Heat

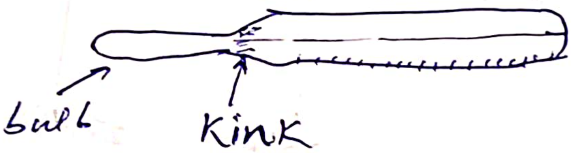
1)* Clinical thermometer range — 35°C to 42°C

Small division range : $\frac{1}{2}$ = 0.5°C

2)* Normal temperature of human body : 37°C = 98.4°F

3)* Maxm and minm temp. measured by : Max-mim thermometer.

4)* Laboratory thermometer range: -10°C to 110°C
↳ can't be used to measure body temp. because when take out the thermometer from mouth the level of mercury begins to falls.

5)* 

6)* Kink near the bulb in clinical thermometer prevents falling of mercury level.

7) Heat Transfer :-

(i) Conduction : In solid, e.g. Fe, Cu, Al
Insulator of Heat : Plastic, wood, wood

(ii) Convection : Water heating, Land breeze, Sea breeze, liquid & Gas

(iii) Radiation : No requirement of meadium, Sun rays.

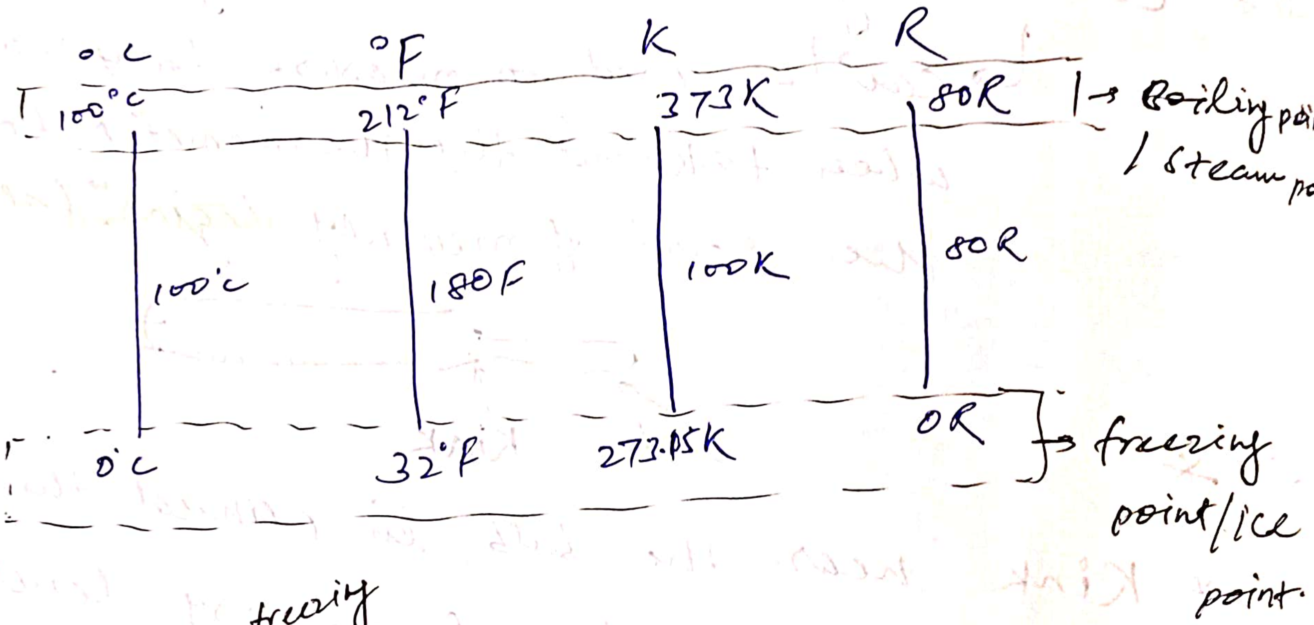
8) Dark color absorbs more heat than light-coloured.

