Question 3
(9 points)

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

A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net-ionic equation above. The 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)
~~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.

$0.100\text{ L} \times \frac{1.0\text{ mol HCl}}{1.0\text{ L}} = 0.10\text{ mol HCl}$ $0.50\text{ g MgO} \times \frac{1\text{ mol MgO}}{40.30\text{ g MgO}} = 0.0124\text{ mol MgO}$ By the stoichiometry of the equation, only 2 × (0.0124 mol) = 0.025 mol HCl is needed to react with the MgO, thus HCl is in excess and MgO is limiting. OR The temperature change depended on the amount of MgO added, indicating that MgO was the limiting reactant.	1 point is earned for the correct choice with justification.

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

Trial 1 is inconsistent. The temperature change should be directly proportional (approximately) to the amount of the limiting reactant present. The ratio $\Delta T/(\text{mass MgO})$ should be constant. In trial 1, the ratio is one-half of trials 2, 3, and 4. Therefore, trial 1 is inconsistent with the other trials.	1 point is earned for identifying trial 1 with a valid justification.

Question 3 (continued)

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.

$q_{calorimeter} = q_{cal} = mc\Delta T$ In trial 2, $q_{cal} = \left[\left(100.0\text{ mL} \times \frac{1.0\text{ g}}{\text{mL}}\right) + 0.50\text{ g}\right]\left(\frac{4.18\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)\left(4.1^\circ\text{C}\right) = 1700\text{ J or } 1.7\text{ kJ}$ OR In trial 3, $q_{cal} = \left[\left(100.0\text{ mL} \times \frac{1.0\text{ g}}{\text{mL}}\right) + 0.25\text{ g}\right]\left(\frac{4.18\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)\left(2.1^\circ\text{C}\right) = 880\text{ J or } 0.88\text{ kJ}$ OR In trial 4, $q_{cal} = \left[\left(100.0\text{ mL} \times \frac{1.0\text{ g}}{\text{mL}}\right) + 0.50\text{ g}\right]\left(\frac{4.18\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)\left(4.0^\circ\text{C}\right) = 1700\text{ J or } 1.7\text{ kJ}$	1 point is earned for the correct mass of the solution. 1 point is earned for the correct calculation of q for any trial with a valid ΔT and correct units.

(d) Determine the student's experimental value of $\Delta H°$ for the reaction between MgO and HCl in units of kJ/mol_{rxn}.

Assuming that no heat was lost to the surroundings, $q_{rxn} = -q_{cal}$. In trials 2 and 4, $\Delta H° = \frac{q_{rxn}}{n_{MgO}} = \frac{-1{,}700\text{ J}}{0.50\text{ g MgO} \times \frac{1\text{ mol MgO}}{40.30\text{ g MgO}}} = -140{,}000\text{ J/mol}_{rxn} \times \frac{1\text{ kJ}}{1000\text{ J}}$ $= -140\text{ kJ/mol}_{rxn}$ In trial 3, $\Delta H° = \frac{-880\text{ J}}{0.25\text{ g MgO} \times \frac{1\text{ mol MgO}}{40.30\text{ g MgO}}} = -140{,}000\text{ J/mol}_{rxn} \times \frac{1\text{ kJ}}{1000\text{ J}}$ $= -140\text{ kJ/mol}_{rxn}$	1 point is earned for the correct calculation of moles of MgO or setup of equation. 1 point is earned for the value of $\Delta H°$ and sign consistent with the setup.

Question 3 (continued)

(e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of $\Delta H°$ for the reaction between $MgO(s)$ and $HCl(aq)$.

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

$\Delta H° = \sum n_p \Delta H_f^\circ \text{ products} - \sum n_r \Delta H_f^\circ \text{ reactants}$ $= \left[\Delta H_f^\circ\, Mg^{2+}(aq) + \Delta H_f^\circ\, H_2O(l)\right] - \left[\Delta H_f^\circ\, MgO(s) + 2\,\Delta H_f^\circ\, H^+(aq)\right]$ $= [-467 \text{ kJ/mol} + (-286 \text{ kJ/mol})] - [-602 \text{ kJ/mol} + 2(0) \text{ kJ/mol}]$ $= -151 \text{ kJ/mol}_{rxn}$	1 point is earned for the correct setup using the ΔH_f° values. 1 point is earned for the correct value and sign consistent with the setup.

(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.

Yes. The experimentally determined value for $\Delta H°$ was less negative than the accepted value. If heat had leaked out of the calorimeter, then the ΔT of the contents would be less than expected, leading to a smaller calculated value for q and a less negative value for $\Delta H°$.	1 point is earned for the correct response with a valid explanation.

3A1

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

3. A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net-ionic equation above. The 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)	ΔT
~~1~~	~~100.0~~	~~0.25~~	~~25.5~~	~~26.5~~	1
2	100.0	0.50	25.0	29.1	4.1
3	100.0	0.25	26.0	28.1	2.1
4	100.0	0.50	24.1	28.1	4.1

(a) Which is the limiting reactant in all four trials, HCl or MgO? Justify your answer.

