

Calorimetry

Calorimetry

the science of measuring heat

**based on observing the temperature change
when a body absorbs or discharges energy as
heat.**

$$q = nC \Delta T$$

moles

grams

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Molar heat capacity

Specific heat capacity

Heat Capacity (C)

the amount of heat needed to raise the temperature of a certain amount of material 1K or 1°C

Molar heat capacity $\frac{\text{J}}{\text{K mol}}$ or $\frac{\text{J}}{^{\circ}\text{C mol}}$

Specific heat capacity $\frac{\text{J}}{\text{K g}}$ or $\frac{\text{J}}{^{\circ}\text{C g}}$

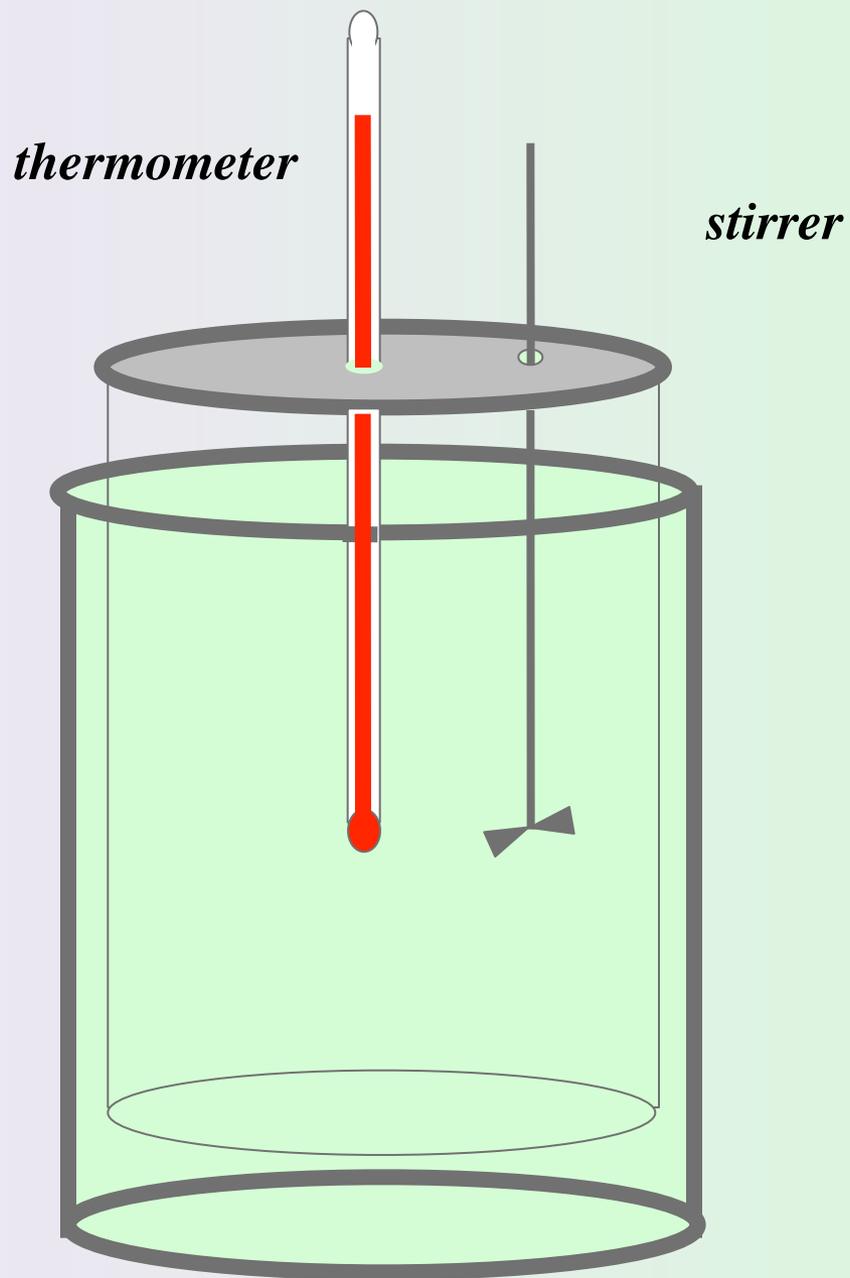
Constant Pressure Calorimeter

takes place at constant pressure

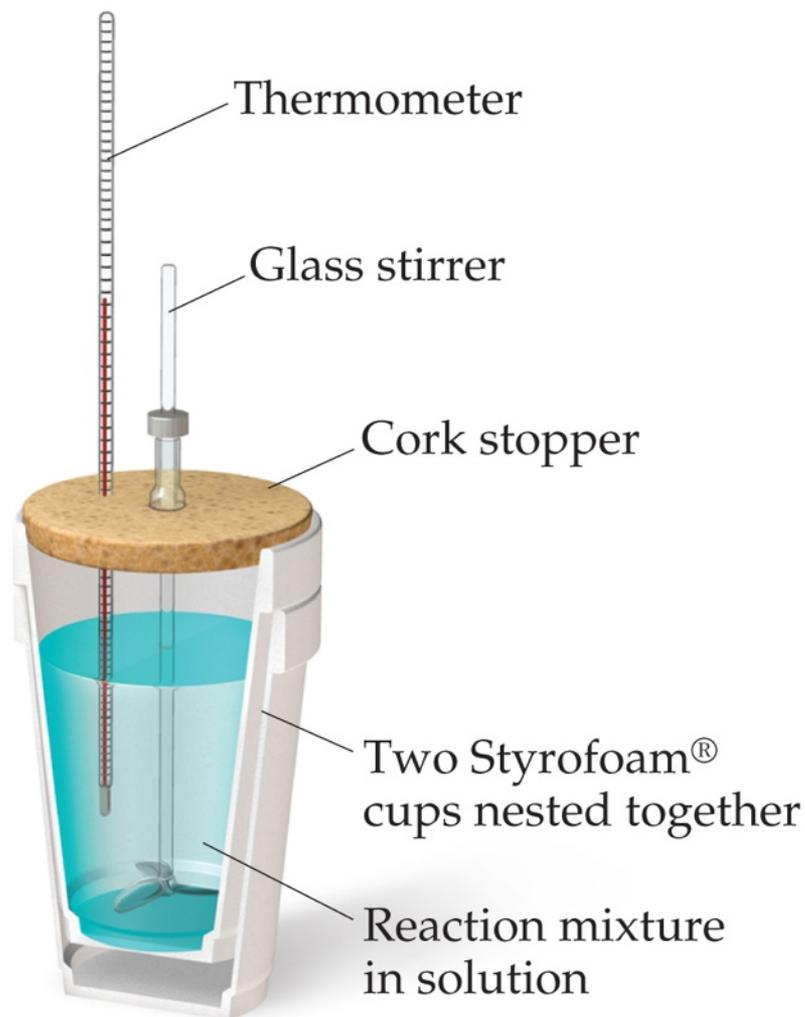
(atmospheric pressure)

**used to determine the changes in enthalpy (ΔH)
occurring in a solution**

Coffee Cup calorimeter



Calorimetry



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Since we cannot know the exact enthalpy of the reactants and products, we measure ΔH through **calorimetry**, the measurement of heat flow.

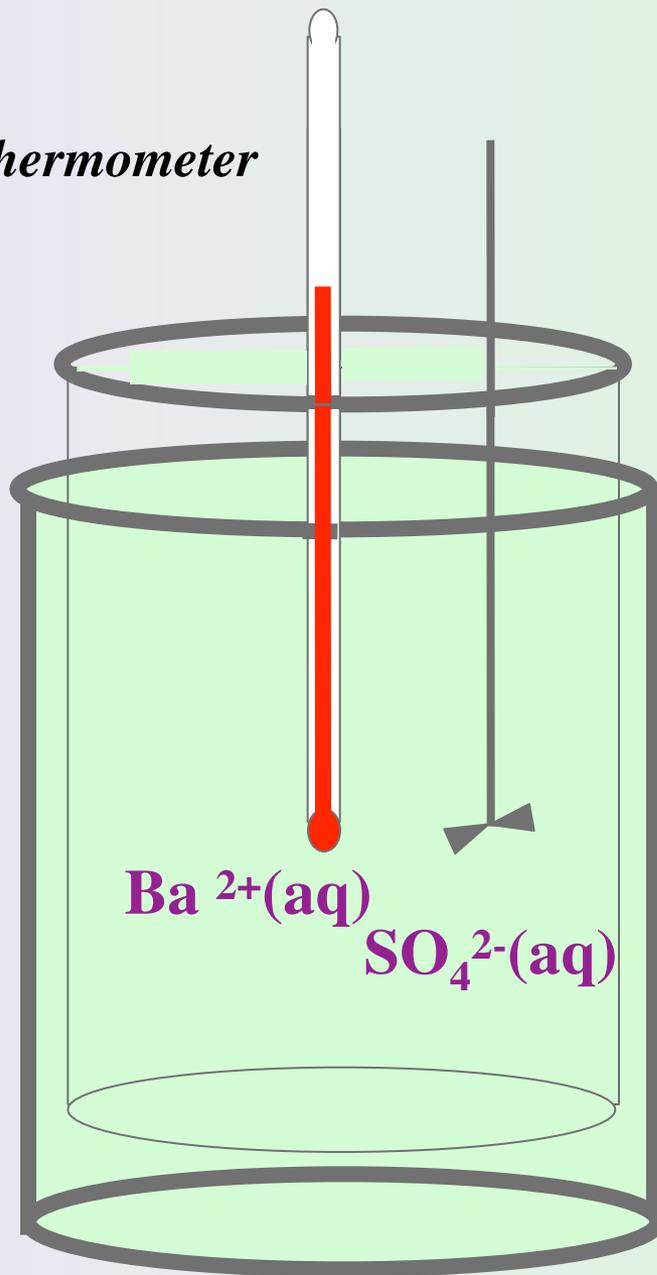
Example:

