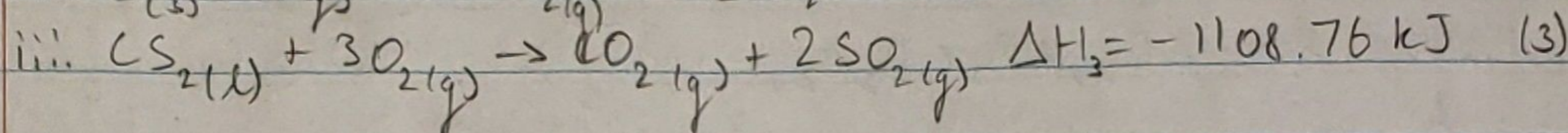
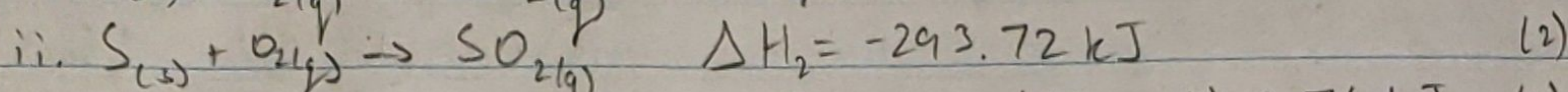
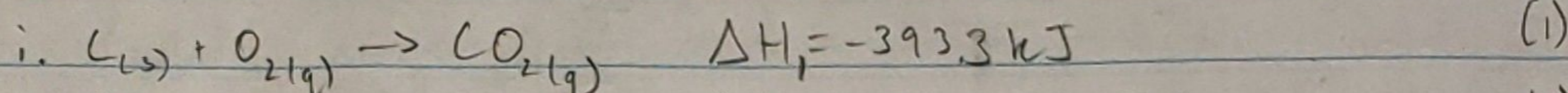
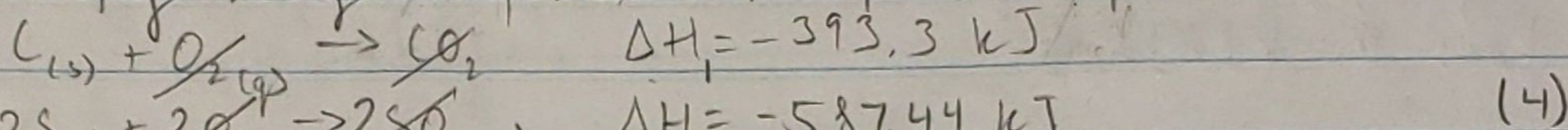


Assignment D

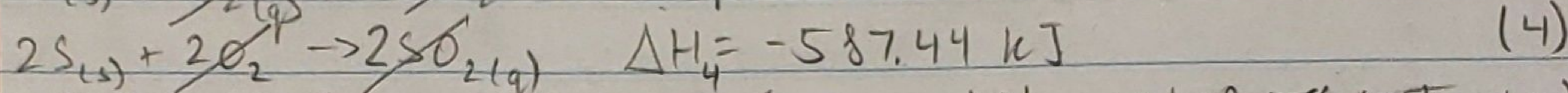
1.



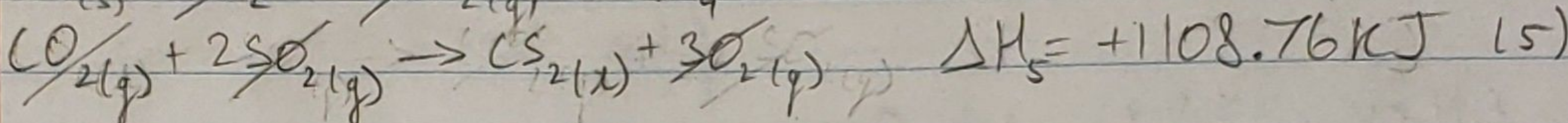
Multiply (2) by 2 to produce (4), flip (3) to produce 5.



Multiplied by 2:



Flipped:



$$\Delta H_r^\circ = \sum (n \Delta H_f^\circ \text{ products}) - \sum (n \Delta H_f^\circ \text{ reactants})$$

$$\Delta H_{(5)} = (\Delta H_{\text{CS}_2} + \Delta H_{\text{O}_2}) - (\Delta H_{\text{CO}_2} + \Delta H_{(4)})$$

$$1108.76 \text{ kJ} = (\Delta H_{\text{CS}_2} + 0 \text{ kJ}) - (-393.3 \text{ kJ} - 587.44 \text{ kJ})$$

$$\Delta H_{\text{CS}_2} = 128.02 \text{ kJ}$$

The heat of formation of CS_2 is 128.0 kJ

2.

