Thermodynamics & Heat Transfer

by
Dr. Eng. Abdel-Nasser Saber

THERMODYNAMICS & HEAT TRANFER

1. **Fundamental Concepts and First Principles**
2. **Ideal gas and First law of thermodynamics**

Chapter 1

Fundamental Concepts

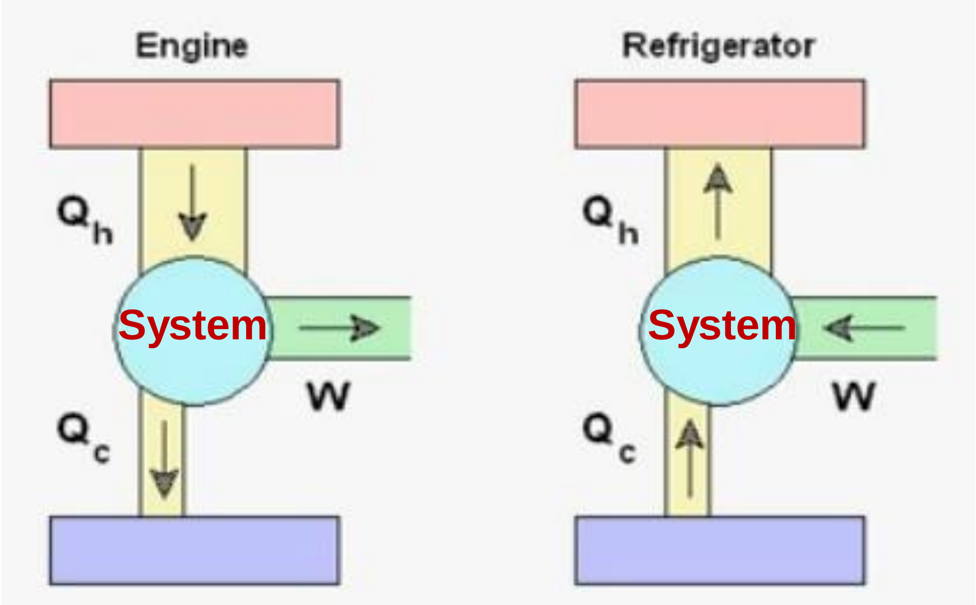
- Electric Power Generation
- Thermal station

Oil Company

❑Definitions

1. Thermodynamic Science:

It is the branch of science and an engineering tool used to describe processes that involve **changes in temperatures**, **transformation of energy**, and **the relationship between heat and work**.

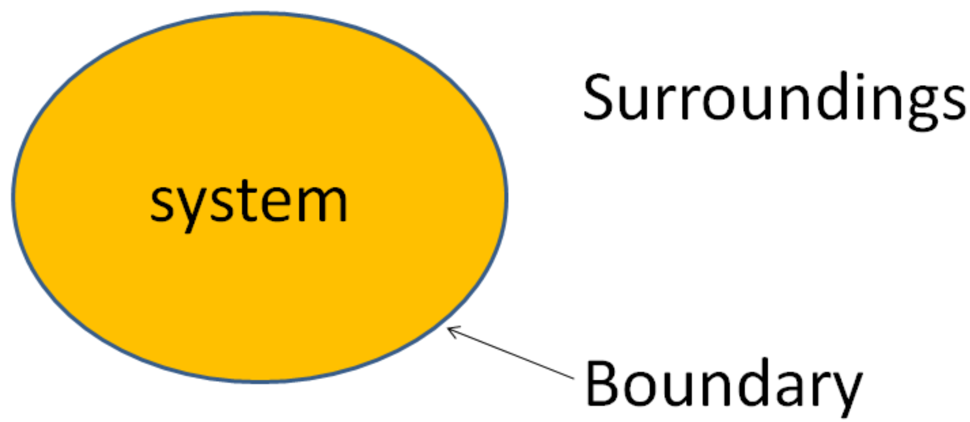


➢ Historically, thermodynamics developed out of a desire to increase the efficiency of early steam engines.

❑Definitions

2. System:

A thermodynamic system is a quantity of matter of fixed identity around which we can draw a boundary.



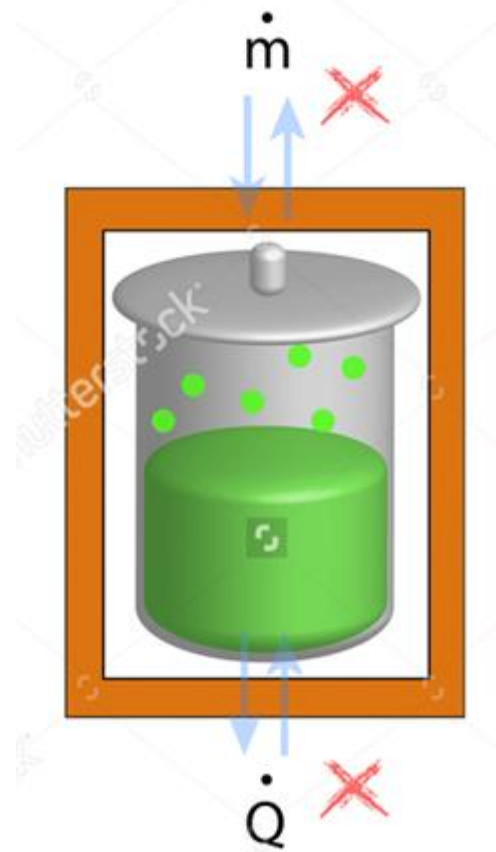
- The boundary may be fixed or movable.
- The surrounding is everything outside the boundary.
- Work and heat can be transferred across the system boundary.

❑Definitions

➢ **Types of thermodynamic systems**

Open
Heat transfer
Mass transfer

Closed
Heat transfer
NO Mass transfer



Isolated
NO Heat transfer
NO Mass transfer

❑Definitions

3. State:

the state of a thermodynamic system is the system condition at any specific time which is defined by specifying values of a set of measurable properties.

- the **measurable properties** are used to determine the values of the **non- measurable properties**.
- **State Variables of a system (Measurable + Non- measurable properties)**

- **the state variables describe the state only when the system is in equilibrium.**

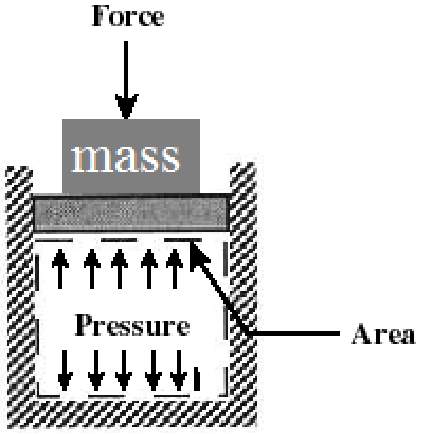
❑Definitions

4. Equilibrium:

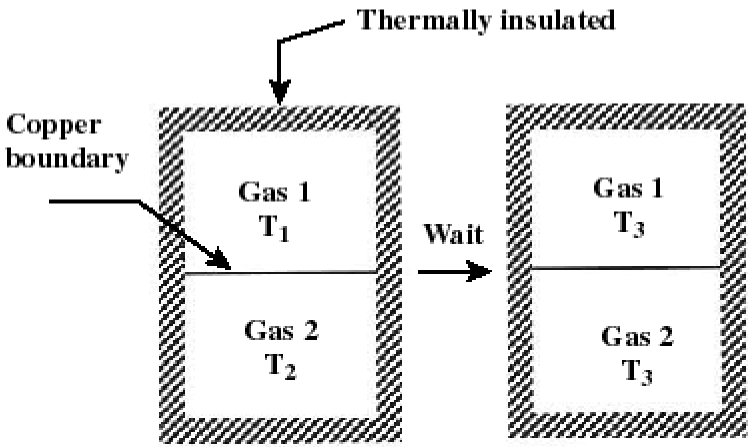
The equilibrium state of a system is a state in which the system properties have definite unchanged values as long as the external conditions are unchanged.

➢ **A system in thermodynamic equilibrium satisfies:**

- Mechanical equilibrium (no Unbalanced forces ΣF= 0)
- Thermal equilibrium (ΔT=0)
- Chemical equilibrium



Mechanical Equilibrium



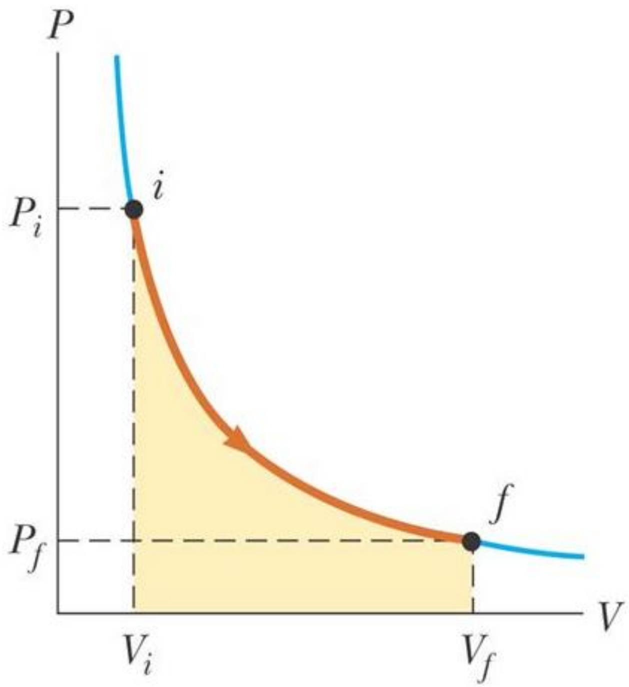
Thermal Equilibrium

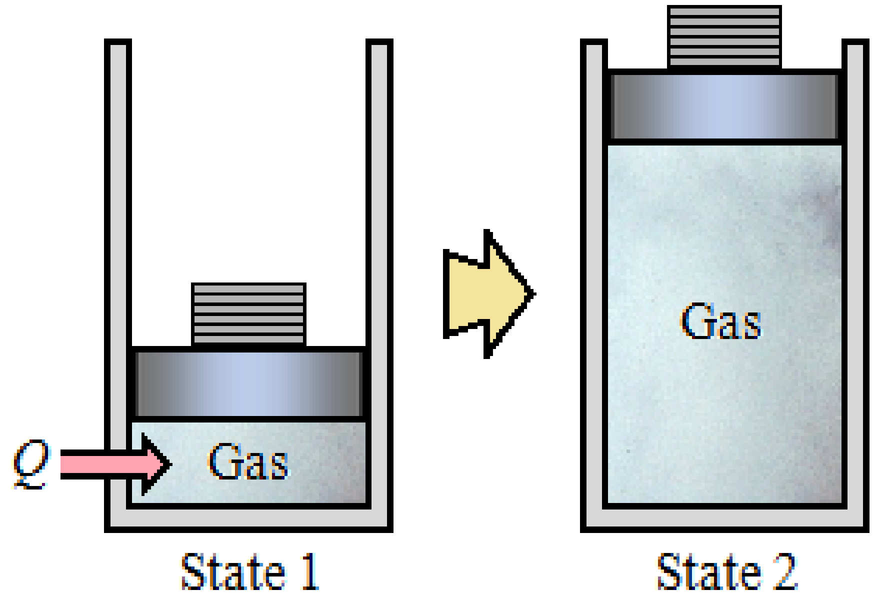
❑Definitions

5. Process:

If the state of a system changes, then it is undergoing a process.

