^{3+} + e^-$ and $Cl_2 + 2e^- \rightarrow 2Cl^-$
8. What is the total oxidation state of a sulfate ion (SO_4^{2-})?
9. Identify what is reduced and what is oxidised in the reaction:
   $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
10. State a rule for assigning oxidation numbers to elements in compounds.

Title: ***Introduction to Organic Chemistry – AQA Chemistry (3.3.1)***	
Key Terms and Definitions Include these: • Homologous series • Functional group • Structural formula • Displayed formula • Skeletal formula • Molecular formula • Empirical formula • Isomer • Reaction mechanism Use a table format to explain each with examples.	Key Terms and Definitions Include these: • Homologous series • Functional group • Structural formula • Displayed formula • Skeletal formula • Molecular formula • Empirical formula • Isomer • Reaction mechanism Use a table format to explain each with examples.
Naming Organic Compounds • Outline IUPAC rules: prefixes, roots, and suffixes. • Provide naming steps for alkanes, alkenes, alcohols, haloalkanes. • Include examples: ○ CH_4 methane ○ $CH_3CH_2CH_3$ propane ○ $CH_2{=}CH_2$ ethene	Types of Formulae • Draw and label examples for: ○ Molecular ○ Empirical ○ Structural ○ Displayed ○ Skeletal
Types of Organic Reactions Define: • Addition • Substitution • Elimination Include brief examples and diagrams if possible.	General Formulae of Homologous Series List common general formulae: • Alkanes: C_nH_{2n+2} • Alkenes: C_nH_{2n} • Alcohols: $C_nH_{2n+1}OH$
Types of Isomerism Explain: • **Structural isomerism** (chain, positional, functional group) • **Stereoisomerism** (brief intro for future topics) Provide examples like: • Butane vs methylpropane (chain) • Butan-1-ol vs butan-2-ol (positional) • Ethanol vs methoxymethane (functional group)	

Retrieval Practice:

1. What is a homologous series?
2. What is the molecular formula of ethane?
3. Give an example of a structural isomer of butane.
4. What is the displayed formula of ethanol?
5. Name this compound: CH_3CH_2OH
6. What is the difference between empirical and molecular formula?
7. What type of reaction involves two molecules combining to form one product?
8. Give the general formula of alkanes.
9. What is a functional group?
10. Define structural isomerism.

→ Required Practicals (1-6)

Information

Paper 1 (7405_1)** and **Paper 2 (7405_2)

You will be assessed on "relevant practical skills" in the form of short and long answer questions.

Paper 3 (7405_3)

*There will be 40 marks of questions on **any** practical techniques and data analysis.*

At least 15% if the overall assessment of A-level Chemistry will assess knowledge, skills and understanding in relation to practical work.

Tasks

Complete a summary for **each** practical.

The resources in this folder on Team A Level Chemistry Class of 2026 (Practical Endorsement) may help you.

Practical 1 Make up a volumetric solution and carry out a simple acid-base titration (3.1.2 Amount of substance)

- List all the equipment required to prepare 250 cm^3 of a standard solution.
- Describe how to perform this task, step by step. Consider the steps required to ensure that your concentration is accurate.
- List all the equipment required to perform an acid-base titration.
- Describe how to perform the titration, step by step and make it clear what steps need to be taken to obtain concordant results.
- What can you use an acid-base titration to determine?

Questions

1. Why are the burette and pipette rinsed with the reagent they are to contain, rather than water, before they are used?
2. Why should the funnel be removed from the burette before it is used?
3. What happens to the value of your titre if you do **not** remove any air beneath the tap in the burette before you start a titration?
4. Why is a conical flask used when performing a titration?
5. Why does rinsing the conical flask with distilled water during the titration not affect the titre value?
6. Each reading on the burette has an uncertainty of ±0.05 cm^3. How is the percentage uncertainty of the mean titre calculated?

Practical 2 Measurement of an enthalpy change (3.1.4 Energetics)

- List all the equipment required to measure the enthalpy change for a reaction.
- Describe how to perform this practical, step by step. Consider the steps required to minimise heat loss and to determine the maximum temperature change for the reaction.
- Show the equations needed to calculate ΔH from the experimental data.

