Marks
5

- Stearic acid, $C_{18}H_{36}O_2$, is a fatty acid common in animal fats and vegetable oils and is a valuable energy source for mammals. The net reaction for its metabolism in humans is:

$$C_{18}H_{36}O_2(s) + 26O_2(g) \rightarrow 18CO_2(g) + 18H_2O(l)$$

Calculate $\Delta H°$ for this reaction given the following heats of formation.

Compound	$C_{18}H_{36}O_2(s)$	$CO_2(g)$	$H_2O(l)$
$\Delta_f H°$ / kJ mol^{-1}	–948	–393	–285

Using $\Delta_{rxn}H° = \Sigma m\Delta_f H°(\text{products}) - \Sigma n\Delta_f H°(\text{reactants})$, the enthalpy of the combustion of the reaction with excess O_2 is:

$$\Delta H° = (18\Delta_f H°(CO_2(g) + 18\Delta_f H°(H_2O(l)) - (\Delta_f H°(C_{18}H_{36}O_2(s))$$

$$= [(18 \times -393 + 18 \times -285) – (-948)] \text{ kJ mol}^{-1} = –11300 \text{ kJ mol}^{-1}$$

As $O_2(g)$ is an element in its standard state, its $\Delta_f H°$ is zero.

$\Delta H°$ = **–11300 kJ mol^{-1}**

If the combustion of stearic acid is carried out in air, water is produced as a vapour. Calculate the $\Delta H°$ for the combustion of stearic acid in air given that:

$$H_2O(l) \rightarrow H_2O(g) \qquad \Delta H° = +44 \text{ kJ mol}^{-1}$$

The reaction now needs vapourisation of $18H_2O(l)$ which requires *input* of (18 × 44) kJ mol^{-1}. The combustion enthalpy is reduced to:

$$\Delta H° = (-11300 \text{ kJ mol}^{-1}) + (18 \times 44 \text{ kJ mol}^{-1}) = –10500 \text{ kJ mol}^{-1}$$

$\Delta H°$ = **–10500 kJ mol^{-1}**

Will ΔS be different for the two oxidation reactions? If so, how will it differ and why?

ΔS will be greater for the air-oxidation as the product $H_2O(g)$ has a much greater entropy than the product $H_2O(l)$ - gases are much more disordered than liquids.

Calculate the mass of carbon dioxide produced by the complete oxidation of 1.00 g of stearic acid.

The molar mass of stearic acid, $C_{18}H_{36}O_2$, is:

$$\text{molar mass} = (18 \times 12.01 \text{ (C)} + 36 \times 1.008 \text{ (H)} + 2 \times 16.00 \text{ (H)}) \text{ g mol}^{-1}$$
$$= 284.47 \text{ g mol}^{-1}$$

ANSWER CONTINUES ON THE NEXT PAGE

The number of moles in 1.00 g is therefore:

number of moles = mass / molar mass
= (1.00 g) / (284.47 g mol^{-1}) = 0.00353 mol

The chemical equation shows that 18 mol of CO_2(g) are produced from every 1 mol of $C_{18}H_{36}O_2$(s). The number of moles of CO_2(g) produced is therefore (18 × 0.00353 mol) = 0.06328 mol.

The molar mass of CO_2 is (12.01 (C) + 2 × 16.00 (O)) g mol^{-1} = 44.01 g mol^{-1}. The mass of CO_2 produced is therefore:

mass of CO_2 = number of moles × molar mass
= (0.06328 mol) × (44.01 g mol^{-1}) = 2.78 g

Answer: **2.78 g**