➢ **The thermodynamic process is a passage of a thermodynamic system from an initial to a final state of thermodynamic equilibrium.**

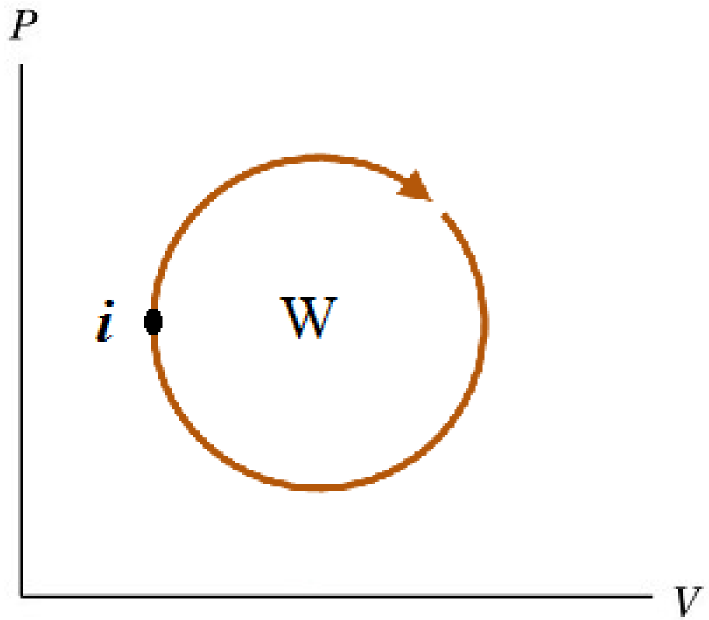


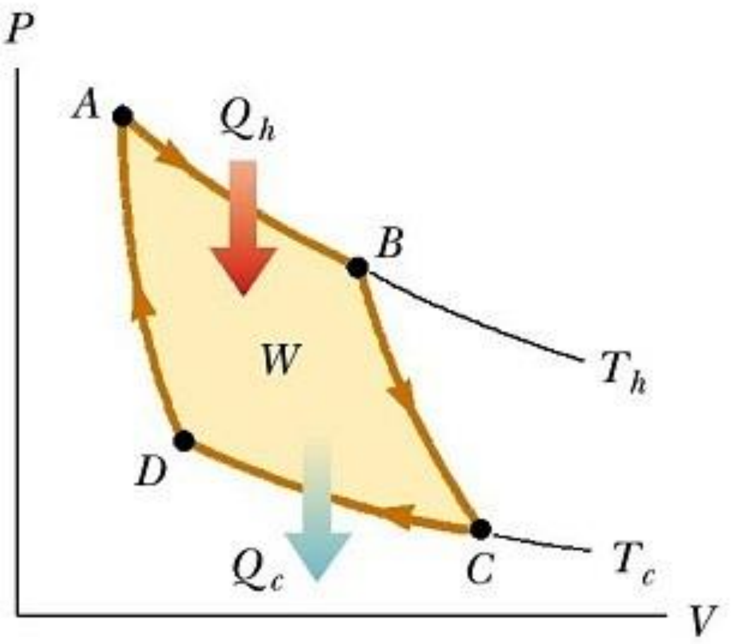


❑Definitions

➢**Cyclic Process:**

It is the process at which the system return to its initial state at its end.





•Represented by a closed loop on the PV diagram.

•The system may pass through a set of equilibrium states (such as A,B,...) during the cyclic process.

❑Definitions

6. Macroscopic scale:

•On the macro-scale the gas is considered a uniform, with properties that are averaged from the individual components.

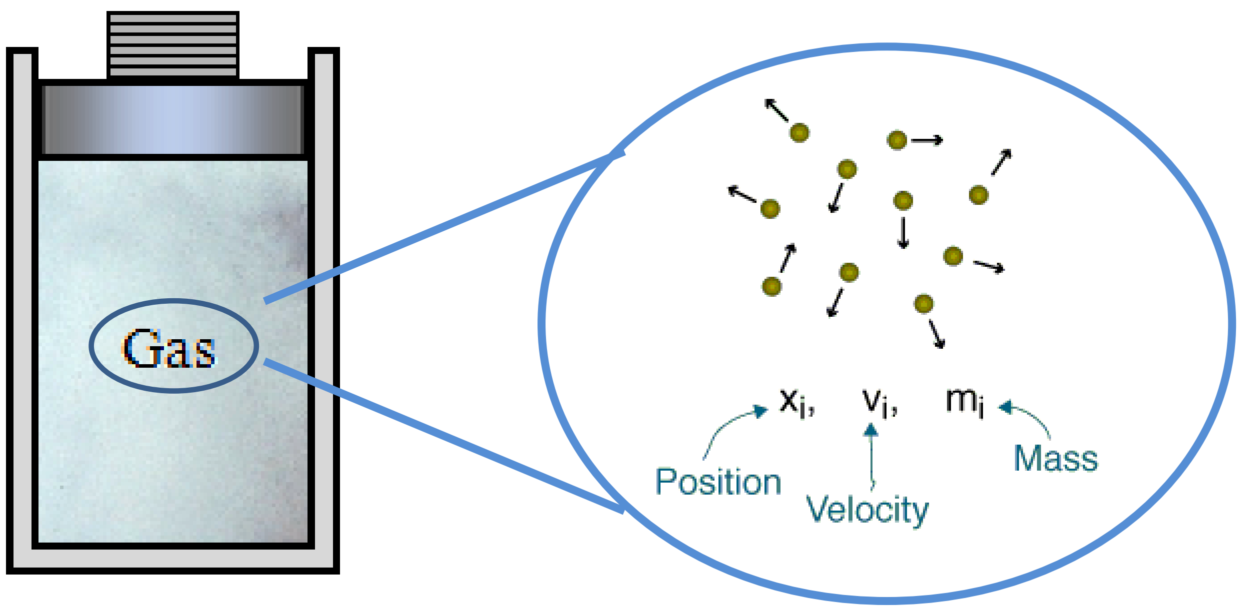
•On the macro-scale, we are dealing with large scale effects that we can measure, such as **pressure**, **temperature**, **flow velocity**,…

❑Definitions

7. Microscopic scale:

•On the micro-scale the gas is modeled by the kinetic theory with molecules size are very small relative to the distance between them.

•Simple theories were developed to relate the macroscopic properties (**pressure, temperature**,...) to the microscopic properties (**mass, momentum, kinetic energy**,..)

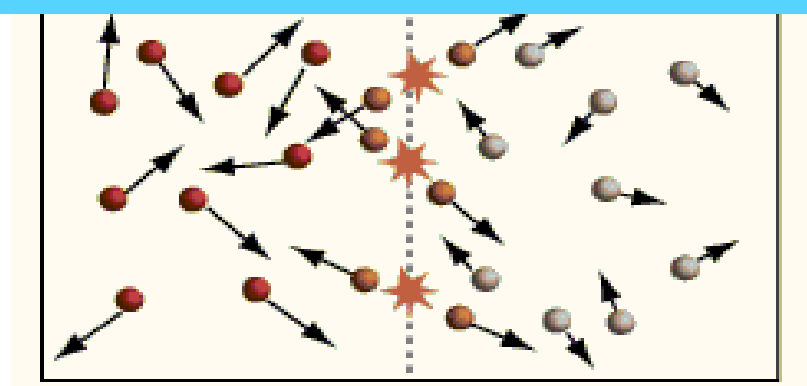


❑Temperature

➢**Temperature**

it is a measure of the system internal energy.

- Temperature is **Macroscopic property**.(Sense that one gas is hotter than another gas and therefore has a higher temperature)
- The temperature of a gas(**Microscopic property)** is a measure of the average translational kinetic energy of the gas molecules.



➢**In hot gas, the molecules move faster than the molecules of the cold gas.**

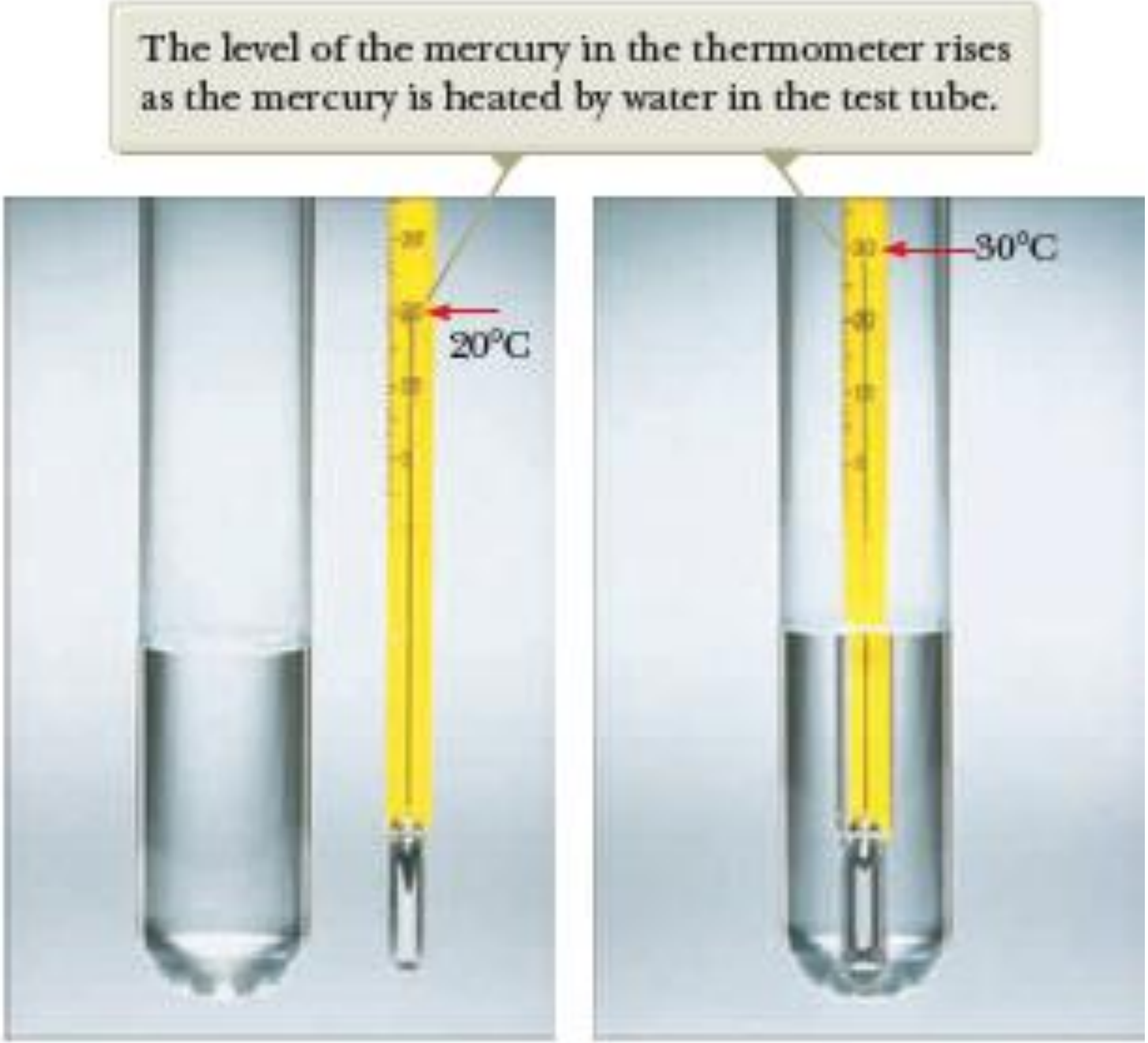
❑Temperature

➢ **Units of Temperature (Temperature scales)**

1. **Celsius Scale (°C)**
2. **Kelvin Scale (K)**
3. **Fahrenheit Scale (F)**
4. **Rankine Scale (°R)**

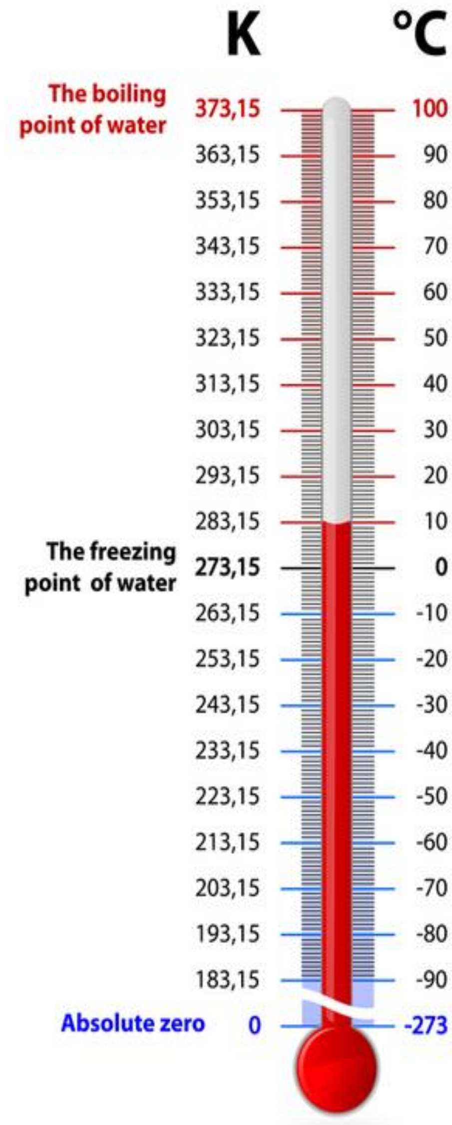
1. Celsius Scale (°C)

- Is the **CGS unit** of temperature
- Use the freezing point of pure water as the zero point (0 °C)
- Use the boiling point of the pure water as 100 °C.
- The scale between the two points was divided into 100 equal parts (Linear scale).



