Thermal Properties of Matter

Learning Outcomes

Candidates should be able to:

(a) describe a rise in temperature of a body in terms of an increase in its internal energy (random thermal energy)

(b) define the terms *heat capacity* and *specific heat capacity*

(c) recall and apply the relationship *thermal energy = mass* × *specific heat capacity* × *change in temperature* to new situations or to solve related problems

(d) describe melting/solidification and boiling/condensation as processes of energy transfer without a change in temperature

(e) explain the difference between boiling and evaporation

(f) define the terms *latent heat* and *specific latent heat*

(g) recall and apply the relationship *thermal energy = mass* × *specific latent heat* to new situations or to solve related problems

(h) explain latent heat in terms of molecular behaviour

(i) sketch and interpret a cooling curve

(a) Internal energy (Distinguish between internal energy, temperature and thermal energy)

- The internal energy of a body is the sum of the **total random kinetic energies and total intermolecular potential energies** of all the molecules inside it.
- **Random** – molecular movements are **disordered** and **unpredictable** (different from KE of a moving car or change in GPE of a falling object)
- If energy is transferred to a body, it gains internal energy and its molecules move faster. This is **measured** as an increased **temperature**.
- **Thermal energy** refers to energy **flowing** from a higher temperature to a lower temperature.

(b) Heat capacity

Heat capacity ***C*** is the amount of thermal energy required to raise the temperature of **a substance** by 1 K (or 1°C).

$$C = \frac{Q}{\Delta\theta}$$

Specific heat capacity

Specific heat capacity ***c*** is the amount of thermal energy required to raise the temperature of **unit mass** of a substance by 1 K (or 1°C).

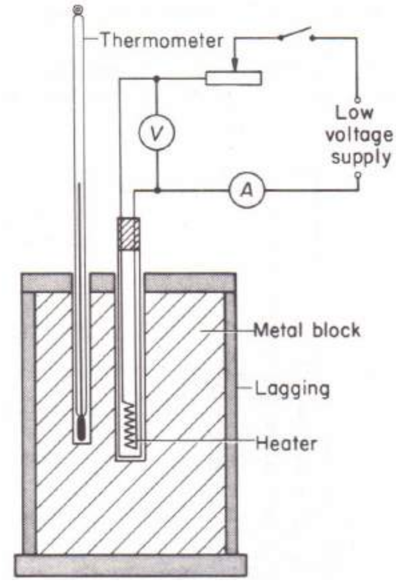
$$c = \frac{C}{m} = \frac{Q}{m\Delta\theta}$$

A substance with a **high c** heats up **more slowly** compared to a substance with a **lower c** (and vice versa)

This is important in heating and cooling systems, e.g. water is used in heating and cooling systems because it has high c. It can store a large amount of heat and also lose its heat slowly.

Question 1
Determine the quantity of heat required to raise the temperature of 100 g of ice from −20°C to −5°C. The specific heat capacity of ice is 2000 J kg^{-1} K^{-1}.

Determining specific heat capacity c of a substance (Calorimetry)



- Supply a known amount of energy from an electric heater placed inside the substance. (immersion heater)
- A **joulemeter** can be used to measure the **energy Q** transferred, or the amount of energy can be calculated using knowledge of electric circuits, from the equation **E = VIt or E = Pt**)
- Assumption – all of the substance was at the **same temperature** and the thermometer recorded the temperature **accurately** at the relevant times
- As some energy will be lost to the surroundings, **insulation** is needed to limit the energy transfers. The process involves surrounding it with a material that traps air (a poor conductor) and is known as **lagging**.
- **Calorimetry** – experiments that try to accurately measure the temperature changes produced by various physical or chemical processes. Calorimeters are used in these experiments as they are designed to limit thermal energy transfer to, or from, the surroundings.

Question 2
Determine the specific heat capacity of aluminium, given that a heater rated 30 W takes 5.0 min to raise the temperature of 450 g of aluminium from 27°C to 50°C.