Questions

1. Why is the enthalpy change calculated using this method less accurate than a value calculated using Hess's law?
2. Which value of n is used to calculate the enthalpy change?
3. Each reading on the thermometer has an uncertainty of ±0.5 °C. How is the percentage uncertainty of the temperature change calculated?

Practical 3 Investigation of how the rate of a reaction changes with temperature (3.1.5 Kinetics)

- List all the equipment required to measure the time taken for hydrochloric acid to react with a solution of sodium thiosulfate at different temperatures.
- Describe how to perform this practical, step by step.
- Show how the data obtained is plotted on a graph.

Questions

1. What are the dependent, independent and control variables in this experiment?
2. Each reading on the thermometer has an uncertainty of ±0.5 °C. How is the percentage uncertainty of the temperature change calculated?

Practical 4 Carry out simple test-tube reactions to identify: cations (group 2, NH_4^+) and anions (halides, OH^-, CO_3^{2-}, SO_4^{2-}) (3.2.2 Group 2) (3.2.3 Group 7)

- List all the equipment required to test for the presence of all the ions.
- Describe how to perform these tests and make it clear in which order they should be completed to reduce the amount of chemicals required.
- Draw a table to summarise the observations that you would make for all these tests.
- Write ionic equations for any reactions which occur.

Practical 5 Distillation of a product from a reaction (3.3.5 Alcohols)

- Draw a labelled diagram of the apparatus required to obtain a sample of ethanal from the reaction between ethanol and acidified potassium dichromate.
- Describe how to perform this practical, step by step and how to minimise the loss of the product obtained at the end.

Practical 6 Test for alcohol, aldehyde, alkene and carboxylic acid (3.3.6 Organic analysis)

- List all the equipment required to test for the presence of each of these.
- Describe how to perform these tests and make it clear in which order they should be completed to reduce the amount of chemicals required.
- Draw a table to summarise the observations that you would make for all these tests.
- Write equations for any reactions which occur.

→ Graph Practice

Graph Plotting (*graph paper at the end of questions*)

Task 1

1. Plot the graph of log(ionisation energy) against electron number removed. Join the individual points using straight lines.

Electron number removed	1	2	3	4	5	6	7	8	9	10	11	12
Ionisation energy / kJ mol^{-1}	738	1451	7733	10 541	13 629	17 995	21 704	25 657	31 644	35 463	169 996	189 371
Log (ionisation energy)	2.87	3.16	3.89	4.02	4.13		4.34	4.41	4.50	4.55	5.23	

2. Identify on the graph, using a circle, the points that represent the removal of the electrons in the outermost energy level of magnesium.

Task 2

When boric acid (H_3BO_3) is applied as a coating on wood, it acts as a fire retardant by decreasing the rate of combustion.

Thermal decomposition of boric acid takes place in two stages.

In an experiment a sample of boric acid was heated in a crucible at 170 °C. The results of this experiment are given in the table.

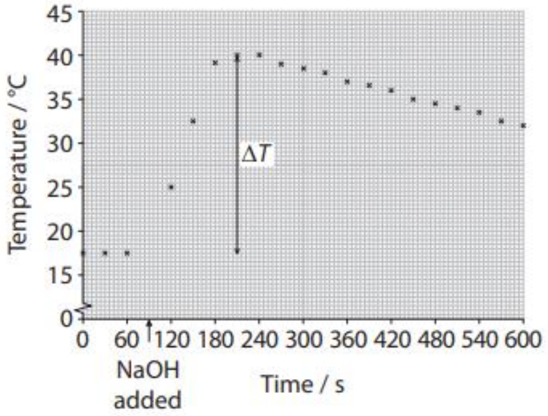
Time of heating / minutes	Mass of crucible and contents / g
0	35.85
5	35.10
10	34.41
15	34.00
20	33.70
25	33.56
30	33.50
35	33.50

Plot a graph of the results from the table above to show the mass of the crucible and boric acid (y-axis) against time of heating on the grid.

