THERMAL PHYSICS
P6

Thermodynamics

OBJECTIVES

- Importance of temperature in the transfer of heat, & Introduce heat as energy in transit.
- Describe how internal energy changes due to absorption or emission of heat and how it relates to the temperature of a gaseous system.
- State and explain the first law of thermodynamics.
- Discuss + and – signs of the quantities ΔQ, ΔU and ΔW
- Describe constant pressure process and show that work done in it is $p. \Delta V$
- Describe constant volume process and show that work done in it is zero.
- Explain isothermal process and show that $\Delta U= 0$ for a gaseous system.
- Explain adiabatic process and point out $\Delta Q= 0$ for it.
- Discuss the temperature changes taken place in adiabatic compression and expansion of a gas.
- Introduce p-V curve and discuss the shapes it can be take.
- Show that ΔW is given by the area under the p-V curve.
- Describe the cyclic process. & Explain how ΔW can be found for a cyclic process.

THERMODYNAMICS (flow of heat)

Thermodynamics is a branch of science that deals with the study of inter conversion of heat with other forms of energy during physical and chemical process

First Law of Thermodynamics

The net heat put into a system is equal to the change in internal energy of the system plus the work done by the system.

Basically, the Law of Conservation of Energy restated.

If we add heat energy to a system, the added energy does one or both of two things:

- increases the internal energy of the system if it remains in the system
- does external work if it leaves the system

So, the first law of thermodynamics states:

Heat added = increase in internal energy + external work done by the system

First Law of Thermodynamics

- $Q = U + W$
  - Q = Heat added to a system
  - U = Increase in internal energy
  - W = Work done by (+) or on (-) the system
- Always measured in Joules

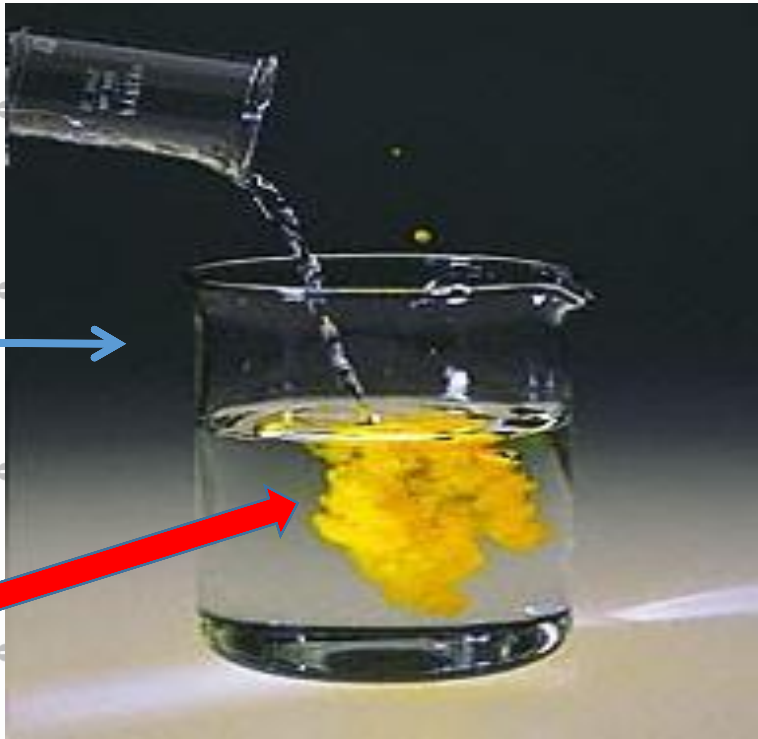
Thermodynamic terms

System It is a specified portion of the universe which is under thermodynamic study and which is separated from the rest of the universe with a definite boundary.

Surrounding

It is the portion of the universe excluding the system and capable of exchanging matter and energy with the system

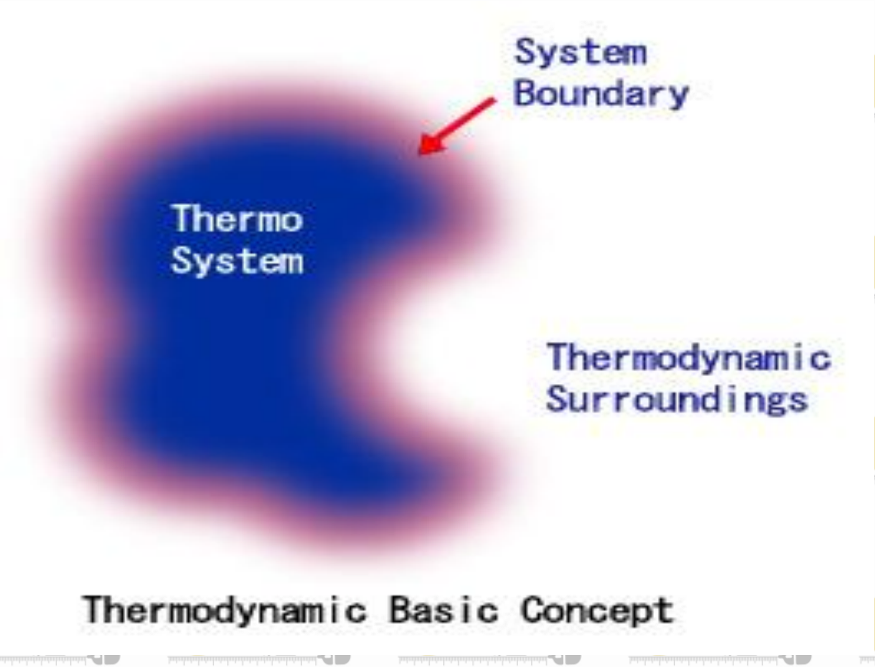
Surrounding

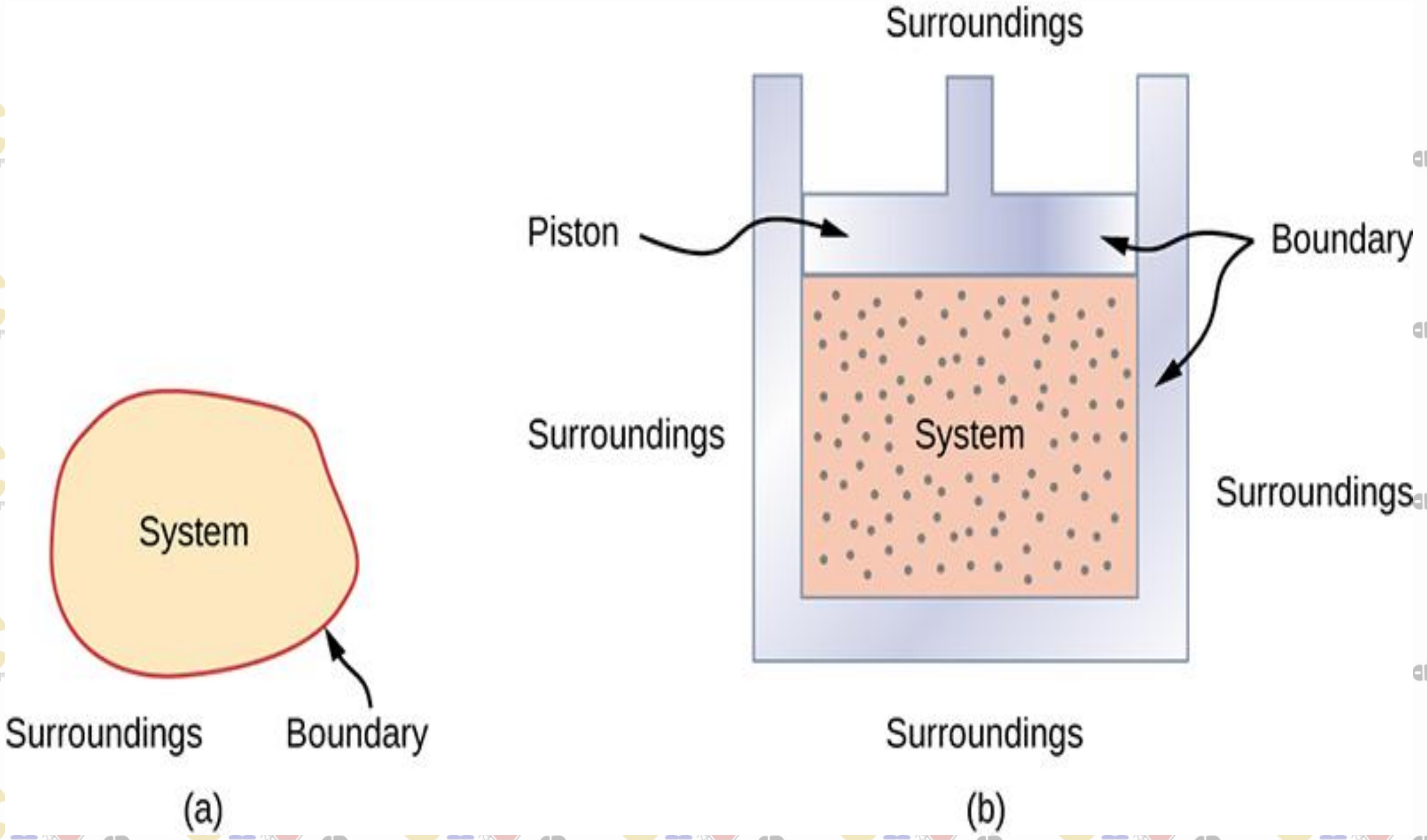


System

Boundary

The real or imaginary surface that separates the system from the surrounding is called boundary





Types of system

1. Open system
2. Closed system
3. Isolated system

Open system

A system which can exchange both matter and energy with the surroundings.

Closed system

A system which can exchange only energy but not matter with the surroundings.

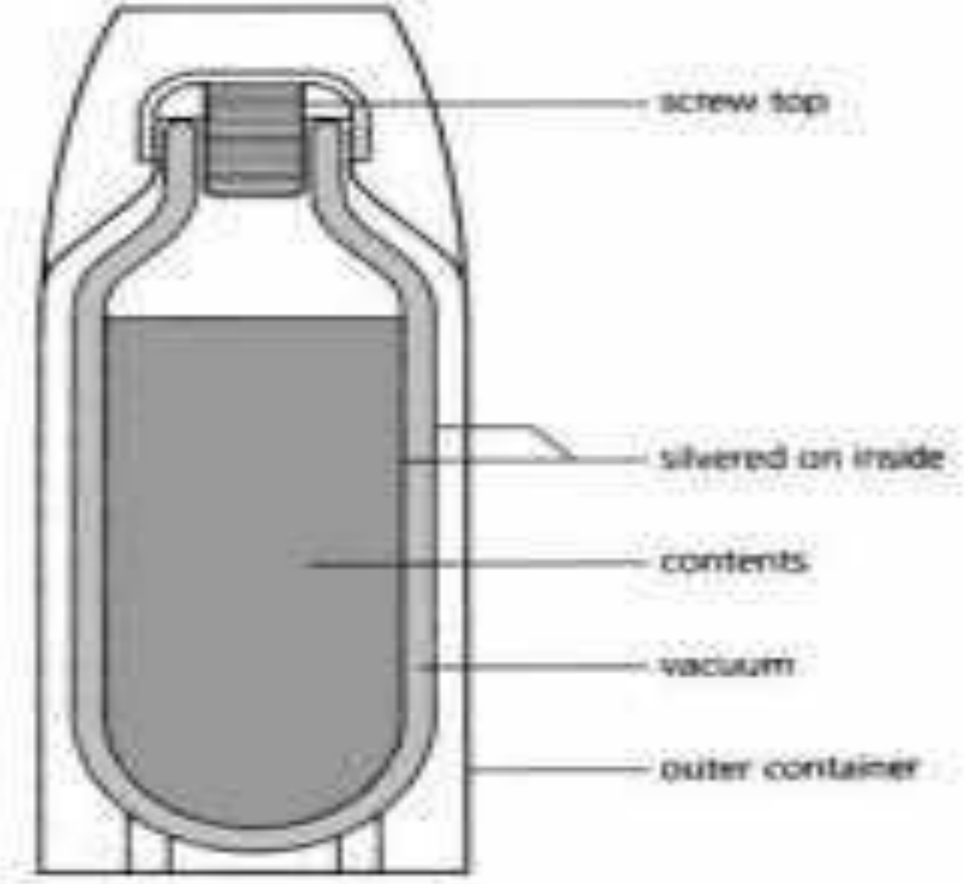
(a) This boiling tea kettle is an open thermodynamic system. It transfers heat and matter (steam) to its surroundings.

(b) A pressure cooker is a good approximati to a closed system. A little steam escapes through the top valve to prevent explosion.

