Honors Chemistry

Chapter 17 – Hess's Law & Assorted Thermochemistry

1. Hess's law of heat summation: you can use ΔH of known reactions to find unknown ΔH

2. When the reverse of a chemical reaction is written, what happens to the sign of ΔH? Why?
   The sign reverses b/c now its the opposite reaction

3. Calculate the enthalpy change, ΔH (in kJ) for the reaction:

   $BaO\ (s) + H_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + H_2O\ (l)$

   Use the following data:

   Flip/Reverse $SO_3\ (g) + H_2O\ (l) \rightarrow H_2SO_4\ (aq)$ $\Delta H = -78.2$ kJ

   $BaO\ (s) + SO_3\ (g) \rightarrow BaSO_4\ (s)$ $\Delta H = -213.4$ kJ

   + $H_2SO_4\ (aq) \rightarrow SO_3\ (g) + H_2O\ (l)$ + $\Delta H = +78.2$ kJ

   $BaO\ (s) + H_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + H_2O\ (l)$ $\Delta H = -135.2$ kJ

4. The following equation shows the combustion of ethane.

   $2C_2H_6\ (g) + 7O_2\ (g) \rightarrow 4CO_2\ (g) + 6H_2O\ (l)$

   Use Hess's law and following data to calculate the enthalpy change. $\Delta H = -2856$ kJ

   $2C_2H_6\ (g) + 7O_2\ (g) \rightarrow 4CO_2\ (g) + 6H_2O\ (l)$

   multiply by 2: $2(C_2H_4\ (g) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 2H_2O\ (g)$ $\Delta H = -1323$ kJ$)2$

   Start here. multiply by 2, then reverse: $2(C_2H_4\ (g) + H_2\ (g) \rightarrow C_2H_6\ (g)$ $\Delta H = -137$ kJ$)2$

   multiply by 2: $2(H_2\ (g) + \frac{1}{2} O_2\ (g) \rightarrow H_2O\ (g)$ $\Delta H = -242$ kJ$)2$

   $2C_2H_6\ (g) \rightarrow 2C_2H_4\ (g) + 2H_2\ (g)$ $\Delta H = +274$ kJ

   $2C_2H_4\ (g) + 6O_2\ (g) \rightarrow 4CO_2\ (g) + 4H_2O\ (g)$ $\Delta H = -2646$ kJ

   + $2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (g)$ + $\Delta H = -484$ kJ

5. Calculate the enthalpy change, ΔH (in kJ) for the reaction:

   $C\ (s) + 2H_2\ (g) \rightarrow CH_4\ (g)$

   Use the following data:

   $C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$ $\Delta H = -393.5$ kJ

   multiply by 2. $2(H_2\ (g) + \frac{1}{2} O_2\ (g) \rightarrow H_2O\ (l)$ $\Delta H = -285.8$ kJ$)2$

   Start here. Reverse $CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (l)$ $\Delta H = -890.3$ kJ

   $CO_2\ (g) + 2H_2O\ (l) \rightarrow CH_4\ (g) + 2O_2\ (g)$ $\Delta H = +890.3$ kJ

   + $2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (l)$ + $\Delta H = -571.6$ kJ

   $C\ (s) + 2H_2\ (g) \rightarrow CH_4\ (g)$ $\Delta H = -74.8$ kJ

6. Calculate the standard molar enthalpy of formation, ΔH_f°, for the following using table 17.4 on page 580.

   $\Delta H_f^\circ = \Sigma \Delta H^\circ_{prod} - \Sigma \Delta H^\circ_{react}$

   Work on another paper

   a) $H_2\ (g) + \frac{1}{2} O_2\ (g) \rightarrow H_2O\ (l)$

   b) $C\ (s) + 2H_2\ (g) \rightarrow CH_4\ (g)$

   c) $2Al\ (s) + 3/2\ O_2\ (g) \rightarrow Al_2O_3\ (s)$

   d) $\frac{1}{2} N_2\ (g) + 3/2\ H_2\ (g) \rightarrow NH_3\ (g)$

$\uparrow D_{H_2O} = 1\ \frac{g}{mL}$ ∴ 250 mL = 250 g

⊖ΔH lost = ⊕ΔH gained

7. If 80.0 g of aluminum initially at 70°C is dropped into 250.0 mL of water the temperature of the water rises to 22°C. What was the initial temperature of the water. The specific heat of aluminum is 0.215 cal/g°C.

⊖ΔH lost Al = ⊕ΔH gain H_2O

Aluminum: m = 80.0 g; Cp = 0.215 cal/g°C; T_i = 70°C; T_f = 22°C

Water: m = 250 g; Cp = 1 cal/g°C; T_i = ?; T_f = 22°C

Find ΔH lost Al: $\Delta H = mc_p\Delta T = 80 \cdot 0.215 \frac{cal}{g°C} \cdot (22-70) = ⊖825.6\ cal$

ΔH gain H_2O = ⊕ 825.6 cal

Find T_i H_2O: $\Delta H = mc_p\Delta T$

$825.6\ cal = 250 \cdot 1 \frac{cal}{g°C} \cdot (22 - T_i)$

$T_i = 18.7°C$

⊖ΔH lost = ⊕ΔH gain

8. A piece of unknown metal with a mass of 23.8 g is heated to 100.0°C and dropped into 50 mL of water at 24.0°C. The final temperature of the system is 32.5°C. What is the specific heat of the metal? (in J/g°C)

