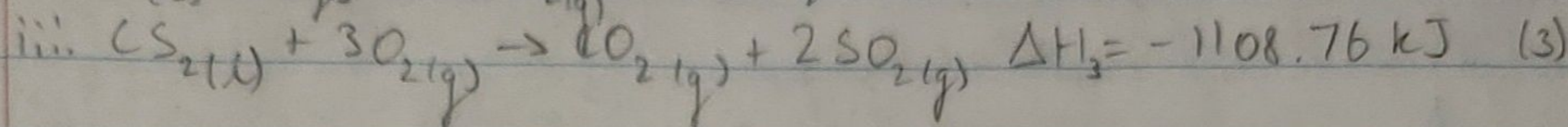
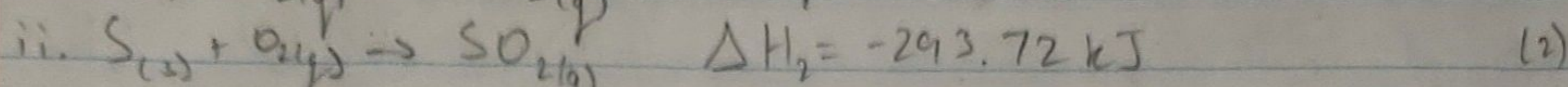
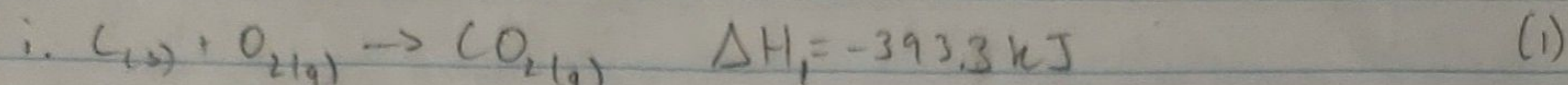
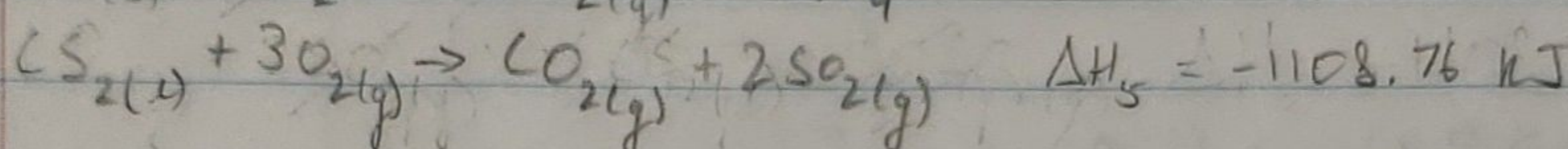
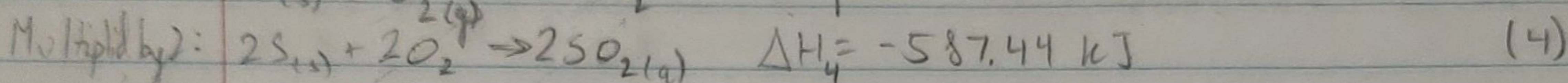
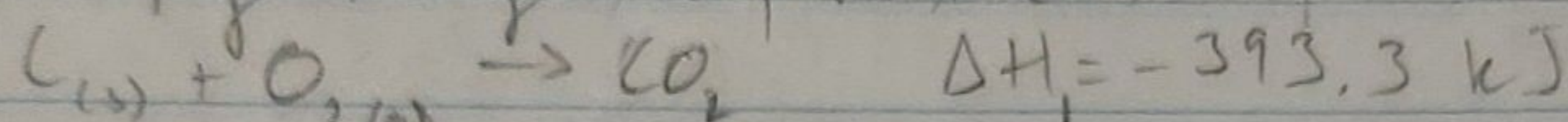


## Asgmt. D

1.



