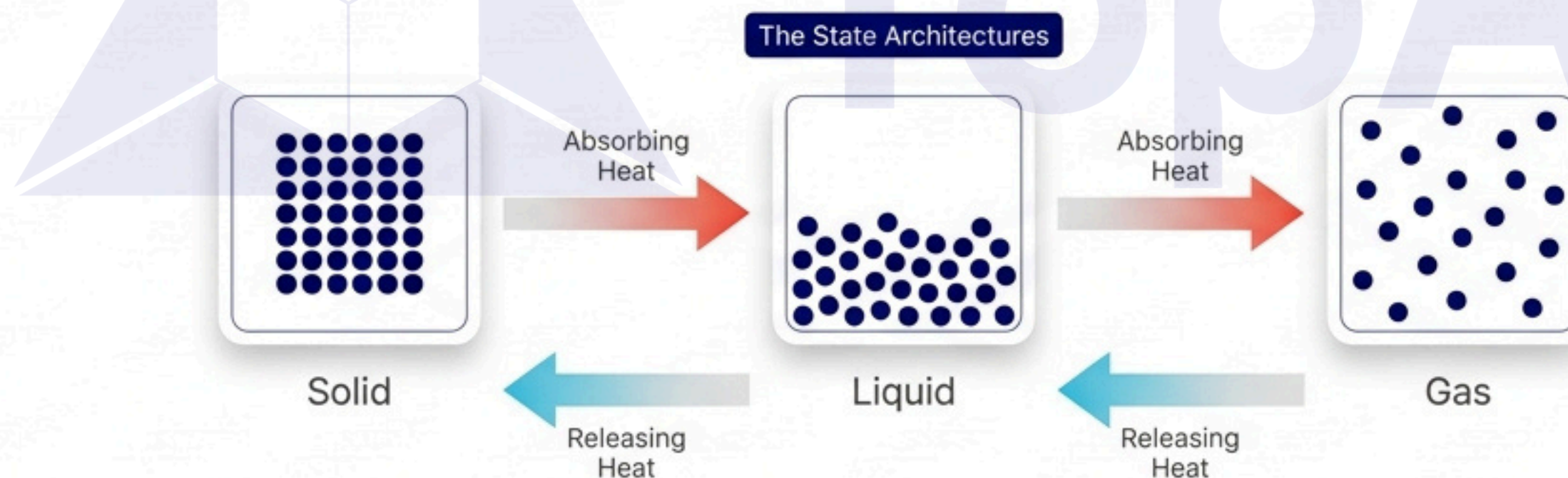


CLASS 11 CHEMISTRY • THERMODYNAMICS

Enthalpy Changes during Phase Transformations

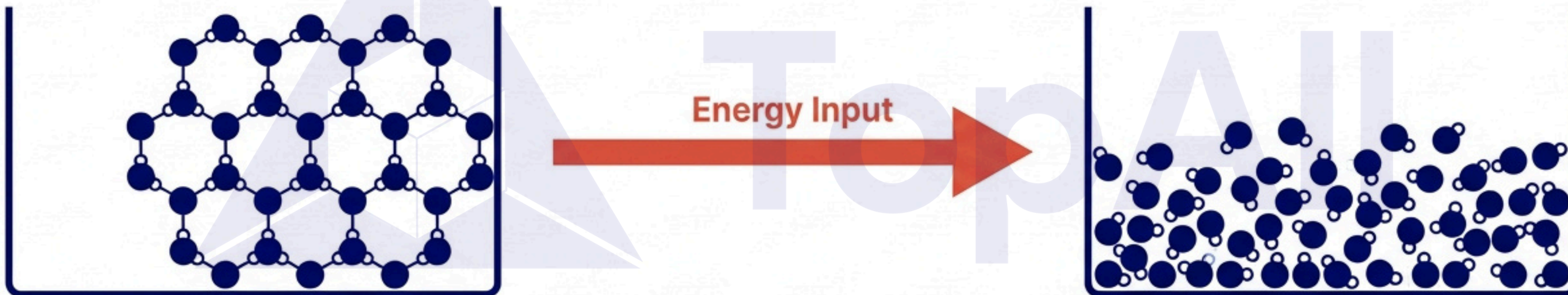
The Thermodynamics of State Changes: Fusion, Vaporisation, and Reversibility



Fusion: Breaking the Crystal Lattice

Standard Enthalpy of Fusion ($\Delta_{\text{fus}}H^\circ$)

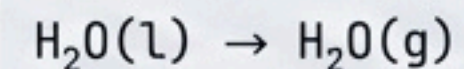
The heat energy required to melt one mole of a solid substance into a liquid at a constant pressure and standard state.



$$\Delta_{\text{fus}}H^\circ = +6.00 \text{ kJ mol}^{-1}$$

Melting is **inherently endothermic**. Because energy must be absorbed to overcome the intermolecular forces holding the solid structure together, the enthalpy of fusion is always a positive value.

