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Warm-Up | Enthalpy and Phase Changes

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 _____ during changes of state (phase) using the molar enthalpies (heats) of fusion, solidification, vaporization, and condensation.

Science Practice: Use appropriate _____ 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.

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

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Lesson Question

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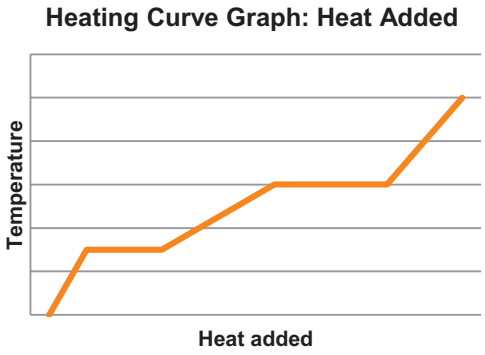
Heat and Phase Changes

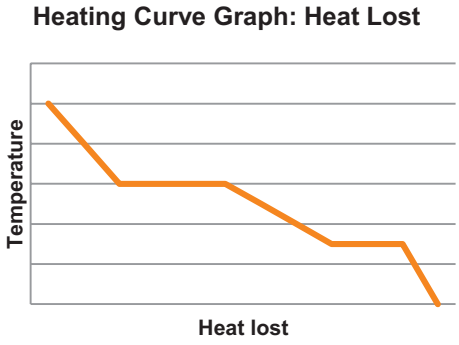
- **Molar heat of** [ ] **(ΔH_{fus})**:
  the amount of heat required to melt one mole of a substance
  - Units are J/mol, [ ].
  - Heat of fusion (ΔH_{fus}) can also be given in [ ] or cal/g.



- **Molar heat of** [ ] **(ΔH_{vap})**: the amount of heat required to vaporize one mole of a substance
  - Heat of vaporization (ΔH_{vap}) can also be given in J/g or cal/g.

Exothermic Phase Changes

- [ ] lost during freezing and condensation
  - Heat released during [ ]: heat of solidification ($-\Delta H_{fus}$)
  - Heat released during [ ]: heat of condensation ($-\Delta H_{vap}$)



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Effects of Heat During Phase Changes

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

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

- Calculate the number of _____ of water present.

$$H_2 = 2\ \text{g/mol}$$
$$O = 16\ \text{g/mol}$$
$$= \boxed{}\ \text{g/mol}$$

$$50.0\ \text{g} \times \frac{1\ \text{mol}}{18\ \text{g}} = \boxed{}\ \text{mol}$$

- Calculate the heat involved.

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

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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.

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

- Calculate the number of moles of [ ] present.

$$491.5\ \text{g} \times \frac{1\ \text{mol}}{18\ \text{g}} = [\ ]\ \text{mol}$$

- Calculate the heat involved.

$$27.28\ \text{mol} \times -40.7\frac{\text{kJ}}{\text{mol}} = [\ ]\ \text{kJ}$$

Calculating Heat during a Phase Change

- Identify the correct constant and equation to use.
  - $q = n\Delta H$
  - [ ]: ΔH_{fus}
  - Freezing: $-\Delta H_{fus}$
  - Evaporation or [ ]: ΔH_{vap}
  - Condensation: $-\Delta H_{vap}$
- Perform any necessary unit conversions.
- Substitute the given values and solve the problem.

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Relationship between Enthalpies

Substance	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)
H_2O	6.01	40.7
CO_2	7.94	25.2
O_2	0.443	6.81
H_2	0.117	0.903
N_2	0.719	5.58
HCl	1.99	16.1

Substance	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)
Fe	14.9	354
Al	10.7	255
Cu	13.0	304
NaCl	30.2	171
CH_4	0.936	8.53
H_2S	2.37	18.7

The heat of [ ] is always greater than the heat of [ ], but the difference between the two varies.

There is no relationship between ΔH_{vap} and ΔH_{fus}. One cannot be used to predict the other.

Sweat and Temperature Regulation

- Evaporation is [ ].
- Evaporation of sweat absorbs heat from the body.

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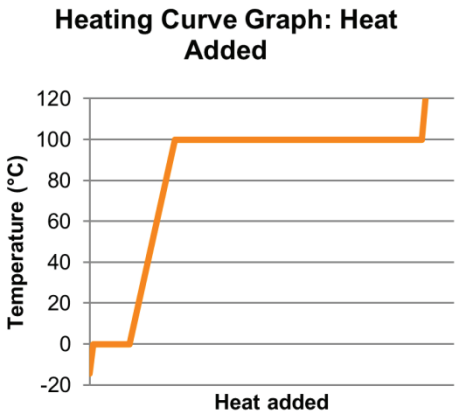
Putting Together Molar Enthalpy and Specific Heat

A 30.0 g sample of ice at –14.5°C is heated to water vapor at 120°C. How much heat is required for this sample to reach its boiling point?



