Beckfoot

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Enthalpy Change

1	Enthalpy change (ΔH) is the heat energy change measured under conditions of constant pressure.
2	Standard enthalpy changes refer to standard conditions ie100 kPa and a stated temperature (eg $\Delta H_{298}^{\ominus}$).
3	Exothermic reaction release energy into the surroundings so the temperature increases. The products have less energy than the reactants so ΔH is negative.
4	Endothermic reaction take in energy from the surroundings so the temperature decreases. The products have more energy than the reactants so ΔH is positive.

Bond Enthalpy

1	Atoms in molecules are held together by strong covalent bonds. It takes energy to break them (endothermic) and energy is released when they are made (exothermic).
2	If more energy is needed to break the bonds in a reaction then ΔH is positive and if more energy is needed to make the bonds then ΔH is negative.
3	In calculations, a mean bond enthalpy is used which is the average energy needed to break a certain type of bond in a range of compounds.
4	Enthalpy change of reaction ▭▭ Total energy absorbed to break the bonds ▭ Total energy released when bonds are formed

Calorimetry

1	Experiments that measure the heat given out in a reaction are called calorimetry experiments.
2	The equation for enthalpy change is: $q = mc\Delta T$ Where q is the heat lost or gained (J), m is the mass of the water in combustion reactions or solution (g), c is the specific heat capacity of the water (4.18 J $g^{-1}K^{-1}$) or solution and ΔT is the change in temperature of the water or solution (K).
3	To calculate the enthalpy change of a reaction: 1. Calculate q then convert from J to kJ. 2. Calculate the number of moles. If combustion, calculate moles of fuel using n=m/Mr where m is the mass of fuel burnt. If solution, use n=cv. 3. Calculate the enthalpy change in kJ mol^{-1} using $\Delta H=q/n$.
4	A temperature increase shows an exothermic reaction so the enthalpy change must be negative.

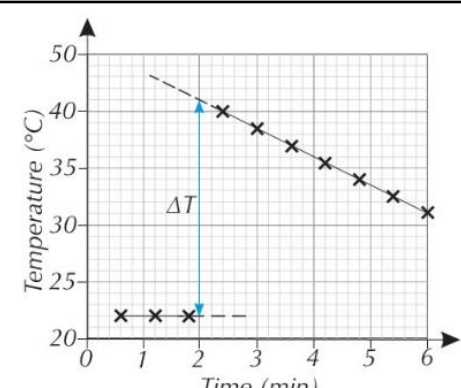## Problems with Calorimetry

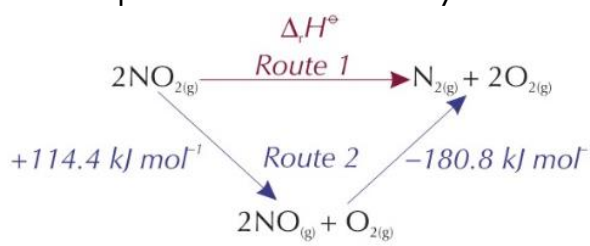| 1	Heat is always lost to the surrounding during the experiment which can make calculated value inaccurate.
2	Combustion may be incomplete which means less energy is given out and calculated value is inaccurate.

Key Vocabulary

1	Standard enthalpy of formation ($\Delta_f H^{\ominus}$)	Enthalpy change when 1 mole of a compound is formed from its elements in their standard states under standard conditions.
2	Standard enthalpy of combustion ($\Delta_c H^{\ominus}$)	Enthalpy change when 1 mole of a substance is completely burned in oxygen under standard conditions with all reactants and products in their standard states.
3	Bond enthalpy	The energy needed to break a bond.
4	Hess's Law	The total enthalpy change for a reaction is independent of the route taken.

Temperature change from graph

1	You can use a graph to find an accurate temperature change for a reaction. You need to extend both lines on the graph so they pass through when the reaction started.
2		The distance between the 2 lines at the point the reaction started is the accurate temperature change (ΔT).

Hess's Law

1	Hess's law is used when you can't measure the enthalpy change directly with an experiment.
2	The example below shows a Hess's cycle:

$$2NO_{2(g)} \xrightarrow{\quad Route\ 1\quad \Delta_rH^\ominus\ } N_{2(g)} + 2O_{2(g)}$$

$+114.4\ kJ\ mol^{-1}$ $Route\ 2$ $-180.8\ kJ\ mol^{-1}$

$$2NO_{(g)} + O_{2(g)}$$

Hess's Law says that the total enthalpy change for route 1 is the same as for route 2.
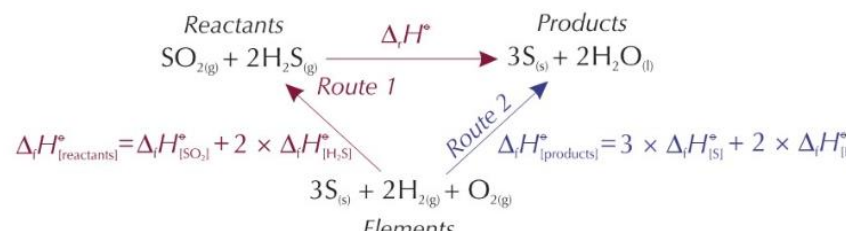
$\Delta_rH^\ominus = 114.4\ kJ\ mol^{-1} + (-180.8\ kJ\ mol^{-1}) = -66.4\ kJ\ mol^{-1}$.

