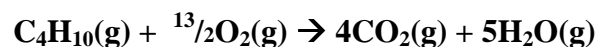
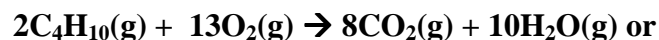


- The complete combustion of butane, C<sub>4</sub>H<sub>10</sub>, in air gives water and carbon dioxide as the products. Write a balanced equation for this reaction.

**Marks**  
**4**



What mass of oxygen is required for the complete combustion of 454 g of butane and what masses of carbon dioxide and water are produced?

**The molar mass of butane is  $(4 \times 12.01 \text{ (C)}) + (10 \times 1.008 \text{ (H)}) = 58.12$ . Therefore, the amount of butane in 454 g is:**

$$\text{number of moles of C}_4\text{H}_{10} = \frac{\text{mass}}{\text{molar mass}} = \frac{454}{58.12} = 7.81 \text{ mol}$$

**From the chemical equation, each mole of C<sub>4</sub>H<sub>10</sub> requires  $\frac{13}{2}$  moles of O<sub>2</sub> and produces 4 moles of CO<sub>2</sub> and 5 moles of H<sub>2</sub>O.**

**Therefore:**

$$\text{number of moles of O}_2 = \frac{13}{2} \times 7.81 = 50.8 \text{ mol}$$

$$\text{number of moles of CO}_2 = 4 \times 7.81 = 31.2 \text{ mol}$$

$$\text{number of moles of H}_2\text{O} = 5 \times 7.81 = 39.1 \text{ mol}$$

**The molar masses of O<sub>2</sub>, CO<sub>2</sub> and H<sub>2</sub>O are:**

$$\text{molar mass of O}_2 = 2 \times 16.00 = 32.00$$

$$\text{molar mass of CO}_2 = 12.01 \text{ (C)} + (2 \times 16.00) = 44.01$$

$$\text{molar mass of H}_2\text{O} = (2 \times 1.008 \text{ (H)}) + 16.00 \text{ (O)} = 18.016$$

**Therefore:**

$$\text{mass of O}_2 = \text{number of moles} \times \text{molar mass} = 50.8 \times 32.00 = 1620 \text{ g} = 1.62 \text{ kg}$$

$$\text{mass of CO}_2 = 31.2 \times 44.01 = 1380 \text{ g} = 1.38 \text{ kg}$$

$$\text{mass of H}_2\text{O} = 39.1 \times 18.016 = 704 \text{ g} = 0.704 \text{ kg}$$