The level of the mercury in the thermometer rises as the mercury is heated by water in the test tube.
20°C
30°C

❑Temperature



2. Kelvin Scale (K) (Absolute Scale)

- Is the **SI unit** of temperature
- According to Kelvin scale,
- the triple point of water(**0.01 °C)**
- chosen as a fixed point and equals 273.16 K.

$$T_K = T_C + 273.15$$

➢**Absolute zero (0 K) or (-273.15 °C)** is where all kinetic motion in the particles in matter is minimum and matter contains no thermal energy.

➢**One degree increment on Kelvin scale has precisely the same magnitude as one degree increment on the Celsius scale.**

❑Temperature

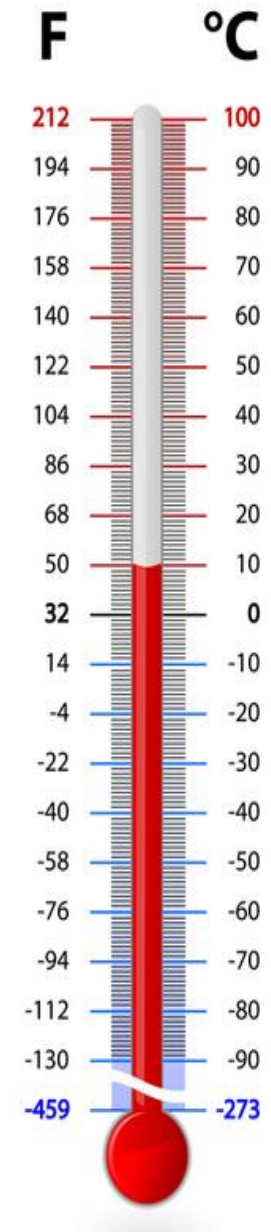
3. Fahrenheit Scale (°F)

- Is the unit of temperature in **the British system.**
- Use the freezing point of **sea water** as the zero point (°F) and the freezing point of pure water as the 32 °F.
- Use the boiling point of the pure water as 212 °F.
- The difference between the two reference points is 180 °F.

$$1\ degree\ {}^{o}F = 1\ degree\ {}^{o}C \times \frac{180}{100} = 1\ degree\ {}^{o}C \times \frac{9}{5}$$

$$T_F = 32 + \frac{9}{5}T_C$$

$$T_C = \frac{5}{9}(T_F - 32)$$



❑Temperature

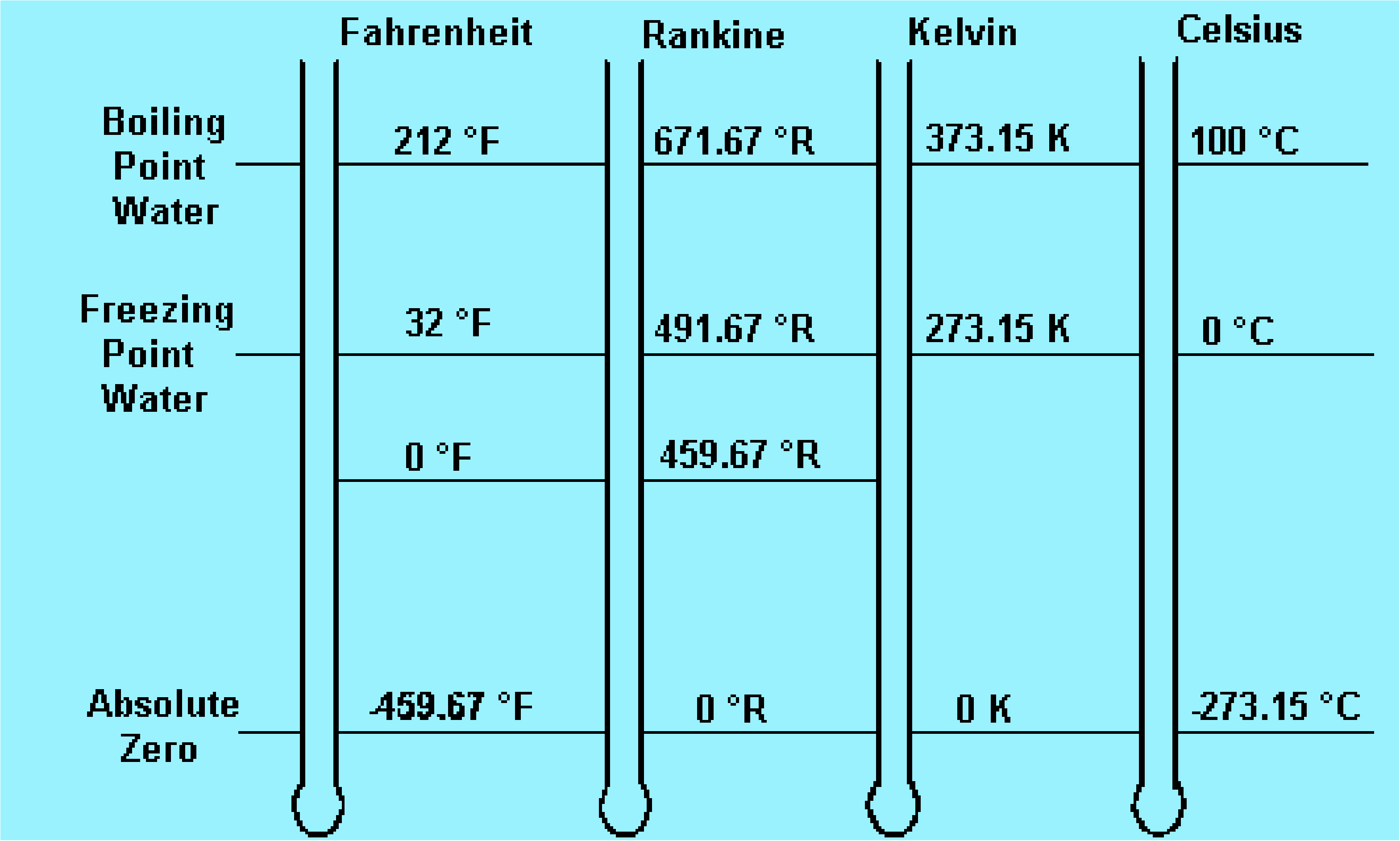
4. Rankine Scale (°R)

- A similar absolute zero corresponding to the Fahrenheit scale.
- Zero on both the Kelvin and Rankine scales is absolute zero,
- the Rankine degree is defined as equal to one degree Fahrenheit.

$$T_R = T_F + 459.67$$

$$T_R = (\frac{9}{5}T_C + 32) + 459.67 \Rightarrow T_C = \frac{5}{9}(T_R - 491.67)$$

❑Temperature



❑Temperature

Example (1)

The normal boiling point of liquid Oxygen is – 183 °C. what is this temperature on (a) Kelvin scale, (b) Rankine scale?

Solution

$T_C = -183°C$

(a) $$T_K = T_C + 273.15$$

$$T_K = (-183) + 273.15 = \mathbf{90.15\ K}$$

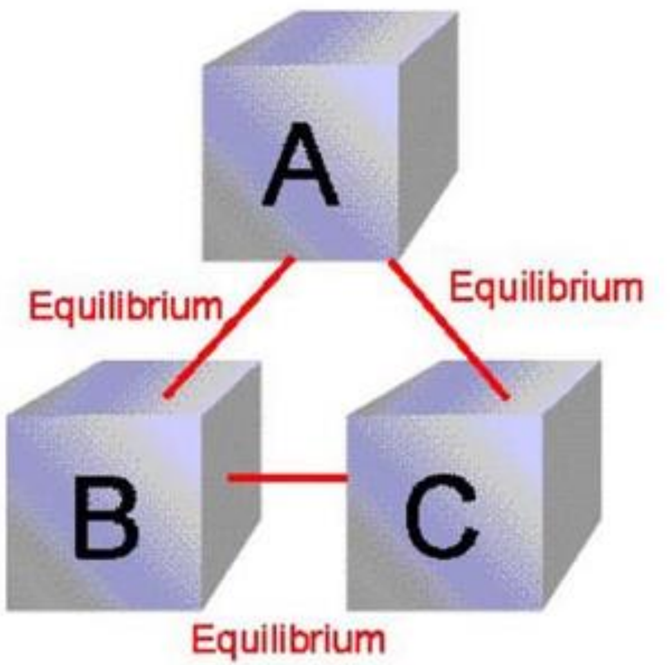
$$(b)\quad T_R = T_F + 459.69 = \frac{9}{5}T_C + 491.67$$

$$T_R = \frac{9}{5} \times -183 + 491.67 = \mathbf{162.27\ {}^{o}R}$$

❑The Zeroth law of thermodynamic

- This law states that:

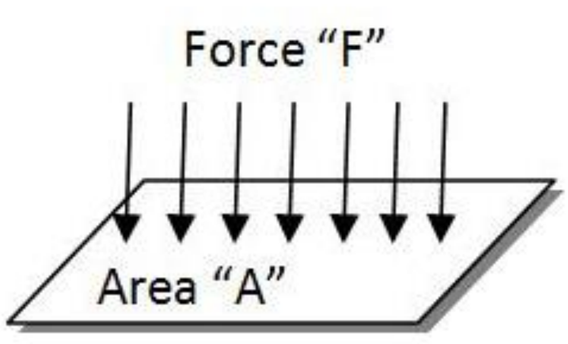
  " if object A is in thermal equilibrium with object B, and object B is in thermal equilibrium with object C, then object C is also in thermal equilibrium with object A".

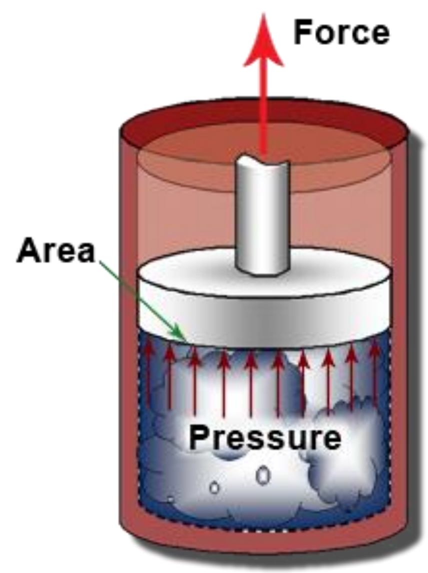


❑Pressure

The pressure of a system can be defined as" **the force exerted by the system on unit area of the boundaries".**

1. Mechanical pressure





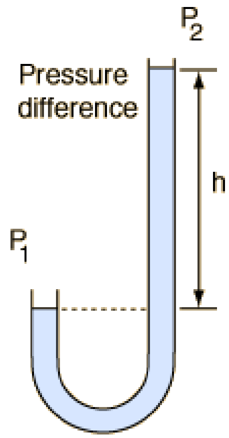
$$P = \frac{F_{\perp}}{A}$$

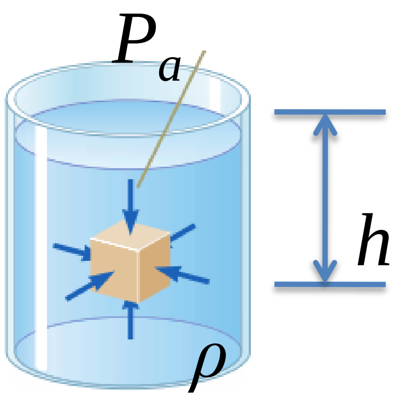
Pressure is a scalar quantity since it proportional to the magnitude of the force.

