Chapter Summary: Stoichiometry

- **Chapter:** 4: Stoichiometry - chemical calculations
- **Book/PDF:** Chapter-04.pdf
- **Pages:** 54–68
- **Exam level:** Cambridge IGCSE (0610)

1) Big-picture overview

This chapter introduces **stoichiometry**, which is the foundation of all chemical calculations. It's about using the relationships between reactants and products in a chemical reaction to figure out "how much." You'll learn about the concept of the **mole**, a unit that lets us count atoms and molecules by weighing them. This is crucial because atoms are too small to count individually. By understanding the mole, relative atomic mass, and balanced chemical equations, you can calculate the exact mass of reactants you need, predict the mass of the product you'll make (the **theoretical yield**), and even determine the chemical formula of an unknown substance. These skills are essential for chemists in labs and in industry to run reactions efficiently and without waste.

2) Syllabus mapping

Outcome code	Outcome description	Where covered (page)
C4.1	Define relative atomic mass, A_r, and relative molecular mass, M_r.	54–55
C4.1	Calculate relative molecular mass and relative formula mass.	54–55
C4.2	Define the mole and the Avogadro constant.	55
C4.3	Use the molar gas volume (24 dm^3 at r.t.p.) in calculations.	58
C4.4	Calculate stoichiometric reacting masses, volumes of gases and solutions, and concentrations of solutions expressed in mol/dm^3 and g/dm^3.	55, 62–64

Outcome code	Outcome description	Where covered (page)
C4.5	Calculate empirical formulae and molecular formulae.	60–62
C4.6	Calculate percentage yield and percentage purity.	65
C4.6	Calculate the percentage composition of an element in a compound.	65
C4.3	Use the concept of the mole to calculate the limiting reactant.	66

3) Key terms and definitions

Term	One-sentence definition	First appears (page)	Example/application
Relative atomic mass (A_r)	The average mass of the isotopes of an element compared to 1/12th of the mass of a carbon-12 atom[cite: 24].	54	The A_r of Magnesium (Mg) is 24[cite: 35, 38].
Relative molecular mass (M_r)	The sum of the relative atomic masses of all atoms in a molecule[cite: 40].	54	For water (H_2O), M_r = (2 × 1) + 16 = 18[cite: 112, 178].
Relative formula mass (M_r)	The sum of the relative atomic masses of all atoms in the formula of an ionic compound[cite: 51].	55	For sodium hydroxide (NaOH), M_r = 23 + 16 + 1 = 40[cite: 112].
Mole (mol)	The unit for the amount of a substance, which contains 6.02×10^{23} particles (the Avogadro constant)[cite: 64].	55	12 g of carbon is 1 mole of carbon atoms[cite: 69].

Term	One-sentence definition	First appears (page)	Example/application
Avogadro constant	The number of particles (atoms, ions, or molecules) in one mole of a substance, equal to 6.02×10^{23}[cite: 65, 81].	55	56 g of iron (1 mole) contains 6.02×10^{23} iron atoms[cite: 83].
Molecular formula	A formula that shows the actual number and type of different atoms in one molecule[cite: 49, 421].	55	The molecular formula for butene is C_4H_8[cite: 417].
Molar mass	The mass of one mole of an element or compound, with units of g/mol[cite: 88].	55	The molar mass of ethanol (C_2H_5OH) is 46 g/mol[cite: 188].
Molar gas volume (V_m)	The volume occupied by one mole of any gas at room temperature and pressure (r.t.p.), which is 24 dm^3[cite: 233, 234].	58	1 mole of CO_2 gas occupies 24 dm^3 at r.t.p.[cite: 233].
Concentration	The amount of solute dissolved in a given volume of solvent, measured in g/dm^3 or mol/dm^3[cite: 261, 282].	59	A solution containing 1 mole of NaOH in 1 dm^3 of water has a concentration of 1 mol/dm^3[cite: 283].
Empirical formula	The simplest whole-number ratio of the different atoms or ions in a compound[cite: 395].	61	The empirical formula for benzene (C_6H_6) is CH[cite: 427, 437].
Theoretical yield	The maximum amount of product that could be formed from the given amounts of reactants,	65	Burning 12 g of carbon should theoretically yield 44 g of CO_2[cite: 570, 571].