$$q = c \Delta T$$

$$m = 0.250 \text{ g}$$

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

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

$$c = 7.78 \text{ kJ/}^\circ\text{C}$$

$$\frac{q}{m} = \frac{(7.78 \text{ kJ/}^\circ\text{C})(20^\circ\text{C})}{0.250 \text{ g}}$$

$$\frac{q}{m} = 623.2 \text{ kJ/g}$$

$$T = 582.4 \text{ kJ/g} \approx 580 \text{ kJ/g}$$

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

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

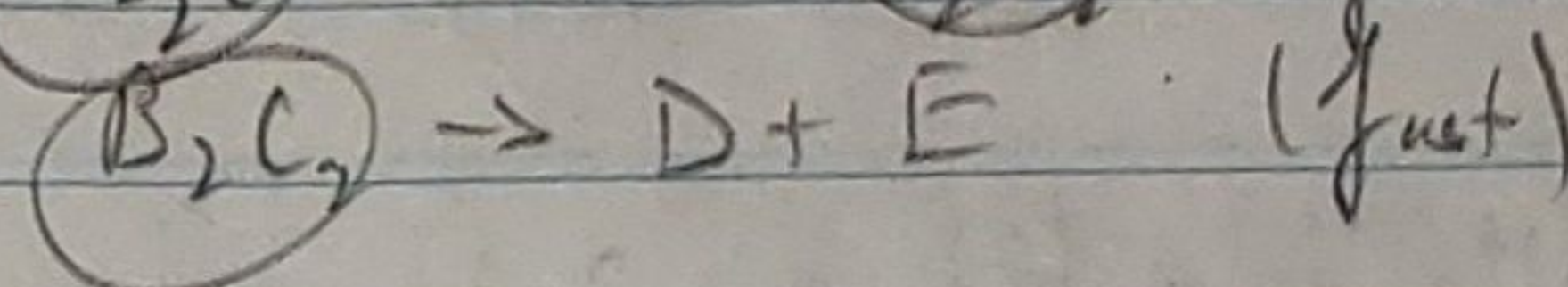
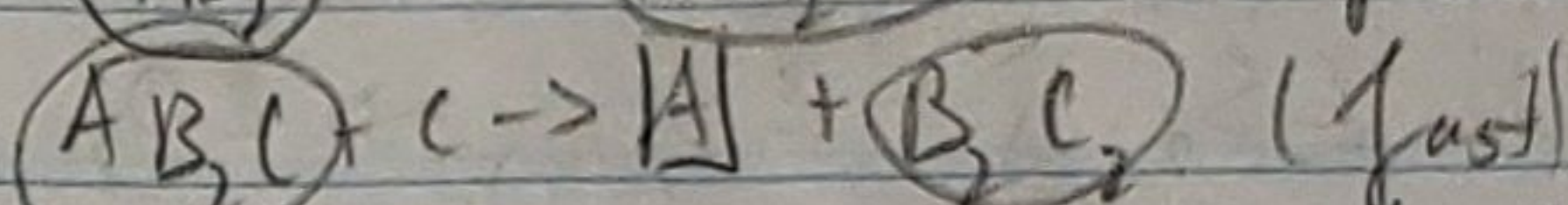
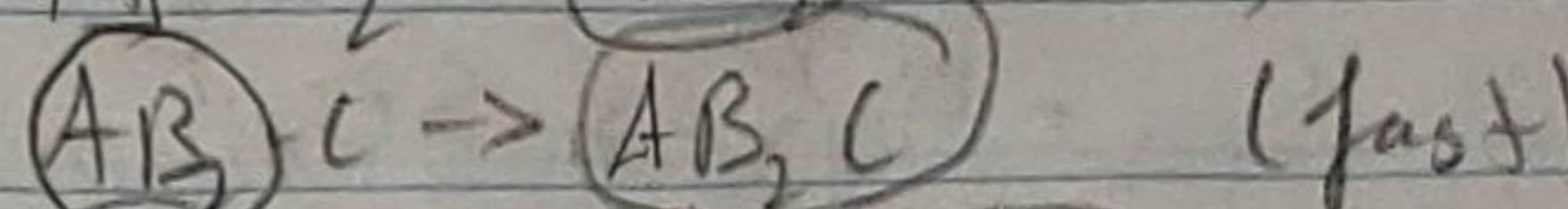
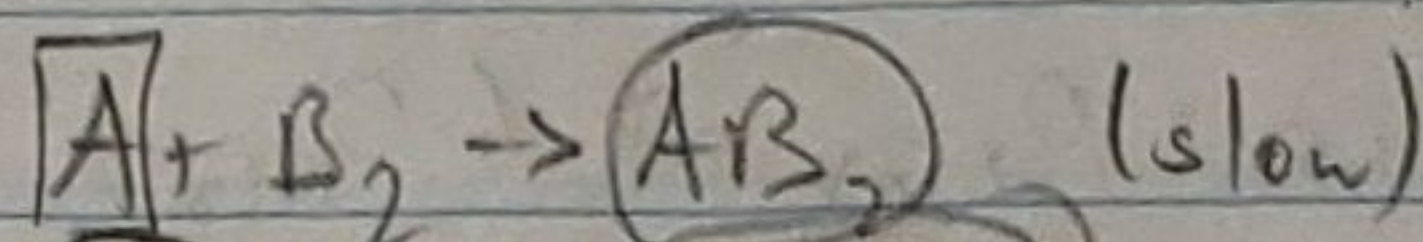
The heat of combustion of fuel per gram is $-5.8 \times 10^2 \text{ kJ/g}$

