

The mole concept: counting particles by weighing, with practice

Atoms and molecules are far too small to count one by one, so chemists count them by weighing. The idea that makes this possible is the mole. These notes are written for Cambridge O Level Chemistry 5070 and Cambridge IGCSE Chemistry 0620. Read the rules and worked examples first, then work through the 25 questions, which start easy and get harder. Full worked answers are at the end.

Relative atomic masses used in this document: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, Al = 27, S = 32, Cl = 35.5, K = 39, Ca = 40, Fe = 56, Cu = 64, Zn = 65. In the exam, always take values from the Periodic Table printed on the paper.

What you need to know

1. Relative atomic mass and relative formula mass

Atoms are compared with an atom of carbon-12, which is given a mass of exactly 12 units. Most elements have more than one isotope, so the relative atomic mass is an average that takes account of how common each isotope is. That is why chlorine has $A_r = 35.5$ and not a whole number.

Relative atomic mass, A_r , is the average mass of the atoms of an element, taking account of its isotopes and their abundances, compared with 1/12 of the mass of one atom of carbon-12.

Relative formula mass, M_r , is the sum of the relative atomic masses of all the atoms shown in the formula. For a substance made of molecules it is also called relative molecular mass.

A_r and M_r are ratios, so they have no units. To find M_r , multiply each A_r by the number of atoms of that element in the formula and add. Numbers outside a bracket multiply everything inside it. In a hydrated salt, the water after the dot is included.

Worked example 1: relative formula mass

Calculate M_r of aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$.

Count the atoms: Al = 2, S = 3, O = $3 \times 4 = 12$.

$M_r = (2 \times 27) + (3 \times 32) + (12 \times 16) = 54 + 96 + 192 = 342$

2. The mole and the Avogadro constant

One mole of any substance contains the same number of particles. This number is the Avogadro constant, 6.02×10^{23} . The particles might be atoms (for example Fe), molecules (for example H_2O) or formula units of an ionic compound (for example NaCl), so always state which particles you mean.

The mole is the amount of substance that contains 6.02×10^{23} particles. This number is the **Avogadro constant**.

The mass of one mole of a substance, in grams, is numerically equal to its A_r or M_r . This is its **molar mass**, with units g/mol.



So 1 mol of carbon atoms has a mass of 12 g, 1 mol of water molecules has a mass of 18 g, and both contain 6.02×10^{23} particles. This is how weighing lets us count.

3. Converting between mass, moles and particles

Rules to learn

number of moles = mass (g) $\div$ molar mass (g/mol)

mass (g) = number of moles $\times$ molar mass

number of particles = number of moles $\times 6.02 \times 10^{23}$

Moles are the bridge. You cannot go directly from mass to particles: always convert to moles first. Watch the particle the question asks for. One mole of CO_2 molecules contains 1 mol of carbon atoms and 2 mol of oxygen atoms, so 3 mol of atoms altogether.

Worked example 2: mass to particles

How many molecules and how many atoms are there in 8.8 g of carbon dioxide?

Step 1: $M_r(\text{CO}_2) = 12 + (2 \times 16) = 44$, so molar mass = 44 g/mol.

Step 2: moles = $8.8 \div 44 = 0.20$ mol of CO_2 molecules.

Step 3: molecules = $0.20 \times 6.02 \times 10^{23} = 1.204 \times 10^{23}$, which is 1.20×10^{23} to 3 significant figures.

Step 4: each molecule has 3 atoms, so atoms = $3 \times 1.204 \times 10^{23} = 3.61 \times 10^{23}$.

4. Molar gas volume

Gas particles are far apart, so the volume of a gas depends on the number of particles, not on what the gas is. At the same temperature and pressure, equal volumes of all gases contain equal numbers of molecules.

One mole of any gas occupies **24 dm³** (24 000 cm³) at room temperature and pressure (r.t.p.). This is the **molar gas volume**.

number of moles of gas = volume (dm³) $\div$ 24 = volume (cm³) $\div$ 24 000

Remember: 1 dm³ = 1000 cm³. The rule only works for gases. It does not apply to liquids or solids, and it does not apply to a solution of a gas in water.

Worked example 3: gas volume

Calculate the volume at r.t.p. of 1.6 g of methane, CH_4 .

$M_r(\text{CH}_4) = 12 + 4 = 16$. Moles = $1.6 \div 16 = 0.10$ mol.

Volume = $0.10 \times 24 = 2.4$ dm³ (or 2400 cm³).