Question 3
It takes the same heater 500 s to heat up a piece of iron by 5.0 K, and 200 s to heat up a piece of copper by 10 K. Which piece of metal has the larger mass? Assume no heat loss and that the specific heat capacity of iron and copper are 460 J kg^{-1} K^{-1} and 400 J kg^{-1} K^{-1} respectively.

Question 4
A piece of hot coal of mass 50 g at a temperature of 200°C is dropped into 150 g of water at a temperature of 25°C.
(a) Determine the final temperature reached. (The specific heat capacity of coal and water are 710 J kg^{-1} K^{-1} and 4200 J kg^{-1} K^{-1} respectively.)
(b) Suggest why your answer could be inaccurate.

Question 5
A large metal bolt, of mass 53.6 g was heated for a long time in an oven at 245°C. The bolt was transferred as quickly as possible from the oven into a beaker containing 257.9 g of water initially at 23.1°C. The water was stirred continuously and the temperature rose to a maximum of 26.5 °C.

(a) Calculate the energy transferred to the water.
(b) Why was the bolt kept in the oven for a long time?
(c) Why was the transfer made quickly?
(d) Calculate the specific heat capacity of the material of which the bolt is made.
(e) Why was it necessary to stir the water?
(f) Is the value for the specific heat capacity of the bolt likely to be an underestimate or an overestimate of its true value? Explain your answer.

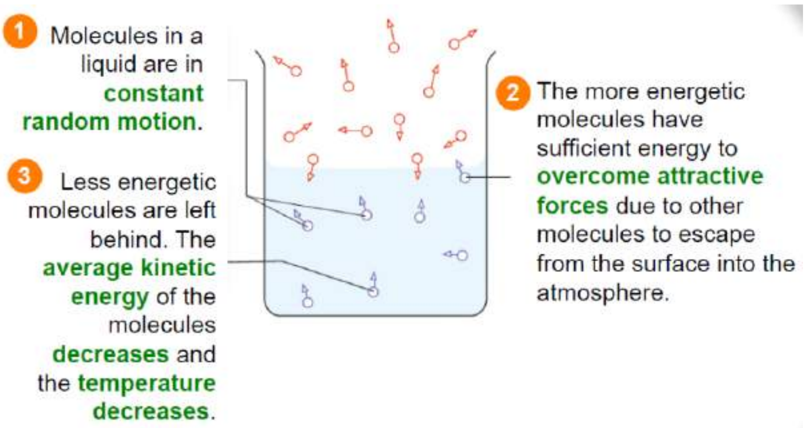
(d) Change of state

- Thermal energy, which causes a change of state, is known as latent heat. (Latent heat means 'hidden heat', in the sense that when ice melts, its temperature is constant even when absorbing energy)
- During melting or boiling, additional energy input is used to overcome **intermolecular forces** in the solid or liquid. The **potential energy** of the molecules **increase**.
- **Boiling** occurs at a precise temperature – the temperature at which the molecules have enough kinetic energy to form bubbles inside the liquid

Melting	Melting is the process in which the thermal energy absorbed by a substance changes it from solid state to liquid state without a change in temperature.
Boiling	Boiling is the process in which the thermal energy absorbed by a substance changes it from liquid state to gaseous state without a change in temperature.
Condensation	Condensation is the process in which the thermal energy taken away from a substance changes it from gaseous state to liquid state without a change in temperature.
Freezing	Freezing is the process in which the thermal energy taken away from a substance changes it from liquid state to solid state without a change in temperature.

