

Class 11 Physical Chemistry - Free Preview

A concept-first route to numericals, mole concept and stoichiometry

- Use the six-line numerical protocol instead of formula hunting.
- See how the mole concept connects mass, particles and chemical amount.
- Learn stoichiometry and limiting-reagent reasoning step by step.

Selected preview - read, evaluate, then choose the complete book

How to Use This Book

Use the chapter in four passes rather than reading from beginning to end only once.

1 • BUILD THE MODEL

Read the theory, definitions, diagrams and derivations. State the condition beside every important equation.

2 • READ THE GRAPH

Identify axes, slope, intercept, limiting behaviour and the physical meaning of every region or curve.

3 • SOLVE BEFORE YOU SEE

Cover the worked solution, predict the governing principle and first step, then compare your method.

4 • DIAGNOSE ERRORS

Classify each mistake and retry it after a delay. Do not turn an answer key into passive reading.

The six-line numerical protocol

- | | | |
|----------|---------------------|---|
| 1 | TRANSLATE | Write the physical/chemical situation in words. |
| 2 | LIST DATA | Convert units and keep constants visible. |
| 3 | CHOOSE MODEL | State the law, equation or equilibrium relation and its condition. |
| 4 | SOLVE | Substitute step by step; do not skip dimensions or signs. |
| 5 | AUDIT | Check direction, limiting case, magnitude and physical possibility. |
| 6 | LEARN | If wrong, identify the cause and reattempt without the solution. |

(A) 12 (B) 7 (C) 4 (D) 3

MCQ 7.9 If the fraction of a heavier isotope increases, average atomic mass generally

(A) must become an integer (B) increases (C) stays fixed by the mass number (D) decreases

MCQ 7.10 Relative atomic mass has

(A) no unit (B) unit g (C) unit u (D) unit mol

BRAIN TWISTER — five reasoning challenges

BT 7.1 Two isotope masses are 10 and 11 u. Can the average be exactly 10 u if both isotopes have non-zero abundance?

BT 7.2 An element has isotope masses 20, 21 and 22 u and average 21 u. Does that prove all isotopes are equally abundant?

BT 7.3 A sample contains equal masses, not equal numbers, of 10 u and 20 u isotope atoms. Find average atomic mass.

BT 7.4 A hydrated salt has molar mass 250 g mol^{-1} ; its anhydrous salt is 160 g mol^{-1} . Find waters of crystallisation.

BT 7.5 Isotopes 35 and 37 u have average 35.8 u. What is the ratio of their numbers?

Answers and reasoning for Topic 1.7 appear in the solutions section.

1.8 Mole concept: mass, particles, atoms and gas volume

The mole is a counting amount, not a mass

A dozen means twelve objects. A mole similarly specifies a number of entities, but the number is enormous: one mole contains exactly $6.02214076 \times 10^{23}$ specified elementary entities. The Avogadro constant N_a has unit mol^{-1} . It connects an amount n in moles with an entity count N through $N = nN_a$.

The entity must be specified: atoms, molecules, ions, electrons or formula units. One mole of O atoms and one mole of O_2 molecules contain the same number of named entities but different numbers of oxygen atoms. One mole of O_2 molecules contains two moles of O atoms.

Equal moles do not mean equal masses. One mole of helium atoms has mass about 4 g, whereas one mole of carbon dioxide molecules has mass about 44 g. Both contain N_a named entities. Conversely, equal masses of different substances usually contain different numbers of entities because their molar masses differ.

Deriving the main conversion relationships

Molar mass is defined by $M = \text{mass}/\text{amount}$. Rearrangement gives $n = \text{mass}/M$ and $\text{mass} = nM$. Combining this with $N = nN_a$ gives $N = (\text{mass}/M)N_a$ and $\text{mass} = (N/N_a)M$. These are connected forms of the same two definitions; memorising them as unrelated formulas hides the logic.

A dependable route is given quantity $\rightarrow$ moles of the specified substance $\rightarrow$ required quantity. Start with the chemical formula and molar mass, not with the largest-looking number in the problem. If mass is in grams, use M in g mol^{-1} . If mass is in kilograms, either convert it or use kg mol^{-1} .

In this chapter use $N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$ for ordinary numerical work unless a question supplies another value. The definition is exact, but the numerical approximation used in arithmetic is not the limiting uncertainty in most school problems. Rounded atomic masses supplied for a question are the intended working values.