5. Concentration of solutions

Concentration tells you how much solute is dissolved in each cubic decimetre of solution. It can be

given in moles or in grams.

$$\text{concentration (mol/dm}^3\text{)} = \text{number of moles of solute} \div \text{volume of solution (dm}^3\text{)}$$

$$\text{concentration (g/dm}^3\text{)} = \text{mass of solute (g)} \div \text{volume of solution (dm}^3\text{)}$$

$$\text{concentration (g/dm}^3\text{)} = \text{concentration (mol/dm}^3\text{)} \times \text{molar mass}$$

Volumes in titrations are usually measured in cm^3 . Divide by 1000 to change cm^3 into dm^3 before you use the formula. A common slip is to forget this step and get an answer 1000 times too big.

Worked example 4: titration

25.0 cm^3 of sodium hydroxide solution is exactly neutralised by 20.0 cm^3 of 0.125 mol/dm^3 hydrochloric acid. Calculate the concentration of the sodium hydroxide in mol/dm^3 and in g/dm^3 .

Equation: $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ (mole ratio 1 : 1)

$$\text{Moles of HCl} = 0.125 \times (20.0 \div 1000) = 0.00250 \text{ mol}$$

$$\text{Moles of NaOH} = 0.00250 \text{ mol (1 : 1 ratio)}$$

$$\text{Concentration of NaOH} = 0.00250 \div (25.0 \div 1000) = 0.100 \text{ mol/dm}^3$$

$$M_r(\text{NaOH}) = 23 + 16 + 1 = 40, \text{ so concentration} = 0.100 \times 40 = 4.00 \text{ g/dm}^3$$

6. Empirical and molecular formulae

The **empirical formula** is the simplest whole number ratio of the atoms of each element in a compound.

The **molecular formula** is the actual number of atoms of each element in one molecule.

Method for the empirical formula: (1) write the mass or percentage of each element; (2) divide each by its A_r to get moles; (3) divide every answer by the smallest; (4) if a value is close to a half or a third (for example 1.5 or 1.33), multiply all values by 2 or 3 to get whole numbers. Do not round 1.5 to 2.

Method for the molecular formula: find the M_r of the empirical formula, divide the given M_r by it, and multiply every subscript in the empirical formula by the answer.

Worked example 5: empirical and molecular formula

A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen by mass. Its M_r is 60. Find its empirical and molecular formulae.

	C	H	O
mass in 100 g	40.0	6.7	53.3
moles ($\div A_r$)	$40.0 \div 12 = 3.33$	$6.7 \div 1 = 6.7$	$53.3 \div 16 = 3.33$
$\div$ smallest (3.33)	1	2.0	1

Empirical formula = CH_2O . Its $M_r = 12 + 2 + 16 = 30$.

$60 \div 30 = 2$, so the molecular formula is $\text{C}_2\text{H}_4\text{O}_2$.

The same method finds the number of molecules of water of crystallisation, x , in a hydrated salt: work out the moles of anhydrous salt and the moles of water lost on heating, then find the ratio.

7. Reacting masses from equations

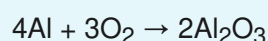
A balanced equation gives the ratio of moles, not of masses. The method is always the same three steps.

Three steps for any reacting mass or volume question

1. Change the quantity you know into moles.
2. Use the balanced equation to find the moles of the substance you want (use the mole ratio).
3. Change those moles into the quantity asked for (mass, gas volume, particles or concentration).

Worked example 6: reacting masses

Calculate the mass of aluminium oxide formed when 5.4 g of aluminium burns completely in oxygen.



Step 1: moles of Al = $5.4 \div 27 = 0.20$ mol

Step 2: ratio Al : $\text{Al}_2\text{O}_3 = 4 : 2 = 2 : 1$, so moles of $\text{Al}_2\text{O}_3 = 0.20 \div 2 = 0.10$ mol

Step 3: $M_r(\text{Al}_2\text{O}_3) = (2 \times 27) + (3 \times 16) = 102$. Mass = $0.10 \times 102 = 10.2$ g

For large-scale questions in tonnes or kilograms, you can work in grams, or use the fact that the mass ratio stays the same whatever unit you use: if 160 g of a reactant gives 112 g of product, then 160 tonnes gives 112 tonnes.

8. Limiting reagent

In many reactions one reactant is used up first. This is the **limiting reagent**, and it decides how much product forms. The other reactant is in **excess**, so some of it is left over.

