Thermochemistry Reference Sheet

Q = mCΔT

Variable	Meaning	Common Units
Q	Energy being transferred	Joules, kJ, calorie, Calorie
m	Mass of sample	grams, kilograms
C	Specific Heat	J/g°C, kJ/g°C
ΔT	Change in Temperature	°C

Note:
- Temperature is in CELSIUS not Kelvins for this topic!
- Q can be positive or negative, energy absorbed or energy released
- 1 kJ = 1000 J
- 1 calorie = 4.18 J
- 1 Calorie = 1 kcal = 1000 calories
- Specific Heat values can vary a little bit from source to source. That is ok. Depends on the exact type that was measured. Feel free to add other specific heats you come across to this list!

Common Specific Heat Values

Substance	J/g°C	cal/g°C
Aluminum	0.90	0.23
Copper	0.39	0.093
Ethanol	2.44	0.58
Glass	0.50	0.12
Gold	0.13	0.031
Graphite	0.71	0.17
Ice	2.09	0.50
Iron	0.45	0.11
Lead	0.13	0.031
Mercury	0.14	0.033
Silver	0.24	0.057
Steam	1.87	0.45
Water (liq)	4.18	1.00
Wood	1.8	0.42

Calorimetry
- The fundamental idea behind calorimetry is that sometimes we cannot measure things directly, we need to measure them indirectly.

- If we know that the energy INTO a system equals the energy OUT of a system then we can set up an equation where we know the amount of energy being released is equal in number but opposite in sign to the energy being absorbed.

$$Q_{substance\ 1} = -Q_{substance\ 2}$$

- Remember that the negative sign just means OPPOSITE sign to the other Q value, not necessarily an actual negative value.

- Remember that T$_{final\ substance\ 1}$ and T$_{final\ substance\ 2}$ are the same! They reach the same ending temperature

For a metal object dropped into water, i.e., a typical calorimetry problem

$Q_{water} = (m_{water})(C_{water})(\Delta T_{water})$
$Q_{metal} = (m_{metal})(C_{metal})(\Delta T_{metal})$
$Q_{water} = -Q_{metal}$

Therefore… $(m_{water})(C_{water})(\Delta T_{water}) = -(m_{metal})(C_{metal})(\Delta T_{metal})$

Same as…

$(m_{water})(C_{water})(T_{final\ temp\ of\ both} - T_{initial\ water}) = -(m_{metal})(C_{metal})(T_{final\ temp\ of\ both} - T_{initial\ metal})$

Heating and Cooling Curves

This is an example of water – can be done with any substance but the temperature values will be different. Also please note that the slope and length of lines are not drawn to scale. It is traditional to just draw the heating curve. A cooling curve would just be the opposite direction!