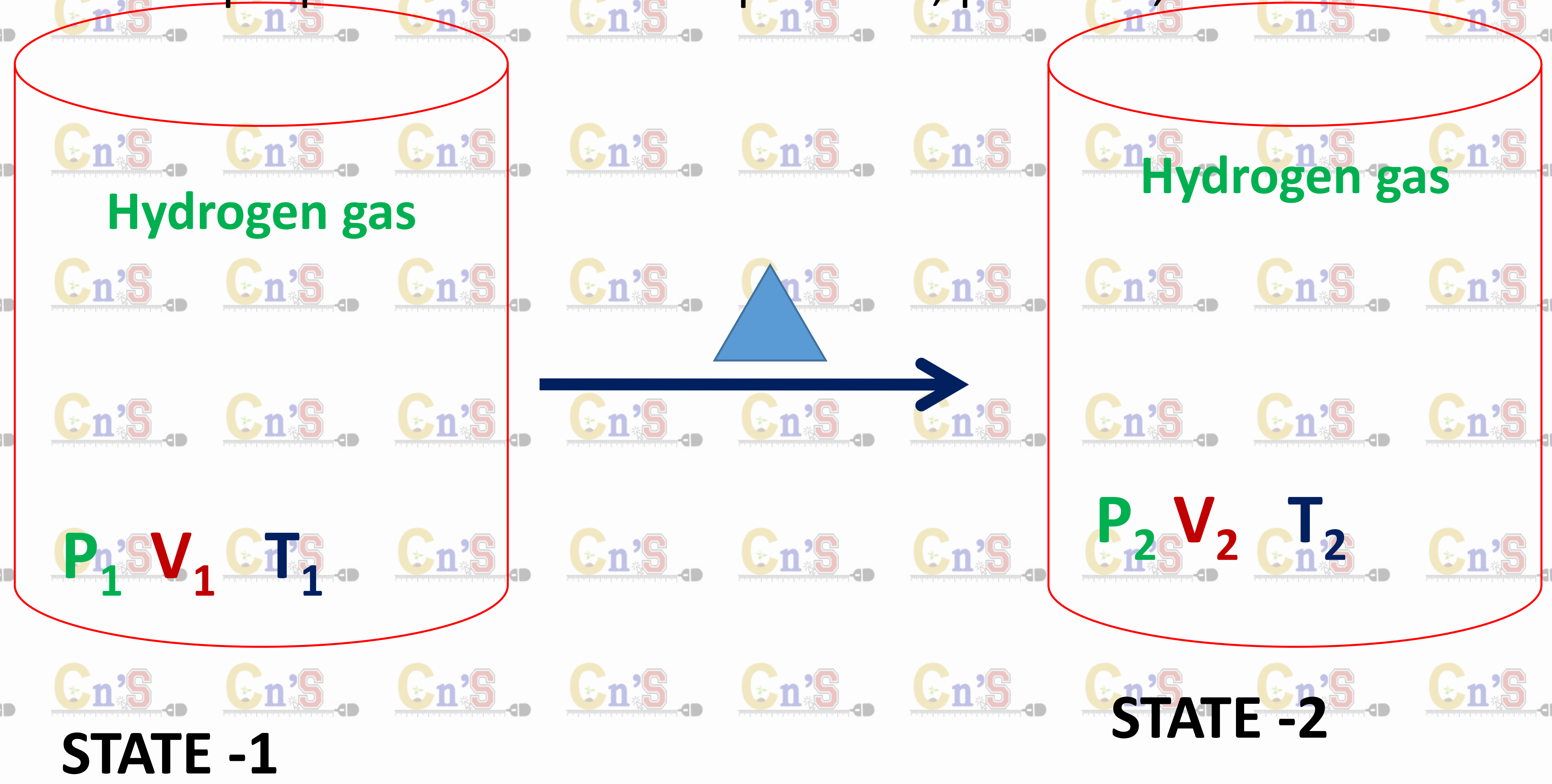
(a) (b)

Isolated system

A system which cannot exchange both energy and matter with the surroundings. Thus, if a system is isolated, its internal energy must remain constant.



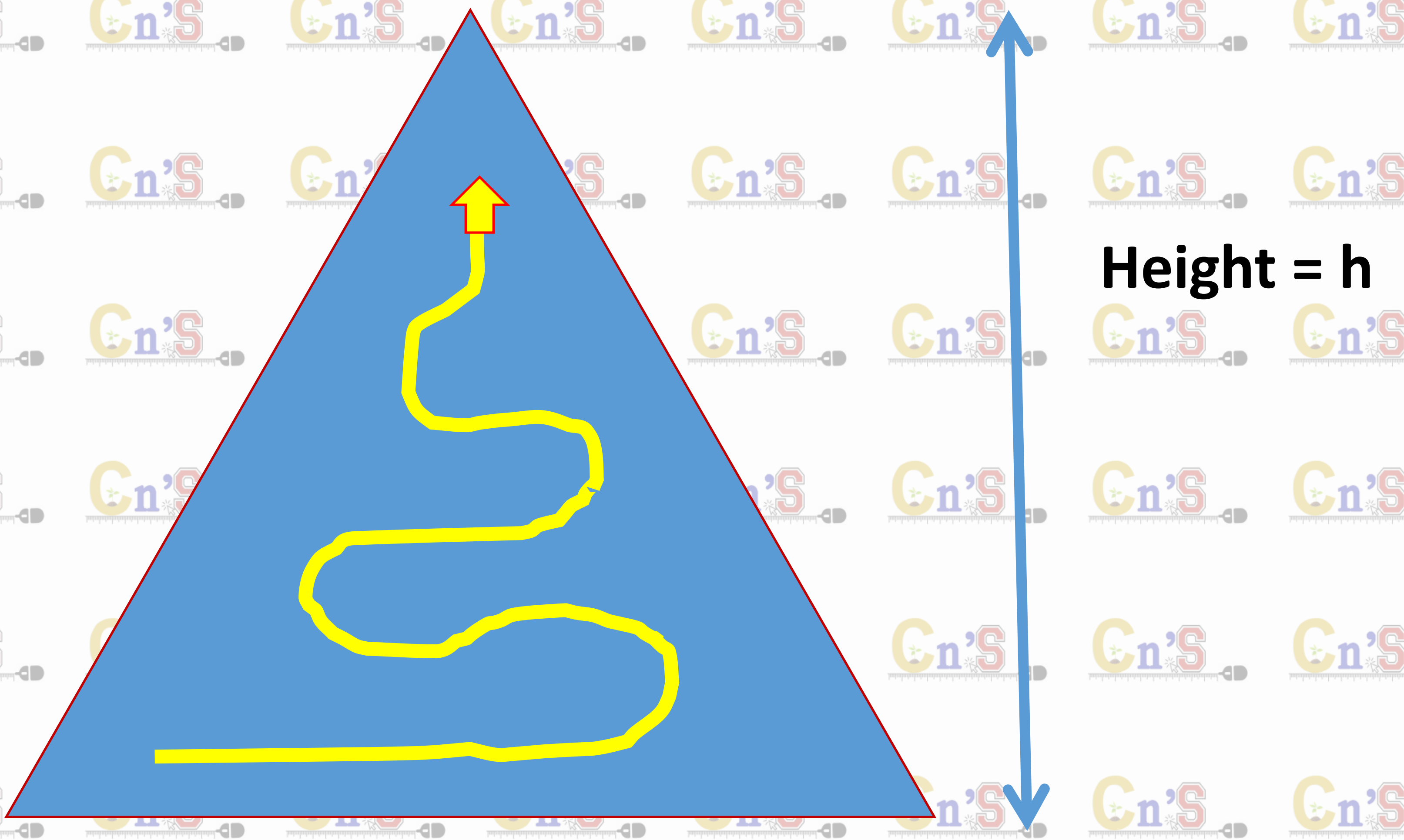
State of a system It is the condition of the system expressed by giving definite values for its properties such as temperature, pressure, volume etc.



State functions

The thermodynamic properties whose values depend only on the <u>initial and final state of the system</u> and are independent of the manner as to how the changes is brought about .

Eg. Pressure,
temperature ,
volume,
internal energy,
enthalpy,
entropy



Height h of a mountain is independent of the path followed in reaching the top of the mountain. h is similar to a state function

Path functions Work as Path functions

Work = force x displacement

The definition of work indicates that work depends on its path it takes, because the movement of an object is dependent upon the path taken to execute that movement.

Eg. Work done by a person for climbing stairs is different from using a lift.

Heat as Path functions

For instance, if a gas expands isothermally, then heat has to be supplied to the system so that the gas maintains its temperature as it expands. But if you do this adiabatically, then the system does work. Same final state (pressure and volume) but different work and heat.

Thermodynamic process

The operation which brings about the change in the state of the system .

<u>Isothermal process</u>: A process which is carried out at constant temperature. $\Delta T = 0$

An isothermal process is one that takes place without any change in *temperature*.

Isothermal processes are often described as "slow".

The pressure of a gas is inversely proportional to its volume *only if* the change takes place isothermally.

Adiabatic:

A process in which there is no heat exchange occurs between system and surrounding $\Delta q = 0$

An adiabatic process is one that takes place without any exchange of *heat*.

Adiabatic processes are often described as "fast".

The pressure of a gas is *not* inversely proportional to its volume if the change takes place adiabatically.

- **Isobaric:**

A process which is carried out at constant pressure.ΔP = 0

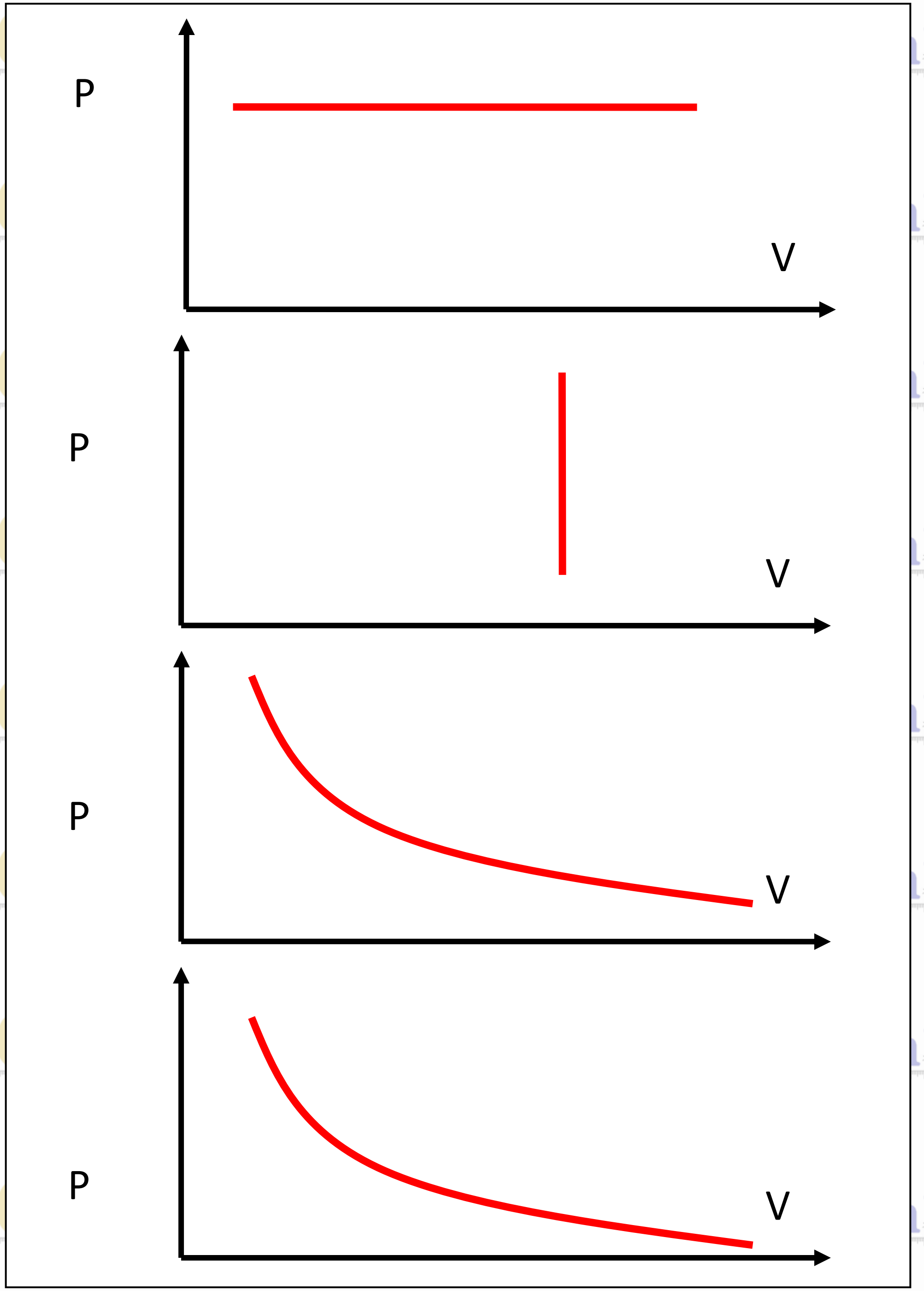
- An isobaric process is one that takes place without any change in *pressure*.

- **Isochoric :**

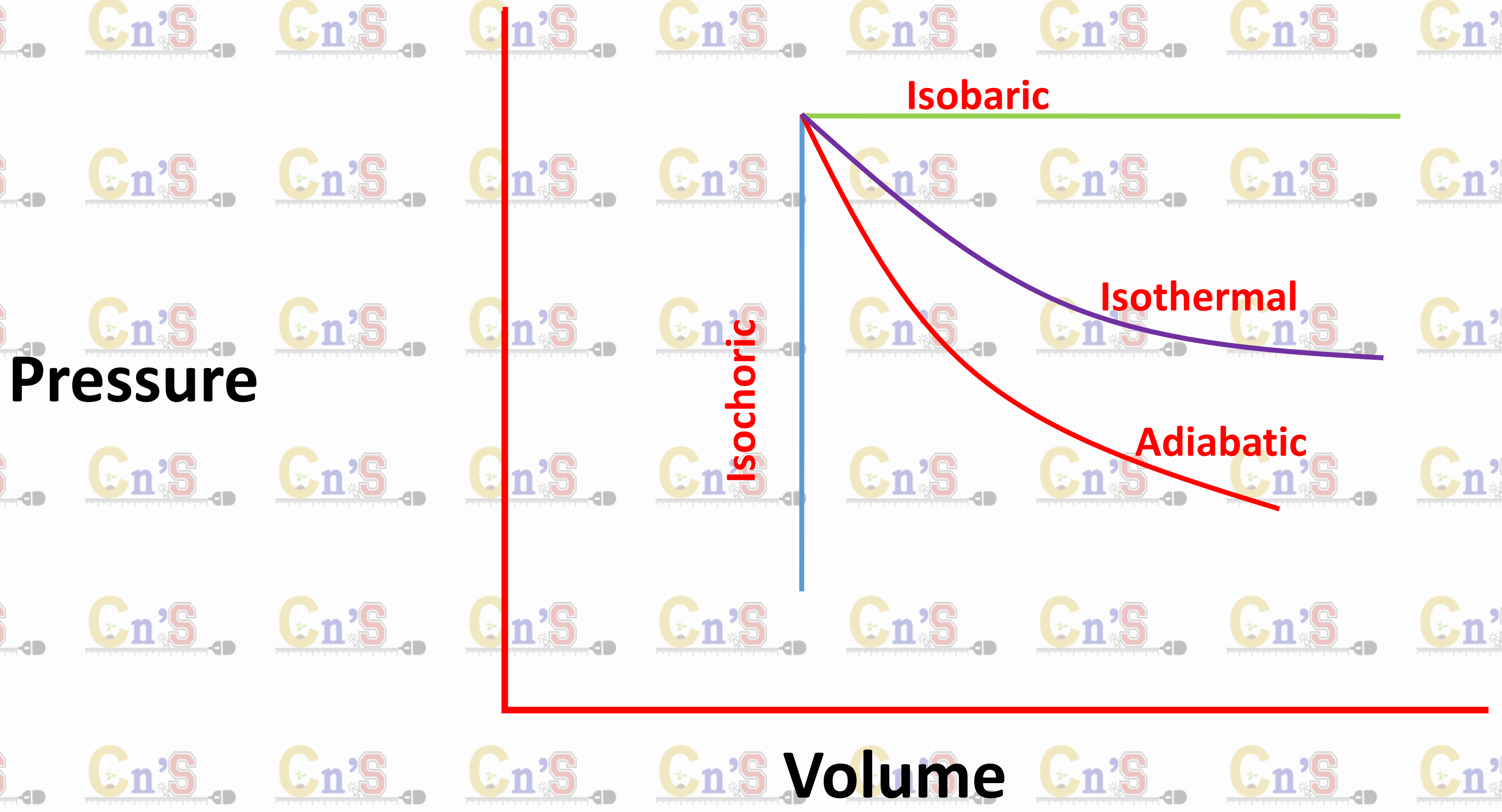
A process which is carried out at constant volume. ΔV = 0

- An isochoric process is one that takes place without any change in *volume*.