❑Pressure

2. Hydrostatic pressure

The pressure due to a column of a liquid





$$\Delta P = P_1 - P_2 = \rho g h$$

$$P = P_a + \rho g h$$

➢If the pressure varies over an area, the infinitesimal force dF on an infinitesimal surface element of area dA is

$$P = \frac{dF}{dA} \quad or \; dF = PdA$$

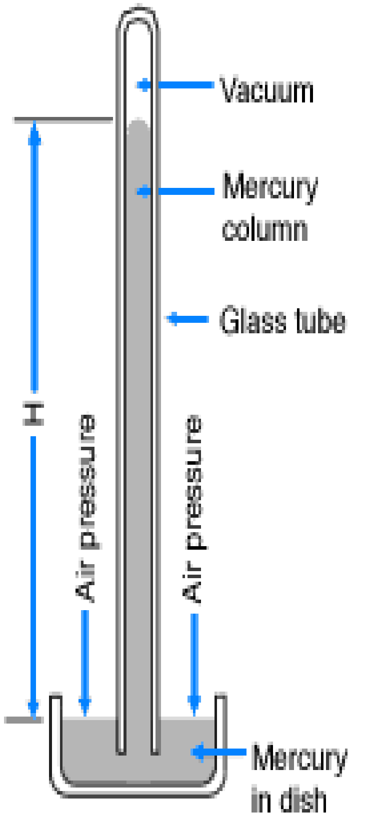
Pressure

A normal force exerted by a fluid per unit area.

$$P = \frac{F}{A} \qquad N/m^2$$

Pressure units:-

Pascal (Pa) = N/m^2 Bar $=10^5$ N/m^2



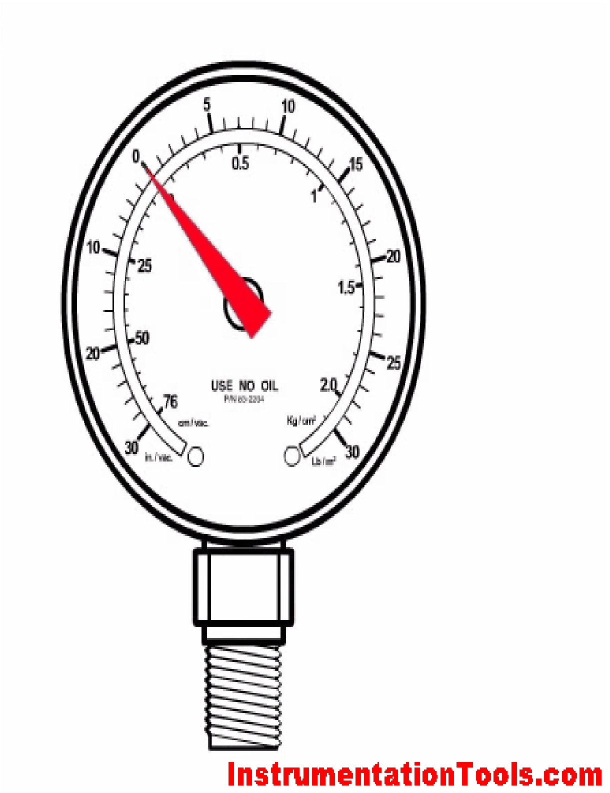
Hydrostatic pressure

Atmospheric pressure

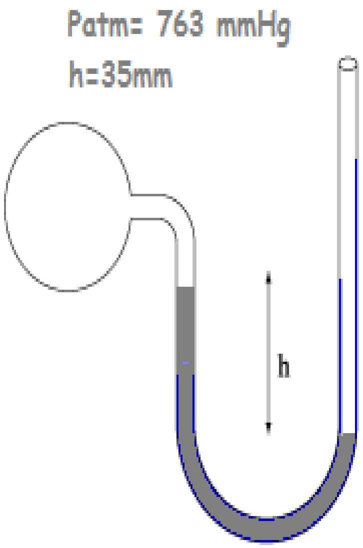
1 Atm = 101325 N/m^2 = 14.7 psi = 76 cm Hg = 760 mm Hg = 760 torr

Pressure measuring instruments

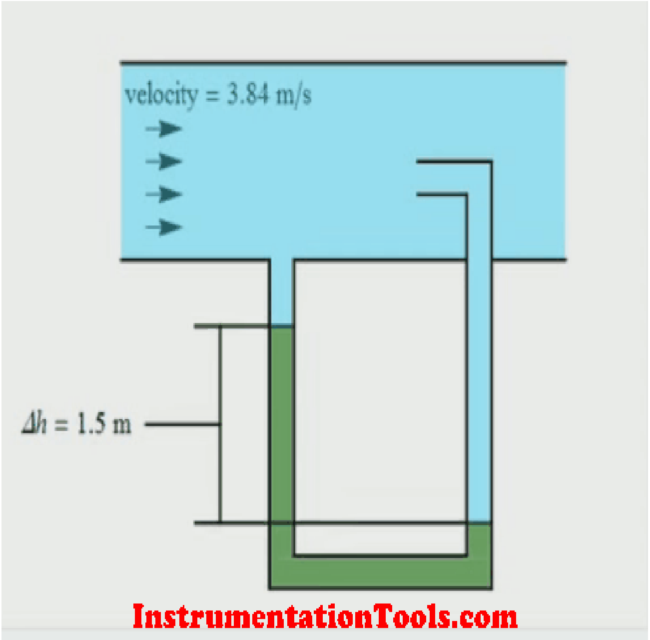
Dynamic pressure



Borden tube



Manometer



pitot tube

❑Pressure

➢**Gauge Pressure**

The pressure difference between the system pressure and the atmospheric pressure.

$$P_{gauge} = P_{system} - P_a$$

Absolute pressure of a system

$$P_{system} = P_{gauge} + P_a$$

❑Pressure

$$P = \frac{F_{\perp}}{A}$$

➢ **Units of pressure**

- **SI unit** → N/ m^2 = Pascal (Pa)
- **CGS unit** → dyne/ cm^2
- **British system** → (lb/in^2 = psi)

Units	Pa	psi	atm	bar	torr
Pa	$1 N/m^2$	1.45×10^{-4}	9.869×10^{-6}	10^{-5}	7.5×10^{-3}
psi	6.894×10^{3}	$1\ lb/in^2$	6.8×10^{-2}	6.894×10^{-2}	51.714
atm	1.01325×10^{5}	14.695	P_0	1.01325	760
bar	10^{5}	14.5	0.9869	10^{6} $dyne/cm^2$	750
torr	133.322	1.93×10^{-2}	1.315×10^{-3}	1.333×10^{-3}	1 mmHg

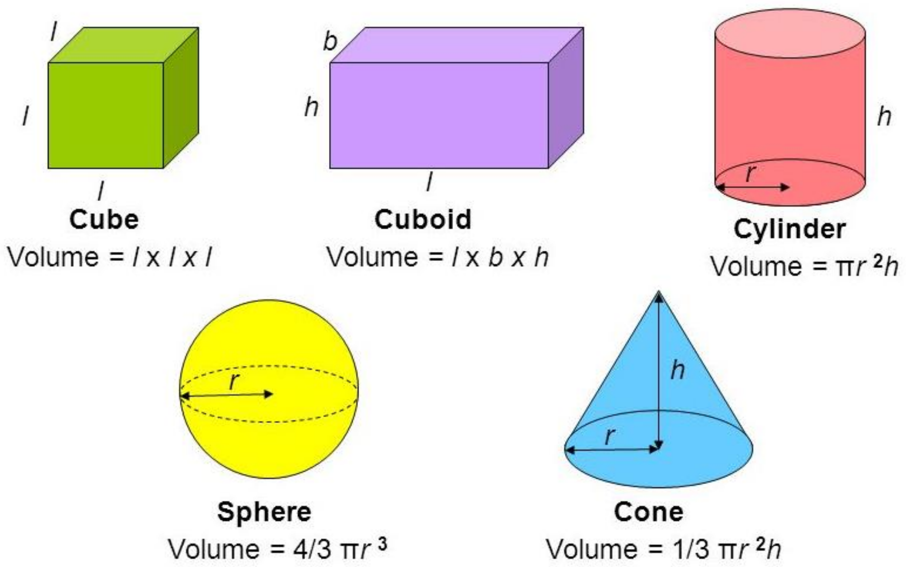
❑Volume

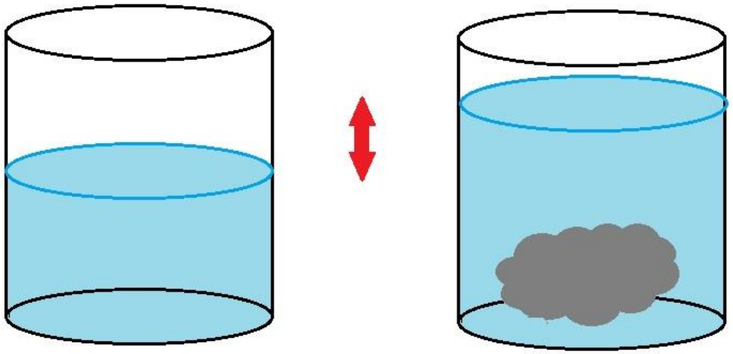
The volume of a solid object is:

➢The three-dimensional space the body occupies.

or

➢The quantity of three-dimensional space enclosed by a closed surface.





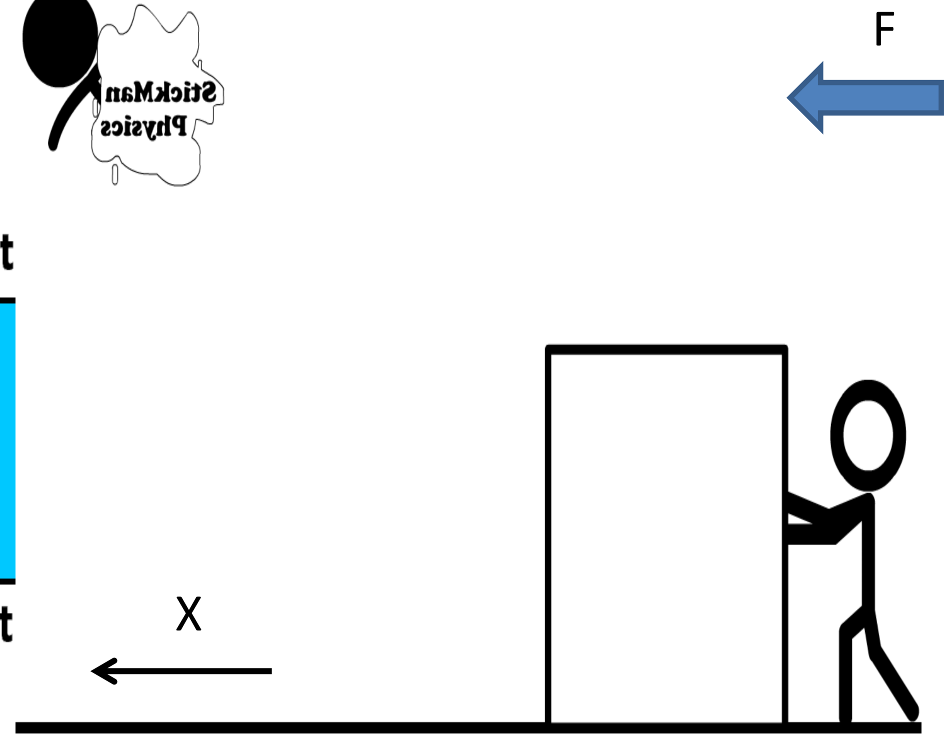
❑Volume

- **Common Units of Volume and Capacity:**
  - SI unit → m^3
  - CGS unit→ $cm^3 = 10^{-6}\ m^3$
  - British system→ cubic inch or cubic foot
  - **Liter= 1000 cm^3= 10^{-3} m^3**

WORK

Energy transfer associated with a force acting through a distance.