Exam Tips

1	You don't have to pick a route that follows the direction of the arrows. If your route goes against an arrow you can change the signs (so negative enthalpies become positive and positive enthalpies become negative).
2	Use brackets when putting values into your calculator to avoid calculation errors.
3	When using combustion data you can ignore the enthalpy change of combustion for oxygen because it doesn't have one.

Hess's Law Using Formation

1	You need to know $\Delta_fH^\ominus$ for all the reactants and products that are compounds. The $\Delta_fH^\ominus$ for elements is zero.
2	Calculate the $\Delta_rH^\ominus$ using the enthalpies of formation in the table.

$$Reactants \xrightarrow{\quad \Delta_rH^\ominus\quad} Products$$
$$SO_{2(g)} + 2H_2S_{(g)} \xrightarrow{\quad Route\ 1\quad} 3S_{(s)} + 2H_2O_{(l)}$$

$\Delta_fH^\circ_{[reactants]} = \Delta_fH^\circ_{[SO_2]} + 2 \times \Delta_fH^\circ_{[H_2S]}$ $Route\ 2$ $\Delta_fH^\circ_{[products]} = 3 \times \Delta_fH^\circ_{[S]} + 2 \times \Delta_fH^\circ_{[H_2O]}$

$$3S_{(s)} + 2H_{2(g)} + O_{2(g)}$$
$$Elements$$

Elements are at the bottom of the Hess's cycle and the arrows point upwards because the elements form the reactants and products.

3		Compound	Δ_fH°		The table shows the values for the $\Delta_fH^\ominus$ that are needed to calculate the enthalpy change for the reaction.
		$SO_{2(g)}$	$-297\ kJ\ mol^{-1}$		
		$H_2S_{(g)}$	$-20.6\ kJ\ mol^{-1}$		
		$H_2O_{(l)}$	$-286\ kJ\ mol^{-1}$		

4	Use Hess's Law, Route 1 = Route 2, and plug the numbers from the table into the equation:

$\Delta_fH^\ominus[SO_2] + 2\Delta_fH^\ominus[H_2S] + \Delta_rH^\ominus = 3\Delta_fH^\ominus[S] + 2\Delta_fH^\ominus[H_2O]$
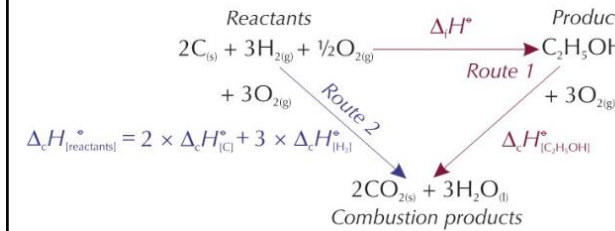$-297 + (2 \times -20.6) + \Delta_rH^\ominus = (3 \times 0) + (2 \times -286)$
$\Delta_rH^\ominus = (3 \times 0) + (2 \times -286) - [-297 + (2 \times -20.6)] = -233.8\ kJ\ mol^{-1}$.

Hess's Law Using Combustion

1	You will be provided with values for $\Delta_cH^\ominus$ for substances required.		Substance	Δ_cH°
			$C_{(s)}$	$-394\ kJ\ mol^{-1}$
			$H_{2(g)}$	$-286\ kJ\ mol^{-1}$
			$C_2H_5OH_{(l)}$	$-1367\ kJ\ mol^{-1}$

2	Calculate $\Delta_fH^\ominus$ of ethanol using the enthalpies of combustion in the table.

$$Reactants \xrightarrow{\quad \Delta_fH^\circ\quad} Product$$
$$2C_{(s)} + 3H_{2(g)} + \tfrac{1}{2}O_{2(g)} \xrightarrow{\quad Route\ 1\quad} C_2H_5OH_{(l)}$$
$+ 3O_{2(g)}$ $Route\ 2$ $+ 3O_{2(g)}$

$\Delta_cH^\circ_{[reactants]} = 2 \times \Delta_cH^\circ_{[C]} + 3 \times \Delta_cH^\circ_{[H_2]}$ $\Delta_cH^\circ_{[C_2H_5OH]}$

$$2CO_{2(s)} + 3H_2O_{(l)}$$
$$Combustion\ products$$

Combustion products are at the bottom of the Hess's cycle and the arrows point down because they are formed in a combustion reaction.

3	Use Hess's Law, Route 1 = Route 2, and plug the numbers from the table into the equation:

$\Delta_fH^\circ[ethanol] + \Delta_cH^\circ[C_2H_5OH] = 2\Delta_cH^\circ[C] + 3\Delta_cH^\circ[H_2]$
$\Delta_fH^\circ[ethanol] + (-1367) = (2 \times -394) + (3 \times -286)$
$\Delta_fH^\circ[ethanol] = -788 + -858 - (-1367)$
$\Delta_fH^\circ[ethanol] = -279\ kJ\ mol^{-1}$.