

Percentage Yield

Mark Scheme



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1. Percentage yield measures a reaction's efficiency by comparing the actual yield to the theoretical yield. The following questions involve percentage yield. Circle the letter of the correct answer of each.

- a) A reaction is expected to produce a maximum of 40.0 grams of a product. If the experiment only produces 32.0 grams, what is the percentage yield?

C – is the only correct answer.	1 mark
	Total: 1 mark

- b) Before an experiment, a student uses the balanced chemical equation to calculate that they should produce a maximum of 4.0 g of copper sulfate. This 4.0 g value is known as the

B – is the only correct answer.	1 mark
	Total: 1 mark

- c) In many chemical reactions, the percentage yield is less than 100%. Which of the following is a valid reason for this?

B – is the only correct answer.	1 mark
	Total: 1 mark

- d) In an experiment, the maximum theoretical yield of a salt was calculated to be 20 g. A student only managed to produce 15 g of the salt. What is the percentage yield?

B – is the only correct answer.	1 mark
	Total: 1 mark

- e) Which statement best describes percentage yield?

C – is the only correct answer.	1 mark
	Total: 1 mark

(Total for Question 1 is 5 marks)

2. The following questions are about percentage yield.

- a) A student has reacted some sodium metal with chlorine gas to produce sodium chloride.

The theoretical yield of sodium chloride from this reaction is 11.70 g.

In the experiment, the student produces an actual yield of 9.90 g.

Calculate the percentage yield of sodium chloride. Give your answer to **one** decimal place.

$(9.90 \div 11.70) \times 100$	1 mark
84.6%	1 mark
	Total: 2 marks

- b) In an experiment, a sample of calcium carbonate is heated strongly.

The theoretical yield of calcium oxide is calculated to be 28.0 g.

The actual mass of calcium oxide produced is 25.5 g.

Calculate the percentage yield of calcium oxide. Give your answer to **three** significant figures.

$(25.5 \div 28.0) \times 100$	1 mark
91.1%	1 mark
	Total: 2 marks

- c) An excess of iron is added to a solution of copper(II) sulfate.

The maximum mass of copper that could possibly be formed is 6.4 g.

After filtering and drying, the student collects 5.8 g of copper.

What is the percentage yield for this reaction? Give your answer to **three** significant figures.

$(5.80 \div 6.40) \times 100$	1 mark
90.6%	1 mark
	Total: 2 marks

- d) Ammonia is produced in the Haber process.

Calculate the percentage yield of ammonia. Give your answer to **three** significant figures.

$(85 \div 500) \times 100$	1 mark
17.0%	1 mark
	Total: 2 marks

- e) A chemist makes a sample of the insoluble salt, silver chloride, by mixing two solutions together.

Based on the amounts of reactants used, the theoretical yield of silver chloride is 14.35 g.

After filtering and drying the product, the chemist obtains an actual yield of 13.80 g of silver chloride.

Calculate the percentage yield of silver chloride. Give your answer to **three** significant figures.

$(13.80 \div 14.35) \times 100$	1 mark
96.2%	1 mark
	Total: 2 marks

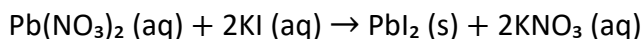
- f) A student burns a piece of magnesium ribbon in a crucible to form magnesium oxide.

$(3.5 \div 4.0) \times 100$	1 mark
87.5%	1 mark
	Total: 2 marks

(Total for Question 2 is 12 marks)

3. A student prepares a sample of lead(II) iodide by mixing a solution of lead(II) nitrate with a solution of potassium iodide. Lead(II) iodide is an insoluble yellow solid.

The balanced chemical equation for the reaction is:



The student calculated that the theoretical yield of lead(II) iodide is 9.2 g.

The student filters the mixture and after drying the solid product, finds the actual yield is 8.5 g.

- a) Calculate the percentage yield of lead(II) iodide. Give your answer to one decimal place.

$(8.5 \div 9.2) \times 100$	1 mark
92.4%	1 mark
	Total: 2 marks

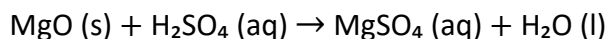
- b) The student's percentage yield was less than 100%.

Give **two** reasons why the actual yield in an experiment is often less than the theoretical yield.

Any two of the following:	
• The reaction may be reversible / did not go to completion. • Some of the product was lost during transfer (e.g., left behind in the beaker). • Some of the product was lost during the filtration process (e.g., stuck to filter paper). • There may have been unwanted side reactions that used up the reactants. • The reactants used may have been impure.	
	Total: 2 marks

(Total for Question 3 is 4 marks)

4. A student prepared magnesium sulfate by reacting magnesium oxide with sulfuric acid.
The balanced equation for the reaction is:



The student started with 4.0 g of magnesium oxide.

They produced 8.5 g of pure, dry magnesium sulfate crystals.

- a) Calculate the relative formula mass (M_r) of magnesium oxide (MgO).

$\text{MgO} = 24 + 16 = 40$	1 mark
	Total: 1 mark

- b) Calculate the number of moles in 4.0 g of magnesium oxide.

$\text{Moles} = \text{mass} \div M_r$	
$\text{Moles} = 4.0 \div 40$	
$\text{Moles} = 0.1 \text{ mols}$	1 mark
	Total: 1 mark

- c) Using the balanced equation, determine the maximum theoretical mass of magnesium sulfate (MgSO_4) that could be produced.