(e) Evaporation

- Evaporation occurs when a liquid turns into a vapour. The liquid requires latent heat of vaporization which is from the liquid or its surroundings. Evaporation therefore results in cooling.
- Molecules in a liquid have a range of different kinetic energies that are continuously transferred between them as they collide.
- The molecules at the surface, which have **enough kinetic energy** to overcome the attractive forces that hold molecules together **escape from the surface**.
- The average kinetic energy of the molecules remaining decrease **resulting in cooling**
- Evaporation occurs at the surface of the liquid and at any temperature, with the **rate** of evaporation increasing with rising temperature



Differences between boiling and evaporation

Boiling	Evaporation
• Occurs at a particular temperature	• Occurs at any temperature
• Relatively fast	• Relatively slow
• Takes place throughout the liquid	• Takes place only at the liquid surface
• Bubbles are formed in the liquid	• No bubbles are formed in the liquid
• Temperature remains constant	• Temperature may change
• External thermal energy source required	• External thermal energy source not required

(f) Latent Heat

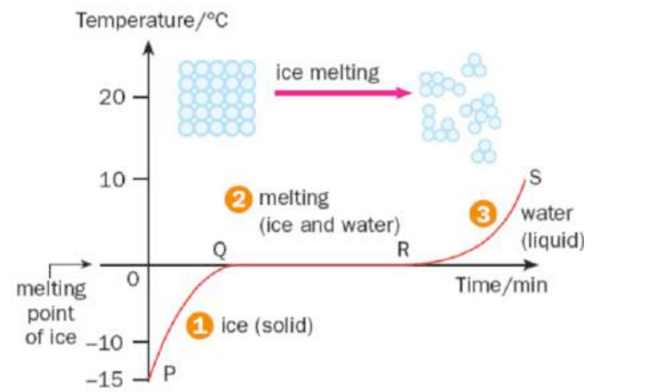
- **The thermal energy involved with changing potential energies during a change of state is known as latent heat.**
- The latent heat associated with melting or freezing is called **latent heat of fusion**. The latent heat associated with boiling or condensation is called **latent heat of vaporization**.

Specific latent heat (*l*)

- The specific latent heat of a substance is the amount of energy transferred when unit mass of the substance changes state at a constant temperature

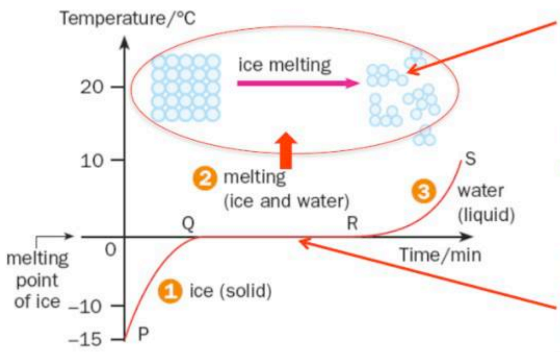
$$l = \frac{Q}{m}$$

(h) Latent heat and Molecular behaviour



1. From P to Q, the temperature of solid ice rises from −15°C to 0°C.
2. From Q to R, the temperature remains steady at 0°C, even though heat is being absorbed as ice melts.
3. From R to S, the temperature of melted ice rises from 0°C to 10°C.

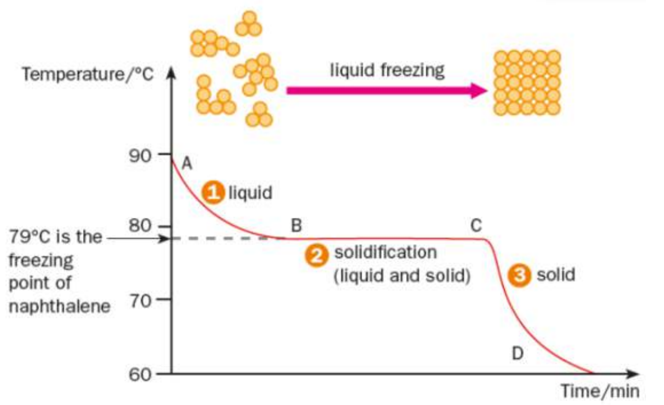
Why is PQ steeper than RS?



