

Calorimetry + Hess's Law Drylab

Part 1: Data Analysis

1. Graph the reactions and determine corrected final temperature.

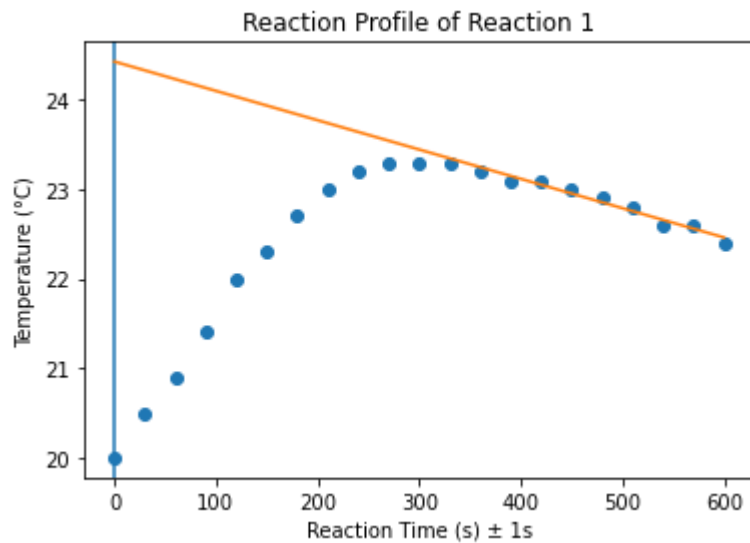
```
import numpy as np
import matplotlib.pyplot as plt

time = np.array([]);
for x in range(0,21):
    time = np.append(time, x * 30)
r1 = np.array([20.0, 20.5, 20.9, 21.4, 22.0, 22.3,
22.7,23,23.2,23.3,23.3,23.3,23.2,23.1,23.1,23.0,22.9,22.8,22.
6,22.6,22.4])
plt.plot(time, r1, 'o');
plt.axvline(x=0)

m, b = np.polyfit(time[12:], r1[12:], 1)
plt.plot(time, m * time + b )

plt.title('Reaction Profile of Reaction 1')
plt.xlabel('Reaction Time (s) ± 1s')
plt.ylabel('Temperature (°C)')

plt.show()
print('Corrected Final Temperature of Reaction 1: ', "
{:.1f}".format(b), '°C')
```



Corrected Final Temperature: 24.4 °C

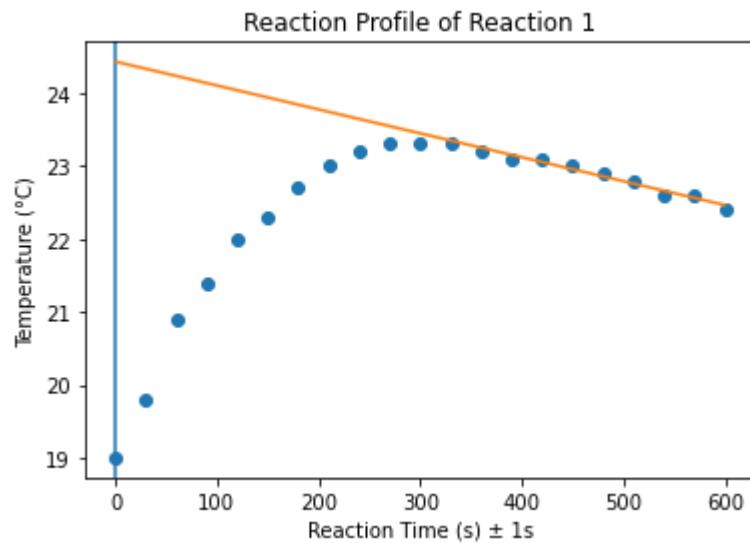
```
import numpy as np
import matplotlib.pyplot as plt

time = np.array([]);
for x in range(0,21):
    time = np.append(time, x * 30)
r1 = np.array([19.0, 19.8, 20.9, 21.4, 22.0, 22.3, 22.7,
23.0, 23.2, 23.3, 23.3, 23.3, 23.2, 23.1, 23.1, 23.0, 22.9,
22.8, 22.6, 22.6, 22.4])
plt.plot(time, r1, 'o');
plt.axvline(x=0)

m, b = np.polyfit(time[12:], r1[12:], 1)
plt.plot(time, m * time + b )

plt.title('Reaction Profile of Reaction 1')
plt.xlabel('Reaction Time (s)  $\pm 1s$ ')
plt.ylabel('Temperature (°C)')

plt.show()
print('Corrected Final Temperature: ', "{:.1f}".format(b),
'°C')
```