$W = F \cdot X$

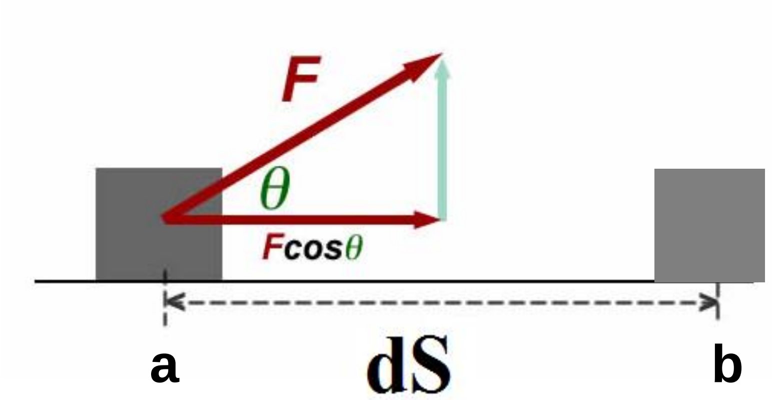


Energy transfer , not due to mass transfer across system boundary

❑Work

- **The work (W) done on a body can be defined as:**
  The product of the force **F** acting on the body through a distance **S**.

$$W = \int_a^b F \cdot dS = \int_a^b F \, dS \cos\theta$$



- **The SI unit of the work is Joule (J):**

  1 Joule = N.m

- **The C.G.S. unit of the work is erg:**

  1 erg= dyne. cm= 10^{-7} Joule

❑Work

➢ Consider a thermodynamic systems such as a gas contained in a cylinder fitted with a movable piston of a cross section area A

The force exerted by the gas molecules on the piston is:

$$F = PA$$

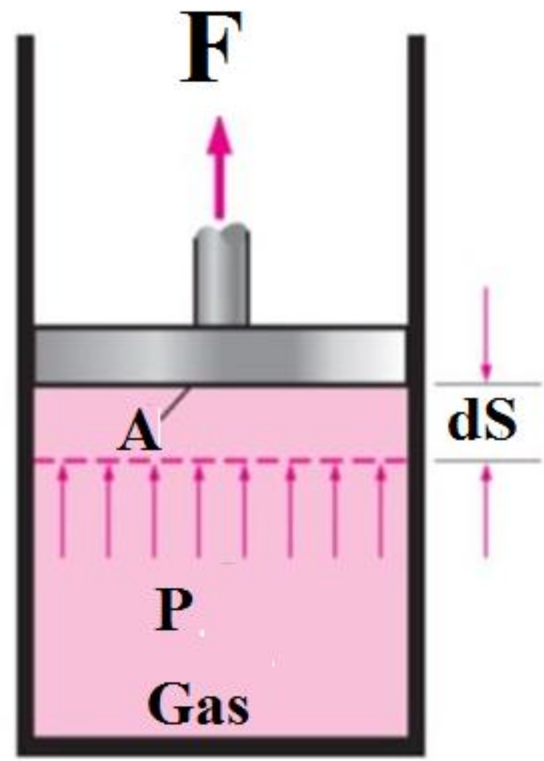
If this pressure causes the piston to move a distance dS

$$\therefore dW = FdS = (PA)dS$$

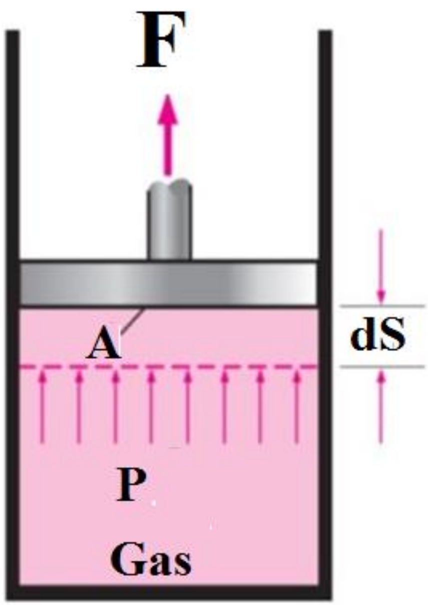
$$dW = P(AdS) = PdV$$

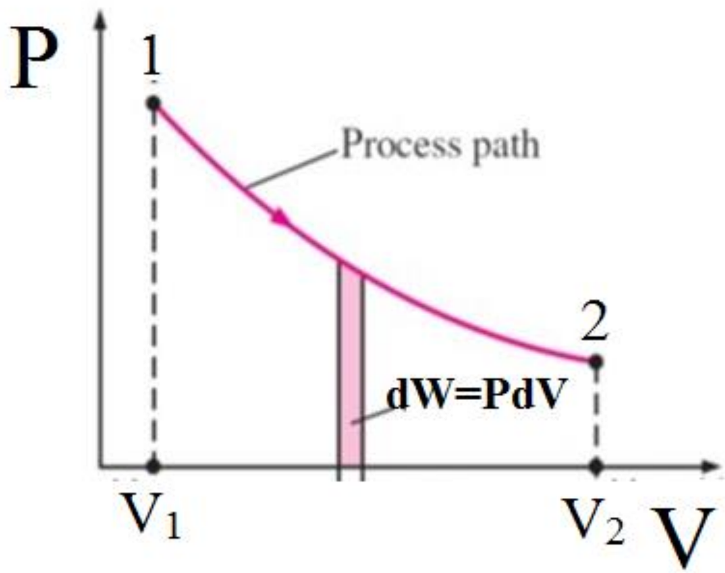
The total work if the volume changes from $V_1 \rightarrow V_2$

$$W = \int_{V_1}^{V_2} PdV = P\Delta V$$



❑Work



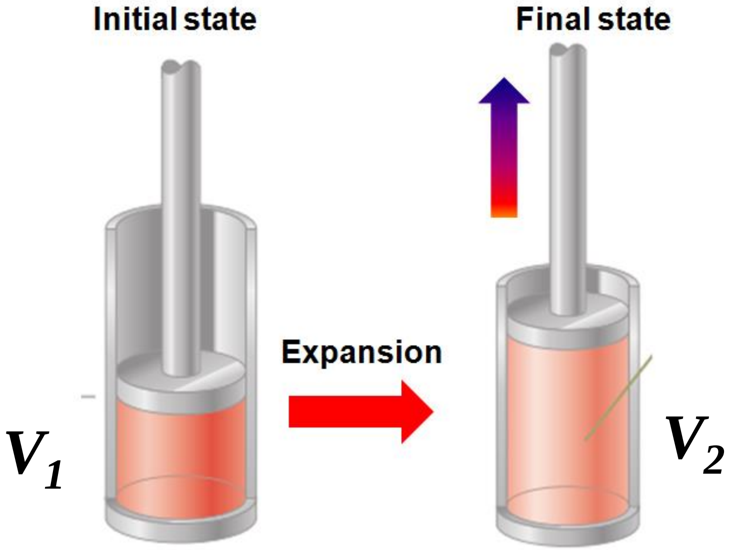


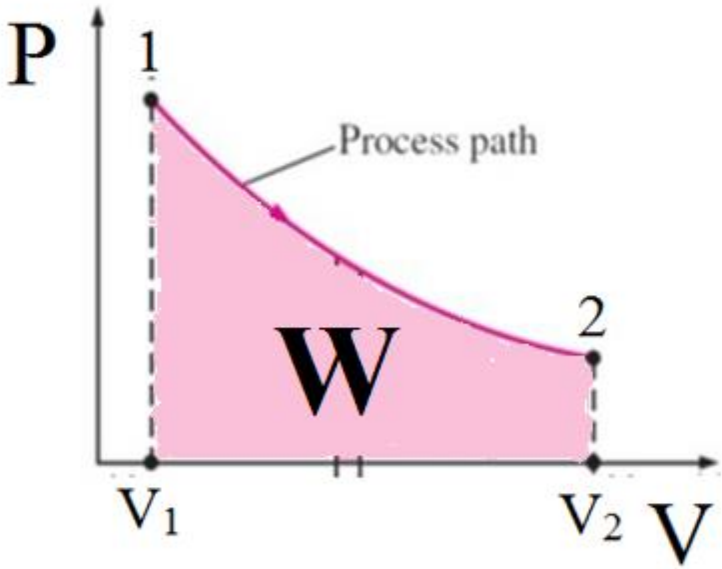
$$dW = P(AdS) = PdV$$

❑Work

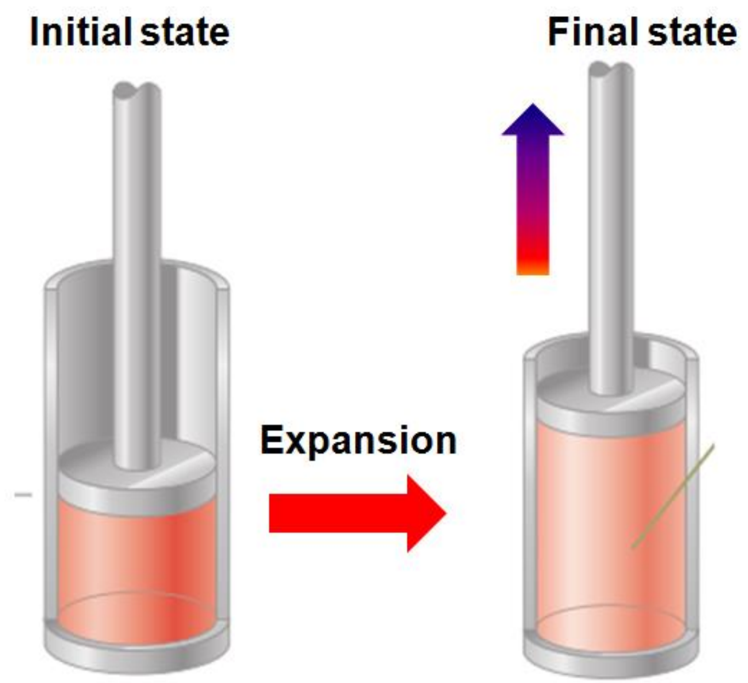
$$W = \int_{V_1}^{V_2} PdV = PV|_{V_1}^{V_2} = P(V_2 - V_1)$$

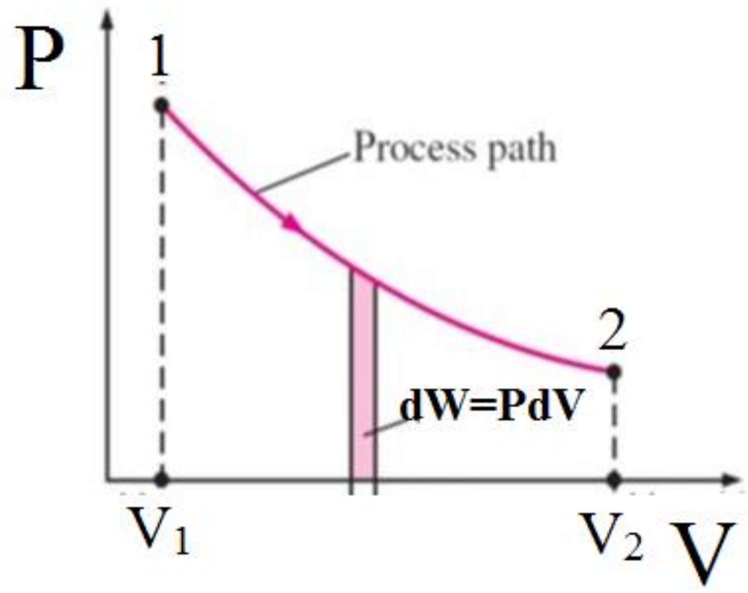
$$W = \int_{V_1}^{V_2} PdV = P\Delta V$$