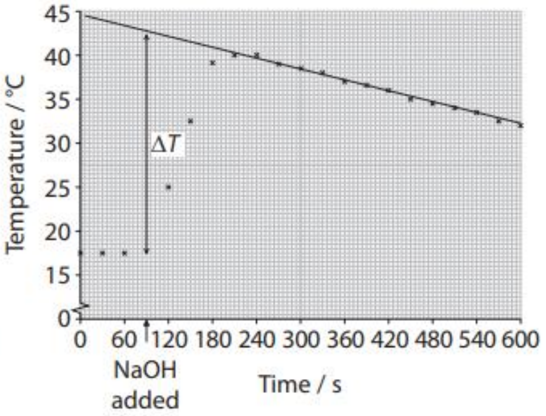
Task 3

Which graph shows the correct way that the maximum temperature rise should be determined?

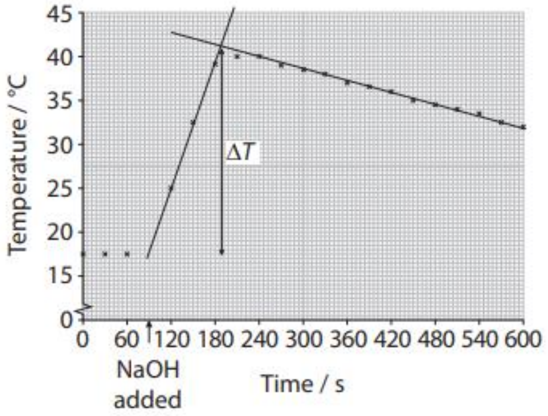
☒ **A**



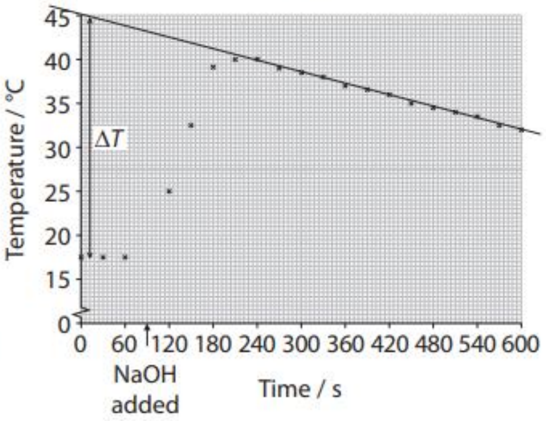
☒ **B**



☒ **C**



☒ **D**



Task 4

A student carried out a reaction between magnesium ribbon and aqueous trichloroethanoic acid in order to determine the enthalpy change. The equation for the reaction is shown:

Mg(s) + 2H+(aq) ⟶ Mg2+(aq) + H2(g)

The student measured the initial temperature of the trichloroethanoic acid and again every minute for 3 minutes before adding the magnesium ribbon at the fourth minute.
The student continued to measure the temperature every minute for a further 10 minutes.

1. Plot a graph for the measurements shown below.

Time /minutes	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Temperature / °C	23.0	22.5	23.0	23.0		33.5	37.5	39.5	39.5	39.5	39.0	39.0	39.0	39.0	38.5

2. The student used 240 mg of magnesium and 10.0 cm3 of aqueous trichloroethanoic acid (an excess).

Use these data and information determined from the graph above to calculate the enthalpy change, in kJ mol−1, for this reaction. (The specific heat capacity of water = 4.18 J K−1 g−1) Show your working. Give your answer to an appropriate precision.

