Discussion (2 pts)

- Summarize and analyze data from Results section
  - o Is your experimental ΔH_r close to the theoretical ΔH_r?
  - o What do you think were the possible reasons that affected your results? Were there any errors or limitations in the procedures?
- How to improve the conditions in order to reach a better result?

Conclusion (1 pt)

- Restate the goal of this investigation
- Summarize the results

Information that might be helpful:

Substance	ΔH^o_f (kJ/mol)
Mg (s)	0
HCl (aq)	-167.2
MgO (s)	-601.8
$MgCl_2$ (aq)	-796.9
H_2 (g)	0
H_2O (l)	-285.8

[1]. How to calculate the theoretical ΔHr?

Find the values of ΔH^o_f in the appendix of your textbook, or refer to this box:

Theoretical $\Delta Hr = \Delta H^o_f = \sum \Delta H^o_{f(product)} - \sum \Delta H^o_{f(reactant)}$

- $\sum$: the sum of

[2]. Experimental ΔH_{r1} or $\Delta H_{r2} = q_r / n$ (q_{r1} and n_{Mg} for ΔH_{r1}; q_{r2} and n_{MgO} for ΔH_{r2})

$$q_r = - q_{sol}$$

$$q_{sol} = m_{sol} \cdot C_s \cdot (T_{final} - T_{initial})$$

- n: # of moles of Mg or MgO
- m_{sol} : mass of the HCl solution used in reaction (1) or (2)
- $C_{s(HCl\ solution)} \approx C_{s(H2O)} = 4.184$ J/(g·°C)

Include reaction (3) when calculating the experimental ΔH_r (*i.e.* the heat of combustion for Mg)!

(3) $H_2 (g) + ½ O_2 (g) \rightarrow H_2O (l)$ $\Delta Hr_3 = -285.8$ kJ/mol H_2 (or ½ mol O_2)

[3]. Standard deviation α:

$$\alpha = \sqrt{\frac{\sum_{i=1}^{n} (X_i - X)^2}{n-1}}$$

X_i: the result of the #i trial; X: the mean of all the results; n: the total number of the trials.

Since each reaction was performed for at least three times, error incorporates in data analysis. The final enthalpy of combustion should be reported with an error margin, *i.e.* standard deviation.

For example:

If reaction (1) was performed 3 times with the results: -460, -452 and -467 kJ/mol, the enthalpy reported will be average ± standard deviation = -459.7 ± 7.5 kJ/mol.

In this case, errors are additive; so if reaction 2 was reported with a standard deviation of 3.5 kJ/mol, the total standard deviation will be 7.5 + 3.5 = ± 11 kJ/mol.