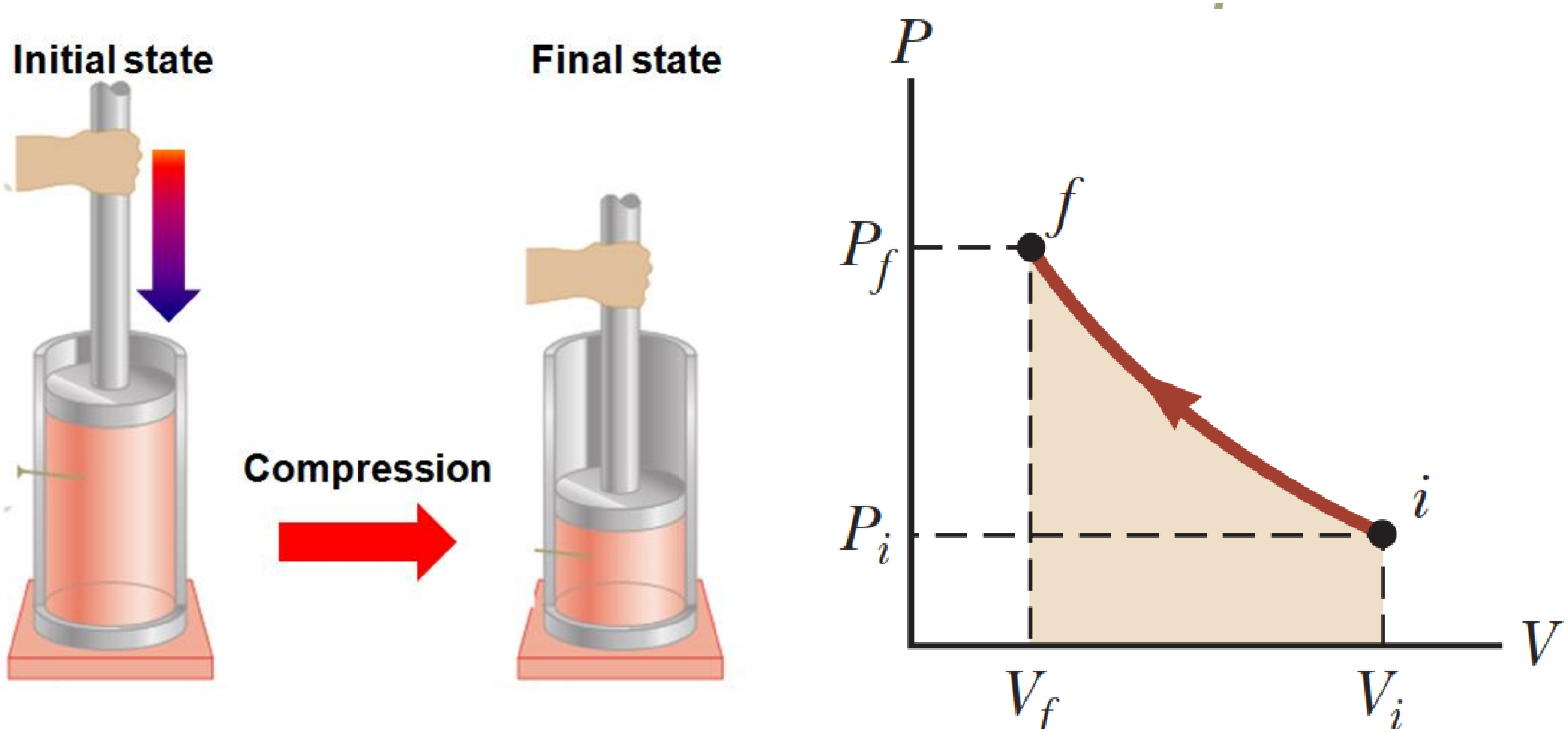
❑Work





$$W = +ve \rightarrow work\ done\ \mathbf{by}\ the\ system$$

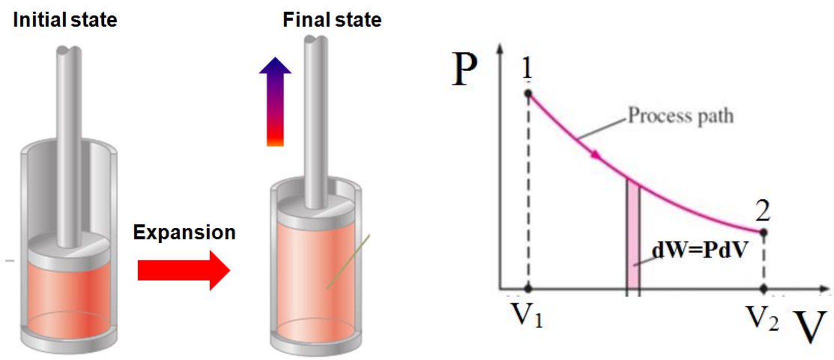
❑Work



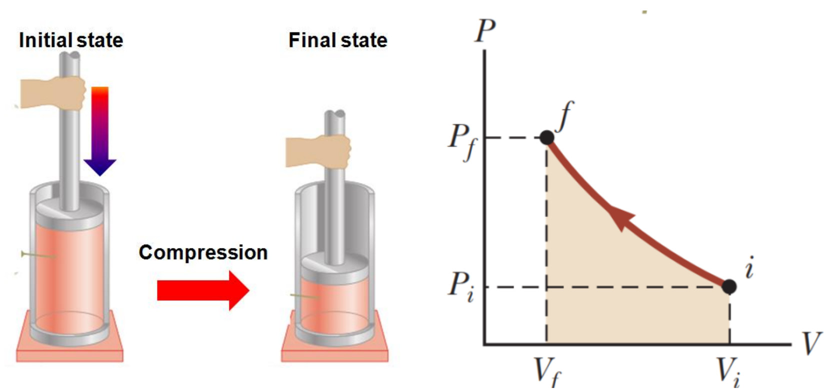
$$W = -\boldsymbol{ve} \rightarrow work\ done\ \boldsymbol{on}\ the\ system$$

❑Work

Work Sign Rule



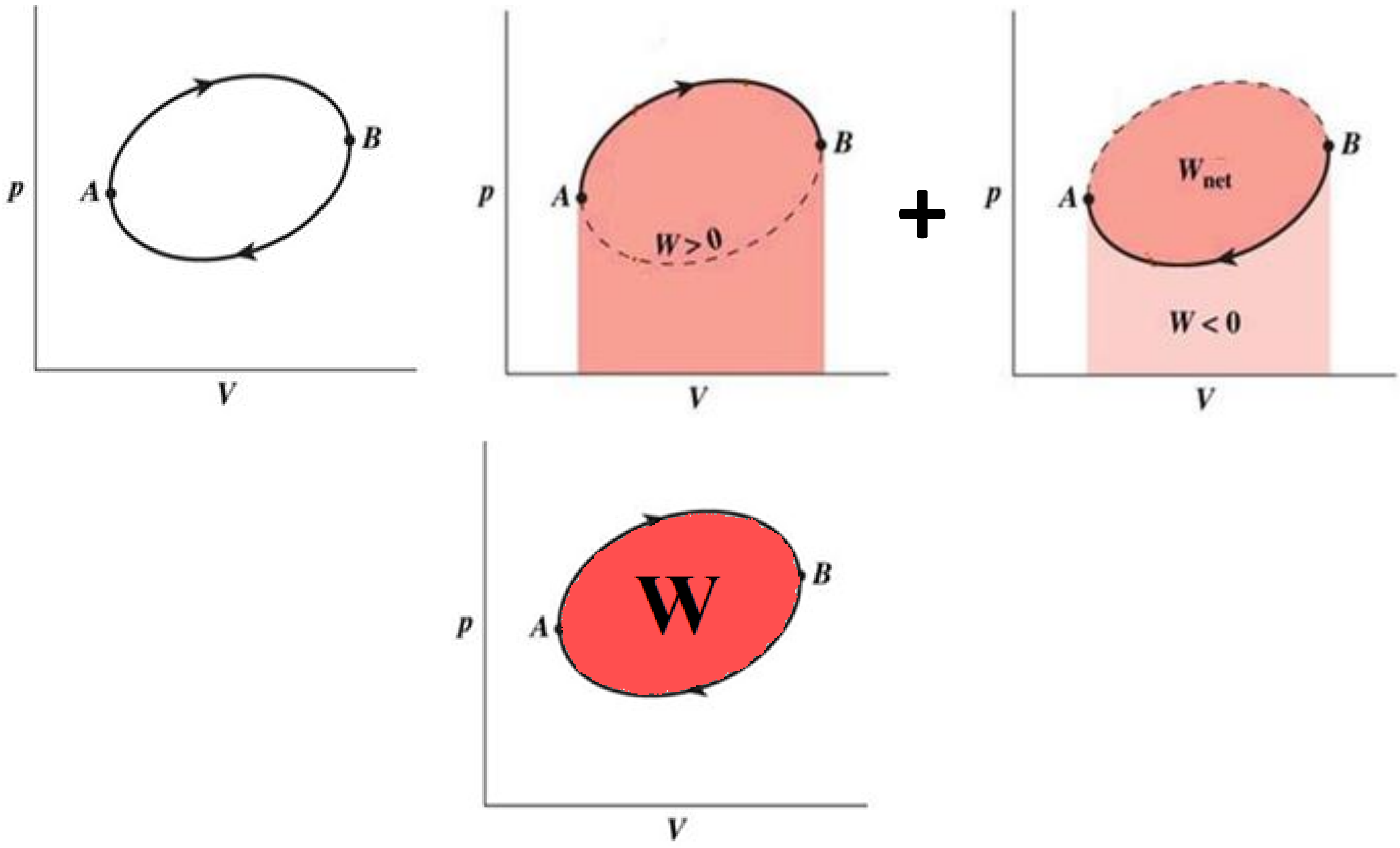
$$W = +\boldsymbol{ve} \rightarrow \textit{work done } \textbf{by} \textit{ the system}$$



$$W = -\boldsymbol{ve} \rightarrow \textit{work done } \textbf{on} \textit{ the system}$$

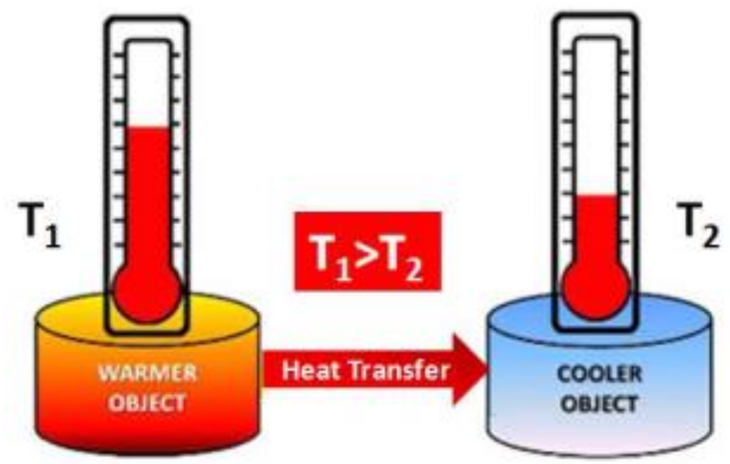
❑Work

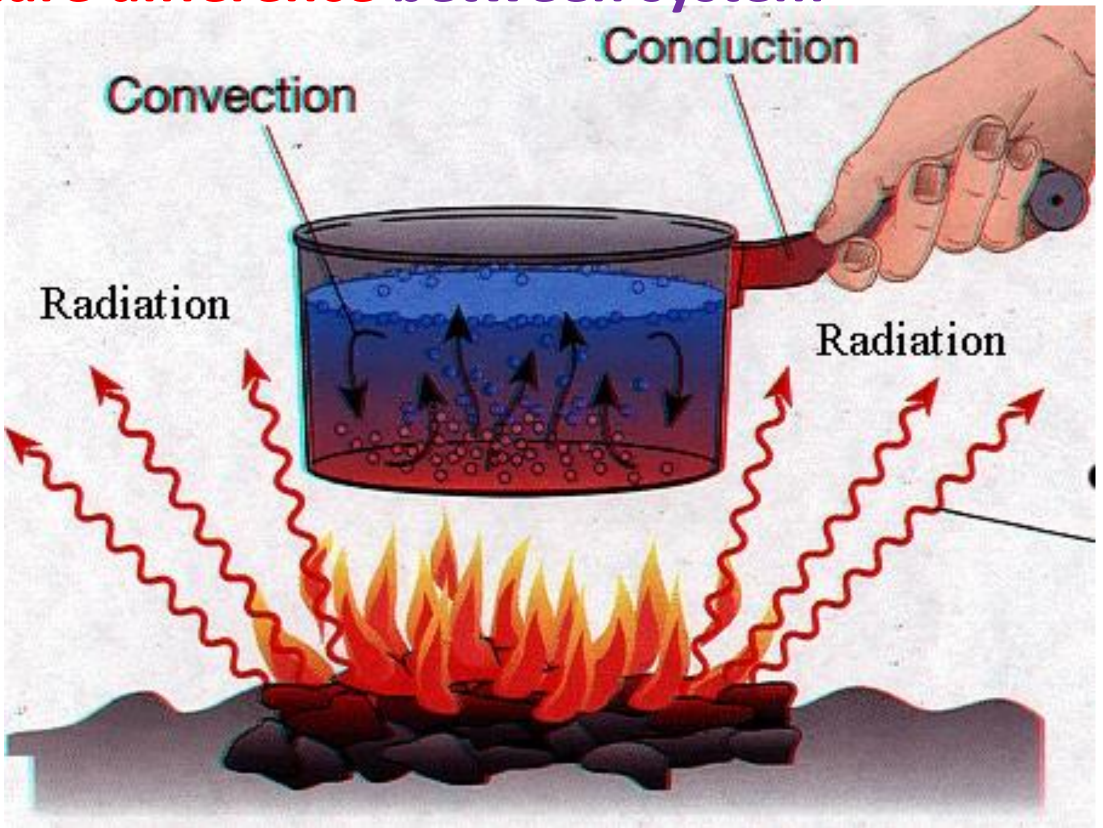
➢ **Net Work from Cyclic Process:**



❑Heat

Energy transfer , not due to mass transfer across system boundary but due to temperature difference between system and surrounding.





