What mass of oxygen is required for the complete combustion of 5.8 g of butane, $C_4H_{10}$. How many moles of $CO_2$ and $H_2O$ are produced?

**Marks**
**4**

**The molar mass of butane, $C_4H_{10}$, is:**

**molar mass = $M$ = (4 × 12.01 (C) + 10 × 1.008 (H)) g mol$^{-1}$ = 58.12 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:**

$$C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

**Hence, 13/2 moles of $O_2$ are required and 4 moles of $CO_2$ and 5 moles of $H_2O$ are produced for every mole of $C_4H_{10}$ which combusts. As 0.10 mol of $C_4H_{10}$ is present:**

$$\text{number of moles of } O_2 \text{ needed} = \frac{13}{2} \times 0.10\text{ mol} = 0.65\text{ mol}$$

**number of moles of $CO_2$ produced = 4 × 0.10 mol = 0.40 mol**

**number of moles of $H_2O$ produced = 5 × 0. 10 mol = 0.50 mol**

**$O_2$ has a molar mass of (2 × 16.00) g mol$^{-1}$ = 32.00 g mol$^{-1}$. Hence the mass of $O_2$ required is:**

**mass = number of moles × molar mass**
**= $nM$ = (0.65 mol) × (32.00 g mol$^{-1}$) = 21 g**