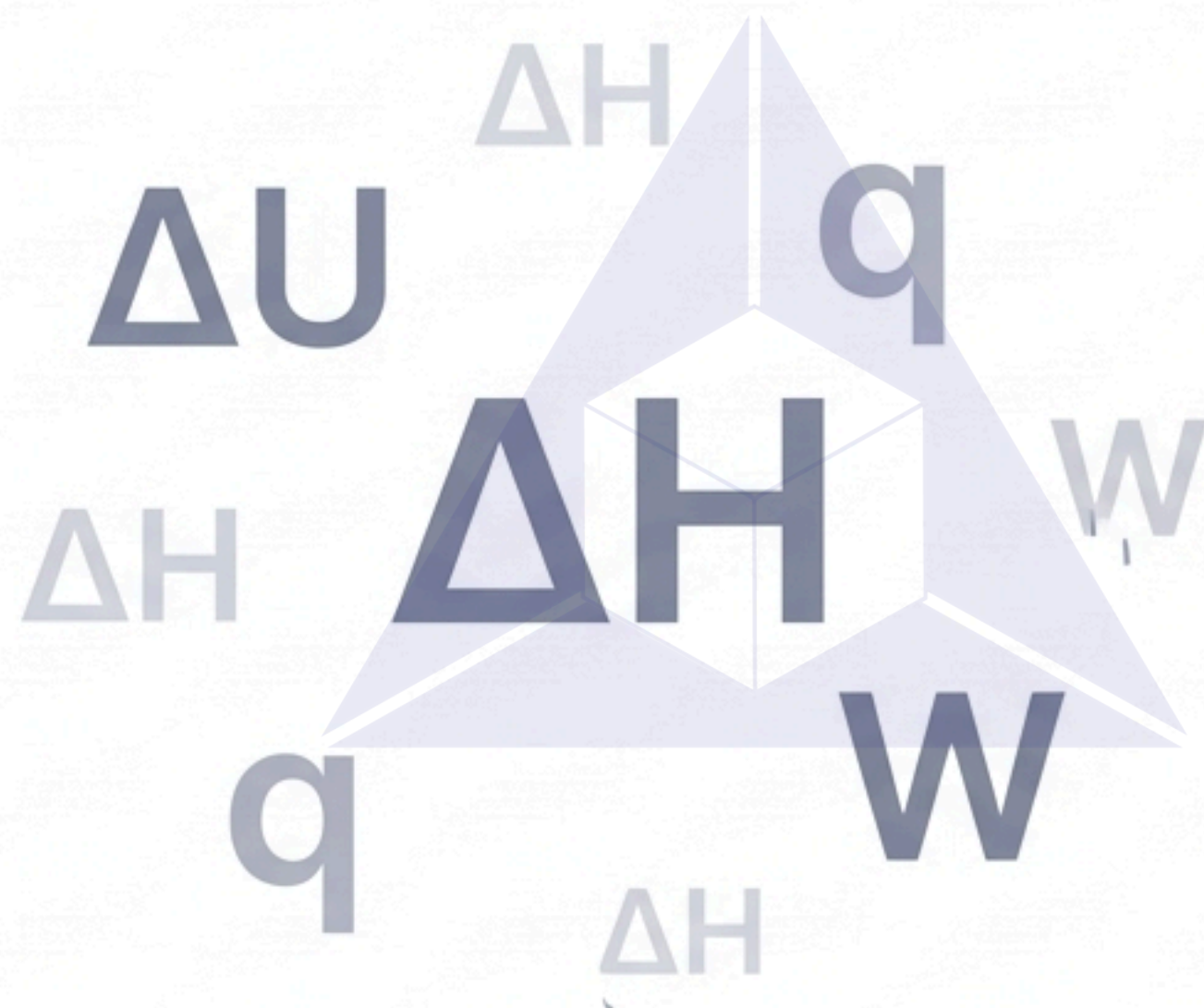