Multiply (2) by 2 to produce (4)



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

$$\Delta H_{(5)} = (\Delta H_{(1)} + \Delta H_{(4)}) - (\Delta H_{\text{CS}_2} + \Delta H_{\text{O}_2})$$

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

$$\Delta H_{\text{CS}_2} = -980.74 \text{ kJ} + 1108.76 \text{ kJ}$$

$$= 128.02 \text{ kJ} \quad \therefore \Delta H_f \text{ of } \text{CS}_2 \text{ is } 128.0 \text{ kJ}$$

2.

$$q = c \Delta T$$

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

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

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

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

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

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

$$\frac{q}{m} = 582.4 \text{ kJ/g} \approx 580 \text{ kJ/g}$$

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

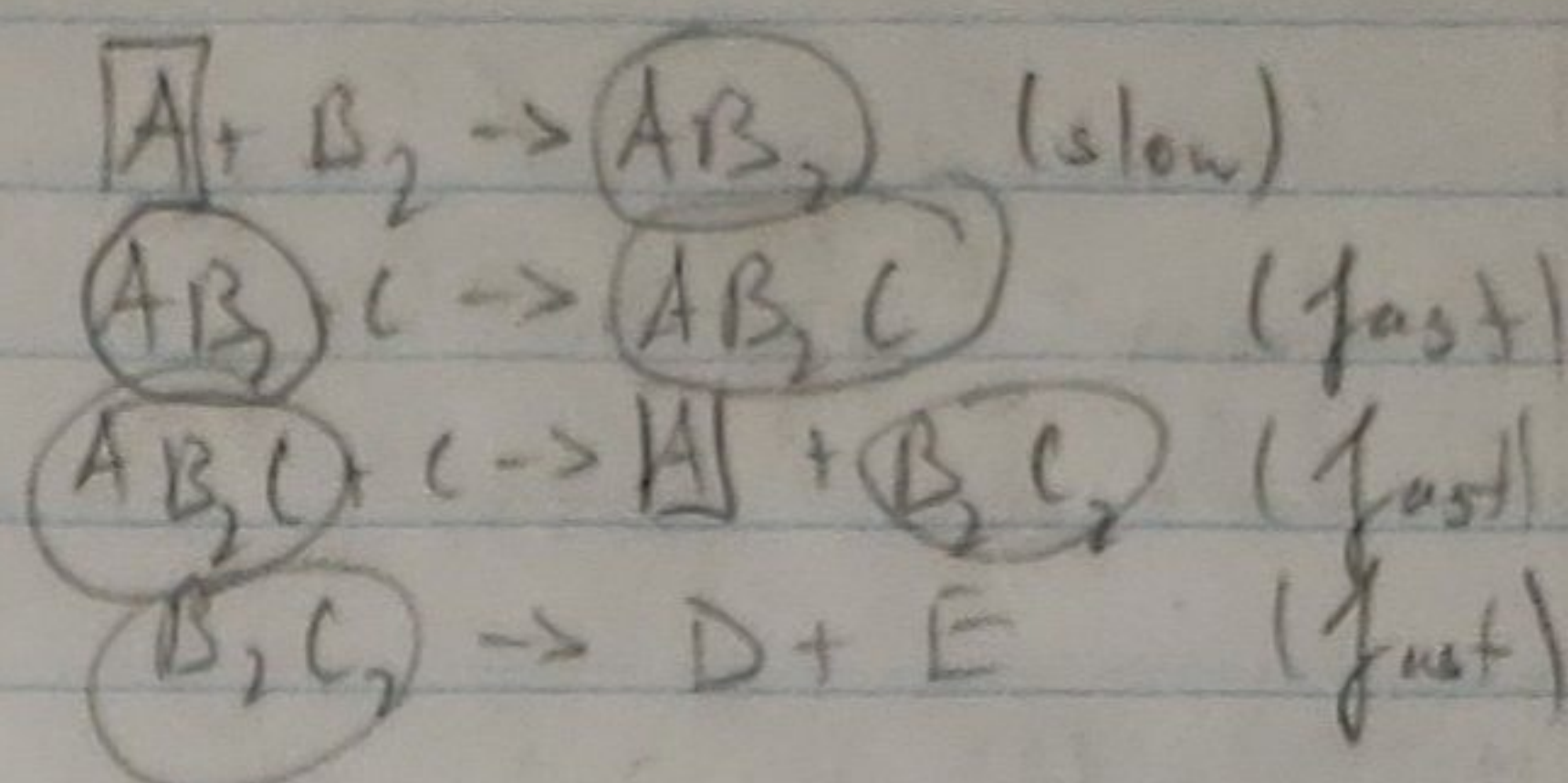
$$\frac{q}{m} = -5.8 \times 10^2 \text{ kJ/g}$$

