PHYS1001 Physics 1 REGULAR
Module 2 Thermal Physics

Chapter 15 Temperature and Heat

SPECIFIC HEAT CAPACITY

PHASE CHANGES

CALORIMETRY

References: 15.6  15.7
Examples 15.7  15.8  15.9  15.10  15.11  15.12

Checklist

• Distinction between temperature \( T \) (K or °C) and heat \( Q \) (joule J)

• Change in temperature of object of mass \( m \) kg

\[
\Delta T = \frac{Q}{mc} \quad \text{assume no change in phase} \quad \text{(Eq. 15.13)}
\]

• Specific heat capacity \( c \) – characteristic of substance

\[
c = \frac{1}{m} \frac{dQ}{dT} \quad \text{(Eq. 15.14)}
\]

Amount of energy transferred to change the temperature by 1 K or 1 °C for a 1 kg object. Assume \( c \) independent of temperature and measured at constant pressure and there is no change in phase.

• Water has a very large specific heat capacity compared to other substances

\[c_{\text{water}} = 4190 \text{ J.kg}^{-1}\text{.K}^{-1}\]

• number of moles \( n \) (mol)

mass of sample \( m \) (kg)

molar mass \( M \) (kg.mol\(^{-1}\) or g.mol\(^{-1}\)) mass of 1 mol of sample

\[
m = nM \quad M = \frac{m}{n} \quad n = \frac{m}{M} \quad \text{(Eq. 15.16)}
\]
**molar heat capacity** \( C \) (J.mol\(^{-1}\).K\(^{-1}\))

\[
Q = n \ C \ \Delta T \quad \text{(Eq. 15.18)}
\]

\[
C = \frac{1}{n} \frac{dQ}{dT} = Mc \quad \text{(Eq. 15.19)}
\]

Amount of energy transferred to change the temperature by 1 K (1 °C) for 1 mol of a substance.

molar heat capacity measured at constant pressure \( C_p \)

molar heat capacity measured at constant volume \( C_V \)

- Dulong and Petit's rule
  
  solid elements \( C \sim 25 \text{ J.mol}^{-1}.\text{K}^{-1} \)

  For solid elements, the energy transferred to change the temperature depends on the number of atoms and not upon the mass of each individual atom.

- **Phase changes**

  \[
  \begin{align*}
  Q &= \pm m \ L_f \\
  &\text{At melting point: } L_f \text{ latent heat of fusion or heat of fusion}
  \\
  Q &= \pm m \ L_v \\
  &\text{At boiling point point: } L_v \text{ latent heat of vaporization or heat of vaporization}
  \end{align*}
  \]

  \( Q > 0 \) energy *absorbed* by substance during phase change

  \( Q < 0 \) energy *released* by substance during phase change

- **Heat (calorimetry) calculations are based upon the principle of conservation of energy**
Notes

Energy

Mechanical energy: kinetic and potential
Thermal energy: internal energy, $\Sigma(KE + PE)$
Chemical energy: chemical bonds
Electrical energy: electrical potential

When a change takes place, there is a transfer or flow of energy.

Heat (energy transferred because of a temperature difference) and work (energy transferred by a force acting over a distance) are words that refer to the amount of energy in transit or transferred.

Specific heat capacity $c$ (J.kg$^{-1}$.K$^{-1}$) – no change in phase

$$\Delta T = T_f - T_i$$

$T_i$ \hspace{2cm} $\Delta T$ \hspace{2cm} $T_f$

Mass of object $m$  \hspace{1cm} Specific heat capacity $c$  \hspace{1cm} $Q$

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

Specific heat capacity, $c$, amount of heat required to change the temp. of 1 kg by 1 °C

Problem

If 2000 J are absorbed by copper plate, aluminium plate, glass plate and water, each of mass 250 g, which would be "hottest"?