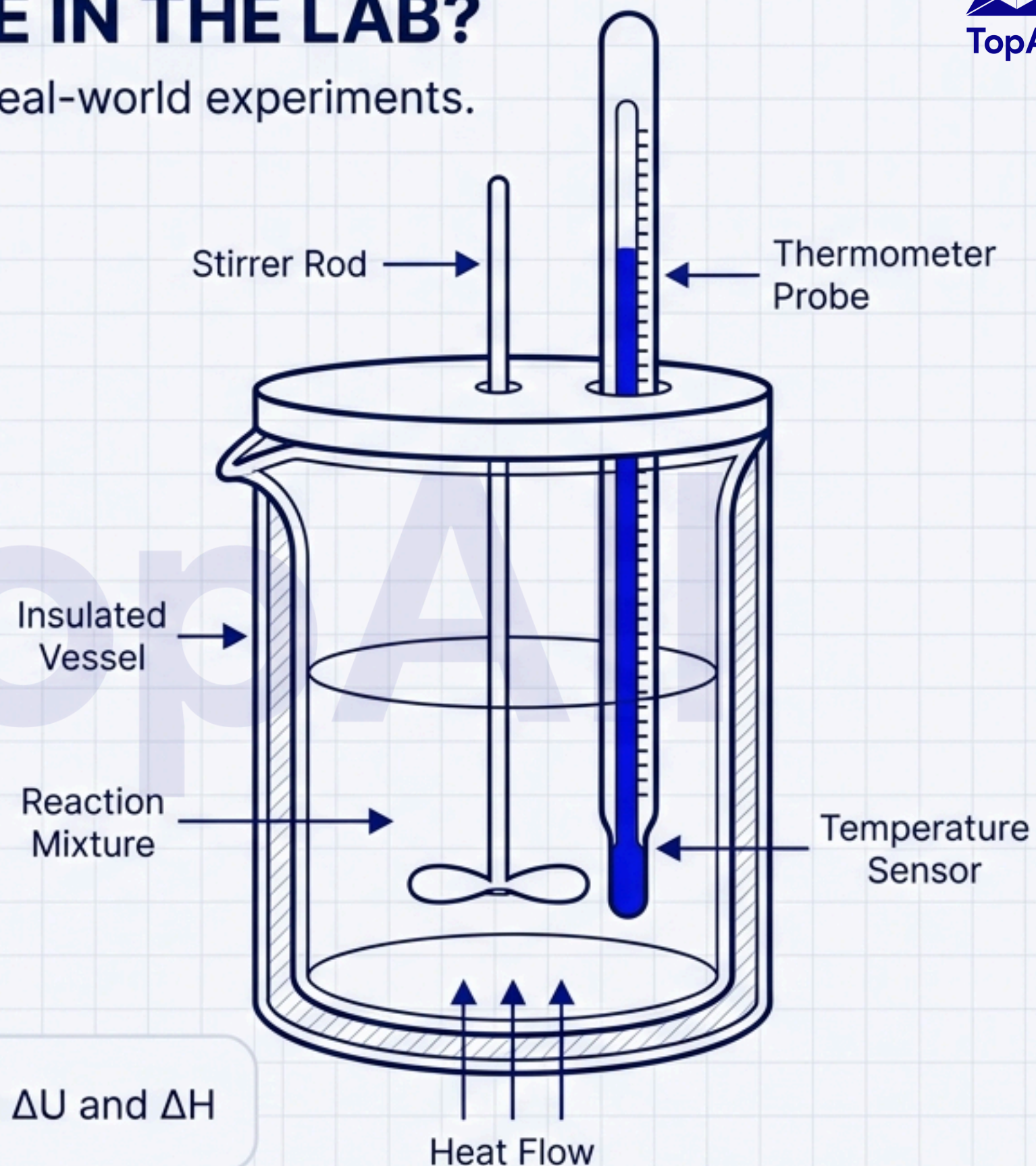


