



Unit 6

Thermodynamics

6.1 Endothermic & Exothermic Processes

6.2 Energy Diagrams

6.3 Heat Transfer & Thermal Equilibrium

Heat vs. Temperature

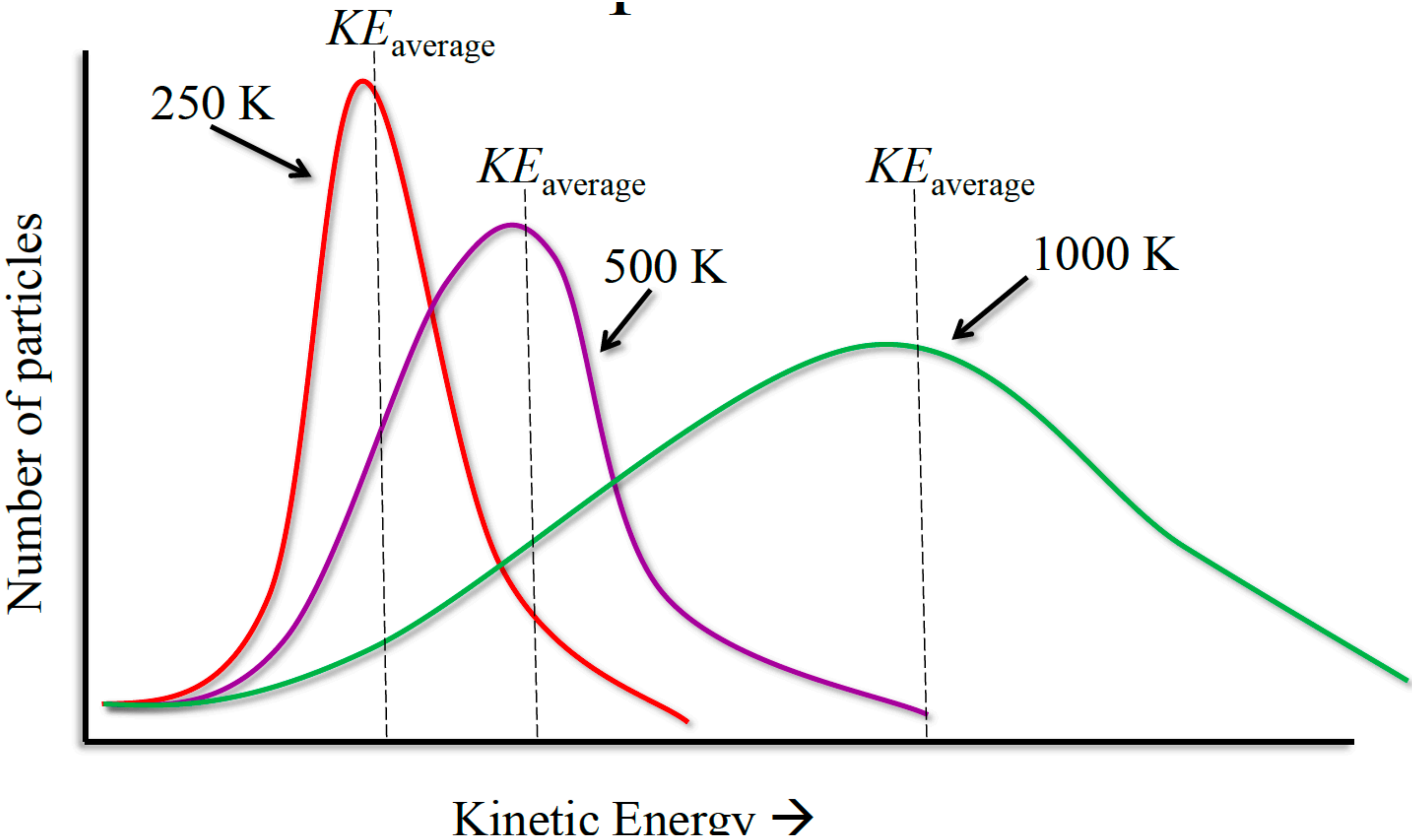
Heat

1. A form of energy measured in Joules (J).

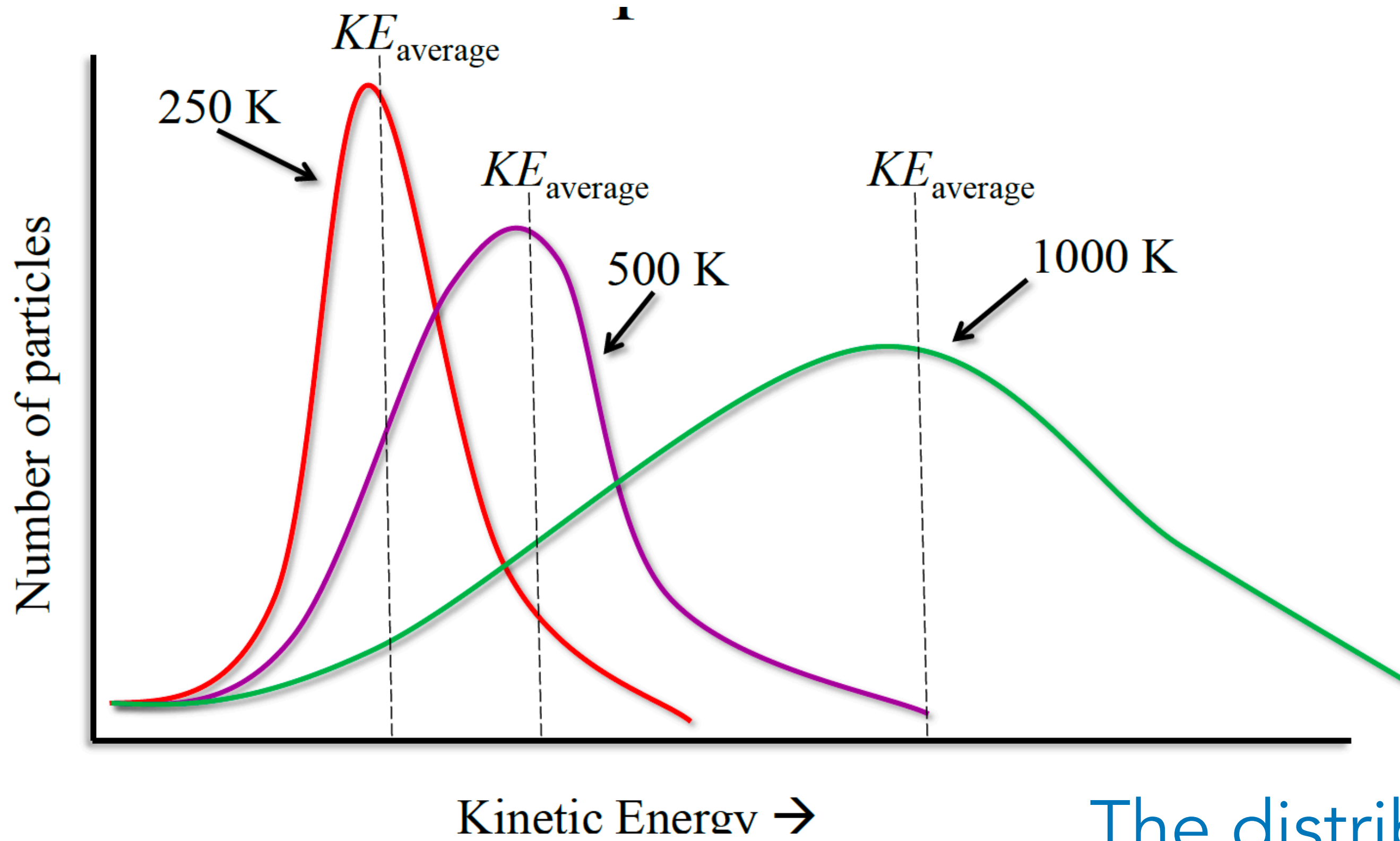
Temperature

1. A measure of the average kinetic energy of atoms and molecules in a system.
 - The Kelvin (K) temperature scale is proportional to this.
 - When KE doubles, the Kelvin temperature doubles.

Boltzmann Distribution & Temperature



Boltzmann Distribution & Temperature



The distribution of kinetic energies increases as temperature increases.

The average kinetic energy of the particles in a system increases as temperature increases.

Indications of Energy Changes

Macroscopic indications that an energy change has occurred:

- **Changes in temperature**

- after mixing two solutions that are at the same temperature, the temperature of the final solution is either higher or lower than the original solution.
- Heating or cooling a substance changes its internal energy.

- **Changes in volume**

- The volume of a gaseous system increases or decreases.

Causes of Energy Changes

- **Heating or cooling a substance**

- Adding heat to, or removing heat from, a substance changes its internal energy.
- If the temperature changes, the average KE of the particles in the system changes

- **Phase changes**

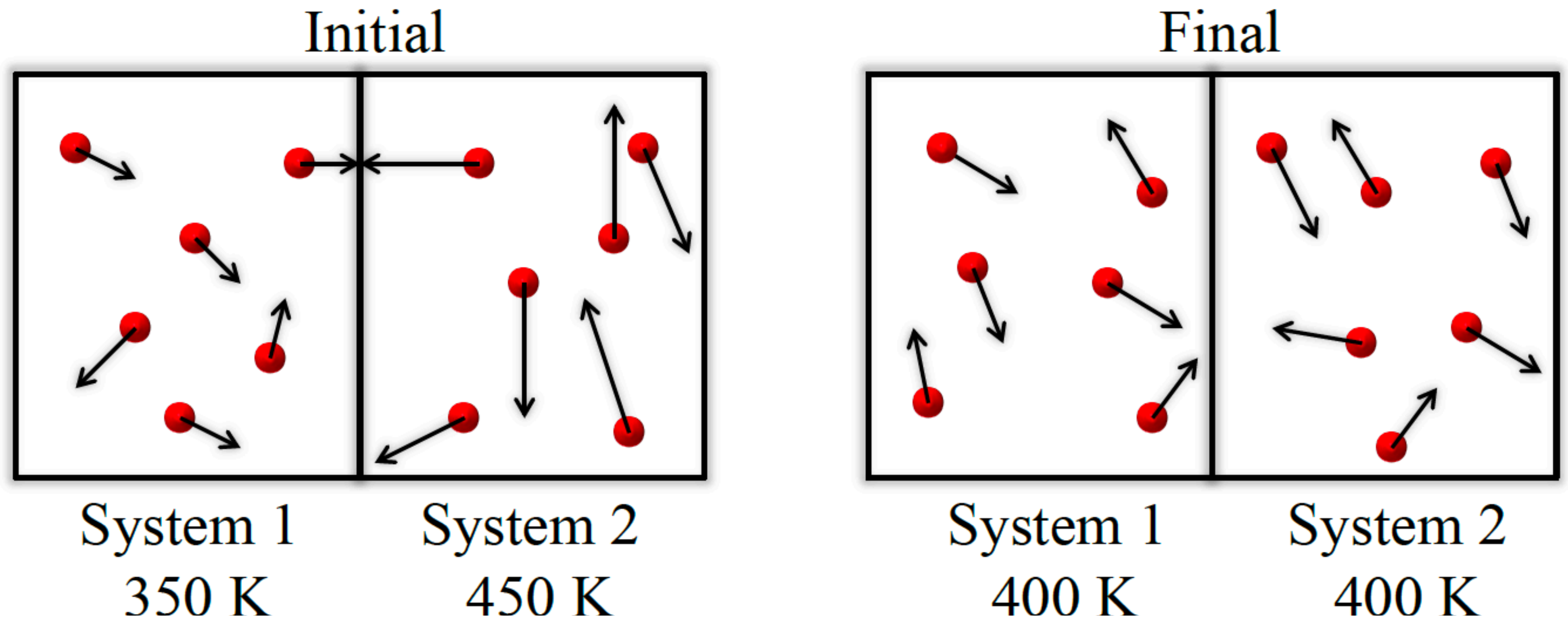
- Temperature does not change during a phase change.
- The energy that is released or absorbed by the system is used to form, alter, or sever intermolecular forces.

- **Chemical reactions**

- During chemical reactions, energy is released or absorbed in the form of heat and/or work.

Heat Transfer (Exchange)

- Two systems at different temperatures that are in thermal contact with one another will exchange energy (heat).
- The energy transferred to system 1 is equal to the energy transferred from system 2.



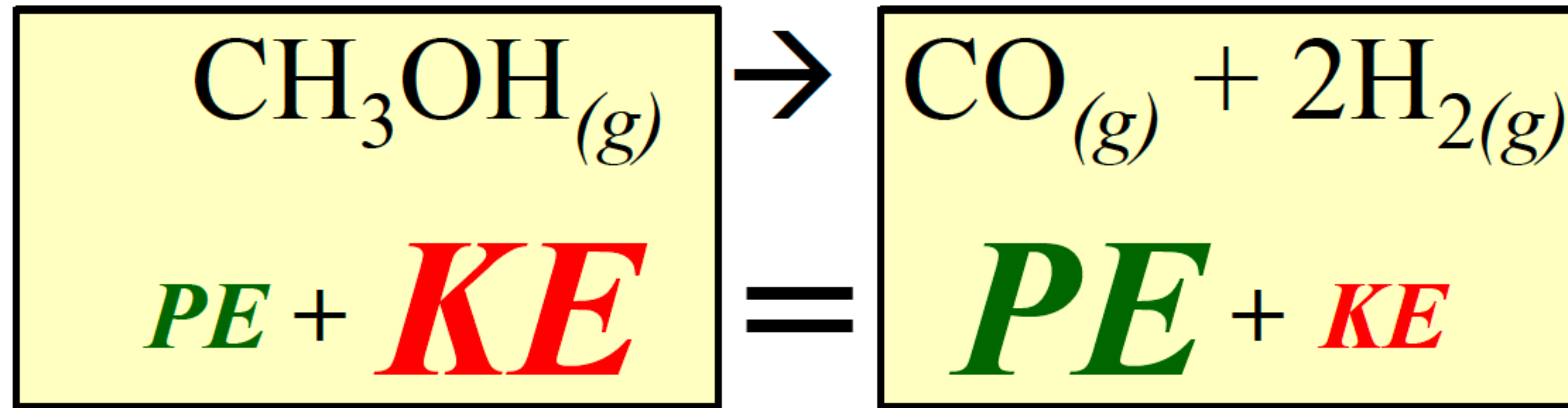
Conservation of Energy

Energy can neither be created nor destroyed, but it can be transformed from one form to another.

- chemical reactions
- phase changes
- heating or cooling
- **System**
 - The actual chemical reaction that is taking place.
- **Surroundings**
 - The entire universe outside of the chemical reactions

Heat and Endothermic Processes

In an isolated system, energy is conserved.

