What mass of oxygen is required for the complete combustion of 5.8 g of butane, \( \text{C}_4\text{H}_{10} \). How many moles of \( \text{CO}_2 \) and \( \text{H}_2\text{O} \) are produced?

The molar mass of butane, \( \text{C}_4\text{H}_{10} \), is:

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

Hence, the 5.8 g corresponds to:

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

The balanced equation for the combustion of butane is:

\[
\text{C}_4\text{H}_{10}(g) + \frac{13}{2} \text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 5\text{H}_2\text{O}(l)
\]

Hence, \( \frac{13}{2} \) moles of \( \text{O}_2 \) are required and 4 moles of \( \text{CO}_2 \) and 5 moles of \( \text{H}_2\text{O} \) are produced for every mole of \( \text{C}_4\text{H}_{10} \) which combusts. As 0.10 mol of \( \text{C}_4\text{H}_{10} \) is present:

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

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

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

\( \text{O}_2 \) has a molar mass of \( (2 \times 16.00) \text{ g mol}^{-1} = 32.00 \text{ g mol}^{-1} \). Hence the mass of \( \text{O}_2 \) required is:

\[
\text{mass} = \text{number of moles} \times \text{molar mass} = nM = (0.65 \text{ mol}) \times (32.00 \text{ g mol}^{-1}) = 21 \text{ g}
\]