1. A sample of 50. mL of a 0.20 *M* solution of HCl was mixed with 50. mL of 0.20 *M* NaOH in a coffee cup calorimeter. The initial temperature of both solutions was 22.2°C. After mixing, the temperature rose to 23.5°C. What is the enthalpy change for the neutralization reaction which occurred?

2. 0.510 g of ethanol is burned in a flame calorimeter containing 1200 g of water. The water is initially at 22.46°C and is warmed up to 25.52°C as a result of the reaction. What is the ΔH for one mole of ethanol?

3. 300 mL of 0.2 M aqueous KOH neutralizes 150 mL of aqueous 0.2 M H_2SO_4. We go from an average initial temperature of 22.3 °C to a maximum of 29.2 C. Calculate the molar heat(enthalpy) of neutralization of KOH.

4. Find the final temperature of the following mixture: $C_{Cu} = 0.39$ J/[g·°C]
   400. g of Cu initially at 99.0 °C
   25 L of water initially at 10.0 °C

5. Example: Use the following standard enthalpies of combustion to calculate the enthalpy change for the following reaction
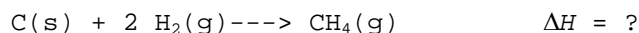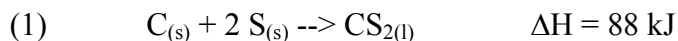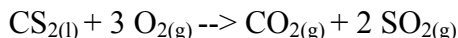
   ```
   C(s) + 2 H₂(g)---> CH₄(g)            ΔH = ?

   1) ΔH°comb(C)   = -393.5 kJ
   2) ΔH°comb(H₂)  = -285.8 kJ
   3) ΔH°comb(CH₄) = -890.4 kJ
   ```

6. Carbon disulfide is a very flammable solvent. Calculate ΔH for the following combustion reaction:

   $$CS_{2(l)} + 3\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ SO_{2(g)}$$

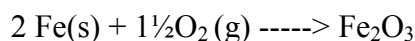 | | | |
   |---|---|---|
   | (1) | $C_{(s)} + 2\ S_{(s)} \rightarrow CS_{2(l)}$ | $\Delta H = 88$ kJ |
   | (2) | $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ | $\Delta H = -394$ kJ |
   | (3) | $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$ | $\Delta H = -297$ kJ |

7. Carbon monoxide is often used in metallurgy to remove oxygen from metal oxides and thereby give the free metal. The thermochemical equation for the reaction of CO with iron(III) oxide, Fe_2O_3, is

   $$Fe_2O_3(s) + 3\ CO(g) \longrightarrow 2\ Fe(s) + 3\ CO_2(g) \qquad \Delta H° = -26.74\ kJ$$
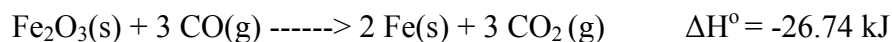
   Use this equation and the equation for the combustion of CO

   $$CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g) \qquad \Delta H° = -283.0\ kJ$$

   to calculate the value of H° for the reaction

   $$2\ Fe(s) + 1\tfrac{1}{2}O_2(g) \longrightarrow Fe_2O_3$$

Solutions

1. $q = 4.18$ J K^{-1}g^{-1} × 100. g × -1.3 K
  $= -540.$ J

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2. $C_2H_5OH(l) + 3\ O_2(g) \text{---}> 2CO_2(g) + 3\ H_2O(l)$
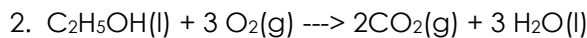
$\Delta T_{water} = 25.52°C - 22.46°C = 3.06°C$

$$\left(\frac{4.18J}{K\ g\ H_2O}\right)(1200\ g\ H_2O)(3.06\ K) = 15.3\ kJ$$

This is the total heat absorbed by the water and is the negative of the heat evolved by the system.
Hence,

$\Delta H' = q = -15.3$ kJ

$\dfrac{0.510\ g\ ethanol}{46.05\ g/mol} = 0.0111$ mol ethanol

$\Delta H = q = \dfrac{-15.3\ kJ}{0.0111\ mol\ ethanol} = -1380$ kJ/mol ethanol

(1 mol ethanol in 1 mole reaction)
$q = -1380$ kJ/mol ethanol or simply -1380 kJ/mol.

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3. $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2\ H_2O$
$q = mc\ \Delta T$
$= (300 + 150\ g)(4.18)(29.2-22.3)$
$=12978.9$ J

$\Delta H = - q = -12978.9$ J

$n = cv = 0.2mole/L(0.300\ L) = 0.06$ moles

$\Delta H/n = -12978.9 J/0.06$ moles $= -216$ kJ/mole of KOH

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4. $q_{H2O} = - q_{Cu}$ (because the heat lost by Cu will be gained by the water)

OR   q water + q copper = 0

$m_{H2O}C_{H2O}\ (T_f - T_{H2O}) = -m_{Cu}C_{Cu}\ (T_f - T_{Cu})$ Note that they reach the same final temperature.
25000g (4.18)(T_f-10) = -400 (0.39)(T_f -99)
$T_f = 10.1$ C = 10. °C

The temperature change is insignificant (especially if we pay attention to significant figures) because the amount of water is much greater than the amount of Cu, and because water has a higher specific heat than Cu.