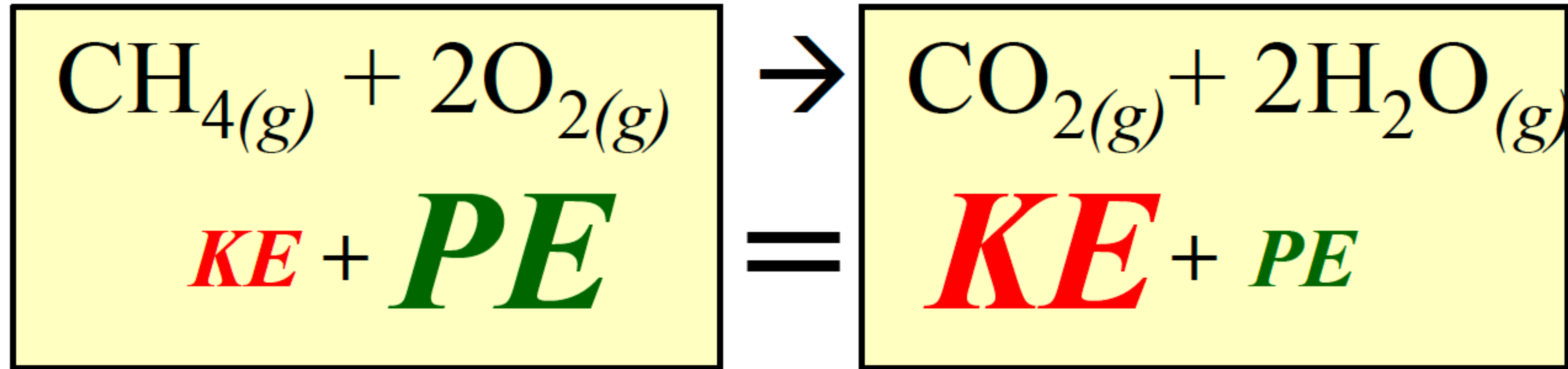


$$\Delta H = +90.7 \text{ kJ}$$

- As the bonds in the products contain more PE, the product molecules must contain less KE.
- This is why the products of an endothermic reaction are at a lower temperature.
- Heat flows from the surroundings into the system until thermal equilibrium is established.

Heat and Exothermic Processes

In an isolated system, energy is conserved.



- As the bonds in the products contain less PE, the product molecules must contain more KE.
- This is why the products of an exothermic reaction are at a higher temperature.
- Heat flows from the system into the surroundings until thermal equilibrium is established.

Energy Changes and Chemical Reactions

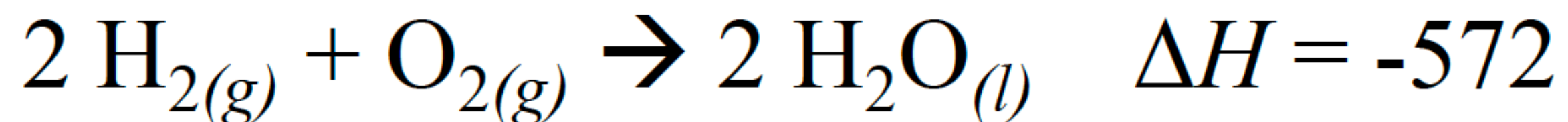
- During chemical reactions, the internal energy of the system can:
 - Decrease (exothermic)
 - Increase (endothermic)
 - Remain the same



Enthalpy Change (ΔH)

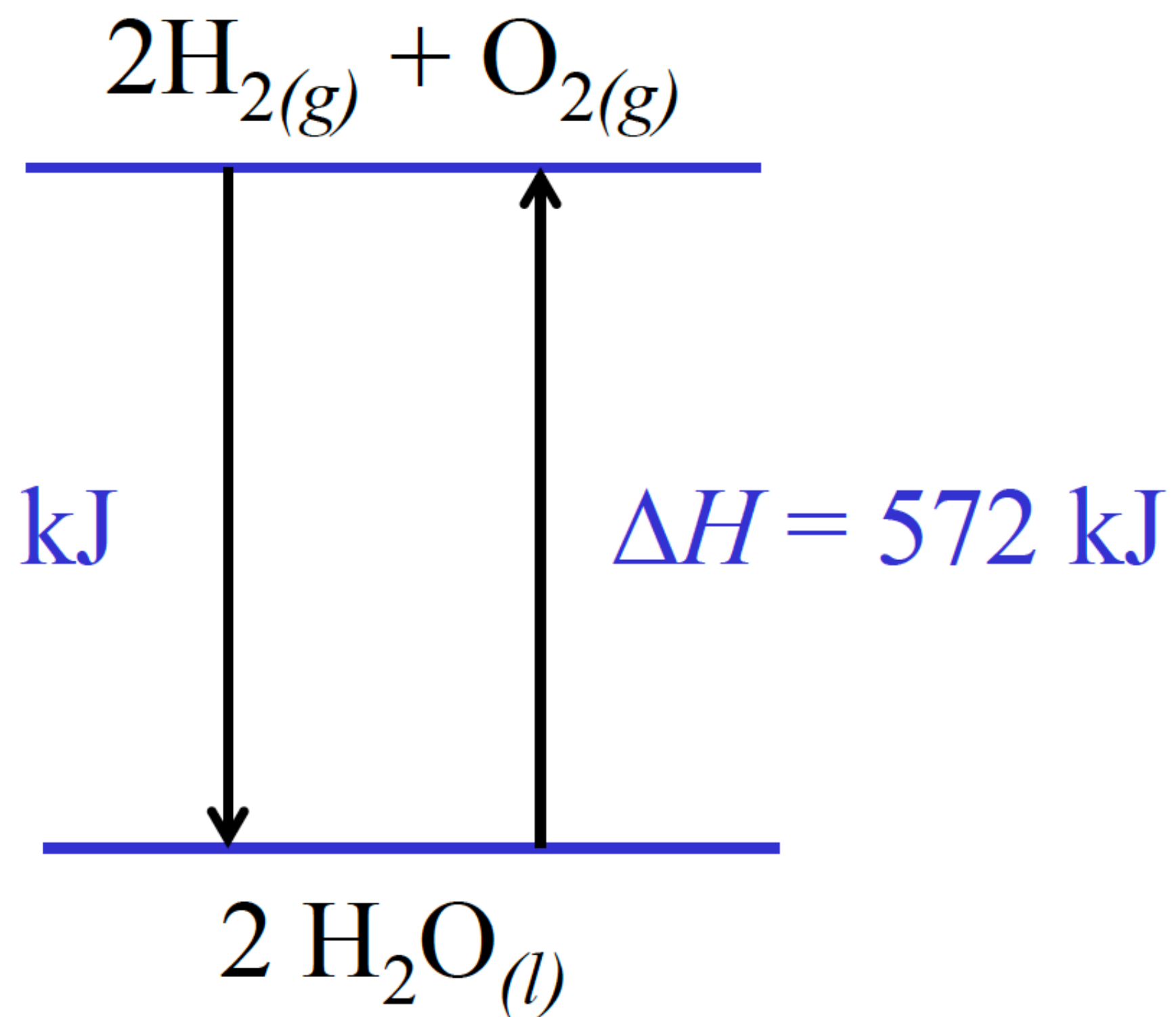
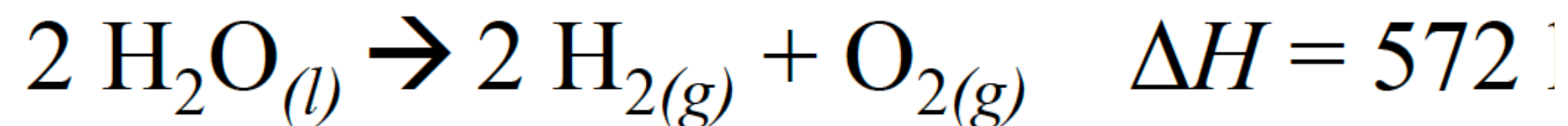
- **Hydrogen Fuel for Your Car**

Burning Hydrogen Fuel



$$\Delta H = -572 \text{ kJ}$$

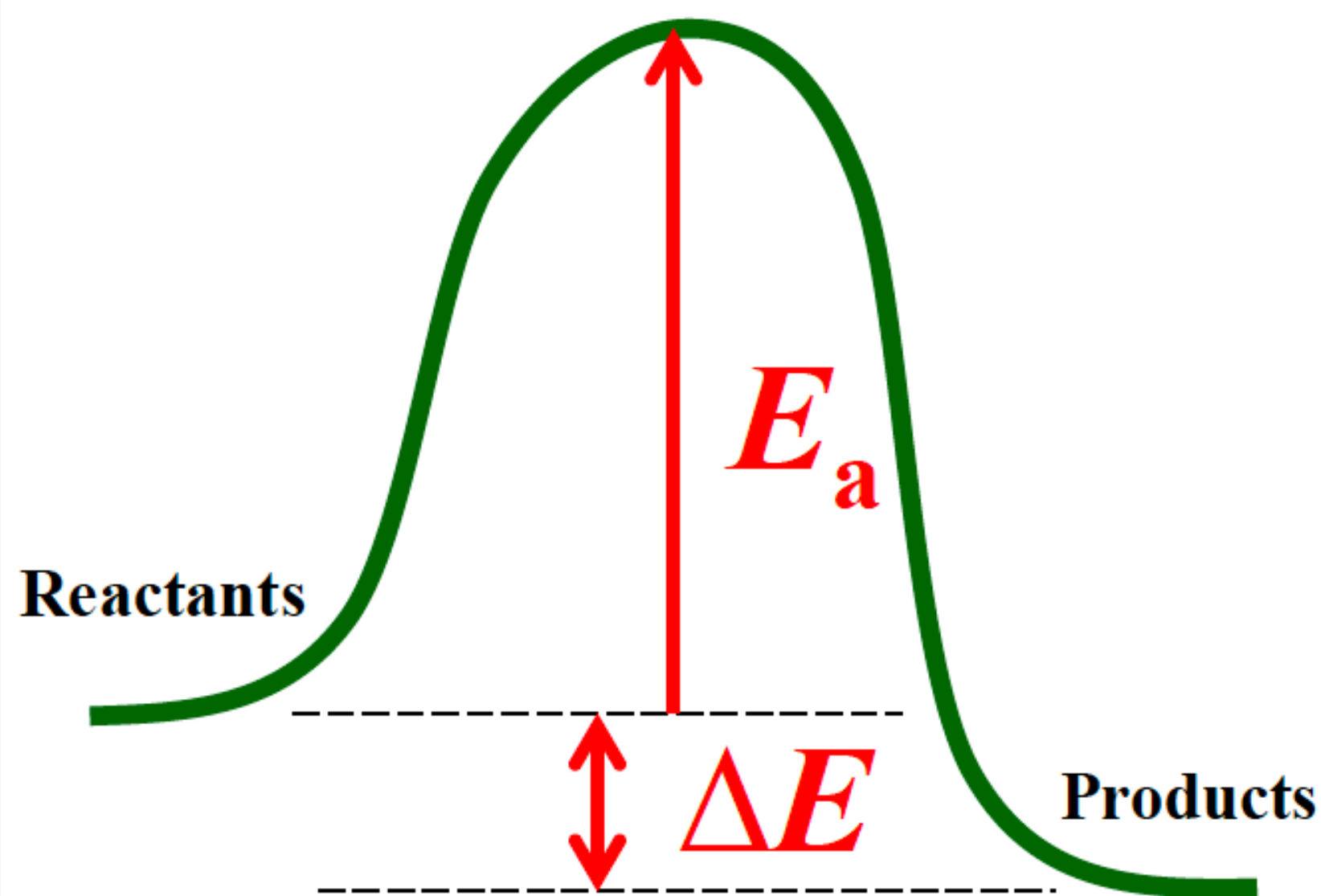
Making Hydrogen Gas



Energy Profile Diagrams

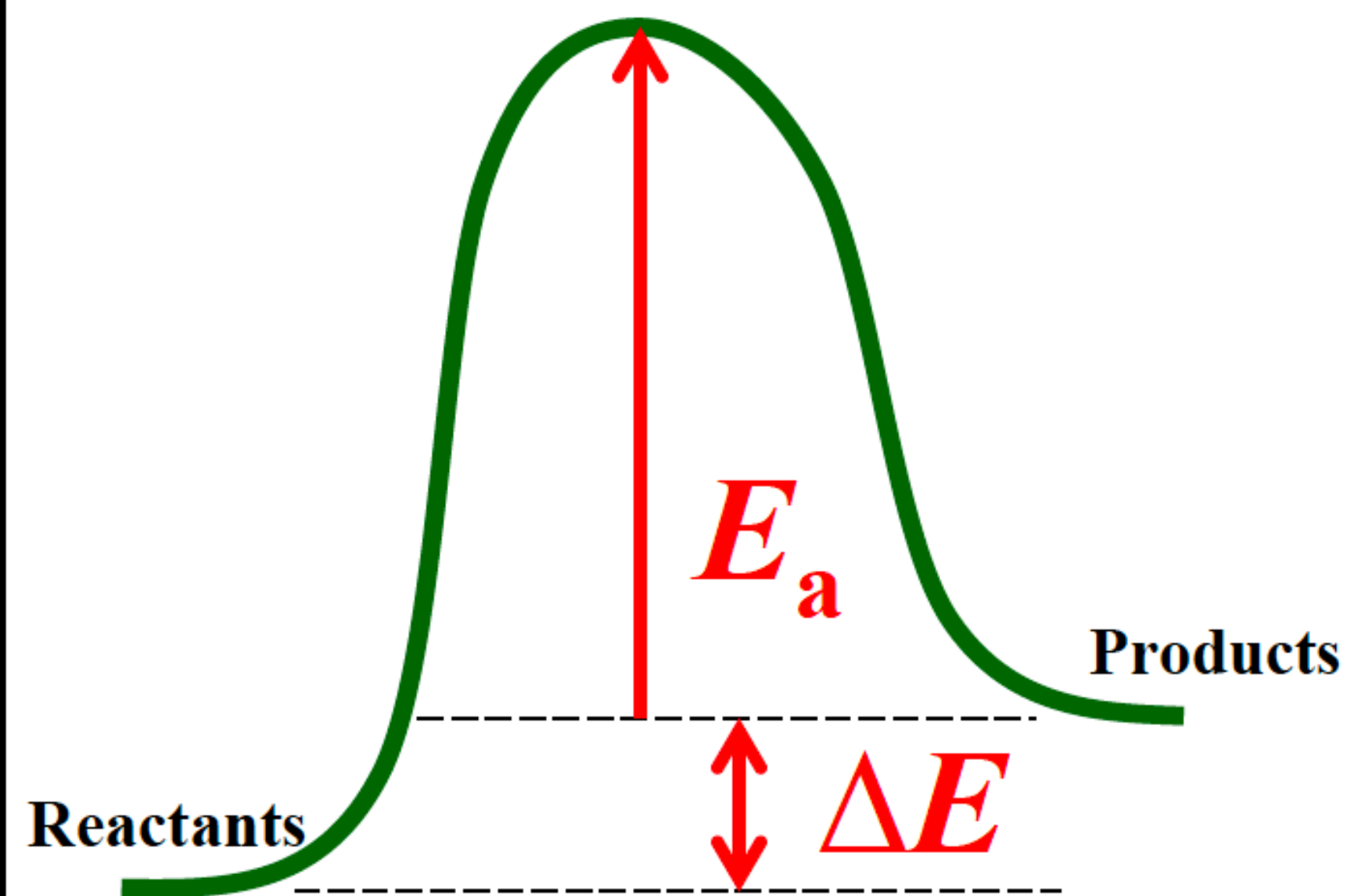
Exothermic Reaction

Bonds in products contain less PE than bonds in reactants.



