5. During complete combustion of one mole of butane, 2658 kJ of heat is released. The thermochemical reaction for above change is

(i) $2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(l)$ $\Delta_c H = -2658.0 \text{ kJ mol}^{-1}$

(ii) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(g)$ $\Delta_c H = -1329.0 \text{ kJ mol}^{-1}$

(iii) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$ $\Delta_c H = -2658.0 \text{ kJ mol}^{-1}$

(iv) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$ $\Delta_c H = +2658.0 \text{ kJ mol}^{-1}$

Solution:

(iii) is the correct answer

Explanation:

As heat is released, hence, heat of combustion will be negative. By convention, heat of combustion is defined to the heat released for the complete combustion of a compound in its standard state to form stable products in their standard state: so water has to be in liquid state