Term	One-sentence definition	First appears (page)	Example/application
	assuming 100% efficiency[cite: 571, 572].		
Limiting reactant	The reactant that is completely used up in a chemical reaction and therefore determines the amount of product formed[cite: 609, 610].	66	If excess carbonate is added to acid, the acid is the limiting reactant[cite: 621, 622].

4) Core concepts explained

Relative Mass (p. 54)

- The actual mass of an atom is incredibly small, so a relative scale is used for convenience[cite: 16, 17].
- The scale is based on the **carbon-12 isotope**, which is assigned a relative atomic mass (A_r) of exactly 12.00[cite: 18].
- The **relative atomic mass (A_r)** of an element is the weighted average mass of its isotopes compared to 1/12th the mass of a carbon-12 atom[cite: 24].
- For molecules (covalently bonded substances), the **relative molecular mass (M_r)** is the sum of the A_r values of all atoms in the molecule[cite: 40].
- For ionic compounds, the term **relative formula mass (M_r)** is used and calculated in the same way[cite: 51, 111].

The Mole and Avogadro's Constant (p. 55)

- The **mole (mol)** is a unit for the amount of substance[cite: 64].
- One mole of any substance contains 6.02×10^{23} particles (atoms, molecules, or ions). This number is the **Avogadro constant**[cite: 64, 65, 81].
- The mass of one mole of a substance in grams is numerically equal to its relative atomic/molecular/formula mass. This is called the **molar mass** (g/mol)[cite: 88].

- Example: The A_r of iron (Fe) is 56. Therefore, 1 mole of iron has a mass of 56 g and contains 6.02×10^{23} atoms[cite: 71, 83].

Moles and Gases (p. 58)

- **Avogadro's Law** states that equal volumes of all gases, at the same temperature and pressure, contain the same number of molecules[cite: 257, 258].
- At **room temperature and pressure (r.t.p.)**, one mole of *any* gas occupies a volume of **24 dm^3** (or 24,000 cm^3)[cite: 233].
- This value, 24 dm^3/mol, is called the **molar gas volume**[cite: 234].
- This relationship allows for easy conversion between the volume of a gas and the number of moles.

Moles and Solutions (p. 59)

- The **concentration** of a solution tells us how much solute is dissolved in a certain volume of solution[cite: 260].
- Concentration is typically measured in grams per cubic decimetre (g/dm^3) or, more commonly in chemistry, **moles per cubic decimetre (mol/dm^3)**[cite: 261].
- A solution with a concentration of 1 mol/dm^3 is called a 1 molar solution[cite: 283].
- Remember that **1 dm^3 = 1000 cm^3**. To convert cm^3 to dm^3, you must divide by 1000[cite: 264].

Calculating Formulae (p. 60-62)

- The **empirical formula** is the simplest whole-number ratio of atoms of each element in a compound[cite: 395]. It can be found experimentally.
- The **molecular formula** gives the actual number of atoms of each element in one molecule[cite: 421]. It is always a whole-number multiple of the empirical formula.
- To find the molecular formula, you need both the empirical formula and the relative molecular mass (M_r) of the compound[cite: 407].
- **Comparison of Formulae:**
Feature	Empirical Formula	Molecular Formula	Exam note
Definition	Simplest ratio of atoms[cite: 395].	Actual number of atoms[cite: 421].	All ionic

formulae are empirical. |
| **Example (Butene)** | CH_2 [cite: 401] | C_4H_8 [cite: 416, 417] | C_4H_8 is (CH_2) $\times$ 4. |
| **How to find** | From % composition or mass data[cite: 358]. | From empirical formula and M_r [cite: 407]. | You must calculate the empirical formula first. |

