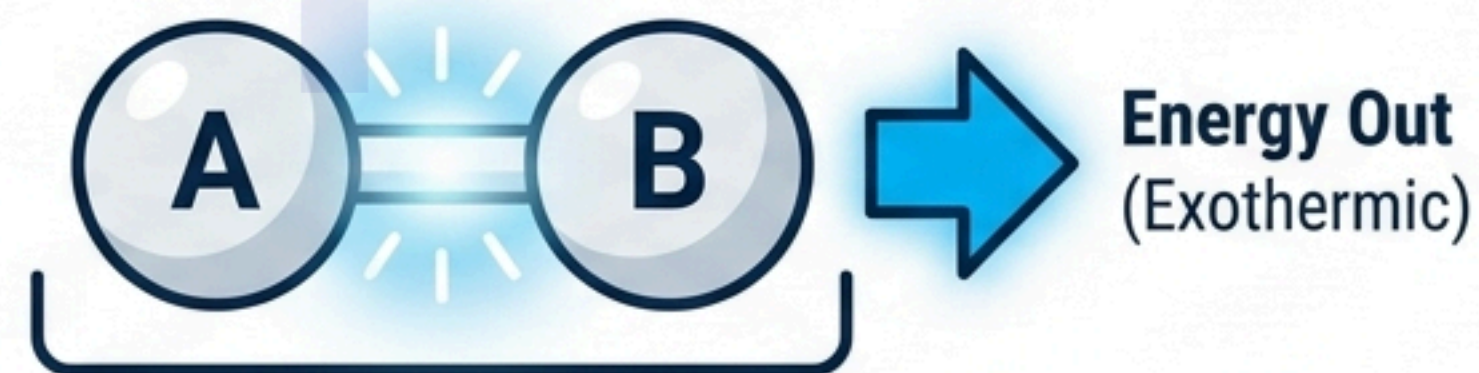


The Molecular Ledger

Bond Enthalpy | Bonds Broken – Bonds Formed

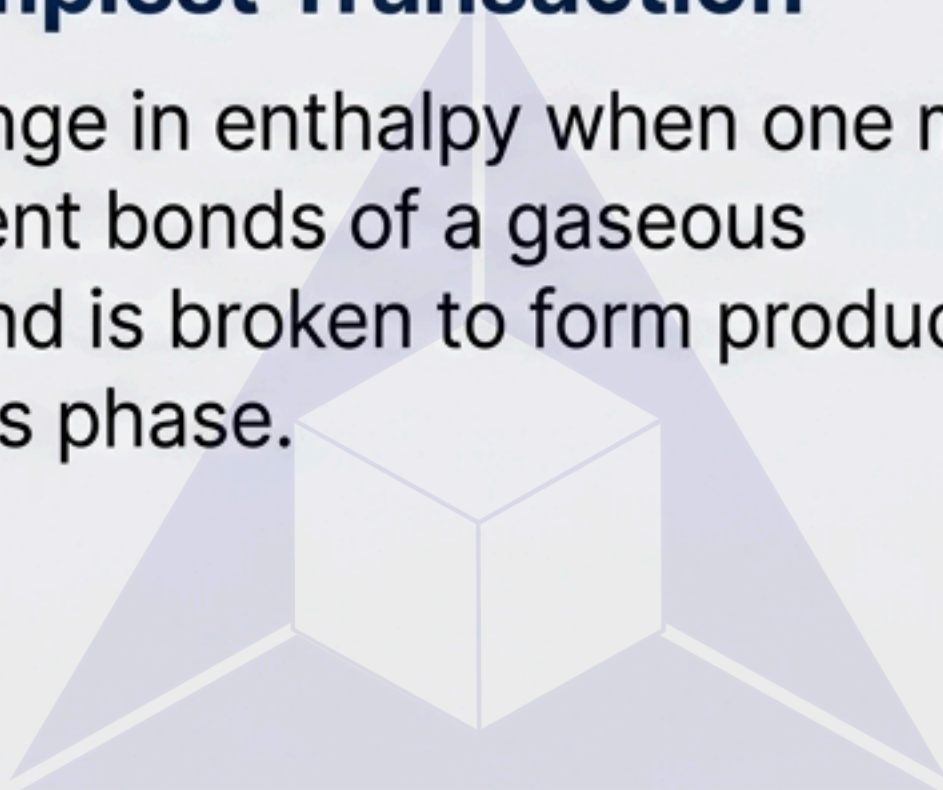
BREAKING COSTS ENERGY. MAKING RELEASES ENERGY.