When 1.0L of 1.0 M $\text{Ba}(\text{NO}_3)_2$ at 25 °C is mixed with 1.0 L of 1.0 M Na_2SO_4 at 25 °C in a calorimeter, the white solid BaSO_4 forms and the temperature increases to 28.1 °C. The specific heat capacity of the solution is $4.18\text{J } ^\circ\text{C}^{-1} \text{g}^{-1}$, and the density of the final solution is 1.0g/ml, calculate the enthalpy (ΔH) change per mole of BaSO_4 .



thermometer

stirrer

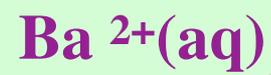
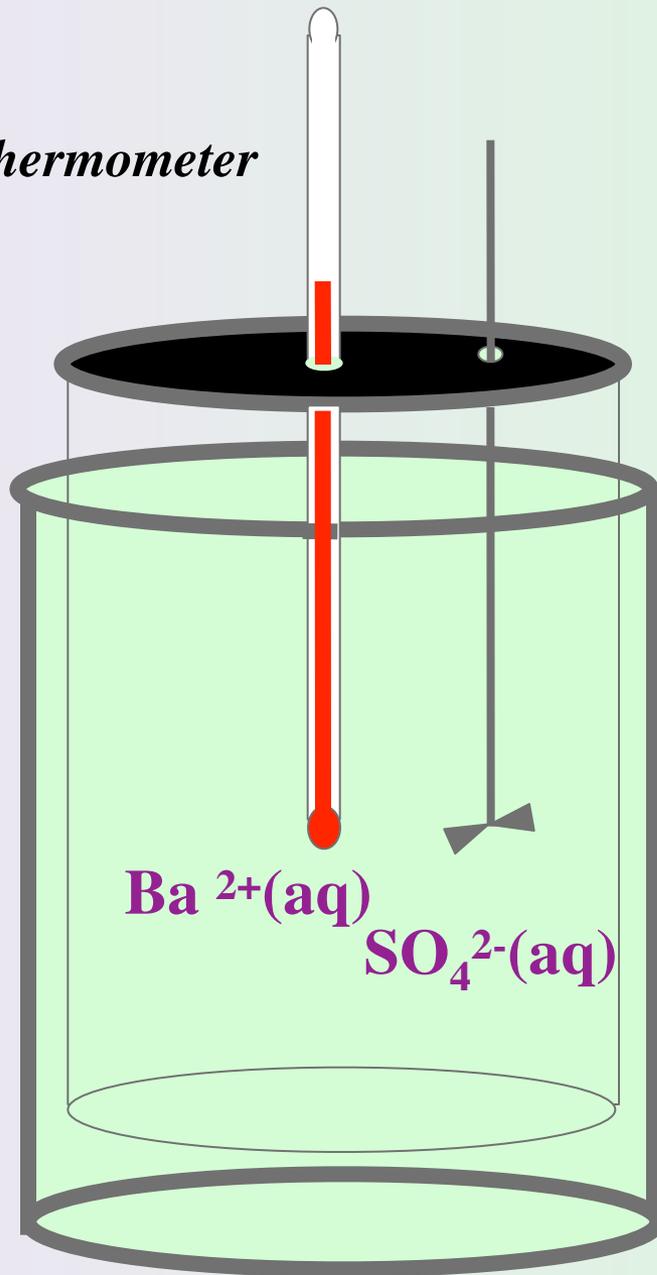