How to find the limiting reagent

1. Work out the moles of each reactant.
2. Use the equation to find how many moles of the second reactant are needed to react with all of the first.
3. If you have less than that, the second reactant is limiting. If you have more, the first reactant is limiting.
4. Always calculate the amount of product from the limiting reagent.

Worked example 7: limiting reagent

6.5 g of zinc is added to 100 cm³ of 1.00 mol/dm³ hydrochloric acid. $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$.

Which reactant is in excess, and what volume of hydrogen forms at r.t.p.?

Moles of Zn = $6.5 \div 65 = 0.10$ mol. Moles of HCl = $1.00 \times 0.100 = 0.10$ mol.

0.10 mol of Zn would need $2 \times 0.10 = 0.20$ mol of HCl. Only 0.10 mol is present, so HCl is the limiting reagent and zinc is in excess.

Moles of $H_2 = 0.10 \div 2 = 0.050$ mol. Volume = $0.050 \times 24 = 1.2 \text{ dm}^3$.

9. Percentage yield and percentage purity

percentage yield = (actual mass of product obtained $\div$ theoretical maximum mass) $\times 100$

percentage purity = (mass of pure substance in the sample $\div$ total mass of the sample) $\times 100$

The theoretical mass is what you calculate from the equation, using the limiting reagent. The actual yield is usually lower because the reaction may be reversible and not go to completion, some product is lost when filtering, transferring or crystallising, and side reactions may form other products. A yield above 100% usually means the product is still wet or impure.

For purity, you usually react the impure sample, measure how much product or gas forms, work back to the mass of pure substance, and compare it with the mass of sample you started with. The method assumes the impurities do not react.

Worked example 8: percentage yield

A student heats 10.0 g of calcium carbonate and obtains 4.9 g of calcium oxide. $CaCO_3 \rightarrow CaO + CO_2$. Calculate the percentage yield.

$M_r(CaCO_3) = 40 + 12 + 48 = 100$. Moles = $10.0 \div 100 = 0.100$ mol.

Ratio 1 : 1, so 0.100 mol of CaO. $M_r(CaO) = 56$. Theoretical mass = $0.100 \times 56 = 5.6$ g.

Percentage yield = $(4.9 \div 5.6) \times 100 = 87.5\%$

Exam tips

Show every step, with units: method marks are given even if the final number is wrong.

Keep at least 3 significant figures in working and round only at the end. Give the final answer to the same number of significant figures as the data, usually 2 or 3.

Always check that the equation is balanced before you use its mole ratio.

Questions

The questions are grouped from easier to harder. Use the relative atomic masses at the top of this document, the Avogadro constant 6.02×10^{23} and the molar gas volume 24 dm^3 at r.t.p.

Section A: formula mass, moles and particles

1. Calculate the relative formula mass of carbon dioxide, CO_2 . [1]
2. Calculate the relative formula mass of calcium hydroxide, $Ca(OH)_2$. [1]

3. Calculate the relative formula mass of hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. [2]
4. Calculate the number of moles in 11 g of carbon dioxide. [1]
5. Calculate the mass of 0.20 mol of sodium hydroxide, NaOH . [2]
6. A beaker contains 9.0 g of water. Calculate (a) the number of water molecules and (b) the total number of atoms. [3]
7. Calculate the percentage by mass of nitrogen in ammonium nitrate, NH_4NO_3 . [2]

Section B: gases and solutions

8. Calculate the volume, in dm^3 , occupied by 3.2 g of oxygen gas, O_2 , at r.t.p. [2]
9. Calculate the mass of 600 cm^3 of methane, CH_4 , at r.t.p. [2]
10. 4.0 g of sodium hydroxide is dissolved in water and the solution is made up to 250 cm^3 . Calculate the concentration in (a) mol/dm^3 and (b) g/dm^3 . [3]
11. Calculate the number of moles, and the mass, of sulfuric acid, H_2SO_4 , in 25.0 cm^3 of 0.200 mol/dm^3 solution. [3]

Section C: formulae from data

12. A hydrocarbon contains 2.4 g of carbon combined with 0.6 g of hydrogen. Its relative molecular mass is 30. Find (a) the empirical formula and (b) the molecular formula. [4]
13. A white solid contains 32.4% sodium, 22.5% sulfur and 45.1% oxygen by mass. Calculate its empirical formula. [3]
14. 5.00 g of hydrated magnesium sulfate, $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$, is heated until all the water is lost. The mass of anhydrous MgSO_4 left is 2.44 g. Calculate the value of x. [4]