➢ **Heat transfer occurs basically in three ways:**

1. Conduction
2. Convection and
3. Radiation

❑Heat

➢ **Unit of Heat:**

1. **SI Unit → Joule (J)**

2. **C.G.S Unit → calorie (cal)**

calorie (cal): the amount of heat energy needed to raise the temperature of **1 g** of water **1 °C (**from **14.5°C** to **15.5°C)**.

$$1\,cal = 4.185\,J$$

$$1\,Joule = \frac{1}{4.185}\,cal = 0.239\,cal$$

❑Heat

Unit of Heat:

1. SI Unit → Joule (J)
2. C.G.S Unit → calorie (cal)
3. British Unit → British Thermal Unit (BTU)

calorie (cal):
the amount of heat energy needed to raise the temperature of **1 g** of water **1 °C** (from **14.5°C** to **15.5°C**).

$$1\ cal = 4.185\ J$$

→

$$1\ Joule = \frac{1}{4.185}\ cal = 0.239\ cal$$

Kilocalorie (Cal):
the amount of heat energy needed to raise the temperature of 1 Kg of water 1 °C (from 14.5°C to 15.5°C).

$$1\ kilocalorie\ (\boldsymbol{Cal\ or\ Kcal}) = 1000\ cal$$

→

$$1\ Cal = 4185\ J$$

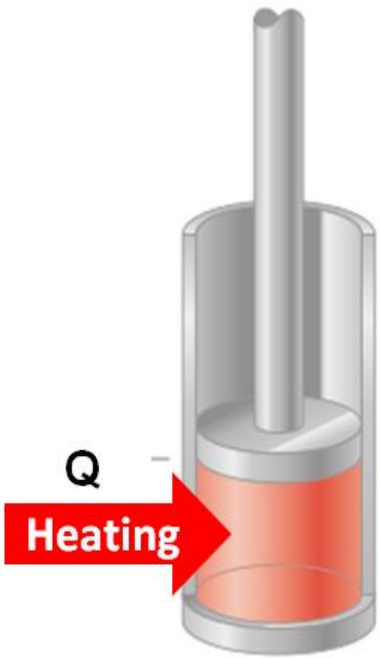
❑Heat

➢Unit of Heat:

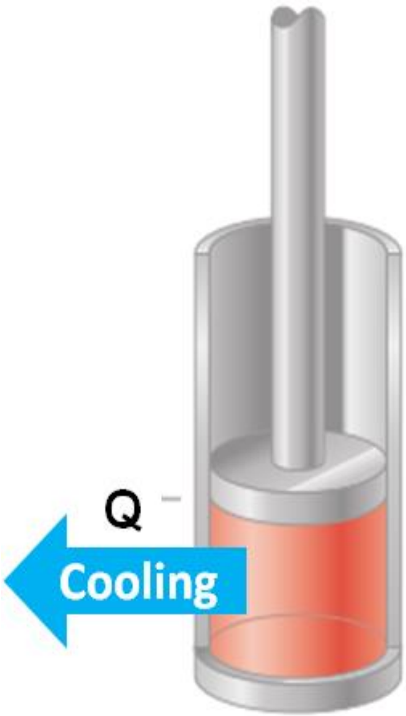
	Joule	Calorie	BTU	Kilowatt-hour	Electronvolt
Joule	XX	0.2390	0.000948	2.77778E-07	6.2383E+18
Calorie	4.184	XX	0.00397	1.16279E-06	2.61097E+19
BTU	1055	252	XX	0.000293	6.57895E+21
Kilowatt-hour	3.6E6	8.6E5	3412	XX	2.24719E+25
Electronvolt	1.603E-19	3.83E-20	1.52E-22	4.45E-26	XX

❑Heat

Heat Sign Rule



$Q = +\boldsymbol{ve} \rightarrow Heat\ \boldsymbol{added\ to}\ the\ system$



$Q = -\boldsymbol{ve} \rightarrow Heat\ \boldsymbol{removed\ from}\ the\ system$

❑Heat

➢**The quantity of heat (Q) produces a change in the temperature (ΔT) of a body:**

$$Q = \int_{T_1}^{T_2} C\,dT = C\Delta T$$

Where

$C \rightarrow \boldsymbol{Heat\ capacity}\ of\ the\ body$

Heat Capacity (C):

It is the amount of heat needed to change the temperature of the substance (**system**) by **1 ºC.**

$$C = \frac{Q}{\Delta T} \rightarrow \begin{cases} \dfrac{Joule}{K} & (SI\ Unit) \\ \dfrac{cal}{^{o}C} & (cgs\ Unit) \\ \dfrac{BTU}{^{o}F} & (British\ Unit) \end{cases}$$

❑Heat

Specific Heat (s or c):

The specific heat c of a substance is the heat capacity per **unit mass.**

The specific heat of a substance is the amount of heat needed to change the temperature of **1 Kg** of the substance by **1 °C.**

$$c = \frac{C}{m} \rightarrow C = mc$$

$$Q = mc\Delta T$$

Specific Heats of Some Substances at 25°C and Atmospheric Pressure

Substance	Specific Heat (cal/g °C)	Specific Heat (J/g °C)
Water	1.00	4.18
Ethanol	0.58	2.4
Aluminum	0.22	0.92
Sand	0.19	0.79
Iron	0.11	0.46
Copper	0.093	0.39
Silver	0.057	0.24
Gold	0.031	0.13

$\frac{BTU}{lb.^{o}F}$

16

$$c = \frac{Q}{m\Delta T} \rightarrow \begin{cases} \frac{Joule}{Kg.K} \ (SI\ Unit) \\ \frac{cal}{g.^{o}C} \ (cgs\ Unit) \\ \frac{BTU}{lb.^{o}F} \ (British\ Unit) \end{cases}$$

❑Heat

Molar Specific Heat (s or c):

It is the amount of heat needed to increase the temperature of **1 mole** of the substance by **1 °C**.

$$Q = nc\Delta T$$

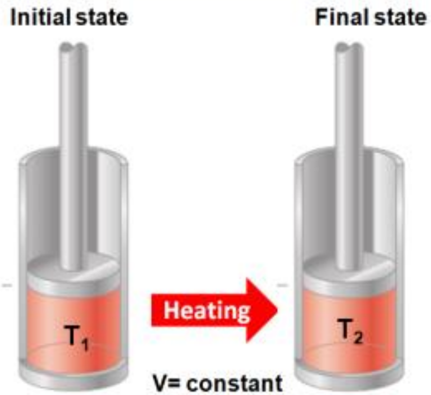
$n \rightarrow number\ of\ moles$

$$c = \frac{Joule}{mole.K}$$

❑Heat

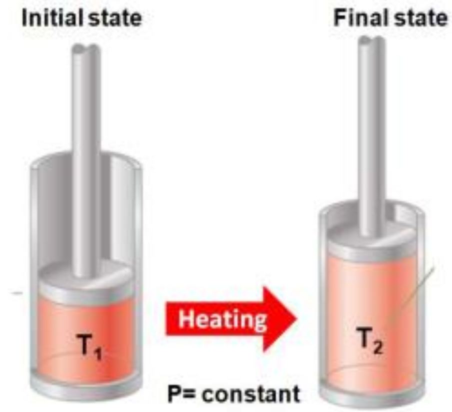
➢ For **gases** and due to the compressibility of there are **two types** of molar specific heat:-

1. **Molar specific heat at constant volume (C_V)** is the amount of heat needed to increase the temperature of **1** mole of the gas by 1 °C at **constant volume.**



$$Q = nc_V \Delta T$$

2. **Molar specific heat at constant pressure (C_P)** is the amount of heat needed to increase the temperature of **1** mole of the gas by 1 °C at **constant pressure.**



$$Q = nc_P \Delta T$$

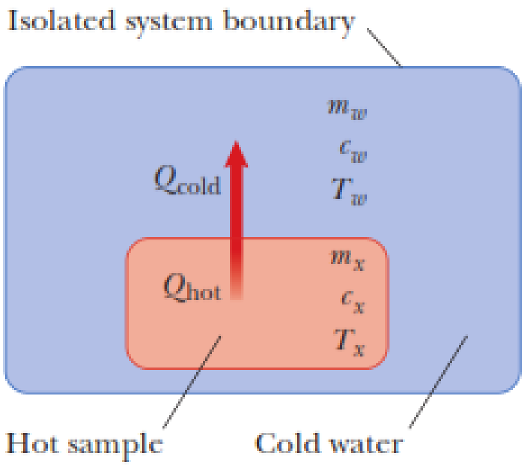
18

Example (4)

A 0.05 kg ingot of metal is heated to 200.0°C and then dropped into a calorimeter containing 0.400 kg of water initially at 20.0°C. The final equilibrium temperature of the mixed system is 22.4°C. Find the specific heat of the metal.

Solution

$m_x = 0.05$ kg, $T_x = 200$ °C, $m_w = 0.4$ kg, $T_w = 20$°C, $T_f = 22.4$ °C, $c_x = ??$



$$\Delta Q_{lost} = \Delta Q_{gained}$$

$$-m_x c_x \Delta T = m_w c_w \Delta T$$

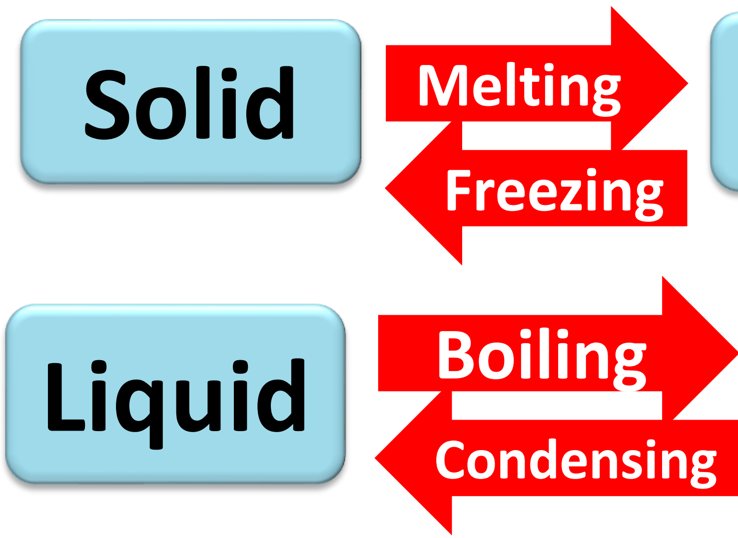
$$-m_x c_x (T_f - T_{x)} = m_w c_w (T_f - T_{w)}$$

