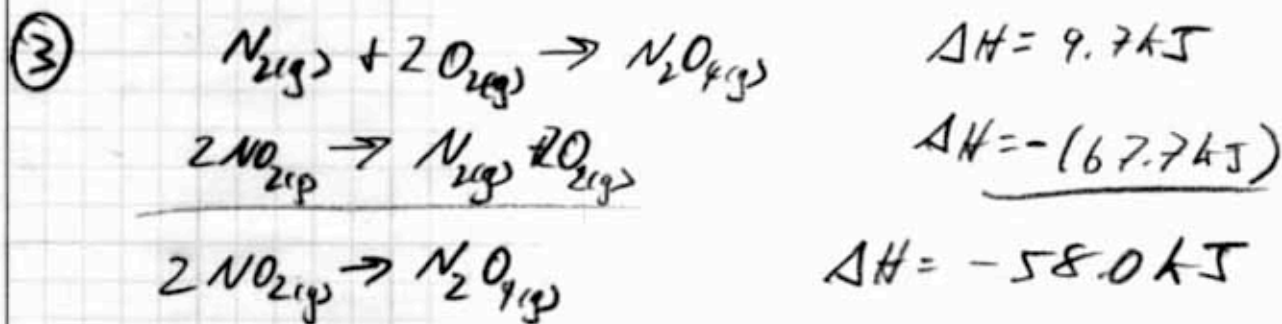


- ② A) -
 B) +



$$\begin{array}{l}
 \textcircled{4} \quad \Delta H^\circ = [\Delta H_{\text{fSiO}_2}^\circ + 4\Delta H_{\text{fHF}}^\circ + 4\Delta H_{\text{fCl}}^\circ] - [\Delta H_{\text{fSiCl}_4}^\circ + 2\Delta H_{\text{fH}_2\text{O}}^\circ] \\
 \Delta H^\circ = [-911 + 4 \cdot 0 + 4 \cdot -167] - [-687 + 2 \cdot -286] = -320 \text{ kJ}
 \end{array}$$

⑤ Find limiting reactant

$$30.5 \text{ g SiCl}_4 \cdot \frac{1 \text{ mol SiCl}_4}{169.8 \text{ g}} \cdot \frac{2 \text{ H}_2\text{O}}{1 \text{ SiCl}_4} \cdot \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 6.47 \text{ g H}_2\text{O needed}$$

so SiCl₄ limiting

$$30.5 \text{ g SiCl}_4 \cdot \frac{1 \text{ mol SiCl}_4}{169.8 \text{ g}} \cdot \frac{-320 \text{ kJ}}{1 \text{ SiCl}_4} = -57.4 \text{ kJ} = \Delta H$$

57.4 kJ released.

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$$\textcircled{6} \text{ A) } \Delta H^\circ = [2\Delta H_{\text{FeCl}_2}^\circ + 6\Delta H_{\text{H}_2\text{O}}^\circ] - [2\Delta H_{\text{Fe}_2\text{O}_3}^\circ + 3\Delta H_{\text{CO}_2}^\circ + 2\Delta H_{\text{Fe}_2\text{H}_2}^\circ]$$

$$\Delta H^\circ = [(2 \cdot 135.1 \text{ kJ}) + (6 \cdot -242 \text{ kJ})] - [(2 \cdot -46) + 3 \cdot 0 + (2 \cdot -75)]$$

$$\Delta H^\circ = -939.8 \text{ kJ}$$

$$\text{B) } \Delta H^\circ = [3\Delta H_{\text{Fe}_2\text{SO}_4}^\circ + 2\Delta H_{\text{H}_2\text{PO}_4}^\circ] - [\Delta H_{\text{Fe}_2(\text{PO}_4)_2}^\circ + 3\Delta H_{\text{H}_2\text{SO}_4}^\circ]$$

$$\Delta H^\circ = [(3 \cdot -1433) + (2 \cdot -1267)] - [(-4126) + (3 \cdot -814)]$$