3.

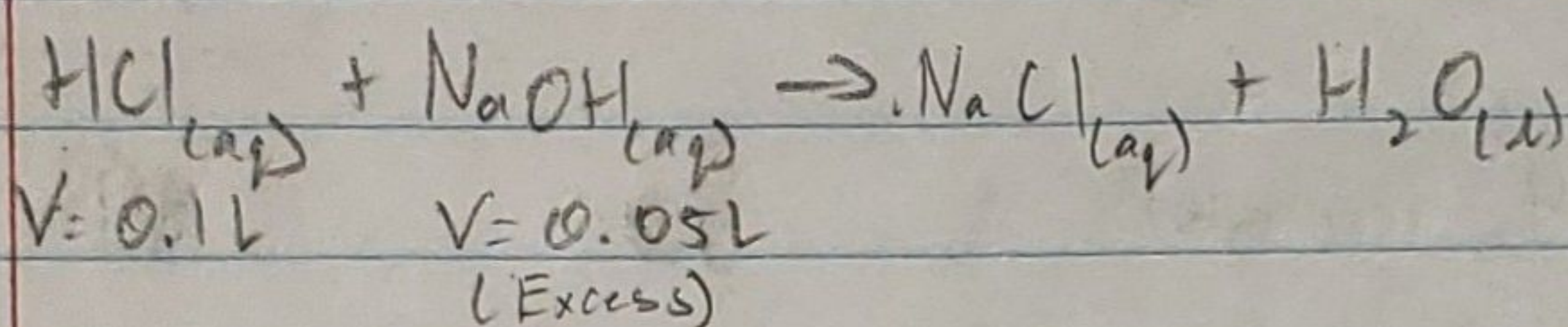
Overall reaction:

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.



$$n_{HCl} = cV$$

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

$$= 0.0355 \text{ mol}$$

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

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

$$= -(150 g)(4.18 J/g^\circ C)(4.2^\circ C) \leftarrow \text{Assume } 1 \text{ mL} = 1 g \text{ of water}$$

$$= -2633.4 J$$

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

$$= -2633.4 J$$

$$(0.0355 \text{ mol})$$

$$= -74180.3 J/mol \approx -74.2 kJ/mol$$

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

5.

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

$$= -\left(\frac{1.6 M - 3.6 M}{180s}\right)$$

$$= 2.6 \times 10^{-3} \text{ mol/L.s}$$

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 .

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

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

$$= 4.0 \times 10^{-3} \text{ mol/L.s}$$

Cont on next page.

5b) cont)

The average rate at which H_2O is formed is $4.0 \times 10^{-3} \text{ mol/L.s}$

5c)

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

$$\begin{aligned}\frac{\Delta[\text{O}_2]}{\Delta t} &= \left(\frac{5}{4}\right) \frac{\Delta[\text{NH}_3]}{\Delta t} \\ &= \left(\frac{5}{4}\right) \left(2.6 \times 10^{-3} \frac{\text{mol}}{\text{L.s}}\right) \\ &= 3.3 \times 10^{-3} \text{ mol/L.s}\end{aligned}$$

The average rate at which O_2 is formed is $3.3 \times 10^{-3} \text{ mol/L.s}$.

6.

a) When $[\text{NO}]$ doubles from row 1 to row 2, the rate also doubles.

Let x = order of rxn w.r.t. $[\text{NO}]$

$$2^x = 2$$

$$x = 1$$

The order of this reaction with respect to $[\text{NO}]$ is 1 (first)

b) When $[\text{Br}_2]$ doubles from row 1 to row 3, the rate quadruples.

Let n = order of rxn w.r.t. $[\text{Br}_2]$

$$2^n = 4$$

$$n = 2$$

The order of this reaction with respect to $[\text{Br}_2]$ is 2 (second)

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

d) Using row 3's data:

$$R = k[\text{NO}][\text{Br}_2]^2$$

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

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

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