$c_{Cu} = 390$ J.kg$^{-1}$.K$^{-1}$  \hspace{1cm} $c_{Al} = 910$ J.kg$^{-1}$.K$^{-1}$

$c_{glass} = 390$ J.kg$^{-1}$.K$^{-1}$  \hspace{1cm} $c_{water} = 4190$ J.kg$^{-1}$.K$^{-1}$
Solution

Setup

\[ Q = 2000 \text{ J} \]
\[ m_{Cu} = m_{Al} = m_{glass} = m_{water} = m = 250 \text{ g} = 0.250 \text{ kg} \]

Specific heat capacity \[ \Delta T = \frac{Q}{mc} \]

Action

- copper \[ \Delta T = \frac{2000}{(0.250)(390)} \text{ °C} = 21 \text{ °C} \]
- aluminium \[ \Delta T = \frac{2000}{(0.250)(910)} \text{ °C} = 8.8 \text{ °C} \]
- glass \[ \Delta T = \frac{2000}{(0.250)(840)} \text{ °C} = 10 \text{ °C} \]
- water \[ \Delta T = \frac{2000}{(0.250)(4190)} \text{ °C} = 1.9 \text{ °C} \]

Notice the small change in the temperature of the water. Why is a copper saucepan better than an aluminium one?

Water

- The large heat capacity of water makes it a good temperature regulator, since a great amount of energy is transferred for a given change in temperature.
- Why is there a bigger difference between the max and min daily temperatures at Campbelltown rather than at Bondi?
- Why is water a good substance to use in a hot water bottle?
- Why is the high water content of our bodies \((c \approx 3500 \text{ J.kg}^{-1}.\text{K}^{-1})\) important in relation to the maintenance of a constant core body temperature?

Problem

For a 70 kg person (specific heat capacity 3500 J.kg\(^{-1}\).K\(^{-1}\)), how much extra released energy would be required to raise the temperature from 37 °C to 40 °C?

Solution

Setup

\[ m = 70 \text{ kg} \quad c = 3500 \text{ J.kg}^{-1}.\text{K}^{-1} \quad \Delta T = (40 - 37) \text{ °C} = 3 \text{ °C} \]

Specific heat capacity \[ Q = mc \Delta T \]

Action

\[ Q = m c \Delta T = (70)(3500)(3) = 7.45 \times 10^5 \text{ J} = 0.74 \text{ MJ} \]
A fever represents a large amount of extra energy released. The metabolic rate depends to a large extent on the temperature of the body. The rate of chemical reactions are very sensitive to temperature and even a small increase in the body's core temperature can increase the metabolic rate quite significantly. If there is an increase of about 1 °C then the metabolic rate can increase by as much as 10%. Therefore, an increase in core temperature of 3% can produce a 30% increase in metabolic rate. If the body's temperature drops by 3 °C the metabolic rate and oxygen consumption decrease by about 30%. This is why animals hibernating have a low body temperature. During heart operations, the person's temperature maybe lowered.

How do we measure a person's metabolic rate?

Phase changes

- As a liquid evaporates it extracts energy from its surroundings and hence the surroundings are cooled.

- When a gas condenses energy is released into the surroundings. Steam heating systems are used in buildings. A boiler produces steam and energy is given out
as the steam condenses in radiators located in rooms of the building. Boiling-condensing processes are used in refrigerators and air conditions.

- **Water** has large values of latent heats at atmospheric pressure, melting point $T_m = 0 \, ^\circ\text{C}$ & boiling point $T_b = 100 \, ^\circ\text{C}$.
  
  \[
  L_f = 3.34 \times 10^5 \, \text{J.kg}^{-1} \quad L_v = 2.26 \times 10^6 \, \text{J.kg}^{-1}
  \]

  You can be badly scolded by steam – more dangerous than an equivalent amount of boiling water.

  It takes about the same amount of energy to melt 1.00 kg of ice at 0 °C as it does to heat 1.00 kg of water by 80 °C (eg 10 °C to 90 °C)

  Melting $Q = m L = (1)(3.34 \times 10^5) \, \text{J} = 3.34 \times 10^5 \, \text{J}$

  Increase in temp $Q = m \cdot c \Delta T = (1)(4190)(80) \, \text{J} = 3.35 \times 10^5 \, \text{J}$

- **Evaporation** rates depend upon temperature, area, humidity and liquid.

  Evaporation rates increase with temperature, volatility of substance, area and lower humidity. You feel uncomfortable on hot humid days because perspiration on the skin surface does not evaporate and the body can't cool itself effectively. The circulation of air from a fan pushes water molecules away from the skin more rapidly helping evaporation and hence cooling. Evaporative cooling is used to cool buildings. Why do dogs pant? When ether is placed on the skin it evaporates so quickly that the skin feels frozen. Ethyl chloride when sprayed on the skin evaporates so rapidly the skin is "frozen" and local surgery can be performed.