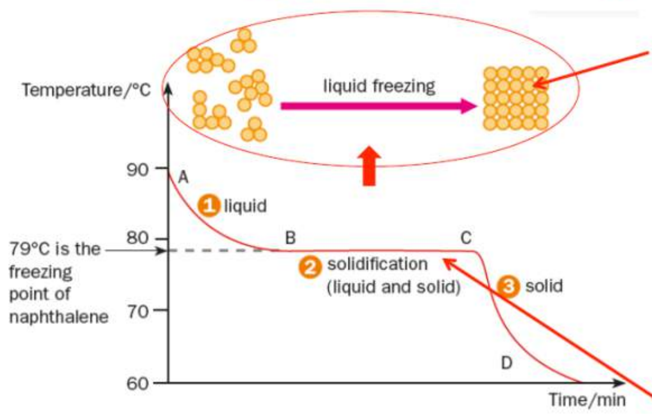
Between **Q and R**, thermal energy is absorbed to weaken the strong bonds between the particles of the solid ice.

Only the **total internal potential energy** of the particles is increased.

Since **total internal kinetic energy** does not increase, the temperature remains constant during melting.



1. From A to B, the temperature of naphthalene falls from 90°C to 79°C.
2. From B to C, the temperature remains steady at 79°C, even though heat is being released as solidification occurs.
3. From C to D, the temperature of solid naphthalene falls from 79°C to room temperature.



Between **B and C**, strong bonds are formed when particles move close together during freezing.

The **total internal potential energy** of the particles decreases as thermal energy is released and lost to the surroundings.

Since **total internal kinetic energy** does not decrease, the temperature remains constant during solidification.

Question 6
2.00 kg of water at 100°C requires 4400 kJ of energy to boil off completely. Determine its latent heat of vaporisation and specific latent heat of vaporisation.

Question 7
A 1.00 kW heater immersed in 550 g of crushed ice at 0°C is switched on for five minutes. In the five minutes, the ice melts and the temperature of the melted water rises to 49°C. Assuming that there is no heat loss to the surroundings and the specific heat capacity of water is 4200 J kg^{-1} K^{-1},
(a) what is the total energy supplied by the heater?
(b) how much energy is needed to raise the temperature of the water from 0°C to 49°C?
(c) using your answers to (a) and (b), what is the specific latent heat of fusion of ice?

Question 8
During the melting process, there is no change in temperature even though thermal energy is absorbed. Where does the thermal energy go?

Solution
The thermal energy absorbed is used to overcome intermolecular bonds in order for the solid to melt. The internal potential energy of the particles increases.

Question 9
1 kg of ice at −10°C is heated until it becomes steam at 100°C. State the **effect** of thermal energy on the water molecules in the mass of ice at each stage of heating **and** the amount of **energy** needed at each stage. (c of ice = 2100 J/(kgK), l_f = 3.4 × 10^5 J kg^{-1}, c of water = 4200 J/(kgK), l_v = 2200 kJ kg^{-1})

Solution
- −10°C to 0°C:

The ice is in solid state and the thermal energy causes the molecules to **vibrate more energetically** as temperature increases.
Q_1 =

- At 0°C:

Melting occurs and the thermal energy absorbed is used to **overcome/weaken** the bonds between the molecules such that they have greater freedom of motion in the liquid state.
Q_2 =

- 0°C to 100°C:

As thermal energy is absorbed by the liquid water, the temperature increases. The kinetic energy of the molecules increases and the molecules move more quickly.
Q_3 =

- At 100°C:

Boiling occurs and the thermal energy absorbed is used to completely break the intermolecular bonds and provide energy to the molecules to push back on the atmosphere and escape into the air.
Q_4 =

- 100°C to 110°C:

The thermal energy absorbed is used to raise the temperature of the water vapour molecules. The molecules gain kinetic energy and move with greater speeds.

