THERMAL PHYSICS.p4

HEAT CAPACITY AND LATENT HEAT

Objectives

The net amount of heat exchanged is proportional to change in temperature.

The heat capacity of a body, $Q = C\,\theta$

Define the specific heat capacity of a substance. $Q=mc\,\theta$

Define Molar heat capacities of a gas at constant pressure and at constant volume

The methods by which heat is lost and methods to reduce the heat lost and how - a correction for heat loss is obtained by changing the initial temperature.

The difference between the cooling of a hot body under natural convection and - forced convection.

Correction for heat loss by changing the initial temperature.

Newton's law of cooling and limits.

Determination of specific heat capacities of solids and liquids by the method of - mixtures

Comparison of specific heat capacities of liquids by the method of cooling.

Latent heat and phase changes

Heat Capacity

Consider a large and small pan.

- The pans are of a similar type but different size. Each pan is filled with water. They are placed on heaters having the *same power*. In which pan would the water boil first? Obviously the smaller one.
- The larger pan of water needs a greater quantity of energy to cause its temperature to change by a given amount. We say that the larger pan has a greater *heat capacity* than the smaller one.

Which has the larger heat capacity?

- ___ water in the pail
- _X_ water in the swimming pool

Why do the metal parts of the car gets really hot while the plastic and other materials stay at more bearable temperature?

-

Heat Capacity

- The heat capacity of an object depends on the
  - __Mass__ of the object: An object with a larger mass will have a __larger__ heat capacity than an object with smaller mass of the same material
  - __Type__ of material: Different materials have different heat capacities.

- The heat capacity of a body depends on
- i)what **substance(s)** it is made of
- ii) the **masses** of the different substances in the body

HEAT CAPACITY

- The heat capacity of a body is the **amount of heat** that must be supplied to the body to **increase its temperature by 1°C.**

- The units for heat capacity are **J°C^{-1} or JK^{-1}**

$$\Delta Q = C\Delta\theta$$

ΔQ is the heat energy added to the body
Δθ is the temperature rise of the body
C is the heat capacity of the body

The Specific Heat Capacity *of a Substance* (c)

Considering again the two pans of water. Suppose that the small pan holds 1kg of water and the larger one holds 3kg of water. It is reasonable to expect that to change the temperature of the 3kg of water, by a given amount, will require *three times* as much energy as to change the temperature of the 1kg of water. We are assuming that 1kg of water always needs the *same quantity* of energy to change its temperature by a given amount.

Comparition of specific heat capacities of different materials

- To change the temperature of a body means to change the **average kinetic energy of its particles.**
- The particles of different substances have different masses. The **number of particles in 1kg** of a substance depends on the mass of those particles. This explains why different substances have different specific heat capacities.
- For example, the mass of an atom of iron is about *twice* the mass of an atom of aluminum. So, 1kg of aluminum must contain about *twice as many atoms* as 1kg of iron. We would therefore expect the specific heat capacity of aluminum to be about twice that of iron.
- (c_{iron} = 460Jkg-1°C-1 $c_{aluminum}$ = 908 Jkg-1°C-1)

Specific heat capacity

- The amount of heat that must be supplied to the substance to increase the temperature by 1 °C for a mass of 1 kg of the substance.

Specific heat capacity

$$c = \frac{Q}{m\,\theta}$$

Symbol	Physical Quantity	Unit
Q	Heat supplied or released	Joule, J
m	Mass of substance	kg
θ	Temperature difference	°C
c	Specific heat capacity	J kg^{-1} °C^{-1}

The units of specific heat capacity are **Jkg^{-1}°C^{-1} or Jkg^{-1}K^{-1}**.

$$Q = m\,c\,\Delta T$$

Heat energy in J

Mass in kg

Specific heat capacity in J/K/kg

Temperature change in K

Exercise 1

- What does it mean by specific heat capacity of aluminium is 900 $Jkg^{-1}\,°C^{-1}$?

- 900 J of heat needs to be supplied to 1 kg of aluminium to produce a 1 °C temperature increase.

Exercise 1, #2

- What does it mean by specific heat capacity of water is 4 200 $Jkg^{-1}\,°C^{-1}$?

- 4 200 J of heat needs to be supplied to 1 kg of water to produce a 1 °C temperature increase.

Exercise 1, #3

A metal of mass 2 kg. Calculate the amount of heat that must be transferred to the metal to raise the temperature from 30 °C to 70 °C. (specific capacity of the metal = 500 $J\,kg^{-1}\,°C^{-1}$)

- $Q = mc\theta = (2)(500)(70 - 30) = 40{,}000$ J

Exercise 1, #4

- A lady takes a watermelon and sandwich from the fridge and leaves them outside. After sometimes she touches the watermelon and sandwich, she feels the watermelon is cooler than the sandwich. Why does the watermelon stay cool for a longer time than the sandwich even though both are taken from the same fridge?

- The water in the melon has a higher specific heat capacity than the sandwich. It requires more heat to raise its temperature and takes a longer time to turn warm.

A gram of water requires 1 calorie of energy to raise the temperature 1°C. • It takes only about one eighth as much energy to raise the temperature of a gram of iron by the same amount.

- Absorbed energy can affect substances in different ways.

- Absorbed energy that increases the translational speed of molecules is responsible for increases in temperature.

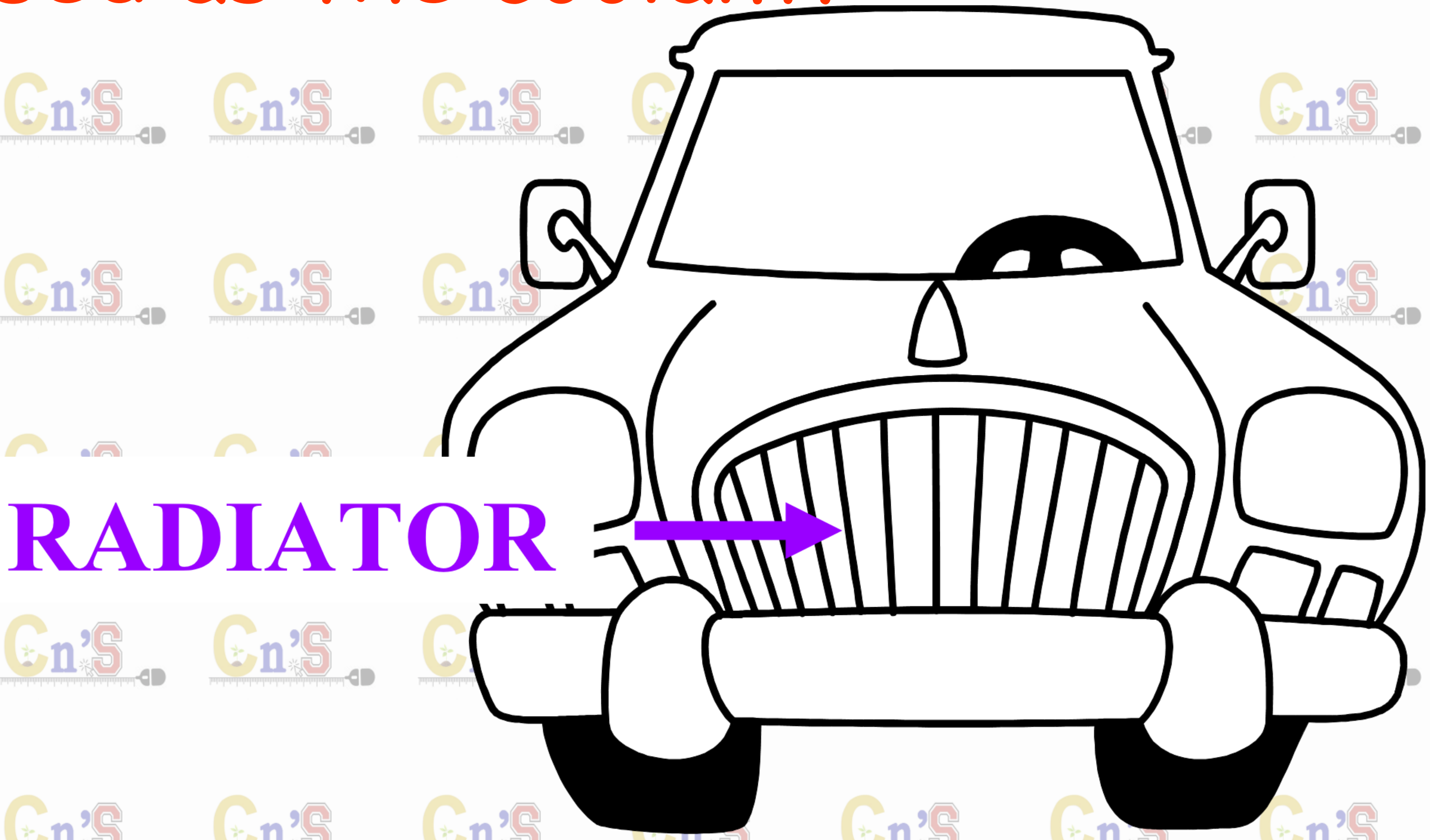
- Temperature is a measure only of the kinetic energy of translational motion.

- Absorbed energy may also increase the rotation of molecules, increase the internal vibrations within molecules, or stretch intermolecular bonds and be stored as potential energy.

- Iron atoms in the iron lattice primarily shake back and forth, while water molecules soak up a lot of energy in rotations, internal vibrations, and bond stretching. • Water absorbs more heat per gram than iron for the same change in temperature. • Water has a higher specific heat capacity (sometimes simply called specific heat) than iron has.

- Which has a higher specific heat capacity—water or sand? Explain.
- Answer: Water has a greater heat capacity than sand. Water is much slower to warm in the hot sun and slower to cool at night. Sand's low heat capacity, shown by how quickly it warms in the morning and how quickly it cools at night, affects local climates.

Why is water used as the coolant?



Because water has a high specific heat capacity, it can take away a lot of energy without boiling away.

The High Specific Heat Capacity of Water

- The property of water to resist changes in temperature improves the climate in many places.
- Water has a much higher capacity for storing energy than most common materials.
- A relatively small amount of water absorbs a great deal of heat for a correspondingly small temperature rise.
- Because of this, water is a very useful cooling agent, and is used in cooling systems in automobiles and other engines.
- For a liquid of lower specific heat capacity, temperature would rise higher for a comparable absorption of heat.
- Water also takes longer to cool.

