CH 11: CALORIMETRY (PART 2)

CHANGE OF PHASE (STATE):

There are three phases (or states): Solid, Liquid, Gas

The process of change from one state to another **at a constant temperature** is called change of phase. It is brought about by the exchange of heat.

- **Melting and Freezing:**

  The change from solid to liquid phase by the absorption of heat at a constant temperature (called melting point), is called melting.

  The reverse change from liquid to solid by the liberation of heat at a constant temperature (called freezing point), is called freezing.

- **Vaporisation and Condensation:**

  The change from liquid to gaseous phase by the absorption of heat at a constant temperature (called boiling point), is called vaporisation.

  The reverse change from gas to liquid by the liberation of heat at a constant temperature (called condensation point), is called condensation.

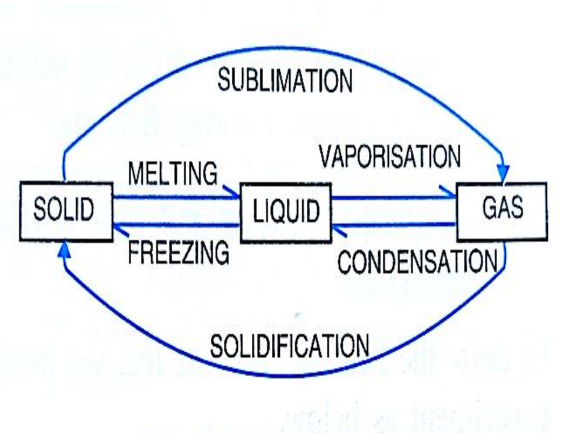


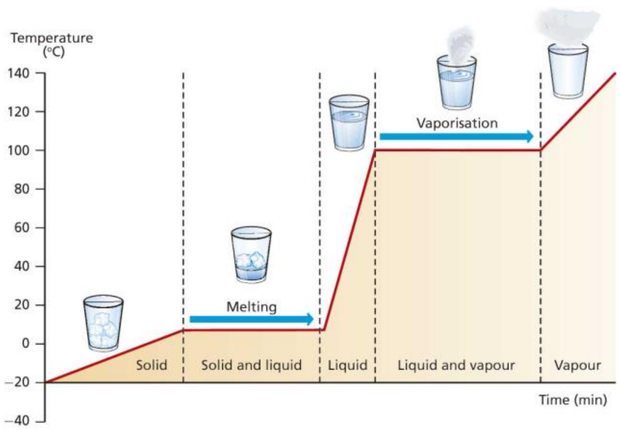
Fig. 11.2 Changes of phase

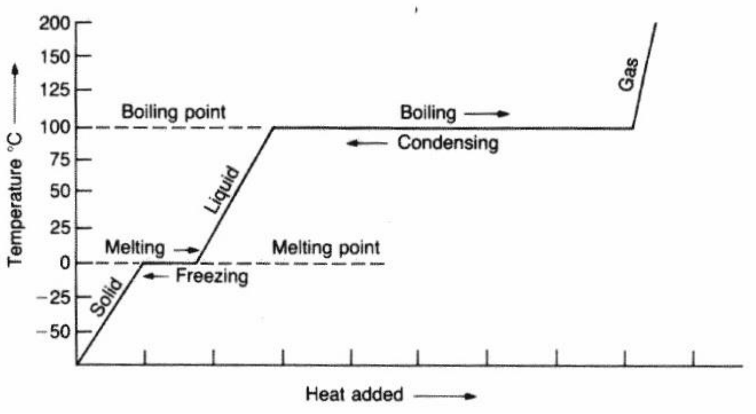
For a pure substance,

- **Melting point = Freezing Point**
  (eg: 0°C for Water, i.e. Water freezes at 0°C and Ice melts at 0°C)
- **Boiling Point = Condensation Point**
  (eg: 100°C for Water, i.e. Water boils at 100°C and steam condenses at 100°C)