Vaporisation: Complete Molecular Separation

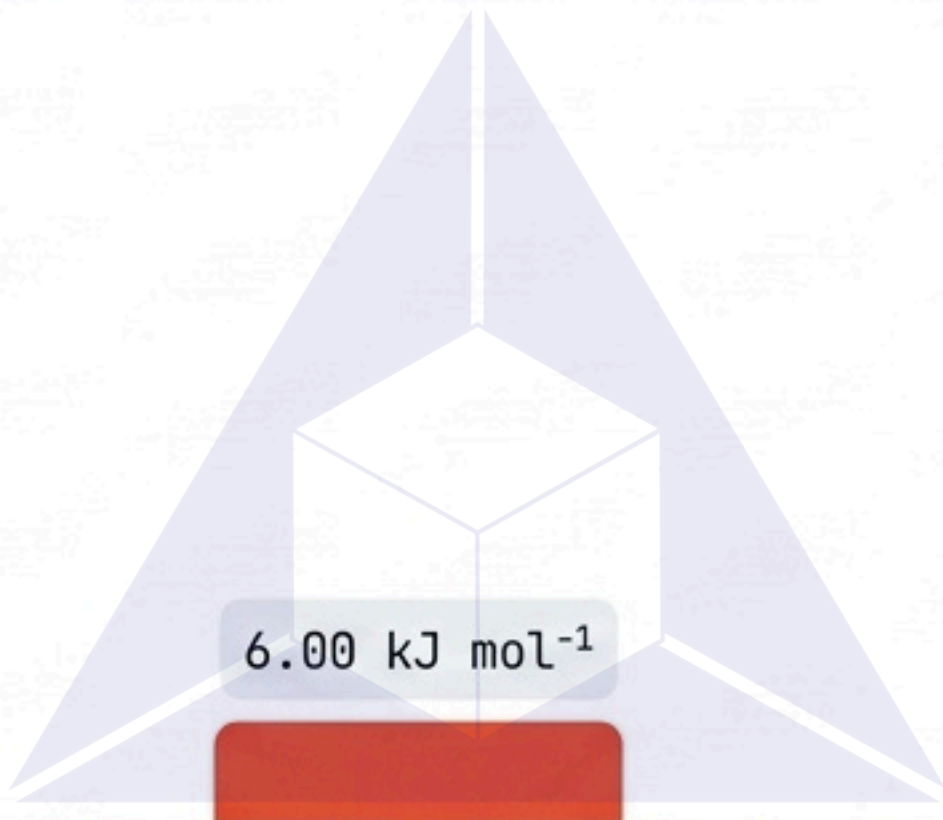


$$\Delta_{\text{vap}}H^\circ = +40.79 \text{ kJ mol}^{-1}$$

The 7x Multiplier

Fusion

Loosening the structure

A diagram showing a white cube representing ice, partially submerged in a red rectangular block representing heat energy. The cube is surrounded by a light blue triangular shape. A label '6.00 kJ mol⁻¹' is placed over the red block.
$$6.00 \text{ kJ mol}^{-1}$$

Vaporisation

Completely separating molecules

A tall, solid red rectangular bar representing the energy required for vaporisation. A label '40.79 kJ mol⁻¹' is placed to its right.
$$40.79 \text{ kJ mol}^{-1}$$

The 'Why'

Why the massive difference? The magnitude of enthalpy change depends on the strength of intermolecular interactions. Water is held tightly together by strong **hydrogen bonds**. While melting only **loosens** these bonds, boiling requires completely **breaking** them, demanding vastly more heat energy.