$M_r \text{ of product (MgSO}_4) = 24 + 32 + (4 \times 16) = 120$	1 mark
Mole ratio $\text{MgO} : \text{MgSO}_4$ is 1 : 1	
So, moles of $\text{MgSO}_4 = 0.1 \text{ mol}$	1 mark
$\text{Mass} = \text{mols} \times M_r$	
$\text{Mass} = 0.1 \times 120$	
$\text{Mass} = 12 \text{ g}$	1 mark
	Total: 3 marks

- d) Calculate the percentage yield of magnesium sulfate.
Give your answer to 3 significant figures.

$(8.5 \div 12) \times 100$	1 mark
70.8%	1 mark
	Total: 2 marks

(Total for Question 4 is 7 marks)

5. Copper(II) carbonate is a green powder that thermally decomposes to produce black copper(II) oxide and carbon dioxide gas. The equation is:



A scientist heats 12.35 g of copper(II) carbonate. She produces 7.16 g of copper(II) oxide.

- a) Calculate the theoretical yield of copper(II) oxide (CuO) from 12.35 g of copper(II) carbonate (CuCO₃).

Moles of reactant = $63.5 + 12 + (3 \times 16) = 123.5$	
Moles = $12.35 \div 123.5 = 0.1 \text{ mol}$	1 mark
Mole ratio CuCO ₃ : CuO is 1 : 1	
So, moles of CuO = 0.1 mol	1 mark
Mass = mols $\times$ Mr	
Mass = 0.1×79.5	
Mass = 7.95 g	1 mark
	Total: 3 marks

- b) Calculate the percentage yield for this experiment.

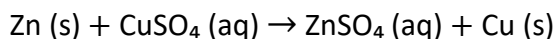
$(7.16 \div 7.95) \times 100$	1 mark
90.1%	1 mark
	Total: 2 marks

- c) Suggest **one** reason why the percentage yield was less than 100%.

Any one of the following:	
- The decomposition was incomplete / reaction did not go to completion. - The reaction is reversible. - Some product was lost during transfer (e.g., when scraping from the crucible). - There may have been unwanted side reactions. - The copper(II) carbonate reactant was not pure.	
	Total: 1 mark

(Total for Question 5 is 6 marks)

6. Zinc powder is added to an excess of copper(II) sulfate solution. The zinc displaces the copper.
The equation for this reaction is:



A student reacts 3.25 g of zinc powder and collects the solid copper produced.
After filtering and drying, the mass of the copper was 2.80 g.

- a) Calculate the number of moles in 3.25 g of zinc.

Moles = mass ÷ Mr	
Moles = 3.25 ÷ 65	
Moles = 0.05 mols	1 mark
	Total: 1 mark

- b) Use the mole ratio in the equation to determine the maximum mass of copper that can be produced.

Mole ratio Zn : Cu is 1 : 1.	
So, moles of Cu = 0.05	1 mark
Mass of Cu = 0.05 × 63.5 = 3.175	1 mark
	Total: 2 marks

- c) Calculate the percentage yield of copper.

(2.80 ÷ 3.175) × 100	1 mark
88.19%	1 mark
	Total: 2 marks

- d) Another student repeats the experiment but reports a percentage yield of 104%.

Explain why a yield greater than 100% is not possible and suggest a likely experimental error.

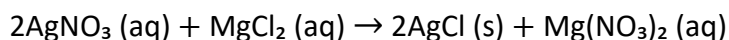
Any explanation to include one of the following:	
A yield that is greater than 100% is impossible because:	
- Mass cannot be created. - It would violate the law of conservation of mass. - The mass of the products cannot be more than the total mass of the reactants.	1 mark
Experimental error:	
- The copper product was not fully dry / was still wet when it was weighed. - The product was impure / contaminated (e.g. with unreacted zinc). - The balance was not zeroed/tared correctly before weighing the product.	1 mark
	Total: 2 marks

(Total for Question 6 is 7 marks)

7. A student prepares a sample of silver chloride, a white precipitate, by reacting silver nitrate with magnesium chloride.

They dissolve 6.80 g of solid silver nitrate (AgNO_3) in water and react it with 30.0 cm^3 of a 0.50 mol/dm^3 magnesium chloride (MgCl_2) solution.

The balanced equation for the reaction is:



After filtering, washing, and drying the precipitate, the student obtains an actual yield of 3.95 g of silver chloride (AgCl).

Calculate the percentage yield of silver chloride for this experiment.
Show all working.

Calculate moles of AgNO_3: Moles = mass $\div$ Mr Moles = $6.80 \div 170$ Moles = 0.04 mol	1 mark
Calculate moles of MgCl_2: Moles = (volume $\div$ 1000) $\times$ concentration Moles = $(30.0 \div 1000) \times 0.50$ Moles = 0.0150 mol Therefore, MgCl_2 is the limiting reactant.	1 mark
The mole ratio $\text{MgCl}_2 : \text{AgCl}$ is 1 : 2	1 mark
Calculate moles of product (AgCl): Moles of AgCl = Moles of $\text{MgCl}_2 \times 2$ Moles of AgCl = $0.0150 \times 2 = 0.0300 \text{ mol}$ Moles of AgCl = 0.0300 mol	1 mark
Calculate the theoretical yield of AgCl: Mr of AgCl = $108 + 35.5 = 143.5$ Mass = $0.0300 \times 143.5 = \mathbf{4.305 \text{ g}}$	1 mark
Calculate the percentage yield: $(3.95 \div 4.305) \times 100 = \mathbf{91.8\%}$	1 mark
Marking note: correct answer with no working shown scores zero marks.	
	Total: 6 marks

(Total for Question 7 is 6 marks)