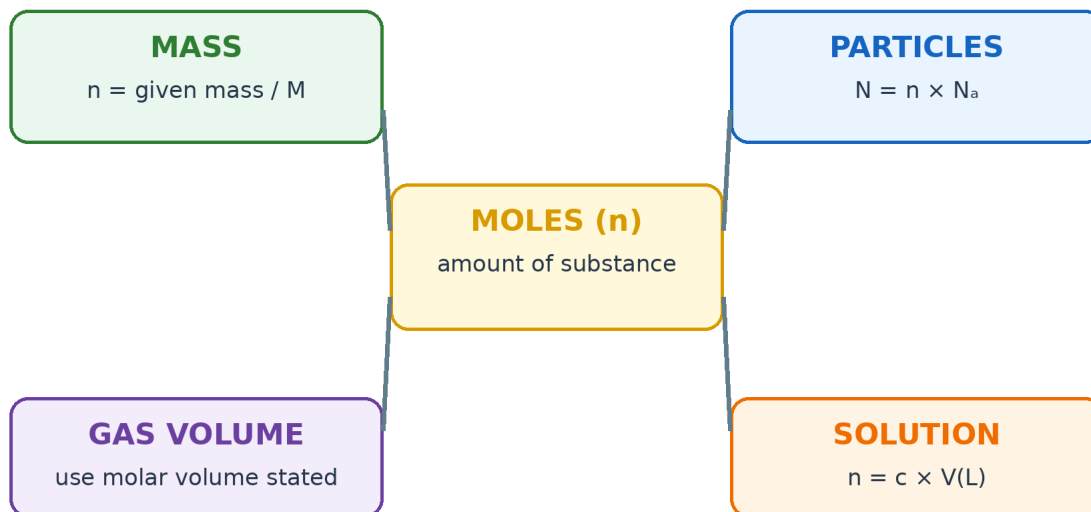


Figure 1.4 • The mole as the conversion hub between mass, entities, gas volume and solution amount.

Amount in moles connects mass, entities, gas volume and solution concentration.

Counting atoms, ions and electrons

First count the entities named by the chemical formula; then multiply by the number of the requested constituent per entity. For 0.20 mol CO_2 molecules, carbon atoms amount to 0.20 mol and oxygen atoms to 0.40 mol. Total atoms amount to 0.60 mol. The phrase “total atoms” includes every element.

For ionic compounds, count formula units and the ions represented by each formula unit. Na_2SO_4 corresponds to two Na^+ ions and one SO_4^{2-} ion per formula unit. If the problem assumes complete dissociation, 0.10 mol Na_2SO_4 gives 0.20 mol Na^+ , 0.10 mol SO_4^{2-} and 0.30 mol ions in total. It does not give separate sulfur and oxygen ions.

For electron-count questions, use the total atomic numbers for a neutral entity and adjust for net charge. A neutral H_2O molecule has $2(1) + 8 = 10$ electrons. Na^+ has one electron fewer than a neutral sodium atom. Electron counting is included here as a labelled extension connecting amount with atomic structure.

Gas volumes and mixture counting

For an ideal gas at specified temperature and pressure, amount can be obtained from $n = V/V_m$, where V_m is the molar volume at those conditions. The commonly used 22.4 L mol^{-1} is an approximation near 273.15 K and 1 atm. A value near 22.7 L mol^{-1} applies at 273.15 K and 1 bar. Always use the conditions or molar volume explicitly supplied; avoid treating “standard” as an unchanging number.

In a mixture, add amounts before adding entity counts. A 4 g He plus 16 g O_2 mixture contains 1 mol He atoms and 0.5 mol O_2 molecules. That is 1.5 mol gas entities but 2 mol

atoms in total. This example shows why “particles” and “atoms” cannot be used interchangeably.

To compare molecule numbers for equal masses, compare $1/M$. To compare total atoms, compare a/M , where a is the number of atoms per molecule. A lighter molecule does not automatically give the most atoms per gram if it contains fewer atoms. Write the actual relationship before deciding.

Memory aid — after understanding

Convert to moles first. Then multiply by the number of the requested entity per molecule or formula unit.

Worked examples: method and reasoning

WORKED EXAMPLE 8.1

Find moles and molecules in 8.80 g CO_2 ($M = 44.0 \text{ g mol}^{-1}$).

$$n = 8.80/44.0 = 0.200 \text{ mol.}$$

$$N = 0.200 \times 6.022 \times 10^{23} = 1.2044 \times 10^{23} \text{ molecules.}$$

Answer: 0.200 mol; 1.20×10^{23} molecules.