Corrected Final Temperature: 24.4 °C

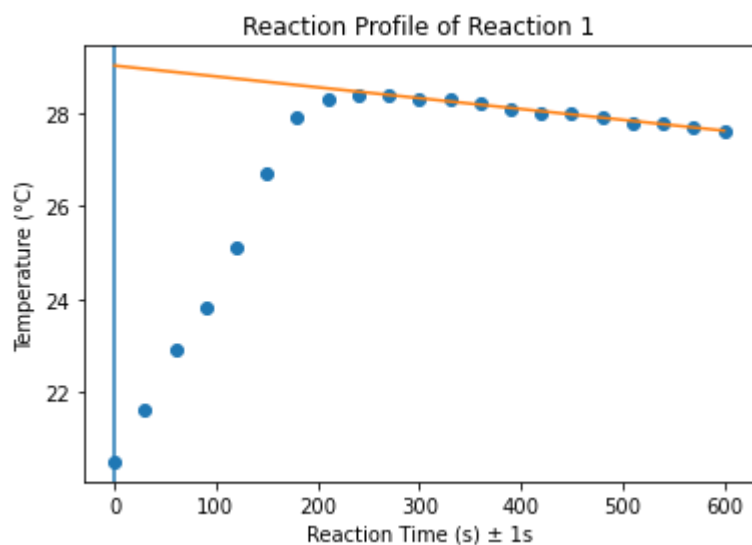
```
import numpy as np
import matplotlib.pyplot as plt

time = np.array([]);
for x in range(0,21):
    time = np.append(time, x * 30)
r1 = np.array([20.5, 21.6, 22.9, 23.8, 25.1, 26.7, 27.9,
28.3, 28.4, 28.4, 28.3, 28.3, 28.2, 28.1, 28.0, 28.0, 27.9,
27.8, 27.8, 27.7, 27.6])
plt.plot(time, r1, 'o');
plt.axvline(x=0)

m, b = np.polyfit(time[12:], r1[12:], 1)
plt.plot(time, m * time + b )

plt.title('Reaction Profile of Reaction 1')
plt.xlabel('Reaction Time (s) ± 1s')
plt.ylabel('Temperature (°C)')

plt.show()
print('Corrected Final Temperature: ', "{:.1f}".format(b),
'°C')
```



Corrected Final Temperature: 29.0 °C

2. Calculate the ΔH of each reaction

$$q_1 = mc \Delta T$$

$$= (100 \text{ ml})(1 \text{ g/1 ml})(4.18 \text{ J/g}^\circ\text{C})(24.3^\circ\text{C} - 20.0^\circ\text{C})$$

$$= 1.80 \times 10^3 \text{ J}$$

$$q_2 = mc \Delta T$$

$$= (50 \text{ ml} + 50 \text{ ml})(1 \text{ g/1 ml})(4.18 \text{ J/g}^\circ\text{C})(24.4^\circ\text{C} - 19.0^\circ\text{C})$$

$$= 2.26 \times 10^3 \text{ J}$$

$$q_3 = mc \Delta T$$

$$= (100 \text{ ml})(1 \text{ g/1 ml})(4.18 \text{ J/g}^\circ\text{C})(29.0^\circ\text{C} - 20.5^\circ\text{C})$$

$$= 3.55 \times 10^3 \text{ J}$$

3. Determine molar enthalpy of relevant reactants or products

$$M_{\text{NaOH}} = 40.0 \text{ g/mol}$$

$$m_{\text{NaOH}} = 2.00 \text{ g/mol}$$

$$\begin{aligned} n_{\text{NaOH}} &= \frac{M_{\text{NaOH}}}{m_{\text{NaOH}}} \\ &= \frac{2.00 \text{ g/mol}}{40.0 \text{ g/mol}} \\ &= 0.05 \text{ mol} \end{aligned}$$

$$\begin{aligned} \Delta H_1 &= -\frac{q_1}{n} \\ &= -\frac{1.80 \times 10^3 \text{ J}}{0.05 \text{ mol}} \\ &= -35.9 \text{ kJ/mol} \end{aligned}$$

$$\begin{aligned} n_{\text{HCl}} &= (50.0 \text{ ml})(1 \text{ mol}/1 \text{ dm}^3)(1 \text{ dm}^3/1000 \text{ cm}^3) \\ &= 0.05 \text{ mol} \end{aligned}$$

$$\begin{aligned} \Delta H_2 &= -\frac{q_2}{n} \\ &= -\frac{2.26 \times 10^3 \text{ J}}{0.05 \text{ mol}} \\ &= -45.2 \text{ kJ/mol} \end{aligned}$$

$$n = 0.05 \text{ mol}$$

$$\begin{aligned}\Delta H_3 &= -\frac{q_3}{n} \\ &= -\frac{3.55 \times 10^3 \text{ J}}{0.05} \\ &= -71.0 \text{ kJ/mol}\end{aligned}$$

4. Determine if Hess's Law stays true based on calculated values

$$\Delta H_1 + \Delta H_2 = \Delta H_3$$

$$(-35.9 \text{ kJ/mol}) + (-45.2 \text{ kJ/mol}) \neq -71.0 \text{ kJ/mol}$$

Based on the experimental data collected after correction for heat loss to environment, the ΔH_1 and ΔH_2 values do not sum up to equal to the ΔH_3 value. Therefore, Hess's Law could not be verified based on this experiment.