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$$E=mc^2$$



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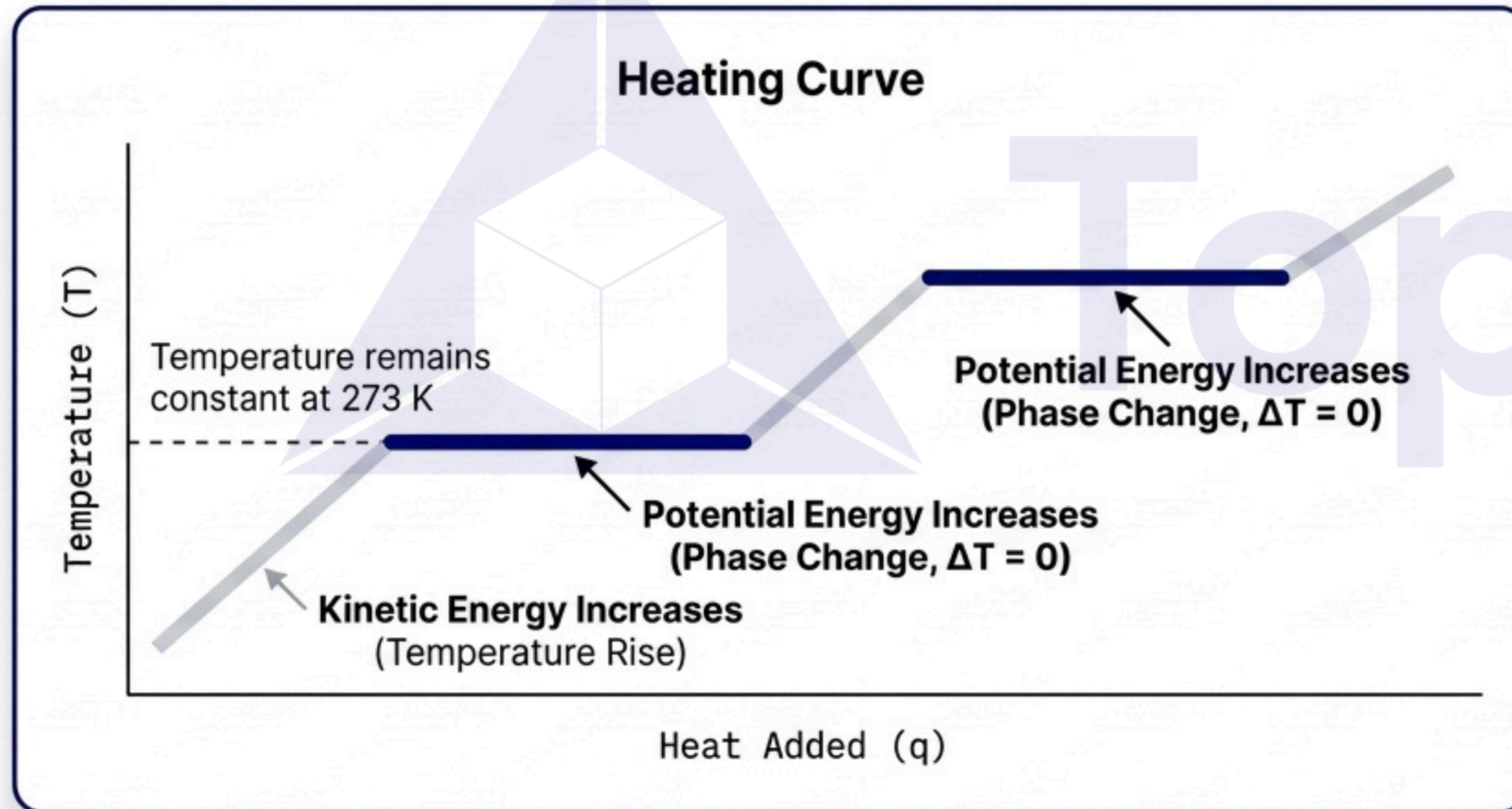
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The Constant Temperature Paradox

Why Temperature Stays Constant During Phase Changes



During a phase transformation, added heat does not increase the kinetic energy (speed) of the molecules. Instead, the energy is entirely consumed as potential energy—used specifically to break intermolecular bonds. Therefore, heat input does not equal temperature increase.

heat $\neq$ temperature $\uparrow$

Reverse Processes: Freezing and Condensation

Endothermic (Forward)

Melting: $\Delta H = +6.00 \text{ kJ mol}^{-1}$

Boiling: $\Delta H = +40.79 \text{ kJ mol}^{-1}$

Exothermic (Reverse)

Freezing: $\Delta H = -6.00 \text{ kJ mol}^{-1}$

Condensation: $\Delta H = -40.79 \text{ kJ mol}^{-1}$

When a physical state-change is reversed, the exact same amount of energy is given off to the surroundings. The magnitude stays the same; only the mathematical sign ($\pm$) flips.

The Phase Change Enthalpy Cheat Sheet

Process	Equation	Energy Flow	Sign of ΔH
Fusion (Melting)	Solid $\rightarrow$ Liquid	Absorbs Heat (Endothermic)	POSITIVE (+)
Vaporisation (Boiling)	Liquid $\rightarrow$ Gas	Absorbs Heat (Endothermic)	POSITIVE (+)
Freezing	Liquid $\rightarrow$ Solid	Releases Heat (Exothermic)	NEGATIVE (-)
Condensation	Gas $\rightarrow$ Liquid	Releases Heat (Exothermic)	NEGATIVE (-)

Golden Rule: Standard phase change enthalpies for **melting** ($\Delta_{\text{fus}}H^\circ$) and **boiling** ($\Delta_{\text{vap}}H^\circ$) are **always positive** because separating molecules fundamentally requires an input of energy.



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