Heating Curve Logic: if heat is added at 10 cal/sec for 1000 sec...how many calories are added?

Mass of water	Temperature change	Calories added
100 g	100°C ÷10	1000 sec x 10 cal/sec = 10000 cal ÷10
10 0 g	10 °C ÷10	1000 cal
100 g ÷100	1 °C ÷10	100 cal
1 g ÷100	1 °C	1 cal ÷100

That means we need 1 cal **of heat to raise the temperature of 1 g of water by 1 °C!**
That is called the specific heat capacity = c = 1 cal/g°C for water.

Knowing that, how many calories do we need to raise the temperature of 10 g of water from 50°C to 75°C?

$\frac{1\ cal}{g \cdot °C} \times 10\ g \times 25\ °C = 250\ cal$ 75°C - 50°C = 25°C

Here is the formal equation for what you just did logically → $Q = mc\Delta T$

Q = heat energy (cal) m = mass (g) c = specific heat (cal/g°C) ΔT = temperature change (°C)

What about the 1st and 3rd sections of the graph? There is no temperature change...so this equation won't work...

1st section: Lasted 800 sec, if we were adding heat at 10 cal/sec → we added 8000 calories
Those 8000 calories were able to melt 100 g of solid water.
So...we need 80 calories to melt 1 g of solid water.
→ energy required to melt (s→l) 1 g of any substance

That is called the heat of fusion = H_f = 80.0 cal/g for water

$Q = mH_f$

Q = heat (cal) m = mass (g) H_f = heat of fusion (cal/g)

3rd section: Lasted 5400 sec! If we added heat at 10 cal/sec → we added 54000 calories
Those 54000 calories were able to vaporize (boil) 100 g of liquid water.
So...we need 540 calories to vaporize 1 g of liquid water (turn it all into gas).
→ energy required to vaporize (l→g) 1 g of any substance

That is called the heat of vaporization = H_v = 540 cal/g for water

$Q = mH_v$

Q = heat (cal) m = mass (g) H_v = heat of vaporization (cal/g)

What if we had a sample of water vapor at 100°C, and cooled it off?

How much heat energy do you think would need to be removed to condense one gram of water vapor at 100°C to one gram of liquid water at 100°C? Why do you think this? 539.4 cal removed
(hint: How much heat energy had to be added to turn it into a vapor? Therefore, how much heat energy would have to be removed to reverse the process?) 539.4 cal to vaporize 1 g of water

How much heat energy do you think would need to be removed to solidify one gram of liquid water at 0°C to one gram of solid ice at 0°C? Why do you think this? ~80 cal removed
(hint: How much heat energy had to be added to melt it? Therefore, how much heat energy would have to be removed to reverse the process?)

This leads us to two more equations:

$Q = mH_c$ (used when a gas **condenses** to become a liquid)

$Q = mH_s$ (used when a liquid freezes to become a solid)

ɔrie Problems

- Theoretical values for energy changes during the heating or cooling of a substance, or during a phase change, can be calculated using five basic equations.

$Q = mc\Delta T$ (used when the temperature is changing and the phase stays the same) Q = heat energy c = specific heat for liquid water = 1 cal/g °C m = mass of sample ΔT = change in temperature of sample in °C	$Q = mH_v$ (used when a liquid **vaporizes** to become a gas) $Q = mH_c$ (used when a gas **condenses** to become a liquid) m = mass of sample H_v = heat of vaporization (for water = 539.4 cal/g) H_c = heat of condensation (for water = 539.4 cal/g) negative b/c energy is removed $Q = mH_f$ (used when a solid melts to become a liquid) $Q = mH_s$ (used when a liquid freezes to become a solid) solidifies m = mass of sample H_f = heat of fusion (for water = 79.72 cal/g) H_s = heat of solidification (for water = 79.72 cal/g) energy is removed

The heat energy (Q) can be calculated in terms of calories (cal), kilocalories (Cal or kcal), or joules. (1 calorie = 4.184 J) ~~cal)~~ A calorie is defined as the amount of energy required to raise 1.0 g of water exactly 1.0 °C.

The value of Q for any substance can be calculated, but note that each substance has unique values for specific heat capacity (c), heat of fusion (H_f), and heat of vaporization (H_v).

hiah specific heat capacity (c) = a large amount of energy must be added in order to increase the temperature
ecific heat capacity (c) = a small amount of energy must be added in order to increase the temperature

Write a definition for **specific heat capacity**:

energy required to raise the temp. of 1 g of a substance by 1°C

Write a definition for **heat of fusion**:

energy it takes to change 1 g of a substance from solid → liquid

Explain the difference between **heat of fusion** and **heat of solidification**:

energy added s → l l → s energy removed

Write a definition for **heat of vaporization**:

energy needed to vaporize 1 g of a given substance

Explain the difference between **heat of vaporization** and **heat of condensation**:

adding energy l → g g → l removing energy

EXAMPLE CALORIE PROBLEMS:

For any calorie problem, you can find the requested information as long as you are given all the other information you need. Some problems require you to use algebra to rearrange the equation to solve it.

$Q = mc\Delta T$

ow much heat is required to raise the temperature of 10.0 g of water from 5.0degC to 25.0degC?

$Q = 10.0\,g \times 1\,\frac{cal}{g^\circ C} \times 20.0^\circ C = 200.\,cal$

What will be the temperature change if 100. cal of heat are added to 25 g of water?

$100.\,cal = 25\,g \times 1\,\frac{cal}{g^\circ C} \times \Delta T \qquad \Delta T = 4.0^\circ C$

$Q = mH_f$

How much heat is needed to melt 5.0 g of ice (frozen water)?

$Q = 5.0\,g \times 79.72\,\frac{cal}{g} \qquad Q = 398.6\,cal = 400\,cal$

$Q = mH_v$

How many grams of water can be vaporized by 750. calories of heat energy?

$750.\,cal = m \times 539.4\,\frac{cal}{g} \qquad m = 1.3904\,g = 1.39\,g$

$Q = mH_s$

How much heat is released when 6.2 g of liquid water freezes to become solid?

$Q = 6.2\,g \times -79.72\,\frac{cal}{g} \qquad Q = -494.264\,cal = -490\,cal$

(l → s)

(− means removed)

$Q = mH_c$

How many grams of water will condense if 375 calories of heat energy are removed from it?

$-375\,cal = m \times -539.4\,\frac{cal}{g} \qquad m = 0.695217\,g = 0.695\,g$

Calorimetry

- Physical changes are often accompanied by the transfer of energy
- To understand the transfer of energy, you must consider both the substance in question and its surroundings
  - Energy transfers that result in a temperature change for the substance and its surroundings
- Examples:
  1.
  2.
  3.
  4.

Substance and surroundings