Q_5 =

4 Fig. 4.1 shows steam from a boiler passing through a turbine connected to a generator.

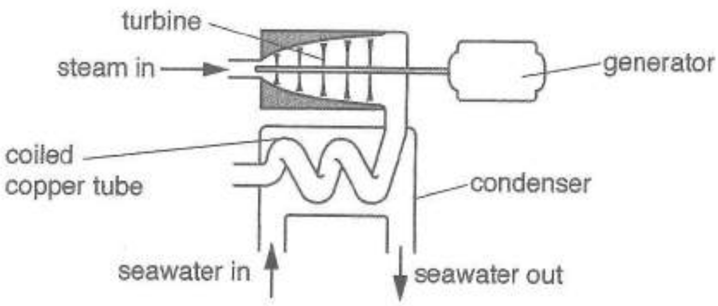


Fig. 4.1

Steam passes through the turbine and condenses in the condenser. The internal energy of the seawater rises.

(a) State what is meant by *condensation*.

.... [1]

(b) The steam is not in contact with the seawater.
Explain how condensation of the steam causes the internal energy of the seawater to rise and state the effect on the molecules of the seawater.

....

....

....

....

.... [3]

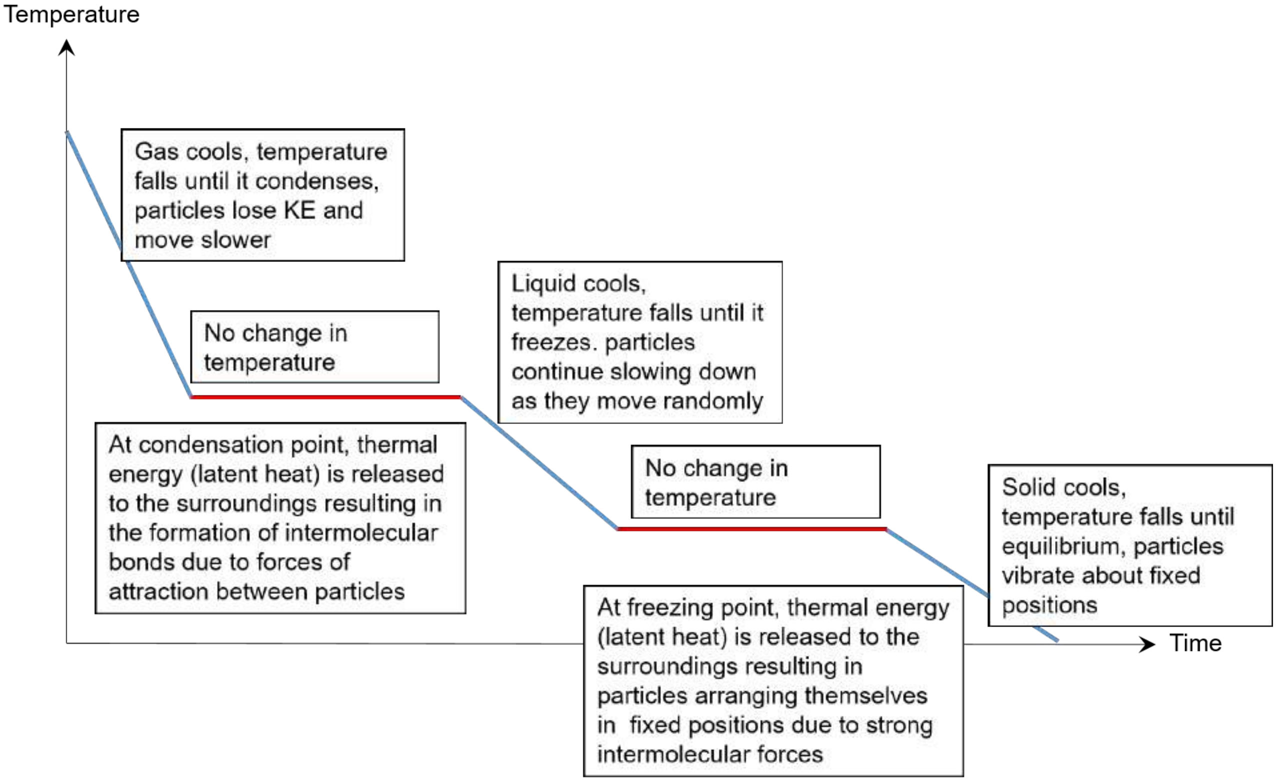
(c) The seawater enters the condenser at a temperature of 28 °C, and leaves at a temperature of 49 °C. In a certain time, 220 MJ of thermal energy passes into the seawater.