WORKED EXAMPLE 8.2

Find the mass of 3.011×10^{22} NH_3 molecules ($M = 17.0$).

$$n = N/N_a = (3.011 \times 10^{22})/(6.022 \times 10^{23}) = 0.05000 \text{ mol.}$$

$$\text{mass} = nM = 0.05000 \times 17.0 = 0.850 \text{ g.}$$

Answer: 0.850 g NH_3 .

WORKED EXAMPLE 8.3

In 0.250 mol CO_2 , find moles of O atoms and the total number of atoms.

$$\text{O atoms: } 2 \times 0.250 = 0.500 \text{ mol. Total atoms: } 3 \times 0.250 = 0.750 \text{ mol.}$$

$$\text{Total atom count} = 0.750 \times 6.022 \times 10^{23} = 4.5165 \times 10^{23}.$$

Answer: 0.500 mol O atoms; 4.52×10^{23} total atoms.

WORKED EXAMPLE 8.4

Find moles of all atoms in 0.100 mol $\text{Al}_2(\text{SO}_4)_3$.

$$\text{Each formula unit contains } 2 + 3 + 12 = 17 \text{ atoms.}$$

$$\text{Amount of atoms} = 17 \times 0.100 = 1.70 \text{ mol. By element: Al } 0.200 \text{ mol, S } 0.300 \text{ mol, O } 1.20 \text{ mol.}$$

Answer: 1.70 mol total atoms.

WORKED EXAMPLE 8.5

5.85 g NaCl ($M = 58.5$) dissolves completely. Find Na^+ ions, Cl^- ions and total ions.

$$\text{Amount NaCl} = 5.85/58.5 = 0.100 \text{ mol. Each formula unit corresponds to one ion of each type.}$$

$$\text{Na}^+ \text{ count} = \text{Cl}^- \text{ count} = 0.100N_a = 6.022 \times 10^{22}. \text{ Total ion count} = 0.200N_a.$$

Answer: 6.02×10^{22} of each ion; 1.20×10^{23} ions altogether.

WORKED EXAMPLE 8.6

A molecular gas occupies 5.60 L under conditions where $V_m = 22.4 \text{ L mol}^{-1}$. Find amount and molecules.

$$n = V/V_m = 5.60/22.4 = 0.250 \text{ mol.}$$

$$N = 0.250N_a = 1.5055 \times 10^{23}.$$

Answer: 0.250 mol; 1.51×10^{23} molecules.

WORKED EXAMPLE 8.7

Compare total atoms in 4.00 g He and 4.00 g H₂ using molar masses 4.00 and 2.00.

He: $4.00/4.00 = 1.00$ mol atoms.

H₂: $4.00/2.00 = 2.00$ mol molecules, each with two atoms, so 4.00 mol atoms.

Answer: H₂ contains four times as many atoms.

WORKED EXAMPLE 8.8

A mixture contains 4.0 g He and 16 g O₂. Find moles of gas entities and moles of all atoms.

He: 1.0 mol monatomic entities. O₂: 0.50 mol molecules. Gas entities total = 1.50 mol.

O atoms from O₂ = 1.0 mol; add 1.0 mol He atoms for 2.0 mol total atoms.

Answer: 1.5 mol gas entities; 2.0 mol atoms.

WORKED EXAMPLE 8.9

Extension: find the number of electrons in 0.100 mol neutral H₂O molecules; atomic numbers H = 1, O = 8.

Each H₂O molecule has $2(1) + 8 = 10$ electrons.

Amount of electrons = $10 \times 0.100 = 1.00$ mol; number = $1.00N_a$.

Answer: 6.02×10^{23} electrons.

Practice: build fluency, then apply

P 8.1 Find moles in 22.0 g CO₂.

P 8.2 Find molecules in 3.60 g H₂O.

P 8.3 Find mass of 0.750 mol NH₃.

P 8.4 Find oxygen atoms in 0.20 mol H₂SO₄.

P 8.5 Find total atoms in 0.0500 mol Ca₃(PO₄)₂.

P 8.6 Find formula units in 14.2 g Na₂SO₄ (M = 142).

P 8.7 For that Na₂SO₄ sample, find total ions assuming complete dissociation.

P 8.8 A gas sample occupies 3.00 L at $V_m = 24.0 \text{ L mol}^{-1}$. Find amount.