Substance	c (J.kg^{-1}.K^{-1})
Aluminum	910
Copper	390
Ice	2100
Water	**4190**
Steam	2010
Air	1000
Soils / sand	~500

Specific Heat Capacities *of Substances*

Substance	Specific Heat, c: SI Units: J/kg•K	Specific Heat, c: cal/g•°C, kcal/kg•°C, or Btu/lb•°F
Aluminum	900	0.215
Concrete	880	0.24
Copper	386	0.0923
Iron	447	0.107
Glass	753	0.18
Mercury	140	0.033
Steel	502	0.12
Stone (granite)	840	0.20
Water:		
Liquid	4184	1.00
Ice, −10°C	2050	0.49
Wood	1400	0.33

* Temperature range 0°C to 100°C except as noted.

Insulators and Conductors

Substances with a low specific heat make good thermal conductors.

- Substances with a high specific heat make good thermal insulators.

Table 2 Common Insulators

Good insulators	
oil	plastic
fur	wood
silk	paper
wool	wax
rubber	ebonite
porcelain, glass	pure water

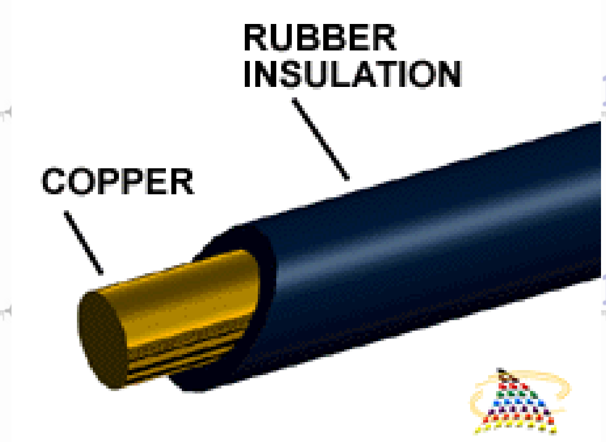
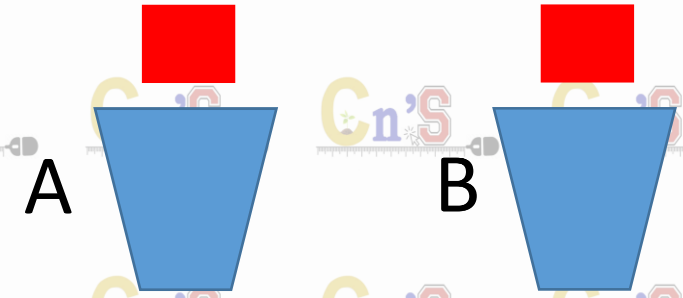


Table 1 Common Conductors

Good conductors	Fair conductors
silver	graphite (carbon)
copper	nichrome
gold	the human body
aluminum	damp skin
magnesium	acid solutions
tungsten	salt water
nickel	Earth
mercury	water vapour in air
platinum	silicon
iron	germanium

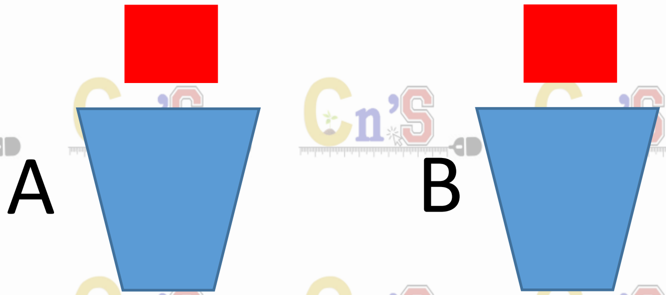
Two identical mass metals at 95°C are placed in separate identical beakers of water at 25°C. You measure the temperature of the water after each metal has cooled by 10°C and find that the water in A is hotter than the water in B. Which metal has the higher specific heat ?



0%

1. Metal A
2. Metal B
3. They are the same
4. Can't tell since not in equilibrium

Two identical mass metals at 95°C are placed in separate identical beakers of water at 25°C. You measure the temperature of the water after each metal has cooled by 10°C and find that the water in A is hotter than the water in B. Which metal has the higher specific heat ?



Heat energy lost by metal = heat energy gained by water

$$Metal\ A : m_{metal}\, c_A\, \Delta T_{metal} = m_{water}\, c_{water}\, \Delta T_A$$

$$Metal\ B : m_{metal}\, c_B\, \Delta T_{metal} = m_{water}\, c_{water}\, \Delta T_B$$

$$\Delta T_A > \Delta T_B \rightarrow c_A > c_B$$

(a) How much heat does it take to bring a 3.5 kg iron frypan from 20℃ to 120℃ ? (b) If a $2kW$ stovetop heats the pan, how long will this take? ($c_{iron} = 447\ J\ kg^{-1}K^{-1}$)

(a) $Q = m\,c\,\Delta T$

$$m = 3.5\ kg$$
$$c = 447\ J\ kg^{-1}K^{-1}$$
$$\Delta T = 100\ K$$

$$Q = 3.5 \times 447 \times 100 = 0.16\ MJ$$

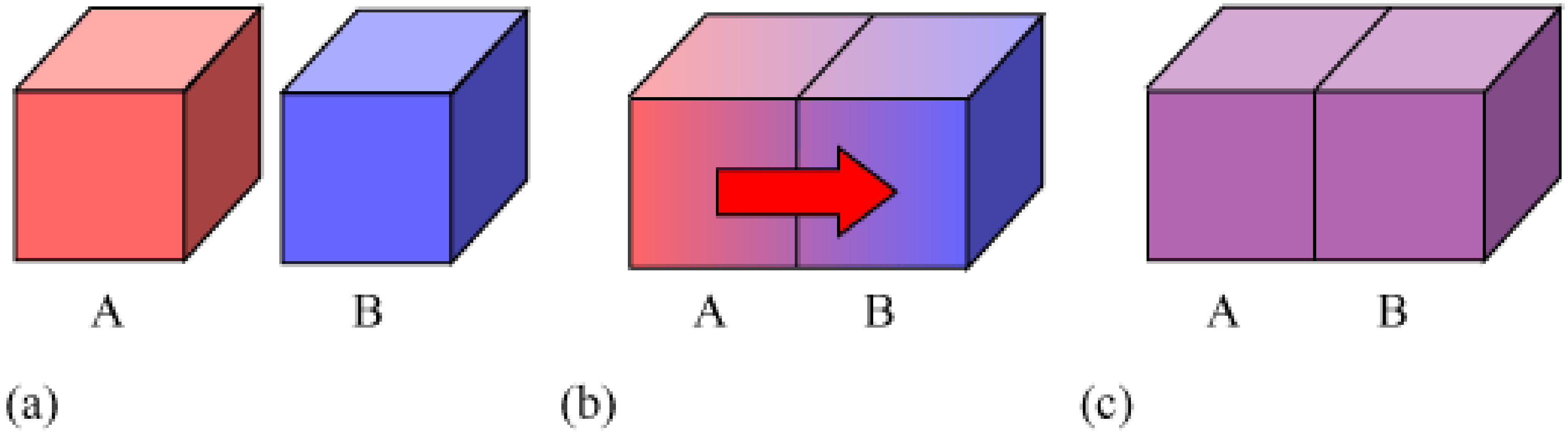
(b) $Power = Energy/Time$

$$Time = Energy/Power$$

$$Energy = 0.16\ MJ = 1.6 \times 10^5\ J$$
$$Power = 2\ kW = 2 \times 10^3 W$$

$$t = \frac{1.6 \times 10^5}{2 \times 10^3} = 78\ s$$

(c) The same 3.5 kg iron frypan at 120℃ is plunged into a sink filled with 2 $litres$ of water at 20℃. What is the equilibrium temperature? ($c_{iron} = 447\, J\, kg^{-1} K^{-1}$, $c_{water} = 4184\, J\, kg^{-1} K^{-1}$)



- Heat energy flows until equilibrium is reached
- Heat energy lost by frypan = Heat energy gained by water
- Total heat energy change = 0 **(conservation of energy)**

(c) The same 3.5 kg iron frypan at 120°C is plunged into a sink filled with 2 $litres$ of water at 20°C. What is the equilibrium temperature? ($c_{iron} = 447\ J\ kg^{-1}K^{-1}, c_{water} = 4184\ J\ kg^{-1}K^{-1}$)

$$Q = mc\Delta T = mc(T_f - T_i) \quad and \quad \sum Q = 0$$

$$m_1c_1(T_{eq} - T_1) + m_2c_2(T_{eq} - T_2) = 0$$

$$T_{eq}(m_1c_1 + m_2c_2) = m_1c_1T_1 + m_2c_2T_2$$

$$T_{eq} = \frac{m_1c_1T_1 + m_2c_2T_2}{m_1c_1 + m_2c_2}$$

$$T_{eq} = \frac{3.5 \times 447 \times 393 + 2 \times 4184 \times 293}{3.5 \times 447 + 2 \times 4184}$$