$$\Delta H^\circ = -265$$

$$\text{C) } \Delta H^\circ = [\Delta H_{\text{NH}_4\text{Cl}}^\circ] - [\Delta H_{\text{NH}_3}^\circ + \Delta H_{\text{HCl}}^\circ]$$

$$\Delta H^\circ = [-314 \text{ kJ}] - [-46 + -92]$$

$$\Delta H^\circ = -176 \text{ kJ}$$

$$\textcircled{7} \left(\frac{1}{3} \text{Fe}_2\text{O}_3 + \frac{1}{3} \text{CO}_2 \rightarrow \frac{1}{3} \text{Fe}_2\text{O}_4 + \frac{1}{3} \text{CO}_2 \right)^{\times 3} \quad \Delta H = -\frac{1}{3} (38 \text{ kJ})$$

$$\left(\frac{1}{2} \text{Fe}_3\text{O}_4 + \frac{1}{6} \text{CO}_2 \rightarrow \frac{1}{2} \text{Fe}_2\text{O}_3 + \frac{1}{6} \text{CO} \right)^{\times 6} \quad \Delta H = -\frac{1}{6} (-596 \text{ kJ})$$

$$\left(\frac{1}{2} \text{Fe}_2\text{O}_3 + \frac{2}{3} \text{CO}_2 \rightarrow \frac{1}{2} \text{Fe}_2\text{O}_4 + \frac{2}{3} \text{CO}_2 \right)^{\times 2} \quad \Delta H = \frac{1}{2} (-28 \text{ kJ})$$

$$\text{Fe}_2\text{O}_3 + \text{CO}_2 \rightarrow \text{Fe}_2\text{O}_4 + \text{CO}_2 \quad \Delta H = -16.8 \text{ kJ}$$

$$\textcircled{8} \Delta H = [0 + -110.5] - [0 + -242] = 131.5 \frac{\text{kJ}}{\text{mol}}$$

Stoich -

$$50.0 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.01 \text{ g}} \cdot \frac{1 \text{ mol H}_2}{1 \text{ mol C}} = 4.16 \text{ mol H}_2$$

$$V = \frac{4.16 \cdot 0.0821 \cdot 373}{1 \text{ atm}} = 127 \text{ L H}_2 \text{ produced}$$

So use mol H₂ to find ΔH

$$4.16 \text{ mol H}_2 \cdot \frac{131.5 \text{ kJ}}{1 \text{ mol H}_2} = 5.47 \cdot 10^2 \text{ kJ absorbed.}$$

- 9 Calculate amount of energy needed.
 $q = m \cdot \Delta T \cdot C_p$

$$q = 500.0g \cdot (100 - 25) \cdot 4.18 \frac{J}{g^\circ C} = 1.57 \cdot 10^5 J$$

157 kJ

Take energy convert to time

$$157 kJ \cdot \frac{1 \text{ min}}{400 J} = 392 \text{ min}$$

- 10 knowing that the energy lost = energy gained
 we can say:

$$q_{\text{metal}} = -q_{\text{H}_2\text{O}}$$

set up equations for the q_{metal} and $q_{\text{H}_2\text{O}}$

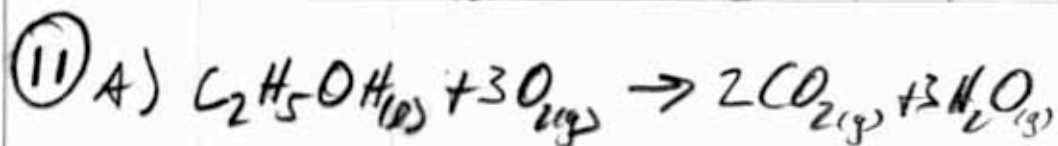
$$q_{\text{metal}} = 33.2g \cdot (24.8 - 58.3) \cdot C_p$$

$$q_{\text{H}_2\text{O}} = 100.0g \cdot (24.8 - 23.7) \cdot 4.18 \frac{J}{g^\circ C}$$

now combine equations

$$33.2g \cdot (24.8 - 58.3) \cdot C_p = -[100.0g \cdot (24.8 - 23.7) \cdot 4.18]$$

$$C_p = 0.413 \frac{J}{g^\circ C}$$



B) The key here is that the energy released by the reaction goes in to the water AND the calorimeter its self, so

$$-q_{C_2H_5OH} = q_{cal} + q_{H_2O}$$

Now we need equations for q_{cal} and q_{H_2O} .