Reacting Quantities, Yield, and Purity (p. 62-66)

- A balanced chemical equation shows the **ratio of moles** of reactants and products[cite: 444, 447].
- The **Law of Conservation of Mass** states that the total mass of reactants equals the total mass of products in a chemical reaction[cite: 452, 453].
- The **limiting reactant** is the reactant that gets completely used up first, limiting the amount of product that can be formed[cite: 609, 610]. Other reactants are said to be in **excess**[cite: 609].
- **Percentage yield** compares the actual amount of product obtained in an experiment to the maximum amount that was theoretically possible[cite: 576, 581]. Reactions are rarely 100% efficient[cite: 566].
- **Percentage purity** calculates what proportion of a sample is the desired chemical, and what proportion is impurities[cite: 592].

5) Diagrams and micrographs (figures)

- **Figure 4.2 (p. 56):** Shows two weighing scales.
  - One scale displays **24.00 g** with a coil of magnesium ribbon on it, representing 1 mole of Mg[cite: 92, 96].
  - The other scale displays **12.00 g** with a sample of powdered carbon in a petri dish, representing 1 mole of C[cite: 93, 103].
  - This visually demonstrates that one mole of different substances has a different mass.
- **Figure 4.4 (p. 60): Apparatus for determining the formula of magnesium oxide.**
  - **What it shows:** A setup for heating magnesium in a crucible to react it with oxygen from the air[cite: 371, 373].
  - **Labels:**
    - **Tripod:** Supports the apparatus over the heat source.
    - **Pipe clay triangle:** Sits on the tripod and holds the hot crucible.
    - **Crucible and lid:** A small ceramic pot with a lid, used to contain the magnesium ribbon while it is heated strongly. The lid prevents the fine magnesium oxide powder from

escaping[cite: 383].

- **Magnesium ribbon:** The reactant being heated.
- **Heat:** Arrow indicating heat applied from below (e.g., from a Bunsen burner).

6) Processes and cycles

Calculating Empirical Formula from Mass Data (p. 61)

This process is used to find the simplest formula of a compound, for example, magnesium oxide (MgO).

1. **Find the mass of each element:**
   - Mass of Mg = (Mass of crucible + Mg) - (Mass of crucible)[cite: 387].
   - Mass of O = (Mass of crucible + MgO) - (Mass of crucible + Mg)[cite: 388].
2. **Convert masses to moles:**
   - Moles of Mg = Mass of Mg / A_r of Mg (24)[cite: 389].
   - Moles of O = Mass of O / A_r of O (16)[cite: 390].
3. **Find the simplest whole-number ratio:**
   - Divide both mole values by the smallest of the two numbers[cite: 391].
   - The resulting numbers give the ratio of atoms in the empirical formula.

- **Inputs:** Mass of reactants.
- **Outputs:** Simplest whole-number ratio of atoms (empirical formula)[cite: 392].
- **Common error:** Forgetting to divide by the *smallest* number of moles to find the ratio.

Calculating Reacting Masses (p. 62-63)

This process determines the mass of product formed or reactant needed. Example: Heating limestone ($CaCO_3$) to make lime (CaO).

1. **Write a balanced chemical equation:**
   - $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$[cite: 472, 473].
2. **Calculate the molar masses (M_r)** of the substances involved.
   - M_r of $CaCO_3 = 40 + 12 + (3 \times 16) = 100$[cite: 474].
   - M_r of $CaO = 40 + 16 = 56$[cite: 477].
3. **Write the reacting masses under the equation** based on the mole ratio.

- The equation shows 1 mole of $CaCO_3$ produces 1 mole of CaO[cite: 474, 476].
- So, 100 g of $CaCO_3$ produces 56 g of CaO[cite: 474, 477].

