Unit 4: AP Free Response Practice #4 [2013 FR #3, 10 points]

$$MgO(s) + 2\,H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$$

4. A student was assigned to the task of determining the enthalpy change for the reaction between solid MgO and aqeous HCl represented by the net ionic equation above. Ther student uses a polystyrene cup calorimeter and performs four trials. Data for each trial are shown in the table below.

Trial	Volume of 1.0 *M* HCl (mL)	Mass of MgO(*s*) Added (g)	Initial Temperature of Solution (°C)	Final Temperature of Solution (°C)
(b) ~~1~~	~~100.0~~	~~0.25~~	~~25.5~~	~~26.5~~
2	100.0	0.50	25.0	29.1
3	100.0	0.25	26.0	28.1
4	100.0	0.50	24.1	28.1

a. Which is the limiting reactant in all four trials, HCl or MgO? Justify your answer. [*1 point*]

b. The data in one of the trials is inconsistent with the data in the other three trials. Identify the trial with the inconsistent data and draw a line through the data from that trial in the table above. Explain how you identified the inconsistent data. [*1 point*]

For parts (c) and (d), use the data from one of the other three trials (i.e., not from the trial you identified in part (b) above.) Assume the calorimeter has a negligible heat capacity and that the specific heat of the contents of the calorimeter is 4.18 J/(g °C). Assume that the density of the HCl(aq) is 1.0 g/mL.

c. Calculate the magnitude of q, the thermal energy change, when the MgO was added to the 1.0 M HCl(aq). Include units with your answer. [*2 points*]

d. Determine the student's experimental value of ΔH° for the reaction between MgO and HCl in units of kJ/mol_{rxn}. [*2 points*]

e. Enthalpies of formation for substances involved in the reaction are shown in table below. Using the information in the table, determine the accepted value of ΔH° for the reaction between MgO(s) and HCl(aq). [*2 points*]

Substance	ΔH_f° (kJ/mol)
$MgO(s)$	−602
$H_2O(l)$	−286
$H^+(aq)$	0
$Mg^{2+}(aq)$	−467

f. The accepted value and the experimental value do not agree. If the calorimeter leaked heat energy to the environment, would it help account for the discrepancy between the values? Explain. [*1 point*]

a.) $1.0\,M \times 0.100\,L = 0.10\,mol\,HCl \times \frac{1\,MgO}{2\,HCl} = 0.050\,mol\,MgO$

$0.50\,g\,MgO \times \frac{1\,mol\,MgO}{40.3\,g\,MgO} = 0.0124\,mol\,MgO < 0.050\,mol\,MgO$

⇒ MgO is limiting! (or: ΔT changed when amount of MgO added was changed, so MgO must be limiting ☺)

b) Trial 1 is inconsistent b/c ΔT should be proportional to the amount of MgO added. In trials 2 & 4, 0.5 g MgO gave a $\Delta T \approx 4°C$, and in trial 3 0.25 g MgO gave $\Delta T \approx 2°C$, but in trial 1 the ratio is $\frac{1}{2}$ of the other trials.

c) $q_{cal} = mC\Delta T = (100.\text{g} + 0.50\text{g})(4.18\ \frac{J}{g°C})(29.1 - 25.0) = 1700\ J$ (using trial 2; b/c d = 1.0 g/mL)

d) $\Delta H^\circ_{rxn} = \frac{q_{rxn}}{mol_{rxn}}$ } $0.0124\ mol\ MgO \times \frac{1\ mol_{rxn}}{1\ mol\ MgO} = 0.0124\ mol_{rxn}$

$q_{rxn} = -q_{cal}$

$\Delta H^\circ_{rxn} = \frac{-1.7\ kJ}{0.0124\ mol_{rxn}} = \boxed{-140\ kJ/mol_{rxn}}$

e) $\Delta H^\circ_{rxn} = \Sigma \Delta H^\circ_f(pr) - \Sigma \Delta H^\circ_f(re)$

$= [Mg^{2+} + H_2O] - [MgO + 2H^+]$

$= [-467 + -286] - [-602 + 2(0)]$

$= -753 + 602 = \boxed{-151\ kJ/mol_{rxn}}$

f) Yes, b/c the experimental value of $-140\ kJ/mol_{rxn}$ was less negative than the accepted value of $-151\ kJ/mol_{rxn}$. If the calorimeter leaked heat energy, calculated ΔT would be less than expected, leading to a smaller calculated q + thus a less negative ΔH°.