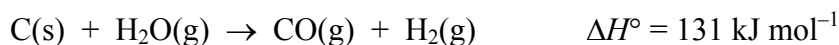
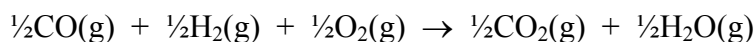


- “Water gas” is a mixture of combustible gases produced from steam and coal according to the following reaction:



The equation for the complete combustion of 1 mol of water gas (*i.e.* 0.5 mol CO(g) and 0.5 mol H<sub>2</sub>(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 H<sub>2</sub>(g) and O<sub>2</sub>(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<sup>-1</sup>

THIS QUESTION CONTINUES ON THE NEXT PAGE.