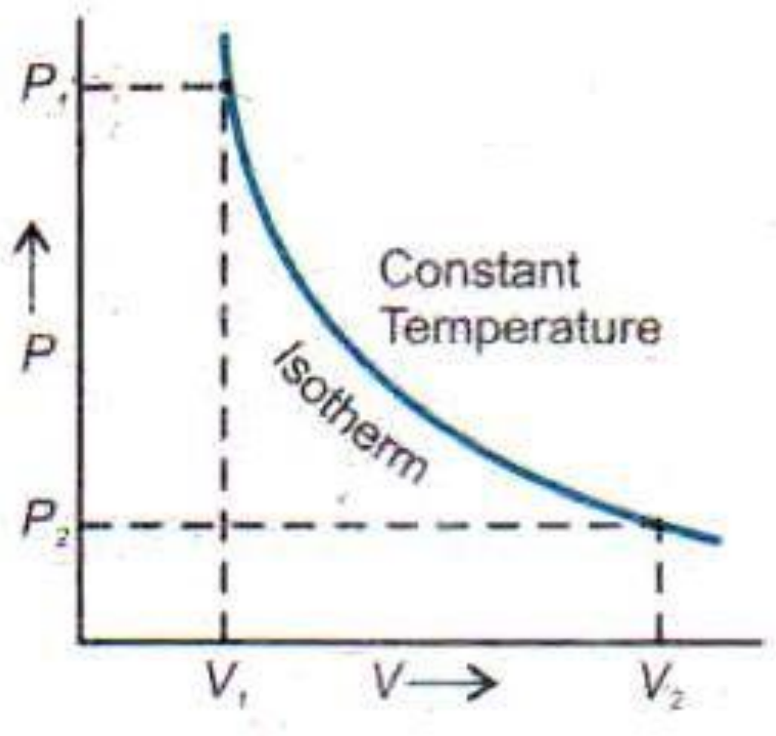
- Isobaric
  - P = constant
- Isovolumetric (Isochoric)
  - V = constant
  - W = 0
- Isothermal
  - T = constant
  - $\Delta U = 0$ (ideal gas)
- Adiabatic
  - Q = 0



For a given amount of ideal gas P – V relation

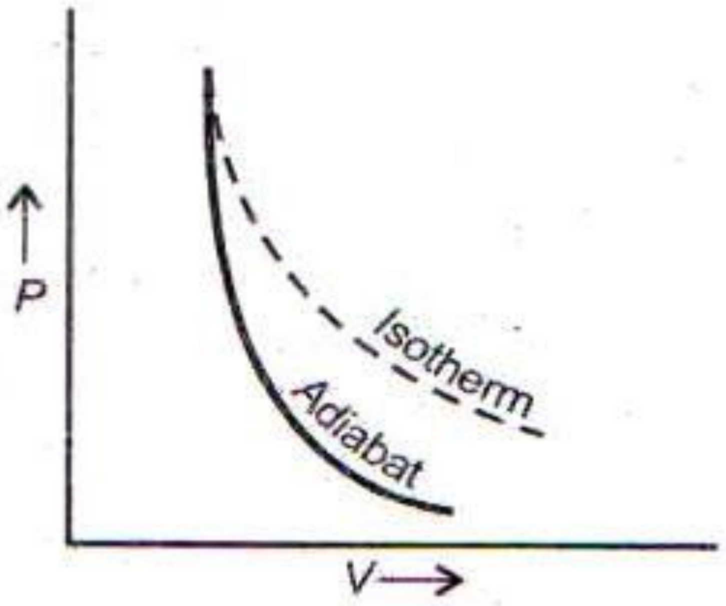


Isothermal Process is a process which is carried out at constant temperature. At constant temperature, the condition of Boyle's law on the gas is fulfilled. Therefore, when gas expands or compresses isothermally, the product of its pressure and volume during the process remains constant. If P_1, V_1 are the initial pressure and volume where as P_2, V_2 are pressure and volume after the isothermal change takes place, then $P_1V_1 = P_2V_2$ The PV-curve representing an isothermal process is called an isotherm. In case of an ideal gas, the P.E. associated with the molecules is zero. And the internal energy of an ideal gas depends only on its temperature, which in this case remains constant, therefore . Hence the first law of thermodynamics reduces to: $\Delta U = Q - W = 0$, $Q = W$ Thus if gas expands and does external work W, an amount of heat Q has to be supplied to the gas in order to produce an isothermal change.



Note: Since the transfer of heat from one place to another requires time, hence, to keep the temperature of the gas constant, the expansion of compression must take place slowly

Adiabatic process is a process in which no heat enters of leaves the system. In case of adiabatic change it has been seen that $PV^{\gamma} = \text{constant}$ Where γ is the ratio molar specific heat of the gas at constant pressure to molar specific heat at constant volume. The PV-curve representing an adiabatic process is called an adiabat. As no heat enters or leave the system during an adiabatic process i.e.



$$\Delta U = Q - W$$
$$\Rightarrow \Delta U = 0 - W$$
$$\Rightarrow W = -\Delta U$$

- If the gas expands and does external work, it is done at the expense of the internal energy of its molecules and, hence, the temperature of the gas falls.
- An adiabatic compression causes the temperature of the gas to rise because of the work done on the gas.

This expression tells that:

 If the gas expands and does external work, it is done at the expense of the internal energy of its molecules and, hence, the temperature of the gas falls.

 An adiabatic compression causes the temperature of the gas to rise because of the work done on the gas.

- Important Note:
- Adiabatic change occurs when the gas expands or is compressed rapidly. The examples of adiabatic processes are:
- The rapid escape of air from a burst tyre
- The rapid expansion and compression of air through which the sound wave is passing
- Cloud formation in atmosphere Important
- Note: Adiabat is steeper than Isotherm

- Q # 1 Is it possible to convert internal energy into mechanical energy?
- Ans. Yes it is possible to convert internal energy into mechanical energy. In adiabatic expansion, the system expands and moves the piston upward at the cost of its own internal energy.
- Q # 2. Can the mechanical energy be converted completely into heat energy?
- Ans. Yes mechanical energy can be converted into heat energy. In adiabatic compression, when the piston of the cylinder is pushed downwards, the temperature of the gas increases.

- 5. A process in which no heat enters or leaves the system is called: a) Isothermal process b) Adiabatic process c) Isochoric process d) Isobaric process
- 6. Gas law is for: a) Isothermal process b) Adiabatic process c) Isobaric process d) Isochoric process
- 7. Cloud formation in the atmosphere is example of: a) Adiabatic process b) Isothermal process c) Isochoric process d) Isobaric process
- 8. Which one is true for internal energy? a) It is sum of all forms of energies associated with molecules of a system. b) It is a state function of a system c) It is proportional to transnational K.E of the molecules d) All are correct
- 9. An adiabatic change is one in which: a) No heat is added to or taken out of a system b) No change of temperature takes place c) Boyle's law is applicable d) Pressure and volume remains constant
- 10. The first law of thermodynamics is an expression of: a) The conservation of energy b) Conservation of mass c) Heat death of the universe d) Degradation of energy
- 11. The expression for isothermal process is: a) Q=U b)Q=W c)U=W d) U=-W
- 12. In adiabatic expansion, first law of thermodynamics becomes: a) Δ Q=U b) Δ Q=W
- c) Δ U=W d) None of these

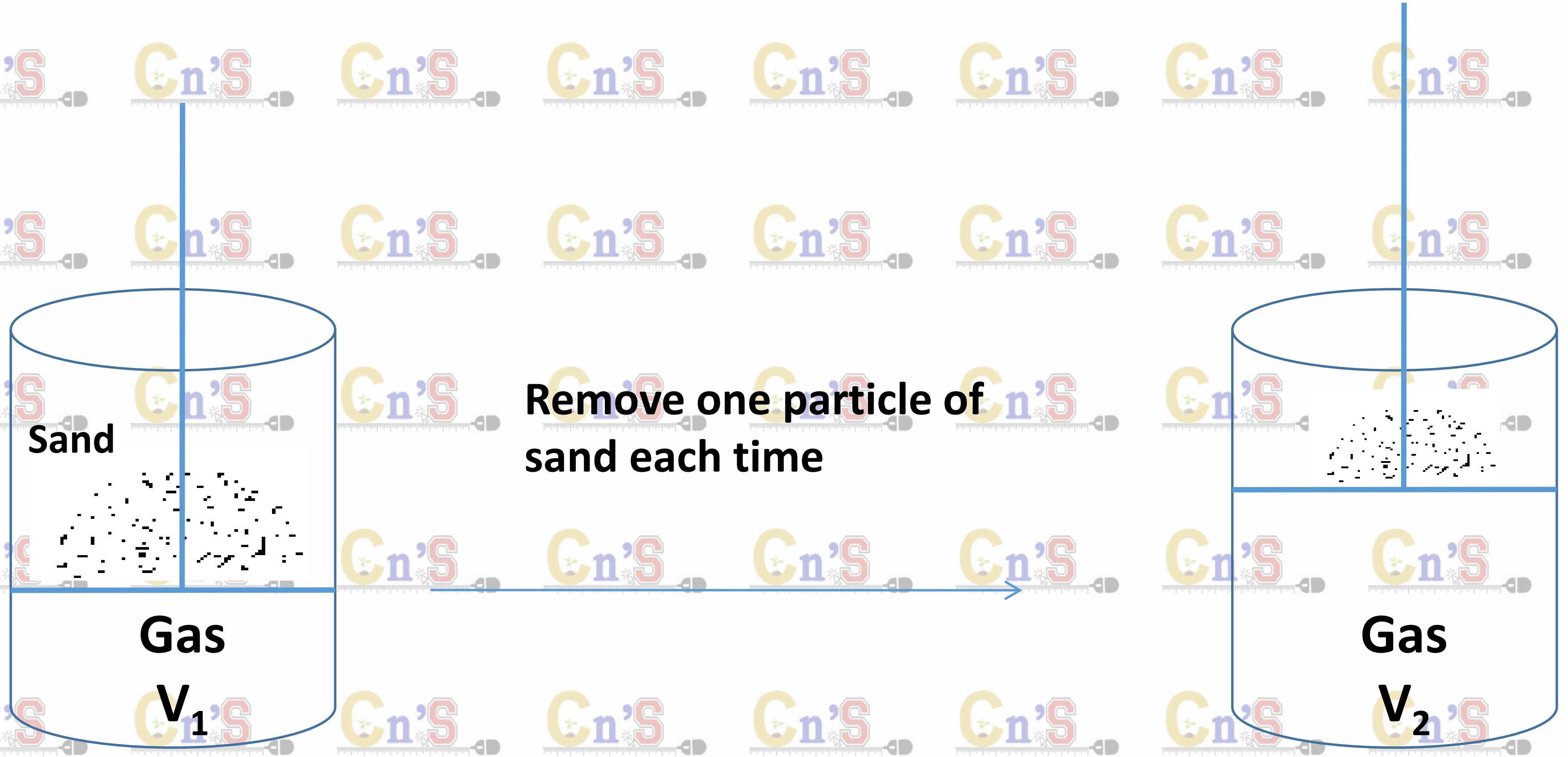
Reversible process

It is a process which is carried out infinitely slowly through a series of steps so that system and surroundings always remain almost in equilibrium state. The process is conducted in such a manner that any moment it could be reversed by a infinitesimal change.

i.e. ▯ The processes of liquefaction and evaporation, performed slowly, are practically reversible.

▯ Slow compression of gas in a cylinder is reversible as the compression can be changed to expansion by slowly decreasing pressure on the piston to reverse the operation.

Reversible process



It is a process which is carried out infinitely slowly through a series of steps so that system and surroundings always remain almost in equilibrium state. The process is conducted in such a manner that any moment it could be reversed by a infinitesimal change.

Reversible expansion process involves infinite number of steps.