CAN WE MEASURE ENERGY CHANGE IN THE LAB?

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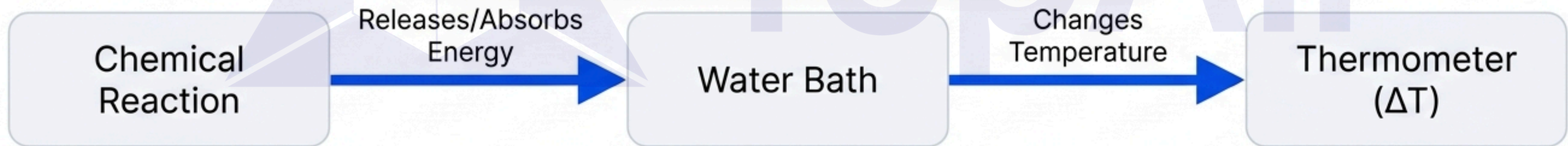


Calorimetry | Measuring ΔU and ΔH



The Calorimetry Proxy: Temperature as a Tool

$$q = C \Delta T$$



Calorimetry relies on a simple heat balance: the heat lost by the reaction is equal to the heat gained by the calorimeter and its surroundings.

$$q_{\text{system}} = -q_{\text{calorimeter}}$$



You've Learned the Theory. Let's Test Them Now.

$$E=mc^2$$



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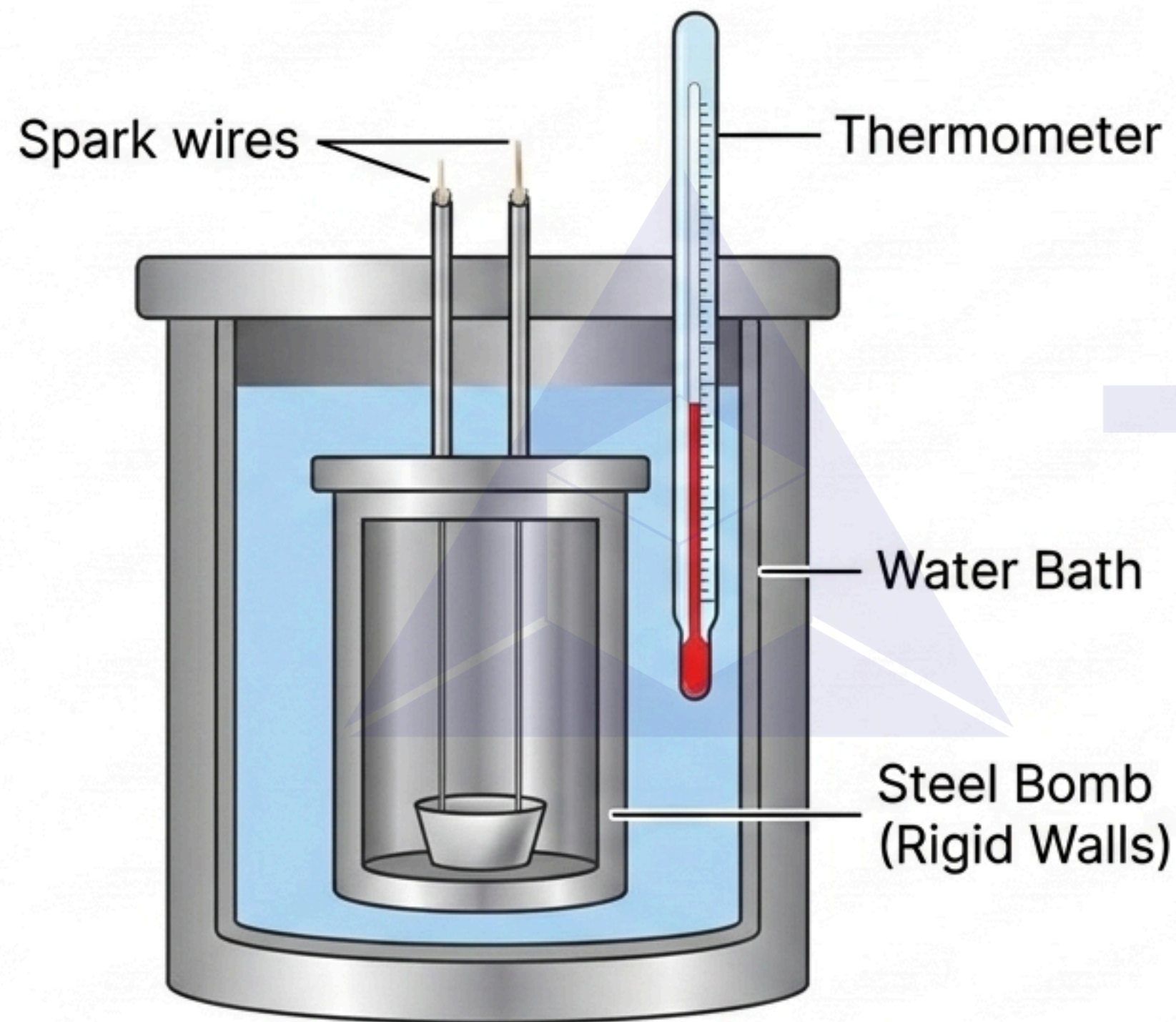
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The Bomb Calorimeter: Measuring ΔU



Used primarily for combustion reactions of combustible substances in pure dioxygen.