P 8.9 What mass of O₂ contains the same molecules as 4.00 g CH₄?

P 8.10 Compare molecules in 9.0 g H₂O and 22 g CO₂.

P 8.11 A mixture contains 0.20 mol CH₄ and 0.30 mol CO₂. Find moles of total atoms.

P 8.12 Find mass of oxygen atoms present in 0.500 mol CO₂.

Topic MCQs: concept and calculation

MCQ 8.1 Number of molecules in 11 g CO₂ is

(A) $0.50N_a$ (B) N_a (C) $0.25N_a$ (D) $4N_a$

MCQ 8.2 Total atoms in 0.50 mol NH₃ are

(A) $1.5N_a$ (B) $2.0N_a$ (C) $0.50N_a$ (D) N_a

MCQ 8.3 Oxygen atom count in 0.20 mol Al₂(SO₄)₃ is

(A) $0.60N_a$ (B) $0.80N_a$ (C) $2.4N_a$ (D) $3.4N_a$

MCQ 8.4 Total ions from 0.10 mol CaCl₂ on complete dissociation are

BT 9.3 A CHO compound of mass 0.460 g yields 0.880 g CO₂ and 0.540 g H₂O. Find empirical formula.

BT 9.4 A hydrate loses 45.0% mass as water. Its anhydrous molar mass is 110. Find x.

BT 9.5 A compound has empirical formula CH₂O but a reported molar mass of 75 g mol⁻¹. Why should the data be rechecked?

Answers and reasoning for Topic 1.9 appear in the solutions section.

1.10 Stoichiometry, limiting reagents, purity and yield

A balanced equation is a quantitative relationship

Stoichiometry relates reactant and product amounts through a balanced chemical equation. Begin by writing correct formulas and balancing atoms of each element. For $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, one mole of N₂ reacts with three moles of H₂ to form two moles of NH₃ when the reaction proceeds as written.

The coefficients do not give direct mass ratios. Using N₂ = 28, H₂ = 2 and NH₃ = 17 g mol⁻¹, the corresponding mass relationship is $28 \text{ g N}_2 + 6 \text{ g H}_2 \rightarrow 34 \text{ g NH}_3$. This mass balance is consistent with conservation. A student who reads the equation as $1 \text{ g} + 3 \text{ g} \rightarrow 2 \text{ g}$ has confused amounts with masses.

For $a\text{A} + b\text{B} \rightarrow c\text{C}$, the reacted or formed amounts satisfy $n(\text{A})/a = n(\text{B})/b = n(\text{C})/c$. These are amounts participating in the reaction, not necessarily all amounts initially placed in the vessel. When reactants are not supplied in the exact ratio, some material can remain.

The five-step calculation route

First balance the equation. Second convert the given quantity to moles of the named substance. Third use the coefficient ratio to find moles of the required substance. Fourth convert to the requested mass, particles or gas volume. Fifth check units, conservation and reasonable magnitude.

This route handles mass-to-mass, mass-to-gas, particle-to-mass and gas-to-gas questions. The middle link is always the chemical mole ratio. The first and last conversions vary with the data. Keep those physical conversions separate from the chemical equation ratio.

A reaction written in reverse must also reverse the interpretation of reactant and product amounts. Multiplying every coefficient by a common factor leaves all ratios unchanged. Neither operation changes the molar masses of the substances.

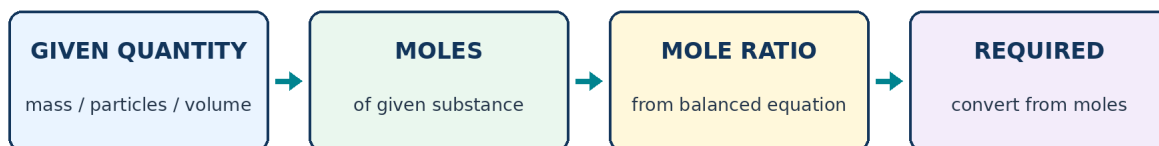


Figure 1.6 • General stoichiometric route: given quantity → moles → mole ratio → required quantity.

The balanced-equation mole ratio is the chemical link in the calculation.

How to identify a limiting reagent

The limiting reagent is exhausted when the reaction proceeds as far as the available reactant amounts permit under the stated model. It determines the maximum product. The excess reagent is supplied beyond the amount required to react with the limiter. “Excess” does not mean that none of it reacts.