Irreversible process

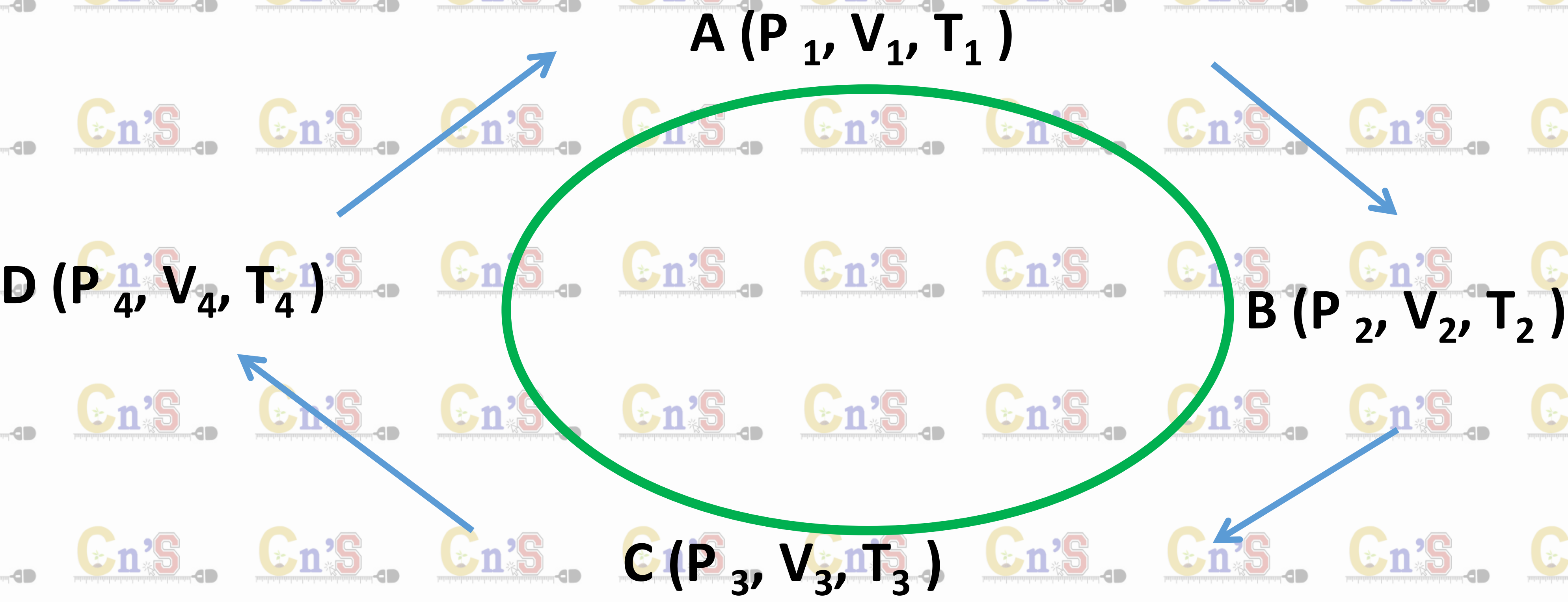
- **A process which is carried out rapidly so that the system does not get a chance to attain equilibrium.**

- All changes which occur suddenly or which involve friction or dissipation of energy through conduction, convection or radiation are irreversible. An example of highly irreversible process is an explosion.

(**Quasi-static process:-** The process in which change in any of the parameters take place at such a slow speed that the values of P,V, and T can be taken to be, practically, constant, is called a quasi-static process.)

Cyclic process

A process during which the system undergoes a series of changes and return to its initial state.



Properties of the system

a. Intensive property

Property of a system which does not depend upon the quantity of substance present in the system.

Eg. Density, <u>temperature</u>, refractive index, viscosity, <u>pressure</u> ,surface tension, specific heat, freezing point, boiling point, melting point, emf, pH, mole fraction, molarity etc.

Intensive is independent of quantity

Properties of the system

b. Extensive property.

Property of a system which depends upon the quantity of substance present in the system.

Eg. Mass , volume, energy, enthalpy, internal energy, the number of moles,**etc.**

HEAT

Form of energy

How can we feel it?

From the change in temperature

Heat is the amount of energy transferred between the system and the surrounding when they are at different temperatures.

International conventions

Symbol of heat = q

Heat absorbed by the system = +q

Example?

Heat liberated by the system = -q

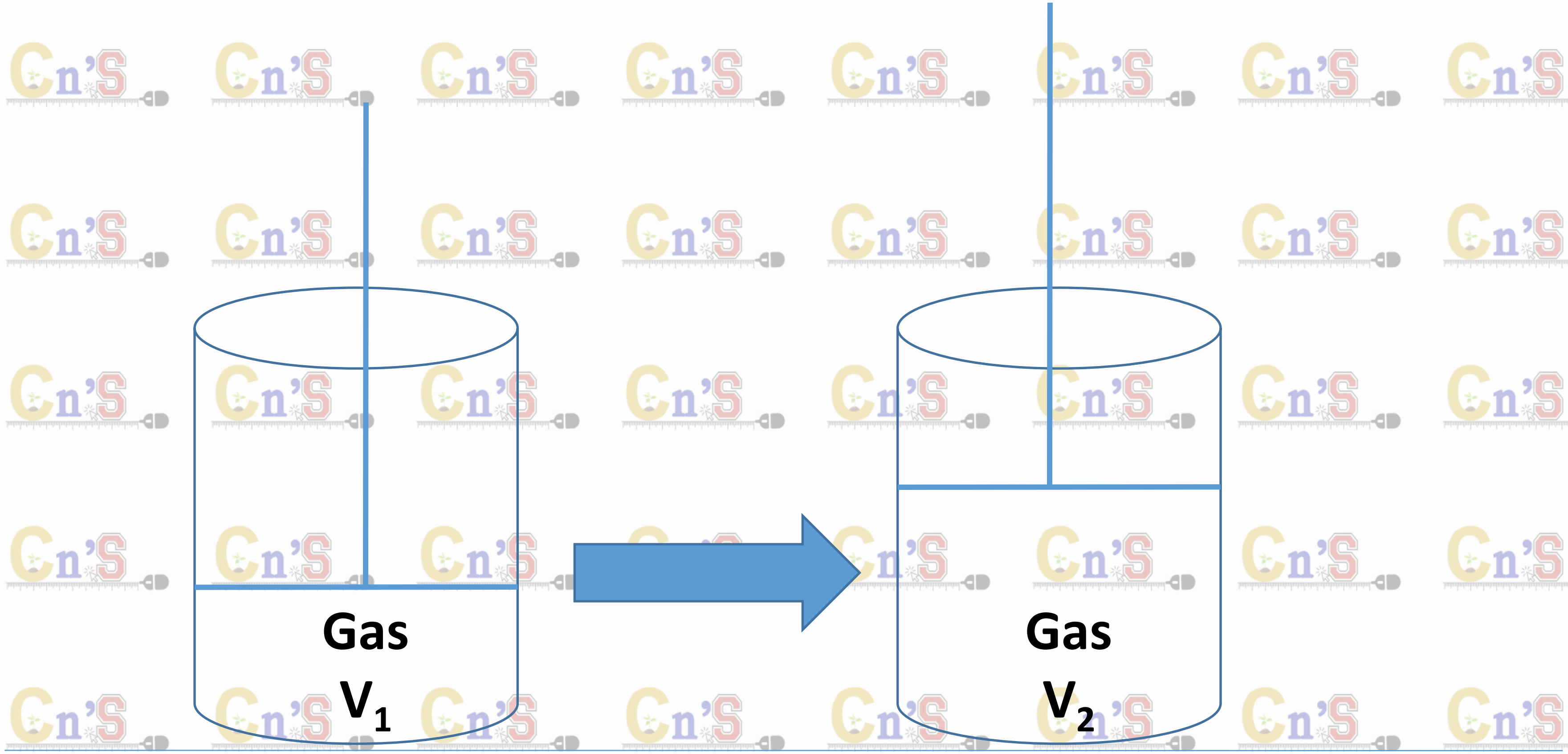
Example?

Other method of exchange of energy between system and surrounding

WORK

1. Mechanical work
2. Electrical work
3. Pressure volume work

Pressure volume work

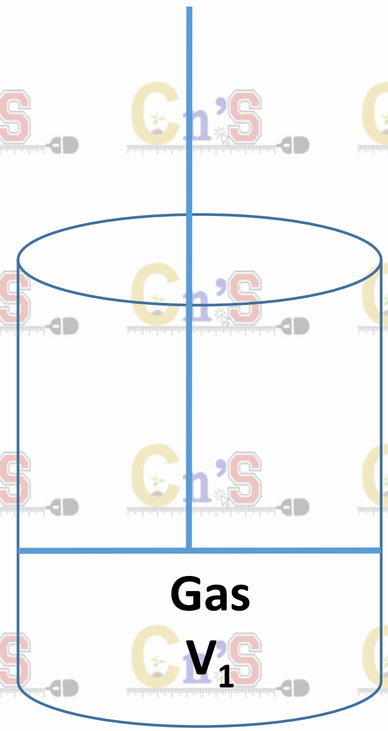


It is also called expansion work. It is significant in system which consists of gases and involve change in volume against external pressure

Compression

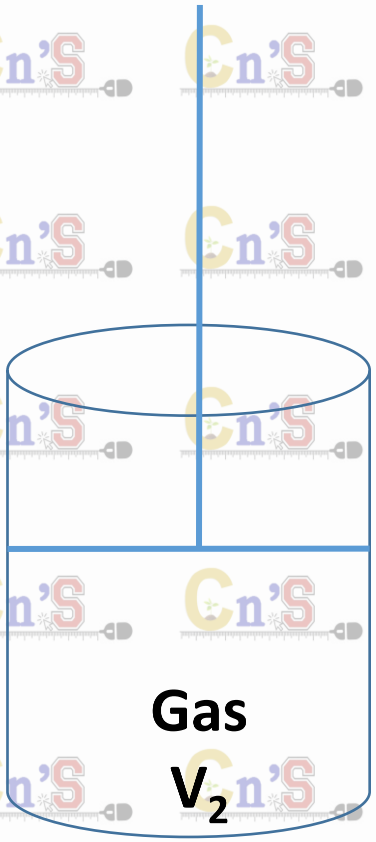
work done

on the system = + w



work done

by the system = - w



Expansion

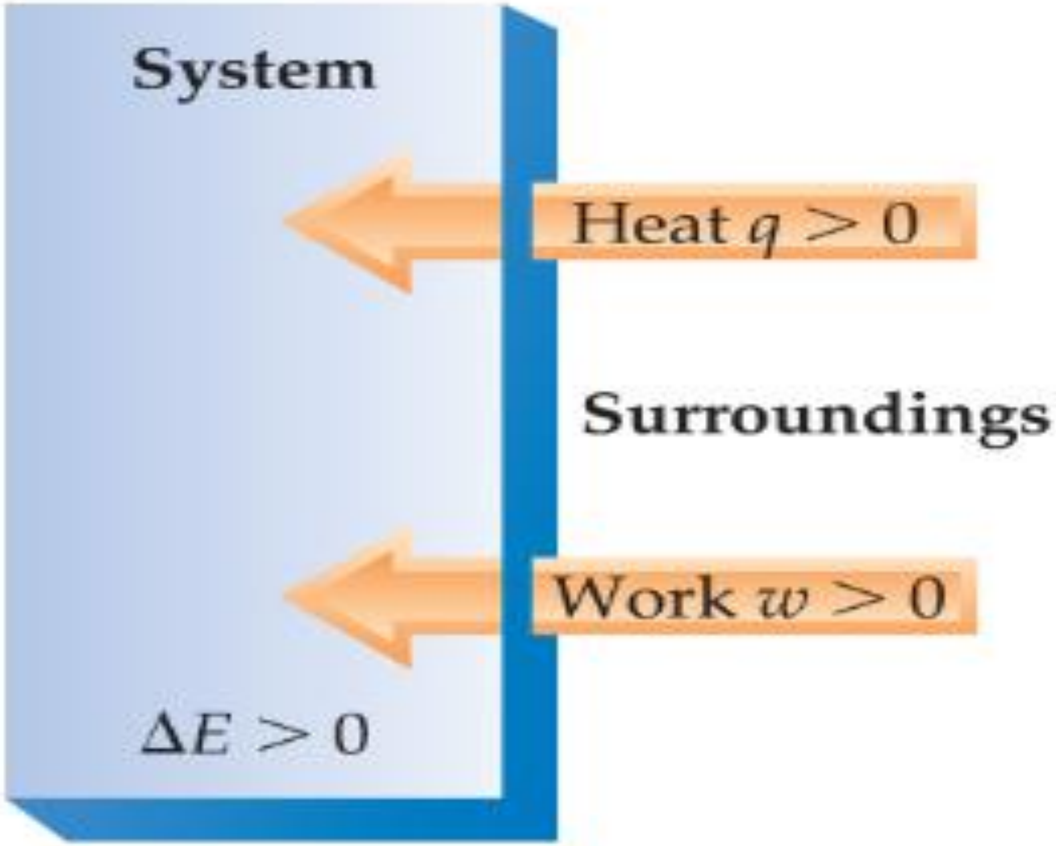
Internal energy

Characteristics:

1. It is the sum of translational E + rotational E + vibration E + Bond E
2. It depends on mass of system
3. It depends on state of system
4. It is indicated by U
5. The absolute value of internal energy cannot be measured.
6. Change in internal energy of a system can be measured
7. $\Delta U = U_2 - U_1$

- The sum of all forms of molecular energies (kinetic and potential) of a substance is termed as its internal energy. Kinetic energy is due to various types of motion (translational, rotational, vibrational).
- The molecules of an ideal gas don't exert forces on one another. So the internal energy of an ideal gas system is generally the translational K.E. of its molecules. Since the temperature of a system is defined as the average K.E of its molecules, thus for an ideal as system, the internal energy is directly proportional to its temperature
- . In case of an ideal gas, the P.E. associated with the molecules is zero. And the internal energy of an ideal gas depends only on its temperature, which in this case remains constant.

Changes in Internal Energy



- When energy is exchanged between the system and the surroundings, it is exchanged as either heat (q) or work (w).
- **That is, $\Delta E = q + w$.**

Internal energy of a system may change when:

1. Heat passes into or out of the system

2. Work is done on or by the system

3. Matter enters or leaves the system

Change in internal energy in terms of both adiabatic work and heat transfer

$$\Delta U = U_2 - U_1 = q + w$$

Mathematical expression for 1st law of thermodynamics

First law of thermodynamics is thus *conventionally* stated as:
"The change in internal energy of a closed system is equal to the energy added to it in the form of heat (Q) plus the work (W) done on the system by the surroundings."