4. **Use simple proportion (scaling)** to find the unknown mass.
    - If 100 tonnes of $CaCO_3$ produces 56 tonnes of CaO, then 10 tonnes will produce (10/100) × 56 = 5.6 tonnes of CaO[cite: 486, 487].

7) Formulae and calculations

Quantity	Formula	Units	Worked example (from text)
Number of moles (from mass)	$moles = \frac{mass}{molar\ mass}$	mol	Calculate moles in 108 g of Aluminium (Al, A_r=27): $moles = \frac{108}{27} = 4$ moles[cite: 164, 165].
Mass (from moles)	$mass = moles \times molar\ mass$	g	Calculate mass of 2 moles of iron (Fe, A_r=56): $mass = 2 \times 56 = 112$ g[cite: 133, 134].
Number of moles (gas volume)	$moles = \frac{volume\ at\ r.t.p.}{24\ dm^3}$	mol	Calculate moles in 72 dm^3 of ammonia (NH_3): $moles = \frac{72}{24} = 3$ moles[cite: 246, 247].
Gas volume (from moles)	$volume = moles \times 24\ dm^3$	dm^3	Calculate volume of 5 moles of CO_2: $volume = 5 \times 24 = 120\ dm^3$[cite: 253].
Concentration (from moles)	$concentration = \frac{moles}{volume}$	mol/dm^3	0.25 mol of NaOH in 250 cm^3 (0.25 dm^3) solution: $conc = \frac{0.25}{0.25} = 1\ mol/dm^3$[cite: 280].
Number of moles (in solution)	$moles = concentration \times volume$	mol	Moles of HCl in 40 cm^3 (0.04 dm^3) of 0.2 mol/dm^3 solution: $moles = 0.2 \times 0.04 = 0.008$ mol[cite: 541, 543].

Quantity	Formula	Units	Worked example (from text)
Percentage Yield	$\%Yield = \frac{actual\ yield}{theoretical\ yield} \times 100$	%	28 g of CO_2 was made when 44 g was expected: $\%Yield = \frac{28}{44} \times 100 = 63.6\%$[cite: 578, 579].
Percentage Purity	$\%Purity = \frac{mass\ of\ pure\ product}{mass\ of\ impure\ sample} \times 100$	%	An 84 g sample contained 80.5 g of pure substance: $\%Purity = \frac{80.5}{84} \times 100 = 95.8\%$[cite: 604].

8) Required practicals / experiments

Experiment to find the formula of magnesium oxide (p. 60-61)

- **Aim:** To determine the empirical formula of magnesium oxide by reacting a known mass of magnesium with oxygen.
- **Apparatus:** Crucible with lid, pipe clay triangle, tripod, Bunsen burner, heat-proof mat, weighing balance (p. 60, Figure 4.4).
- **Method:**
  i. Weigh an empty, clean crucible with its lid[cite: 385].
  ii. Add a coil of magnesium ribbon and reweigh the crucible, lid, and magnesium[cite: 385].
  iii. Place the crucible on the pipe clay triangle and heat it strongly with a Bunsen burner[cite: 373].
  iv. Periodically lift the lid slightly with tongs to allow oxygen (air) to enter, but replace it quickly to prevent the white magnesium oxide powder from escaping[cite: 378, 383].
  v. Heat until the magnesium has completely turned into a white powder[cite: 383].
  vi. Allow the crucible to cool completely, then reweigh it with the lid and contents[cite: 380].
  vii. Heat the crucible again for a few minutes, cool, and reweigh. Repeat this process (**heating to constant mass**) until two consecutive mass readings are the same, ensuring the reaction is complete[cite: 379, 381].
- **Variables:**
  - **Independent Variable (IV):** Mass of magnesium used.
  - **Dependent Variable (DV):** Mass of oxygen that reacts (and thus mass of magnesium oxide formed).
  - **Controlled Variables:** Heating time (initially), ensuring sufficient oxygen supply.