Constant Volume

$$\Delta V = 0$$

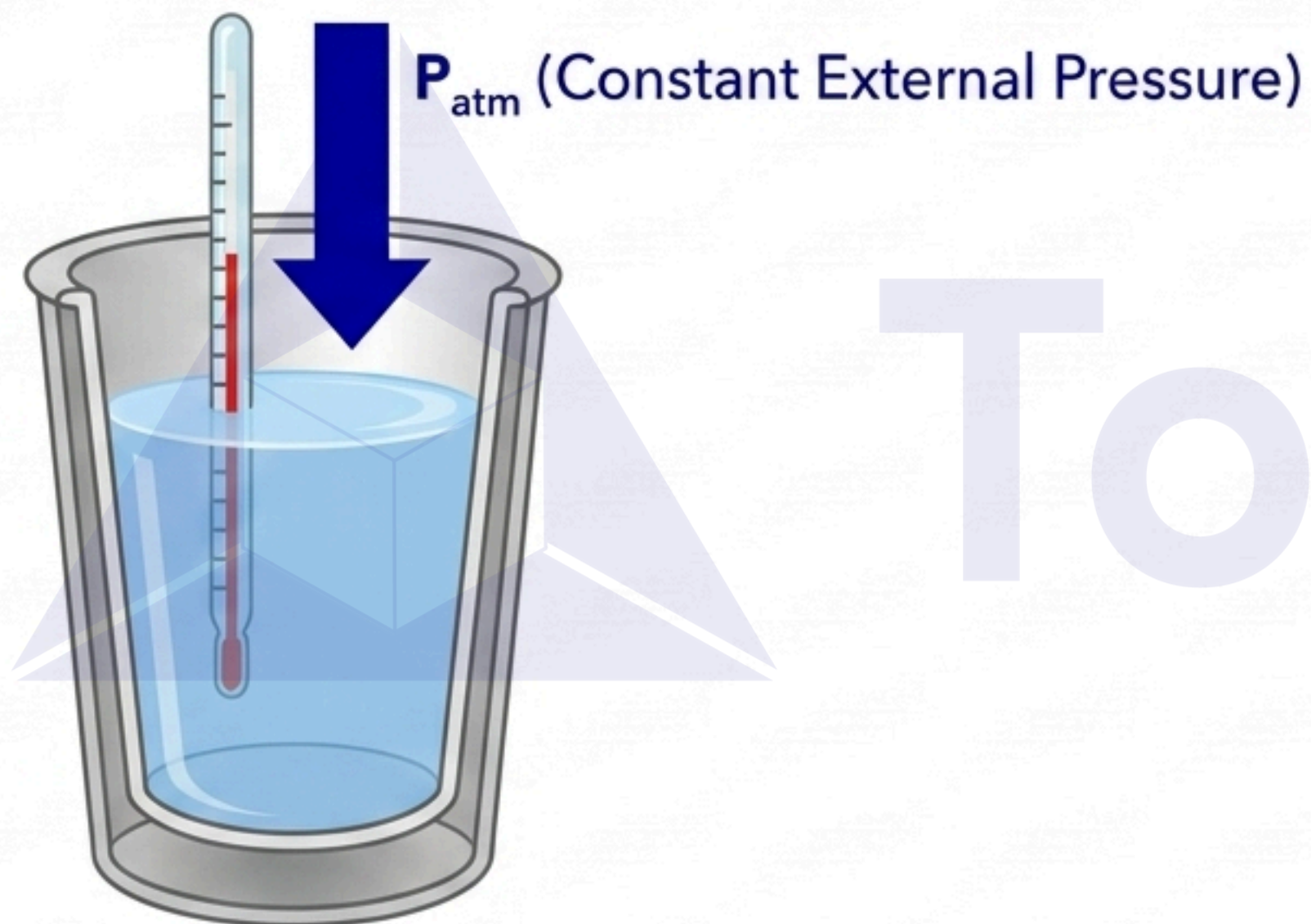
No Work Done

$$W = -p_{\text{ex}}\Delta V = 0$$

The Result (First Law: $\Delta U = q + W$)

$$q_V = \Delta U$$

Constant-Pressure Calorimeter: Measuring ΔH



Commonly used for reactions in solution where volume changes are negligible.