$\text{Ba}^{2+}(\text{aq})$

$\text{SO}_4^{2-}(\text{aq})$

thermometer

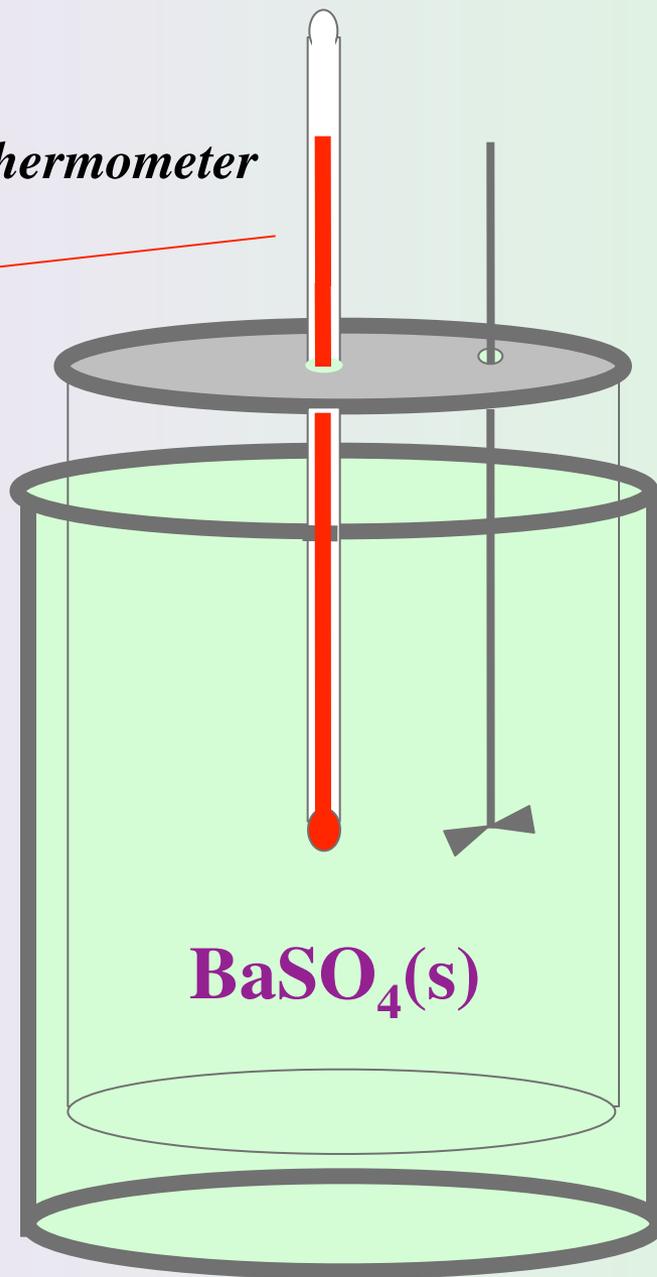
stirrer



ΔT

thermometer

stirrer



$\text{BaSO}_4(\text{s})$

Example:



**Energy released
by the reaction = Energy absorbed
by the solution**

**= Mass of the
solution x Specific heat
capacity x Increase in
temperature**

$$q = nC \Delta T$$

Example:



Volume solution

Density solution

Mass of the
solution

$$2.0\text{L} \times \frac{1000 \text{ ml}}{1\text{L}} \times \frac{1.0 \text{ g}}{1\text{ml}} = 2.0 \times 10^3\text{g}$$

Increase in
temperature

$$T_{\text{final}} - T_{\text{int}} \\ 28.1^\circ \text{C} - 25^\circ \text{C} = 3.1^\circ \text{C}$$

Example:



$$q = nC \Delta T$$

$$= (2.0 \times 10^3 \text{g}) (4.18 \text{J } ^\circ \text{C}^{-1} \text{g}^{-1}) (3.1 \text{ } ^\circ \text{C})$$

$$= 2.6 \times 10^4 \text{J}$$

Heat absorbed by the solution

$$= q_p = \Delta H = -2.6 \times 10^4 \text{J} \times \frac{1 \text{kJ}}{1000 \text{J}} = -26 \text{ kJ}$$