5. Calculate the % error based on literature values

$$\text{Percent Error} = \left| \frac{\text{Experimental} - \text{Literature}}{\text{Literature}} \right|$$

$$\text{Reaction 1} = \frac{-35.9 \text{ kJ/mol} - (-44.5 \text{ kJ/mol})}{-44.5 \text{ kJ/mol}}$$

$$= 19.3 \%$$

$$\text{Reaction 2} = \left| \frac{-45.2 \text{ kJ/mol} - (-55.8 \text{ kJ/mol})}{-55.8 \text{ kJ/mol}} \right|$$

$$= 19.0 \%$$

$$\text{Reaction 3} = \Delta H_2 + \Delta H_3$$

$$= -44.5 \text{ kJ/mol} + (-55.8 \text{ kJ/mol})$$

$$= -100.3 \text{ kJ/mol}$$

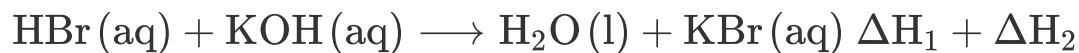
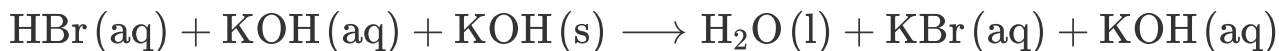
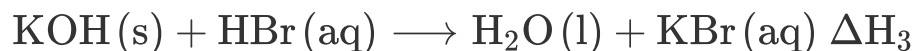
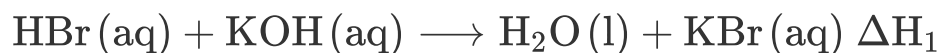
$$\% \text{ Error} = \left| \frac{-71.0 \text{ kJ/mol} + (-100.3 \text{ kJ/mol})}{-100.03 \text{ kJ/mol}} \right|$$

$$= 29.2 \%$$

The error range affected the ability to verify Hess's Law as the percentage error between the experimental and literature values did not match. Based on the experimental data, Hess's law could not be verified.

Part 2: Practice Questions

1. Use Hess Law to show ΔH_3 as a result of ΔH_1 and ΔH_2



The third equation can be achieved by adding up the first and second equation, therefore, according to Hess's Law, the enthalpy change of the third chemical equation must be the sum of the first two ΔH values.

2. Calculate the molar enthalpy for each of the following reactions

$$q = mc\Delta T$$

$$q_1 = (100 \text{ g} + 100 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(22.5^\circ\text{C} - 20.0^\circ\text{C})$$

$$= 2.09 \text{ kJ}$$

$$n_{\text{HBr}} = (100 \text{ ml})(1 \text{ mol}/1 \text{ dm}^3)(1 \text{ dm}^3/1000 \text{ ml})$$

$$= 0.1 \text{ mol}$$

$$\Delta H = \frac{2.09 \text{ kJ}}{0.1 \text{ mol}}$$

$$= 20.9 \text{ kJ}$$

$$q_2 = (200 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(24.1^\circ\text{C} - 20.0^\circ\text{C})$$

$$= 3.43 \text{ kJ}$$

$$n_{\text{KOH}} = \frac{5.61 \text{ g}}{56.10 \text{ g/mol}}$$

$$= 0.1 \text{ mol}$$

$$\Delta H = \frac{3.43 \text{ kJ}}{0.1 \text{ mol}}$$

$$= 34.3 \text{ kJ}$$

$$q_3 = (200 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(26.7^\circ\text{C} - 20.0^\circ\text{C})$$

$$= 5.60 \text{ kJ}$$

$$n_{\text{KOH}} = \frac{5.62 \text{ g}}{56.10 \text{ g/mol}}$$

$$= 0.1 \text{ mol}$$

$$\Delta H = \frac{5.60 \text{ kJ}}{0.1 \text{ mol}}$$

$$= 56.0 \text{ kJ}$$

3. Find the percentage error of the experiment

$$\Delta H_1 + \Delta H_2 = 20.9 \text{ kJ} + 34.3 \text{ kJ}$$

$$= 55.2 \text{ kJ}$$

$$\Delta H_3 = 56.0 \text{ kJ}$$

$$\% \text{ Error} = \frac{56.0 \text{ kJ} - 55.2 \text{ kJ}}{55.2 \text{ kJ}} \times 100 \%$$

$$= 1.45 \%$$