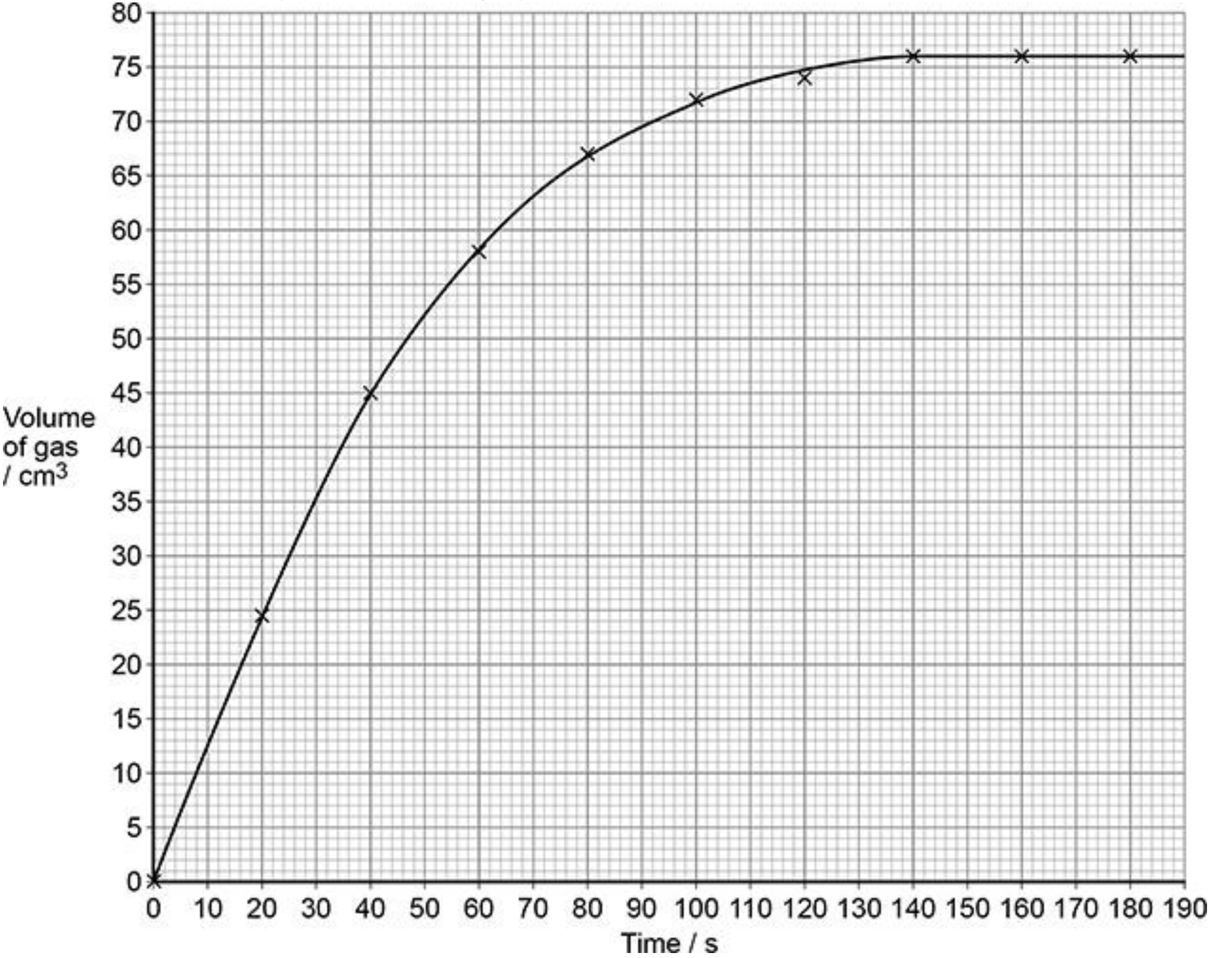
Gradients

Task 1

A student investigated the reaction between magnesium and hydrochloric acid.