Endothermic Reaction

Bonds in products contain more PE than bonds in reactants.

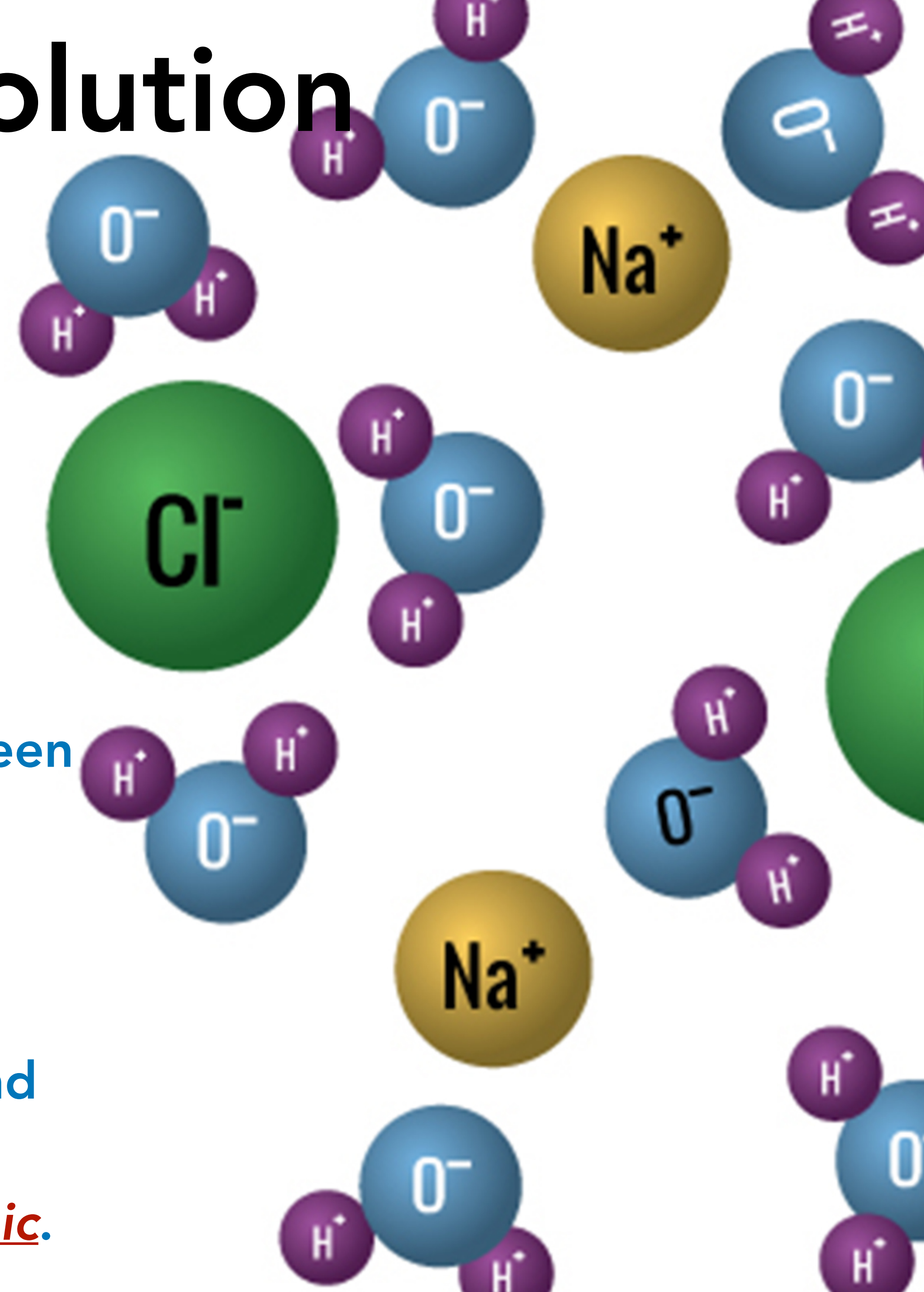


ΔE is the energy lost or gained in a reaction.

$$\Delta E = q \cdot w \quad \text{If the volume does not change, } \Delta E = \Delta H$$

Ex: Forming a $\text{NaCl}_{(aq)}$ Solution

- Ionic bonds must be broken
- Hydrogen bonds between some of the water molecules must be broken.
 - Both of these are **ENDOTHERMIC** processes.
- Ion-dipole forces of attraction are established between Na^+ and H_2O and between Cl^- and H_2O .
 - These are **EXOTHERMIC** processes.
- More energy is required to break the ionic bonds and H-bonds than is released when the ion-dipole forces are established, so the overall process is **endothermic**.

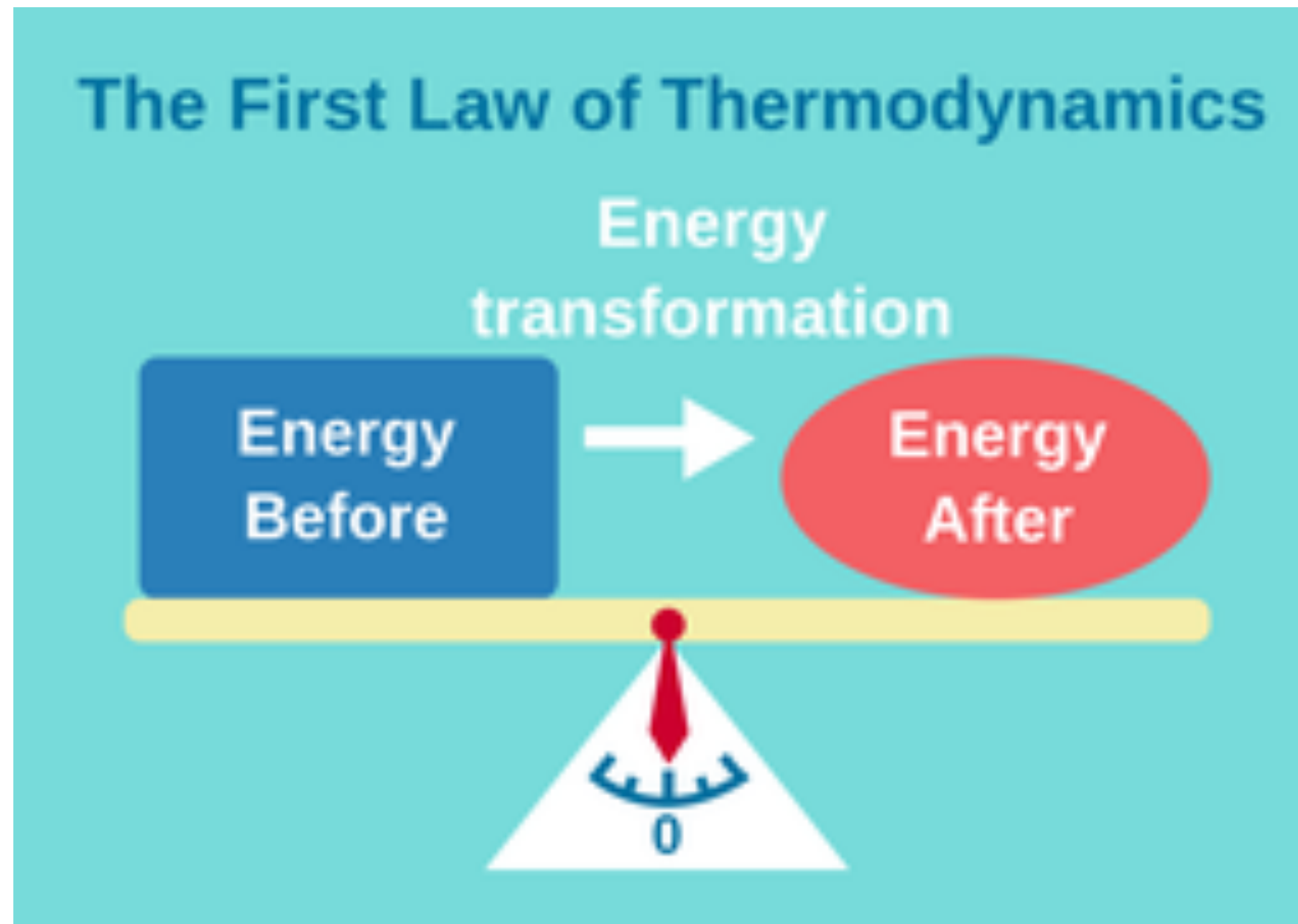


6.4 Heat Capacity & Calorimetry

6.6 Introduction to Enthalpy of Reaction

First Law of Thermodynamics

- Energy is conserved in chemical and physical processes.
- The internal energy of a system may be transferred into or out of that system in the form of heat or work, but it cannot be created or destroyed.



Heating & Cooling

- When heat flows into a system, the energy of that system increases.



The internal energy of the system increases as heat flows into the system.

- When heat flows out of a system, the energy of that system decreases.



The internal energy of the system decreases as heat flows out of the system.

Specific Heat Capacity (C)

- Specific heat capacity, C , is the amount of heat required to raise the temperature of 1.0 g of a substance by 1.0 K.
- A large value for water (4.184 J/g•K)
 - It takes a lot of heat to increase the temperature of 1.0 g of water by 1.0 K.
- A small value for iron (0.45 J/g•K)
 - It takes much less heat to raise the temperature of 1.0 g of Fe by 1.0 K.

Example: Specific Heat Capacity (C)

- 10.0 J of heat are absorbed by a 1.0 g sample of water at 25.0°C and 10.0 J of heat are absorbed by a 1.0 g sample of iron at 25.0°C. Which sample will experience the greatest temperature? Justify your answer.



Specific Heat Capacity (C)

- The heat transfer equation shows that the transfer of equal quantities of heat to two materials of equal masses with different specific heat capacities will produce different temperature changes.
- The material with the higher specific heat capacity will experience a smaller increase in temperature (logically).

$$q = mC\Delta T$$

q = heat lost or gained

m = mass of that substance

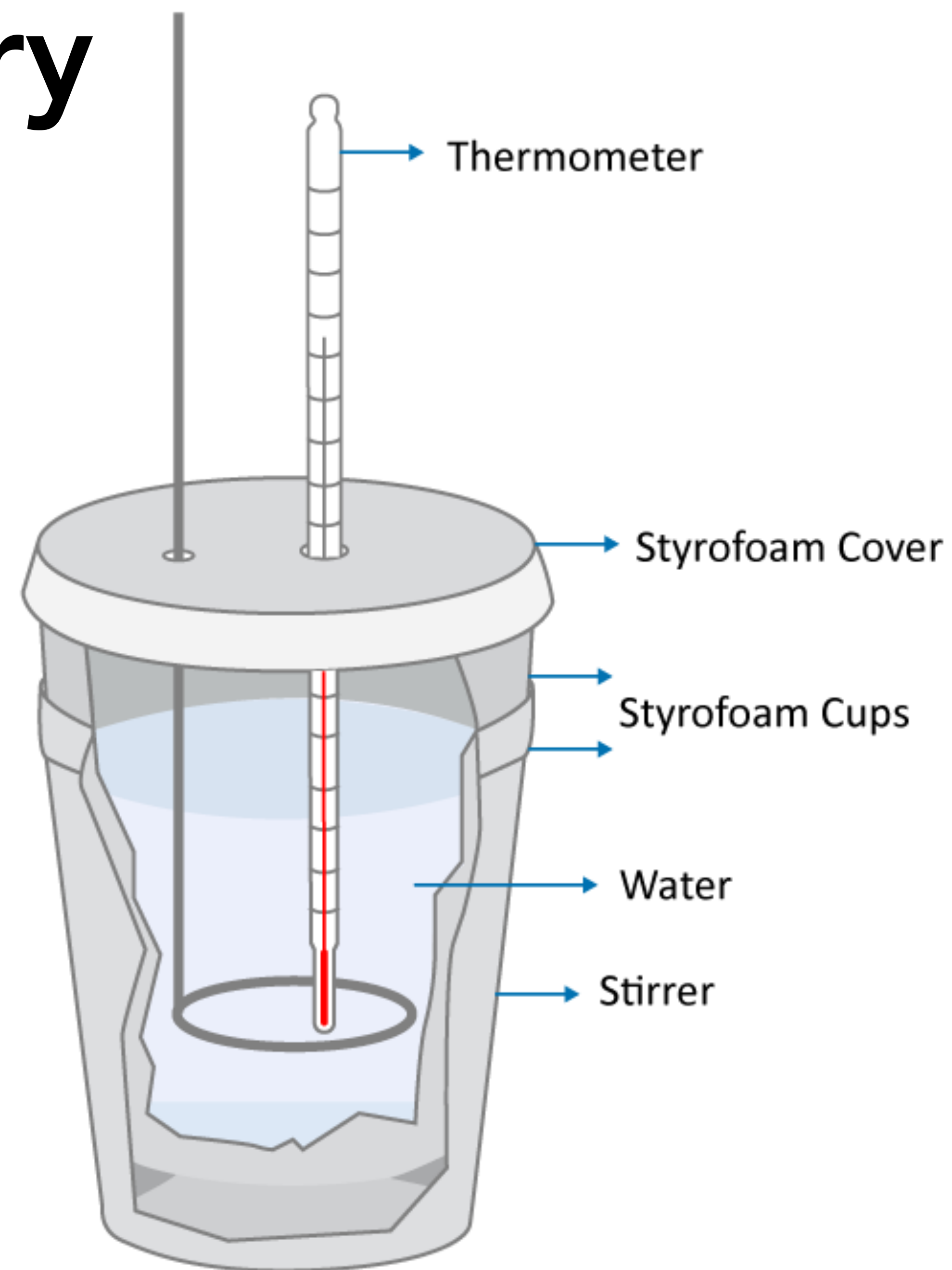
C = specific heat capacity

ΔT = Temperature change of that substance ($\Delta T = T_f - T_i$)