50 mL → 50 g

Metal: m = 23.8 g; Cp = ?; T_i = 100°C; T_f = 32.5°C

Water: m = 50.0 g; Cp = 4.18 J/g°C; T_i = 24°C; T_f = 32.5°C

⊖ΔH lost metal = ⊕ΔH gain H_2O

Find ΔH gain H_2O: $\Delta H = mc_p\Delta T = 50 \cdot 4.18 \frac{J}{g°C} \cdot (32.5-24) = ⊕1776.5\ J$

ΔH lost metal = ⊖1776.5 J

Find Cp metal: $\Delta H = mc_p\Delta T$

$⊖1776.5\ J = 23.8 \cdot c_p \cdot (32.5-100)$

$C_p = 1.11 \frac{J}{g°C}$

Phase Change

9. How much heat is released when 108 g of water at 0°C freezes to ice.
$\Delta H_{solid} = -6.02\ kJ/mol$

$\frac{108 g\ H_2O}{1} \cdot \frac{1\ mol\ H_2O}{18\ g\ H_2O} = 5.99\ mol\ H_2O$

$\Delta H = mol \cdot \Delta H_{solid}$

$= 5.99\ mol \cdot ⊖6.02 \frac{kJ}{mol}$

$\Delta H = ⊖36.1\ kJ$

⊖ΔH lost = ⊕ΔH gain

10. You burn 1.5 g of benzoic acid, $C_7H_6O_2$ (aq), in a combustion calorimeter and find that the temperature of the calorimeter increases from 22.5°C to 31.69°C. The calorimeter contains 775 g of water. Calculate the molar heat of combustion, ΔH (in kJ/mol), of benzoic acid.

$$2C_7H_6O_2\,(aq) + 15O_2\,(g) \rightarrow 14CO_2\,(g) + 6H_2O\,(g)$$

Benzoic Acid: m = 1.5 g; Cp; T_i; T_f = 31.69°C

Water: m = 775 g; Cp = 4.18 J/g°C; T_i = 22.5°C; T_f = 31.69°C

Find ΔH gain H_2O: $\Delta H = mc_p\Delta T = 775 \cdot 4.18 \frac{J}{g°C} \cdot (31.69-22.5) = ⊕29771\ J$

ΔH lost Benzoic Acid = ⊖29771 J. This is for 1.5 g $C_7H_6O_2$: $\frac{1.5 g}{1} \cdot \frac{1\ mol}{122\ g} = 0.0123\ mol\ C_7H_6O_2$

* For the full reaction *: $\frac{⊖29771\ J}{0.0123\ mol} = 2420406.5 \frac{J}{mol} \times \frac{1\ kJ}{1000\ J}$

$\Delta H = ⊖2420 \frac{kJ}{mol}$

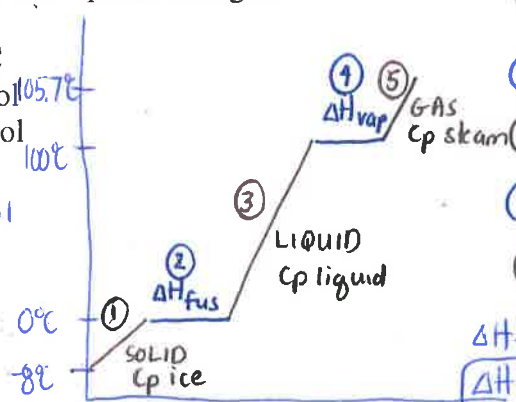
Multi-step

11. Calculate the heat energy required to raise the temperature of 4.5 g of water from -8°C to 105.7°C. Draw a diagram to show phase changes.

$C_{ice} = 2.06\ J/g°C$
$C_{steam} = 2.02\ J/g°C$
$\Delta H_{fus} = 6.02\ kJ/mol$
$\Delta H_{vap} = 40.7\ kJ/mol$

$C_p = 4.18 \frac{J}{g°C}$ liquid

$\frac{4.5 g\ H_2O}{1} \cdot \frac{1\ mol}{18.02 g} = 0.25\ mol$



① $\Delta H = 4.5 g \cdot 2.06 \frac{J}{g°C} \cdot (0 - -8) = ⊕74.2\ J$ or ⊕0.0742 kJ

② $\Delta H = 0.25\ mol \cdot 6.02 \frac{kJ}{mol} = ⊕1.51\ kJ$ or ⊕1510 J

③ $\Delta H = 4.5 \cdot 4.18 \frac{J}{g°C} \cdot (100-0) = ⊕1881\ J$ or ⊕1.881 kJ

④ $\Delta H = 0.25\ mol \cdot 40.7 \frac{kJ}{mol} = ⊕10.18\ kJ$ or ⊕10,175 J

⑤ $\Delta H = 4.5 g \cdot 2.02 \frac{J}{g°C} \cdot (105.7-100) = ⊕51.8\ J$ or ⊕0.0518 kJ

$\Delta H_{Total} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5$

ΔH_{Total} = 13,692 J or 13.7 kJ

Thermochemical Eqn. Use fencepost

12. How much heat energy is released when 80.0 of oxygen gas reacts?

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) + 2220\ kJ$$

$\frac{80.0 g\ O_2}{1} \cdot \frac{1\ mol\ O_2}{32.0\ g\ O_2} \cdot \frac{2220\ kJ}{5\ mol\ O_2}$ = ⊖1110 kJ

Temperature Change

13. The temperature of a 6.0-g sample of glass changed from 20°C to 45°C when it absorbed 550 J of heat. What is the specific heat of this glass sample?

m = 6.0 g; Cp = ?; T_i = 20°C; T_f = 45°C; ΔH = 550 J

$\Delta H = mc_p\Delta T$

$550\ J = 6.0 g \cdot c_p \cdot (45-20)$

$C_p = 3.67 \frac{J}{g°C}$