Section D: reacting masses and gas volumes

15. Calcium carbonate decomposes on heating: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. Calculate the mass of calcium oxide formed from 25 g of calcium carbonate. [3]
16. Using the same reaction, calculate the volume of carbon dioxide, at r.t.p., given off from 25 g of calcium carbonate. [2]
17. In the blast furnace, $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$. Calculate the maximum mass of iron, in tonnes, that can be obtained from 1.6 tonnes of iron(III) oxide. [3]
18. 25.0 cm^3 of sodium hydroxide solution is exactly neutralised by 20.0 cm^3 of 0.100 mol/dm^3 sulfuric acid. $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. Calculate the concentration of the sodium hydroxide in (a) mol/dm^3 and (b) g/dm^3 . [4]
19. 0.48 g of magnesium ribbon reacts with excess dilute hydrochloric acid. $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. Calculate the volume of hydrogen formed, in cm^3 , at r.t.p. [3]

Section E: limiting reagent, yield and purity

20. 2.4 g of magnesium is added to 50.0 cm^3 of 1.00 mol/dm^3 hydrochloric acid. $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. (a) Show which reactant is limiting. (b) Calculate the volume of hydrogen, in cm^3 , at r.t.p. (c) Calculate the mass of the reactant left over. [5]
21. 5.6 g of iron is heated with 4.0 g of sulfur. $\text{Fe} + \text{S} \rightarrow \text{FeS}$. (a) Identify the limiting reagent. (b)

Calculate the maximum mass of iron(II) sulfide formed. (c) Calculate the mass of the excess reactant that remains. [5]

22. 12.0 g of calcium carbonate is heated and 5.6 g of calcium oxide is collected. Calculate the percentage yield. Suggest one reason why the yield is less than 100%. [4]

23. In a test run of the Haber process, $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, 14 kg of nitrogen reacts with excess hydrogen and 2.55 kg of ammonia is produced. Calculate the percentage yield and explain why the yield is low. [4]

24. A 10.0 g sample of impure limestone reacts with excess hydrochloric acid. 1.92 dm³ of carbon dioxide is collected at r.t.p. $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$. Calculate the percentage purity of the limestone. State one assumption you have made. [5]

25. A student makes copper(II) sulfate crystals by adding excess copper(II) oxide to 50.0 cm³ of 1.00 mol/dm³ sulfuric acid, filtering, and crystallising the filtrate. $\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$. Each mole of CuSO₄ gives one mole of CuSO₄·5H₂O crystals. (a) Explain why copper(II) oxide is added in excess. (b) Calculate the theoretical mass of CuSO₄·5H₂O crystals. (c) The student obtains 9.5 g of dry crystals. Calculate the percentage yield. (d) Evaluate whether the method could be improved to give a higher yield. [7]

Answers

1. $M_r(\text{CO}_2) = 12 + (2 \times 16) = 12 + 32 = 44$

2. $\text{Ca}(\text{OH})_2$ contains 1 Ca, 2 O and 2 H. $M_r = 40 + (2 \times 16) + (2 \times 1) = 40 + 32 + 2 = 74$

3. CuSO_4 : $64 + 32 + (4 \times 16) = 160$. Water: $5 \times (2 + 16) = 5 \times 18 = 90$. $M_r = 160 + 90 = 250$

4. $M_r(\text{CO}_2) = 44$. Moles = $11 \div 44 = 0.25 \text{ mol}$

5. $M_r(\text{NaOH}) = 23 + 16 + 1 = 40$. Mass = $0.20 \times 40 = 8.0 \text{ g}$

6. (a) $M_r(\text{H}_2\text{O}) = 18$. Moles = $9.0 \div 18 = 0.50 \text{ mol}$. Molecules = $0.50 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$.
(b) Each H₂O molecule has 3 atoms. Atoms = $3 \times 3.01 \times 10^{23} = 9.03 \times 10^{23}$.

7. $M_r(\text{NH}_4\text{NO}_3) = (2 \times 14) + (4 \times 1) + (3 \times 16) = 28 + 4 + 48 = 80$. There are 2 nitrogen atoms, mass 28. Percentage = $(28 \div 80) \times 100 = 35\%$

8. $M_r(\text{O}_2) = 32$. Moles = $3.2 \div 32 = 0.10 \text{ mol}$. Volume = $0.10 \times 24 = 2.4 \text{ dm}^3$. (Using $A_r = 16$ instead of $M_r = 32$ is a common error, because oxygen gas is diatomic.)

