

What mass of oxygen is required for the complete combustion of 5.8 g of butane, C₄H₁₀. How many moles of CO₂ and H₂O are produced?

Marks
4

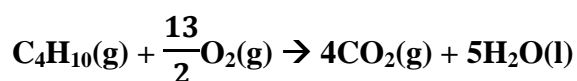
The molar mass of butane, C₄H₁₀, is:

$$\text{molar mass} = M = (4 \times 12.01 \text{ (C)} + 10 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 58.12 \text{ g mol}^{-1}$$

Hence, the 5.8 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{number of moles}} = \frac{m}{M} = \frac{5.8 \text{ g}}{58.12 \text{ g mol}^{-1}} = 0.10 \text{ mol}$$

The balanced equation for the combustion of butane is:



Hence, 13/2 moles of O₂ are required and 4 moles of CO₂ and 5 moles of H₂O are produced for every mole of C₄H₁₀ which combusts. As 0.10 mol of C₄H₁₀ is present:

$$\text{number of moles of O}_2 \text{ needed} = \frac{13}{2} \times 0.10 \text{ mol} = 0.65 \text{ mol}$$

$$\text{number of moles of CO}_2 \text{ produced} = 4 \times 0.10 \text{ mol} = 0.40 \text{ mol}$$

$$\text{number of moles of H}_2\text{O produced} = 5 \times 0.10 \text{ mol} = 0.50 \text{ mol}$$

O₂ has a molar mass of (2 × 16.00) g mol⁻¹ = 32.00 g mol⁻¹. Hence the mass of O₂ required is:

$$\begin{aligned} \text{mass} &= \text{number of moles} \times \text{molar mass} \\ &= nM = (0.65 \text{ mol}) \times (32.00 \text{ g mol}^{-1}) = 21 \text{ g} \end{aligned}$$