A safe method is to calculate the product that each reactant could form if all others were available. The smaller possible product identifies the limiting reagent. An equivalent method is to compare n_i divided by its stoichiometric coefficient. The smallest value gives the maximum reaction extent.

LIMITING-REAGENT CHECK Convert each reactant to the amount of the same product (or compare $n/\text{coefficient}$). The reactant predicting the smaller reaction extent is limiting. Do not choose the smaller mass by inspection.

Do not choose the limiter by smallest mass, smallest mole amount or smallest coefficient alone. For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, 1 mol O_2 needs 2 mol H_2 . If 1.5 mol H_2 and 1 mol O_2 are supplied, hydrogen limits even though the oxygen mole amount is smaller.

To find excess material left, first determine how much of it reacts with the limiting reagent. Subtract that consumed amount from its initial amount. The remaining mass is part of the final mass balance and must not be counted as product.

Purity changes the available reacting amount

A commercial sample can contain inert impurities. If a sample is $p\%$ pure by mass, pure reacting mass = sample mass $\times p/100$. Convert that pure mass, not the whole sample mass, to moles. This assumes impurities do not participate in the reaction being calculated.

For a required product amount, work backwards through the balanced equation to obtain the pure reactant mass. Divide by the purity fraction to obtain the commercial sample mass. Multiplying in this reverse situation would make the impure sample smaller than the pure amount required, which is physically unreasonable.

Theoretical yield, actual yield and percentage yield

Theoretical yield is the maximum product predicted from the limiting reagent and reaction stoichiometry. Actual yield is the product obtained under the experimental conditions. Percentage yield = actual/theoretical $\times 100$. Use the same units and chemical product in numerator and denominator.

A lower actual yield can result from incomplete reaction, side reactions or loss during isolation. A reported yield above 100% usually signals issues such as wet or impure product, an incorrect assumed formula or measurement error; it does not show that more pure product was formed than atom conservation permits.

Purity and yield answer different questions. Purity describes the composition of the starting sample or collected product. Reaction yield describes the fraction of the theoretical product recovered. Apply them at the correct stages; if both are present, identify the pure reacting amount first, determine theoretical product, then apply the yield fraction.

Purity, yield and linked-reaction questions here are extensions where appropriate. They deepen the same mole reasoning without replacing the core concepts.

Memory aid — after understanding

Balance → convert to moles → identify limiter if needed → use ratio → convert units. Apply purity before theoretical yield.

Worked examples: method and reasoning**WORKED EXAMPLE 10.1**

Find O₂ mass required to burn 8.00 g CH₄: CH₄ + 2O₂ → CO₂ + 2H₂O.

$n(\text{CH}_4) = 8.00/16.0 = 0.500$ mol. The coefficient ratio requires 2 mol O₂ per mol CH₄.

$n(\text{O}_2) = 1.00$ mol; mass = $1.00 \times 32.0 = 32.0$ g.

Answer: 32.0 g O₂.

WORKED EXAMPLE 10.2

Find CaO mass from complete decomposition of 25.0 g pure CaCO₃: CaCO₃ → CaO + CO₂. Use molar masses 100 and 56.

$n(\text{CaCO}_3) = 25.0/100 = 0.250$ mol. The CaCO₃:CaO ratio is 1:1.

$n(\text{CaO}) = 0.250$ mol; mass = $0.250 \times 56 = 14.0$ g. CO₂ mass = 11.0 g, providing a mass-balance check.

Answer: 14.0 g CaO.

WORKED EXAMPLE 10.3

Find O₂ volume from 12.25 g KClO₃: 2KClO₃ → 2KCl + 3O₂. M(KClO₃) = 122.5; V_m = 22.4 L mol⁻¹.

$n(\text{KClO}_3) = 12.25/122.5 = 0.1000$ mol.

$n(\text{O}_2) = 0.1000 \times 3/2 = 0.1500$ mol; V = 0.1500×22.4 .

Answer: 3.36 L O₂.

WORKED EXAMPLE 10.4

3.0 mol N₂ and 6.0 mol H₂ react by N₂ + 3H₂ → 2NH₃. Find limiter, NH₃ and excess left.

From N₂, possible NH₃ = $3.0 \times 2 = 6.0$ mol. From H₂, possible NH₃ = $6.0 \times 2/3 = 4.0$ mol. Hydrogen limits.

