§3.2 Thermal Properties of Matter

What happens to the internal energy of a system if you transfer internal energy into the system? (via → conduction/convection/radiation)

↑ total internal energy of the system

So why does the temperature increase?

↑ kinetic energy of the particles (translational)

Is the temperature increase the same for every substance?

depends on
- how many particles (i.e. mass)
- how many other forms of energy
- how the energy is distributed
- the substance

Specific Heat Capacity (c)

units: $J\ kg^{-1}\ °C^{-1}$

or $J\ kg^{-1}\ K^{-1}$

↑ kg is the SI unit for mass

$C = mc$

$$\Delta Q = mc\Delta T$$

Thermal Capacity (C)

units: $J\ °C^{-1}$

$J\ K^{-1}$

$$\Delta Q = C\Delta T$$

$$mc\cancel{\Delta T} = C\cancel{\Delta T}$$

$$c = \frac{C}{m}$$

For mixtures:

of two or more substances undergoing the same temperature change!

$$\Delta Q = C\Delta T = m_1c_1\Delta T + m_2c_2\Delta T$$

$$C = m_1c_1 + m_2c_2$$

Example

Calculate the energy required to raise the temperature of 250 g of copper from 20°C to 80°C

$c = 0.39\ kJ\ kg^{-1}\ K^{-1}$

$c = 390\ J\ kg^{-1}\ °C^{-1}$

$\Delta Q = mc\Delta T$

$5.8 \times 10^3\ J$ (5.8 kJ)

$5.9 \times 10^3\ J$

Example

If $3.0 \times 10^5\ J$ of heat energy is conducted into 5.0 kg water at 10°C, what is the final temperature?

$\Delta Q = mc\Delta T$

$\Delta T = \frac{\Delta Q}{mc}$

$\Delta T = \frac{3.0 \times 10^5\ \cancel{J}}{(5.0\ \cancel{kg})(4200\ \cancel{J}/\cancel{kg}\ °C)}$

$\Delta T = 14°C$

$\Delta T = T_f - T_i$

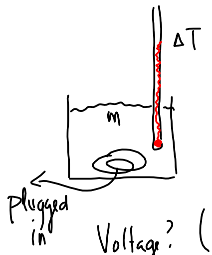
$T_f = \Delta T + T_i$

$T_f = 14°C + 10°C$

$T_f = 24°C$

Methods of Measuring Specific Heat Capacity

① Direct measurement - add a known quantity of heat + measure temperature rise.



How much energy is going in is related to the power.

$$P = \frac{\Delta Q}{\Delta t} \leftarrow \left(P = \frac{\Delta W}{\Delta t}\right)$$

Voltage? (V)
Current? (I)
Power? (P)

$$\Delta Q = P \Delta t$$

$$\Delta Q = IV \Delta t \quad (P = IV)$$

more on this later

$$\Delta Q = mc\Delta T$$

$$IV\Delta t = mc\Delta T$$

$$c = \frac{\overset{P}{IV}\Delta t}{m\Delta T}$$

Example

A small electric heater which delivers 20W of power takes 10 minutes to heat 150mL (cm^3) of water from 15°C to 34°C. Calculate the specific heat capacity of the water ($d = 1.0 \times 10^3\ kg\,m^{-3}$)

1g = 1mL

$$x\,kg = 150mL\left(\frac{1cm^3}{1mL}\right)\left(\frac{1m}{100cm}\right)^3\left(\frac{1.0\times10^3 kg}{1m^3}\right)$$

$$x\,kg = 0.150\ kg$$

$$P = \frac{\Delta Q}{\Delta t}$$

$$\Delta Q = P\Delta t$$

$$\Delta Q = 20\,Js^{-1}(600s)$$

$$\Delta Q = 12000\,J$$

$$\Delta Q = mc\Delta T$$

$$c = \frac{\Delta Q}{m\Delta T} \leftarrow ??$$

$$c = \frac{12000\,J}{(0.150kg)(19°C)}$$

$$c = 4.2\times10^3\ J\,kg^{-1}\,°C^{-1}$$

Calorimeters

- heat transfer from the "bomb" to the water
- a low thermal capacity (so that the heat it absorbs is small)
- a good insulator (minimal transfer of heat to the surroundings)

② Indirect measurement (using mixtures)

- two bodies at different temperatures are brought into "thermal contact"
- the mass of each and the final and initial temperatures are measured.
- assuming no heat is lost to surroundings =>

heat lost = heat gained.

- if we know one specific heat capacity, we can find the other.

consider taking a piece of hot metal (85°C) and placing it in cold water at 5°C → eventually, the metal and the water will be at the same temperature.

heat lost ↙ ↙ heat gained

$$\Delta Q_{metal} = -\Delta Q_{water}$$

Example

A 0.050 kg mass of brass at 100°C is plunged into 0.10 kg of water at 15°C. When the brass has come into thermal equilibrium with the cold water, the temperature is measured and found to be 18°C. Calculate the specific heat capacity of brass.