$$-0.05 \times c_x (22.4 - 200) = 0.04 \times 4186 (22.4 - 20)$$

$$c_x = 453 \frac{J}{Kg.K}$$

❑Heat of Transformation

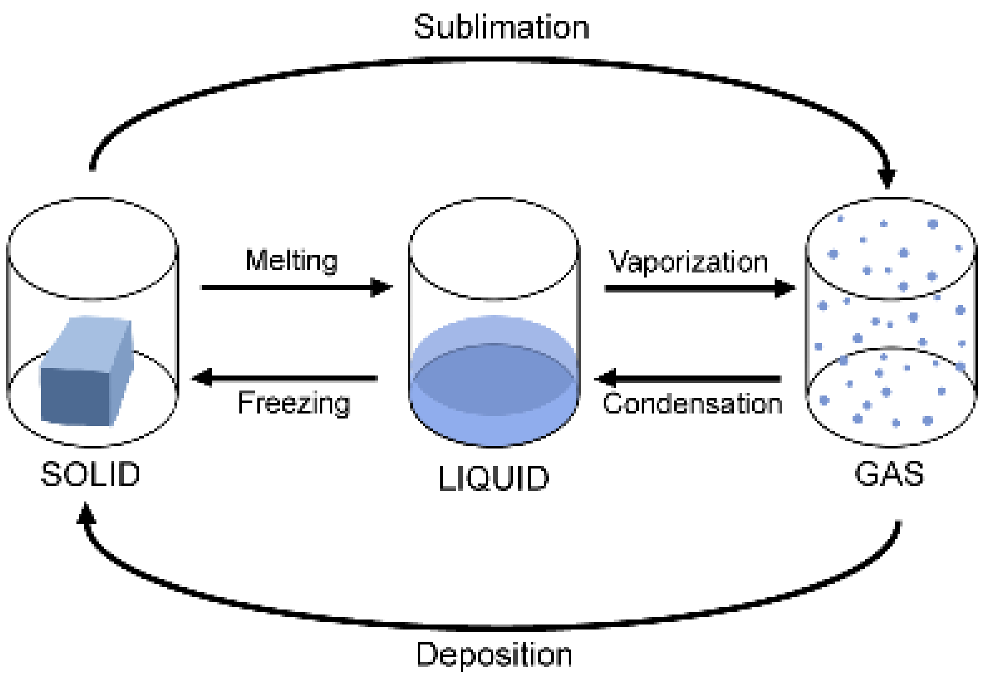
In some situations, the transfer of energy does not result in a change in temperature → **Phase change**

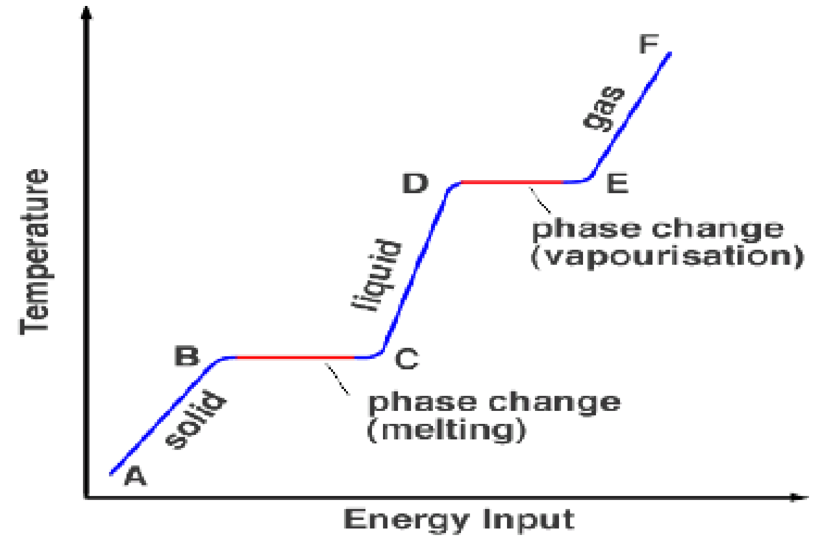


All phase change processes involve a change in the system's internal energy but no change in its temperature.

For example, The increase in internal energy in boiling process , is represented by the breaking of bonds between molecules in the liquid state.

❑Heat of Transformation





Phase change diagram.

$$Q \propto m \rightarrow Q = mL$$

$L \rightarrow Latent\ Heat\ or\ Heat\ of\ transformation$

❑Heat of Transformation

Latent heat of Fusion (L_F) of a substance:

Is the amount of heat must be supplied to change **1 Kg** of the substance at its melting point from **solid** to **liquid.**

The same amount of heat must be removed from 1 kg of the substance to change it from liquid to solid

Latent heat of Vaporization (L_V) of a substance:

Is the amount of heat must be supplied to change **1 Kg** of the substance at its boiling point from **liquid** to **vapor.**

The same amount of heat must be removed from 1 kg of the substance to change it from gas to liquid

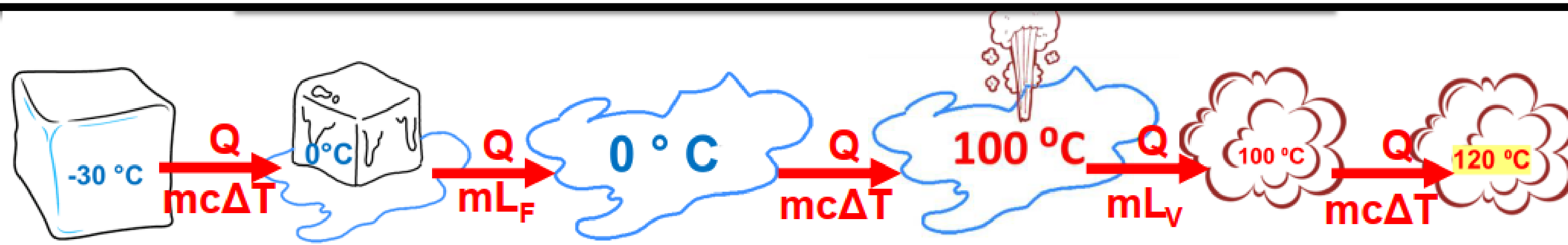
❑Heat of Transformation

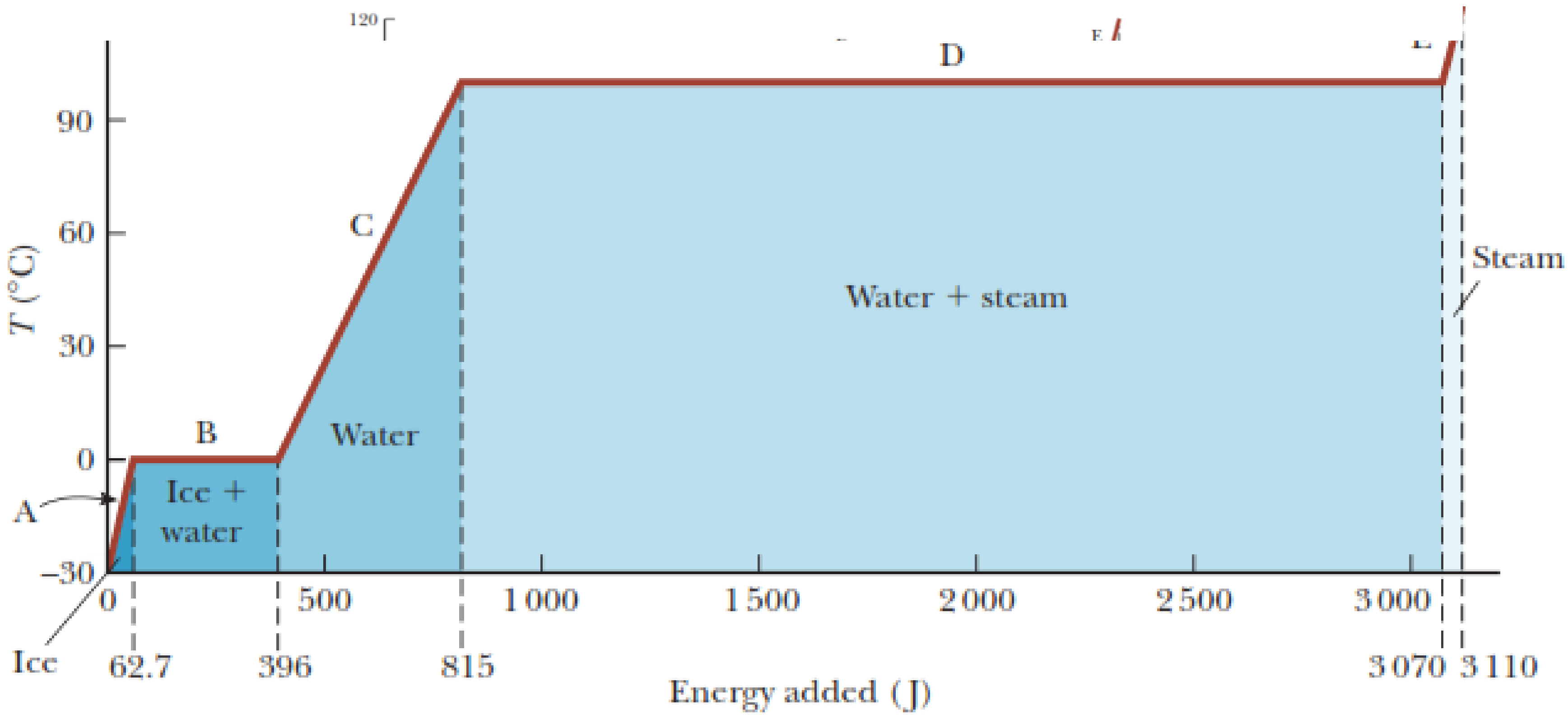
Latent Heats of Fusion and Vaporization

Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium	−269.65	5.23×10^3	−268.93	2.09×10^4
Nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
Oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
Ethyl alcohol	−114	1.04×10^5	78	8.54×10^5
Water	0.00	3.33×10^5	100.00	2.26×10^6
Sulfur	119	3.81×10^4	444.60	3.26×10^5
Lead	327.3	2.45×10^4	1 750	8.70×10^5
Aluminum	660	3.97×10^5	2 450	1.14×10^7
Silver	960.80	8.82×10^4	2 193	2.33×10^6
Gold	1 063.00	6.44×10^4	2 660	1.58×10^6
Copper	1 083	1.34×10^5	1 187	5.06×10^6

Example

What is the amount of energy required to convert a system consisting of a 1.00-g cube of ice at -30.0°C to steam at 120.0°C?





❑Internal Energy

The internal energy of a thermodynamic system (U)

Is the total kinetic energy and potential energy of the atoms and molecules consisting the system.

$$U = K.E. + P.E.$$

Due to the motion of molecules (translational, rotational, vibrational)

Due to binding energy of chemical bonding between molecules.

The Kinetic energy is due to the motion of molecules (translational, rotational, vibrational)

The potential energy is associated with electric potential energy of atoms, chemical bonding.

❑Internal Energy

Any thermodynamic system or body at any temperature includes internal energy

$$U = U(T)$$

$$U \approx 0 \; if \; T = 0 \; K \; (absolute \; temperature)$$

So, the internal energy of any thermodynamic system is usually expressed as a change (ΔU)

Example (5)

Electric current 3A passing through a heater at a potential difference 220V during a time 15 minutes. Calculate the energy lost by the electric energy in (a) calorie and (b) KWh. If this energy is used to heat 30 liter of water, (c) How much the water temperature increased?

Solution

I= 3 A, V= 220V, t= 15 min= 900 s, V= 30 lit → m_w=30 kg, Q=?? , ΔT=??

$$Electric\ power = VI$$

$$Energy = VIt = 220 \times 3 \times 900 = 594\ KJ$$

$$(a)\ Q = \frac{594 \times 10^3}{4.18} = 1421052.6\ cal = 142.1\ Kcal$$

$$(b)\ Q = \frac{594 \times 10^3}{3.6 \times 10^6} = 0.165\ KWh \qquad (KWh = 3.6 \times 10^6\ J)$$

$(c)\ Q = mc\Delta T$

$$1421052.6 = (30 \times 10^3)(1)\Delta T$$

$$\boxed{\Delta T = 4.7\ ^\circ C}$$