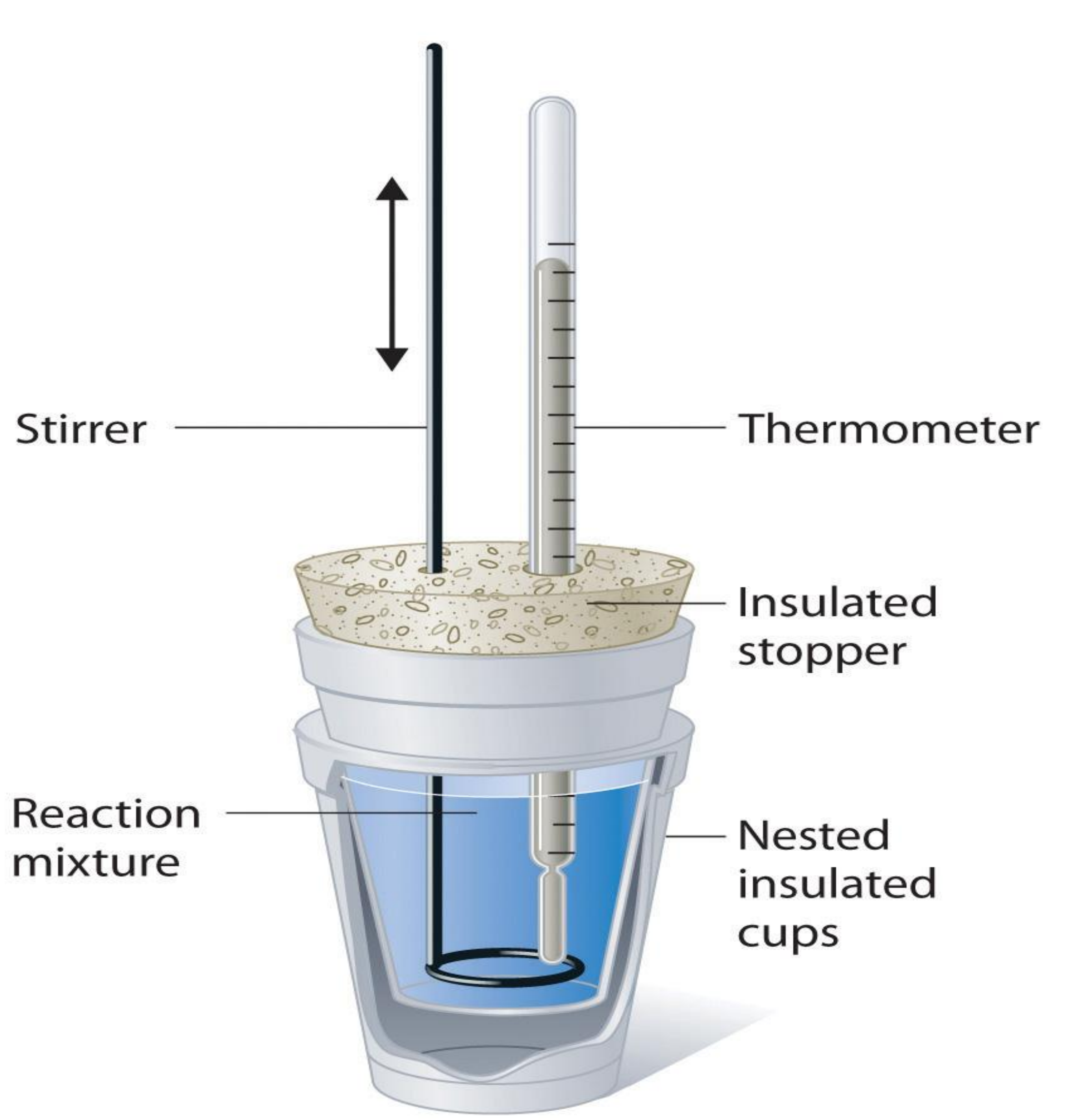
$$T_{eq} = 308.75°K = 36°C$$

CONVERSION OF ELECTRICAL ENERGY INTO THERMAL ENERGY

$$Pt = mc\theta$$

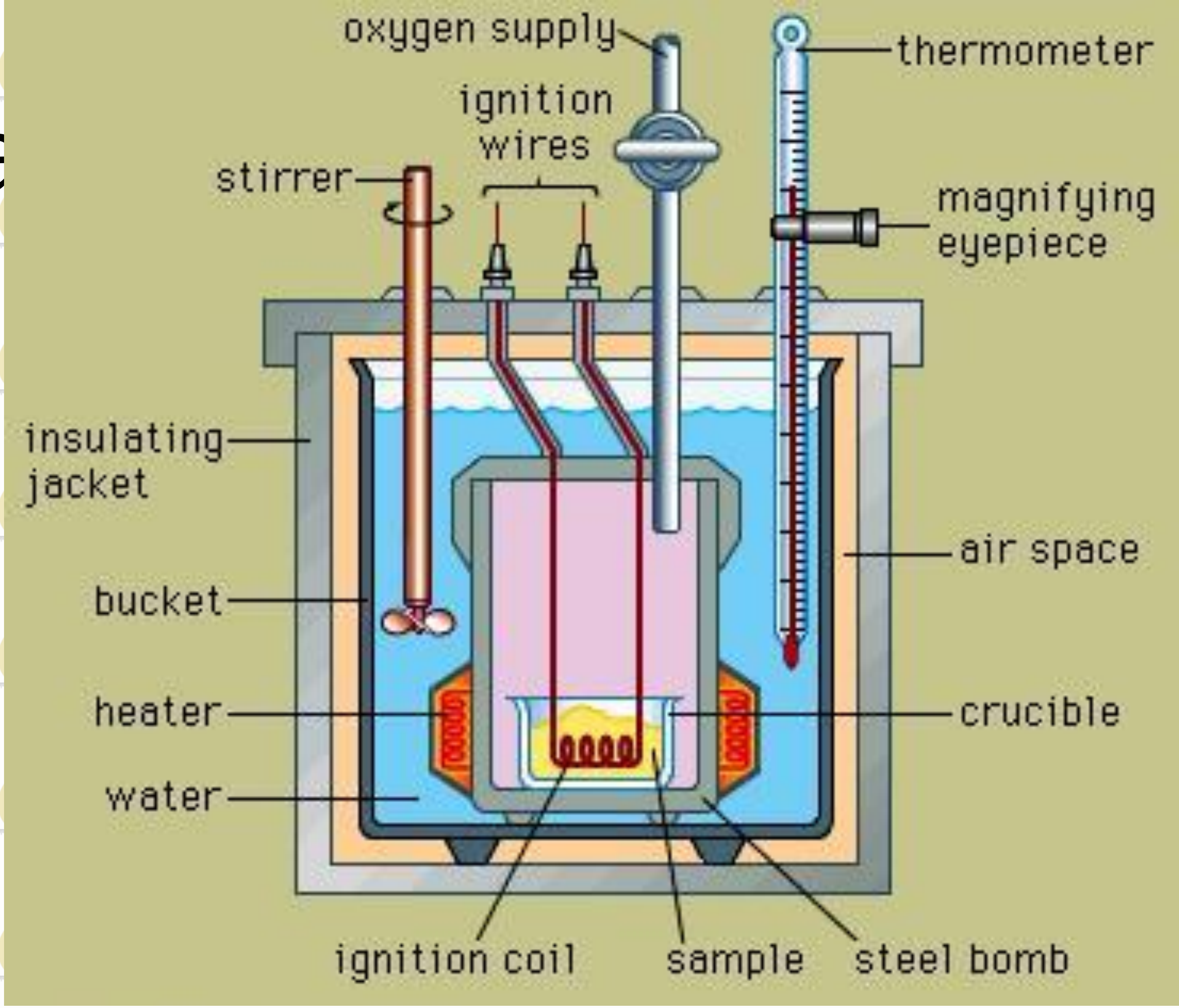
P = Power of the electric heater	(W)
t = time (in second)	(s)
m = mass	(kg)
c = specific heat capacity	($J\ kg^{-1}\ {}^{o}C^{-1}$)
θ = temperature change	(${}^{o}C$)

Calorimeter



Measuring Hea

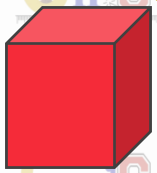
- **Calorimeter** – a device used to measure the energy given off or absorbed during a chemical or physical change.



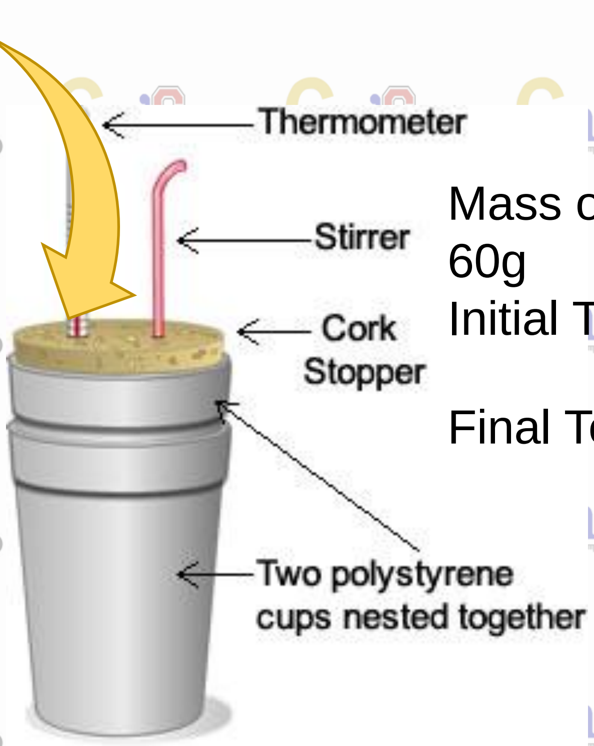
Calorimetry

- Can be used to determine the identity of a sample based on temperature changes and mass of sample.

Unknown metal block



Mass of block: 226g
Initial Temp: 100°C



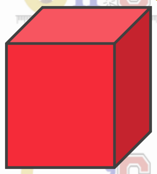
Mass of water: 60g
Initial Temp: 30°C

Final Temp: 50°C

Calculate the specific heat of the unknown metal.

Unknown metal block

Mass of block: 226g
Initial Temp: 100°C



Mass of water: 60g
Initial Temp: 30°C

Final Temp: 50°C

The heat (Q) lost by the block is gained by water

$$Q_{block} = Q_{water}$$

THERMOCHEMISTRY

Heat gained or lost = (mass)$\begin{pmatrix}\text{specific}\\ \text{heat}\end{pmatrix}\begin{pmatrix}\text{change in}\\ \text{temperature}\end{pmatrix}$ $\qquad Q = mc_p\Delta T$

C_{water} = 4.18 J/g°C

(226g)(C)(50°C-100°C)= (60g)(4.18 J/g°C)(50°C-30°C)

C = 0.444 J/g°C

Comparing Specific Heat

- What happens to the change in temperature or equal masses of copper and water when equal amounts of heat energy are given?
- c for Cu = 0.387 J/(g x °C)
- Use mass = 1.0g   Q=10.0J

Cu $$10.0J = (1.0g)\left(0.387\frac{J}{g\cdot{}^{o}C}\right)(\Delta T)$$ **ΔT = 25.8°C**

H_2O $$10.0J = (1.0g)\left(4.18\frac{J}{g\cdot{}^{o}C}\right)(\Delta T)$$ **ΔT = 2.4°C**

Practice example

- A piece of unknown metal with mass 23.8 g is heated to 100.0^oC and dropped into 50mL of water at 24.0^oC. The final temperature of the system is 32.5^oC. What is the specific heat of the metal? The density of water is 1g/mL.