- **Safety:**
  - Wear eye protection at all times[cite: 364].
  - Do not look directly at the burning magnesium as it is very bright.
  - Handle the hot crucible and apparatus with tongs[cite: 365].
- **Sources of Error:**
  i. Some magnesium oxide powder may escape as smoke when the lid is lifted[cite: 383]. This would lead to a lower final mass and an incorrect Mg:O ratio.
  ii. The reaction may be incomplete, meaning some magnesium remains. This would lead to a lower mass of oxygen being recorded.
  iii. Magnesium can also react with nitrogen in the air to form magnesium nitride (Mg_3N_2), which would affect the final mass.
- **Expected Results:** The calculations using the masses should lead to a mole ratio of magnesium to oxygen of approximately 1:1, giving the empirical formula **MgO**[cite: 391, 392].

9) Data handling and graphing

This chapter primarily uses **tables** for data handling, especially for calculating empirical formulae.

- **What data is used:** Experimental masses of elements in a compound, or percentage composition data[cite: 358, 422].
- **Typical table structure for empirical formula calculations:**

	Element 1 (e.g., C)	**Element 2 (e.g., H)**
Mass (g) or %	Mass of C	Mass of H
Molar Mass (A_r)	12	1
Moles (Mass / A_r)	Moles of C	Moles of H
Mole Ratio	Moles of C / smallest moles	Moles of H / smallest moles
Simplest Ratio	Round to nearest whole number	Round to nearest whole number

- **Typical exam prompts:** "Use the data from the experiment to calculate the empirical formula of the compound." (p. 61, 68).

10) Common misconceptions and exam tips

- **Misconception:** The terms "relative molecular mass" and "molar mass" are the same.
  - **Correct understanding:** Relative molecular mass (M_r) has no units[cite: 177, 189]. Molar mass has units of **g/mol**[cite: 88, 176]. They are numerically equal, but conceptually different.
  - **Quick tip:** If the question asks for "mass," your answer needs units (g). If it asks for "M_r", there are no units.
- **Misconception:** The big numbers in front of a formula in an equation (coefficients) are part of the M_r calculation.
  - **Correct understanding:** The coefficient tells you the **number of moles**. The M_r is calculated for a single formula unit only. You multiply the M_r by the coefficient later to find the total reacting mass.
  - **Quick tip:** For $2H_2O$, first find M_r of H_2O (18), *then* multiply by 2 for the reacting mass (36 g)[cite: 117].
- **Misconception:** You can just use the ratio of masses from an experiment to find the formula.
  - **Correct understanding:** You MUST convert mass to **moles** first. Chemical formulae represent the ratio of atoms (moles), not the ratio of masses.
  - **Quick tip:** Always follow the steps: Mass → Moles → Ratio.
- **Misconception:** Forgetting to convert volumes in cm^3 to dm^3 for concentration calculations.
  - **Correct understanding:** Concentration is in mol/dm^3. All volumes must be in dm^3.
  - **Quick tip:** Remember **1 dm^3 = 1000 cm^3**. To convert $cm^3 \rightarrow dm^3$, divide by 1000. For 250 cm^3, use 0.250 dm^3 in your formula[cite: 264].

11) Exam-style practice

Multiple Choice Questions (MCQs)

1. What is the relative formula mass (M_r) of calcium carbonate, $CaCO_3$? (A_r: Ca=40, C=12, O=16)
   a) 56
   b) 68
   c) 100
   d) 116
   Answer: c) (Explanation: 40 + 12 + (3 × 16) = 100)