$$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

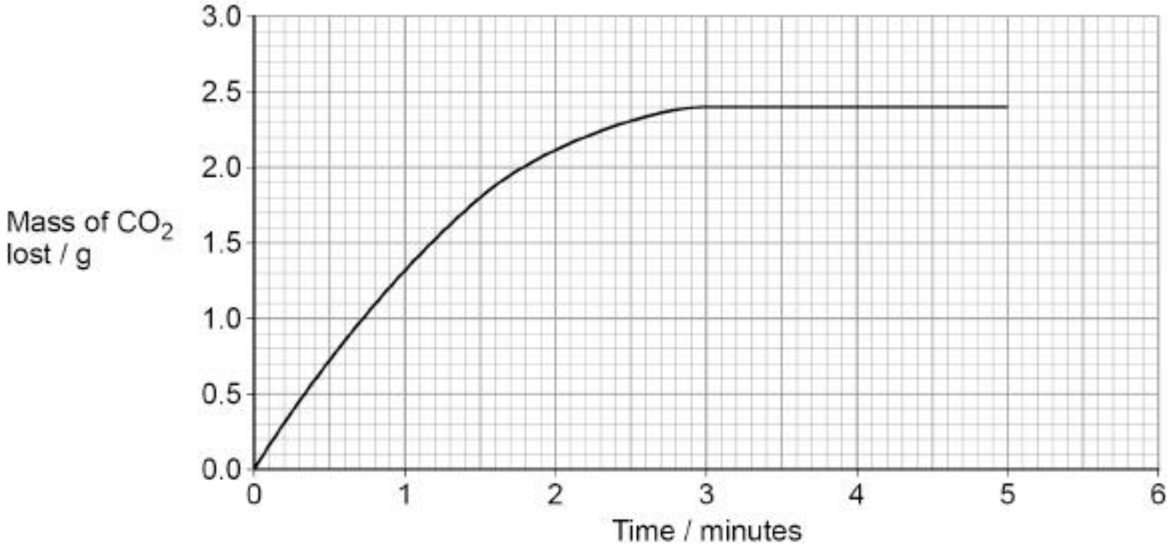
The results are plotted on this graph.



1. Determine the rate of reaction at 20s, 70s and 140s. You should draw a tangent on the curve and show your workings. Give the units.
   - Rate at 20s: _________________________
   - Rate at 70s: _________________________
   - Rate at 140s: ________________________
2. Explain, in terms of collision theory how the rate of reaction changes during the reaction.
3. Sketch a line on the graph to show how the reaction would differ if the mass of Mg was halved.
4. Sketch a dotted line (or in a different colour) to show how the reaction would differ if the concentration of HCl was increased.

Task 2

The graph shows the results of this experiment when 25.0 cm^3 of a 2.23 mol dm^{-3} solution of ethanoic acid reacts with an excess of $NaHCO_3$