Metal

m = 23.8g

ΔT = 100 – 32.5 = 67.5°C

Water

m = 50.0g

ΔT = 32.5 – 24 = 8.5°C

c = 4.18 J/(g x °C)

$$density = \frac{mass}{volume} = \frac{1g}{1mL} = \frac{x}{50mL}$$

X = 50.0g

For water D = 1g/mL

Heat gained by water = Heat lost by metal

$$(50.0\cancel{g})\left(4.18\frac{J}{g\cdot{}^{o}C}\right)(8.5^{o}\cancel{C}) = (23.8\cancel{g})(c)(67.5^{o}\cancel{C})$$

c = 1.1 J/(g x °C)

Gases have two types of Molar heat capacities.

1. The molar heat capacity at constant volume (Cv).

The heat required to raise the temperature of one mole of the gas by 1^0C when the volume remains constant.

2. The molar heat capacity at constant pressure (Cp).

The heat required to raise the temperature of one mole of the gas by 1^0C when the pressure remains constant .

- When a gas is heated at constant P it expands, and therefore some of the heat which is supplied to the gas is used;

1. to do external work,
2. to increase the potential energy of its molecules.

* But when a gas is heated at constant V, all the heat is used to increase the T, So

The molar heat capacity at constant pressure (Cp) is greater than The molar heat capacity at constant volume (CV), i.e., $C_p > C_V$

- For molar heats $C_p - C_V = R$ where R = gas constant. (Mayer's formula.)

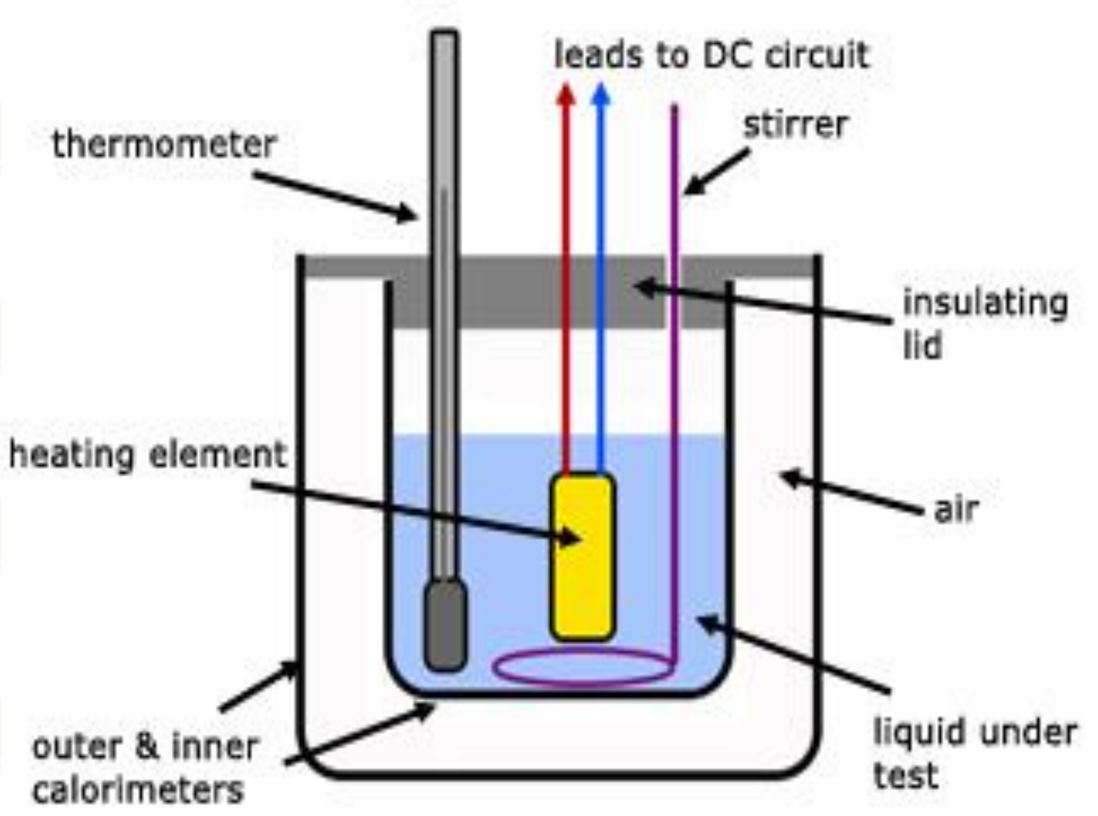
- The ratio of two principal heats of a gas is represented by γ.
- ϒ = Cp/Cv
- The value of γ depends on atomicity of the gas.

ϒ = 1.67 for a monoatomic gas

ϒ = 1.40 for a diatomic gas

ϒ = 1.33 for a polyatomic gas

Specific Heat Capacity of a **Liquid** by an Electrical Method



The heat energy supplied by the electrical element is given to the liquid and its container, producing a temperature rise **$\Delta\theta$**.

The heater current (I) and voltage (V) are monitored for a time (**t**).

energy supplied by heater = **VIt**

energy absorbed by liquid and container = $\mathbf{m_L c_L \Delta\theta + m_C c_C \Delta\theta}$

where,

$\mathbf{m_L}$ mass of liquid
$\mathbf{m_C}$ mass of container
$\mathbf{c_L}$ specific heat capacity of liquid
$\mathbf{c_C}$ specific heat capacity of container

Equating the two quantities,

$$VIt = m_L c_L \Delta\theta + m_C c_C \Delta\theta$$

$\mathbf{m_L}$, $\mathbf{m_C}$, $\mathbf{c_C}$ are known and **V**, **I**, **t**, **Δθ** are all measured.

So the specific heat capacity of the liquid ($\mathbf{c_L}$)can be calculated.

Qns

(1) How much heat (energy) is needed to supply to an iron rod of mass 5 kg to increase its temperature from 25 C to 60 C?

(Specific heat capacity of iron = 452 J kg-1 C-1)

(2) A pail contains 8 kg of hot water at 85 C. What is the temperature of the water after 1.68 106 J of heat is released?

(Specific heat capacity of water = 4200 J kg-1 C-1)

(3) An electrical kettle of power 2400 W contains water of mass 4 kg and temperature 25 C. What is the time needed to heat the water until its boiling point at 100 C?

(Specific heat capacity of water = 4200 J kg-1 C-1)

(4) 200 g of water at 100 C is poured into a cup made of glass. The initial temperature of the glass cup is 30 C and the mass of the glass cup is 150 g. What is the temperature of the water when thermal equilibrium is achieved between the water and the glass cup?

(Specific heat capacity of water = 4200 J kg-1 C-1;

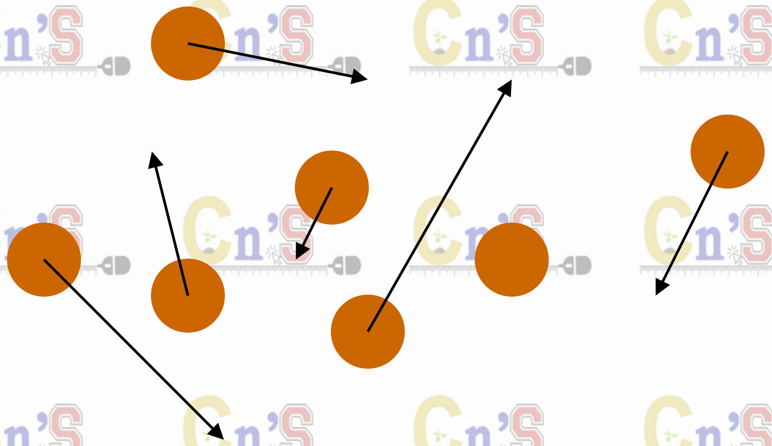
(Specific heat capacity of glass = 650 J kg-1 C-1)

Exercises

1. A liquid of mass m kg is cooled from 70^oC to 30^oC. The total heat energy lost is 8400 J. What is the value of m if the specific heat capacity of the liquid is 4200 J $kg^{-1o}C^{-1}$
2. An electric heater which has power 700 W, is used to heat up 2 kg of water. If the specific heat capacity of the water is 4200 J $kg^{-1o}C^{-1,}$ what is the time taken to raise the temperature from 28^oC to 53^oC
3. A cooler is connected to a container which holds 2.0 kg of water at 50oC. If the cooler absorbs heat at a rate of 600 Js-1, what is the time taken to cool the water from 50^oC to 0^oC? [specific heat capacity of the water is 4200 J $kg^{-1o}C^{-1}$]
4. An immersion heater which has a power of 300 W, is inserted into a metal block. The metal block has a mass of 1.5 kg and its specific heat capacity is 420 J $kg^{-1o}C^{-1}$. what is increase temperature in every 5s when the heater is on?
5. 800 g of liquid X at 30^oC is mixed with the same liquid of mass M kg at 80^oC . The final temperature after mixing is 60^oC. What is value of M?