2. How many molecules are in 0.5 moles of carbon dioxide, CO_2?
   a) 6.02×10^{23}
   b) 3.01×10^{23}
   c) 1.204×10^{24}
   d) 6.02×10^{22}
   Answer: b) (Explanation: $0.5 \times (6.02 \times 10^{23}) = 3.01 \times 10^{23}$ [cite: 104])
3. What is the volume occupied by 0.25 moles of nitrogen gas at r.t.p.?
   a) 6 dm^3
   b) 12 dm^3
   c) 24 dm^3
   d) 96 dm^3
   Answer: a) (Explanation: $0.25 \times 24\ dm^3 = 6\ dm^3$)
4. A compound has an empirical formula of CH_2 and a relative molecular mass of 42. What is its molecular formula? (A_r: C=12, H=1)
   a) C_2H_4
   b) C_3H_6
   c) C_4H_8
   d) CH_2
   Answer: b) (Explanation: M_r of CH_2 is 14. 42 / 14 = 3. So the formula is $(CH_2)_3 = C_3H_6$)
5. In the reaction $2Mg + O_2 \rightarrow 2MgO$, which mass of oxygen reacts completely with 48 g of magnesium? (A_r: Mg=24, O=16)
   a) 16 g
   b) 32 g
   c) 48 g
   d) 64 g
   Answer: b) (Explanation: 48 g Mg is 2 moles. Ratio is 2 Mg : 1 O_2, so 1 mole of O_2 is needed. Mass of 1 mole O_2 = 32 g.)
6. What is the concentration of a solution containing 20 g of NaOH in 500 cm^3 of water? (M_r of NaOH = 40)
   a) 0.5 mol/dm^3
   b) 1.0 mol/dm^3
   c) 2.0 mol/dm^3
   d) 40 mol/dm^3
   Answer: b) (Explanation: Moles = 20/40 = 0.5 mol. Volume = 0.5 dm^3. Conc = 0.5/0.5 = 1.0 mol/dm^3.)
7. The simplest whole-number ratio of atoms in a compound is called the:
   a) Molecular formula
   b) Ionic formula

c) Structural formula

d) Empirical formula

Answer: d) (Explanation: This is the definition of empirical formula[cite: 395].)

8. In an experiment, the theoretical yield was 10 g, but the actual yield was 8 g. What is the percentage yield?

   a) 125%

   b) 100%

   c) 80%

   d) 20%

   Answer: c) (Explanation: $(8/10) \times 100 = 80\%$[cite: 581].)

9. Which reactant is the limiting reactant when 2 moles of Hydrogen (H_2) react with 2 moles of Oxygen (O_2) according to the equation $2H_2 + O_2 \rightarrow 2H_2O$?

   a) Hydrogen

   b) Oxygen

   c) Water

   d) Both are completely used

   Answer: a) (Explanation: 2 moles of H_2 require only 1 mole of O_2. Since there are 2 moles of O_2 available, the H_2 will run out first.)

10. What is the percentage by mass of magnesium in magnesium oxide, MgO? (A_r: Mg=24, O=16)

    a) 40%

    b) 50%

    c) 60%

    d) 67%

    Answer: c) (Explanation: M_r of MgO = 40. %Mg = $(24/40) \times 100 = 60\%$[cite: 588].)

Short-Answer Questions

1. **Define** the term 'mole'.
   - **Answer:** The mole is the unit for the amount of substance [cite: 64] which contains 6.02×10^{23} particles (atoms, ions or molecules)[cite: 64].
2. Calculate the mass of 0.2 moles of copper(II) sulfate, $CuSO_4$. (A_r: Cu=64, S=32, O=16)
   - **Answer:**
     - M_r of $CuSO_4 = 64 + 32 + (4 \times 16) = 160$.
     - Mass = moles $\times M_r = 0.2 \times 160 = 32$ g.
3. A sample of limestone ($CaCO_3$) is impure. When 12 g of the sample is heated, 4.8 g of lime (CaO) is produced. The equation is $CaCO_3 \rightarrow CaO + CO_2$. Calculate the percentage purity

of the limestone. (A_r: Ca=40, C=12, O=16)

- **Answer:**
  - M_r: $CaCO_3 = 100$, $CaO = 56$.
  - From the equation, 100 g of pure $CaCO_3$ produces 56 g of CaO.
  - Mass of pure $CaCO_3$ needed to make 4.8 g of CaO = $(100/56) \times 4.8 = 8.57$ g.
  - Percentage purity = (mass of pure / mass of impure) $\times 100 = (8.57/12) \times 100 = 71.4\%$.