$$q_{cal} = 654 \frac{J}{^{\circ}C} \cdot (29.49 - 25.12) = 2.86 \cdot 10^3 J$$

$$q_{H_2O} = 1.100 \cdot 10^3 J \cdot \Delta T \cdot 4.18 \frac{J}{^{\circ}C} = 1.100 \cdot 10^3 \cdot (29.49 - 25.12) \cdot 4.18$$

$$= 2.01 \cdot 10^4 J$$

$$\text{so } q_{C_2H_5OH} = -(q_{cal} + q_{H_2O}) = -(2.86 \cdot 10^3 J + 2.01 \cdot 10^4 J)$$

$$q_{C_2H_5OH} = -2.30 \cdot 10^4 J$$

now divide by grams and convert to kJ

$$\frac{-2.30 \cdot 10^4 J}{0.853 g} \cdot \frac{1 kJ}{1000 J} = -26.9 \frac{kJ}{g}$$

C) This is the Enthalpy of combustion ΔH .

$$-26.9 \frac{kJ}{g} \cdot \frac{46.07 g}{1 mol} = -1.24 \cdot 10^3 \frac{kJ}{mol}$$

⑪ D) Don't Do

E) we know $\Delta H_{\text{comb}}^{\circ}$ from C

$$\Delta H_{\text{comb}}^{\circ} = [2\Delta H_{\text{f,CO}_2}^{\circ} + 3\Delta H_{\text{f,H}_2\text{O}}^{\circ}] - [\Delta H_{\text{f,C}_2\text{H}_5\text{OH}}^{\circ} + 3\Delta H_{\text{f,O}_2}^{\circ}]$$

$$-26.9 \text{ kJ} = [2 \cdot (-395.5) + 3 \cdot (-241.8)] - [\Delta H_{\text{f,C}_2\text{H}_5\text{OH}}^{\circ} + 3 \cdot 0]$$

$$\Delta H_{\text{f,C}_2\text{H}_5\text{OH}}^{\circ} = -2.76 \cdot 10^2 \frac{\text{kJ}}{\text{mol}}$$

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(12) To determine this value we would mix 1.0L of 1.0M of acetic acid with 1.0L of 1.0M of NaOH, we would measure the initial temperatures and the final temp of the solution. Assume volumes are additive and solutions have same density as water. we could then calculate the energy change of the water with.

$$q_{H_2O} = m \cdot \Delta T \cdot C_p$$

where $C_p = 4.18 \frac{J}{g \cdot ^\circ C}$ and mass can be found from volume and density of water ($\frac{1g}{mL}$).

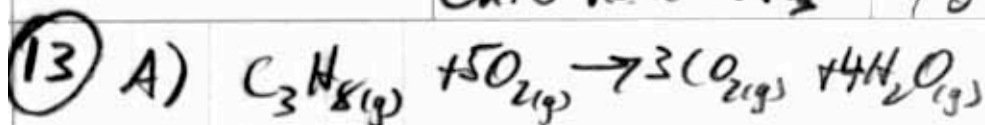
So we know the amount of energy change of the water we also know this is related to the energy change of the reaction by.