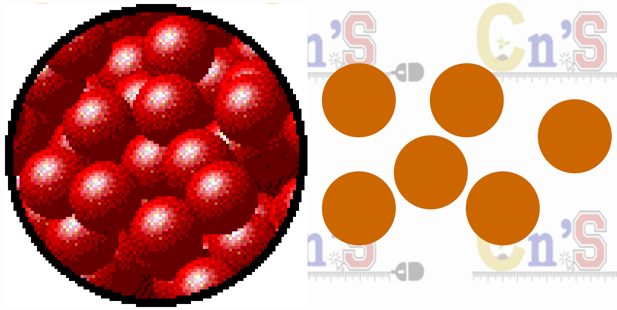
Phases of matter

Gas - very weak intermolecular forces, rapid random motion



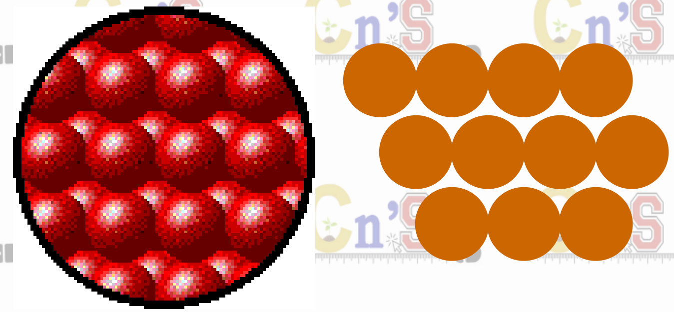
high temp
low pressure

Liquid - intermolecular forces bind closest neighbours



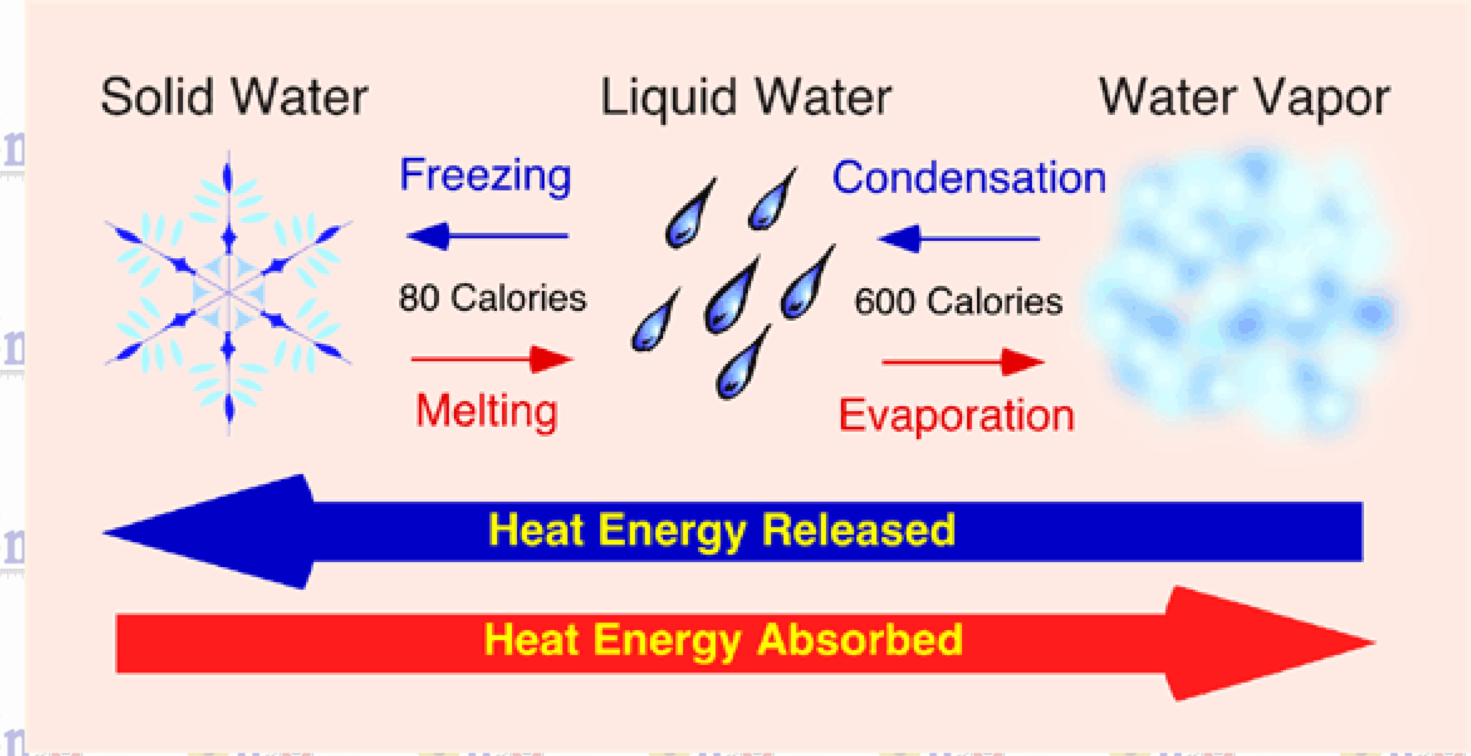
low temp
high pressure

Solid - strong intermolecular forces



Phase Changes

- *Phase changes* (solid, liquid, gas) are accompanied by **release** or **absorption** of heat energy



phase change	action	symbol
solid to liquid	melting	L_F
liquid to solid	fusion	L_F
liquid to vapour	vaporization	L_V
vapour to liquid	condensation	L_V
solid to vapour	sublimation	L_S
vapour to solid	sublimation	L_S

Specific Latent Heat of Fusion

- When we heat a solid, it will melt and become a liquid. At this point, energy is required but the temperature does not rise.

- The specific latent heat of fusion of a substance is the heat energy required to change 1kg of solid at its melting point to 1 kg of liquid – without a change in temperature.

Specific Latent Heat of Vaporisation

- When we heat a liquid, it will evaporate and become a gas. At this point, energy is required but the temperature does not rise.

- The specific latent heat of vaporisation of a substance is the heat energy required to change 1kg of liquid at its boiling point to 1kg of vapour (gas) – without a change in temperature.