Bond Dissociation Enthalpy

The Simplest Transaction

The change in enthalpy when one mole of covalent bonds of a gaseous compound is broken to form products in the gas phase.



For diatomic molecules, the Enthalpy of Atomization ($\Delta_a H^\circ$) is exactly equal to the Bond Dissociation Enthalpy.



$$\Delta_{\text{H-H}} H^\circ = +435.0 \text{ kJ mol}^{-1}$$

↗ Endothermic Process (Energy Investment Required)

Polyatomic Molecules: The Need for Mean Values

In polyatomic molecules like Methane (CH_4), all four C–H bonds are structurally identical. Yet, the energy required to break them successively differs as the molecular environment changes.



Total Enthalpy of Atomization ($\Delta_a H^\circ$): 1665 kJ/mol

Mean C–H Bond Enthalpy = $1665 / 4 = 416 \text{ kJ mol}^{-1}$

Insight: For polyatomic calculations, we always use these average values.



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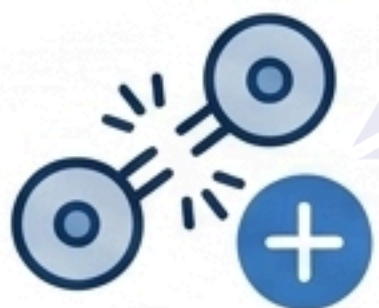
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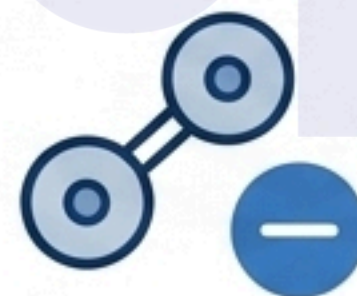
Calculating Reaction Enthalpy ($\Delta_r H^\circ$)

$$\Delta_r H^\circ \approx \sum E_{\text{bonds broken (Reactants)}} - \sum E_{\text{bonds formed (Products)}}$$



The Cost (Reactants)

- Bonds Broken
- Endothermic Process
- Positive (+) Energy Value



The Payoff (Products)

- Bonds Formed
- Exothermic Process
- Negative (–) Energy Value

High-Yield JEE Traps



The Sign Trap

Misconception: $\Delta H = \text{Products} - \text{Reactants}$.

Reality: Bond Enthalpy is the ONLY major thermodynamic formula that is $\text{Reactants} - \text{Products}$ (Broken minus Formed).

Do not reverse them!



The State Trap

Bond enthalpy data strictly applies to gaseous species.

If a species is liquid or solid (e.g., $\text{H}_2\text{O}(\text{l})$ or $\text{C}(\text{s})$), you must account for the enthalpy of vaporization/sublimation in your calculation first.



The Approximation Trap

Because we use mean bond enthalpies for polyatomic molecules, calculating $\Delta_r H^\circ$ this way yields an approximation.

It is less exact than using Standard Enthalpies of Formation ($\Delta_f H^\circ$).

