Exercise walk-through

3.2
Moles – amount of substance and reacting masses (4.3.2)

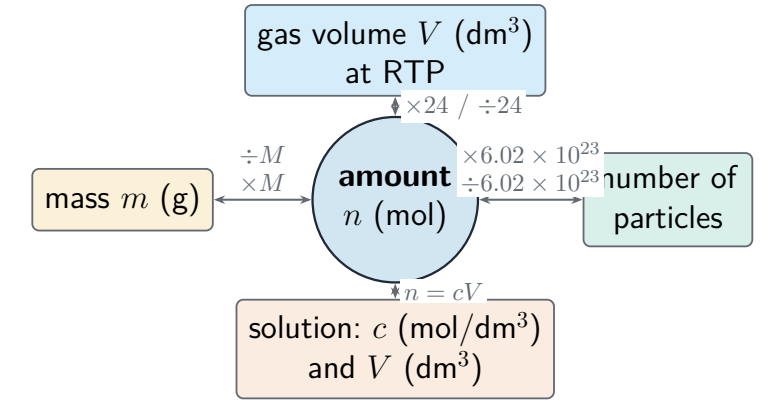
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Moles and reacting masses (higher tier)

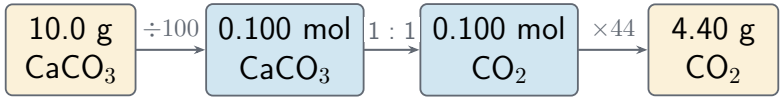
- Atoms are far too small to count one by one, so chemists count them in moles.
- One **mole** of a substance contains 6.02×10^{23} particles: this number is the **Avogadro constant**.
- The mass of one mole in grams is equal to the Mr, so one mole of carbon atoms (Ar 12) weighs 12 g and one mole of water (Mr 18) weighs 18 g.
- This mass per mole is the **molar mass**, in g/mol.

$$\text{moles} = \frac{\text{mass (g)}}{M_r}$$

Figure: Moles and reacting masses (higher tier)



worked chain: $CaCO_3 \rightarrow CaO + CO_2$



the equation ratio links amounts in moles, never masses

- Mass $\div$ M gives moles; moles $\times$ M gives mass.
- Use the equation's ratio in moles.

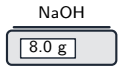
Concentrations in mol/dm^3 (chemistry only, higher tier)

- **Molar concentration** is the number of moles of **solute** in each dm^3 of solution:
- To change it into **mass concentration** , multiply by the Mr: 0.20 mol/dm^3 NaOH is $0.20 \times 40 = 8.0$ g/dm^3.
- In a **titration** , you measure the exact volumes of an acid and an alkali that react.
- If you know one **concentration**, you can calculate the other, using the mole ratio in the equation.

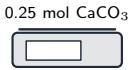
$$\text{moles} = \text{concentration (mol/dm}^3\text{)} \times \text{volume (dm}^3\text{)}$$

Exercises 2.1·2.3

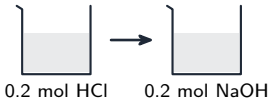
2.1 How many moles are in 8.0 g of sodium hydroxide, NaOH (Na = 23, O = 16, H = 1)? [2]



2.2 What mass is 0.25 mol of $CaCO_3$ (Mr = 100)? [1]



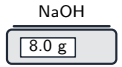
2.3 0.2 mol of HCl is mixed with 0.2 mol of NaOH. Why does neither remain in excess? [1]



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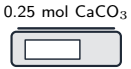
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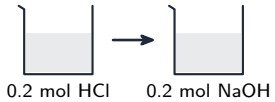
2.1 Mr(NaOH) = 40; moles = 8.0 ÷ 40 = 0.2.

2.2 What mass is 0.25 mol of $CaCO_3$ (Mr = 100)? [1]



2.2 0.25 × 100 = 25 g.

2.3 0.2 mol of HCl is mixed with 0.2 mol of NaOH. Why does neither remain in excess? [1]

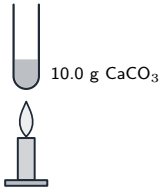


2.3 The equation ratio NaOH : HCl is 1 : 1, and the moles are equal, so both are fully used.

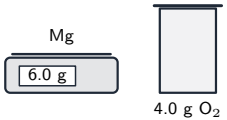
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Exercises 3.1–3.2

3.1 Calculate the mass of carbon dioxide produced when 10.0 g of calcium carbonate decomposes: $CaCO_3 \rightarrow CaO + CO_2$ (Ca = 40, C = 12, O = 16). [4]

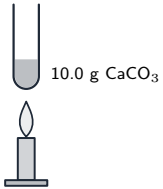


3.2 6.0 g of magnesium reacts with 4.0 g of oxygen: $2\ Mg + O_2 \rightarrow 2\ MgO$. Identify the limiting reactant, if any, and calculate the mass of MgO formed (Mg = 24, O = 16). [5]



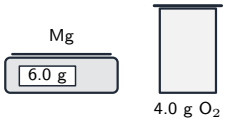
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3.1 $Mr(CaCO_3) = 100$, moles $= 10.0 \div 100 = 0.10$; ratio 1 : 1, so 0.10 mol CO_2; $Mr(CO_2) = 44$; mass $= 0.10 \times 44 = 4.4$ g.

