

**Lesson Objectives**

By the end of this lesson, you should be able to:

- Analyze conceptually the flow of energy during changes of state (phase).
- Analyze quantitatively the flow of **energy** during changes of state (phase) using the molar enthalpies (heats) of fusion, solidification, vaporization, and condensation.

**Science Practice:** Use appropriate **scientific** tools and techniques to gather and analyze data.

**Words to Know**

Fill in this table as you work through the lesson. You may also use the glossary to help you.

<b>molar heat of vaporization</b>	the amount of heat required to vaporize one mole of a substance
<b>molar heat of fusion</b>	the amount of heat required to melt one mole of a substance

## Instruction

## Enthalpy and Phase Changes

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## Lesson Question

How can enthalpies be used to calculate the amount of energy absorbed or released during changes of state?

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## Heat and Phase Changes

- Molar heat of **fusion** ( $\Delta H_{\text{fus}}$ ):

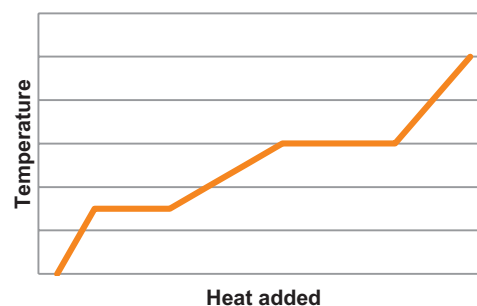
the amount of heat required to melt one mole of a substance

- Units are J/mol, **cal/mol**.
- Heat of fusion ( $\Delta H_{\text{fus}}$ ) can also be given in **J/g** or cal/g.

- Molar heat of **vaporization** ( $\Delta H_{\text{vap}}$ ): the amount of heat required to vaporize one mole of a substance

- Heat of vaporization ( $\Delta H_{\text{vap}}$ ) can also be given in J/g or cal/g.

Heating Curve Graph: Heat Added

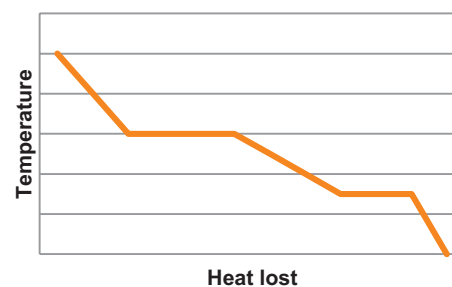


## Exothermic Phase Changes

- **Energy** lost during freezing and condensation

- Heat released during **freezing**: heat of solidification ( $-\Delta H_{\text{fus}}$ )
- Heat released during **condensation**: heat of condensation ( $-\Delta H_{\text{vap}}$ )

Heating Curve Graph: Heat Lost



## Instruction

## Enthalpy and Phase Changes

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## Effects of Heat During Phase Changes

- Balance between **intermolecular** forces (chemical potential energy) and kinetic energy determines phase.
- During endothermic phase change, **potential** energy increases.
  - **Kinetic** energy does not change, so the temperature does not change.
- During **exothermic** phase change, potential energy decreases.

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## Calculating Heat using Molar Enthalpy

The molar heat of fusion for water is 6.01 kJ/mol. How much energy must be added to a 50.0 g block of ice at 0°C to change it to 50.0 g of liquid water at 0°C?

- Write the appropriate **equation**.

$$q = n \times H_{\text{fus}}$$

- Calculate the number of **moles** of water present.

$$\begin{aligned} H_2 &= 2 \text{ g/mol} \\ O &= 16 \text{ g/mol} \\ &= \mathbf{18} \text{ g/mol} \end{aligned} \qquad 50.0 \text{ g} \times \frac{1 \text{ mol}}{18 \text{ g}} = \mathbf{2.78} \text{ mol}$$

- Calculate the heat involved.

$$2.78 \text{ mol} \times 6.01 \frac{\text{kJ}}{\text{mol}} = \mathbf{16.7} \text{ kJ}$$

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**Calculating Heat using Molar Enthalpy**

The heat of vaporization for water is 40.7 kJ/mol. How much heat is released when 491.5 g of water vapor at 100°C condenses to liquid water at 100°C?

- Write the appropriate equation.

$$q = n \times -H_{\text{vap}}$$

- Calculate the number of moles of **water** present.

$$491.5 \text{ g} \times \frac{1 \text{ mol}}{18 \text{ g}} = \boxed{27.28} \text{ mol}$$

- Calculate the heat involved.

$$27.28 \text{ mol} \times -40.7 \frac{\text{kJ}}{\text{mol}} = \boxed{-1110} \text{ kJ}$$

**Calculating Heat during a Phase Change**

- Identify the correct constant and equation to use.