9) Conversion

$$\boxed{\frac{C}{100} = \frac{F-32}{180} = \frac{K-273}{\cancel{373}\ 100} = \frac{R}{80}}$$

$$\frac{C}{5} = \frac{F-32}{9} = \frac{K-273}{5} = \frac{R}{4}$$

$$\boxed{F = \frac{9}{5}C + 32}$$



$$\frac{C - 0^{\swarrow \text{freezing}}}{\underset{\nearrow \text{Boiling}}{100 - 0}} = \frac{F-32}{212-32} = \frac{K-273}{373-273}$$

$$\boxed{\frac{C}{100} = \frac{F-32}{180} = \frac{K-273}{100}}$$

10) $-40°C$ to $F = ?$

$$F = \frac{9}{5} \times \overset{8}{\cancel{-40}} + 32$$

$$-72 + 32 = -40°F$$

$$\boxed{-40°C = -40°F}$$ *

11) * Kelvin does not have negative value

12) * $\boxed{160°C = 320°F}$ *

Evaporation :-

water evaporation — surface phenomena

water boiling — whole water phenomena

↳ convection method.

↳ Bulk phenomena.

↳ fixed temp.

↑ Rate of evaporation ∝ Temp. ↑

↑ evaporation ∝ wind speed ↑

evaporation ↓ ∝ Humidity ↑

* Evaporation cools the body because for evaporation heat energy absorbs from body. hence body becomes cools.

Latent Heat :

ice $\xrightarrow[\text{fusion}]{\text{Latent heat of}}$ water $\xrightarrow[\text{Vaporisation}]{\text{Latent heat of}}$ Vapour

0°C 100°C

$L = \frac{Q}{m} \frac{J}{kg}$; Q → heat, m → mass

(J/kg)

Heat Capacity :-

$$s = \frac{\Delta Q \text{ (J)}}{\Delta T \text{ (K)}}$$ heat absorb/Reject ; during change in temp.

(J/K)

Specific heat capacity : (specify the mass of substance, temp)

$$\boxed{s = \frac{\Delta Q \text{ J}}{\Delta T \cdot m \text{ (K, kg)}}} \left(\frac{\text{cal.}}{g - {}^\circ C}\right)$$

mass = 1g

temp = 1K

s (J/Kg-K)

$\boxed{1 \text{ cal.} = 4.2 \text{J}}$

* water has Specific heat capacity : 4186 J/Kg-K
& used as hot bag

Thermal Expansion :

Temp. ↑ ∝ Volume ↑

· 10°c ⟶ 9°c ⟶ 8°c ⟶ 5°c ⟶ 4°c

volume ↓↓↓

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

water density max^m at 4°c

Anomalous behaviour of water (0° — 4°c)

4°c — 3°c — 2°c — 0°c

Volume of water ↑↑↑

Density of water ↓↓↓

3 0°c to 4°C water,
density = ?
volume ↓, density ↑ (∵ anomalous behaviour of water)

Day 2

Thermodynamic:

Adiabatic : No transfer of heat

Diabatic : Heat transfer possible

1.) Isobaric process ⇒ same pressure
(ii) Isothermal process ⇒ same temperature
(iii) Isochoric process ⇒ same volume

Zeroth Law:

A
// \\
B = C

Thermodynamic state variable

① Intensive ⇒ Not changeable e.g. resistivity, Refractive, Boiling point
(ii) Extensive ⇒ change eg mass, weight, length,

Internal energy :- Sum of K.E + P.E of all molecules

- Heat and work can change the internal energy

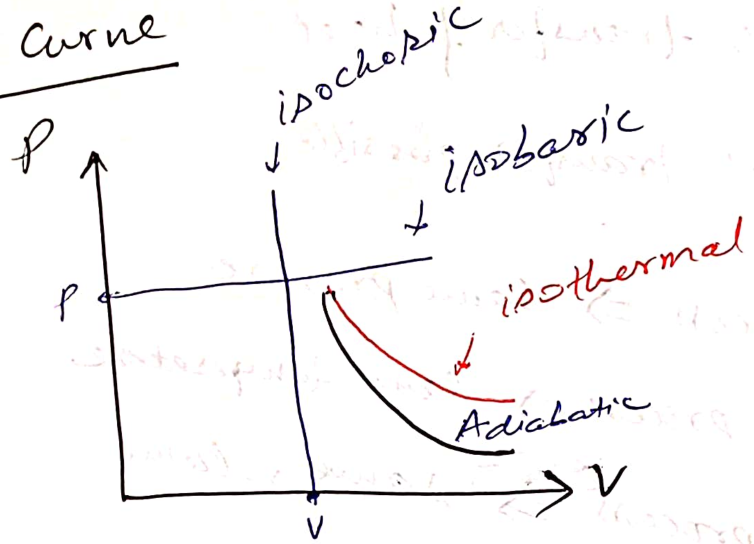
First law of thermodynamics → energy conservation

