Specific heat
Phase change & latent heat
Evaporation
Relative humidity
What is heat?

Heat (denoted Q) is a flow of energy between 2 objects as a result of temperature difference between them.

The direction of heat flow is from the object at higher temperature to the object at lower temperature.

This flow will reduce the higher temperature and increase the lower temperature until the two objects have a common temperature $T$.

If two objects are at the same temperature, no heat energy is transferred between them when they are brought into contact.
Specific heat

To increase the temperature of an object
• increase the thermal energy of the molecules

Quantity of heat \( Q \) (thermal energy) needed to change the temperature depends on:

1. temperature change \( \Delta T \)
2. mass \( m \)
3. specific heat of substance

Specific Heat
• Depends on the nature of the substance
Specific heat

A specific amount of energy is required by the same mass of every substance to change its temperature by 1\(^o\)C

**Definition:**

If a quantity of energy \( Q \) transferred to an object of mass \( m \), changes its temperature by an amount \( \Delta T \) the specific heat \( c \) of the object is defined as:

\[
c = \frac{Q}{m \cdot \Delta T}
\]

\[
Q = c \cdot m \cdot \Delta T
\]

**SI Units**

- \( Q \text{—(Joules)} \)
- \( c \text{—(Joules.kg}^{-1}.k}^{-1}) \)
- \( m \text{—(kg)} \)
- \( \Delta T \text{—(k)} \)

**Units**

- \( Q \text{—(kcal)} \)
- \( c \text{—(kcal.kg}^{-1}.{\degree}C}^{-1}) \)
- \( m \text{—(kg)} \)
- \( \Delta T \text{—(°C)} \)
The specific heat \( c \) varies strongly from substance to substance:

**Air:** \( c = 0.25 \text{ kcal.kg}^{-1}.\text{°C}^{-1} = 1047 \text{ J.kg}^{-1}.\text{°C}^{-1} \)

**Water:** \( c = 1.00 \text{ kcal.kg}^{-1}.\text{°C}^{-1} = 4186 \text{ J.kg}^{-1}.\text{°C}^{-1} \)

**Lead:** \( c = 0.03 \text{ kcal.kg}^{-1}.\text{°C}^{-1} = 130 \text{ J.kg}^{-1}.\text{°C}^{-1} \)

Water: Large specific heat

>> moderate temperatures near large quantities of water
Specific heats

Dental Applications

Important during melting and casting.

Determines the quantity of heat required to reach melting point

<table>
<thead>
<tr>
<th>Specific Heat (J.kg(^{-1}.^\circ C(^{-1}))</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Enamel</td>
<td>712</td>
</tr>
<tr>
<td>Dentine</td>
<td>1591</td>
</tr>
<tr>
<td>Amalgam</td>
<td>206</td>
</tr>
<tr>
<td>Gold</td>
<td>130</td>
</tr>
<tr>
<td>Human Body</td>
<td>3500</td>
</tr>
</tbody>
</table>

Relatively low specific heat value for amalgam
- little thermal energy required to change its temperature

Drinking iced water → thermal shock
Specific Heats

Specific Heat depends on the structure of the substance

**Same** substances in **different phases** have different specific heats.

Example

Water: \( c = 1.00 \text{ kcal.kg}^{-1}.^{\circ}\text{C}^{-1} = 4186 \text{ J.kg}^{-1}.^{\circ}\text{C}^{-1} \)

Ice: \( c = 0.50 \text{ kcal.kg}^{-1}.^{\circ}\text{C}^{-1} = 2093 \text{ J.kg}^{-1}.^{\circ}\text{C}^{-1} \)

Steam: \( c = 0.48 \text{ kcal.kg}^{-1}.^{\circ}\text{C}^{-1} = 2009 \text{ J.kg}^{-1}.^{\circ}\text{C}^{-1} \)
How much thermal energy (in Joules) is needed to increase, by 5°C, the temperature of
(a) 200g of water \((c = 1.00 \text{ kcal.kg}^{-1}\text{.°C}^{-1})\)
(b) 200g of lead \((c = 0.03 \text{ kcal.kg}^{-1}\text{.°C}^{-1})\)

water: \[ Q = c \times m \times \Delta T \]
\[
Q = (1 \times 0.2 \times 5) \text{ kcal} = 1 \text{ kcal}
\]
Converting to SI units 1 kcal = 4186 Joules

lead: \[ Q = c \times m \times \Delta T \]
\[
Q = (0.03 \times 0.2 \times 5) \text{ kcal} = 0.03 \text{ kcal}
\]
in SI units = 125.6 J