$\left(C_{H_2O} = 4.2 \times 10^3 \, J \, kg^{-1} \, °C^{-1}\right)$

$$\Delta Q_{brass} = -\Delta Q_{water}$$

$$m_b c_b \Delta T_b = -m_w c_w \Delta T_w$$

$$c_b = \frac{-m_w c_w \Delta T_w}{m_b \Delta T_b}$$

$$c_b = \frac{-(0.10kg)(4.2 \times 10^3 J \, kg^{-1} \, °C^{-1})(18°C - 15°C)}{(0.050kg)(18°C - 100°C)}$$

$$c_b \doteq 3.1 \times 10^2 \, J \, kg^{-1} \, °C^{-1}$$

Example

Determine the final temperature of the mixture when 50 g of water at 80 is poured into a calorimeter cup containing 30 g of water at 20 . The thermal capacity of the calorimeter is 60 J K^{-1} and the specific heat capacity of water is 4.2 x 10^3 J kg^{-1} K^{-1}.

$4.2\ J\ g^{-1}\ {}^\circ C^{-1}$ $\quad {}^\circ C^{-1}$ $\quad {}^\circ C^{-1}$

let x be the final temperature.

heat lost by 50g = heat gained by 30g + heat gained by calorimeter.

$$-\Delta Q_{\text{lost by water}} = \Delta Q_{\text{water}} + \Delta Q_{\text{calorimeter}}$$

$\Delta Q = mc\Delta T$

$\Delta Q = C\Delta T$

$$-(50g)(4.2\,J\,g^{-1}\,{}^\circ C^{-1})(x - 80^\circ C) = (30g)(4.2\,J\,g^{-1}\,{}^\circ C^{-1})(x - 20^\circ C) + (60\,J\,{}^\circ C^{-1})(x - 20^\circ C)$$

HINT! Your answer must be between 20°C and 80°C

~~54°C~~ ✓51.8° (52°C) ← $\frac{20500}{396}$

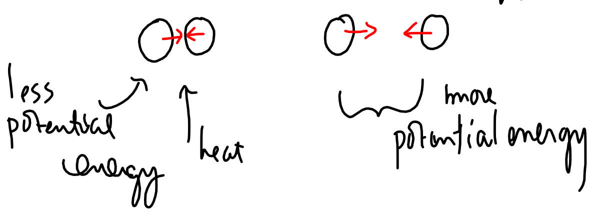
✓51.8° (52°C)

$$\Delta Q_{\text{lost}} = -\Delta Q_{\text{gained}}.$$

Melting + Freezing

Consider ice below freezing + gradually add thermal energy

- temperature increases (increasing the kinetic energy of the particles)
- reach melting pt, then the particles can no longer "stick" together → ice melts
- all energy goes into melting. (increasing potential energy due to bonding)

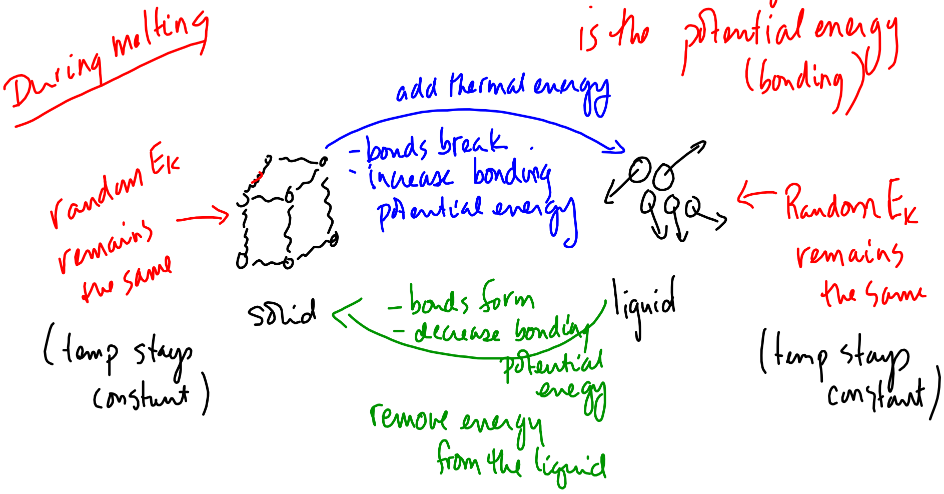


- opposite occurs during freezing (thermal energy is released potential energy decreases)

- no temperature change during a phase change

↓

the average random kinetic energy remains same → all that changes is the potential energy (bonding)



Boiling + Condensing .... same line of thought as for melting / freezing.

Read over 3.2.1 to 3.2.6