

Bond Enthalpies

How can bond strengths be used to estimate the enthalpy change of a reaction?

●●○ Intermediate

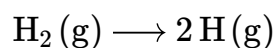
Bonding & Molecular Structure

Last reviewed 5 October 2026

Before you start: it helps to understand [Lewis Structures](#) and [Enthalpy and Calorimetry](#).

What is it?

The **bond enthalpy** (bond energy) is the energy needed to **break one mole** of a particular bond in gaseous molecules, forming gaseous atoms. For hydrogen:



$$\Delta H = +436 \text{ kJ/mol}$$

- **Breaking** a bond always **needs** energy (endothermic, positive).
- **Forming** a bond always **releases** the same amount of energy (exothermic, negative).

Bond	Bond enthalpy (kJ/mol)	Bond	Bond enthalpy (kJ/mol)
H-H	436	C-C	348
Cl-Cl	242	C=C	614
H-Cl	431	C≡C	839
C-H	413	O=O	498
O-H	463	C=O (in CO ₂)	799
N-H	391	N≡N	945

Except for diatomic molecules such as H-H, these are **average** values: a C-H bond in methane is not exactly as strong as one in ethanol, so tables give an average over many compounds.

KEY IDEA

A reaction breaks the old bonds and makes new ones. The enthalpy change is the energy **taken in** to break bonds minus the energy **given out** when new bonds form:

$$\Delta H \approx \sum(\text{bonds broken}) - \sum(\text{bonds formed})$$

If the new bonds are stronger than the old ones, more energy is released than taken in, and the reaction is **exothermic**.

Why does it matter?

- **Explaining energy changes.** Fuels release energy because the bonds in CO_2 and H_2O are stronger than those in the fuel and oxygen.
- **Quick estimates.** You can estimate ΔH for a reaction even when no enthalpies of formation are available.
- **Bond strength and reactivity.** The very strong $\text{N}\equiv\text{N}$ bond (945 kJ/mol) explains why nitrogen gas is so unreactive, and why making ammonia needs a catalyst.

How does it work?

1. Bond order, length and strength

The **bond order** is the number of shared pairs: 1 for single, 2 for double, 3 for triple. As bond order rises, the bond gets **shorter** and **stronger**:

Bond	Order	Length (pm)	Enthalpy (kJ/mol)
C–C	1	154	348
C=C	2	134	614
C≡C	3	120	839

A double bond is not twice as strong as a single bond, because the π bond is weaker than the σ bond.

2. Estimating ΔH : the steps

1. Draw the structures and **list every bond** in the reactants and the products, using the coefficients.
2. Add up the bond enthalpies of the bonds **broken** (reactants).
3. Add up the bond enthalpies of the bonds **formed** (products).
4. $\Delta H \approx \text{broken} - \text{formed}$.

3. Why it is only an estimate

The answer is approximate because the bond enthalpies are **averages**, and because they apply to **gases**: if a reactant or product is a liquid or solid, the energy of changing state is left out. Enthalpies of formation (Hess's law) give more accurate values.

∞ THINK OF IT LIKE THIS

Think of bonds as money. Breaking bonds is spending: you pay energy to pull atoms apart. Making bonds is earning: you are paid energy back. If you earn more from the new bonds than you spent on the old ones, you finish with a profit: the reaction gives out energy (exothermic).

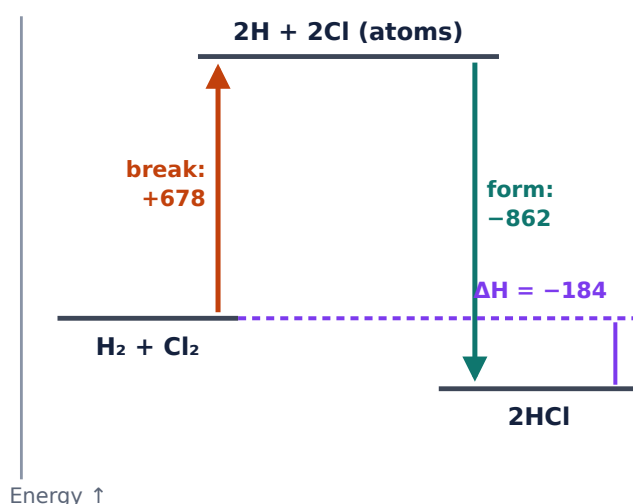
Q MORE PRECISELY

Bond enthalpies are measured for the gas phase, so the estimate works best when every species is a gas. For a diatomic molecule the bond enthalpy is an exact value (the **bond dissociation enthalpy**); for bonds in larger molecules it depends on the rest of the molecule. The value for C=O in CO₂ (799 kJ/mol) differs from the average C=O in aldehydes and ketones (about 745 kJ/mol), which is why the table says "in CO₂". Some tables give 495 kJ/mol for O=O or 941 kJ/mol for N≡N; small differences like these change the estimate by a few kJ/mol.

