Answer on Question#59876 - Chemistry | General Chemistry

A) calculate the missing ΔH values (using Hess's Law) in the table below:
use $\Delta H_f\ CO_2(g) = -394 kj\ mol^{-1}$ AND $\Delta H_f\ H_2O(l) = -286 kj\ mol^{-1}$

	ΔH_f/k J mol^{-1}	ΔH_c/k J mol^{-1}
CH_4		-890
$C_2H_4(g)$		-1409
$C_4H_{10}(g)$	-125	
$CH_3OH(l)$		-715
$C_2H_5OH(l)$	-278	
CH_3COOH		-876
$CH_3COOC_2H_5$	-481	

B) use the data to calculate ΔH for the reaction
$CH_3COOH(l) + C_2H_5OH(l) \longrightarrow CH_3COOC_2H_5(l) + H_2O(l)$

Answers:

A) The equation we want for the formation of methane from its elements would be the following:

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

You are given the three equations:

1. $C(s) + O_2(g) \rightarrow CO_2(g)$..... $\Delta H_1 = -394\ kJmol^{-1}$
2. $H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(g)$..... $\Delta H_2 = -286\ kJmol^{-1}$
3. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$..... $\Delta H_3 = -890\ kJmol^{-1}$

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Equation 1 already has carbon on the left side. Equation 2 already has hydrogen on the left side, but we need two hydrogen molecules, so we can double the second equation. Equation 3 has methane on the left side instead of on the right side, so we need to reverse it.

1. $C(s) + O_2(g) \rightarrow CO_2(g)$..... $\Delta H_1 = -394\ kJmol^{-1}$
2. $2H2(g) + O2(g) \rightarrow 2\ H2O(g)$..... $\Delta H_2 = -572\ kJmol^{-1}$
3. $CO_2(g) + 2\ H_2O(g) \rightarrow CH_4(g) + 2\ O_2(g)$..... $\Delta H_3 = +890\ kJmol^{-1}$

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Eliminating oxygen, carbon dioxide, and water on both sides in all the three equations, we end up with:

Total: $C(s) + 2\ H_2(g) \rightarrow CH_4(g)$.....$\Delta H = -76\ kJmol^{-1}$

The standard heat of formation for methane is $-76\ kJmol^{-1}$.

Answer: $\Delta H_f\ (CH_4) = -76\ kJmol^{-1}$

B) The combustion reaction is

$C_2H_4(g) + 3O_2(g) => 2CO_2(g) + 2H_2O(g)$

On the right hand side of the above expression we know the enthalpy of formation of the CO_2 and the H_2O from their atoms.

It is 2(-394) + 2(-286) = - (788 + 572) = -1360 kJ. This means that in forming 2 moles of $CO_2(g)$ and two moles of $H_2O(l)$ from their atoms, 1360 kJ are released.
To get there from the left hand side of the equation, we must first form one mole of C_2H_4 from its atoms and then burn it with 3 moles of O_2. ΔH_c is the heat of combustion.
$\Delta Hf[C_2H_4(g)] + \Delta Hc[C_2H_4(g)] = 2\Delta H°f[CO_2(g)] + 2\Delta H°f[H_2O(l)]$
$\Delta Hf[C_2H_4(g)] - 1411kJ = -1358.6$ kJ.
$\Delta Hf[C_2H_4(g)] = 52.4$ kJ.....and since we were talking about one mole of ethylene, that's **52.4 kJ/mole.**
Answer: ΔH_f (C_2H_4) =52.4 kJmol^{-1}

C) Complete combustion of butane:
$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$
Using standard enthalpies of formation, we imagine the reaction takes place by one mole of each starting molecule breaking apart into their elements in their standard states, and then reforming into the products. Obviously this is not really what happens but thermodynamically it is equivalent.
Here's some data:
ΔHf butane(g) = -125 kJ/mol $4C(s) + 5H_2(g) \rightarrow C4H10(g)$
ΔHf $CO_2(g)$ = -394 kJ/mol $C(s) + O_2(g) \rightarrow CO_2(g)$
ΔHf $H_2O(g)$ = -286.0 kJ/mol $H_2(g) + 0.5O_2(g) \rightarrow H_2O(l)$
We can simply reverse the sign of the ΔHf for the breaking apart of the molecule, as opposed to its formation.
So our reaction comprises:
$2C_4H_{10}(g) \rightarrow 8C(s) + 10H_2(g)$ +125 kJ/mol x 2 = +250.0 kJ/mol
$8O_2(g) + 8C(s) \rightarrow 8CO_2(g)$ -394 kJ/mol x 8 = -3152 kJ/mol
$10H_2(g) + 5O_2(g) \rightarrow 10H_2O(l)$ -286 kJ/mol x 10 = -2860.0 kJ/mol
Adding these values together:
+250.0 kJ/mol + -3152.0 kJ/mol + -2860.0 kJ/mol = -5762.0 kJ/mol
We then half this to calculate the standard enthalpy of combustion, since in the above equation we used 2 moles of butane to keep whole numbers:
-5762 kJ/mol /2 = **-2881 kJ/mol**
Answer: ΔH_c (C_4H_{10}) =-2881 kJmol^{-1}

