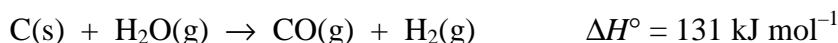
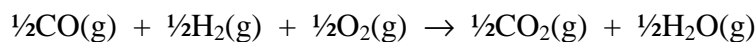


Marks
3

- “Water gas” is a mixture of combustible gases produced from steam and coal according to the following reaction:



The equation for the complete combustion of 1 mol of water gas (*i.e.* 0.5 mol CO(g) and 0.5 mol H₂(g)) can be written as:



Calculate the standard enthalpy of combustion of water gas, given the following thermochemical data.

$$\Delta H^\circ_{\text{vap}}(\text{H}_2\text{O}) = 44 \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_{\text{f}}(\text{H}_2\text{O(l)}) = -286 \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_{\text{f}}(\text{CO}_2\text{(g)}) = -393 \text{ kJ mol}^{-1}$$

Using $\Delta_{\text{rxn}}H^\circ = \sum m\Delta_{\text{f}}H^\circ(\text{products}) - \sum n\Delta_{\text{f}}H^\circ(\text{reactants})$ for the vaporization of water ($\text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{O(g)}$) gives

$$\begin{aligned} \Delta_{\text{vap}}H^\circ &= [\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(g)})] - [\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(l)})] \\ &= [\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(g)})] - (-286) = +44 \end{aligned}$$

Hence $\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(g)}) = (+44) + (-286) = -242 \text{ kJ mol}^{-1}$

Using $\Delta_{\text{rxn}}H^\circ = \sum m\Delta_{\text{f}}H^\circ(\text{products}) - \sum n\Delta_{\text{f}}H^\circ(\text{reactants})$ for the reaction,

$\text{C(s)} + \text{H}_2\text{O(g)} \rightarrow \text{CO(g)} + \text{H}_2\text{(g)}$ gives

$$\begin{aligned} \Delta_{\text{rxn}}H^\circ &= [\Delta_{\text{f}}H^\circ(\text{CO(g)})] - [\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(g)})] \\ &= [\Delta_{\text{f}}H^\circ(\text{CO(g)})] - (-242) = +131 \end{aligned}$$

as $\Delta_{\text{f}}H^\circ(\text{H}_2\text{(g)})$ and $\Delta_{\text{f}}H^\circ(\text{C(s)})$ are both zero for elements in their standard states. Hence $\Delta_{\text{f}}H^\circ(\text{CO(g)}) = -111 \text{ kJ mol}^{-1}$

Using $\Delta_{\text{rxn}}H^\circ = \sum m\Delta_{\text{f}}H^\circ(\text{products}) - \sum n\Delta_{\text{f}}H^\circ(\text{reactants})$ for the reaction,

$\frac{1}{2}\text{CO(g)} + \frac{1}{2}\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \frac{1}{2}\text{CO}_2\text{(g)} + \frac{1}{2}\text{H}_2\text{O(g)}$ gives

$$\Delta_{\text{comb}}H^\circ = \left[\frac{1}{2}\Delta_{\text{f}}H^\circ(\text{CO}_2\text{(g)}) + \frac{1}{2}\Delta_{\text{f}}H^\circ(\text{H}_2\text{O(g)})\right] - \left[\frac{1}{2}\Delta_{\text{f}}H^\circ(\text{CO(g)})\right]$$

as the enthalpy of formation of H₂(g) and O₂(g) are both zero for elements in their standard states. Hence,

$$\Delta_{\text{comb}}H^\circ = \left[\left(\frac{1}{2} \times -393\right) + \left(\frac{1}{2} \times -242\right)\right] - \left[\left(\frac{1}{2} \times -111\right)\right] = -262 \text{ kJ mol}^{-1}$$

Answer: -262 kJ mol^{-1}

THIS QUESTION CONTINUES ON THE NEXT PAGE.