$$q = mC\Delta T$$

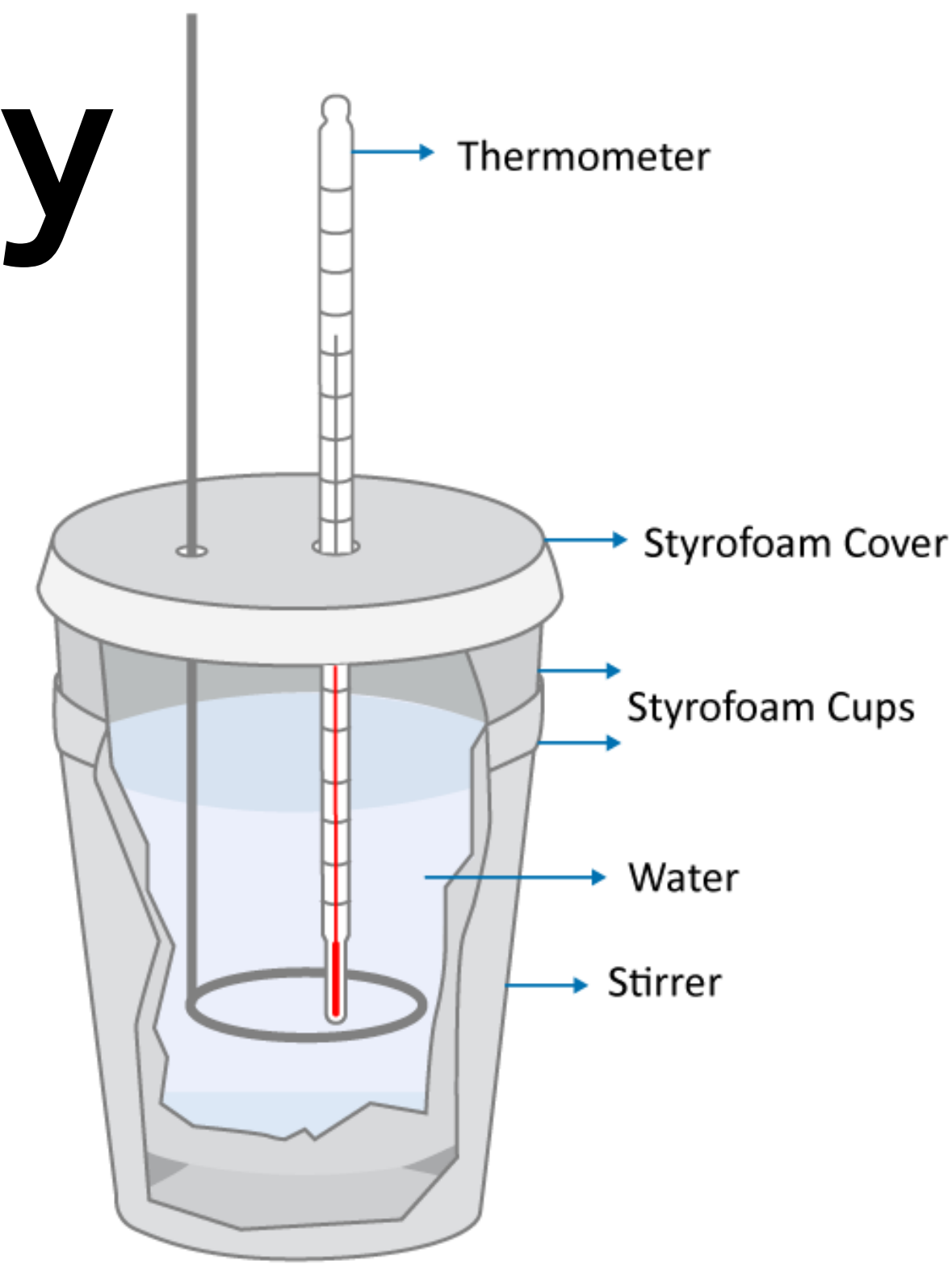
Calorimetry

- The measure of heat transfer using a **calorimeter**.
- A device used to determine the amount of heat transferred.
- Heat flow cannot be measured directly.
- We measure the temperature before and after a reaction, and use that data to calculate heat transfer.



Coffee Cup Calorimetry

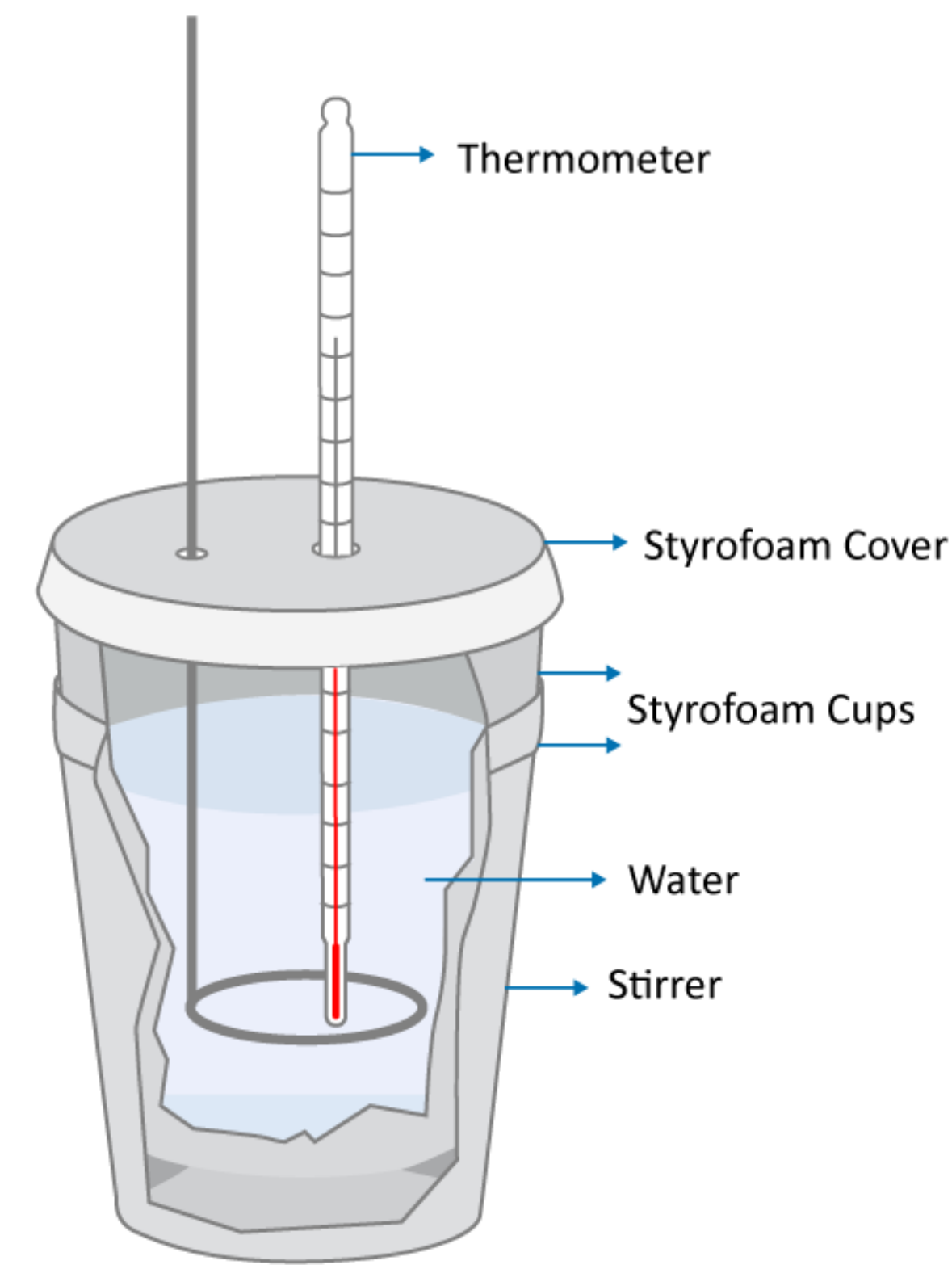
- The reactants and products are the system.
- The water in which they are dissolved and the calorimeter make up the surroundings.
- The heat lost or gained by the solution is equal to the heat lost or gained by the reaction.
- e.g. - if you calculated q_{solution} to be +568 J, the solution gained heat. This means the reaction was exothermic, as the heat gained by the solution is equal to the heat lost by the reaction ($q_{\text{rxn}} = -568 \text{ J}$).



Ex. Calorimetry

- A 97 g iron bar was heated to 143°C and placed in a coffee cup calorimeter containing 350.0 g of water at 25.00°C. The maximum temperature of the water was measured to be 28.40°C.

- a) How much heat was gained by the water?
- b) How much heat was lost by the iron bar?
- c) What is the specific heat of iron?



$$q_{\text{Fe}} = mc\Delta T$$

$$c = \frac{q_{\text{Fe}}}{m\Delta T}$$

$$c = \frac{-4979 \text{ J}}{(97 \text{ g})(28.40 \text{ }^\circ\text{C} - 143 \text{ }^\circ\text{C})}$$

$$c = 0.45 \text{ J/g}\cdot^\circ\text{C}$$

Enthalpy Change

- The enthalpy change, ΔH , of a reaction is the amount of heat energy that is released (exothermic) or absorbed (endothermic) by a chemical reaction at constant pressure.
- Four methods for finding ΔH :
 - 1. Find it by performing calorimetry experiments.**
 2. Calculate it using average bond enthalpies.
 3. Calculate it using enthalpies of formation.
 4. Calculate it using Hess's Law.

Example 2: Calorimetry

- When 20.0 mL of 1.00 M NaCl at 22.00°C is mixed with 20.0 mL of 1.00 M AgNO₃ at 22.0°C in a calorimeter, a white precipitate forms and the temperature of the mixture reaches 29.60°C. Assume that the specific heat capacity of the mixture is 4.184 J/g•K, its density is 1.00 g/mL and the volumes are additive.

- Write the net ionic equation for this reaction.
- Calculate ΔH for the reaction.

Step 1) Calculate q_{sol} .

$$q_{sol} = mc\Delta T$$

$$q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g}\cdot^{\circ}\text{C})(29.60^{\circ}\text{C} - 22.00^{\circ}\text{C})$$

$$q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g}\cdot^{\circ}\text{C})(7.60^{\circ}\text{C})$$

$$q_{sol} = +1270 \text{ J}$$

Step 2) Calculate q_{rxn} .

$$q_{rxn} = -q_{sol}$$

$$q_{rxn} = -1270 \text{ J}$$

The reaction is **exothermic** as heat is released to the water.

Example 2: Calorimetry

Step 3) Calculate Moles of Ag^+ or Cl^- .

$$0.0200 \text{ L} \times \frac{1.0 \text{ moles Ag}^+}{1 \text{ L}} = 0.0200 \text{ moles Ag}^+$$

Step 4) Calculate ΔH_{rxn} .

$$\Delta H_{rxn} = \frac{q_{rxn}}{n_{\text{AgCl}}}$$

$$\Delta H_{rxn} = \frac{-1270 \text{ J}}{0.0200 \text{ moles AgCl}}$$

$$\Delta H_{rxn} = -63500 \text{ J/mol} = -63.5 \text{ kJ/mol}$$

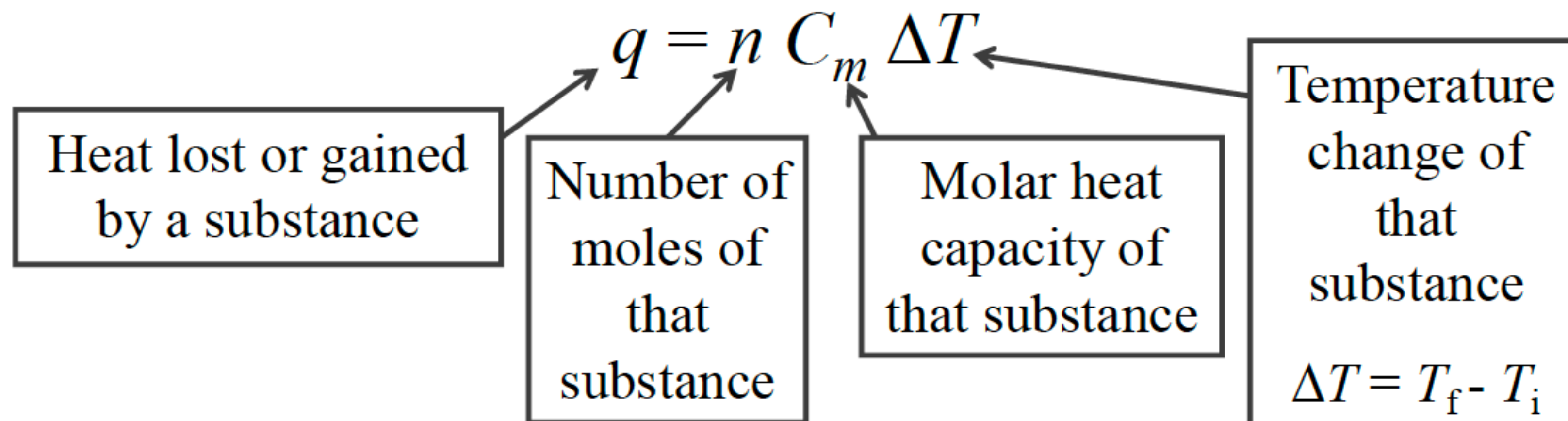
Molar Heat Capacity

- C_m is the amount of heat required to raise the temperature of 1.0 mole of a substance by 1.0 K.
- The transfer of equal quantities of heat to equal numbers of moles of two substances with different molar heat capacities will produce different temperature changes.
- The substance with the higher molar heat capacity will experience a smaller increase in temperature.