Heat released by the system

$$1 \text{L Ba}^+ \times \frac{1 \text{mol Ba}^+}{1 \text{L}} \times \frac{1 \text{mol BaSO}_4}{1 \text{mol Ba}^+} = 1 \text{mol BaSO}_4$$

$$\frac{-26 \text{ kJ BaSO}_4}{\text{mol}}$$

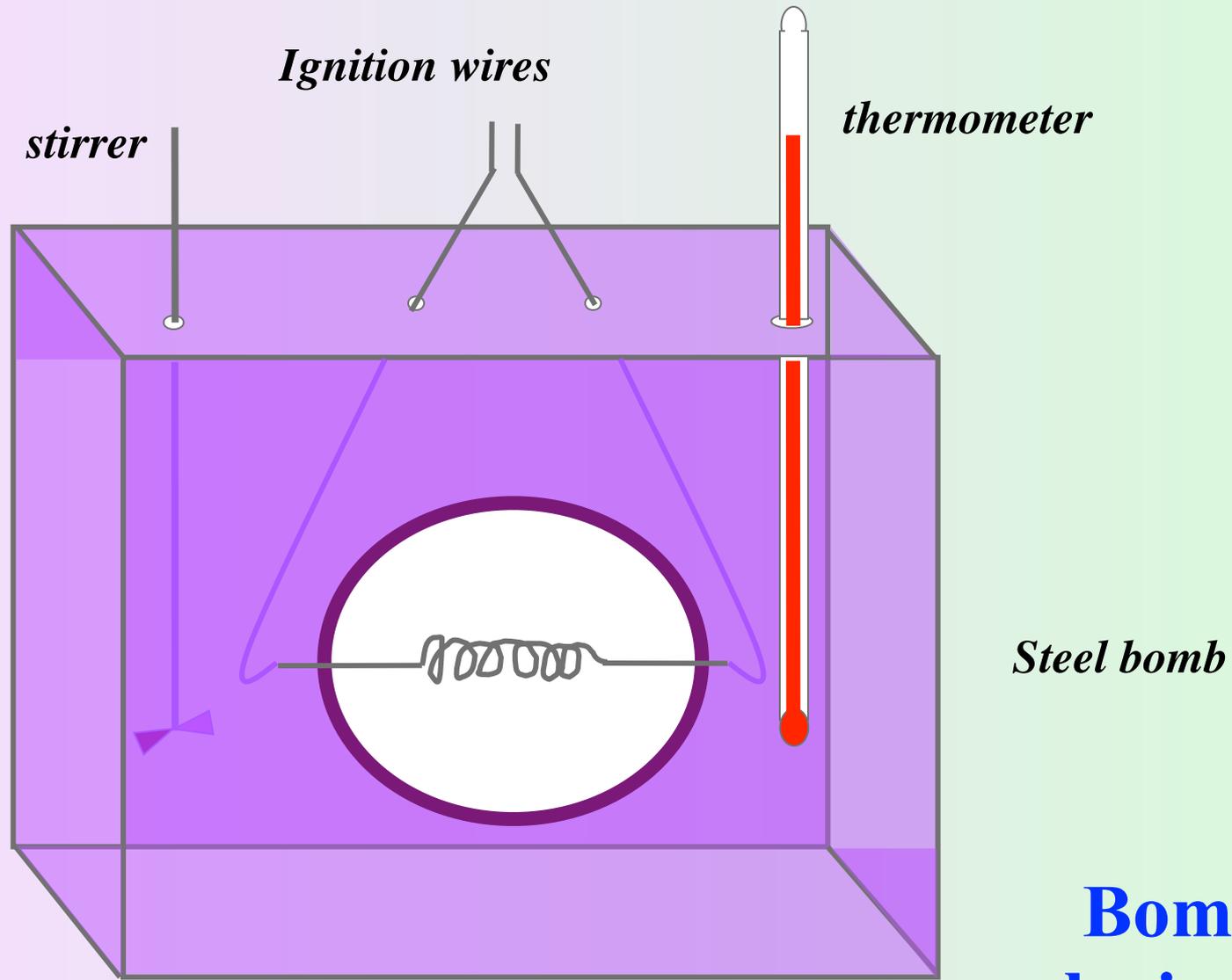
Constant Volume Calorimeter

takes place at constant volume

A bomb calorimeter

Weighed reactants are placed inside a rigid steel container and ignited