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

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.

(d) Determine the student's experimental value of $\Delta H°$ for the reaction between MgO and HCl in units of kJ/mol_{rxn}.

(e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of $\Delta H°$ for the reaction between MgO(*s*) and HCl(*aq*).

Substance	$\Delta H_f°$ (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.

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$M_{MgO} = 24.30\frac{g}{mol} + 16.00\frac{g}{mol} = 40.30\frac{g}{mol}$

3.a. MgO is limiting in all four trials. 0.50g, the largest MgO sample, would react with:

$\frac{0.50g\ MgO}{40.30\frac{g}{mol}} \cdot \frac{2\ mol\ H^+}{1\ mol\ MgO} = 2.5 \times 10^{-2}\ mol\ HCl$

While in each trial, $1.0M \cdot \frac{100.0\ mL}{1000\ \frac{mL}{L}} = 0.1000\ mol\ HCl$ were present.

3.b. Trial I is inconsistent with the others. Both 0.50g MgO trials had a ΔT of 4.1°C, so with half the amount of MgO, the ΔT should be about half of that. This is the case with trial 3 ($\Delta T = 2.1 \approx \frac{4.1}{2}$), but not with trial 1. ($\Delta T = 1 \neq \frac{4.1}{2}$).

3.c. Using trial 2:

$q = mc_p\Delta T = (100mL \cdot 1.0\frac{g}{mL})(4.18\frac{J}{g°C})(29.1°C - 25.0°C) = 1700J$

3.d. $\Delta H° = -1700J \cdot \left(\frac{0.50g\ MgO}{40.30\frac{g}{mol}}\right)^{-1} \cdot \frac{1\ kJ}{1000J} = -140\ \frac{kJ}{mol}$

3.e. $\Delta H° = (-467\frac{kJ}{mol} - 286\frac{kJ}{mol}) - (-602\frac{kJ}{mol}) = -151\frac{kJ}{mol}$

3.f. Yes, it would explain the discrepancy. If the calorimeter lost heat, there would be a lower measured ΔT, which would result in a lower calculated q, which would result in a lower calculated $\Delta H°$ (since q is proportional to ΔT, and $\Delta H°$ is proportional to q).

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1 mol | mol MgO | 2 mol HCl |

40.35 /mol 3B

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

3. A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net-ionic equation above. The 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)
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.

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

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.~~

(d) Determine the student's experimental value of $\Delta H°$ for the reaction between MgO and HCl in units of kJ/mol_{rxn}.

(e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of $\Delta H°$ for the reaction between MgO(s) and HCl(aq).

Substance	ΔH_f° (kJ/mol)
MgO(s)	−602
$H_2O(l)$	−286
$H^+(aq)$	0
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(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.

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3B2

a) $\frac{.1 L\ H^+}{} \Big| \frac{1\ mol\ H^+}{1 L} \Big| \frac{1\ mol\ Mg^{2+}}{2\ mol\ H^+} \Big| = .05\ mol\ Mg^{2+}$

$\frac{.25 g\ MgO}{} \Big| \frac{1\ mol\ MgO}{40.3 g} \Big| \frac{1\ mol\ Mg^{2+}}{1\ mol\ MgO} \Big| = .006\ mol\ Mg^{2+}$

$\frac{.5 g\ MgO}{} \Big| \frac{1\ mol\ MgO}{40.3 g\ MgO} \Big| \frac{1\ mol\ Mg^{2+}}{1\ mol\ MgO} \Big| = .012\ mol\ Mg^{2+}$

in all experiments, MgO is the limiting reactant because it produces fewer moles of product.

b) 1: $\Delta t = 1°c$ 2: $\Delta t = 4.1$ 3: $\Delta t = 2.1$ 4: $\Delta t = 4$

for trials: 2,3 and 4, the Δt doubled when the mass of MgO doubled. trial 1 did not follow that pattern and $q = mc\Delta t$ is a proportional relationship.

c) $\frac{100 mL\ HCl}{} \Big| \frac{1 g\ HCl}{1 mL} \Big| = 100 g\ HCl$

trial 2: $q_{H_2O} = mc(\Delta t)$

$= (100.5 g)(4.18)(4.1)$

$= 1722.37\ J$

d) $\frac{-1722.37 J}{.0124\ mol\ MgO} \Big| \frac{1 kJ}{1000 J} \Big| = -138.82\ kJ/mol\ MgO$

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e) $\Delta H_f^\circ = [-467 + -286] - [-602 + 2(0)]$
$= -148$ kJ/mol

f) yes, that explanation would help account for the discrepancy. If the true value of ΔH_f° is -148 kJ/mol, the measured value has too small a magnitude. This means not as much of a heat release was registered in the experiment as it should have had. The ΔT could have been too low because some energy was lost to the environment.

3C1

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

3. A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net-ionic equation above. The 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)
~~1~~	~~100.0~~	~~0.25~~	~~25.5~~	~~26.5~~
2	100.0	0.50	25.0	29.1
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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.

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

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.