Ex. Molar Heat Capacity

- The specific heat capacity of gold is $0.129 \text{ J/g}\cdot\text{K}$.
 - a) Find the molar heat capacity of gold.
 - b) If 743 J of heat are added to 3.4 moles of gold at 2°C , what is the final temperature of the gold?

$$C_{m_{\text{Au}}} = \frac{0.129 \text{ J}}{\text{gK}} \times \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 25.4 \text{ J}/(\text{molK})$$



Ex. Molar Heat Capacity

$$q = n C_{m_{\text{Au}}} \Delta T$$

$$q = n C_{m_{\text{Au}}} (T_f - T_i)$$

$$T_f = \frac{q}{n C_{m_{\text{Au}}}} + T_i$$

$$T_f = \frac{743\text{J}}{(3.4 \text{ moles})(25.4 \text{ J}/(\text{molK}))} + (2 + 273)\text{K}$$
$$= 8.6\text{K} + 275\text{K} = 284\text{K}$$

Heat lost or gained
by a substance

Number of
moles of
that
substance

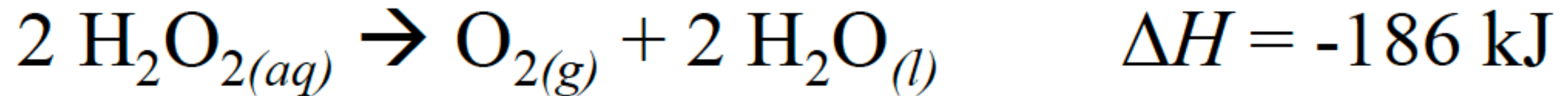
Molar heat
capacity of
that substance

Temperature
change of
that
substance
 $\Delta T = T_f - T_i$

$$q = n C_m \Delta T$$

Enthalpy & Stoichiometry

- How much heat is released when 3.56 g of $\text{H}_2\text{O}_{2(aq)}$ decomposes into water and oxygen gas?



$$3.56 \text{ g H}_2\text{O}_2 \times \frac{1 \text{ mol H}_2\text{O}_2}{34.02 \text{ g H}_2\text{O}_2} \times \frac{-186 \text{ kJ}}{2 \text{ mol H}_2\text{O}_2} = -9.73 \text{ kJ}$$

There are 2 moles H_2O_2 in the balanced equation.

Thermodynamics

9.1 Introduction to Entropy

9.2 Absolute Entropy & Entropy Change

Laws of Thermodynamics

- First Law:

- The energy contained within the universe is constant.

- Second Law:

- The entropy of the universe is constantly increasing.

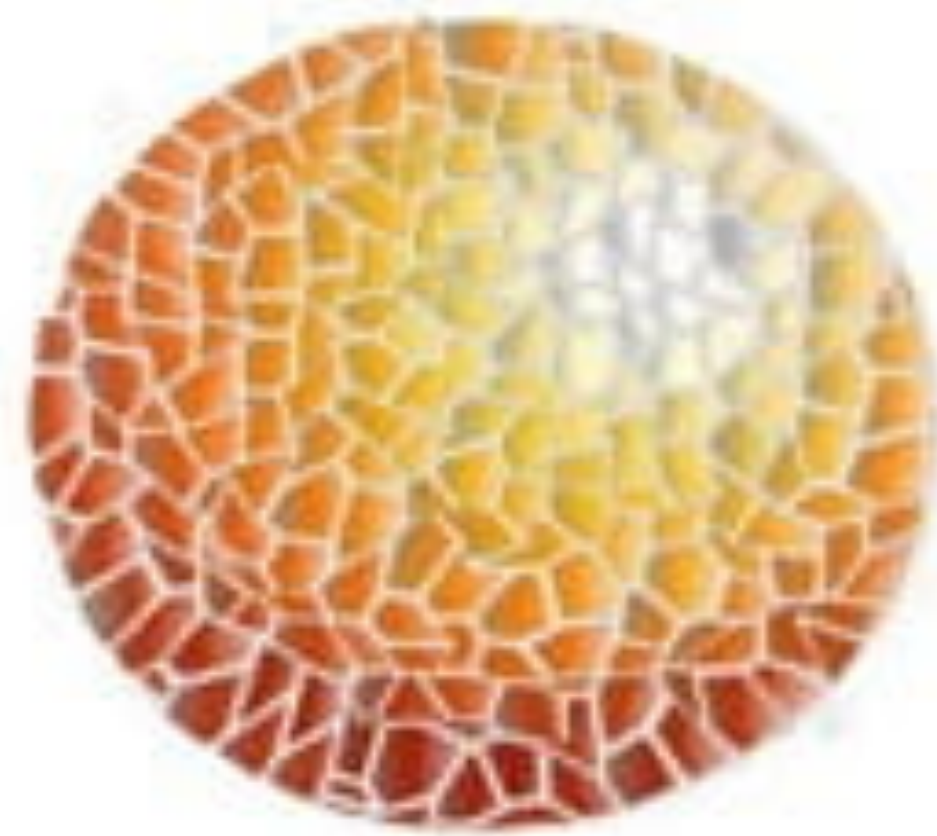
Entropy (S)

- A measure of the disorder of a system.
- Entropy increases when matter becomes more dispersed.
- A greater degree of disorder is favorable.
- A positive value for ΔS is favorable.

$$\Delta S = S_{(\text{products})} - S_{(\text{reactants})}$$

Increases in Entropy are Favored

- Your house gets dirty
- Making a stir-fry
- The impossibility of maintaining a perfect lawn
- Spilling a glass of milk
- Breaking a window



**Highly
Ordered**

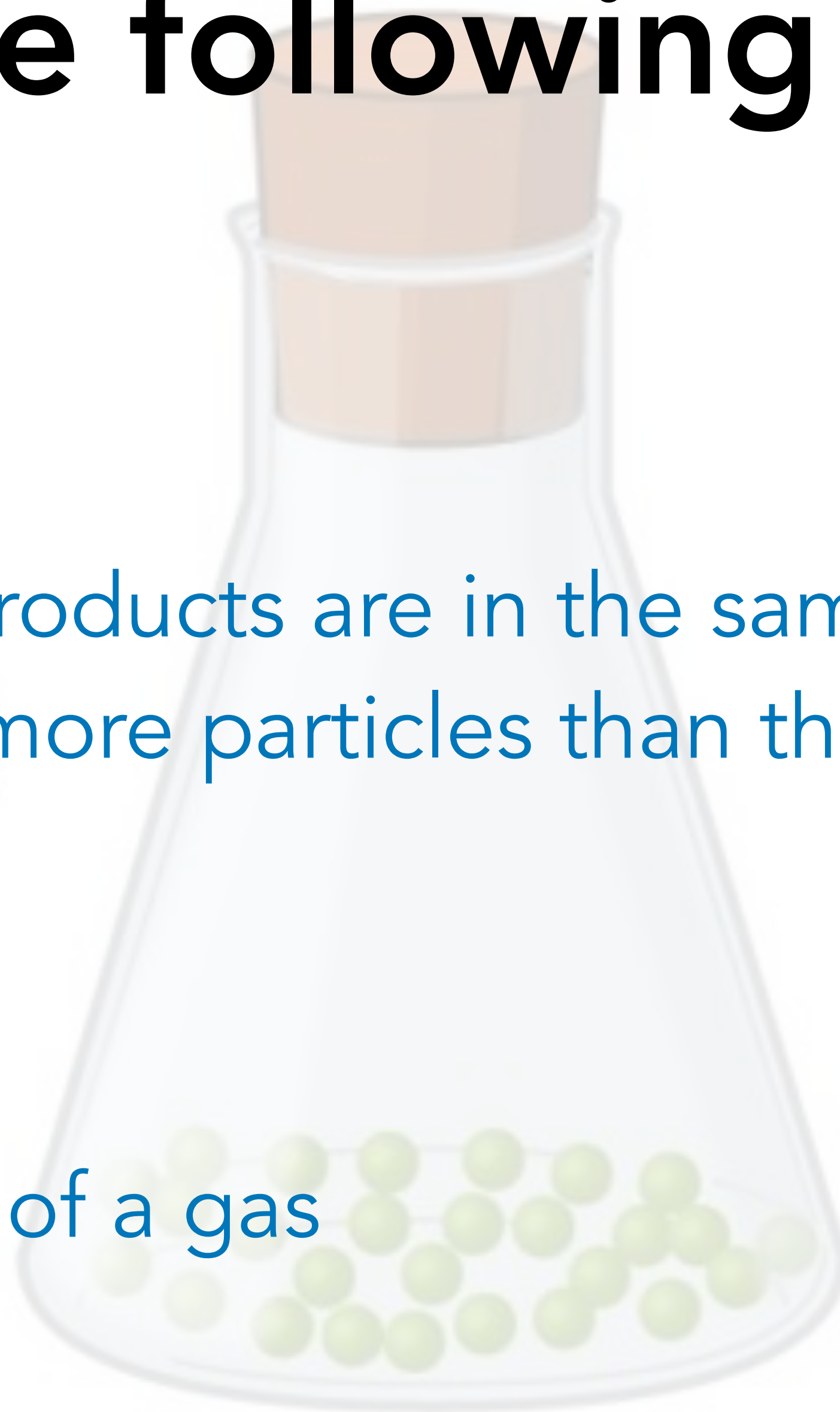
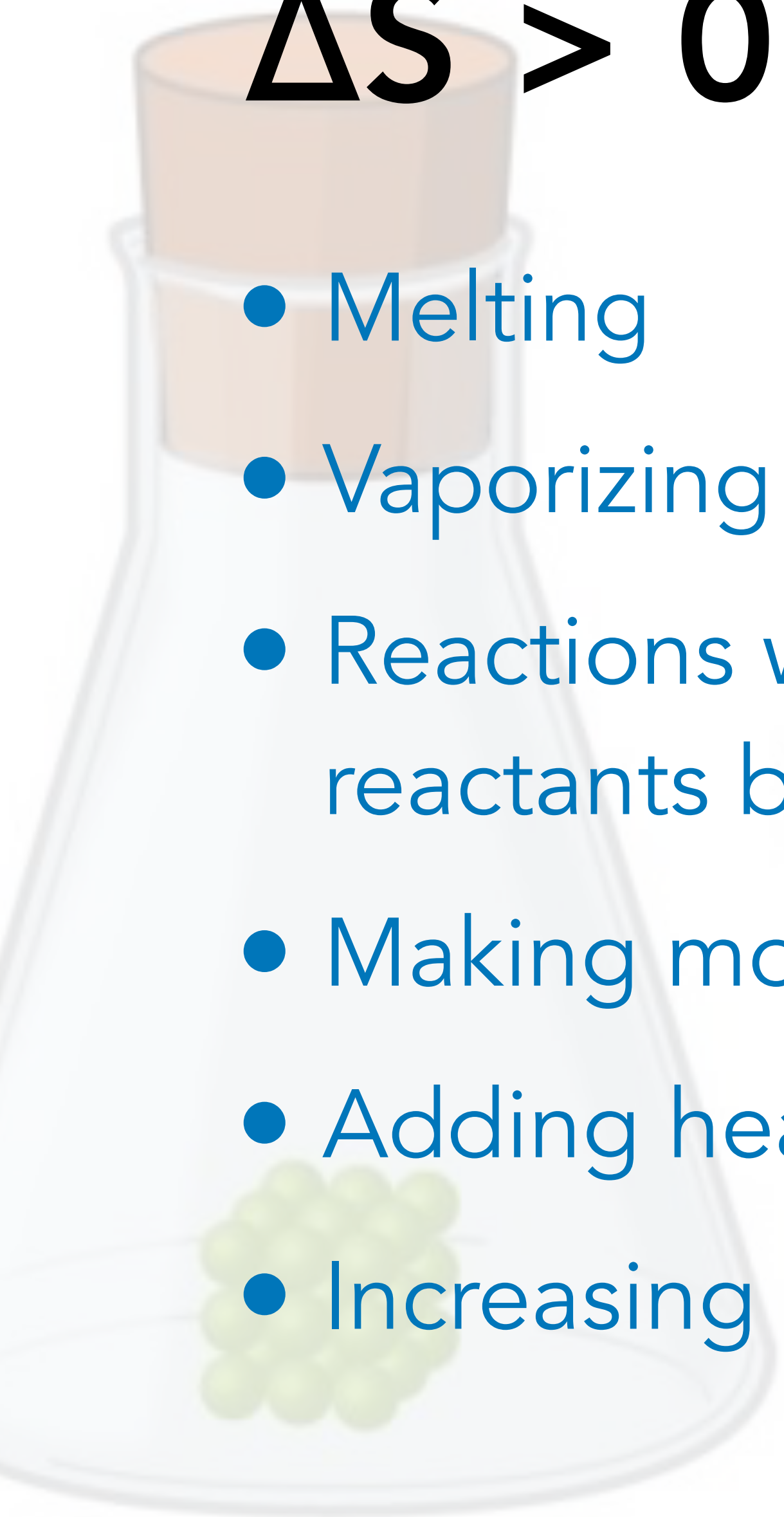
Increase
→
in Entropy



More disordered

$\Delta S > 0$ for the following processes:

- Melting
- Vaporizing
- Reactions where the products are in the same phase as the reactants but contain more particles than the reactants.
- Making most solutions
- Adding heat
- Increasing the volume of a gas

