SCH4U

Dhrumil Patel

Asgmt. D

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)

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

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

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

$\Delta H^\circ_r = \Sigma(n\Delta H^\circ_f \text{ products}) - \Sigma(n\Delta H^\circ_f \text{ reactants})$ (Hess' Law)

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

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

$\Delta H_{CS_2} = -980.74\ kJ + 1108.76\ kJ$

$= 128.02\ kJ$ $\therefore \Delta H_f$ of CS_2 is $128.0\ kJ$

2.

$q = c\Delta T$ $m = 0.250\ g$

$\frac{q}{m} = \frac{c\Delta T}{m}$ $\Delta T = 20.^\circ C$ (From our discussion in class)

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

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

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

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 $-5.8 \times 10^2\ kJ/g$

Dhrumil Patel

3.

$\boxed{A} + B_2 \rightarrow AB_2$ (slow)

$AB_2 + C \rightarrow AB_2C$ (fast)

$AB_2C + C \rightarrow \boxed{A} + B_2C_2$ (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: 0.1L V= 0.05L (Excess)

$n_{HCl} = cV$

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

$= 0.0355$ mol

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

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

$= -(150\,g)(4.18\,J/g°C)(4.2°C)$ ← Assume 1mL = 1g, Use c = 4.18 J/g°C

$= -2633.4$ J

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

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

$= -74180.3$ J/mol ≈ -74.2 kJ/mol

∴ The molar enthalpy change is -74.2 kJ/mol

5.

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

$= \frac{-\left(\frac{1.6}{4.0}M - \frac{3.5}{4.0}M\right)}{180\,s}$

$= 2.6 \times 10^{-3}$ mol/L·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.

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

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

$= 4.0 \times 10^{-3}$ mol/L·s

Con't on next page.

Dhrumil Patel

5b. (con't)

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

5d)

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

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

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

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

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

$2^n = 4$

$n = 2$

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:

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

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

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

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

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