4. Calculate the number of moles in 480 cm^3 of carbon dioxide gas at r.t.p.
   - **Answer:**
     - Convert volume to dm^3: 480 cm^3 / 1000 = 0.48 dm^3.
     - Moles = volume / 24 = 0.48 / 24 = 0.02 mol.
5. What is the difference between an empirical formula and a molecular formula?
   - **Answer:** The empirical formula shows the simplest whole-number ratio of atoms in a compound [cite: 395], while the molecular formula shows the actual number of atoms of each element in one molecule[cite: 421].

Structured Questions

Question 1

An organic compound, **X**, was found to contain 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen by mass.

a) **Calculate** the empirical formula of compound **X**. (A_r: C=12, H=1, O=16) [3]

b) The relative molecular mass, M_r, of compound **X** is 180. **Determine** the molecular formula of **X**. [2]

- **Marking points (a):**
  - **(1)** Divide % by A_r for each element:
    - C: 40.0 / 12 = 3.33
    - H: 6.7 / 1 = 6.7
    - O: 53.3 / 16 = 3.33
  - **(1)** Divide by the smallest value (3.33) to find the ratio:
    - C: 3.33 / 3.33 = 1
    - H: 6.7 / 3.33 = 2
    - O: 3.33 / 3.33 = 1
  - **(1)** State the empirical formula: CH_2O.
- **Marking points (b):**
  - **(1)** Calculate the mass of the empirical formula unit: $12 + (2 \times 1) + 16 = 30$.

- **(1)** Find the multiple (n) and the molecular formula:
  - n = (M_r of molecule) / (M_r of empirical formula) = 180 / 30 = 6.
  - Molecular formula = $(CH_2O)_6 = C_6H_{12}O_6$.

Question 2
Iron(III) oxide reacts with carbon monoxide in a blast furnace to produce iron.
$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
a) **Calculate** the maximum mass of iron that can be produced from 800 tonnes of iron(III) oxide (Fe_2O_3). (A_r: Fe=56, O=16) [3]
b) In the reaction, 2.4 g of iron(III) oxide was heated with excess carbon monoxide, and 1.4 g of iron was produced. **Calculate** the percentage yield. [2]

- **Marking points (a):**
  - **(1)** Calculate molar masses:
    - M_r of $Fe_2O_3 = (2 \times 56) + (3 \times 16) = 112 + 48 = 160.$
    - A_r of Fe = 56. The equation has 2Fe, so mass is $2 \times 56 = 112.$
  - **(1)** State the mass ratio from the equation: 160 tonnes of Fe_2O_3 produces 112 tonnes of Fe.
  - **(1)** Calculate the mass of iron produced from 800 tonnes:
    - Mass Fe = $(112/160) \times 800 = 560$ tonnes.
- **Marking points (b):**
  - **(1)** Calculate the theoretical yield of iron from 2.4 g of Fe_2O_3:
    - Theoretical Mass Fe = $(112/160) \times 2.4 = 1.68$ g.
  - **(1)** Calculate the percentage yield:
    - % Yield = (actual yield / theoretical yield) $\times 100 = (1.4/1.68) \times 100 = 83.3\%.$

12) Quick revision checklist

- [ ] I can define Relative Atomic Mass (A_r) based on the carbon-12 scale.
- [ ] I can calculate the Relative Molecular/Formula Mass (M_r) of any given compound.
- [ ] I can define the mole and state the value of the Avogadro constant (6.02×10^{23}).
- [ ] I can convert between mass, moles, and molar mass for any substance.
- [ ] I can use the molar gas volume (24 dm^3 at r.t.p.) to convert between moles and volume for any gas.
- [ ] I can calculate the concentration of a solution in both mol/dm^3 and g/dm^3.
- [ ] I can calculate the number of moles of solute in a solution of known volume and concentration.
- [ ] I can determine the empirical formula of a compound from mass or percentage composition data.