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5.   Knowing the standard format for the combustion reaction (1 mol of reactant burns to completion in oxygen to produce carbon dioxide and water) we can easily write down the three reactions.
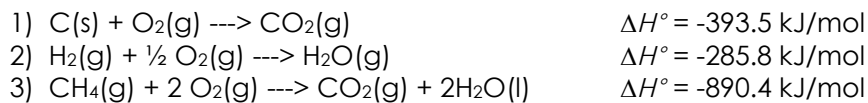
1) $C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H° = -393.5$ kJ/mol
2) $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(g)$ $\Delta H° = -285.8$ kJ/mol
3) $CH_4(g) + 2 O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$ $\Delta H° = -890.4$ kJ/mol
Now lets add up the three reactions to give the desired reaction

1) $C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H° = -393.5$ kJ/mol
2)$2\times[H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(g)$ $\Delta H° = -285.8$ kJ/mol]
3) $CO_2(g) + 2H_2O(l) \longrightarrow CH_4(g) + 2 O_2(g)$ $\Delta H° = +890.4$ kJ/mol

_____

   $C(s) + 2 H_2(g) \longrightarrow CH_4(g)$ $\Delta H° = -74.7$ kJ/mol

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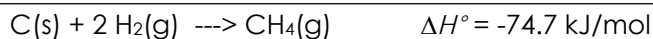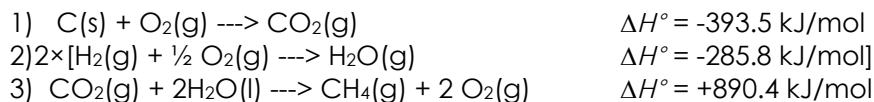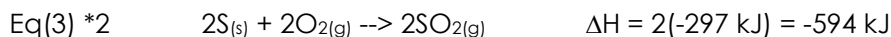
6.   Equation (1) is the only one with $CS_{2(l)}$, but it has the compound as a product. Our target equation has $CS_{2(l)}$ on the left hand side. So we reverse equation (1) and change the sign of DH.

reverse(1) $CS_{2(l)} \longrightarrow C_{(s)} + 2 S_{(s)}$ $\Delta H = - 88$ kJ
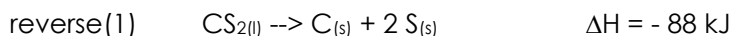
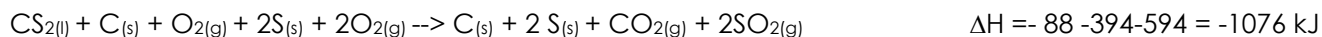   Equation (2) is the only one with $CO_{2(g)}$, and, compared to the target equation, it has the correct amount on the desired side of the equation. So we keep equation (2) as is.

    (2) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H = -394$ kJ
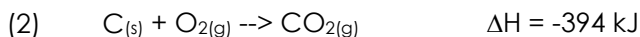
            Equation (3) is the only one with $SO_{2(g)}$. It 's on the correct side of the equation, but the target equation needs twice as much. So we multiply equation (3) by 2.

Eq(3) *2 $2S_{(s)} + 2O_{2(g)} \longrightarrow 2SO_{2(g)}$ $\Delta H = 2(-297$ kJ$) = -594$ kJ

Now if we add the above three equations:

$CS_{2(l)} + C_{(s)} + O_{2(g)} + 2S_{(s)} + 2O_{2(g)} \longrightarrow C_{(s)} + 2 S_{(s)} + CO_{2(g)} + 2SO_{2(g)}$ $\Delta H =- 88 -394-594 = -1076$ kJ
Notice that $C_{(s)}$ and $2S_{(s)}$ cancel. We are left with the target equation
$CS_{2(l)} + 3 O_{2(g)} \longrightarrow CO_{2(g)} + 2 SO_{2(g)}$ and of course the ΔH which we needed.

Answer: $\Delta H = -1076$ kJ

7.
Step 1   The target equation must have 2 Fe on the left, but the first equation above has 2 Fe on the right. To move it left, reverse the entire equation and remember to reverse the sign of H^o. When the equation is flipped over the Fe_2O_3 also falls into the correct position.

Step 2 There must be $1\frac{1}{2}O_2$ of the left, and we must be able to cancel *three* CO and *three* CO_2 when the equations are added. Multiply the second equation by 3 and we get the necessary coefficients. Multiply the H^o values for the second equation by 3 as well. The adjusted equations are:

        $2 Fe(s) + 3 CO_2(g) \longrightarrow Fe_2O_3(s) + 3 CO(g)$ $\Delta H^o = +26.74$ kJ
        $\underline{3 CO(g) + 1\frac{1}{2}O_2(g) \longrightarrow 3 CO_2(g)}$ $\underline{\Delta H^o = 3(-283.0 \text{ kJ}) = -849.0 \text{ kJ}}$
        $2 Fe(s) + 1\frac{1}{2}O_2 (g) \longrightarrow Fe_2O_3$ $\Delta H^o = -822.26$ kJ