Much less heat is required to increase the temperature of lead, than for the same mass of water.
Specific heat

650g of water at 33 °C are poured into an aluminium pot of mass 250g at 15 °C. If no heat is lost to the surroundings, find the final temperature of the system.

\[ c_A = 900 \text{ J.kg}^{-1}.\text{oC}^{-1} \]

\[ Q_w = c_w m_w (\Delta T_w) = c_w m_w (T_f - T_w) \]

\[ Q_A = c_A m_A (\Delta T_A) = c_A m_A (T_f - T_A) \]

no heat is lost to the surroundings

Conservation of energy

heat lost by the water = Heat gained by pot

\[-Q_w = Q_A \]

\[ Q_w + Q_A = 0 \]

\[ c_w m_w (T_f - T_w) + c_A m_A (T_f - T_A) = 0 \]

\[ 4186 \times 0.65 (T_f - 33) + 900 \times 0.25 \times (T_f - 15) = 0 \]

\[ T_f = 31.6 \text{ °C} \]
Other causes of temperature change

Heat is not the only cause of temperature change

Hands (both at same temperature)
Rub together to increase temperature

Work done by friction goes into thermal energy in the hands

**Thermal energy** (or temperature) of an object
• may be increased by adding heat.
• **may be increased by mechanical means**

In general
**loss of mechanical energy** related to **friction**
appears as **thermal energy**
(increased temperature)
During this phase change:
- temperature remains constant
- heat is required

This heat, called **latent heat**, is needed to **break the bonds** between atoms or molecules, as a substance **changes** from **solid to liquid** or from **liquid to gas**.

- Breaking of the bonds
There are 2 distinct phase changes:

1. Solid-liquid, concerns the “latent heat of fusion” (denoted $L_f$): it is the heat required to change a certain mass from solid to liquid at the same temperature:
   \[ Q = mL_f \]
   \[ L_f = Q/m \]

2. Liquid-gas, concerns the “latent heat of vaporization” (denoted $L_v$): it is the heat required to change a certain mass from liquid to gas at the same temperature:
   \[ Q = mL_v \]

**Units**

- $Q$: Joules
- $m$: kg
- $L_f$: Joules.kg$^{-1}$
- $L_v$: Joules.kg$^{-1}$

**SI**

**Units**

- $Q$: kcal
- $m$: kg
- $L_f$: kcal.kg$^{-1}$
Latent heat

Energy input but no increase in temperature

Energy required to overcome attractive forces between molecules

Typical values
Heat of vaporisation and fusion

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting Point (°C)</th>
<th>Boiling Point (°C)</th>
<th>L_\text{v} kJ/kg</th>
<th>L_\text{f} kJ/kg</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>0</td>
<td>100</td>
<td>2260</td>
<td>335</td>
</tr>
<tr>
<td>Copper</td>
<td>1083</td>
<td>2595</td>
<td>4810</td>
<td>205</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>-210</td>
<td>-196</td>
<td>200</td>
<td>25.5</td>
</tr>
</tbody>
</table>
Latent Heat

Reversing direction, i.e. steam-to-water and water-to-ice results in a release of this heat.

Energy must be removed from substance

Example
How much heat is released from 50g of water as it
(a) Changes from liquid to ice at 0°C.
(b) Changes from steam to liquid water at 100°C.

Water \( L_v = 2260\text{kJ/kg} \quad L_f = 335\text{kJ/kg} \)

**Liquid to ice**

\[ Q = mL_f = 0.05\text{kg} \times 335 \text{kJ/kg} = 16,750\text{J} \]

**Steam to Liquid**

\[ Q = mL_v = 0.05\text{kg} \times 2260\text{kJ/kg} = 113,000\text{J} \]

Steam to water transition liberates more heat
than water-ice transition.
Ice \(<0^\circ C\) receives heat at a constant rate:

1/ Temperature increase  2/ Phase changes

$$Q = c_i m \Delta T$$

$$Q = mL_f$$

$$Q = c_w m \Delta T$$

$$Q = mL_v$$

$$Q = c_s m \Delta T$$

Heat added
Latent heats are large quantities:

For the same mass (1kg), show that the same thermal energy is needed to melt ice and to increase the temperature of water by 80°C.

Melt ice

\[ Q = mL_f \]

\[ Q = (1\text{kg})(3.35 \times 10^5\text{Jkg}^{-1}) = 3.35 \times 10^5\text{J} \]

Raise water temperature

\[ Q = c \times m \times \Delta T \]

\[ Q = 4186\text{ J.kg}^{-1}.^\circ\text{C}^{-1}(1\text{kg})(80^\circ\text{C}) = 3.35 \times 10^5\text{J} \]
Evaporation

Water slowly evaporates from a container

Water will evaporate from pool after rainfall

This evaporation occurs at temperatures lower than boiling point of water. Why?