- **Most substances like lead, wax, etc. expand on melting, but some substances like ice contract on melting.**
- **The melting point of the substances which contract on melting (like ice) decreases by the increase in pressure.**
- **On the other hand, the melting point of the substances which expand on melting (like lead, wax, etc.) increases by the increase in pressure.**
- **The melting point of a substance decreases by the presence of impurities in it. For example, the melting point of Ice decrease from 0°C to −22°C on mixing salt in a proper proportion and is called a freezing mixture that is used to prepare 'kulphies'.**
- **All liquids expand on boiling.**
- **The boiling point of a liquid increases with the increase in pressure and decreases with the decrease in pressure.**
- **The boiling point of a liquid increases by addition of impurities to it. For example, on addition of salt to water, it boils at a temperature higher than 100°C.**

HEATING CURVE FOR WATER:





Here is a very good and simple explanation: https://www.youtube.com/watch?v=hklSXPv2vrQ

LATENT HEAT:

- Heat energy is absorbed by a solid during melting and an equal amount is liberated by the liquid during freezing, without showing any change in temperature.
- Heat energy is absorbed by a liquid during vaporisation and an equal amount is liberated by the vapour during condensation, without showing any change in temperature.
- Since the heat energy absorbed (or liberated) in change of phase is not externally manifested by a rise or fall in temperature, it is called ***latent heat***.
- Latent heat when expressed for unit mass of a substance is called specific latent heat and is denoted by *L*.

$$\boxed{L = \frac{Q}{m} = \frac{heat\ absorbed\ (or\ liberated) for\ the\ change\ of\ phase}{mass}}$$

- Specific Latent Heat of a phase is the quantity of heat energy absorbed (or liberated) by the unit mass of the substance for the change in its phase at a constant temperature.

UNIT OF SPECIFIC LATENT HEAT:

S.I. Unit: J Kg^{-1}

Other units: cal g^{-1} & kcal kg^{-1}

SPECIFIC LATENT HEAT OF FUSION:

- The specific latent heat of fusion of ice is the heat energy required to melt unit mass of ice at 0°C to water at 0°C without any change in temperature.
- For water changing to ice i.e., freezing, it is called specific latent heat of freezing.
- For a pure substance, specific latent heat of fusion is same as the specific latent heat of freezing.
- For ice, specific latent heat of fusion is 336000 J Kg^{-1} (= 80 cal g^{-1}). This means that 1 Kg of ice at 0°C absorbs 336000 J of heat energy to convert into water at 0°C (or 1 g of ice at 0°C absorbs 80 cal of heat energy to convert into water at 0°C).
- On the other hand, 1 kg of water at 0°C will liberated 336000 J of heat energy to become ice at 0°C (1 g of water at 0°C will liberated 80 cal of heat energy to become ice at 0°C).

11.26 NATURAL CONSEQUENCES OF HIGH SPECIFIC LATENT HEAT OF FUSION OF ICE

(1) Snow on mountains does not melt all at once : The reason is that ice has a high specific latent heat of fusion (equal to 336000 J kg^{-1}). It is due to this fact that it changes into water slowly as it gets heat energy from the sun. If latent heat would have been low, all the snow would have melted in a short time on getting heat from the sun and there would have been flood in the rivers.

(2) In cold countries water in lakes and ponds does not freeze all at once : The reason is that the specific latent heat of fusion of ice is sufficiently high (= 336 J g^{-1}). The water in lakes and ponds will have to liberate a large quantity of heat to the surrounding before freezing. The layer of ice formed over the water surface, being a poor conductor of heat, will also prevent the loss of heat from the water of lake, hence it does not freeze all at once.

(3) Drinks get cooled more quickly by adding pieces of ice than the ice-cold water at 0°C : This is because 1 g of ice at 0°C takes 336 J of heat energy from the drink to melt into water at 0°C. Thus the drink libraes an additional 336 J of heat energy to 1 g ice at 0°C than to 1 g ice-cold water at 0°C. Therefore cooling produced by 1 g ice at 0°C is much more than that by 1 g water at 0°C.

(4) When ice in a frozen lake starts melting, its surrounding becomes very cold : The reason is that quite a large amount of heat energy is required for melting the frozen lake which is absorbed from the surrounding atmosphere. As a result, the temperature of the surrounding falls and it becomes very cold.

(5) It is generally more cold after a hail-storm (when ice melts) than during or before the hail-storm : The reason is that after the hail storm, ice absorbs the heat energy required for its melting from the surroundings, so the temperature of the surroundings falls further down and we feel more cold.