When $q = 0$ and $w = 0$ (a state possible in an isolated system)

$$\Delta U = 0$$

Statement of 1st law of thermodynamics

The energy of an isolated system in constant

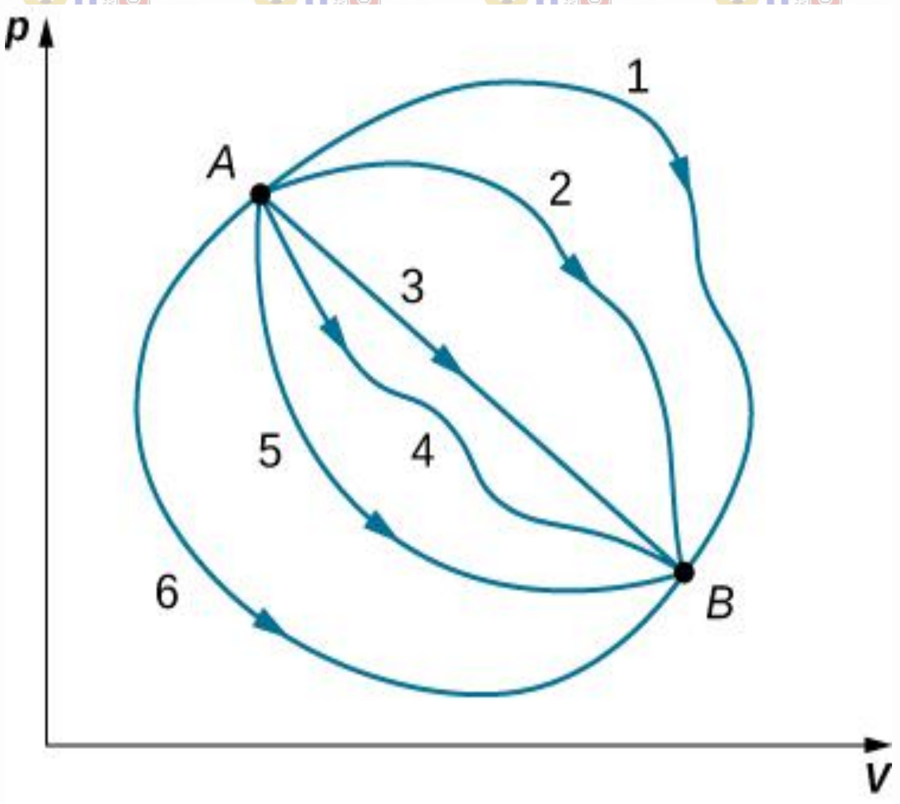
Free expansion:- Such an expansion in which no external work is done and the total internal energy of the system remains constant is called free expansion.

Different thermodynamic paths taken by a system in going from state A to state B.

For all transitions, the change in the internal energy of the system $\Delta E_{int} = Q - W$ *is the same.*

Although Q and W both depend on the thermodynamic path taken between two equilibrium states, their difference Q−W does not. Along path 1, the system absorbs heat Q1 and does work W1 along path 2, it absorbs heat Q2 and does work W2 and so on. But,

$Q_1 - W_1 = Q_2 - W_2$ or $\Delta E_{int\,1} = \Delta E_{int\,2}$



Principles of Thermodynamics

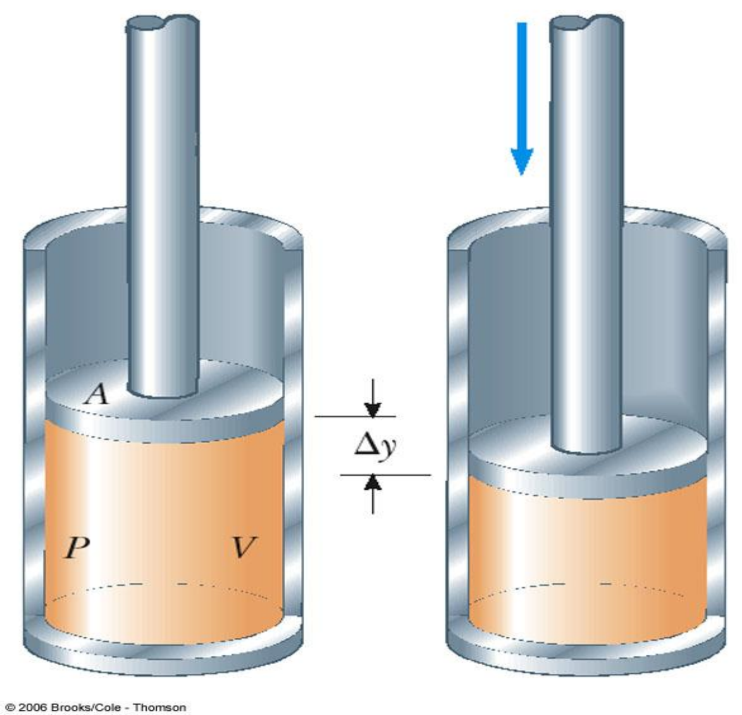
- Energy is conserved
  - FIRST LAW OF THERMODYNAMICS
  - Examples:
    - Engines (Heat -> Mechanical Energy)
    - Friction (Mechanical Energy -> Heat)
- All processes must increase *entropy*
  - SECOND LAW OF THERMODYNAMICS
  - Entropy is measure of disorder
  - Engines can not be 100% efficient

Work done on a gas

- Adding heat Q can:
  - Change temperature
  - Change state of matter
- Can also change ΔU by doing work on the gas

Change of
Internal Energy ΔU

$$W = \vec{F} \cdot \Delta\vec{x} = F(-\Delta y) = (PA)(-\Delta y)$$



$$W = -P\Delta V$$

Work done on the gas

First Law of Thermodynamics

$$\Delta U = Q + W$$

- Conservation of Energy
- Can change internal energy ΔU by
  - Adding heat to gas: Q
  - Doing work on gas:

$$W = -P\Delta V$$

By **considering dimensions** it can be shown that :

$$W = p\Delta V$$

$$[J] = \left[\frac{N}{m^2}\right] \times [m^3] = [Nm] = [J]$$

Note: (Work done <u>by the gas</u>) = - (Work done <u>on the gas</u>)

$$W_{\text{by the gas}} = +P\Delta V$$

$$Q = \Delta U + W_{\text{by the gas}}$$

Add heat => Increase Int. Energy & Gas does work

(1 L atm = 101.3 J).

When heat energy ***ΔQ*** is supplied to the gas :

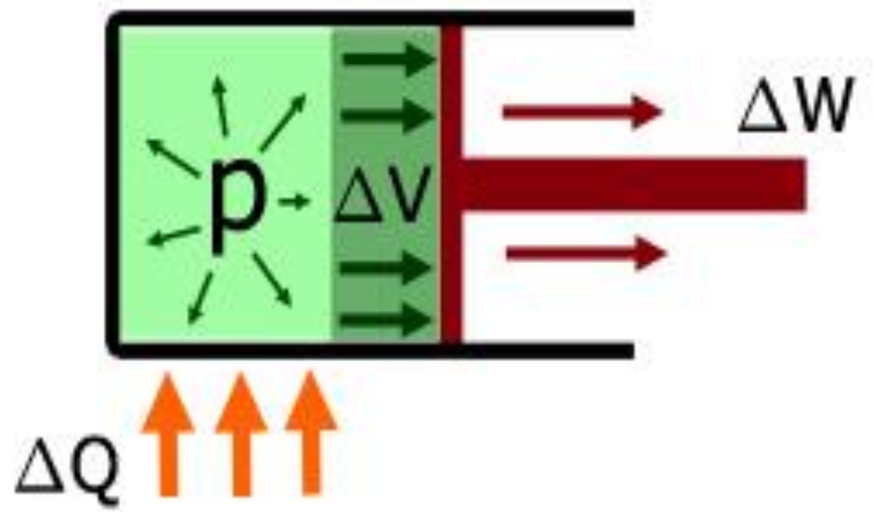
the **temperature** of the gas **increases**

work ***ΔW*** is done by the gas expanding to move the piston

the **internal energy** of the gas **increases *ΔU***

the **volume** of the gas **increases *ΔV***

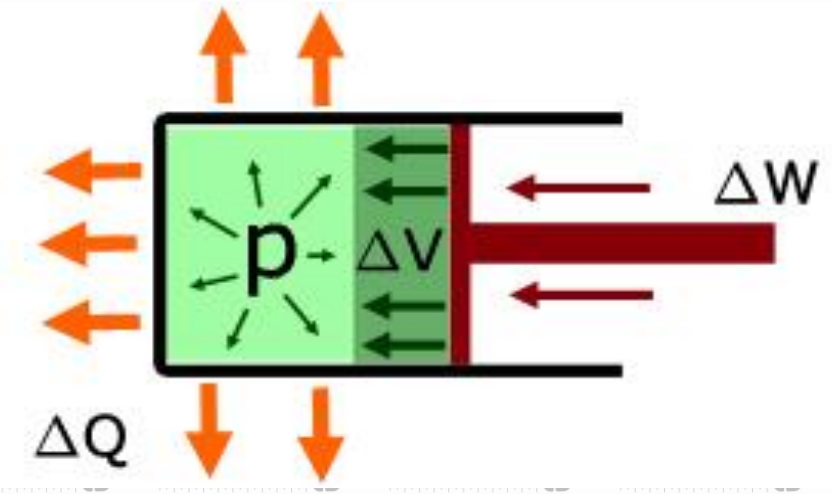
$$\Delta U = \Delta Q - \Delta W$$



ΔW is negative - work is removed from the system

ΔQ is positive - heat is supplied to the system

When no heat is applied and work is done externally by pushing the piston inwards to compress the gas:



the **temperature** of the gas **increases**

work ***ΔW*** is done compressing the gas

the **internal energy** of the gas **increases *ΔU***

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

the volume of the gas decreases ***ΔV***

ΔW is positive - work is done on the system

ΔQ is negative - heat is removed from the system

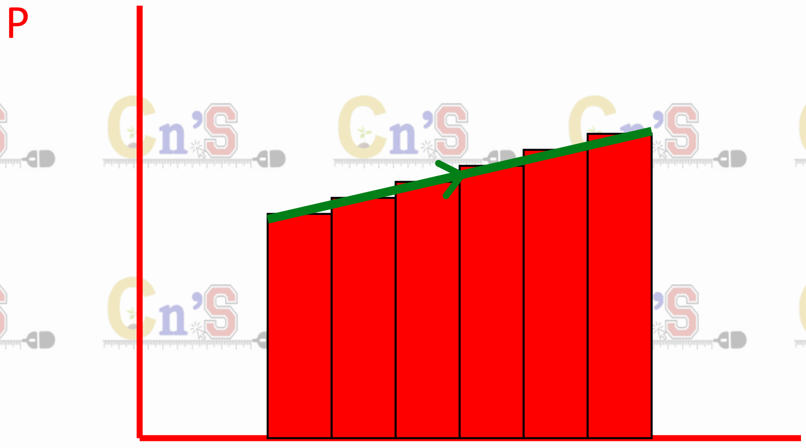
Q # 20. A gas is enclosed in a container fitted with a piston of cross-sectional area $0.10\ m^2$. The pressure of the gas is maintained at $8000\ Nm^{-2}$. When heat is slowly transferred, the piston is pushed up through a distance of $4\ cm$. If $42\ J$ heat is transferred to the system during the expansion, what is the change in internal energy of the system?

Q # 21. A thermodynamic system undergoes a process in which its internal energy decreases by 300 J. If at the same time 120 J of work is done on the system, find the heat lost by the system.

Calculations: Work $W = P\,\Delta V = PA\Delta y = 8000 \times 0.10 \times 0.04 = 32\ J$

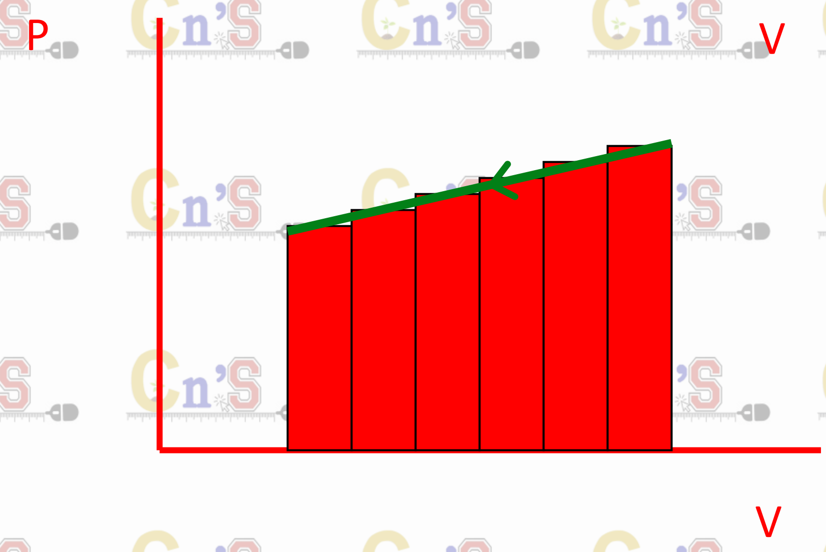
By First Law of Thermodynamics $\Delta U = Q - W = 42 - 32 = 10\ J$

Calculations: By First Law of Thermodynamics $Q = \Delta U + W = -300 - 120 = -420\ J$



Path moves to right:

$W_{\text{by the gas}}$ = Area under curve



Path moves to left:

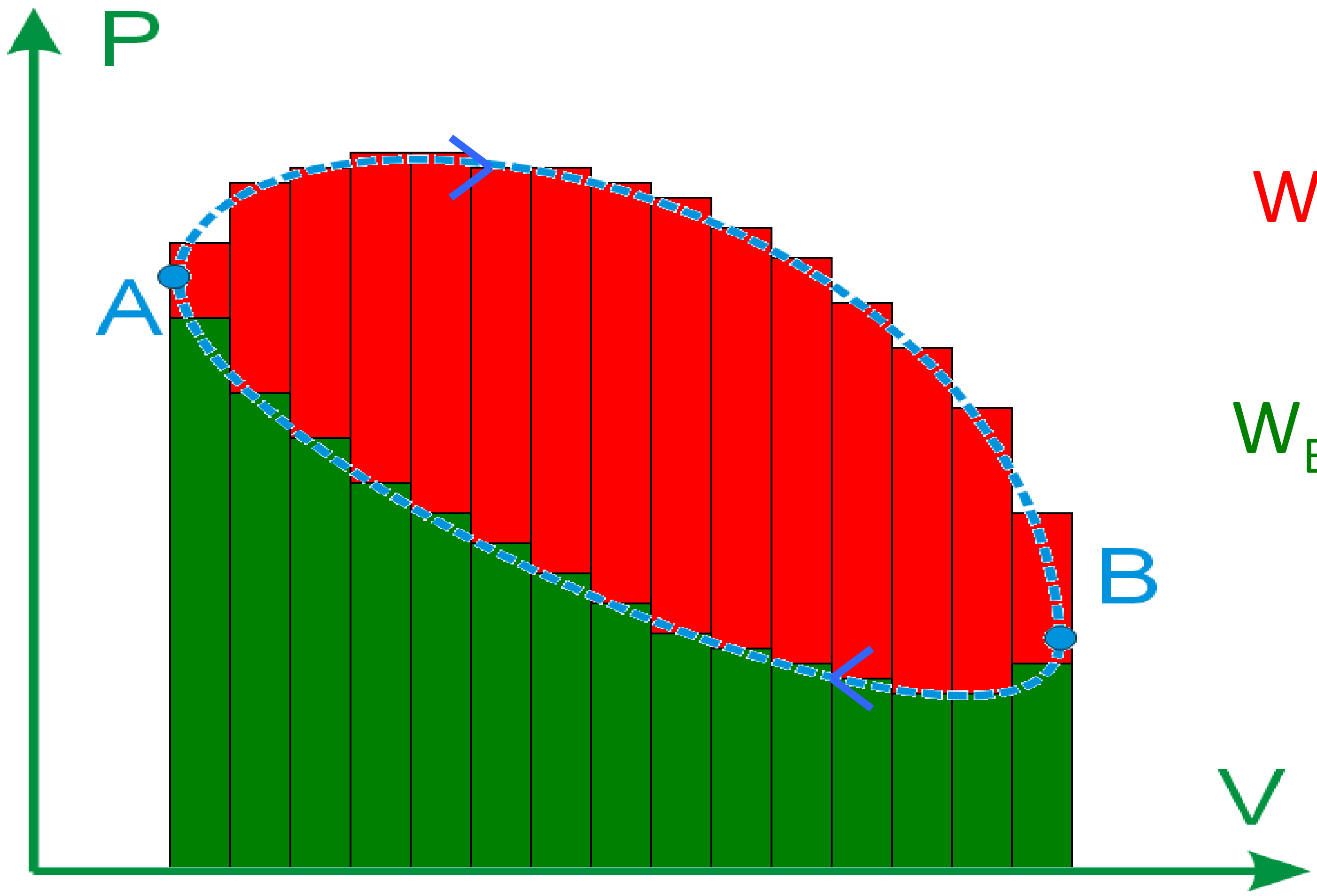
$W_{\text{by the gas}}$ = - Area under curve