- $q = n\Delta H$

- **Melting** :  $\Delta H_{\text{fus}}$

- Freezing:  $-\Delta H_{\text{fus}}$

- Evaporation or **boiling** :  $\Delta H_{\text{vap}}$

- Condensation:  $-\Delta H_{\text{vap}}$

- Perform any necessary unit conversions.
- Substitute the given values and solve the problem.

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## Enthalpy and Phase Changes

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## Relationship between Enthalpies

Substance	$\Delta H_{\text{fus}}$ (kJ/mol)	$\Delta H_{\text{vap}}$ (kJ/mol)
H <sub>2</sub> O	6.01	40.7
CO <sub>2</sub>	7.94	25.2
O <sub>2</sub>	0.443	6.81
H <sub>2</sub>	0.117	0.903
N <sub>2</sub>	0.719	5.58
HCl	1.99	16.1

Substance	$\Delta H_{\text{fus}}$ (kJ/mol)	$\Delta H_{\text{vap}}$ (kJ/mol)
Fe	14.9	354
Al	10.7	255
Cu	13.0	304
NaCl	30.2	171
CH <sub>4</sub>	0.936	8.53
H <sub>2</sub> S	2.37	18.7

The heat of **vaporization** is always greater than the heat of **fusion**, but the difference between the two varies.

There is no relationship between  $\Delta H_{\text{vap}}$  and  $\Delta H_{\text{fus}}$ . One cannot be used to predict the other.

## Sweat and Temperature Regulation

- Evaporation is **endothermic**.
- Evaporation of sweat absorbs heat from the body.

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## Enthalpy and Phase Changes

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## Putting Together Molar Enthalpy and Specific Heat

A 30.0 g sample of ice at  $-14.5^{\circ}\text{C}$  is heated to water vapor at  $120^{\circ}\text{C}$ . How much heat is required for this sample to reach its boiling point?

$$C_{p,\text{ice}} = 2.09 \text{ J/g} \cdot ^{\circ}\text{C}$$

$$C_{p,\text{water}} = \boxed{4.18} \text{ J/g} \cdot ^{\circ}\text{C}$$

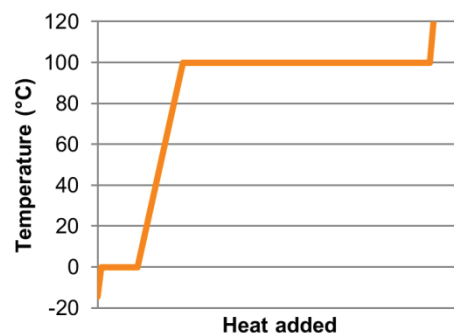
$$C_{p,\text{vapor}} = 1.84 \text{ J/g} \cdot ^{\circ}\text{C}$$

$$\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$$

$$\Delta H_{\text{vap}} = 40.7 \boxed{\text{kJ/mol}}$$

- First, the **solid** heats up, then completely melts it. The liquid heats up all the way up to its **boiling** point.

Heating Curve Graph: Heat Added



## Molar Enthalpy and Specific Heat: Heating the Solid

$$m = \boxed{30.0} \text{ g}$$

$$\Delta T = \boxed{14.5} \text{ }^{\circ}\text{C}$$

$$C_{p,\text{ice}} = 2.09 \text{ J/g} \cdot ^{\circ}\text{C}$$

- Equation:  $q = mC_p\Delta T$
- Substitute the given values and solve the problem.

$$q_1 = mC_p\Delta T$$

$$= 30.0 \text{ g} \times 2.09 \frac{\text{J}}{\text{g} \cdot ^{\circ}\text{C}} \times 14.5^{\circ}\text{C}$$

$$= 909 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$= \boxed{.909} \text{ kJ}$$

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## Enthalpy and Phase Changes

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**Molar Enthalpy and Specific Heat: Melting the Solid**

$$q_1 = 0.909 \text{ kJ}$$

$$\bullet \text{ Equation: } q = nH_{\text{fus}}$$

$$m = 30.0 \text{ g}$$

$$q_2 = nH_{\text{fus}}$$

$$\boxed{\Delta H_{\text{fus}}} = 6.01 \text{ kJ/mol}$$

$$= 1.67 \text{ mol} \times 6.01 \frac{\text{kJ}}{\text{mol}}$$

- Convert grams to moles.

$$= \boxed{10.0} \text{ kJ}$$

$$30.0 \text{ g} \times \frac{1 \text{ mol}}{18 \text{ g}} = \boxed{1.67} \text{ mol}$$

**Molar Enthalpy and Specific Heat: Heating the Liquid**

$$q_1 = 0.909 \text{ kJ}; q_2 = 10.0 \text{ kJ}$$

$$m = 30.0 \text{ g}$$

$$\Delta T = 100.0^\circ\text{C}$$

$$C_{\text{p,water}} = 4.18 \text{ J/g} \cdot ^\circ\text{C}$$

$$\text{Equation: } q = mC_p\Delta T$$

$$q_3 = mC_p\Delta T$$

$$= 30.0 \text{ g} \times 4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times 100^\circ\text{C}$$

$$= 12,540 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$= \boxed{12.5} \text{ kJ}$$