- **Melting** and **boiling points** depend upon external pressure exerted upon substance. Higher altitude lower air pressure, water boils at a lower temperature.

- **Sublimation** – food preservation (frozen food placed into a low-pressure chamber and heated, ice sublimes and vapour drawn off. The drying process does not damage the food. Food reconstituted by adding water).
• **Supercooled water** – liquid or vapour below freezing point (melting point). Supercooled vapour can form fog droplets by condensing on dust particles. If there was no pollution, water vapour could not condense to form rain droplets. This principle is used in seeding cloud.

• **Superheated liquid** – liquid above its boiling point. A small disturbance eg ionised particle can cause boiling (bubble formation). Principle used in bubble chambers to detect radioactive particles.
Problem
The energy released when water condenses during a thunderstorm can be very large.
Calculate the energy released into the atmosphere for a typical small storm.

Solution
Setup
Assume 10 mm of rain falls over a circular area of radius 1 km

\[ h = 10 \text{ mm} = 10^{-2} \text{ m} \quad r = 1 \text{ km} = 10^{3} \text{ m} \]

volume of water \( V = \pi r^2 h = \pi (10^6)(10^{-2}) = 3 \times 10^4 \text{ m}^3 \)

mass of water \( m = ? \text{ kg} \quad \text{density of water } \rho = 10^3 \text{ kg.m}^{-3} \)

\[ m = \rho V = (10^3)(3 \times 10^4) = 3 \times 10^7 \text{ kg} \]

*latents heats – change of phase*

\[ Q = mL \quad L_v = 2.26 \times 10^6 \text{ J.kg}^{-1} \]

Action

\[ Q = m L_v = (3 \times 10^7)(2.26 \times 10^6) = 7 \times 10^{13} \text{ J} \]

The energy released into the atmosphere by condensation for a small thunder storm is more than 10 times greater then the energy released by one of the atomic bombs dropped on Japan in WW2. This calculation gives an indication of the enormous energy transformations that occur in atmospheric processes.
Problem
How much ice at –10.0 °C must be added to 4.00 kg of water at 20.0 °C to cause the resulting mixture to reach thermal equilibrium at 5.0 °C. Assume no energy transfer to the surrounding environment, so that energy transfer occurs only between the water and ice.

Solution

Setup

\[ m_{\text{ice}} = ? \text{ kg} \quad m_{\text{water}} = 4.00 \text{ kg} \]

temperature rise for ice to melt \[ \Delta T_{\text{ice1}} = 0 - (-10) \text{ °C} = 10 \text{ °C} \]

temperature rise of melted ice \[ \Delta T_{\text{ice2}} = (5 - 0) \text{ °C} = 5 \text{ °C} \]

temperature fall for water \[ \Delta T = (20 - 5) \text{ °C} = 15 \text{ °C} \]

\[ c_{\text{ice}} = 2000 \text{ J.kg}^{-1}.\text{K}^{-1} \quad c_{\text{water}} = 4190 \text{ J.kg}^{-1}.\text{K}^{-1} \]

\[ L_f = 3.33 \times 10^5 \text{ J.kg}^{-1} \]

Conservation of energy \[ \text{Energy lost by water} = \text{Energy gained by ice} \]

Specific heat capacity \[ Q = m \cdot c \cdot \Delta T \]

Phase change \[ Q = m \cdot L \]

Action

energy lost water (fall in temperature) = 
energy gained by ice (rise in temp ice + melting + rise in temp water)

\[ m_{\text{water}} \cdot c_{\text{water}} \cdot \Delta T_{\text{water}} = m_{\text{ice}} \cdot c_{\text{ice}} \cdot \Delta T_{\text{ice1}} + m_{\text{ice}} \cdot L_f + m_{\text{ice}} \cdot c_{\text{water}} \cdot \Delta T_{\text{ice2}} \]

\[ m_{\text{ice}} = \frac{m_{\text{water}} \cdot c_{\text{water}} \cdot \Delta T_{\text{water}}}{c_{\text{ice}} \cdot \Delta T_{\text{ice1}} + L_f + c_{\text{water}} \cdot \Delta T_{\text{ice2}}} \]

\[ m_{\text{ice}} = \frac{(4)(4190)(15)}{(2000)(10) + (3.33 \times 10^5) + (4190)(5)} = 0.67 \text{ kg} \]

In such problems you have to be careful to include all changes in temperature and changes in phase.