(d) Determine the student's experimental value of $\Delta H°$ for the reaction between MgO and HCl in units of kJ/mol_{rxn}.

(e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of $\Delta H°$ for the reaction between MgO(*s*) and HCl(*aq*).

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(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.

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3C 2

PAGE FOR ANSW

3a. $MgO + 2H^+ \rightarrow Mg^{2+} + H_2O$

Mols of H^+ used = $.1L \times 1.0M = .1$ mols $H^+ \times \frac{1 mol H_2O}{2 mol H^+} = .05$ mol H_2O

mols of MgO = $.25g \times \frac{1 mol}{43g} = .0058$ mols $MgO \times \frac{1 mol H_2O}{1 mol MgO} = .0058$ mols H_2O

MgO is the limiting reactant. By comparing the yields based solely off of the mols of each reactant (as shown above), it is clear that the MgO will run out first in this reaction

3b. The trial with inconsistent data is Trial 1. Trial 1 uses the same volume of HCl and mass of Mg as Trial 3, so it is expected that the change in temp is approximately 2°C, but Trial 1 only has a 1°C change. Also, in trials 2 & 4, double the mass of MgO is used and the temperature change is 4°C, so it follows that a trial using half of that mass of MgO would have a temp change of 1/2 that of ~~trial~~ 2 & 4. this happens in trial 3, leaving Trial 1 as the inconsistent one.

3c. $q = mC\Delta T$ *Using data from trial 4.

$C = \frac{4.18 J}{g \cdot C°}$

$\Delta T = 28.1°C - 24.1°C = 4.0°C$

$m = .50g$

$q = (.50g)(\frac{4.18J}{g°C})(4.0°C)$

$q = 8.36 J$

3d. $\Delta H° = \frac{4.18 J}{g°C} \times (4.0°C)$

$= 16.72 J/g \times \frac{1 kJ}{1000 J} \times \frac{43 g}{1 mol} = .719$ kJ/mol

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3C3

3e $\Delta H^\circ = \Sigma$ products H°_f $- \Sigma$ reactants H°_f

$= (H^\circ_{Mg^{2+}} + H^\circ_{H_2O}) - (H^\circ_{MgO} + 2H^\circ_{H^+})$

$(-602\ kJ/mol + -286\ kJ/mol) - (-467\ kJ/mol)$

$\Delta H^\circ = -421\ kJ/mol$

3f If the calorimeter leaked heat into the environment, the change in temperature recorded would have been smaller than what it was supposed to be, which would cause the value of q or the heat given off, to be smaller. This would help account for the discrepancy

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Question 3

Overview

This question assessed ability to use collected experimental data for calculations involving stoichiometry, thermodynamics, error analysis, and reporting values with the correct number of significant digits. In part (a) students used experimental data to determine the limiting reactant for the reaction given. In part (b) students examined the data to determine which of the four trials contained inconsistent data and justified their choice. Part (c) required students to use the experimental data to determine the heat released (q) in the reaction. In part (d) students then used the data from part (c) to calculate the experimental value of $\Delta H°$. In part (e) students calculated the accepted value of $\Delta H°$ from enthalpies of formation given in a table. Part (f) asked students to examine the experimental and accepted $\Delta H°$ values and a possible source of error to determine if that error could have accounted for the discrepancy between the accepted and experimental values.

Sample: 3A
Score: 9

This response earned 9 points. For part (a) the number of moles of HCl present is calculated and the number of moles needed to react with the largest sample of MgO is calculated. The mole comparison is correct and the correct limiting reactant is predicted. The point was earned in part (b) because the student uses an argument comparing all four trials. Part (c) earned 2 points. The mass of the solution in the calorimeter is calculated correctly and the final answer is correct (correct math, number of significant figures, and unit). Part (d) earned 2 points. The setup of $\Delta H° = -q/n$ is correct. Numbers are substituted into the equation correctly to get the answer with the correct number of significant figures and the negative sign. Part (e) earned 2 points. A correct setup and number substitution with correct math and significant figures earned both points. The point was earned in part (f). A comparison of absolute values of $\Delta H°$ was accepted.

Sample: 3B
Score: 7

This response earned 7 points. In part (c) the correct mass of the solution in the calorimeter is determined, the correct formula is used, and the correct unit is used. The answer contains six significant figures rather than the appropriate two, so 1 point was earned. Part (d) earned 2 points. The answer given also has too many significant figures, but because this response had already lost the significant figure point in part (c), no point was deducted for this error. Part (e) earned 1 point for the correct setup and number substitution. The calculated value is incorrect.

Sample: 3C
Score: 5

This response earned 5 points. Part (c) earned 1point. The total mass of the contents of the calorimeter is not used, so the first point was not earned. The incorrect mass of the calorimeter is used to calculate a consistent value of q, earning the second point. The points were not earned in part (d) because the student used a formula that was not correct. Using this setup, it is not possible to correctly compute a value of $\Delta H°$. Part (e) earned 1 point. A correct formula is used but incorrect numbers are substituted into the formula, so the point was not earned. Using the incorrect values, an answer with correct math and sign earned the second point.