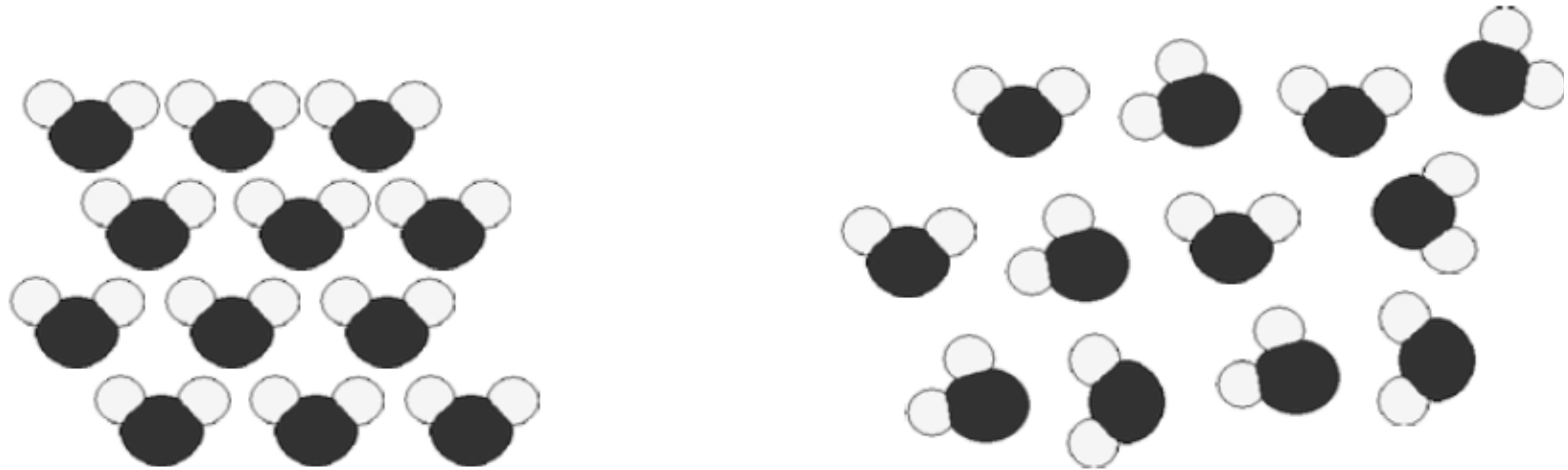


Crystalline solid

Liquid

Gas

$\Delta S > 0$ when Melting



$\text{H}_2\text{O}(s)$

$\text{H}_2\text{O}(l)$

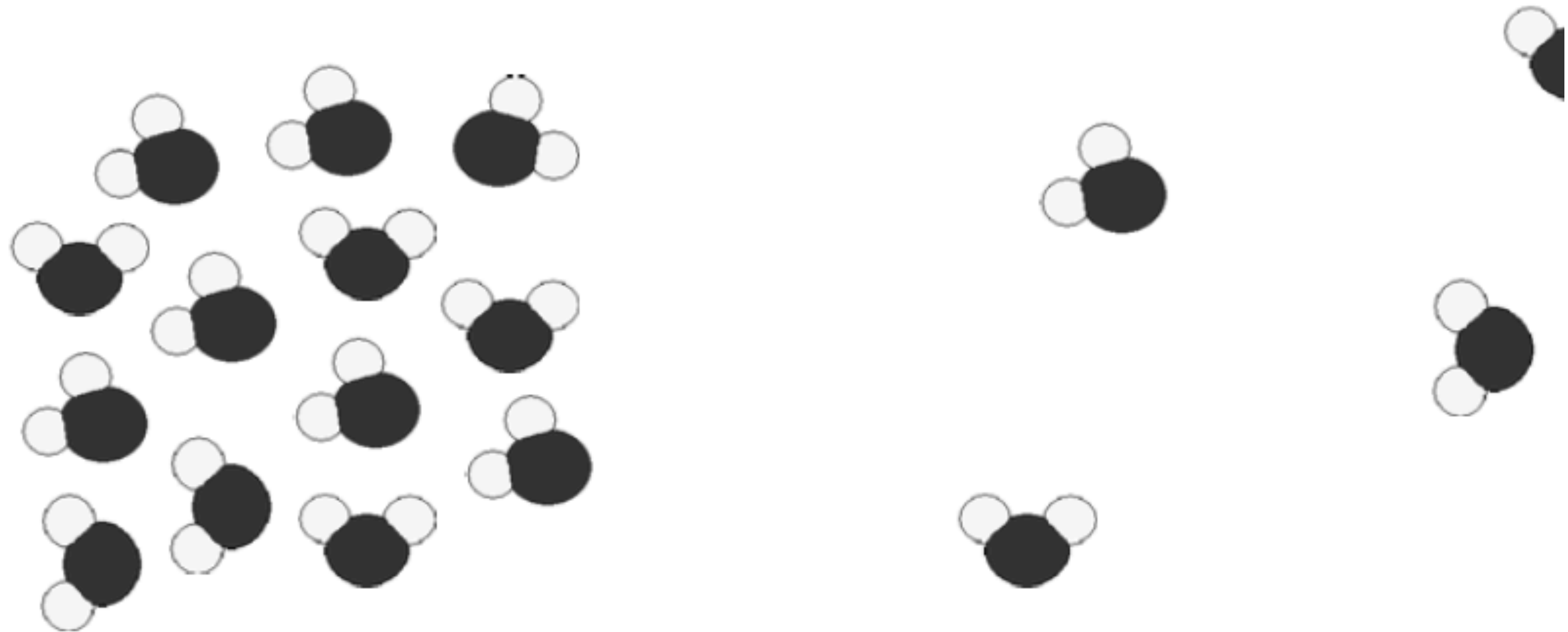
Very Organized

Less Organized

Small S

Larger S

$\Delta S > 0$ when Vaporizing



$\text{H}_2\text{O}(l)$

$\text{H}_2\text{O}(g)$

Relatively Low S

High S

$\Delta S > 0$ when Products have more Particles

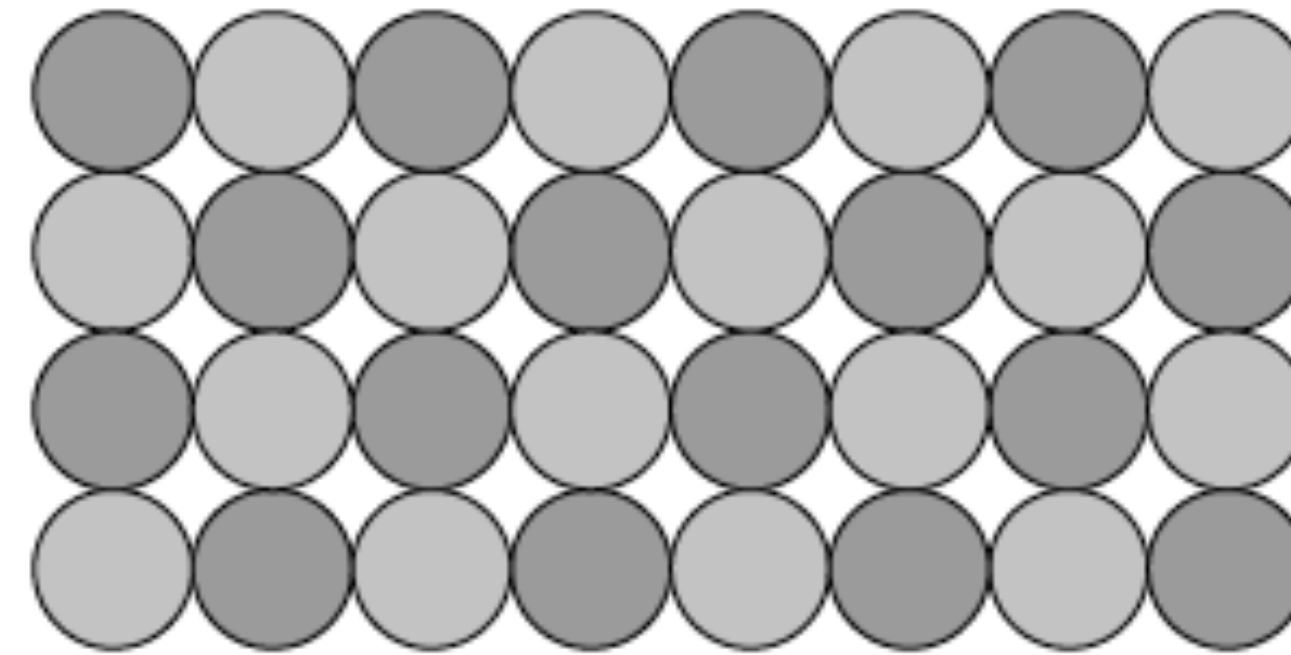


- 4 liquid molecules are converted into 29 gaseous molecules.
- Increasing the number of moles increases the number of potential arrangements.
- Changing from liquid to gas increases the number of potential arrangements.

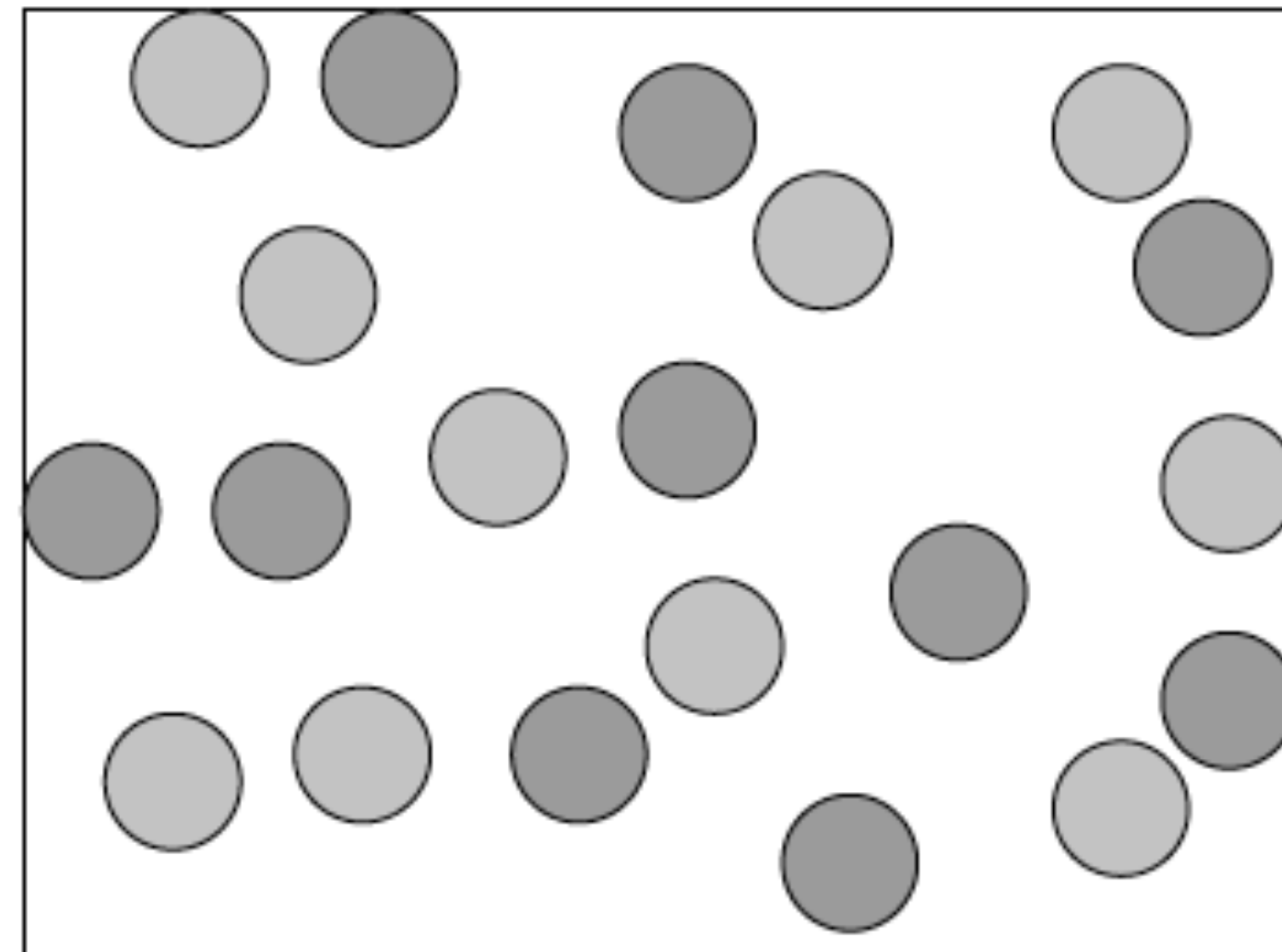
$$\Delta S > 0$$

$\Delta S > 0$ when (*usually*) when making solutions with solids and/or liquids

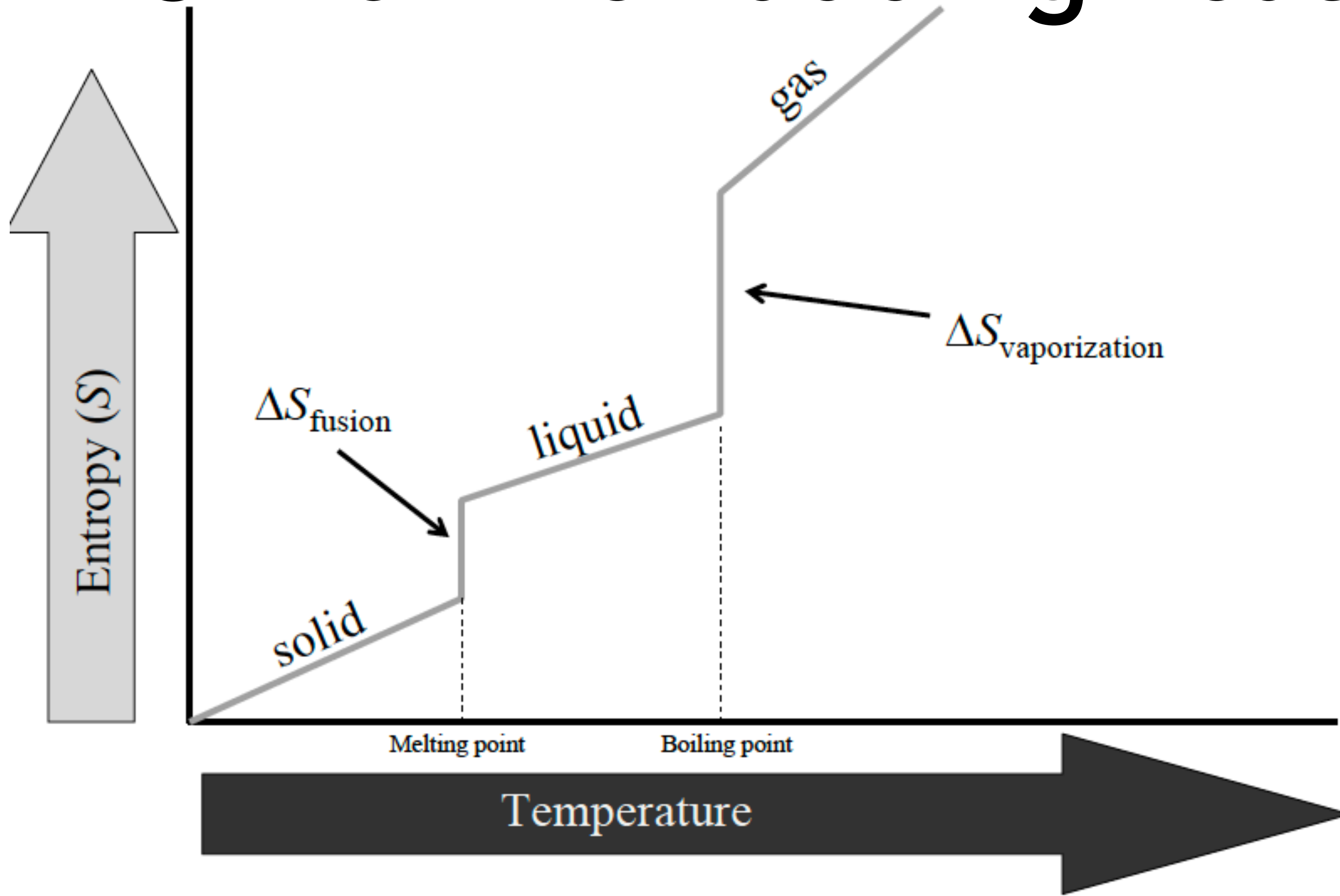
- An ionic compound is very organized.



