Answer on Question #59760 – Chemistry – General Chemistry

Task:

Calculate the enthalpy change of the reaction below by two different methods.

$C_3H_8(g) + 5O_2(g)$ ------> $3CO_2(g) + 4H_2O(g)$

a) Use the standard enthalpies of formation:

$C_3H_8(g)$	-104 kJ mol^{-1}
$CO_2(g)$	-394 kJ mol^{-1}
$H_2O(g)$	-242 kJ mol^{-1}

b) Use the following mean bond enthalpies quoted in kJ mol^{-1}.

C-C 348, C-H 412, O=O 496, C=O 804 and O-H 463

c) Account for any difference between the two values.

Solution:

Propane combustion reaction:

$$C_3H_8(g) + 5O_2(g) = 3CO_2(g) + 4H_2O(g).$$

a) Enthalpy change ΔH: The net heat energy transferred to a system from the surroundings or from the surroundings to a system at constant pressure.

The standard-state enthalpy of reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants:

$$\Delta H^o(reaction) = \sum \Delta H_f^o(products) - \sum \Delta H_f^o(reac\tan ts).$$

NOTE: The heat of formation of O2 is zero because this is the form of the oxygen in its most thermodynamically stable state.

Then,

$$\Delta H^o(reaction) = 3 \times \Delta H_f^o(CO_2) + 4 \times \Delta H_f^o(H_2O) - \Delta H_f^o(C_3H_8).$$

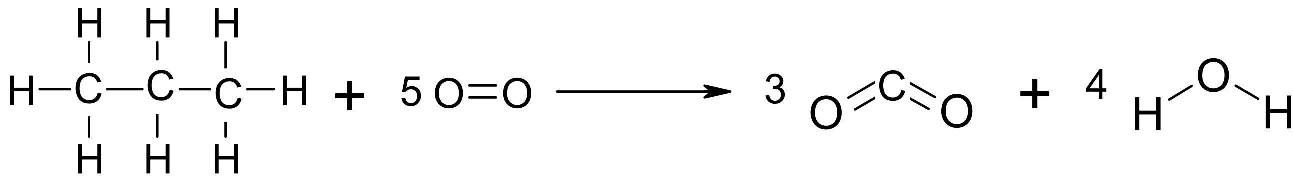
$$\Delta H^o(reaction) = 3 \times (-394) + 4 \times (-242) - (-104) = -2046\,(kJ \times mol^{-1}).$$

$$\Delta H^o(reaction) = -2046\,kJ \times mol^{-1}.$$

Answer (a): $\Delta H^o(reaction)(a) = -2046\,kJ \times mol^{-1}.$

b) Bond Energy: The energy required to break a bond. Bond energy is always a positive number because the breaking of a bond requires an input of energy (endothermic). When a bond is formed, the amount of energy equal to the bond energy is released.

$$\Delta H^o(reaction) = \sum Bond\ energies\ of\ bonds\ broken - \sum Bond\ energies\ of\ bonds\ formed.$$



$$C_3H_8 = 2\times(C-C) + 8\times(C-H);$$
$$O_2 = 1\times(O=O);$$
$$CO_2 = 2\times(C=O);$$
$$H_2O = 2\times(O-H).$$

Then,

$$\Delta H^o(reaction) = E(C_3H_8) + 5\times E(O_2) - (3\times E(CO_2) + 4\times E(H_2O)).$$

$$\Delta H^o(reaction) = 2\times(C-C) + 8\times(C-H) + 5\times 1\times(O=O) - (3\times 2\times(C=O) + 4\times 2\times(O-H)).$$

$$\Delta H^o(reaction) = 2\times 348 + 8\times 412 + 5\times 496 - (6\times 804 + 8\times 463) = -2056\,(kJ\times mol^{-1}).$$

$$\Delta H^o(reaction) = -2056\,kJ\times mol^{-1}.$$

Answer (b): $\Delta H^o(reaction)(b) = -2056\,kJ\times mol^{-1}.$

c) Account for any difference between the two values:

$$\Delta = \Delta H^o(reaction)(a) - \Delta H^o(reaction)(b) = -2046\,kJ\times mol^{-1} - (-2056\,kJ\times mol^{-1}) = 10\,kJ\times mol^{-1}.$$
$$\Delta = 10\,kJ\times mol^{-1}.$$

Answer (c): $\Delta = 10\,kJ\times mol^{-1}.$

Answer:

a) $\Delta H^o(reaction)(a) = -2046\,kJ\times mol^{-1}.$

b) $\Delta H^o(reaction)(b) = -2056\,kJ\times mol^{-1}.$

c) $\Delta = 10\,kJ\times mol^{-1}.$