9. Moles = $600 \div 24\,000 = 0.025 \text{ mol}$. $M_r(\text{CH}_4) = 16$. Mass = $0.025 \times 16 = 0.40 \text{ g}$

10. (a) Moles of NaOH = $4.0 \div 40 = 0.10 \text{ mol}$. Volume = $250 \div 1000 = 0.250 \text{ dm}^3$. Concentration = $0.10 \div 0.250 = 0.40 \text{ mol/dm}^3$.

(b) Concentration = $4.0 \text{ g} \div 0.250 \text{ dm}^3 = 16 \text{ g/dm}^3$ (check: $0.40 \times 40 = 16$).

11. Moles = $0.200 \times (25.0 \div 1000) = 0.00500 \text{ mol}$. $M_r(\text{H}_2\text{SO}_4) = 2 + 32 + 64 = 98$. Mass = $0.00500 \times 98 = 0.49 \text{ g}$

12. (a) Moles of C = $2.4 \div 12 = 0.20$. Moles of H = $0.6 \div 1 = 0.60$. Divide by the smallest (0.20): C = 1, H = 3. Empirical formula = **CH₃**.

(b) $M_r(\text{CH}_3) = 12 + 3 = 15$. $30 \div 15 = 2$. Molecular formula = **C₂H₆** (ethane).

13. Na: $32.4 \div 23 = 1.41$. S: $22.5 \div 32 = 0.703$. O: $45.1 \div 16 = 2.82$.

Divide by the smallest (0.703): Na = 2.00, S = 1, O = 4.01.

Empirical formula = **Na₂SO₄** (sodium sulfate).

14. Mass of water lost = 5.00 – 2.44 = 2.56 g.

$M_r(\text{MgSO}_4) = 24 + 32 + 64 = 120$. Moles of $\text{MgSO}_4 = 2.44 \div 120 = 0.0203$ mol.

Moles of $\text{H}_2\text{O} = 2.56 \div 18 = 0.142$ mol.

Ratio $\text{H}_2\text{O} : \text{MgSO}_4 = 0.142 \div 0.0203 = 7.0$. So **x = 7** and the formula is $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.

15. Step 1: $M_r(\text{CaCO}_3) = 100$. Moles = $25 \div 100 = 0.25$ mol.

Step 2: ratio $\text{CaCO}_3 : \text{CaO} = 1 : 1$, so 0.25 mol of CaO.

Step 3: $M_r(\text{CaO}) = 40 + 16 = 56$. Mass = $0.25 \times 56 = \mathbf{14\text{ g}}$

16. Ratio $\text{CaCO}_3 : \text{CO}_2 = 1 : 1$, so 0.25 mol of CO_2 . Volume = $0.25 \times 24 = \mathbf{6.0\text{ dm}^3}$

17. $M_r(\text{Fe}_2\text{O}_3) = (2 \times 56) + (3 \times 16) = 160$. From the equation, 1 mol of Fe_2O_3 (160 g) gives 2 mol of Fe ($2 \times 56 = 112$ g).

So 160 tonnes gives 112 tonnes, and 1.6 tonnes gives $1.6 \times (112 \div 160) = \mathbf{1.12\text{ tonnes}}$ of iron.

Alternative in grams: 1.6 tonnes = 1.6×10^6 g. Moles of $\text{Fe}_2\text{O}_3 = 1.6 \times 10^6 \div 160 = 1.0 \times 10^4$ mol.

Moles of Fe = 2.0×10^4 mol. Mass = $2.0 \times 10^4 \times 56 = 1.12 \times 10^6$ g = 1.12 tonnes.

18. (a) Moles of $\text{H}_2\text{SO}_4 = 0.100 \times (20.0 \div 1000) = 0.00200$ mol.

Ratio $\text{H}_2\text{SO}_4 : \text{NaOH} = 1 : 2$, so moles of NaOH = $2 \times 0.00200 = 0.00400$ mol.

Concentration of NaOH = $0.00400 \div (25.0 \div 1000) = \mathbf{0.160\text{ mol/dm}^3}$.

(b) $M_r(\text{NaOH}) = 40$. Concentration = $0.160 \times 40 = \mathbf{6.40\text{ g/dm}^3}$.

19. Moles of Mg = $0.48 \div 24 = 0.020$ mol. Ratio $\text{Mg} : \text{H}_2 = 1 : 1$, so 0.020 mol of H_2 .

Volume = $0.020 \times 24\,000 = \mathbf{480\text{ cm}^3}$

20. (a) Moles of Mg = $2.4 \div 24 = 0.10$ mol. Moles of HCl = $1.00 \times (50.0 \div 1000) = 0.050$ mol.

