SCH4U

Asgmt. D

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1.

i. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H_1 = -393.3\ kJ$ (1)

ii. $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$ $\Delta H_2 = -293.72\ kJ$ (2)

iii. $CS_{2(l)} + 3O_{2(g)} \rightarrow CO_{2(g)} + 2SO_{2(g)}$ $\Delta H_3 = -1108.76\ kJ$ (3)

Multiply (2) by 2 to produce (4), flip (3) to produce 5.

$C_{(s)} + O_{2(g)} \rightarrow CO_2$ $\Delta H_1 = -393.3\ kJ$

Multiplid by 2: $2S_{(s)} + 2O_2 \rightarrow 2SO_{2(g)}$ $\Delta H_4 = -587.44\ kJ$ (4)

Flipped: $CO_{2(g)} + 2SO_{2(g)} \rightarrow CS_{2(l)} + 3O_{2(g)}$ $\Delta H_5 = +1108.76\ kJ$ (5)

$\Delta H^\circ_r = \sum (n\Delta H^\circ_f \text{ products}) - \sum (n\Delta H^\circ_f \text{ reactants})$

$\Delta H_{(5)} = (\Delta H_{CS_2} + \Delta H_{O_2}) - (\Delta H_{(1)} + \Delta H_{(4)})$

$1108.76\ kJ = (\Delta H_{CS_2} + 0\ kJ) - (-393.3\ kJ - 587.44\ kJ)$

$\Delta H_{CS_2} = 128.02\ kJ$

The heat of formation of CS_2 is 128.0 kJ

2.

$q = c\Delta T$

$\frac{q}{m} = \frac{c\Delta T}{m}$

$m = 0.250\ g$

$\Delta T = 20.^\circ C$ (From our discussion in class)

$c = 7.28\ kJ/^\circ C$

$\frac{q}{m} = \frac{(7.28\ kJ/^\circ C)(20^\circ C)}{0.250\ g}$

$= 582.4\ kJ/g \simeq 580\ kJ/g$

Since the temperature of a calorimeter increases, the heat of combustion is negative.

$\frac{q}{m} = -5.8 \times 10^2\ kJ/g$

The heat of combustion of fuel per gram is $-5.8 \times 10^2\ kJ/g$

3.

$A + B_2 \rightarrow AB_2$ (slow)

$AB_2 + C \rightarrow AB_2C$ (fast)

$AB_2C + C \rightarrow A + B_2C_2$ (fast)

$B_2C_2 \rightarrow D + E$ (fast)

a) Reaction intermediates: AB_2, AB_2C, B_2C_2

b) Catalyst: A

c) Overall eq: $B_2 + 2C \xrightarrow{A} D + E$

d) Rate Law: $R = k[A]^1[B_2]^1$

4.

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$V = 0.1\,L$ $\quad V = 0.05\,L$ (Excess)

$n_{HCl} = cV$

$= (0.355\,M)(0.1\,L)$

$= 0.0355\,mol$

$q_{system} = -q_{surroundings}$

$q_{system} = -mc\Delta T$

$= -(150\,g)(4.18\,J/g^\circ C)(4.2^\circ C)$ ← Assume $1\,mL = 1\,g$ (of water)

$= -2633.4\,J$

$\frac{q_{system}}{mol} = \frac{q_{system}}{mol}$

$= \frac{-2633.4\,J}{(0.0355\,mol)}$

$= -74180.3\,J/mol \cong -74.2\,kJ/mol$

The molar enthalpy change is $-74.2\,kJ/mol$

5.

a) $\frac{\Delta[NH_3]}{\Delta t} = -\frac{([NH_3]_f - [NH_3]_i)}{\Delta t}$

$= -\frac{(\frac{1.6}{4.0}M - \frac{3.8}{4.0}M)}{180\,s}$

$= 2.6 \times 10^{-3}\,mol/L \cdot s$

b) Since 6 mol of H_2O are produced for 4 mol of NH_3, the rate of production of H_2O is $\frac{3}{2}$ times that of NH_3.

$\frac{\Delta[H_2O]}{\Delta t} = (\frac{3}{2})\frac{\Delta[NH_3]}{\Delta t}$

$= (\frac{3}{2})(2.6 \times 10^{-3}\,mol/L \cdot s)$

$= 4.0 \times 10^{-3}\,mol/L \cdot s$

Con't on next page.

5b (cont'd)

The average rate at which H_2O is formed is 4.0×10^{-3} mol/L·s

5c)

mols of for O_2 are consumed in a 5:4 proportion with NH_3.

$$\frac{\Delta[O_2]}{\Delta t} = \left(\frac{5}{4}\right) \frac{\Delta[NH_3]}{\Delta t}$$

$$= \left(\frac{5}{4}\right)\left(2.6 \times 10^{-3} \frac{mol}{L \cdot s}\right)$$

$$= 3.3 \times 10^{-3} \text{ mol/L·s}$$

The average rate at which O_2 is formed is 3.3×10^{-3} mol/L·s.

6.

a) When [NO] doubles from row 1 to row 2, the rate also doubles.

Let x = order of rxn w.r.t. [NO]

$2^x = 2$

$x = 1$

The order of this reaction with respect to [NO] is 1 (first)

b) When $[Br_2]$ doubles from row 1 to row 3, the rate quadruples.

Let x = order of rxn w.r.t. $[Br_2]$

$2^x = 4$

$x = 2$

The order of this reaction with respect to $[Br_2]$ is 2 (second)

c) The overall order of the reaction is 1+2 = 3 (third).

d) Using row 3's data:

$$R = k[NO][Br_2]^2$$

$$0.56 \frac{mol}{L \cdot s} = k(0.8\ M)(1.20\ M)^2$$

$$k = \frac{0.56 \frac{mol}{L \cdot s}}{(0.8\ M)(1.20\ M)^2}$$

$$= 0.49 \frac{L^2}{mol^2 \cdot s}$$

∴ The rate law constant, $k = 0.49 \frac{L^2}{mol^2 \cdot s}$