$$q_{rxn} = -q_{H_2O}$$

We know that to find ΔH we need to take q_{rxn} and divide by the moles used which would be 1.00 moles so

$$\Delta H = \frac{q_{rxn}}{1.00m}$$

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B) So we need to take the amount of energy generated and divide by the moles
So find moles

$$1.000g C_3H_8 \cdot \frac{1 \text{ mol } C_3H_8}{44.09g C_3H_8} = 2.27 \cdot 10^{-2} \text{ mol}$$

$$\Delta H = \frac{-50.33 \text{ kJ}}{2.27 \cdot 10^{-2} \text{ mol}} = -2.22 \cdot 10^3 \frac{\text{kJ}}{\text{mol}}$$

C) So we will use the combustion equation
where we know

$$\Delta H_{\text{comb}} = [3 \cdot \Delta H_{f,CO_2}^{\circ} + 4 \Delta H_{f,H_2O}^{\circ}] - [\Delta H_{f,C_3H_8}^{\circ} + 5 \Delta H_{f,O_2}^{\circ}]$$

$$-2.22 \cdot 10^3 \frac{\text{kJ}}{\text{mol}} = [(3 \cdot -392 \text{ kJ}) + (4 \cdot -285)] - [\Delta H_{f,C_3H_8}^{\circ} + 0]$$

$$\Delta H_{f,C_3H_8}^{\circ} = -104.5 \text{ kJ/mol}$$

⑭ A) Coffee cup calorimeter
graduated cylinders
thermometers

B) Volume of HCl + NaOH
Temp. of HCl + NaOH at start
Temp of mix after reaction

C) Energy change = mass \cdot specific heat \cdot change in Temp

To calculate q take the volume of each solution and assume the density of the solutions is 1 g/mL then the volume = mass.

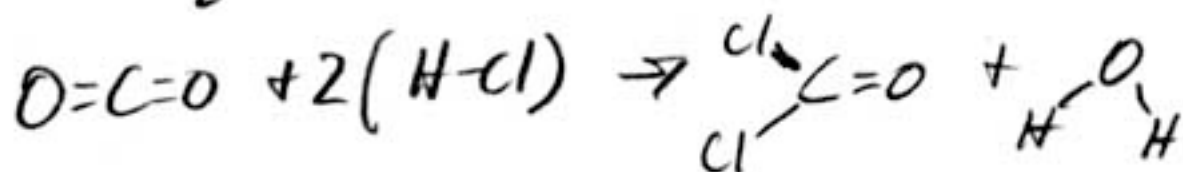
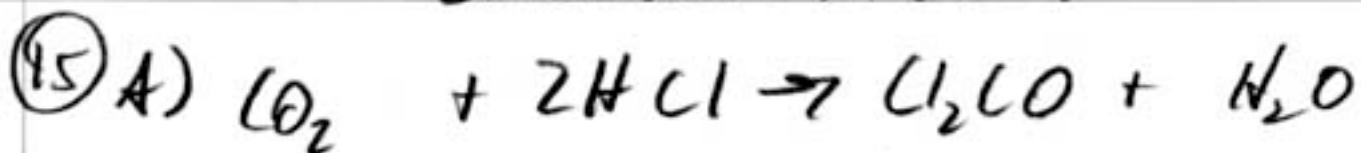
Assume the specific heat capacity is $4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$

The change in temperature would be the final temperature minus the initial temperature

Then $q = m \cdot s \cdot \Delta T$. Finally $\Delta H = \frac{q}{\text{mol}}$ The mol = $M \cdot V$

D). The measured value is lower than the accepted value. This means the amount of energy transferred to the calorimeter is less than it should be this could be due to heat lost to the surroundings. Another possibility is that the solutions were actually less concentrated than 1 M. This would mean less reaction took place

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B) $\Delta H = [2(\text{C}=\text{O}) + 2(\text{H}-\text{Cl})] - [2(\text{C}-\text{C}) + (\text{C}=\text{O}) + 2(\text{H}-\text{O})]$

$\Delta H = [2(715) + 2(431)] - [2(331) + (715) + 2(464)]$

$\Delta H = -13 \text{ kJ/mol}$

C) $\Delta H = -$ **Exothermic**