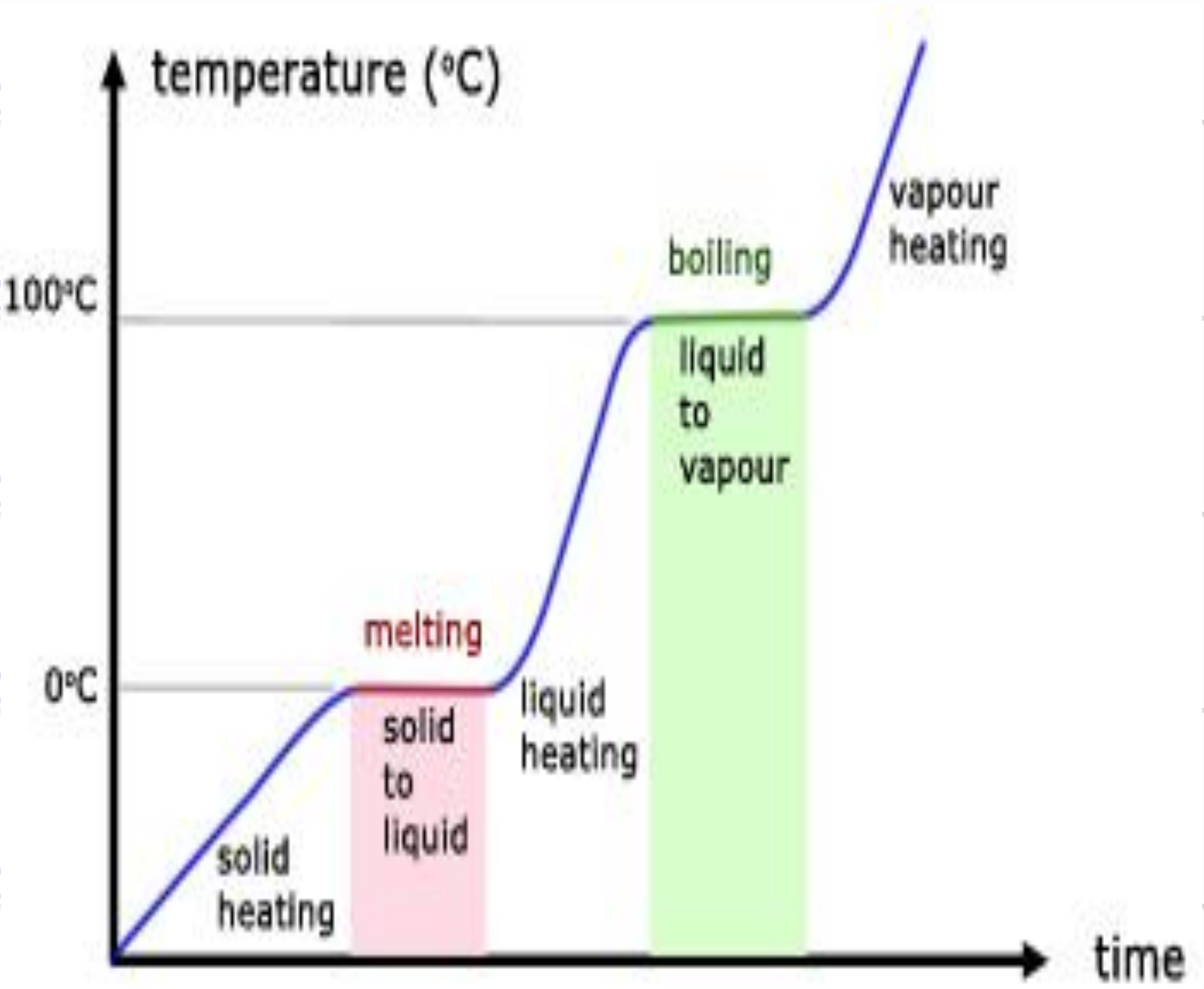
Phase Changes

- The amount of energy involved, per unit mass, is known as the **latent heat**

$$Q = m\,L$$

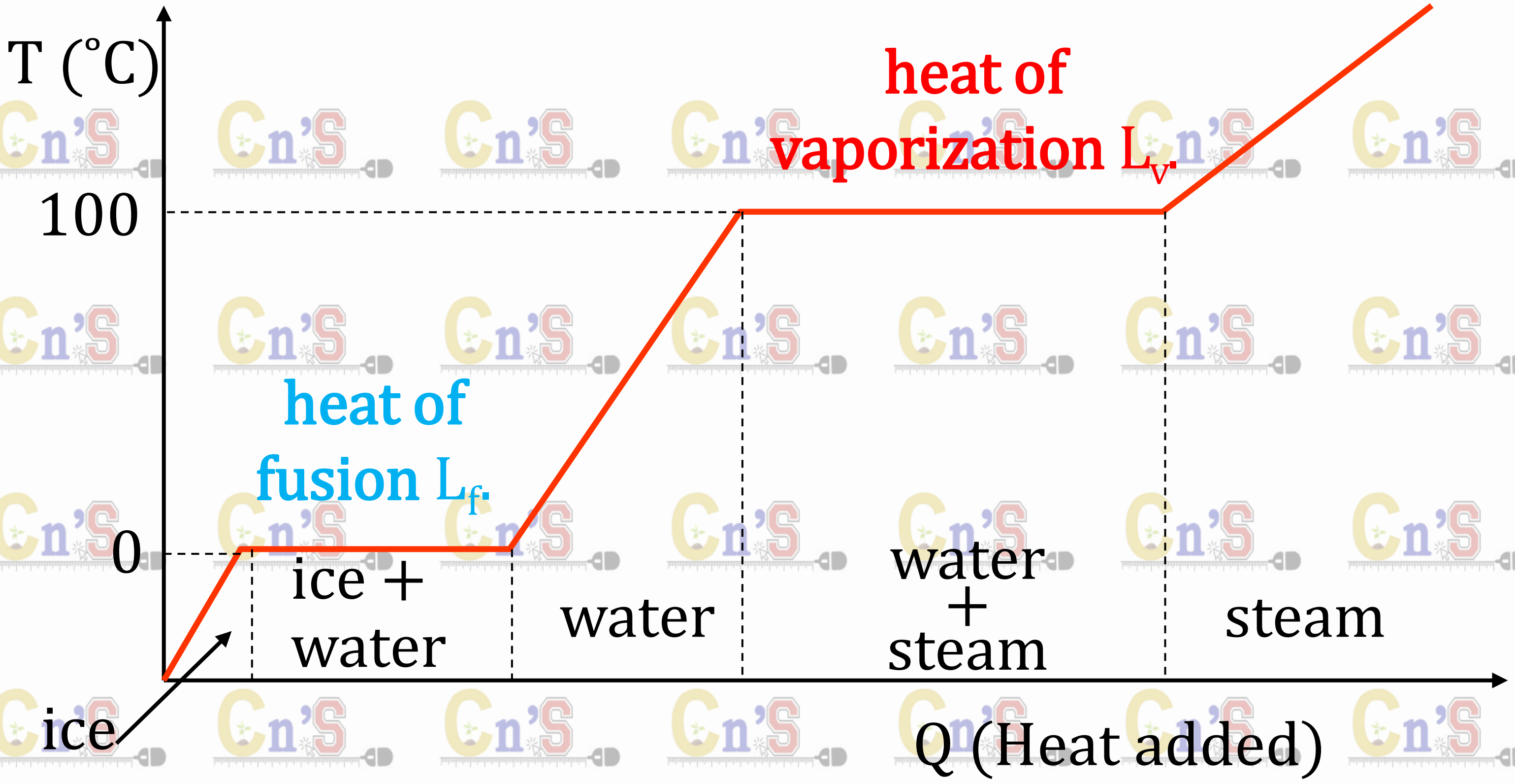
Q = heat energy absorbed/released [J]
m = mass of substance [kg]
L = latent heat [J/kg]

The graph illustrates the temperature changes when a solid(eg ice) is heated from below its melting point, to above boiling.



Phase Changes

eg. adding heat to ice, initially with T=-20 ˚C at 1 atm.



The energy per unit mass to change a phase is called the **latent heat of transformation** L. For solid-liquid change its heat of fusion L_f, for liquid-gas change its heat of vaporisation L_v.

Phase Changes

TABLE 17.1 Heats of Transformation (at Atmospheric Pressure)

Substance	Melting Point (K)	L_f (kJ/kg)	Boiling Point (K)	L_v (kJ/kg)
Alcohol, ethyl	159	109	351	879
Copper	1357	205	2840	4726
Lead	601	24.7	2013	858
Mercury	234	11.3	630	296
Oxygen	54.8	13.8	90.2	213
Sulfur	388	38.5	718	287
Water	273	334	373	2257
Uranium	1406	82.8	4091	1875

$$Q = mL$$

Latent heat – phase change

(formation or breakage of chemical bonds requires or releases energy)

Water - large values of latent heats at atmospheric pressure

L_f = 3.34x10^5 J.kg^{-1} (273 K)

L_v = 2.26x10^6 J.kg^{-1}(373 K)

Calculate the heat required to

a) change 2.0 kg of ice at 0 °C to water at 0 °C

b) change 0.60 kg of water at 100 °C to steam at 100 °C.

a) $E_h = ml$ where l is the specific latent heat of fusion of water

List all the values and their units: $m = 2.0$ kg $l = 3.34 \times 10^5$ J/kg

$E_h = 2 \times 3.34 \times 10^5$

$= 668000$

$= 670$ kJ

$= 6.7 \times 10^5$ J (To 2 significant figures)

b) $E_h = ml$ where l is the specific latent heat of vaporisation of water

List all the values and their units: $m = 0.60$ kg $l = 22.6 \times 10^5$ J/kg

$E_h = 0.60 \times 22.6 \times 10^5$

$= 1356000$

$= 1.4$ MJ

Notice that the specific latent heat of vaporisation is larger than the specific latent heat of fusion.

- A 30g sample of water is heated from 75°C to 135°C. How much energy is needed?

$$Q = m \cdot c \cdot \Delta T \qquad Q = m \cdot \Delta H_v$$

$$Q = (30g)\left(4.18\frac{J}{g \cdot {}^oC}\right)(25^oC)$$

Q = 3,135 J

This accounts for liquid portion of water.

$$Q = (30g)\left(2260\frac{J}{g}\right)$$

Q = 67,800 J

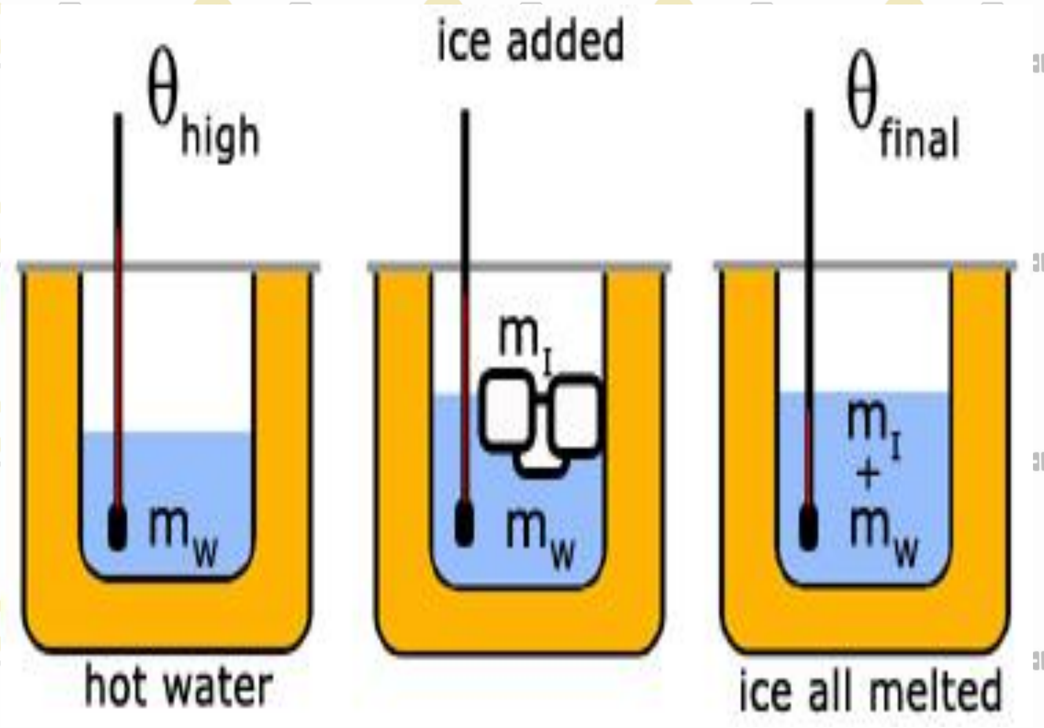
This accounts for the phase change (liquid to gas)

$$Q = (30g)\left(2.02\frac{J}{g \cdot {}^oC}\right)(35^oC)$$

Q = 2,121 J

Accounts for gas (steam) portion of water.

The Specific Latent Heat of Ice by the 'Method of Mixtures'



Ice cubes are added to water of known temperature in a copper calorimeter.

The mixture is stirred until all the ice has melted and a final reading of temperature taken.

heat energy lost from hot water in cooling from θ_{high} to θ_{final} + heat energy lost from calorimeter in cooling from θ_{high} to θ_{final}

$$m_W c_W(\theta_{high} - \theta_{final}) + m_C c_C(\theta_{high} - \theta_{final})$$

= heat energy used to melt ice at 0°C + heat energy used to increase temperature of melted ice from 0°C to θ_{final}

$$= m_I l + m_I c_W(\theta_{final} - 0)$$

where,

m_L mass of water

m_I mass of ice

m_C mass of calorimeter

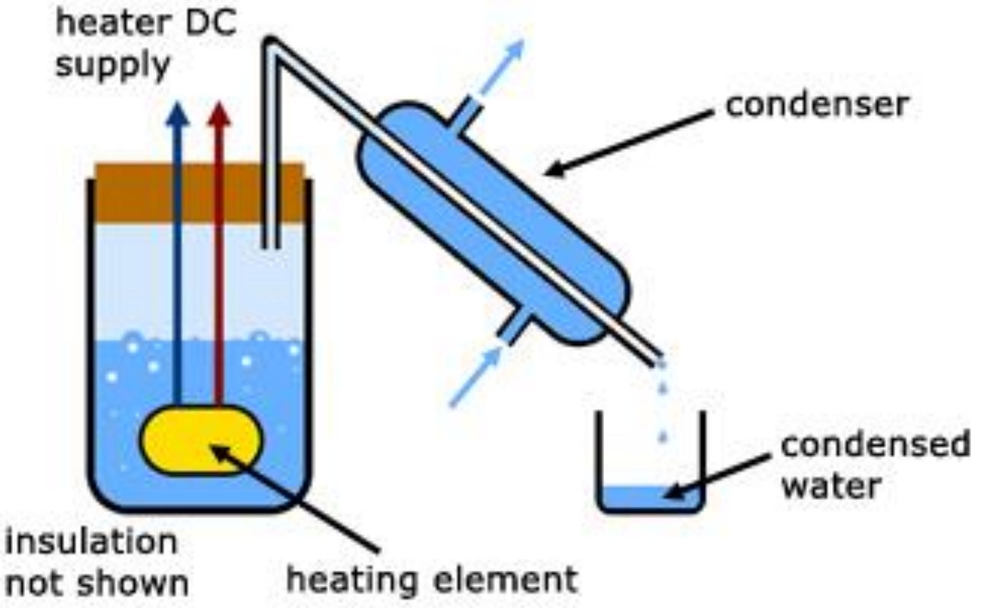
c_L specific heat capacity of liquid water

c_C specific heat capacity of calorimeter

θ_{high} temperature of the hot water

θ_{final} temperature of mixture

The Specific Latent Heat of Vaporization of a Liquid



Water is heated electrically until it boils. The condensed water (m) is collected over time (t).