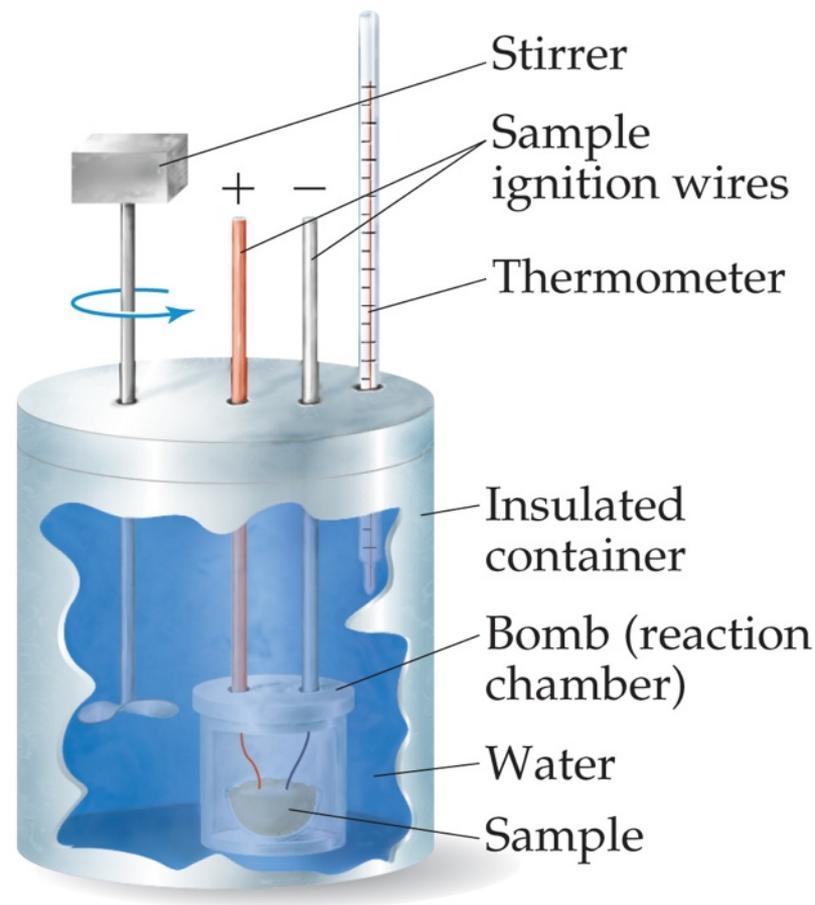
The energy change is determined by measuring the increase in temperature of the calorimeter



**Bomb
calorimeter**

Bomb Calorimetry

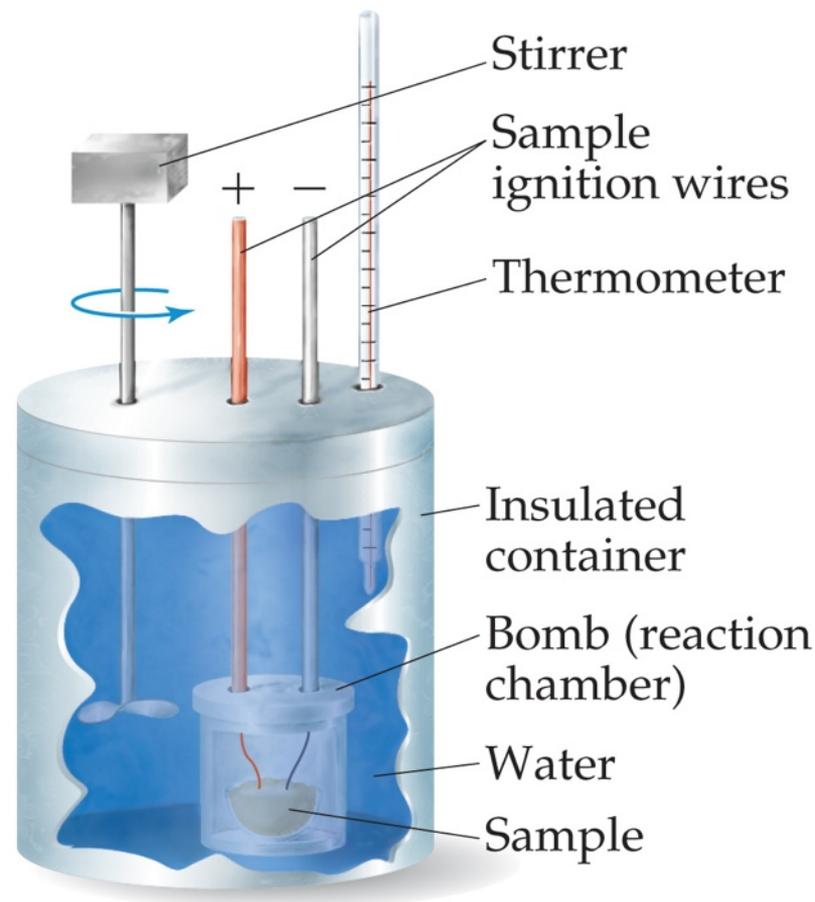
- Reactions can be carried out in a sealed “bomb” such as this one.
- The heat absorbed (or released) by the water is a very good approximation of the enthalpy change for the reaction.