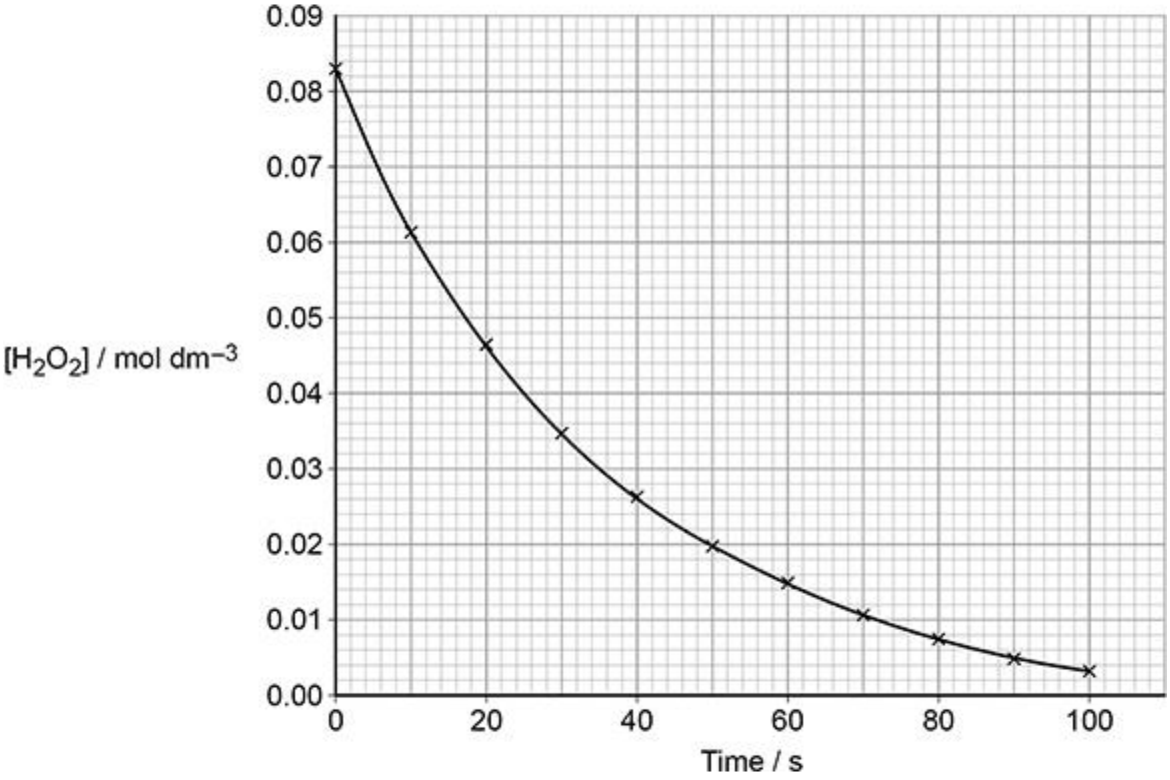
1. Determine the rate at 0.5minutes and 2.0minutes. Draw a tangent on the curve, show your workings and include the correct units.
    - Rate at 0.5 minutes: ________________________
    - Rate at 2.0minutes: ________________________
2. Chloroethanoic acid is a stronger acid than ethanoic acid.
   Sketch, on the graph, the curve you would expect when 25.0 cm^3 of a 2.23 mol dm^{-3} solution of chloroethanoic acid reacts with an excess of $NaHCO_3$
3. Suggest why chloroethanoic acid is a stronger acid than ethanoic acid.
4. Explain the shape of the graph.

Task 3

Hydrogen peroxide solution decomposes to form water and oxygen.

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

The reaction is catalysed by manganese(IV) oxide.



1. Why is the rate fastest at the start?

2. Determine the rate of reaction for the following concentrations. You should show your workings including the tangent on the graph and the units.

   - 0.04 mol dm^{-3}
   - 0.02 mol dm^{-3}

3. Explain the effect of concentration on the rate of reaction in terms of collision theory.

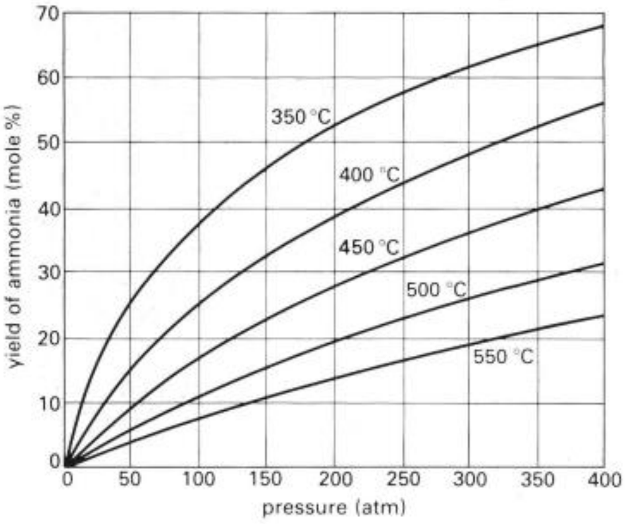
Equilibria

Task 1

Use the graph to answer the following questions:

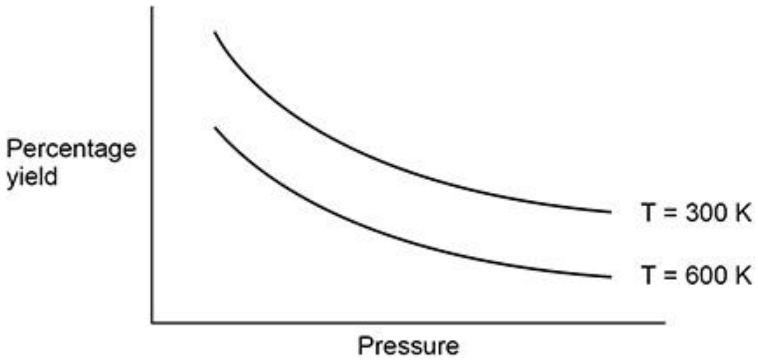
1. How can you tell from the graph that the forward reaction is exothermic?
2. How does pressure effect the yield of ammonia? How can you tell?
3. Explain why 450°C is used in industry rather than 350°C?

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ ΔH = exothermic



Task 2

1. Explain how the diagram shows that the forward reaction in this equilibrium is exothermic.
2. State whether the forward reaction in this equilibrium results in an increase, decrease or no change in the amount, in moles, of gas.



3. Explain why using a catalyst has no effect on the percentage yield.

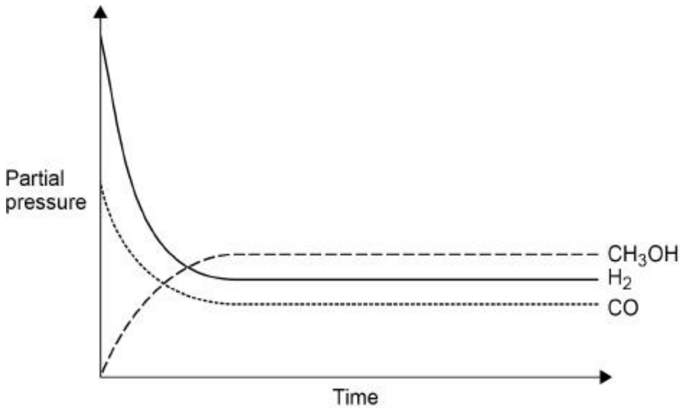
Task 3

Methanol can be manufactured in a reversible reaction as shown.

$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^\ominus = -91\ kJ\ mol^{-1}$

The graph shows how the partial pressures change with time at a constant temperature.

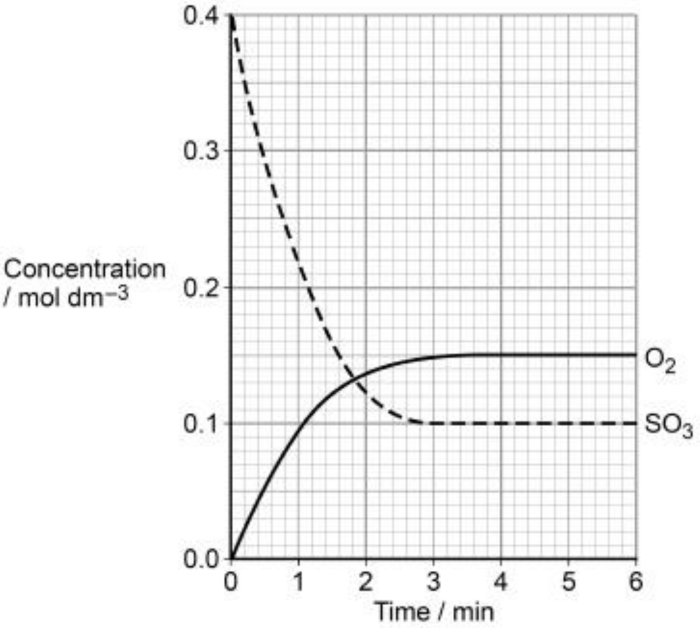
1. Draw a cross (x) on the appropriate axis of the graph when the mixture reaches equilibrium.



Task 4

Sulfur trioxide decomposes to form sulfur dioxide and oxygen at temperature ***T***$_1$ according to the equilibrium shown.