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## Enthalpy and Phase Changes

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**Molar Enthalpy and Specific Heat: Calculating the Total Heat**

$$q_1 = 0.909 \text{ kJ}; q_2 = 10.0 \text{ kJ};$$

$$q_3 = 12.5 \text{ kJ}$$

$$m = 30.0 \text{ g} \quad m = 30.0 \text{ g} \times \frac{1 \text{ mol}}{18 \text{ g}} = \boxed{1.67} \text{ mol}$$

$$\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$$

- Equation:  $\boxed{q} = nH_{\text{vap}}$

$$\begin{aligned} q_4 &= nH_{\text{vap}} \\ &= 1.67 \text{ mol} \times 40.7 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

$$= \boxed{68.0} \text{ kJ}$$

**Molar Enthalpy and Specific Heat: Calculating the Total Heat**

$$q_1 = \boxed{0.909} \text{ kJ}; q_2 = 10.0 \text{ kJ};$$

$$q_3 = 12.5 \text{ kJ}; q_4 = 68.0 \text{ kJ}$$

$$m = 30.0 \text{ g}$$

$$C_{\text{p,vapor}} = 1.84 \text{ J/g} \cdot ^\circ\text{C}$$

- Equation:  $q = mC_p\Delta T$

$$\begin{aligned} q_5 &= mC_p\Delta T \\ &= 30.0 \text{ g} \times 1.84 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times 20^\circ\text{C} \\ &= 1100 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \end{aligned}$$

$$= \boxed{1.10} \text{ kJ}$$

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## Enthalpy and Phase Changes

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**Molar Enthalpy and Specific Heat: Calculating the Total Heat**

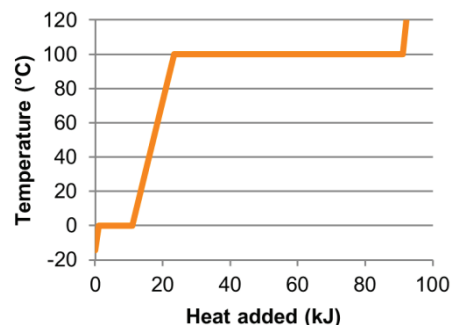
$$q_1 = 0.909 \text{ kJ}; q_2 = 10.0 \text{ kJ};$$

$$q_3 = 12.5 \text{ kJ}; q_4 = 68.0 \text{ kJ};$$

$$q_5 = 1.10 \text{ kJ}$$

- Convert joules to kilojoules.
- Add values.

$$\begin{array}{r} .909 \\ 10.0 \\ 12.5 \\ 68.0 \\ + 1.10 \\ \hline 92.5 \end{array}$$

**Heating Curve Graph: Heat Added**

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**Solving Heat Problems**

- Identify the steps of the process.
  - Heating/ cooling vs. phase change
- Identify the constant required for each step.
  - $C_p$  for various phases (heating/cooling)
  - $\Delta H_{fus}$ ,  $\Delta H_{vap}$  (phase changes)
- Write the correct equation for each step.
- Substitute the given values for each step and solve the problem.
- Add all values together to obtain total heat.

## Summary

## Enthalpy and Phase Changes

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## Lesson Question

How can enthalpies be used to calculate the amount of energy absorbed or released during changes of state?

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## Answer

(Sample answer) Enthalpy, the total heat in a system, can be calculated using the general equation  $q = n\Delta H$ . In cases where the change of state is freezing or condensation, the equation  $q = -n\Delta H$  is used. In cases where the change of state is melting, boiling or evaporation, the equation  $q = n\Delta H$  is used.

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## Heat Flow and Phase Changes

- Phase changes involve the absorption or release of heat.
  - Melting, boiling, evaporation: **endothermic**
  - Freezing, condensation: **exothermic**
- Molar heat of fusion is the heat required to melt one mole of a substance.
- Molar heat of vaporization is the heat required to boil one mole of a substance.
- Temperature does not change during a phase change because the heat absorbed or released is involved in intermolecular forces, not the kinetic energy of the particles.

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**Calculation of Heat Flow During Phase Changes**

- Heats of vaporization and fusion can be used to calculate the heat involved in a phase change.
- Heat equations include heat of fusion or heat of vaporization.
  - Melting:  $q = n\Delta H_{\text{fus}}$
  - Boiling or evaporation:  $q = n\Delta H_{\text{vap}}$
  - Freezing:  $q = -n\Delta H_{\text{fus}}$
  - Condensation:  $q = -n\Delta H_{\text{vap}}$
- Heat of fusion, heat of vaporization, and specific heat can be combined to calculate the amount of heat involved in more complex situations.

*Use this space to write any questions or thoughts about this lesson.*