D) $CH_3OH + 3/2O_2 \rightarrow CO_2 + 2H_2O$ $\Delta H = -726.4$ kJ/mol
$C(graphite) + O_2 \rightarrow CO_2$ $\Delta H = -393.5$ kJ/mol
$H_2 + 1/2\ O_2 \rightarrow H_2O$ $\Delta H = -285.8$ kJ/mol
Calculate the enthalpy of formation of methanol (CH_3OH) from its elements:
$C(graphite) + 2H_2 + 1/2O_2 \rightarrow CH_3OH$
As for Hess's law:
-715 kJ mol-1 = 2x(-286 kJ mol^{-1}) + -394 kJ mol^{-1} - $\Delta H(CH_3OH)$
$\Delta H(CH_3OH)$ **= -251 kJ mol^{-1}**
Answer: ΔH_f (CH_3OH) =-251 kJmol^{-1}

E) The balanced equation I we for ethanol is: $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$ And from that, the three equations we have are:
$C + O_2 \rightarrow CO_2$ ... $\Delta H = -394$ kJ/mol

H_2 + 0.5OH -> H_2O ... ΔH = -286 kJ/mol
2C + 3H_2 + 0.5O_2 -> C_2H_5OH ... Δ H = -278 kJ/mol
Then we got:
(278) - (2 x 394) - (3 x 286)= **-1368 kJ/mol**
Answer: ΔH_c (C_2H_5OH) =-1368 kJmol^{-1}

J) C+O_2 -> CO_2 ΔH_1=-394 kJ/mol
H_2+1/2O_2 -> H_2O ΔH_2 = -286 kJ/mol
CH_3COOH+2O_2 -> 2CO_2+2H_2O ΔH_3 = 876 kJ/mol
2C+O_2+2H_2 -> CH_3COOH ΔH_4 -?
Therefore H_4=2*H_1 + 2*H_2 - H_3 = -394 – 286+876 =196 **kJ/mol**
Answer: ΔH_f (CH_3COOH) = 196 kJ mol^{-1}

F) The balanced equation I we for ethyl acetate is:
$CH_3COOC_2H_5$ (l) + 3O_2(g) -> 4CO_2(g) + 4H_2O(g)
And from that, the three equations we have are:
C+O_2 -> CO_2 ΔH_1=-394 kJ/mol
H_2+1/2O_2 -> H_2O ΔH_2 = -286 kJ/mol
$CH_3COOC_2H_5$+2O_2 -> 4CO_2(g) + 4H_2O(g) ΔH_3 =- 481 kJ/mol
$CH_3COOC_2H_5$ ->4C+O_2+4H_2 ΔH_4 -?
Therefore H_4=2*H_1 + 2*H_2 - H_3 =(-4·394)+(4·-286)+481**=-2239 kJ/mol**
Answer: ΔH_c ($CH_3COOC_2H_5$) =-2239 kJmol^{-1}

B) CH_3COOH (l) + C_2H_5OH(l) -> $CH_3COOC_2H_5$ (l) + 4H_2O (l) ΔH_1 - ?
This could be thought of as the result of two hypothetical reactions, going via the elements:
CH_3COOH (l) + C_2H_5OH(l) -> 4 C (s)+5H_2 (g)+1.5O_2 (g) **ΔH_2**
followed by 4 C (s)+5H_2 (g)+1.5O_2 (g) -> $CH_3COOC_2H_5$ (l) + 4H_2O **ΔH_3**
ΔH_2 is –[ΔH_f C_2H_5OH (l) + ΔH_f CH_3COOH (l)] as it is the reverse of the formation of the compounds from the element, and ΔH_3 is even more obviously ΔH_f ($CH_3COOC_2H_5$ (l)) + ΔH_f (H_2O (l)) as it is the formation of the elements from their compounds. Applying Hess's Law:
ΔH_1 = ΔH_2 + ΔH_3 = –[ΔH_f C_2H_5OH (l) + ΔH_f CH_3COOH (l)]+[ΔH_f ($CH_3COOC_2H_5$ (l)) + ΔH_f (H_2O (l))]
Substituting in appropriate values ΔH_f (C_2H_5OH (l)) =-1367 kJ/mol; ΔH_f CH_3COOH (l) = -874 kJ/mol; ΔH_f ($CH_3COOC_2H_5$ (l))=-2239 kJ/mol; ΔH_f (H_2O (l)) = -286 kJ/mol:
ΔH_1 = -[(-1367)+(-874)]+[(-2239)+(-286)]=2241 – 2525 = -284 kJ/mol
Answer: -284 kJ/mol