The specific heat capacity of seawater is 3900 J/(kg °C).

Calculate the mass of seawater that enters the condenser in this time.
Give your answer to an appropriate number of significant figures.

mass = ... [3]

(i) Cooling curve summary



5 A liquid, at a temperature of 100 °C, is placed in a test-tube. The sample is allowed to cool for 120 s in a laboratory where the temperature is 20 °C.

Fig. 5.1 shows the cooling curve obtained.

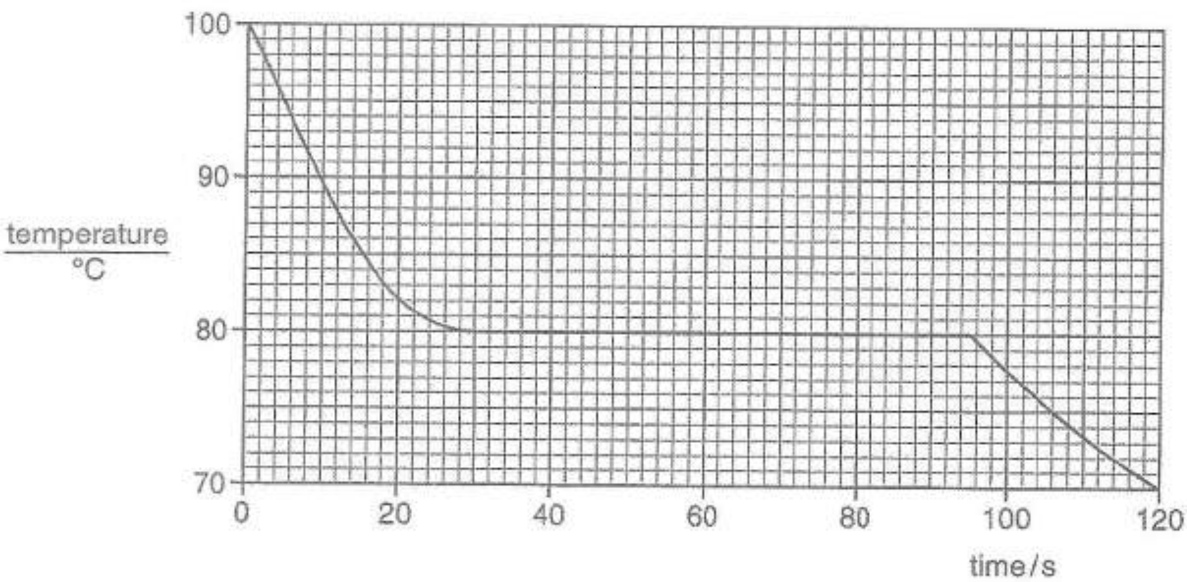


Fig. 5.1

(a) State the time when the sample becomes completely solid.

time = ... [1]

(b) State why the sample is losing thermal energy (heat) throughout the experiment.

....

.... [1]

(c) Explain why the temperature of the sample is constant for some of the time, even though thermal energy (heat) is lost to the laboratory.

....

....

.... [2]

(d) A test-tube containing an identical sample of the liquid at 100 °C is placed in a beaker of water that is kept at a temperature of 90 °C.