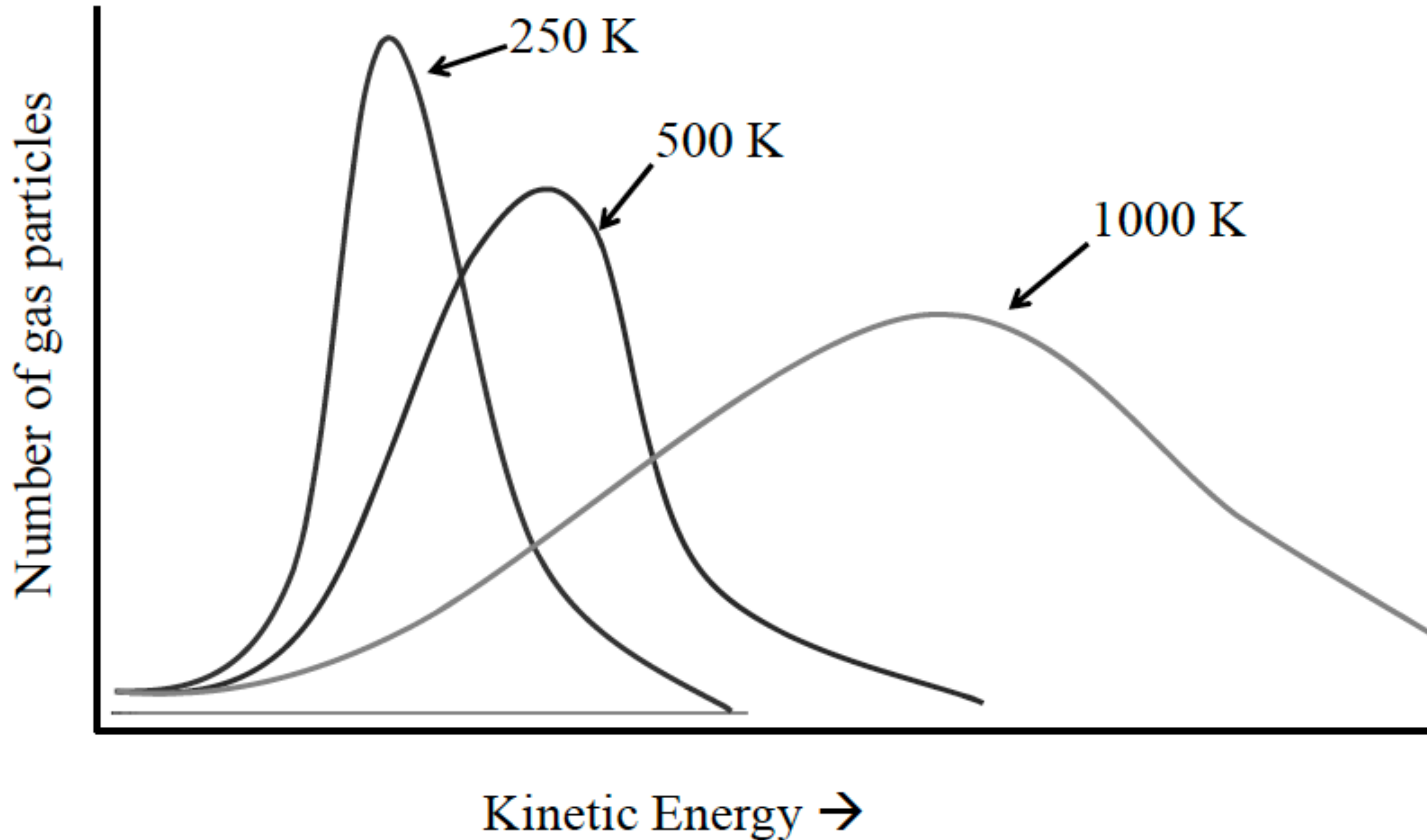
- A solution containing the same ions is much more disordered.



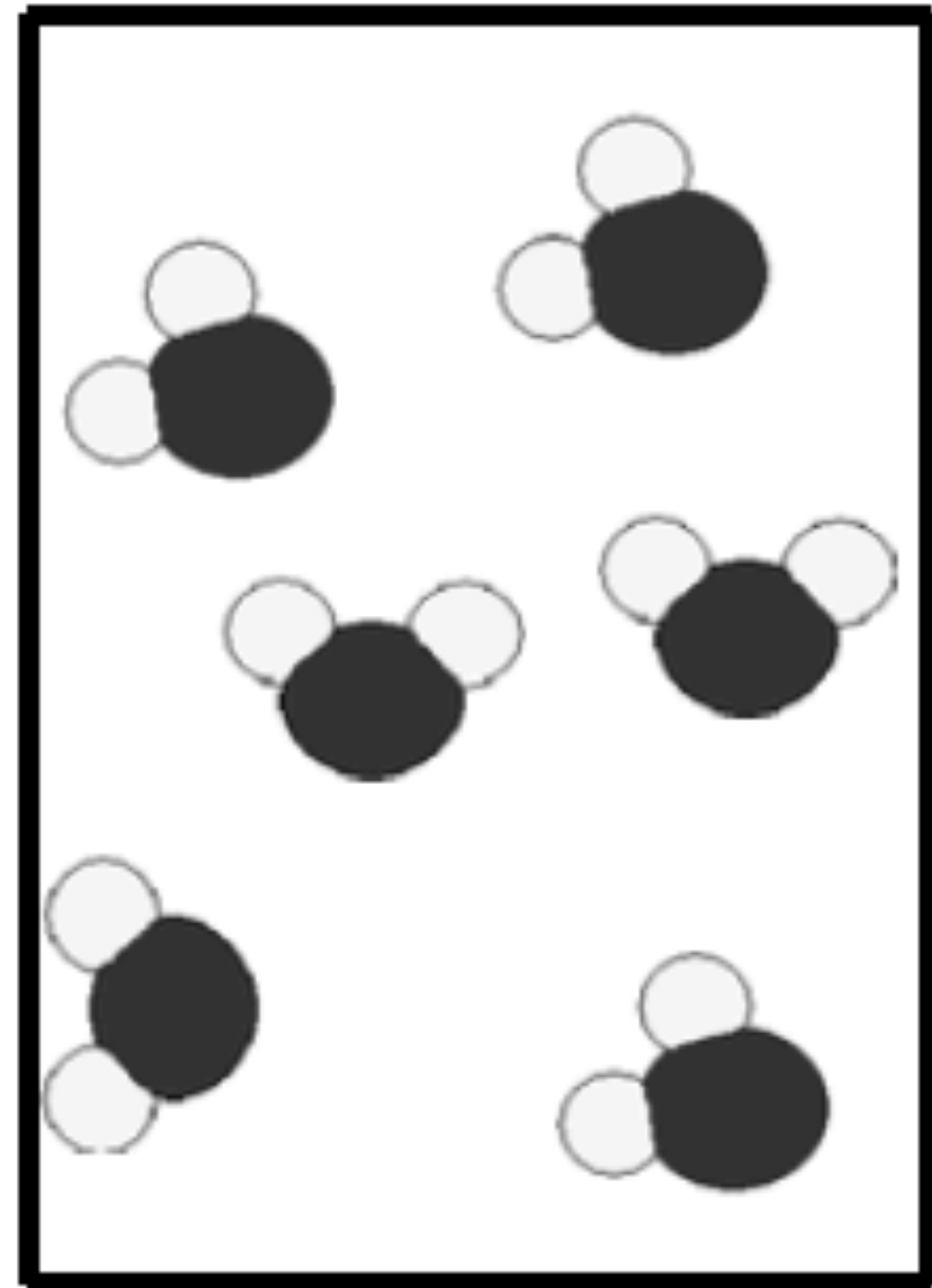
$\Delta S > 0$ when adding heat



$\Delta S > 0$ when adding heat, as the distribution of
KE increases

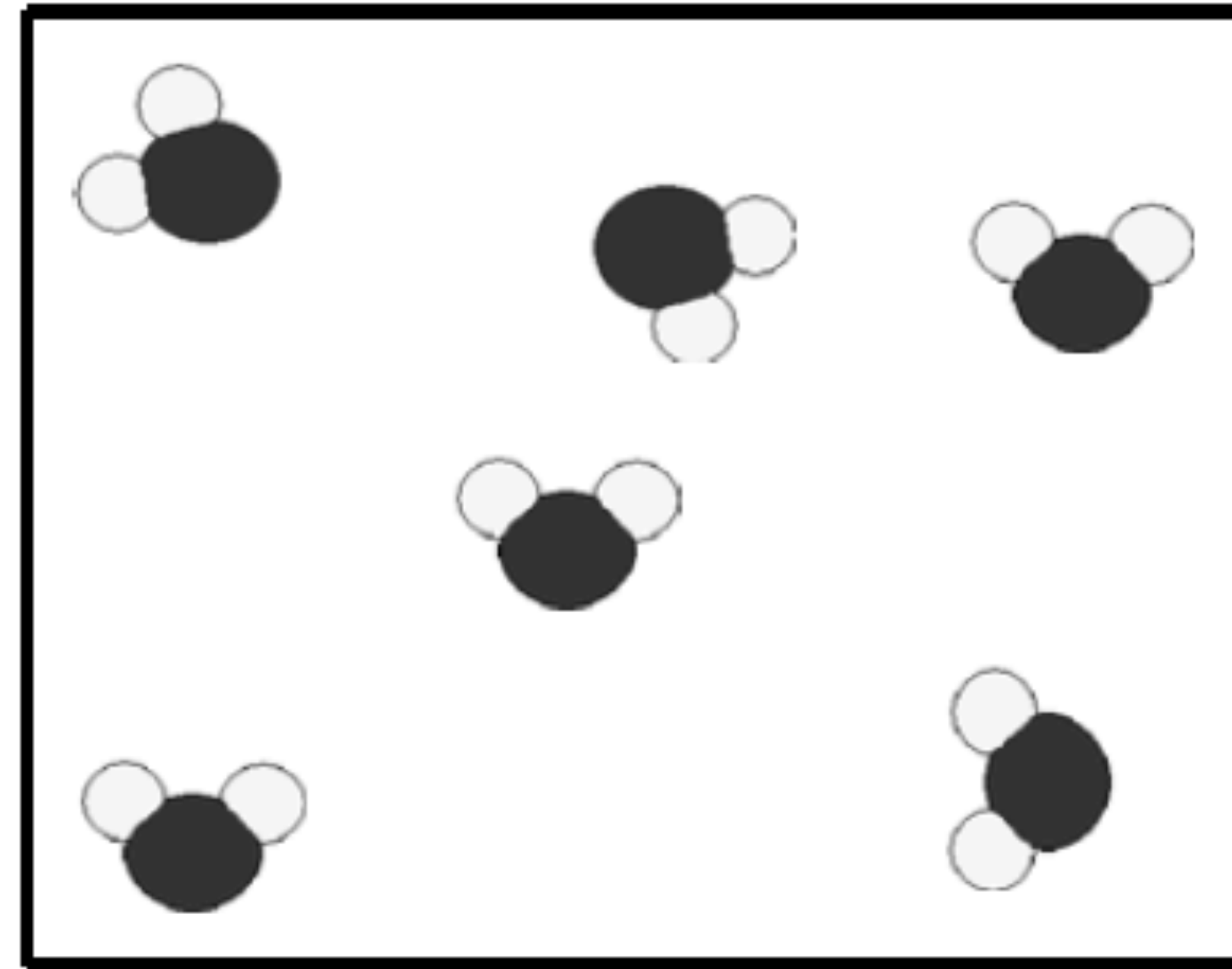


$\Delta S > 0$ when volume increases (at constant temperature)



$\text{H}_2\text{O}(g)$

High S



$\text{H}_2\text{O}(g)$

Higher S

Calculating ΔS°

$$\Delta S^{\circ}_{\text{rxn}} = \sum n S^{\circ}_{\text{(products)}} - \sum n S^{\circ}_{\text{(reactants)}}$$

Sum of

stoichiometric
coefficients

Example: Calculating ΔS

$$\Delta S_{\text{rxn}}^{\circ} = \sum n S^{\circ}_{\text{(products)}} - \sum n S^{\circ}_{\text{(reactants)}}$$

- Calculate ΔS for the following reaction at 298K.



$$\Delta S^{\circ} = \sum n S^{\circ}_{\text{(products)}} - \sum n S^{\circ}_{\text{(reactants)}}$$

$$\Delta S^{\circ} = [1(S^{\circ}(\text{N}_2\text{O}_{4(\text{g})}))] - [1(S^{\circ}(\text{N}_{2(\text{g})})) + 2(S^{\circ}(\text{O}_{2(\text{g})}))]$$

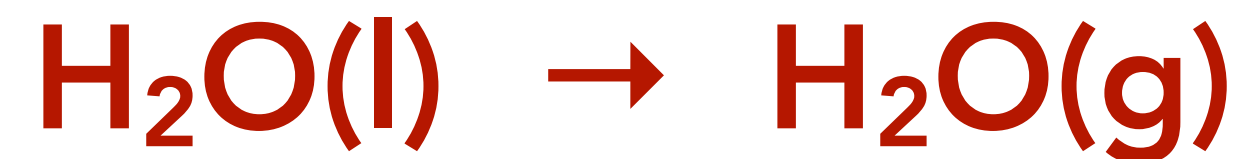
$$\Delta S^{\circ} = [1(304.2)] - [1(191.5) + 2(205.0)]$$

$$\Delta S^{\circ} = -297.3 \text{ J/K}\cdot\text{mol}$$

Example #2: Calculating ΔS

$$\Delta S_{\text{rxn}}^{\circ} = \sum n S^{\circ}_{\text{(products)}} - \sum n S^{\circ}_{\text{(reactants)}}$$

- Calculate ΔS for the following phase change at 298K.



$$\Delta S^{\circ} = \sum n S^{\circ}_{\text{(products)}} - \sum n S^{\circ}_{\text{(reactants)}}$$

$$\Delta S^{\circ} = [1(S^{\circ}(\text{H}_2\text{O}(g)))] - [1(S^{\circ}(\text{H}_2\text{O}(l)))]$$

$$\Delta S^{\circ} = [1(188.72)] - [1(69.94)]$$

$$\Delta S^{\circ} = 118.78 \text{ J/K}\cdot\text{mol}$$

Applications of Thermodynamics

9.3 Gibbs Free Energy & Thermodynamic Favorability

Entropy or Enthalpy?