$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ $\Delta H = +196$ kJ mol^{-1}

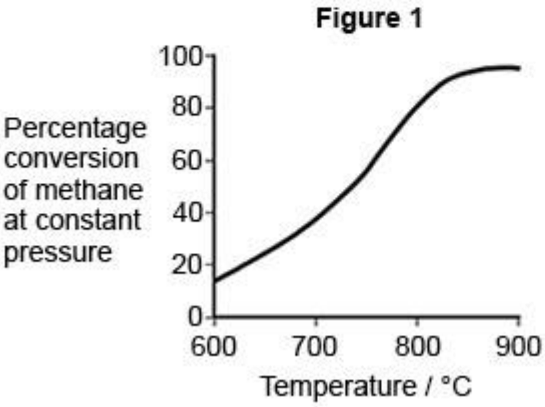


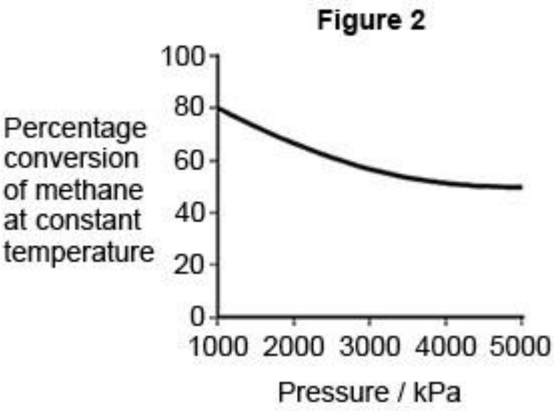
The graph shows the concentrations of sulfur trioxide and of oxygen over a period of 6 minutes at temperature ***T***$_1$

1. State the time, to the nearest minute, when equilibrium is first established.
2. Sketch on the graph above how the concentration of sulfur dioxide changes over these 6 minutes at temperature T1

Task 5

The first stage involves a gaseous equilibrium between the reactants (methane and steam), and some gaseous products. **Figures 1** and **2** show the percentage conversion of methane into the gaseous products under different conditions at equilibrium.





1. Deduce the optimum **temperature** for the industrial conversion of methane and steam into the gaseous products.
2. Deduce the optimum **pressure** for the industrial conversion of methane and steam into the gaseous products.
3. Is the reaction between methane and steam, endothermic or exothermic? Explain how you can tell from the graphs.

→ Past Paper Questions and Exam Practice

Click on each link to take you to the past paper questions.

You can print or complete the answers on paper. Work through the papers in timed conditions using a black pen. Where you have left gaps, use your notes to help you answer these in a different colour pen.

Paper 1 Questions

Paper 2 Questions

Paper 3 Questions

→ New Content: Kp

Watch the videos using the links below:

Equilibrium Constant Kp | A level Chemistry | Explained
AQA 1.10 Equilibrium Constant Kp

- Produce detailed Cornell Notes

- Answer the following core questions:

1. What is Kp?
2. In which state do the reactants and products have to be when calculating Kp?
3. $A + B \rightleftharpoons 2D + E$
   Write an expression for Kp in the reaction above?
4. What values for A, B, D and E are put into the expression to calculate Kp?
5. What is the equation for calculating mole fraction?
6. What is the equation for calculating partial pressure?
7. What is the sum of all of the mole fractions equal to?
8. How is the total pressure of the gas mixture calculated?
9. $A + B \rightleftharpoons 2D + E$
   At equilibrium there are 0.5 mol A, 0.2 mol B, 0.6 mol D and 0.6 mol E. What are the mole fractions of each gas A, B, D and E?
10. Using the above equation, calculate the partial pressures of A, B, D, and E if the total pressure of the reaction mixture is 715 kPa?
11. Using the equation above calculate the value of Kp and give its units (show your working out)?
12. What effect does a change in pressure at equilibrium have on the value of Kp?
13. What effect does the addition of a catalyst have on the value of Kp?
14. $A + B \rightleftharpoons 2D + E \quad \Delta H = -100$ kJ/mol
    What would happen to the value of Kp if the temperature was increased? Explain your answer.
15. $A + B \rightleftharpoons 2D + E \quad \Delta H = -100$ kJ/mol
    What would happen to the value of Kp if the temperature was decreased? Explain your answer.