$C_{p,ice} = 2.09$ J/g • °C

$C_{p,water} =$ [ ] J/g • °C

$C_{p,vapor} = 1.84$ J/g • °C

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

$\Delta H_{vap} = 40.7$ [ ]

- First, the [ ] heats up, then completely melts it. The liquid heats up all the way up to its [ ] point.

Molar Enthalpy and Specific Heat: Heating the Solid

$m =$ [ ] g

$\Delta T =$ [ ] °C

$C_{p,ice} = 2.09$ J/g • °C

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

$$q_1 = mC_p\Delta T$$

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

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

$=$ [ ] kJ

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Molar Enthalpy and Specific Heat: Melting the Solid

$q_1 = 0.909$ kJ

$m = 30.0$ g

[ ] $= 6.01$ kJ/mol

- Convert grams to moles.

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

- Equation: $q = nH_{fus}$

$$q_2 = nH_{fus}$$

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

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

Molar Enthalpy and Specific Heat: Heating the Liquid

$q_1 = 0.909$ kJ; $q_2 = 10.0$ kJ

$m = 30.0$ g

$\Delta T = 100.0$°C

$C_{p,water} = 4.18$ J/g • °C

Equation: $q = mC_p\Delta T$

$$q_3 = mC_p\Delta T$$

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

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

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

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Molar Enthalpy and Specific Heat: Calculating the Total Heat

$q_1 = 0.909$ kJ; $q_2 = 10.0$ kJ;

$q_3 = 12.5$ kJ

$m = 30.0$ g $\quad m = 30.0\text{ g} \times \dfrac{1\text{ mol}}{18\text{ g}} =$ [ ] mol

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

- Equation: [ ] $= nH_{vap}$

$$q_4 = nH_{vap}$$

$$= 1.67\text{ mol} \times 40.7\frac{\text{kJ}}{\text{mol}}$$

$$= \text{[ ]}\ \text{kJ}$$

Molar Enthalpy and Specific Heat: Calculating the Total Heat

$q_1 =$ [ ] kJ; $q_2 = 10.0$ kJ;

$q_3 = 12.5$ kJ; $q_4 = 68.0$ kJ

$m = 30.0$ g

$C_{p,vapor} = 1.84$ J/g • °C

- Equation: $q = mC_p\Delta T$

$$q_5 = mC_p\Delta T$$

$$= 30.0\text{ g} \times 1.84\frac{\text{J}}{\text{g}^\circ \bullet\ \text{C}} \times 20^\circ\text{C}$$

$$= 1100\text{ J} \times \frac{1\text{ kJ}}{1000\text{ J}}$$

$$= \text{[ ]}\ \text{kJ}$$

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Molar Enthalpy and Specific Heat: Calculating the Total Heat

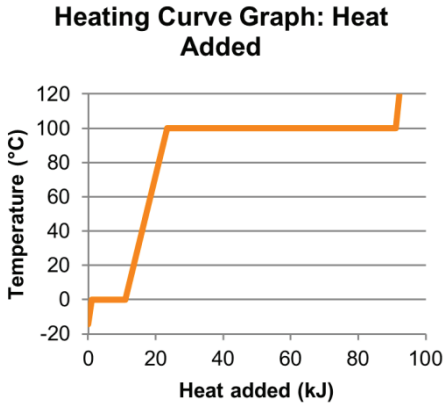
$q_1 = 0.909$ kJ; $q_2 = 10.0$ kJ;

$q_3 = 12.5$ kJ; $q_4 = 68.0$ kJ;

$q_5 = 1.10$ kJ

- Convert [ ] to kilojoules.
- Add values.

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



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Solving Heat Problems

- Identify the steps of the process.
  - Heating/[ ] vs. phase change
- Identify the constant required for each step.
  - C_p for various [ ] (heating/cooling)
  - [ ], ΔH_{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.

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Summary | Enthalpy and Phase Changes

Lesson Question

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

Answer

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Heat Flow and Phase Changes

- Phase changes involve the absorption or release of heat.
  - Melting, boiling, evaporation: [ ]
  - Freezing, condensation: [ ]
- 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: $q = n\Delta H_{fus}$
  - Boiling or evaporation: $q = n\Delta H_{vap}$
  - [ ]: $q = -n\Delta H_{fus}$
  - Condensation: $q = -n\Delta H_{vap}$
- Heat of fusion, heat of [ ], 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.

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