P-V Diagrams

($W_{\text{on the gas}} = - W_{\text{by the gas}}$)

Work from closed cycles

Consider cycle A -> B -> A



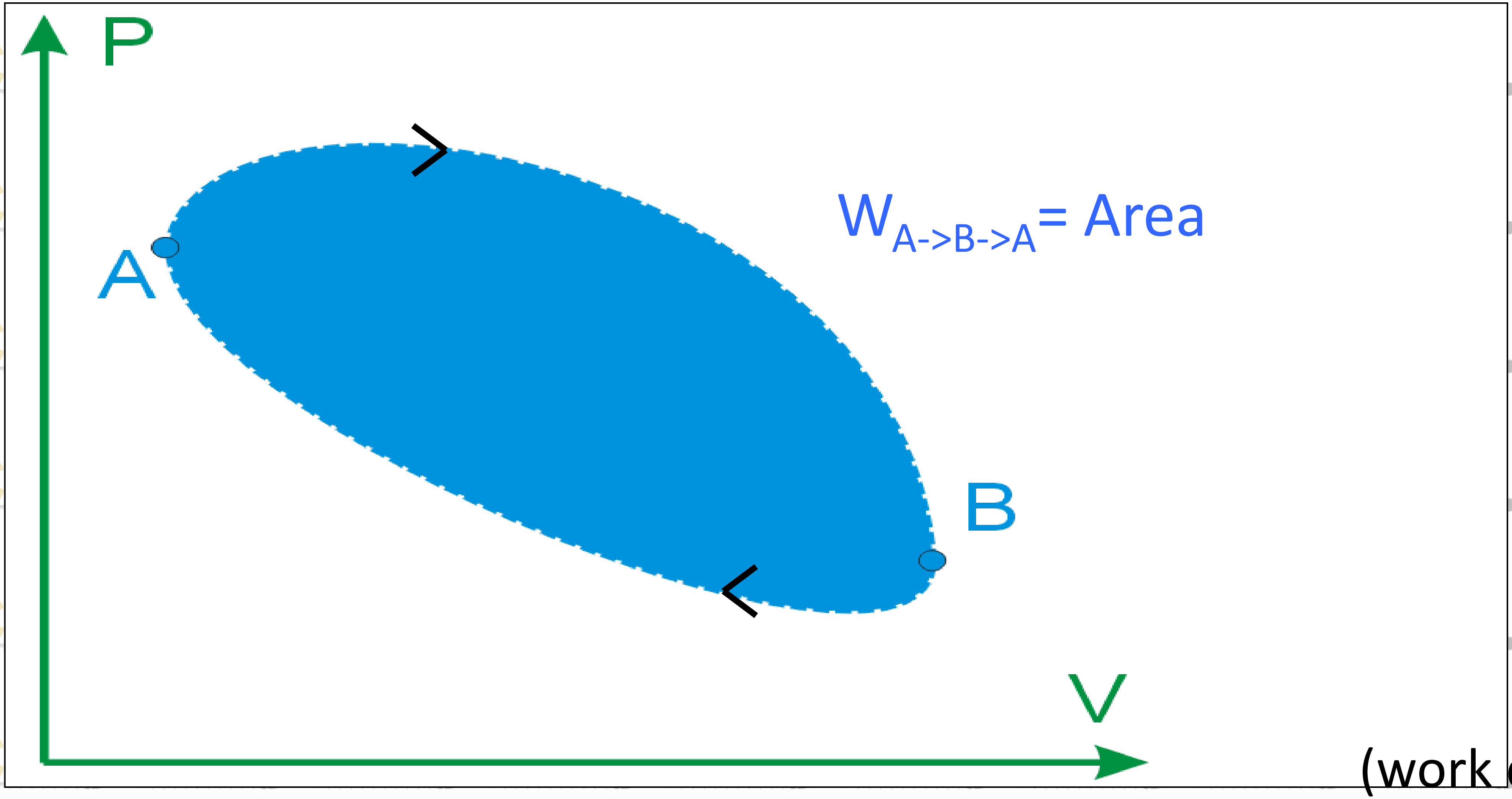
$W_{A->B}$ = Area total

$W_{B->A}$ = - Area

(work done by gas)

Work from closed cycles

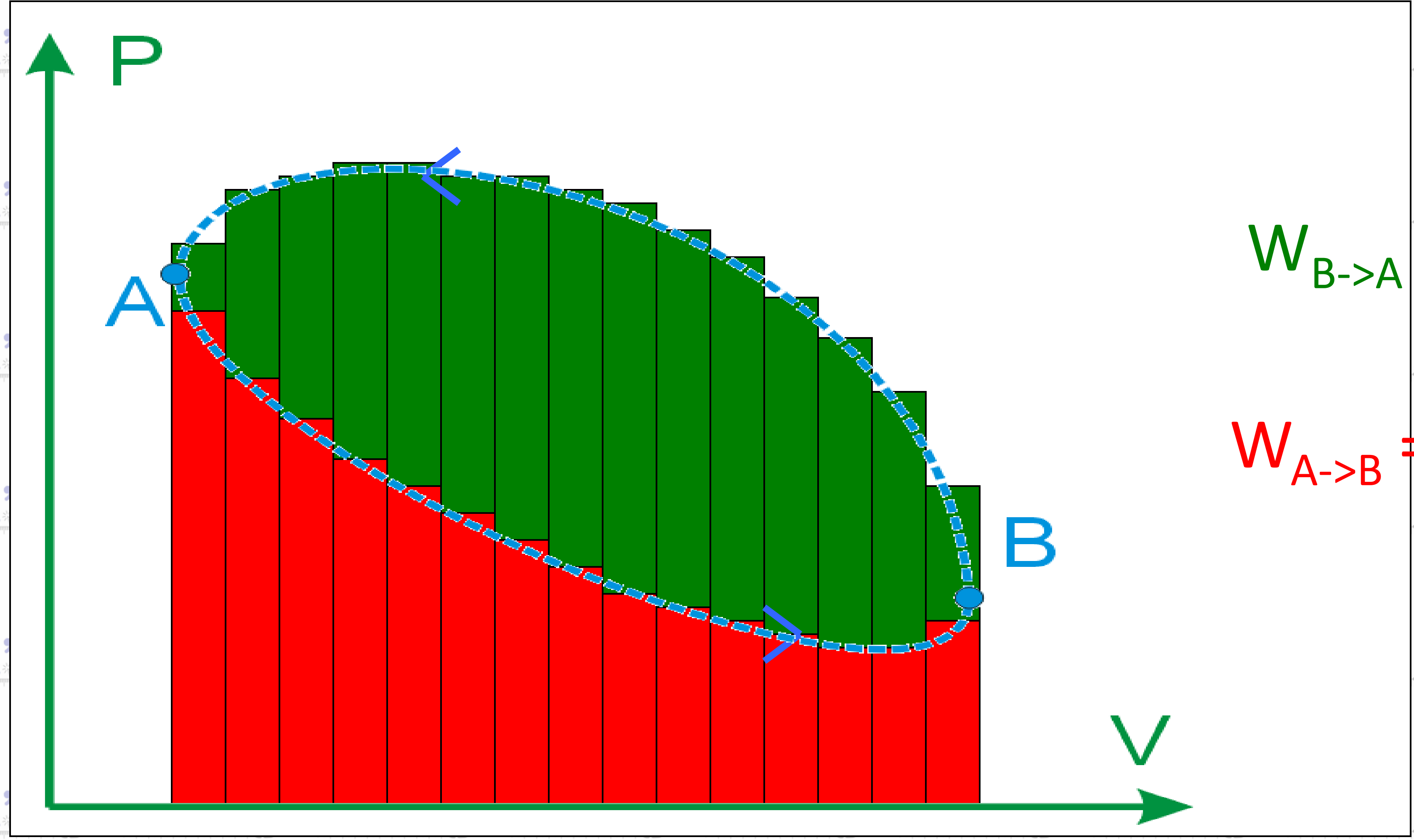
Consider cycle A -> B -> A



(work done by gas)

Work from closed cycles

Reverse the cycle, make it counter clockwise



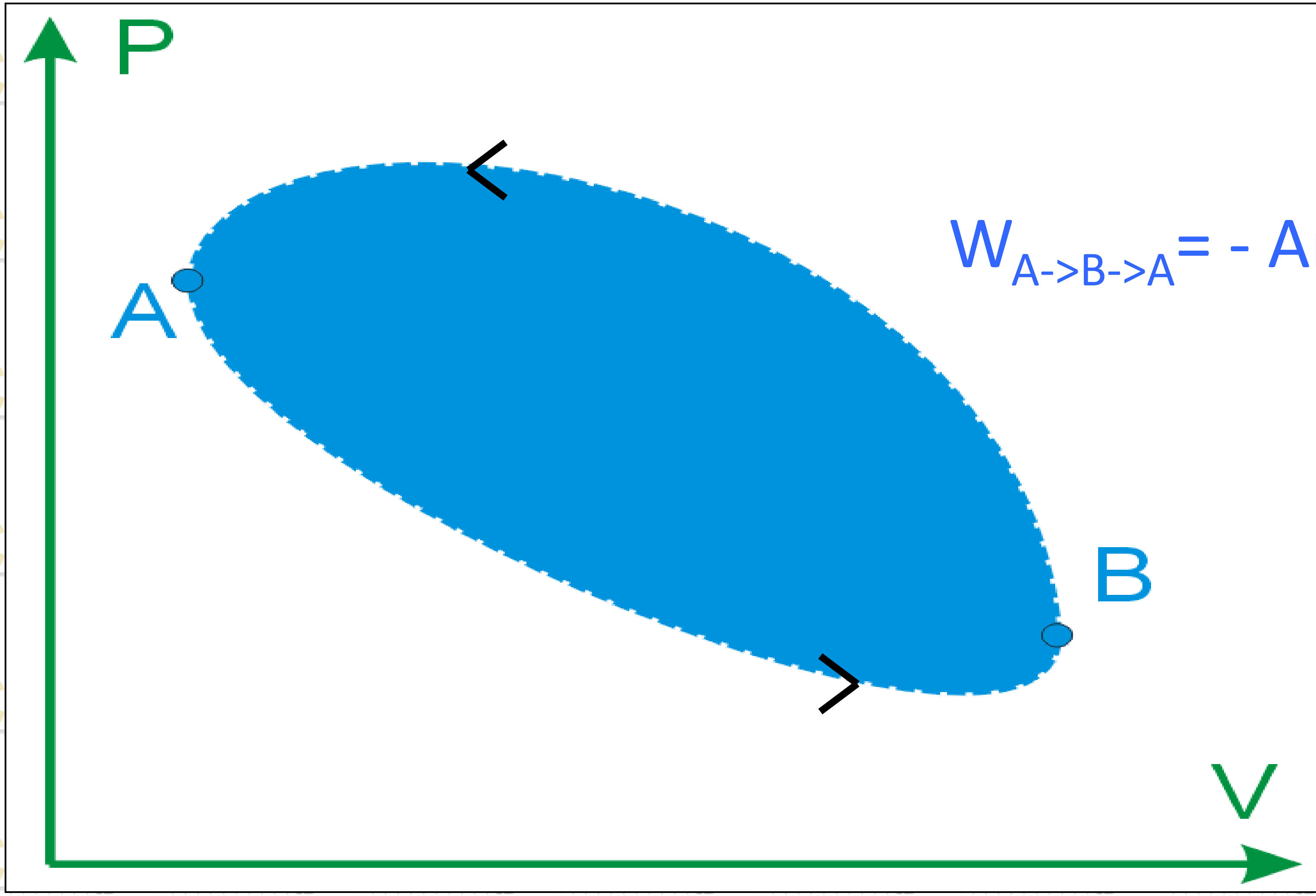
$W_{B->A}$ = - Area

$W_{A->B}$ = Area

(work done by gas)