Nitrogen consumed = $6.0/3 = 2.0$ mol; nitrogen left = $3.0 - 2.0 = 1.0$ mol.

Answer: H₂ limiting; 4.0 mol NH₃; 1.0 mol N₂ left.

WORKED EXAMPLE 10.5

5.00 g H₂ and 32.0 g O₂ react to form water. Use H₂ = 2.00, O₂ = 32.0, H₂O = 18.0.

Initial amounts: H₂ = 2.50 mol, O₂ = 1.00 mol. Compare n/coefficient: H₂ 1.25; O₂ = 1.00. Oxygen limits.

Water formed = 2.00 mol = 36.0 g. Hydrogen consumed = 2.00 mol = 4.00 g; hydrogen left = 1.00 g.

Check total mass: 37.0 g initially; 36.0 g water plus 1.00 g hydrogen finally.

Answer: 36.0 g water; 1.00 g H₂ remains.

WORKED EXAMPLE 10.6

A 20.0 g limestone sample is 75.0% CaCO₃, with inert impurities. Find theoretical CO₂ mass on complete heating (CaCO₃ = 100, CO₂ = 44).

Pure CaCO₃ = $20.0 \times 0.750 = 15.0$ g = 0.150 mol.

CO_2 amount = 0.150 mol; mass = $0.150 \times 44 = 6.60$ g.

Answer: 6.60 g CO_2 .

WORKED EXAMPLE 10.7

A reaction theoretically produces 12.0 g product but gives 9.00 g pure product. Find percentage yield.

Yield = actual/theoretical $\times 100$.

$9.00/12.0 \times 100 = 75.0\%$.

Answer: 75.0%.

WORKED EXAMPLE 10.8

What mass of 80.0%-pure CaCO_3 is needed to obtain 8.80 g CO_2 at 50.0% yield? Use $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$.

Required theoretical $\text{CO}_2 = 8.80/0.500 = 17.6$ g = 0.400 mol.

Required pure $\text{CaCO}_3 = 0.400 \times 100 = 40.0$ g. Commercial sample = $40.0/0.800 = 50.0$ g.

Answer: 50.0 g impure CaCO_3 sample.

WORKED EXAMPLE 10.9

10.0 g NaOH reacts with 7.30 g HCl: $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$. Use $M = 40.0, 36.5, 58.5$. Find NaCl and excess reactant.

$n(\text{NaOH}) = 0.250$ mol; $n(\text{HCl}) = 0.200$ mol. The ratio is 1:1, so HCl limits.

$\text{NaCl} = 0.200 \times 58.5 = 11.7$ g. NaOH left = $(0.250 - 0.200) \times 40.0 = 2.00$ g.

Answer: 11.7 g NaCl; 2.00 g NaOH left.

WORKED EXAMPLE 10.10

Extension: 20.0 g CaCO_3 gives CaO, which reacts completely with water: $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$. Find final product mass if both steps have complete conversion.

First reaction gives $n(\text{CaO}) = 20.0/100 = 0.200$ mol.

Second ratio $\text{CaO}:\text{Ca}(\text{OH})_2 = 1:1$, so product = 0.200 mol. $M[\text{Ca}(\text{OH})_2] = 74$; mass = 14.8 g.

Product mass can differ from initial carbonate mass because CO_2 leaves in step 1 and water enters in step 2.

Answer: 14.8 g $\text{Ca}(\text{OH})_2$.

Practice: build fluency, then apply

P 10.1 Find CO_2 mass from burning 4.00 g CH_4 with excess O_2 .

P 10.2 Find H_2 required for 14.0 g N_2 in ammonia formation.

P 10.3 Find NH_3 mass from 3.00 g H_2 with excess N_2 .

P 10.4 Find O_2 moles from 0.400 mol KClO_3 in $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$.

P 10.5 Find CaO mass from 40.0 g CaCO_3 .

P 10.6 Find limiter and water moles for 3 mol H_2 plus 2 mol O_2 .

P 10.7 Find NH_3 moles and excess left for 2 mol N_2 plus 9 mol H_2 .

P 10.8 A 50.0 g sample is 90.0% CaCO_3 . Find theoretical CO_2 mass.

P 10.9 Theoretical yield is 25.0 g; actual is 18.0 g. Find percentage yield.

P 10.10 Theoretical product is 40.0 g at 65.0% yield. Find actual product.

P 10.11 What mass of 70%-pure CaCO_3 contains 35 g pure CaCO_3 ?

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