Latent Heat Values for H₂O

$$\Delta H_{fus} = L_{fus} = 334\,\frac{J}{g};\ 6.01\,\frac{kJ}{mol}$$

$$\Delta H_{vap} = L_{vap} = 2260\,\frac{J}{g};\ 40.7\,\frac{kJ}{mol}$$

Boiling point

100°C

Freezing point

0°C

Heating Gas ↗
Cooling Gas ↙
$Q = mC_{steam}\Delta T$
5

Vaporizing →
Condensing ←
$Q = mL_{vaporization}$
4

Heating Liquid ↗
Cooling Liquid ↙
$Q = mC_{liquid}\Delta T$
3

Melting →
Freezing ←
$Q = mL_{fussion}$
2

Heating Solid ↗
Cooling Solid ↙
$Q = mC_{ice}\Delta T$
1

Cooling Curve
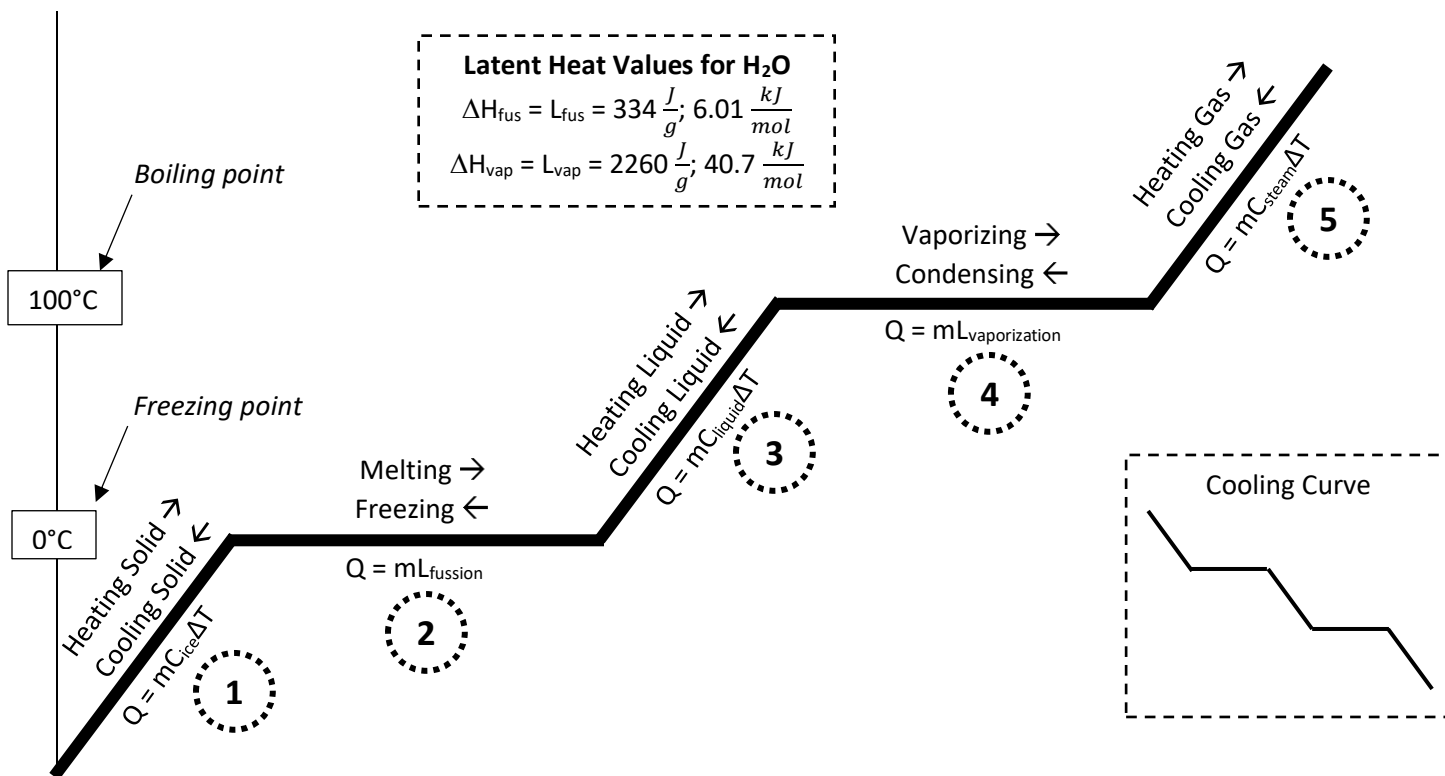
Calorimetry Problems Involving Phase Changes

If a calorimetry problem involves phase changes you need to be careful to use the right equations in your calorimetry calculations – you may not use Q = mCΔT on both sides, you may need to use Q = mL for part of it!

Example: 75g of ice (0°C) are dropped into 300g of water (25°C). What is the final temperature of the mixture when the ice is finished melting?

You would need to use Q=mL for the ice melting values and Q = mCΔT for the water values

$$mL_{fus} = -\,(mC_{liq}\Delta T)$$

$$(75g)(334\,\tfrac{J}{g}) = -(300g)(4.18\tfrac{J}{g°C})(Tf - 25°C)$$