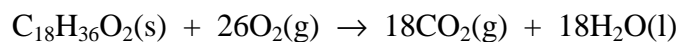


- Stearic acid,  $C_{18}H_{36}O_2$ , is a fatty acid common in animal fats and vegetable oils and is a valuable energy source for mammals. The net reaction for its metabolism in humans is:



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

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

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

$$\begin{aligned} \Delta H^\circ &= (18\Delta_f H^\circ(CO_2(g)) + 18\Delta_f H^\circ(H_2O(l)) - (\Delta_f H^\circ(C_{18}H_{36}O_2(s))) \\ &= [(18 \times -393 + 18 \times -285) - (-948)] \text{ kJ mol}^{-1} = -11300 \text{ kJ mol}^{-1} \end{aligned}$$

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

$$\Delta H^\circ = -11300 \text{ kJ mol}^{-1}$$

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



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

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

$$\Delta H^\circ = -10500 \text{ kJ mol}^{-1}$$

Will  $\Delta S$  be different for the two oxidation reactions? If so, how will it differ and why?

$\Delta S$  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:

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

ANSWER CONTINUES ON THE NEXT PAGE

**The number of moles in 1.00 g is therefore:**

$$\begin{aligned}\text{number of moles} &= \text{mass} / \text{molar mass} \\ &= (1.00 \text{ g}) / (284.47 \text{ g mol}^{-1}) = 0.00353 \text{ mol}\end{aligned}$$

**The chemical equation shows that 18 mol of CO<sub>2</sub>(g) are produced from every 1 mol of C<sub>18</sub>H<sub>36</sub>O<sub>2</sub>(s). The number of moles of CO<sub>2</sub>(g) produced is therefore (18 × 0.00353 mol) = 0.06328 mol.**

**The molar mass of CO<sub>2</sub> is (12.01 (C) + 2 × 16.00 (O)) g mol<sup>-1</sup> = 44.01 g mol<sup>-1</sup>. The mass of CO<sub>2</sub> produced is therefore:**

$$\begin{aligned}\text{mass of CO}_2 &= \text{number of moles} \times \text{molar mass} \\ &= (0.06328 \text{ mol}) \times (44.01 \text{ g mol}^{-1}) = 2.78 \text{ g}\end{aligned}$$

Answer: **2.78 g**