Work from closed cycles

Reverse the cycle, make it counter clockwise



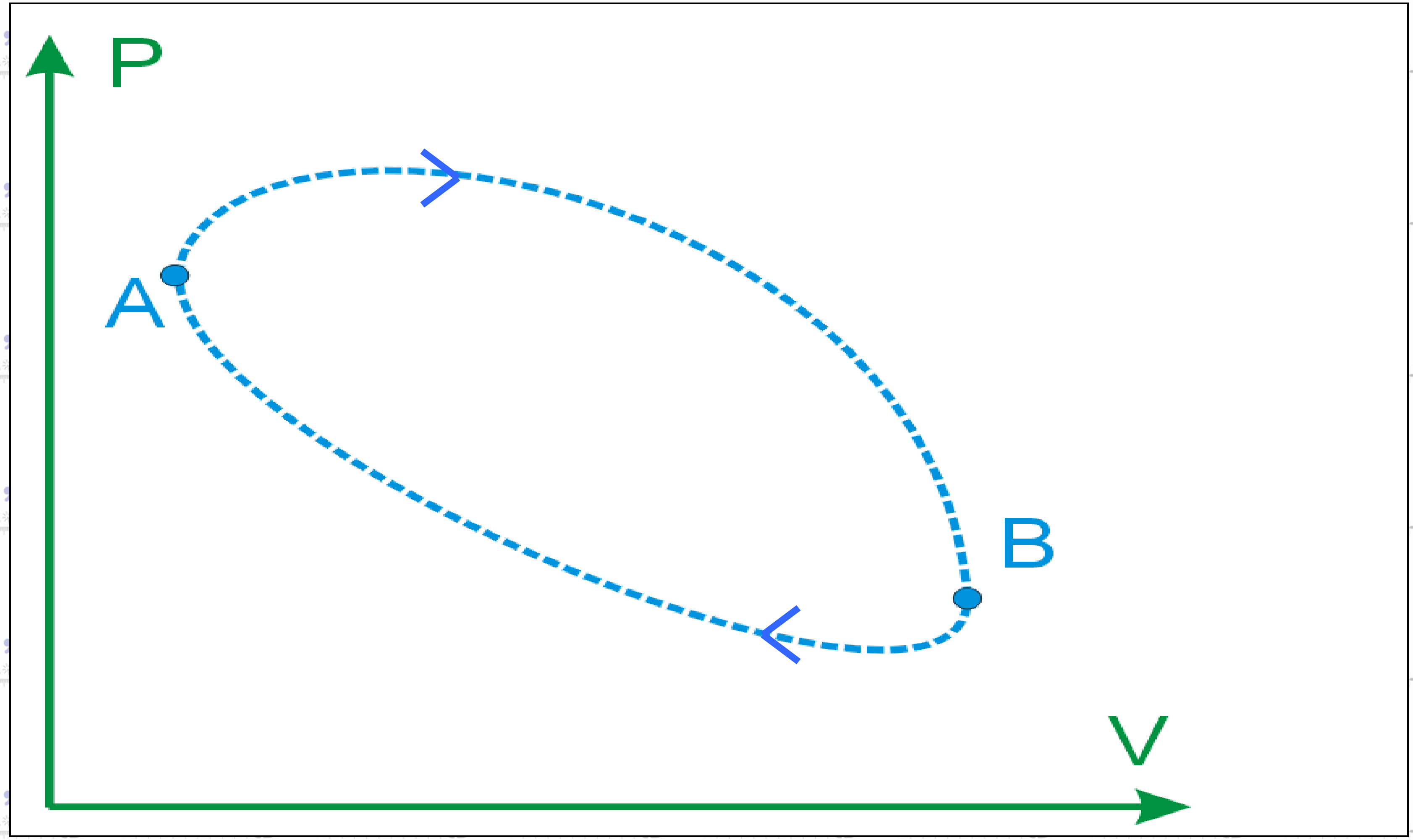
$W_{A\text{->}B\text{->}A} = -\ \text{Area}$

(work done by gas)

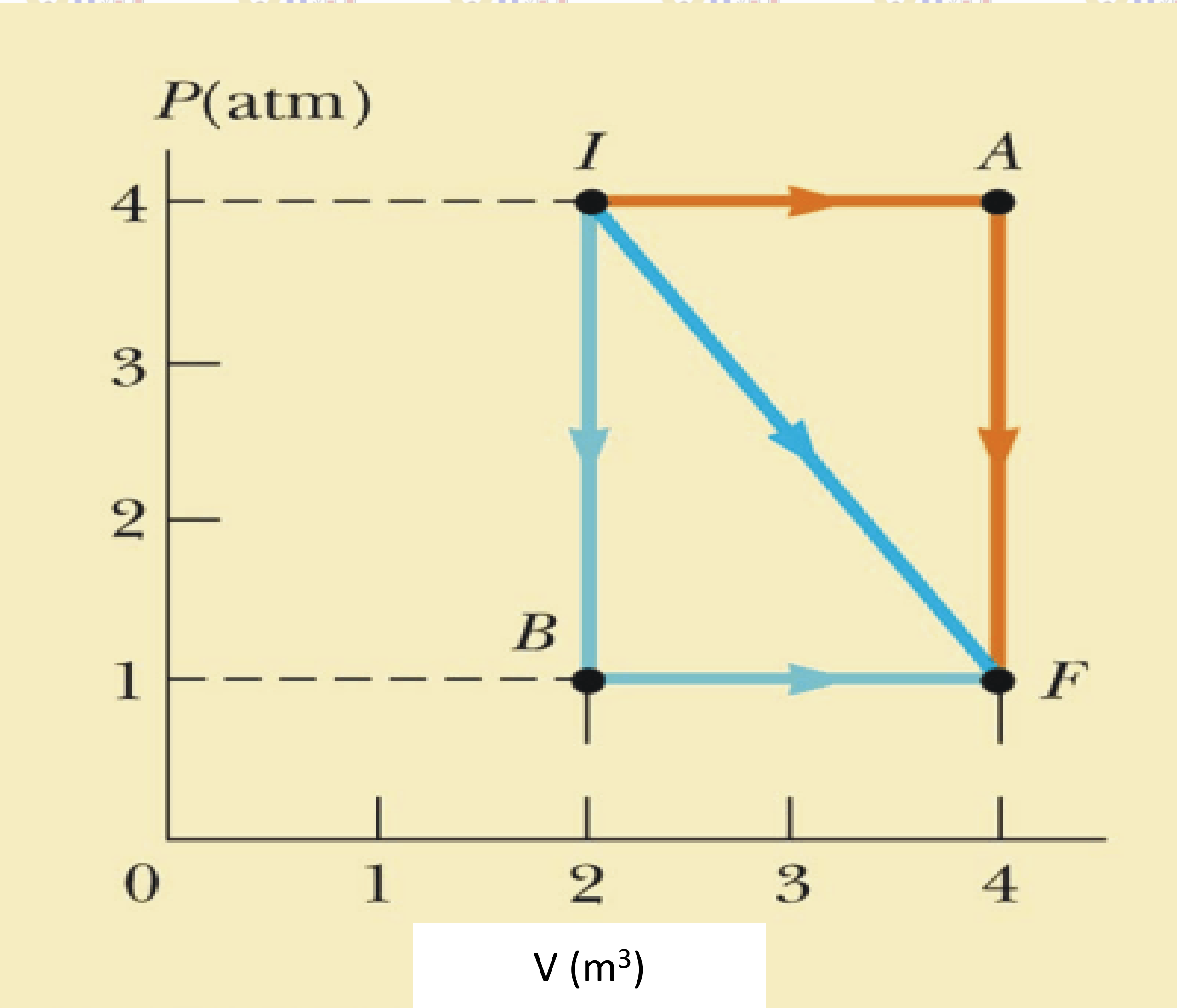
Internal Energy in closed cycles

In closed cycles

$$\Delta U = 0$$



Example



a) What amount of work is performed by the gas in the cycle IAFI?

W=3.04x10^5 J

b) How much heat was inserted into the gas in the cycle IAFI?

Q = 3.04x10^5 J

c) What amount of work is performed by the gas in the cycle IBFI?

W = -3.04x10^5 J

Example 12.7

Consider a monotonic ideal gas.

a) What work was done by the gas from A to B?

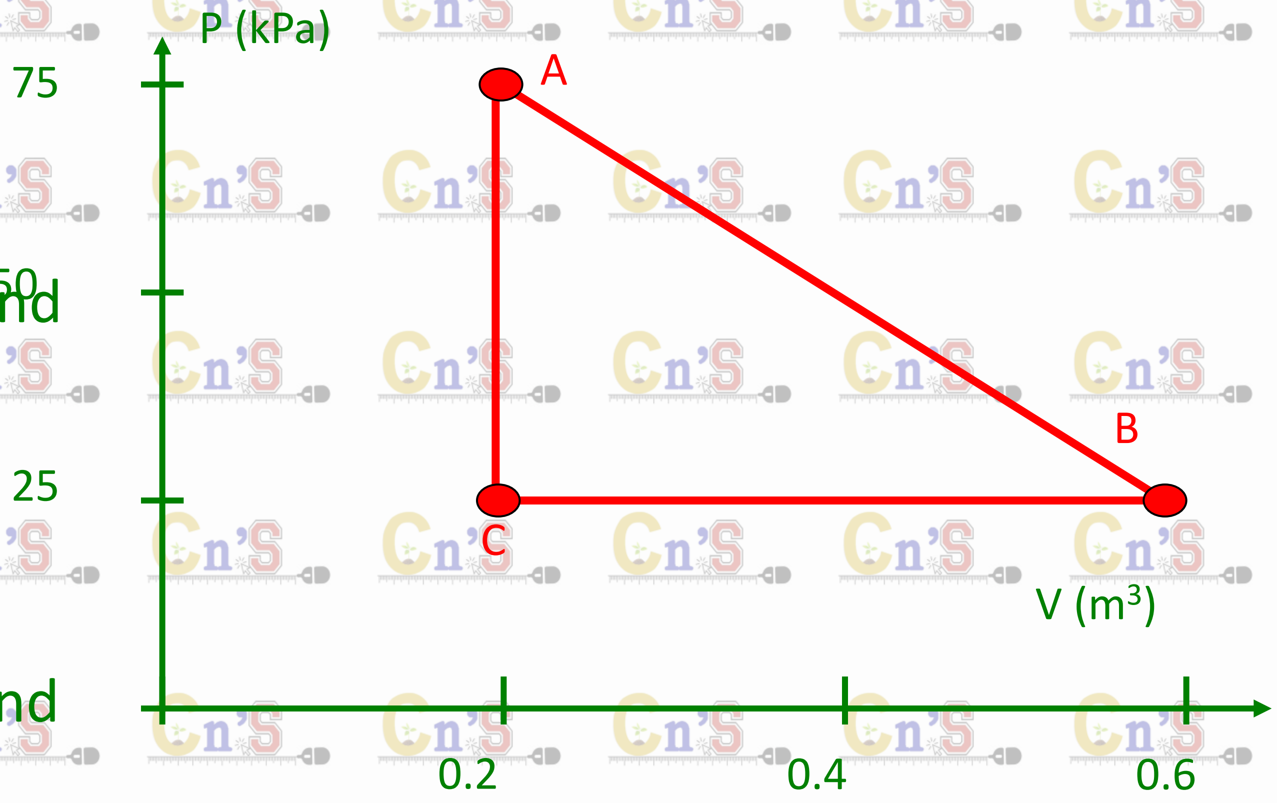
b) What heat was added to the gas between A and B?

c) What work was done by the gas from B to C?

d) What heat was added to the gas between B and C?

e) What work was done by the gas from C to A?

f) What heat was added to the gas from C to A?



Example Continued

Take solutions from last problem and find:

a) Net work done by gas in the cycle
b) Amount of heat added to gas

$W_{AB} + W_{BC} + W_{CA} = 10{,}000\ J$
$Q_{AB} + Q_{BC} + Q_{CA} = 10{,}000\ J$

This does NOT mean that the engine is 100% efficient!

2.Work done during free expansion

$$W = 0$$

3.Work done during irreversible process.

$$W = -p\Delta V$$

Different equations for 1st law of thermodynamics

1.A process carried out at constant volume

$$\Delta U = q_v$$

2.Isothermal process

$$q = -w$$

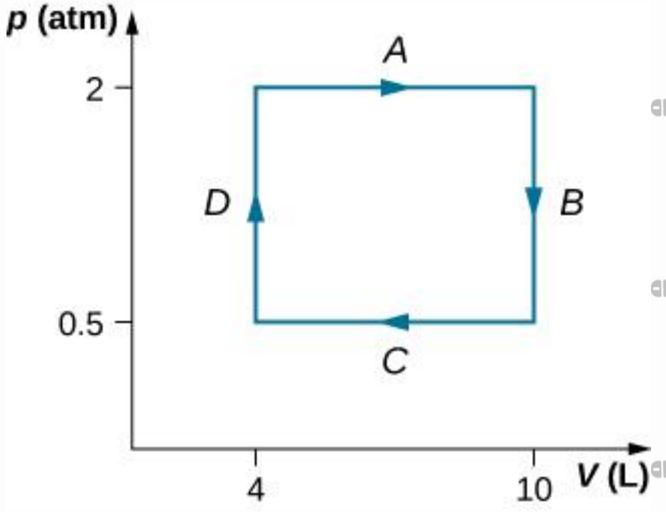
4.Isothermal irreversible process

$$q = P_{ex}(V_f - V_i)$$

5.Adiabatic process

$$\Delta U = W_{ad}$$

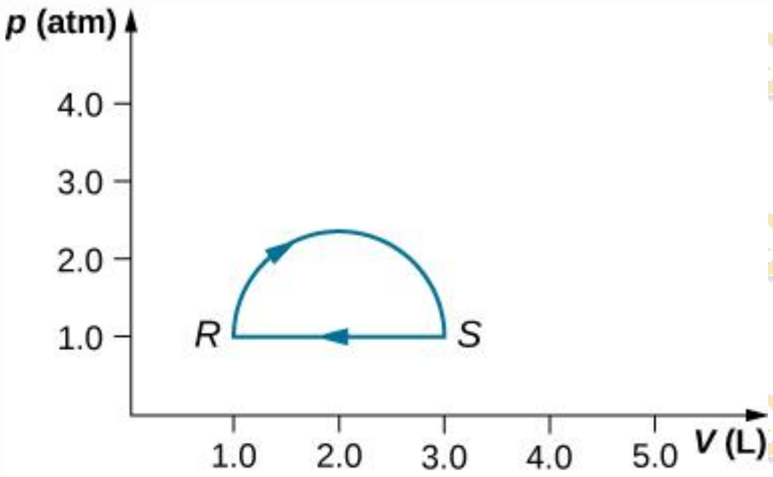
- **31.** A dilute gas at a pressure of 2.0 atm and a volume of 4.0 L is taken through the following quasi-static steps: (a) an isobaric expansion to a volume of 10.0 L, (b) an isochoric change to a pressure of 0.50 atm, (c) an isobaric compression to a volume of 4.0 L, and (d) an isochoric change to a pressure of 2.0 atm. Show these steps on a pV diagram and determine from your graph the net work done by the gas.