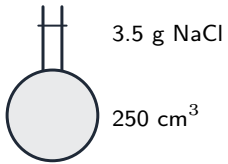
3.2 6.0 g of magnesium reacts with 4.0 g of oxygen: $2\ Mg + O_2 \rightarrow 2\ MgO$. Identify the limiting reactant, if any, and calculate the mass of MgO formed (Mg = 24, O = 16). [5]



3.2 Moles Mg $= 6.0 \div 24 = 0.25$; moles $O_2 = 4.0 \div 32 = 0.125$; ratio needs 2 Mg per 1 O_2, so 0.125 mol O_2 requires 0.25 mol Mg – exactly what is present: neither is limiting / both fully react; moles MgO = 0.25; mass $= 0.25 \times 40 = 10.0$ g.

Setup for exercise 3.3

3.3 *(HT)* 250 cm^3 of sodium chloride solution contains 3.5 g of solute. Calculate the concentration in g/dm^3 and in mol/dm^3 (Na = 23, Cl = 35.5). [4]



Exercise 3.3

3.4 A reaction uses 4.8 g Mg and 6.4 g O_2. Use $2Mg + O_2 \rightarrow 2MgO$, Mg = 24 and O = 16. Identify the limiting reactant, calculate the mass of MgO and the mass of oxygen remaining. [6]



3.5 11.2 g iron reacts with 4.8 g oxygen to make 16.0 g iron(III) oxide. Use Fe = 56 and O = 16; reactant formulae Fe and O_2, product Fe_2O_3. Convert all three masses to moles and deduce the balanced equation. [4]

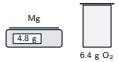
Fe	O_2	Fe_2O_3
11.2 g	4.8 g	16.0 g

3.6 A sample contains 0.25 mol O_2 molecules. Use 6.02×10^{23} particles per mole and O = 16. Calculate the number of O_2 molecules, the amount and number of oxygen atoms, and the sample mass. [4]



Exercise 3.3

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Fe	O_2	Fe_2O_3
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3.3 g/dm³: 3.5 g in 0.250 dm³ → 3.5 ÷ 0.25 = 14 g/dm³; Mr(NaCl) = 58.5; mol/dm³ = 14 ÷ 58.5 = 0.24 mol/dm³.

3.4 n(Mg) = 4.8/24 = 0.20 mol; n(O_2) = 6.4/32 = 0.20 mol. The 2:1 equation ratio means Mg is limiting: 0.20 mol Mg needs only 0.10 mol O_2. MgO amount = 0.20 mol, so mass = 0.20 × 40 = 8.0 g. O_2 left = (0.20 − 0.10) × 32 = 3.2 g.

3.5 Amounts: Fe = 11.2/56 = 0.200 mol; O_2 = 4.8/32 = 0.150 mol; Fe_2O_3 = 16.0/160 =

Full solution on next slide

Full solution 3.3

3.3 g/dm^3: 3.5 g in 0.250 dm^3 $\rightarrow$ $3.5 \div 0.25 = 14$ g/dm^3; Mr(NaCl) = 58.5; mol/dm^3 = $14 \div 58.5 = 0.24$ mol/dm^3.

3.4 n(Mg) = $4.8/24 = 0.20$ mol; n(O_2) = $6.4/32 = 0.20$ mol. The 2:1 equation ratio means Mg is limiting: 0.20 mol Mg needs only 0.10 mol O_2. MgO amount = 0.20 mol, so mass = $0.20 \times 40 = 8.0$ g. O_2 left = $(0.20 - 0.10) \times 32 = 3.2$ g.

3.5 Amounts: Fe = $11.2/56 = 0.200$ mol; O_2 = $4.8/32 = 0.150$ mol; Fe_2O_3 = $16.0/160$ = 0.100 mol. Ratio 2:1.5:1 becomes 4:3:2, hence $4Fe + 3O_2 \rightarrow 2Fe_2O_3$; all atoms and total mass balance.

3.6 O_2 molecules = $0.25 \times 6.02 \times 10^{23} = 1.505 \times 10^{23}$. Each molecule contains two O atoms: amount of O atoms = $2 \times 0.25 = 0.50$ mol, atom count = $0.50 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$. O_2 molar mass = 32 g/mol, so sample mass = $0.25 \times 32 = 8.0$ g. Accept consistent sensible rounding.