Heating element readings of voltage (V) and current (I) are recorded.

In the steady state,

electrical energy supplied = heat energy to produce steam

$$VIt = ml$$

Latent heat and 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.

Evaporation and **cooling**

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.

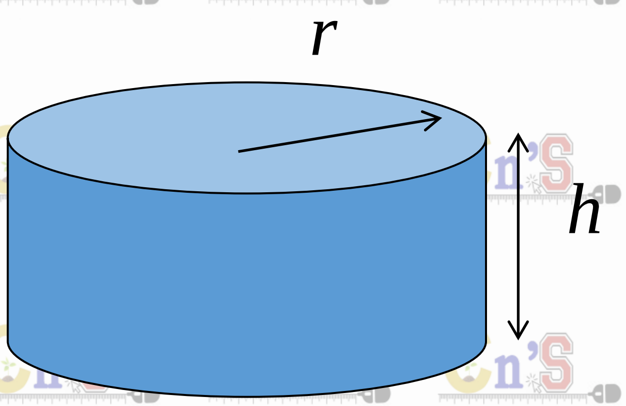
Why do dogs pant?

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.

Calculate the energy released into the atmosphere when water condenses during a thunderstorm?

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

h = 10 mm = 10^{-2} m r = 1 km = 10^3 m



volume of water $V = \pi\, r^2\, h = \pi\,(10^6)(10^{-2}) = 3\times10^4$ m^3

mass of water m = ? kg   density of water $\rho = 10^3$ kg.m^{-3}

$m = \rho\, V = (10^3)(3\times10^4) = 3\times10^7$ kg

Latent heat – change of phase

$Q = m\,L$ $L_v = 2.26\times10^6$ J .kg^{-1}

Energy released in atmosphere due to condensation of water vapour

$Q = m\,L_V = (3\times10^7)(2.26\times10^6)$ J = 7×10^{13} 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.

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.

Sketch two graphs showing the change in temperature of the ice and the temperature of the water as functions of time.

Assume no energy transfer to the surrounding environment, so that energy transfer occurs only between the water and ice.

Solution

ICE gains energy from the water

$\boxed{\text{Ice } -10\ ^{o}\text{C}} \longrightarrow \boxed{\text{Ice/water } 0\ ^{o}\text{C}} \longrightarrow \boxed{\text{Water } 5\ ^{o}\text{C}}$

WATER losses energy to the ice

$\boxed{\text{Water } 20\ ^{o}\text{C}} \longrightarrow \boxed{\text{Water } 5\ ^{o}\text{C}}$

Heat gained by ice Q_{ice} = heat lost by water Q_{water}

$$\boxed{Q = m\,c\,\Delta T \qquad Q = m\,L}$$

m_{ice} = ? kg m_{water} = 4.00 kg

temperature rise for ice to melt

ΔT_{ice1} = 0 – (–10) °C = 10 °C

temperature rise of melted ice

ΔT_{ice2} = (5 – 0) °C = 5 °C

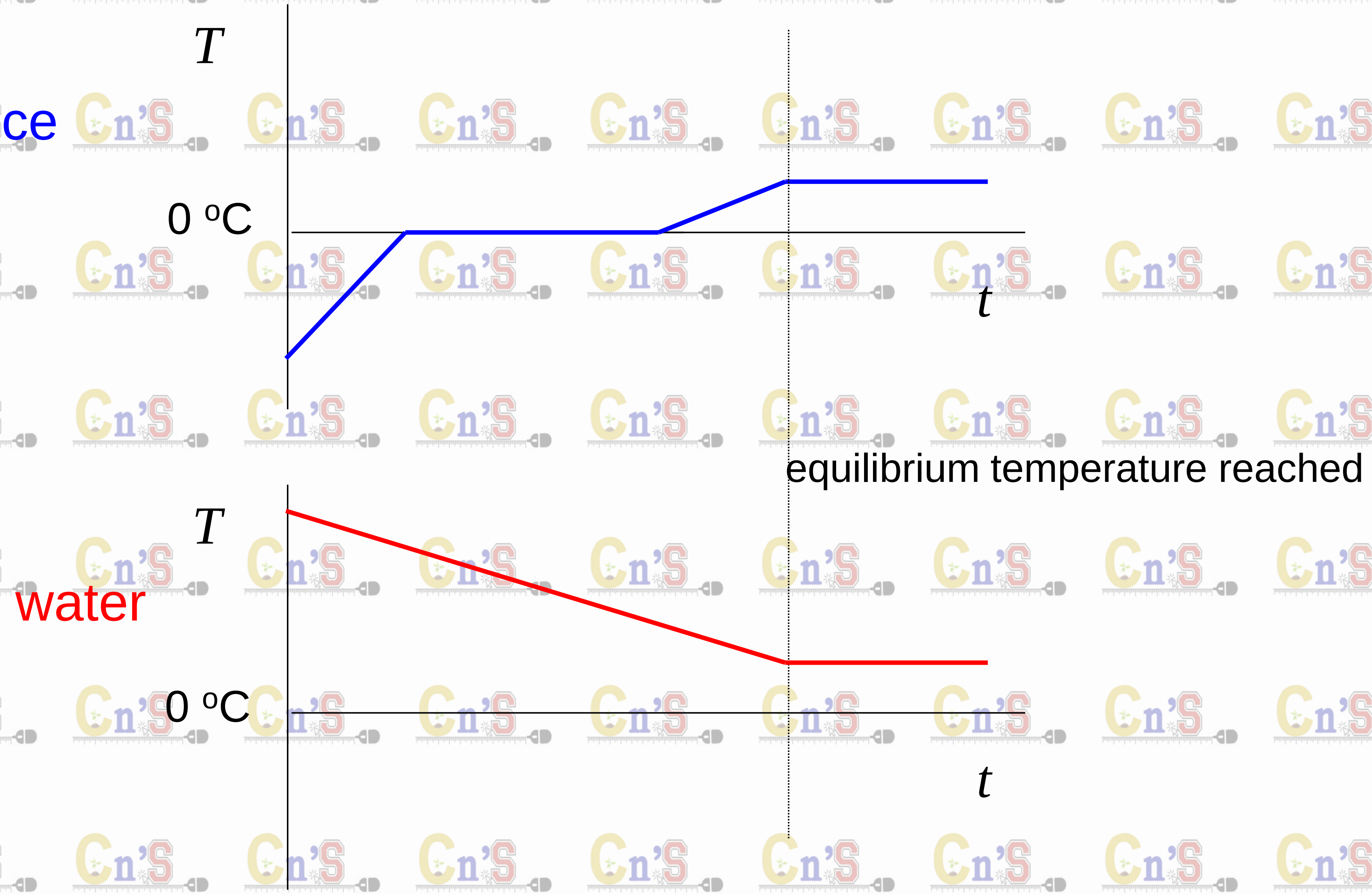
temperature fall for water

ΔT_{water} = (20 – 5) °C = 15 °C

c_{ice} = 2100 J.kg^{-1}.K^{-1} c_{water} = 4190 J.kg^{-1}.K^{-1}

L_f = 3.34×10^5 J.kg^{-1}

Execute



heat gained by ice

$$Q_{\text{ice}} = m_{ice}\, c_{\text{ice}}\, \Delta T_{\text{ice1}} + m_{\text{ice}}\, L_{\text{f}} + m_{ice}\, c_{\text{water}}\, \Delta T_{\text{ice2}}$$

heat lost by water

$$Q_{\text{water}} = m_{\text{water}}\, c_{\text{water}}\, \Delta T_{\text{water}}$$

conservation of energy

$$Q_{\text{ice}} = Q_{\text{water}}$$

$$m_{\text{ice}}\, c_{\text{ice}}\, \Delta T_{\text{ice1}} + m_{\text{ice}}\, L_{\text{f}} + m_{ice}\, c_{\text{water}}\, \Delta T_{\text{ice2}} = m_{\text{water}}\, c_{\text{water}}\, \Delta T_{\text{water}}$$

$$m_{\text{ice}} = \frac{m_{\text{water}}\, c_{\text{water}}\, \Delta T_{\text{water}}}{c_{\text{ice}}\, \Delta T_{\text{ice1}} + L_{\text{f}} + c_{\text{water}}\, \Delta T_{\text{ice2}}}$$

$$m_{\text{ice}} = \frac{(4)(4190)(15)}{(2100)(10) + (3.34 \times 10^{5}) + (4190)(5)}\ \text{kg} = 0.67\,\text{kg}$$

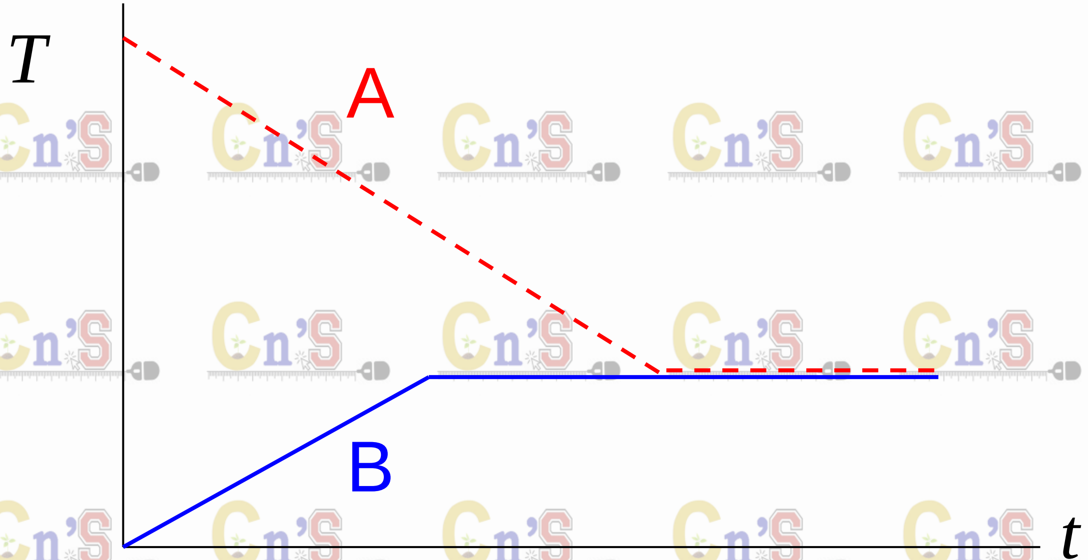
Problem C.4 June 2007 Exam Question (5 mark)

A sample of liquid water A and a sample of ice B of identical masses, are placed in a thermally isolated container and allowed to come to thermal equilibrium. The diagram below is a sketch of the temperature T of the samples verses time t. Answer each of the following questions and justify your answer in each case.

