Answer on Question #81926, Chemistry/ General Chemistry

The standard enthalpies of formation, at 25.00degrees celsius, of methanol (CH4O (l)), water (H2O (l)), and carbon dioxide (CO2 (g)) are respectively -238.7 kJ / mol, - 285.8 kJ / mol, and - 393.5 kJ / mol. Calculate the change in the surrounding entropy (in J / K) when burning 15.4 g of methanol under a constant pressure of 1,000 atm at 25.00 degrees celsius (combustion is the reaction of a substance with oxygen molecular to produce water and carbon dioxide).

Solution

$2CH_3OH\ (l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$

Calculate the standard enthalpy of this reaction

ΔH_f^0($CH_4O(l)$) = -238.7 kJ/mol

ΔH_f^0($CO_2(g)$) = -393.5 kJ/mole.

ΔH_f^0($H_2O(l)$) = -285.8 kJ/mol

ΔH_f^0($O_2(g)$) = 0 kJ/mol

$$\Delta \boldsymbol{H}^{\boldsymbol{0}}_{\boldsymbol{rxn}} = \sum \boldsymbol{H}^{\boldsymbol{0}}_{\boldsymbol{f(products)}} - \sum \boldsymbol{H}^{\boldsymbol{0}}_{\boldsymbol{f(reactants)}}$$

$$\begin{aligned}\Delta \boldsymbol{H}^{\boldsymbol{0}}_{\boldsymbol{rxn}} &= \left[2\ mol \times H_f^0\left(CO_2(g)\right) + 4\ mol \times H_f^0\left(H_2O(l)\right)\right] \\ &\quad - \left[2mol \times\ H_f^0\left(CH_4O(l)\right) + 3\ mol{\times}H_f^0\left(O_2(g)\right)\right] \\ &= \left[2mol \times \left(-393.5\ \frac{kJ}{mol}\right) + 4\ mol{\times}\left(-285.8\frac{kJ}{mol}\right)\right] \\ &\quad - \left[2\ mol \times (\ -238.7\frac{kJ}{mol}) + 1\ mol{\times}(0\frac{kJ}{mol})\right] = -1452.8\ kJ\end{aligned}$$

Calculate the standard enthalpy of this reaction when 15.4 g of methanol is burned:

n=m/M

$M(CH_4O)$= 32 g/mol

$n(CH_4O)$= $\frac{15.4\ g}{32\frac{g}{mol}} = 0.48\ mol$

Solve the proportion:

When 2 mol of methanol is burned $\Delta H^0_{rxn}\ = -1452.8\ kJ$

When 0.48 mol of methanol is burned $\Delta H^0_{rxn}\ = x\ kJ$

$$\frac{2}{0.48} = \frac{-1452.8}{x}$$

$$x = -348.7\ kJ$$

$$\Delta S_{surroundings} \ = -\frac{\Delta H}{T}$$

$$\Delta S_{surroundigs} \ = -\frac{-348700\ J}{(273 + 25)K} = 1170\,\frac{J}{K}$$

Answer: 1170 J/K