$$\Delta Q = \Delta U + \Delta W$$

ΔQ → Heat

ΔU → Int. energy

P-V curve



Adiabatic process, $\Delta Q = 0$

Isothermal process $\Delta T = 0$

- Specific heat capacity

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

, can be zero, negative or inf

Specific Heat

at constant pressure C_p

at constant volume C_v

Specific heat ratio $\gamma = \frac{C_p}{C_v}$

$C_p - C_v = R$ where R = Gas constant mono atomic gas (He, Ar)

$C_p = \frac{5}{2}R$, $C_v = \frac{3}{2}R$ $\gamma = \frac{5}{3}$

Diatomic gas

$C_p = \frac{7}{2}R$, $C_v = \frac{5}{2}R$ $\gamma = \frac{C_p}{C_v} = \frac{7}{5}$

Q. Find amount of heat required to increase the temp of 50g water by 30°C

$\boxed{\Delta T = 30K = 30°K}$

$\downarrow$

$\boxed{\text{Diff. Kelvin} = \text{Diff. °C}}$

$\Delta Q = mc\Delta T$

$\downarrow$ specific heat

$= 50g \times \frac{1\,cal}{g \cdot °C} \times 30°C$

$Q = 1500$ cal.

Note $\Delta T = 63°F$

$\boxed{\begin{matrix} \Delta T°C & & \Delta T°F \\ 100°C & = & 180°F \end{matrix}}$

$1°F = \frac{100°C}{180}$

$63°F = \frac{100}{180} \times 63°C$

Specific Latent heat (Joule/Kg)

- Specific Latent heat of Fusion

$L_{Hf} = 80$ cal/g

- Specific Latent heat of Vapo[rization]

$L_{Hv} = 540$ cal/g

Q Find heat required to change 25g ice into water

1g ice = 80 cal

$25 \times 80 = \underline{2000 \text{ cal}}$

Q Find heat required to change 25g water into vapour

1g water = 540 cal/g

$= 25 \times 540 = \checkmark$ cal

$$\boxed{\begin{array}{l} Q = mL_{Hf} \leftarrow \text{water} \\ Q = mL_{Hv} \leftarrow \text{vapour} \\ Q = mC\Delta T \end{array}}$$

$L_{Hf} = 80$ cal/g

$L_{Hv} = 540$ cal/g

ΔT °c

$C_w = 1$ cal/g°c

1 cal = 4.2 J

Q $T_F = \dfrac{m_1T_1 + m_2T_2}{m_1 + m_2}$

If two different material then

Heat gain = Heat loss

$\Delta Q_{gain} = \Delta Q_{loss}$

$$\boxed{mC_1\Delta T = mC_2\Delta T}$$

Q. ice (20g) $\longrightarrow$ 100° c (vapour)
−10°c

$$\underset{-10^\circ C}{ice} \xrightarrow{mC\Delta T} \underset{0^\circ C}{ice} \xrightarrow{mLH_f} \underset{0^\circ C}{water} \xrightarrow{mC\Delta T} \underset{100^\circ C}{water} \xrightarrow{mLHv} \underset{100^\circ C}{Vapour}$$

Soln

$\cancel{mLH_f}$

$mC\Delta T + mLH_f + mC\Delta T + mLHv$

$m(C_{ice}\Delta T + LH_f + C_w\Delta T + LHv)$

$m[0.5 \times 10 + 80 + 1 \times 100 + 540]$

$= 20 \times 725 = 14500$ cal.

$C_{ice} = 0.5 \times 1 \frac{C}{g\text{-}°C}$