- [ ] I can find the molecular formula of a compound using its empirical formula and M_r.
- [ ] I can use a balanced chemical equation to calculate reacting masses and volumes.
- [ ] I can identify the limiting reactant in a reaction and use it to calculate the amount of product.
- [ ] I can calculate the percentage yield and percentage purity of a product.

13) Flashcards (ready-to-use)

Question	Answer
Q1: What is relative atomic mass, A_r?	The average mass of an element's isotopes compared to 1/12th the mass of a carbon-12 atom[cite: 24].
Q2: What is the value of the Avogadro constant?	6.02×10^{23} particles per mole[cite: 65, 81].
Q3: What is the formula linking mass, moles, and molar mass?	$moles = \frac{mass}{molar\ mass}$[cite: 151, 152].
Q4: What volume does one mole of any gas occupy at r.t.p.?	24 dm^3[cite: 233].
Q5: What is the unit for concentration in chemistry?	mol/dm^3 (moles per cubic decimetre)[cite: 261].
Q6: How do you convert a volume from cm^3 to dm^3?	Divide by 1000[cite: 264].
Q7: What is an empirical formula?	The simplest whole-number ratio of atoms in a compound[cite: 395].
Q8: What two pieces of information are needed to find a molecular formula from an empirical formula?	The empirical formula and the relative molecular mass (M_r)[cite: 407].
Q9: What does the coefficient (big number) in a balanced equation represent?	The mole ratio of the reactants and products[cite: 444].
Q10: What is the formula for percentage yield?	$\%Yield = \frac{actual\ yield}{theoretical\ yield} \times 100$[cite: 581].

Question	Answer
Q11: What is the limiting reactant?	The reactant that is completely used up first and determines how much product is made[cite: 609, 610].
Q12: What is the molar mass of H_2O? (A_r: H=1, O=16)	18 g/mol[cite: 175].
Q13: What is the difference between relative molecular mass (M_r) and molar mass?	M_r has no units; molar mass has units of g/mol[cite: 88, 177].
Q14: What is the purpose of heating to a constant mass in an experiment?	To ensure the reaction has gone to completion[cite: 379].
Q15: State the Law of Conservation of Mass.	The total mass of the reactants is equal to the total mass of the products[cite: 452, 453].

14) 60-second recap

This chapter is all about the **mole**, the chemist's unit for counting particles. We start with **relative atomic mass (A_r)** to compare the masses of different atoms. We learn that one mole of any substance contains 6.02×10^{23} particles and its mass in grams is its **molar mass**. Using this, we can convert between mass and moles for solids, between volume and moles for gases (using 24 dm^3), and between concentration, volume, and moles for solutions. Balanced equations give us the mole ratio, allowing us to calculate reacting masses. We can also work backwards from experimental mass data to find a compound's **empirical and molecular formula**. Finally, we assess a reaction's efficiency using **percentage yield** and determine a sample's purity.

15) References to pages

- **Avogadro's Constant/Law:** 55, 59
- **Concentration:** 59, 64
- **Empirical Formula:** 60, 61, 62
- **Experiments (MgO):** 60, 61
- **Limiting Reactant:** 66

- **Mole Concept:** 55, 57, 62
- **Molecular Formula:** 55, 61, 62
- **Molar Gas Volume:** 58
- **Percentage Composition/Purity/Yield:** 65
- **Reacting Mass/Volume Calculations:** 55, 62, 63, 64
- **Relative Atomic/Molecular/Formula Mass:** 54, 55

16) Excluded "Going further" sections (not summarized)

Section title	Pages
None present in the provided chapter.	-
Total excluded:	0