Molecules in a liquid are held tightly together by strong forces.

Molecules have a range of kinetic energies. Some molecules will have large kinetic energy----far above average.

Molecules with enough energy can escape from the surface and change from the liquid state to the gaseous state----evaporation

Energetic molecules escape from liquid hence average temperature of molecules left behind decreases as evaporation proceeds.
Humidity

As evaporation progresses the surrounding air contains an increasing amount of water vapour and begins to feel humid.

Humidity effects the net evaporation rate of water

Higher the humidity – lower the evaporation rate

The capacity of air to hold water vapour increases with air temperature

vapour density is a measure of the amount of water vapour in the air.

Air can become saturated with water vapour. It will then hold no more water unless the temperature is increased
Relative humidity

The amount of vapour in a saturated air is called the saturation vapour density.

This value increases strongly with temperature (hot air can contain much more vapour than cold air).

When air is **not saturated**, its actual vapour density is lower than the saturation vapour density.

The **relative humidity** is defined as the ratio of the actual vapour density to the saturation vapour density, expressed as a percentage:

\[
\text{relative humidity} = \frac{\text{actual vapor density}}{\text{saturated vapor density}} \times 100
\]

Relative humidity of 100% --- air saturated and can hold no more water
Relative humidity

Relative humidity affects the net evaporation rate

Relative humidity : 100%
Water vapour condenses out of the air at the same rate as evaporation causes it to enter

Relative humidity depends on temperature
Higher the air temperature the more vapour it can hold without it condensing out.

Saturation density of water vapour in air

<table>
<thead>
<tr>
<th>Temp $^0$C</th>
<th>Density $\text{kg/m}^3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>-10</td>
<td>2.36 x $10^{-3}$</td>
</tr>
<tr>
<td>0</td>
<td>4.85 x $10^{-3}$</td>
</tr>
<tr>
<td>15</td>
<td>12.83 x $10^{-3}$</td>
</tr>
<tr>
<td>20</td>
<td>17.30 x $10^{-3}$</td>
</tr>
</tbody>
</table>
Relative humidity

Relative humidity 40%--50%..... feels comfortable

**Low relative humidity** dries out skin and mucous membranes, may cause sinus and other respiratory problems.

**High relative humidity** suppresses the evaporation rate of perspiration; may be uncomfortable even at moderate temperatures.
The relative humidity of the air in a house is 100%. The temperature drops from 20°C during the day down to 15°C during the night. If the volume of the house is 500m³, what quantity of water will condense out?

Saturation vapour density at 20°C = 17.30g.m⁻³
Saturation vapour density at 15°C = 12.83g.m⁻³

The actual vapour density is equal to 17.30g.m⁻³ as the relative humidity is 100%. As the temperature decreases, the relative humidity cannot increase, therefore vapour will thus condense out ensuring the humidity remains at 100%, i.e. leaving an actual vapour density of 12.83g.m⁻³.

Initial vapour quantity
= (17.30g.m⁻³*500m³) = 8.65kg

Final vapour quantity
= (12.83g.m⁻³*500m³) = 6.41kg

Water condensed = 8.65-6.41kg = 2.24kg = 2.24l
Example

How much heat is needed to boil away 2kg of water at 10°C contained in an aluminum pot of mass 200g?

- \(c_{water} = 1.00 \text{ kcal.kg}^{-1}.{\circ}C^{-1}\),
- \(c_{aluminium} = 0.217 \text{ kcal.kg}^{-1}.{\circ}C^{-1}\),
- heat of vaporization of water = 540kcal.kg\(^{-1}\)

Heat > pot, water (10°C > 100°C), boil off water.

1/ Heating of the pot:

\[
Q = c \cdot m \cdot \Delta T \\
Q = (0.217 \times 0.2 \times 90) \text{ kcal} = 3.9\text{kcal}
\]

2/ Heating of the water:

\[
Q = c \cdot m \cdot \Delta T \\
Q = (1.00 \times 2 \times 90) \text{ kcal} = 180\text{kcal}
\]

3/ Vaporization of the water:

\[
Q = m \cdot L_v \\
Q = 2 \times 540 \text{ kcal} = 1080\text{kcal}
\]

Total heat required = (180 + 1080 + 3.9) kcal

= 1263.9kcal