Visualise it

Bonds broken, bonds formed

$\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ (energies in kJ/mol)



Breaking bonds takes energy in; forming bonds gives more out. The difference is ΔH .

Worked example

WORKED EXAMPLE: Hydrogen and chlorine

Question: Estimate ΔH for $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \longrightarrow 2 \text{HCl}(\text{g})$.

Step 1 Bonds broken: one H-H and one Cl-Cl:

$$\begin{aligned} &436 \text{ kJ/mol} \\ &+ 242 \text{ kJ/mol} \\ &= 678 \text{ kJ/mol} \end{aligned}$$

Step 2 Bonds formed: two H-Cl:

$$\begin{aligned} &2 \times 431 \text{ kJ/mol} \\ &= 862 \text{ kJ/mol} \end{aligned}$$

Step 3 Enthalpy change:

$$\begin{aligned} \Delta H &\approx 678 \text{ kJ/mol} \\ &- 862 \text{ kJ/mol} \\ &= -184 \text{ kJ/mol} \end{aligned}$$

Exothermic: the H-Cl bonds are stronger in total than the bonds broken.

 WORKED EXAMPLE: Burning methane

Question: Estimate ΔH for $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$.

Step 1 Broken: 4 C-H and 2 O=O:

$$\begin{aligned} &4 \times 413 \text{ kJ/mol} \\ &+ 2 \times 498 \text{ kJ/mol} \\ &= 2648 \text{ kJ/mol} \end{aligned}$$

Step 2 Formed: 2 C=O (in CO_2) and 4 O-H (in 2 H_2O):

$$\begin{aligned} &2 \times 799 \text{ kJ/mol} \\ &+ 4 \times 463 \text{ kJ/mol} \\ &= 3450 \text{ kJ/mol} \end{aligned}$$

Step 3 Enthalpy change:

$$\begin{aligned} \Delta H &\approx 2648 \text{ kJ/mol} \\ &- 3450 \text{ kJ/mol} \\ &= -802 \text{ kJ/mol} \end{aligned}$$

This agrees closely with the measured value for gaseous water.

 WORKED EXAMPLE: Making ammonia

Question: Estimate ΔH for $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \longrightarrow 2 \text{NH}_3(\text{g})$ and compare it with the measured value of -92 kJ/mol .

Step 1 Broken: 1 N≡N and 3 H-H: $945 \text{ kJ/mol} + 3 \times 436 \text{ kJ/mol} = 2253 \text{ kJ/mol}$

Step 2 Formed: 2 NH_3 contain 6 N-H: $6 \times 391 \text{ kJ/mol} = 2346 \text{ kJ/mol}$

Step 3 Enthalpy change:

$$\begin{aligned} \Delta H &\approx 2253 \text{ kJ/mol} \\ &- 2346 \text{ kJ/mol} \\ &= -93 \text{ kJ/mol} \end{aligned}$$

Step 4 Within about 1 kJ/mol of the measured -92 kJ/mol . The difference comes from using an average N-H value.

Common mistake

COMMON MISTAKE: Subtracting the wrong way round

It is **broken – formed** (reactants – products). This is the opposite order to the enthalpy-of-formation formula (products – reactants), because bond enthalpies are energies **taken in**. Check: forming strong bonds should make ΔH negative.

COMMON MISTAKE: Forgetting the coefficients or missing bonds

In $\text{CH}_4 + 2 \text{O}_2$ there are **two** $\text{O}=\text{O}$ bonds to break, and two water molecules contain **four** $\text{O}-\text{H}$ bonds. Draw every molecule and count every bond.

COMMON MISTAKE: Using a single-bond value for a double bond

$\text{C}=\text{O}$, $\text{C}=\text{C}$ and $\text{N}\equiv\text{N}$ have their own values. Using $\text{C}-\text{O}$ (358 kJ/mol) for the $\text{C}=\text{O}$ in CO_2 gives a badly wrong answer.

NOTATION NOTE

- Bond enthalpy is sometimes written D or E , e.g. $D(\text{H}-\text{H}) = 436 \text{ kJ/mol}$.
- 1 pm (picometre) = 10^{-12} m .

Remember this

REMEMBER THIS

- Bond enthalpy: energy to break 1 mol of a bond in the gas phase (always positive).
- Breaking bonds is endothermic; forming bonds is exothermic.
- $\Delta H \approx \sum(\text{broken}) - \sum(\text{formed})$.
- Higher bond order: shorter and stronger ($\text{C}-\text{C}$ 348, $\text{C}=\text{C}$ 614, $\text{C}\equiv\text{C}$ 839 kJ/mol).
- Results are estimates: values are averages and apply to gases.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

Practise this topic with flashcards and a quiz at chemistryclarity.com/chemistry/bond-enthalpies/