On Fig. 5.1, sketch the cooling curve obtained. [1]

June 1986

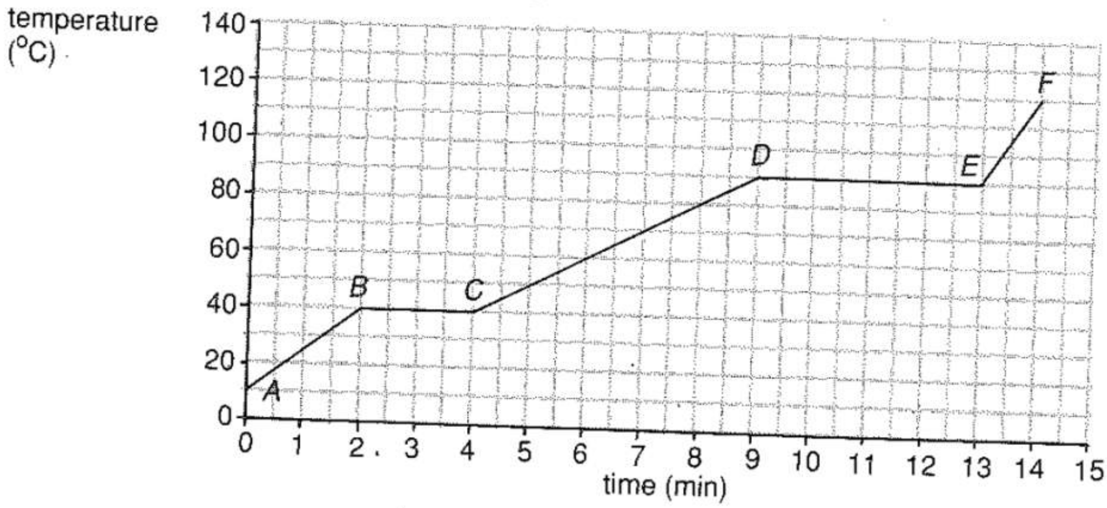
2.0 kg of ice is placed in a vacuum flask, both ice and flask being at 0□C. It is found that exactly 14 hours elapse before the contents of the flask are entirely water at 0□C. Given that the specific latent heat of fusion of ice is 3.4×10^5 J/kg calculate the average rate at which the contents gain heat from the surroundings. Suggest a reason why the rate of gain of heat gradually decreases after all the ice has melted.

Nov 81

Liquid nitrogen boils at a very low temperature, at normal atmospheric pressure. Explain why liquid nitrogen contained in an open vacuum flask in a laboratory boils steadily and continuously? Why does the liquid nitrogen boil more rapidly when contained in a glass beaker?

The following information applies to questions 7 to 12.

A 50 g solid originally at 10°C absorbs energy at the rate of 100 J per minute. The graph of Fig. 7.5 shows the resulting change in temperature with time.



7. The changes of state of the material are represented by lines
   A. *AB* and *DE*
   B. *BC* and *DE*
   C. *BC* and *EF*
   D. *CD* and *EF*
   E. *AC* and *DF*

8. The specific heat capacity of the liquid in J g^{-1} $°C^{-1}$ is closest to
   A. 0.1
   B. 0.2
   C. 0.5
   D. 5.0
   E. 10.0

9. The latent heat of fusion of the material in J g^{-1} is closest to
   A. 0.25
   B. 0.50
   C. 2.0
   D. 4.0
   E. 8.0

10. Compared with the specific heat capacity of the gas, the specific heat capacity of the solid is
    A. one third as great
    B. the same
    C. twice as great
    D. four times as great
    E. one fifth as great

11. Heating the material from 60°C to 80°C requires
    A. 100 J
    B. 200 J
    C. 600 J
    D. 800 J
    E. 10 J

12. The melting point and boiling point of the material are respectively
    A. 10°C and 120°C
    B. 10°C and 40°C
    C. 40°C and 120°C
    D. 40°C and 90°C
    E. 90°C and 120°C