0.10 mol of Mg needs $2 \times 0.10 = 0.20$ mol of HCl, but only 0.050 mol is present. So **hydrochloric acid is the limiting reagent** and magnesium is in excess.

(b) Moles of $\text{H}_2 = 0.050 \div 2 = 0.025$ mol. Volume = $0.025 \times 24\,000 = \mathbf{600\text{ cm}^3}$.

(c) Moles of Mg that react = $0.050 \div 2 = 0.025$ mol. Moles left = $0.10 - 0.025 = 0.075$ mol. Mass left = $0.075 \times 24 = \mathbf{1.8\text{ g}}$ of magnesium.

21. (a) Moles of Fe = $5.6 \div 56 = 0.10$ mol. Moles of S = $4.0 \div 32 = 0.125$ mol. The ratio is 1 : 1, so 0.10 mol of Fe needs only 0.10 mol of S. Sulfur is in excess and **iron is the limiting reagent**.

(b) Moles of FeS = 0.10 mol. $M_r(\text{FeS}) = 56 + 32 = 88$. Mass = $0.10 \times 88 = \mathbf{8.8\text{ g}}$.

(c) Sulfur left = $0.125 - 0.10 = 0.025$ mol. Mass = $0.025 \times 32 = \mathbf{0.80\text{ g}}$ of sulfur.

Check: mass of reactants used = $5.6 + (4.0 - 0.80) = 8.8$ g, which equals the mass of FeS, as conservation of mass requires.

22. Moles of $\text{CaCO}_3 = 12.0 \div 100 = 0.120$ mol, giving 0.120 mol of CaO.

Theoretical mass of CaO = $0.120 \times 56 = 6.72$ g.

Percentage yield = $(5.6 \div 6.72) \times 100 = \mathbf{83\%}$ (83.3%).

Reason, any one: the calcium carbonate was not heated strongly enough or for long enough, so some did not decompose; some solid was lost when it was transferred; the calcium carbonate was impure.

23. Moles of $\text{N}_2 = 14\,000 \text{ g} \div 28 = 500$ mol. Ratio $\text{N}_2 : \text{NH}_3 = 1 : 2$, so 1000 mol of NH_3 .

$M_r(\text{NH}_3) = 14 + 3 = 17$. Theoretical mass = $1000 \times 17 = 17\,000 \text{ g} = 17 \text{ kg}$.

Percentage yield = $(2.55 \div 17) \times 100 = \mathbf{15\%}$.

The reaction is reversible ($\rightleftharpoons$), so it reaches equilibrium and does not go to completion: some ammonia breaks down into nitrogen and hydrogen again as fast as it forms. Only part of the nitrogen is converted in one pass.

24. Moles of $\text{CO}_2 = 1.92 \div 24 = 0.0800 \text{ mol}$.

Ratio $\text{CaCO}_3 : \text{CO}_2 = 1 : 1$, so 0.0800 mol of CaCO_3 reacted.

Mass of $\text{CaCO}_3 = 0.0800 \times 100 = 8.00 \text{ g}$.

Percentage purity = $(8.00 \div 10.0) \times 100 = \mathbf{80.0\%}$.

Assumption, any one: the impurities do not react with the acid to give carbon dioxide; all the carbon dioxide was collected (none dissolved in the water or escaped).

25. (a) Excess copper(II) oxide makes sure that all the sulfuric acid reacts, so the solution contains only copper(II) sulfate and no acid. The unreacted solid can be removed by filtering.

(b) Moles of $\text{H}_2\text{SO}_4 = 1.00 \times (50.0 \div 1000) = 0.0500 \text{ mol}$. The acid is the limiting reagent, so 0.0500 mol of CuSO_4 forms, giving 0.0500 mol of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

$M_r(\text{CuSO}_4 \cdot 5\text{H}_2\text{O}) = 250$. Theoretical mass = $0.0500 \times 250 = \mathbf{12.5 \text{ g}}$.

(c) Percentage yield = $(9.5 \div 12.5) \times 100 = \mathbf{76\%}$.

(d) Some loss is unavoidable, but the yield could be raised: rinse the beaker and filter paper with a little distilled water so that less solution is left behind; after the first crop of crystals, evaporate the remaining solution further to obtain a second crop; avoid overheating, which drives off water of crystallisation and changes the product. Some copper(II) sulfate always stays dissolved in the final solution, so a yield of 100% is not realistic, and a yield near 76% is reasonable for a crystallisation.