

MEAN BOND ENTHALPY

This is defined as 'the energy required to break **one mole** of bonds between atoms in molecules in the gas phase'

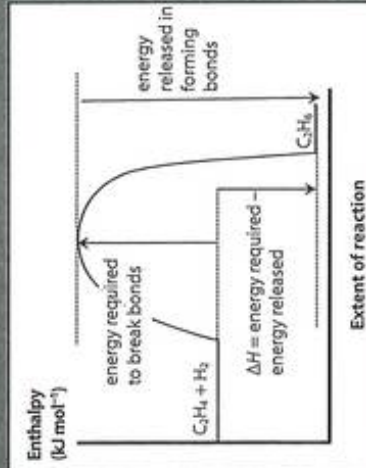
- e.g. $\text{CH}_4(\text{g}) \rightarrow \text{CH}_3(\text{g}) + \text{H}(\text{g}) \quad \Delta H = +435 \text{ kJ mol}^{-1}$
- Breaking bonds is always endothermic ($\Delta H = +ve$) and forming bonds is always exothermic ($\Delta H = -ve$)
- For the **same** bonds, the same amount of energy is taken in to break them as is given out when they form (ΔH is the same number – but different sign)
- The bond between the **same** atoms has different energies in different compounds, e.g. the C–H bond in methane requires more energy to break than the C–H bond in ethane. Since you could not list the energy of every bond in every molecule, a **mean average** is worked out from a representative set of molecules.

Because most bond enthalpy calculations use **mean** (average) values and not values for the specific bonds in the actual compounds in the reaction, enthalpy changes calculated this way are **approximate**. Enthalpy changes calculated using Hess's law will give different values (which will be more **accurate** if the enthalpy data used has been measured accurately).

ENTHALPY CHANGES AND BOND ENERGIES

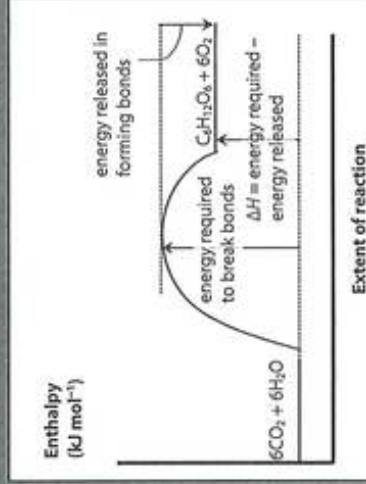
Exothermic reactions

- Less energy is needed to break bonds in the reactants than is given out forming bonds in the products
- Energy given out: $\Delta H = -ve$
e.g. $\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$



Endothermic reactions

- More energy is needed to break bonds in the reactants than is given out forming bonds in the products
- Energy taken in: $\Delta H = +ve$
e.g. $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$



Calculating ΔH using bond energies

- Draw out the structures of all of the reactants and all of the products
- Identify all of the bonds in the reactants and products and look up their energy values
- Add up the energies of all of the bonds in the reactants
- Add up the energies of all of the bonds in the products
- Take the total energy of the bonds in the products from the total energy of the bonds in the reactants

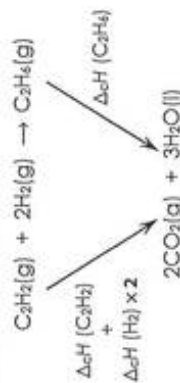
EXAM TIP

In an exam, you will be given the relevant data for any enthalpy calculation

Write out the equation of the reaction for which you want to find the enthalpy change then follow these steps:

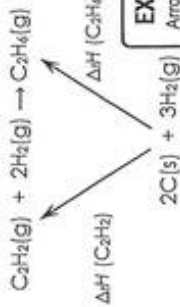
1 CHECK THE DATA YOU ARE GIVEN AND DRAW A CYCLE

For enthalpy changes of **combustion**, draw a cycle using **combustion products** of all reactants and products, e.g. for the hydrogenation of ethyne:



EXAM TIP
Arrows always point down in calculations using enthalpy change of combustion data

For enthalpy changes of **formation**, draw a cycle using **elements** in all reactants and products, e.g. for the hydrogenation of ethyne:

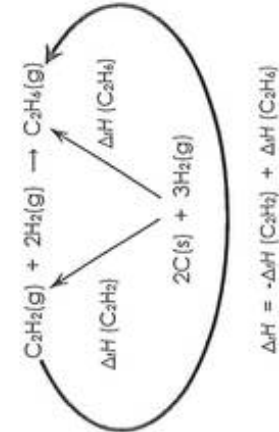
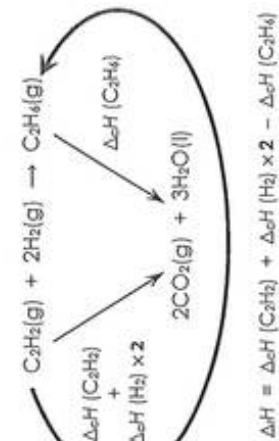


EXAM TIP
 $\Delta_f H^\circ = 0$ for elements in standard states such as H_2

EXAM TIP
Arrows always point up in calculations using enthalpy change of formation data

2 FOLLOW THE CYCLE FROM TOP LEFT TO TOP RIGHT

If you move along an arrow in the direction it points in, then **add** these enthalpies; if you follow it in reverse, then **subtract** these enthalpies



$$\Delta H = \Delta_c H^\circ [\text{C}_2\text{H}_2] + \Delta_c H^\circ [\text{H}_2] \times 2 - \Delta_c H^\circ [\text{C}_2\text{H}_6]$$

3 PUT THE DATA IN AND WORK OUT THE FINAL ANSWER

Remember, some enthalpies are **negative** – subtracting a negative number is the same as adding a positive number

Data

Compound	$\Delta_c H^\circ (\text{kJ mol}^{-1})$	$\Delta_f H^\circ (\text{kJ mol}^{-1})$
C_2H_2	-1301	+224
H_2	-285.8	0
C_2H_6	-1560	-84.6

$$\Delta H = \Delta_c H^\circ [\text{C}_2\text{H}_2] + \Delta_c H^\circ [\text{H}_2] \times 2 - \Delta_c H^\circ [\text{C}_2\text{H}_6]$$

$$= -1301 + (-285.8 \times 2) - (-1560)$$

$$= -1301 - 571.6 + 1560$$

$$= -312.6 \text{ kJ mol}^{-1}$$

Two negatives make a positive

Remember the units