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

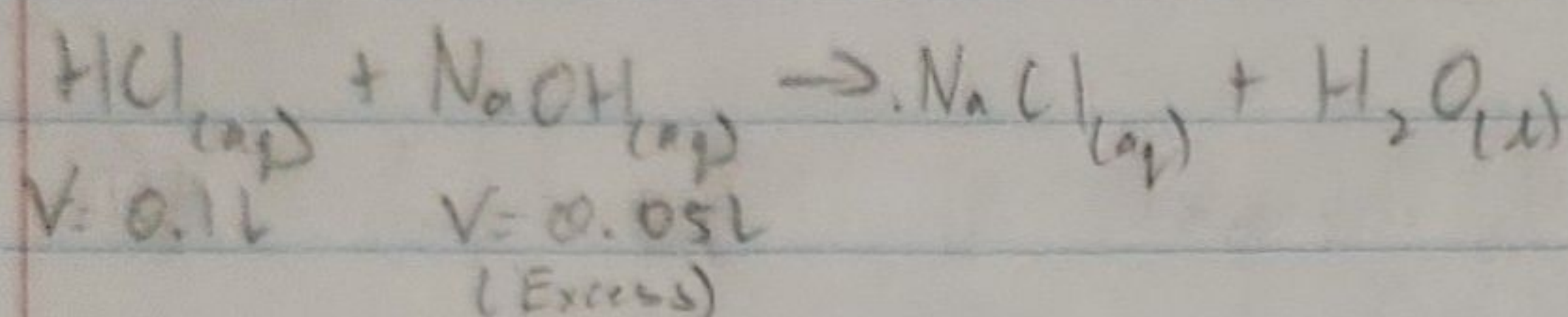


3.

- 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.



$$\begin{aligned} n_{HCl} &= cV \\ &= (0.355 M)(0.1L) \\ &= 0.0355 \text{ mol} \end{aligned}$$

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

$$\begin{aligned} q_{\text{system}} &= -mc\Delta T \\ &= -(150 g)(4.18 J/g^\circ C)(4.2^\circ C) \quad \leftarrow \text{Assume } 1 \text{ mL} = 1 g, \text{ Use } c = 4.18 J/g^\circ C \\ &= -2633.4 J \end{aligned}$$

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

$$= -2633.4 J$$

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

$$= -74180.3 J/\text{mol} \approx -74.2 kJ/\text{mol}$$

$\therefore$  The molar enthalpy change is  $-74.2 kJ/\text{mol}$

5.

$$\begin{aligned} \frac{\Delta[NH_3]}{\Delta t} &= \frac{([NH_3]_f - [NH_3]_i)}{\Delta t} \\ &= \frac{-(\frac{1.6}{4.0} M - \frac{3.0}{4.0} M)}{180s} \\ &= 2.6 \times 10^{-3} \text{ mol/L.s} \end{aligned}$$

- 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  $NH_3$ .

$$\begin{aligned} \frac{\Delta[H_2O]}{\Delta t} &= \left(\frac{3}{2}\right) \frac{\Delta[NH_3]}{\Delta t} \\ &= \left(\frac{3}{2}\right) (2.6 \times 10^{-3} \text{ mol/L.s}) \\ &= 4.0 \times 10^{-3} \text{ mol/L.s} \end{aligned}$$

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5b. cont)

The average rate at which  $\text{H}_2\text{O}$  is formed is  $4.0 \times 10^{-3} \text{ mol/l.s}$

5d)

mols of  $\text{O}_2$  are consumed in a 5:4 proportion with  $\text{NH}_3$ .

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

The average rate at which  $\text{O}_2$  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  $k$  = order of rxn w.r.t.  $[\text{Br}_2]$

$$2^k = 4$$

$$k = 2$$

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:

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

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

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

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