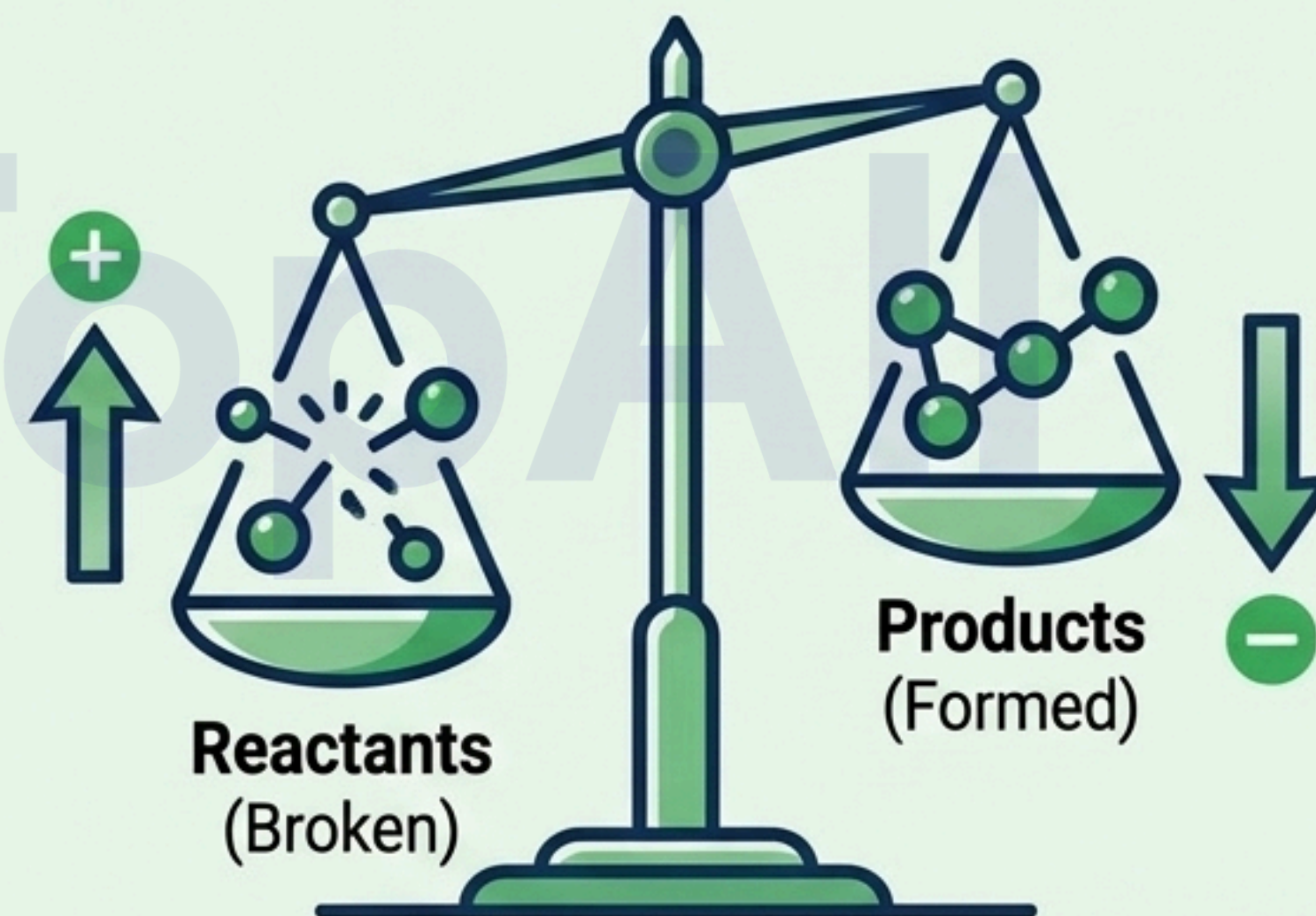
Rules & Formula

- **Rule 1:** Breaking bonds = Endothermic (+).
Forming bonds = Exothermic (-).
- **Rule 2:** Diatomic molecules have exact bond enthalpies. Polyatomic molecules rely on Mean Bond Enthalpies.

The Golden Formula:

$$\Delta_r H^\circ \approx \sum \text{B.E.}(\text{Reactants}) - \sum \text{B.E.}(\text{Products})$$

Visual Summary





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