1 Is the equilibrium temperature above, below or at the freezing point of water?

2 Has the liquid water partly frozen, fully frozen, or not at all?

3 Does the ice partly melt, or does it undergo no melting?



Solution

Phase change – temperature remains constant

$Q = \pm m L$

Ice melts at 0 ^{o}C and liquid water freezes at 0 ^{o}C

Temperature change – no change in phase

$Q = m c \Delta T$

Ice warms and liquid water cools

Energy lost by liquid water (drop in temperature)

= Energy gained by ice (rise in temperature + phase change)

Execute

(1)

Ice increase in temperature initially and then remains constant when there is a change in phase. Therefore, the equilibrium temperature reached is the freezing point.

(2)

The ice reaches the freezing point first and the then the temperature remains constant. As the water cools, the ice melts. The temperature never rises above the freezing point, therefore, only part of the ice melts.

(3)

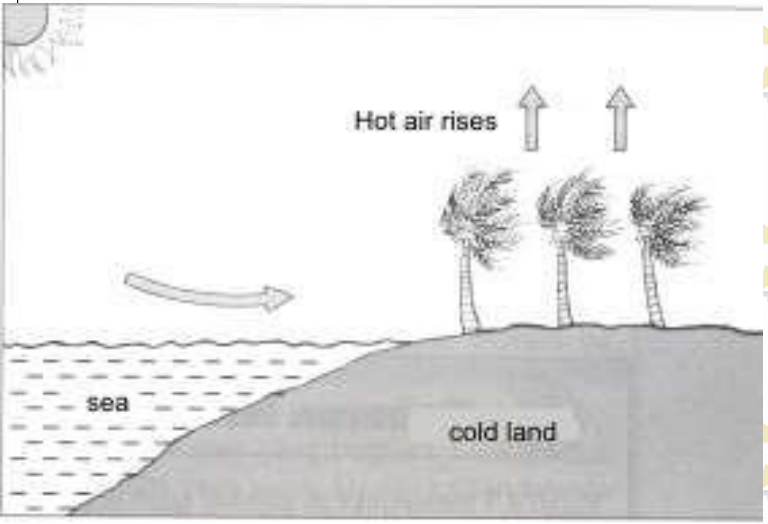
The temperature of the water falls to its freezing point and never falls below this and hence it is most likely that no liquid freezes.

- In the process of melting or boiling, heat supplied is used to increase the internal potential energy of the substance and also in doing work against external pressure while internal kinetic energy remains constant. This is the reason that internal energy of steam at 100°C is more than that of water at 100°C.
- It is more painful to get burnt by steam rather than by boiling water at same temperature. This is so because when steam at 100°C gets converted to water at 100°C, then it gives out 536 calories of heat. So, it is clear that steam at 100°C has more heat than water at 100°C (i.e., boiling of water).
- In case of change of state if the molecules come closer, energy is released and if the molecules move apart, energy is absorbed.
- Latent heat of vaporisation is more than the latent heat of fusion. This is because when a substance gets converted from liquid to vapour, there is a large increase in volume. Hence more amount of heat is required. But when a solid gets converted to a liquid, then the increase in volume is negligible. Hence very less amount of heat is required. So, latent heat of vaporisation is more than the latent heat of fusion.

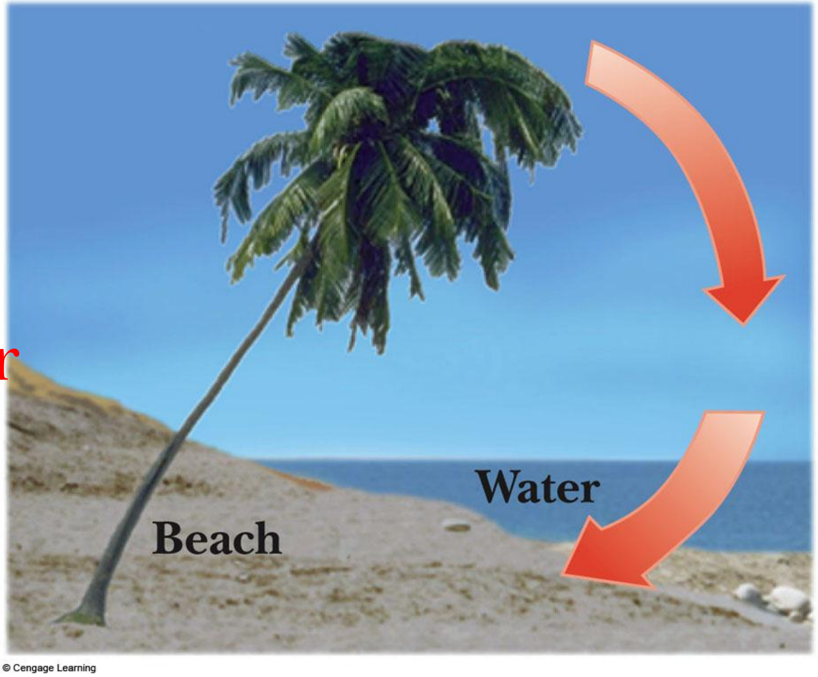
- After snow falls, the temperature of the atmosphere becomes very low. This is because the snow absorbs the heat from the atmosphere to melt down. So, in the mountains, when snow falls, one does not feel too cold, but when ice melts, he feels too cold.
- There is more shivering effect of ice-cream on teeth as compared to that of water (obtained from ice). This is because, when ice-cream melts down, it absorbs large amount of heat from teeth.
- Freezing mixture : If salt is added to ice, then the temperature of mixture drops down to less than 0°C. This is so because, some ice melts down to cool the salt to 0°C. As a result, salt gets dissolved in the water formed and saturated solution of salt is obtained; but the ice point (freeing point) of the solution formed is always less than that of pure water. So, ice cannot be in the solid state with the salt solution at 0°C. The ice which is in contact with the solution, starts melting and it absorbs the required latent heat from the mixture, so the temperature of mixture falls down.

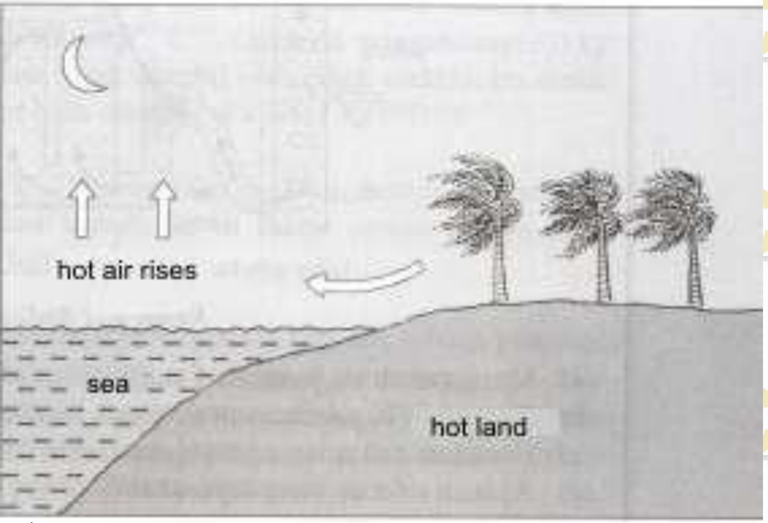
A Consequence of Different Specific Heats

Sea breeze and land breeze



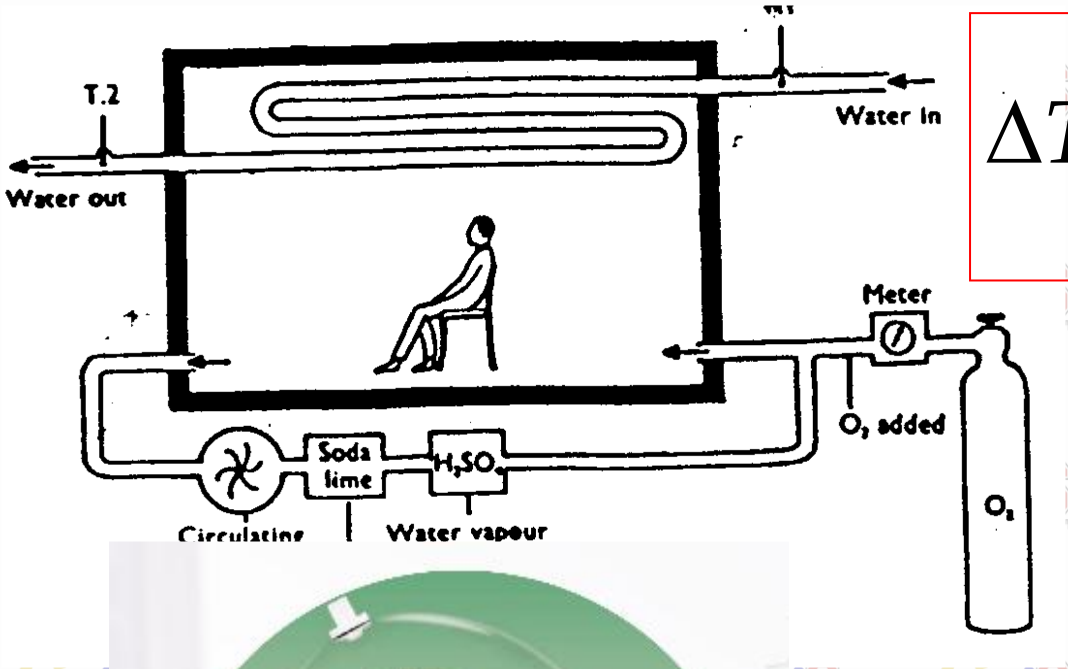
Specific heat capacity of water
= 4200 J kg-1 C-1
Specific heat capacity of land
= 840 J kg-1 C-1





- Water has a high specific heat compared to land
- On a hot day, the air above the land warms faster
- The warmer air flows upward and cooler air moves toward the beach

How do we measure a person's metabolic rate?



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

Santorio Santorio weighed himself before and after a meal, conducting the first controlled test of metabolism, AD 1614.

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.