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Bomb Calorimetry

- Because the volume in the bomb calorimeter is constant, what is measured is really the change in internal energy, ΔE , not ΔH .
- For most reactions, the difference is very small.

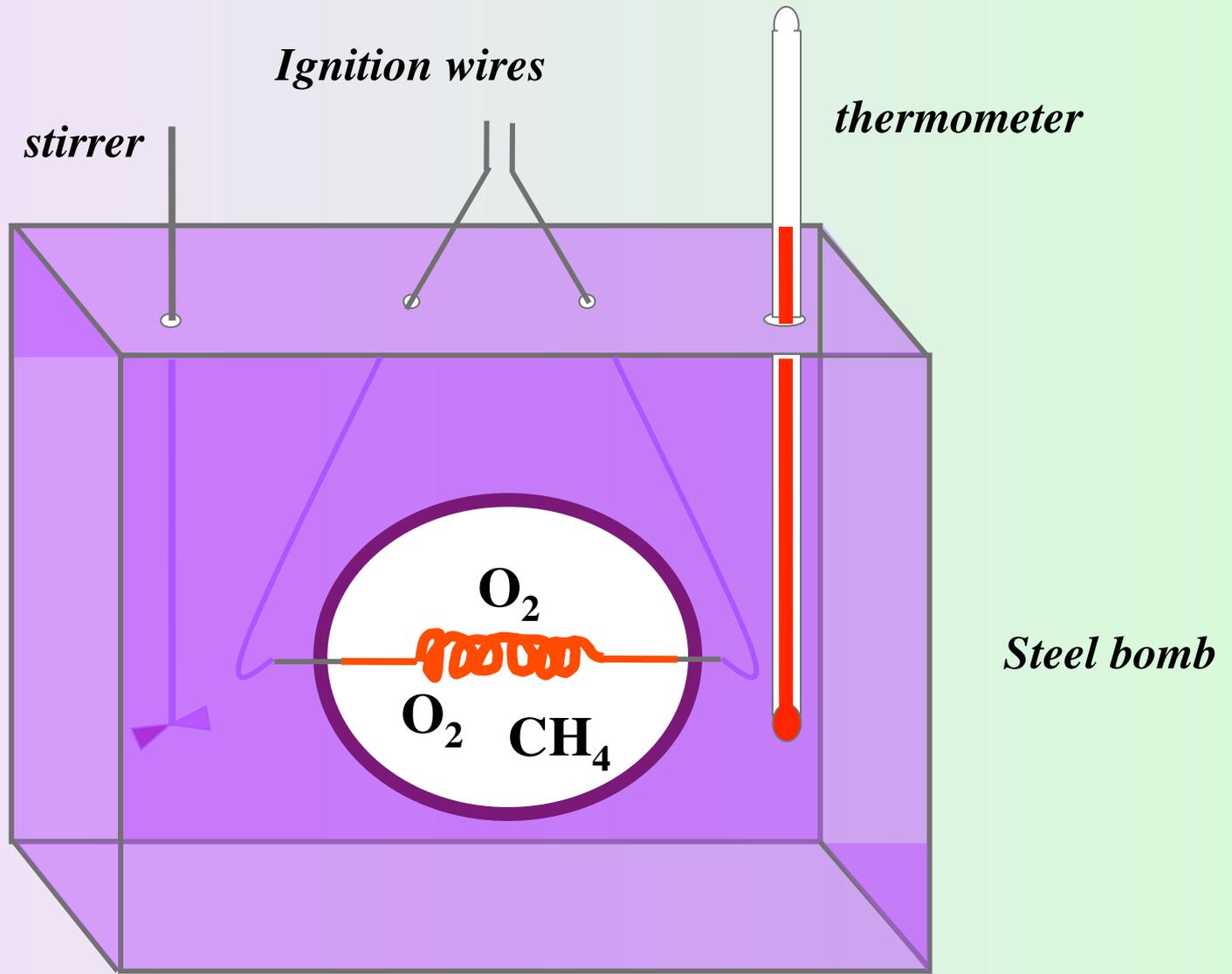


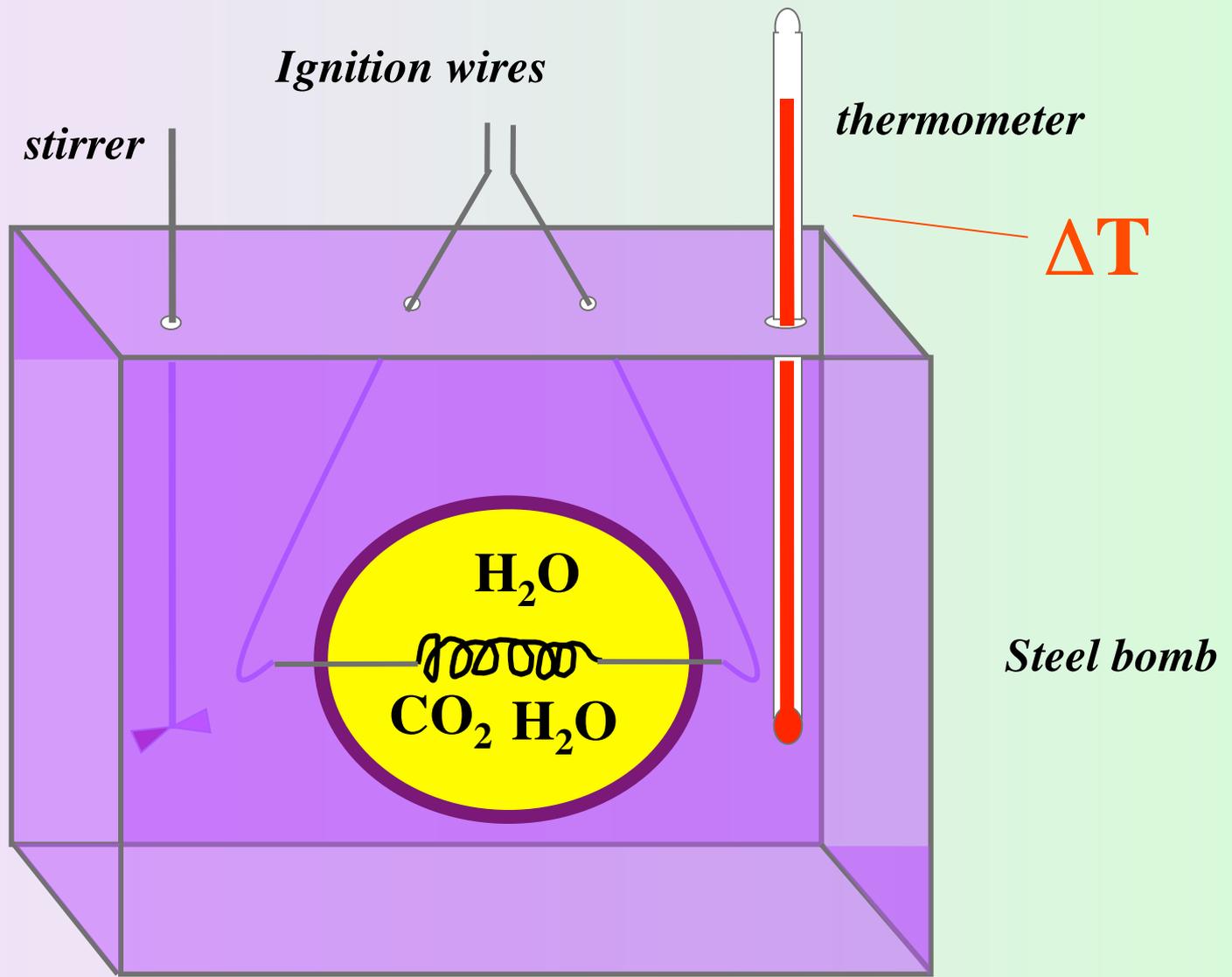
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Example:

1.50g sample of methane gas was burned with excess oxygen in a bomb calorimeter causing a temperature change of 7.3 °C. Calculate the energy of combustion (per gram) if the heat capacity of the calorimeter was 11.3 kJ/ °C.







Example:



Energy released in
the combustion of
1.50g of CH_4 = $(11.3 \text{ kJ}/^\circ\text{C})(7.3^\circ\text{C}) = 83 \text{ kJ}$

$$\frac{83 \text{ kJ}}{1.50\text{g}} = 55 \text{ kJ/g}$$