Solid water

$$\Delta S^\circ = +22.0 \text{ J}/(\text{K} \cdot \text{mol})$$
$$\Delta H^\circ = +6.01 \text{ kJ}/\text{mol}$$

(Entropy increases)
(Endothermic)



Spontaneous above 0°C



Spontaneous below 0°C

$$\Delta S^\circ = -22.0 \text{ J}/(\text{K} \cdot \text{mol})$$
$$\Delta H^\circ = -6.01 \text{ kJ}/\text{mol}$$

(Entropy decreases)
(Exothermic)



Liquid water

Entropy or Enthalpy?

- A reaction is **more likely to be spontaneous** if it is **exothermic** ($-\Delta H$), but endothermic reactions can still be spontaneous.
- A reactions **more likely to be spontaneous** if there is an **increase in entropy, $+\Delta S$** , but a reaction that decreases disorder can still be spontaneous.
- Temperature is also crucial in determining whether a reaction is spontaneous.
- We need a way to incorporate all 3 terms!

Multiple Entropy Changes?

- We have 2 entropy changes, ΔS_{system} and $\Delta S_{\text{surroundings}}$, and if the **overall entropy increases** then the reaction will be spontaneous.

$$\Delta S^0_{\text{total}} = \Delta S^0_{\text{syst}} + \Delta S^0_{\text{surr}}$$

- The **second law of thermodynamics** says, in simple terms, that entropy *always* increases which we translate to be spontaneous, a reaction must lead to an increase in entropy.

Determination of ΔS

- $\Delta S_{\text{surroundings}}$ is determined by the ΔH_{system} and by **temperature.**

$$\Delta S^{\circ}_{\text{surr}} = -\frac{\Delta H^{\circ}_{\text{syst}}}{T}$$

$$\Delta S^{\circ}_{\text{total}} = \Delta S^{\circ}_{\text{syst}} - \frac{\Delta H^{\circ}_{\text{syst}}}{T}$$

- We have methods of calculating ΔS and ΔH , but we actually use a modified version of this equation which introduces **Gibbs (Free) Energy, ΔG .**

Free Energy: ΔG

- **Gibbs** energy is a useful quantity as it represents the energy of a reaction that is available 'to do work' and is given by

$$\Delta G^0 = T\Delta S^0_{\text{total}}$$

- If we multiply both sides of our **previous** equation by $-T$, we get:

$$-T\Delta S^0_{\text{total}} = \Delta H^0_{\text{syst}} - T\Delta S^0_{\text{syst}}$$

- If we combine these two equations we get Gibbs Free Energy Equation

$$\Delta G^0 = \Delta H^0_{\text{syst}} - T\Delta S^0_{\text{syst}}$$

Free Energy: ΔG

$\Delta H < 0$

$\Delta H > 0$

$\Delta S > 0$

Spontaneous at all T
($\Delta G < 0$)

Spontaneous at high T
(when $T\Delta S$ is large)

$\Delta S < 0$

Spontaneous at low T
(when $T\Delta S$ is small)

Non-spontaneous at all T
($\Delta G > 0$)

Calculating ΔG

- At Standard Conditions

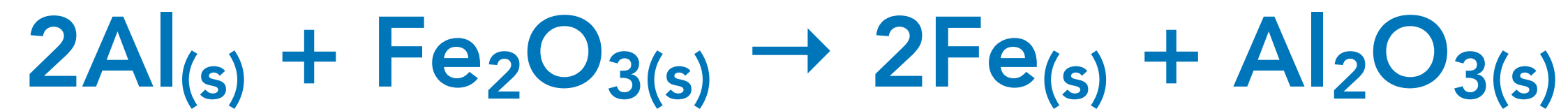
$$\Delta G_{\text{rxn}} = \sum \Delta G^{\circ}_{\text{f(products)}} - \sum \Delta G^{\circ}_{\text{f(reactants)}}$$

- At Temperatures other than 298K

$$\Delta G_{\text{syst}} = \Delta H_{\text{syst}} - T\Delta S_{\text{syst}}$$

Example

- Calculate $\Delta G^\circ_{\text{rxn}}$ for the following reaction:



COMPOUND	ΔG°_f kJ mol ⁻¹
Fe ₂ O ₃ (s)	-742
Al ₂ O ₃ (s)	-1582

Example

- Calculate the ΔG_{rxn} for the thermal decomposition of calcium carbonate @ 298K from the following data

Compound	ΔH°_f kJ mol ⁻¹	ΔS°_f J K ⁻¹ mol ⁻¹
CaCO ₃ (s)	-1207	92.9
CaO(s)	-635	39.7
CO ₂ (g)	-394	214

Hint: calculate ΔH , ΔS ... then ΔG

6.7 Bond Enthalpies

- Calculating ΔH with Average Bond Enthalpies

Bond Enthalpy

- Energy is always released during the formation of a bond.
- The same amount of energy must be added in order to break that specific bond.

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

ΔH from Bond Enthalpies

- **Exothermic Process**

- More energy is released during formation of bonds in the products than is required to break bonds in the reactants.
- The products are at a lower potential energy than the reactants.

ΔH from Bond Enthalpies

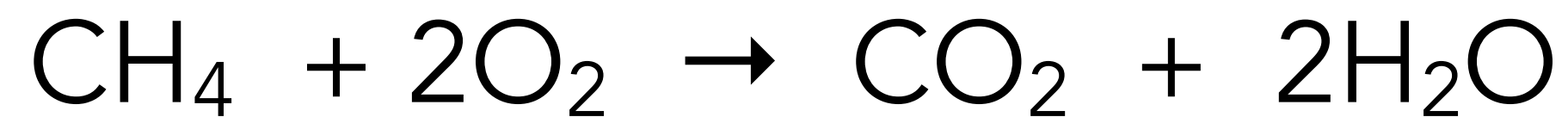
- **Endothermic Process**

- Less energy is released during formation of bonds in the products than is required to break bonds in the reactants.
- The products are at a higher potential energy than the reactants.

Example: ΔH from Bond Enthalpies

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

- Calculate the enthalpy change for the following reaction using the table of values.



Bond	Average Bond Enthalpy
C – H	413 kJ/mol
O = O	495 kJ/mol
C – O	358 kJ/mol
C = O	799 kJ/mol
O – H	467 kJ/mol

Example: ΔH from Bond Enthalpies

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

- Calculate the average bond energy in an H-H bond.



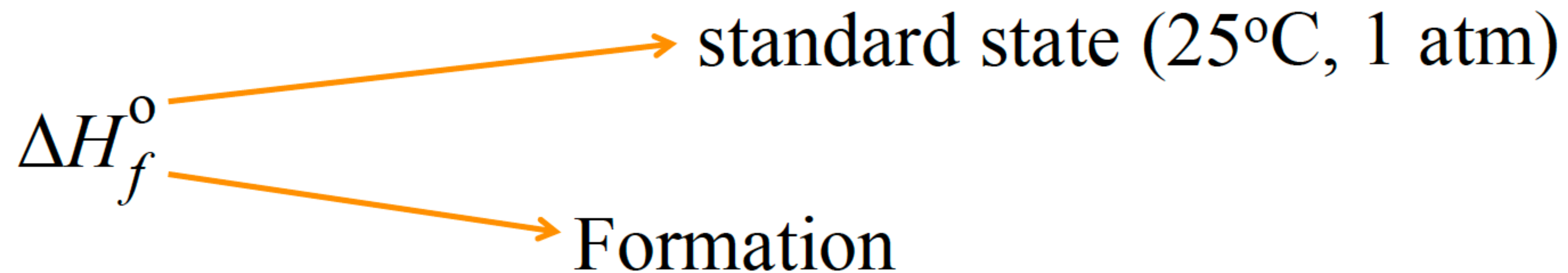
Bond	Average Bond Enthalpy
O = O	495 kJ/mol
O - H	467 kJ/mol

6.8 Enthalpy of Formation

- Calculating ΔH with Standard Enthalpies of Formation

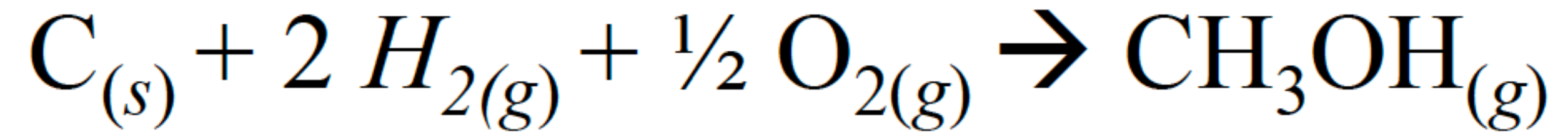
Standard Enthalpy of Formation

- A hypothetical value that indicates how much heat would be lost or gained during the formation of one mole of a compound from the most common form of its elements in their standard states.



Standard Enthalpy of Formation

- For Methanol:



$$\Delta H_f^\circ = -201 \text{ kJ/mol}$$

- Heat of formation reactions are always written so that all reactants exist as they would under standard conditions, and there is one mole of product.

Enthalpy of Formation

The ΔH_f° value for the most stable form of any element in its standard state is zero.

Element	ΔH_f°	Element	ΔH_f°
Ca(s)	0	Cl ₂ (g)	0
Ag(s)	0	H ₂ (g)	0
Na(s)	0	Fe(s)	0
O ₂ (g)	0	N ₂ (g)	0

Formula for Calculating the Enthalpy of a Reaction Under Standard Conditions

$$\Delta H_{\text{rxn}}^{\circ} = \sum n \Delta H_f^{\circ}(\text{products}) - \sum n \Delta H_f^{\circ}(\text{reactants})$$

Sum of

stoichiometric
coefficients

Example: Enthalpy of Formation

$$\Delta H_{\text{rxn}}^{\circ} = \sum n \Delta H_{\text{f}}^{\circ} (\text{products}) - \sum n \Delta H_{\text{f}}^{\circ} (\text{reactants})$$

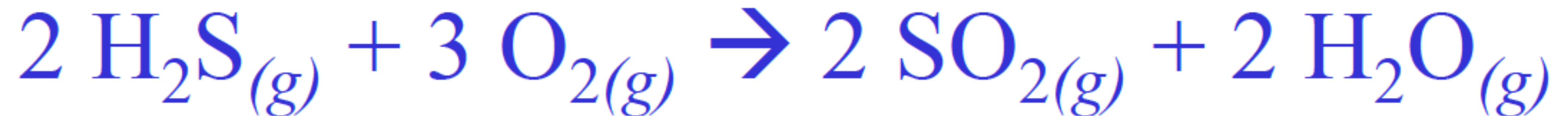
- Find $\Delta H_{\text{rxn}}^{\circ}$ for the thermite reaction under standard conditions.



Example: Enthalpy of Formation

$$\Delta H_{\text{rxn}}^{\circ} = \sum n \Delta H_{\text{f}}^{\circ} (\text{products}) - \sum n \Delta H_{\text{f}}^{\circ} (\text{reactants})$$

- Find $\Delta H_{\text{rxn}}^{\circ}$ for this reaction under standard conditions.



6.9 Hess's Law

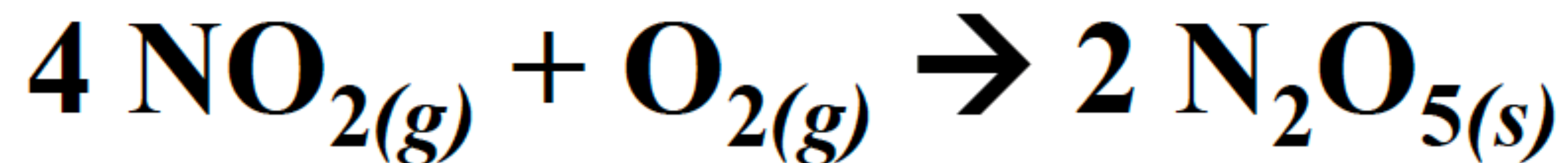
- Calculating ΔH with Hess's Law

Hess's Law

- If a reaction is carried out in a series of steps, the overall change in enthalpy will be equal to the sum of the enthalpy changes for the individual steps.
- The overall enthalpy change will be the same if a reaction is carried out in one step or in several steps.

Example: Hess's Law

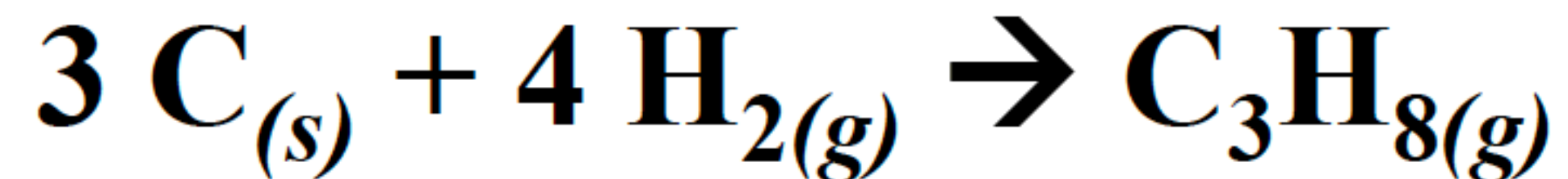
- Calculate ΔH for the following reaction using the information in the table below.



Reaction	Chemical Equation	Enthalpy Change
1	$2 \text{NO}_{(g)} + \frac{3}{2} \text{O}_{2(g)} \rightarrow \text{N}_2\text{O}_{5(s)}$	$\Delta H = - 223.7 \text{ kJ}$
2	$2 \text{NO}_{(g)} + \text{O}_{2(g)} \rightarrow 2 \text{NO}_{2(g)}$	$\Delta H = - 114.2 \text{ kJ}$

Example: Hess's Law

- Calculate ΔH for the following reaction using the information in the table below.



Reaction	Chemical Equation	Enthalpy Change
1	$2 \text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2 \text{H}_2\text{O}_{(l)}$	$\Delta H = - 571.7 \text{ kJ}$
2	$\text{C}_3\text{H}_{8(g)} + 5\text{O}_{2(g)} \rightarrow 3\text{CO}_{2(g)} + 4\text{H}_2\text{O}_{(l)}$	$\Delta H = - 2220.1 \text{ kJ}$
3	$\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$	$\Delta H = - 393.5 \text{ kJ}$

Example: Hess & Enthalpy of Formation

- Use the enthalpy of formation value below to determine ΔH_{rxn} for the reaction between $\text{C}(s)$ and $\text{Fe}_2\text{O}_3(s)$ to produce $\text{Fe}(s)$.