Constant Pressure

$$P = \text{atm}$$

Expansion Work

System can expand/contract against atmosphere, meaning $p\Delta V$ work is done

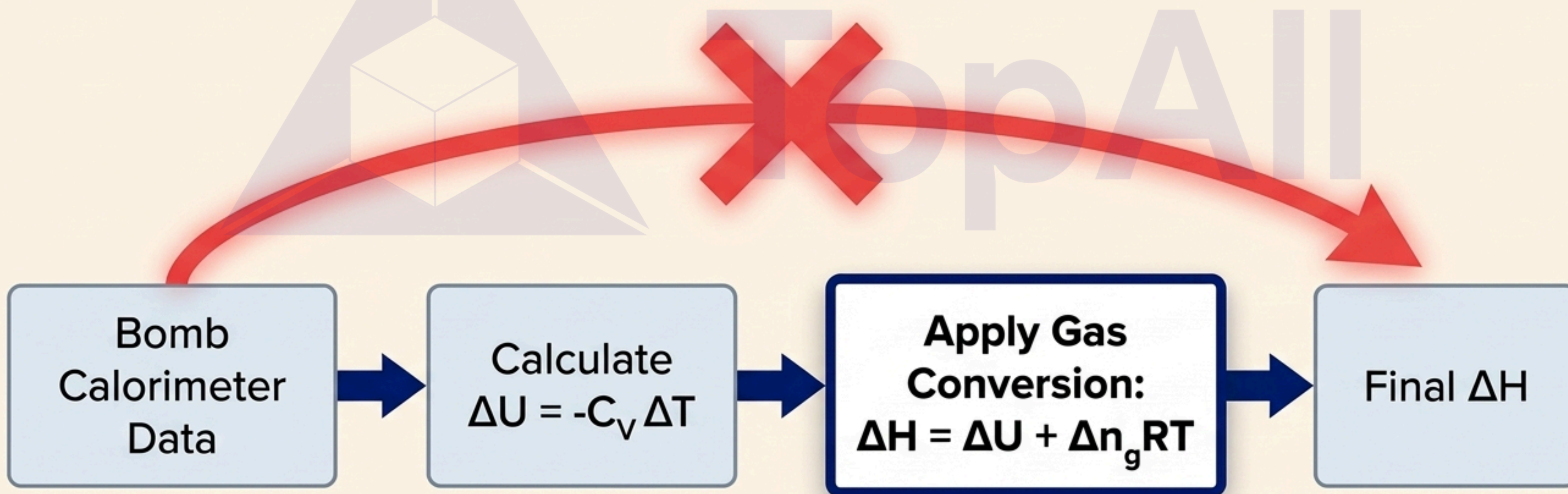
The Result

$$q_p = \Delta H$$

The Classic Exam Trap



Examiners will give you Bomb Calorimeter data and ask for the Enthalpy of reaction (ΔH). Bomb data gives you ΔU , **NOT** ΔH !



Synthesis: Calorimetry at a Glance

Feature	Bomb Calorimeter	Coffee-Cup Calorimeter
Condition	$\Delta V = 0$ (Constant Volume)	$\Delta P = 0$ (Constant Pressure)
Work ($p\Delta V$)	Zero ($W = 0$)	Non-zero ($W \neq 0$)
State Function Measured	Internal Energy (ΔU)	Enthalpy (ΔH)



- Heat is a path function, but under specific constraints (constant V or P), it directly measures state functions (ΔU or ΔH).
- Always balance the heat: $q_{\text{reaction}} = -(C \times \Delta T_{\text{calorimeter}})$.



You've Learned the Theory.

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