31. *W*=900*J*

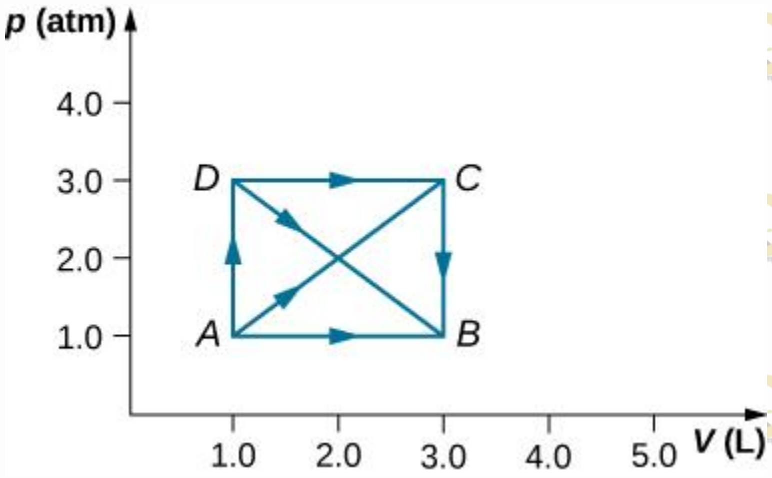
29. (a) Calculate the work done by the gas along the closed path shown below. The curved section between R and S is semicircular.
(b) If the process is carried out in the opposite direction, what is the work done by the gas?

- **29.** a. 160 J; b. –160 J

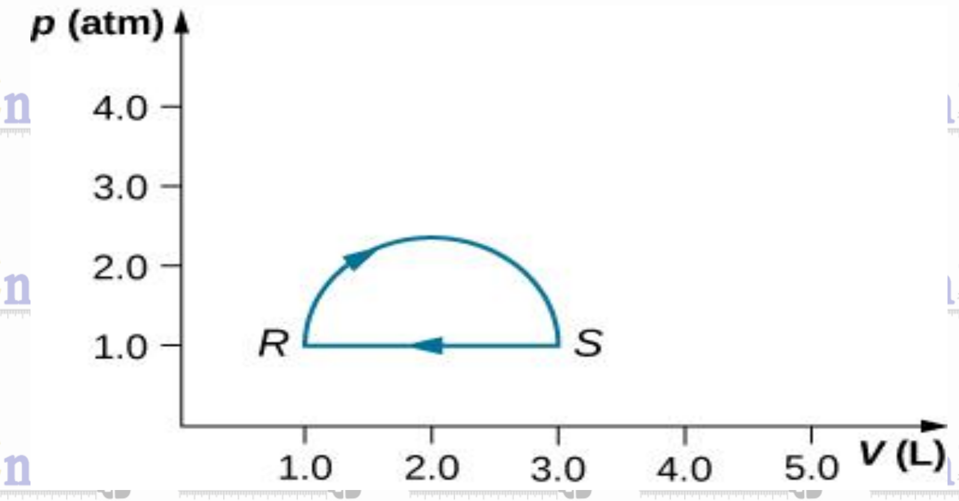


41. a. 600 J;
b. 600 J;
c. 800 J

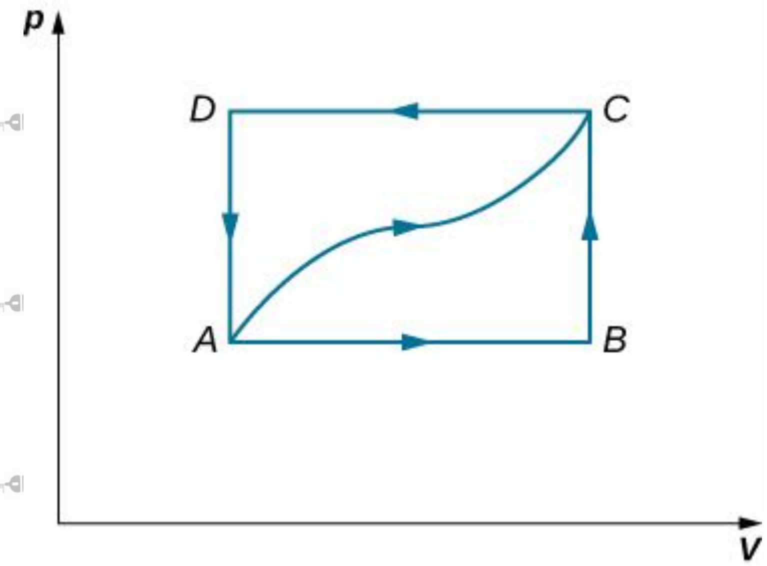
- **41.** As shown below, if the heat absorbed by the gas along AB is 400 J, determine the quantities of heat absorbed along
- (a) ADB; (b) ACB; and (c) ADCB.



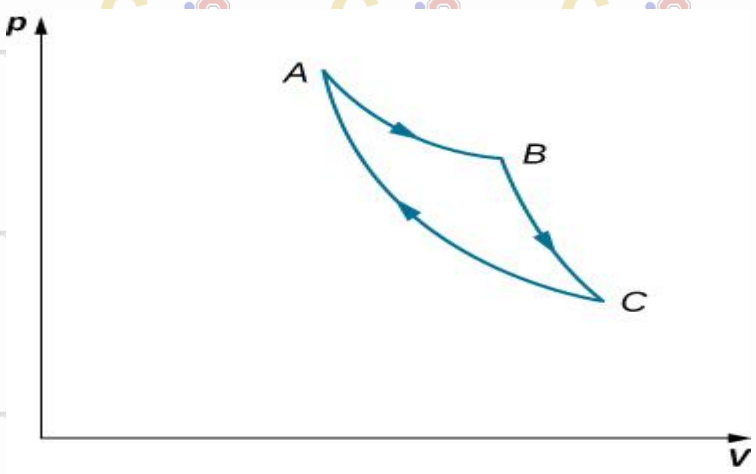
43. (a) What is the change in internal energy for the process represented by the closed path shown below? (b) How much heat is exchanged? (c) If the path is traversed in the opposite direction, how much heat is exchanged? **43.** a. 0;
b. 160 J;
c. –160 J

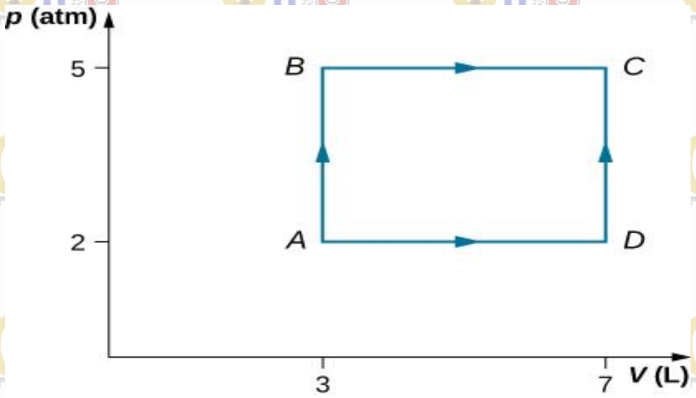


- **45.** When a gas expands along AB (see below), it does 500 J of work and absorbs 250 J of heat. When the gas expands alongAC, it does 700 J of work and absorbs 300 J of heat.
- (a) How much heat does the gas exchange along BC?
- (b) When the gas makes the transmission from C to A along CDA, 800 J of work are done on it from C to D. How much heat does it exchange . **45.** a. –150 J;
- b. –400 J



- **57.** An ideal gas expands isothermally along AB and does 700 J of work (see below).
- (a) How much heat does the gas exchange along AB?
- (b) The gas then expands adiabatically along BC and does 400 J of work. When the gas returns to A along CA, it exhausts 100 J of heat to its surroundings. How much work is done on the gas along this path?
- **57.** a. 700 J;
- b. 500 J



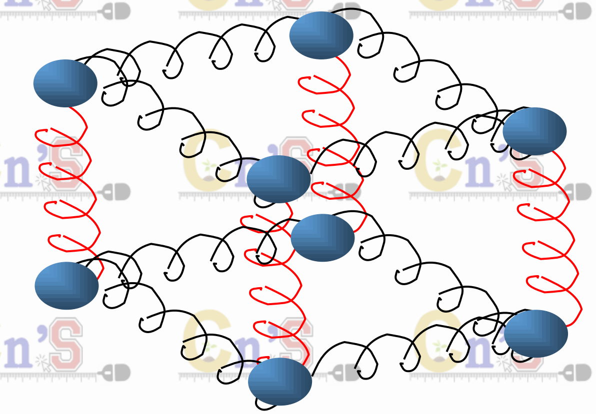


- **79.** Consider the process shown below. During steps AB and BC, 3600 J and 2400 J of heat, respectively, are added to the system.
- (a) Find the work done in each of the processes AB, BC, AD, and DC.
- (b) Find the internal energy change in processes AB and BC.
- (c) Find the internal energy difference between states C and A.
- (d) Find the total heat added in the ADCprocess.
- (e) From the information given, can you find the heat added in process AD? Why or why not?

- **Additional Problems**
- **79.** a. $W_{AB}=0, W_{BC}=2026J, W_{AD}=810.4J, W_{DC}=0$;
- b. $\Delta E_{AB}=3600J, \Delta E_{BC}=374J$
- ;
- c. $\Delta E_{AC}=3974J$
- ;
- d. $Q_{ADC}=4784J$
- ;
- e. No, because heat was added for both parts *AD* and *DC*. There is not enough information to figure out how much is from each segment of the path.

Internal Energy

Internal energy (also called thermal energy) is the energy an object or substance is due to the kinetic and potential energies associated with the random motions of all the particles that make it up. The kinetic energy is, of course, due to the motion of the particles. To understand the potential energy, imagine a solid in which all of its molecules are bound to its neighbors by springs. As the molecules vibrate, the springs are compressed and stretched. (Liquids and gases are not locked in a lattice structure like this.)



The hotter something is, the faster its molecules are moving or vibrating, and the higher its temperature. Temperature is proportional to the average kinetic energy of the atoms or molecules that make up a substance.

Internal Energy vs. Heat

The term *heat* refers is the energy that is transferred from one body or location due to a difference in temperature. This is similar to the idea of *work*, which is the energy that is transferred from one body to another due to forces that act between them. Heat is internal energy when it is transferred between bodies.

Technically, a hot potato does not possess heat; rather it possesses a good deal of internal energy on account of the motion of its molecules. If that potato is dropped in a bowl of cold water, we can talk about heat: There is a heat flow (energy transfer) from the hot potato to the cold water; the potato's internal energy is decreased, while the water's is increased by the same amount.

Temperature vs. Internal Energy

Temperature and internal energy are related but not the same thing. Temperature is directly proportional to the average molecular kinetic energy*. Note the word *average* is used, not *total*.

Consider a bucket of hot water and a swimming pool full of cold water. The hot water is at a higher temperature, but the pool water actually has more internal energy! This is because, even though the average kinetic energy of the water molecules in the bucket is much greater than that of the pool, there are thousands of times more molecules in the pool, so their total energy is greater. A swarm of 1000 slow moving bees could have more total kinetic energy than a dozen fast moving, hyperactive bees buzzing around like crazy. One fast bee has more kinetic energy than a slow one, but there are a lot more slow ones.

* *true for gases, approximately true for solids and liquids whose molecules interact with each other more.*

Temperature vs. Internal Energy (cont.)

Which has more internal energy, a bucket of hot water or a bucket of cold water? *answer:*

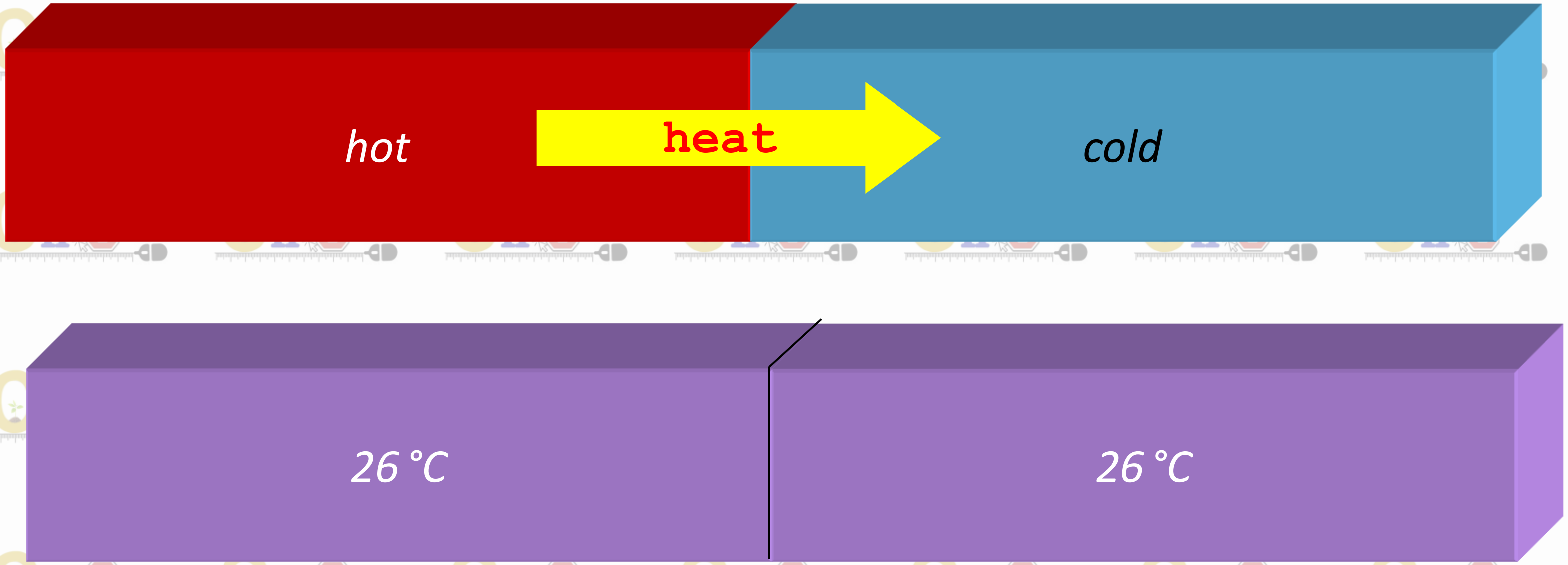
The bucket of hot water has more internal energy, at least if the buckets contain the same amount of water.

Internal energy depends on the amount (mass) of substance and the kinetic energy of the molecules of the substance.

Temperature only depends on the molecules' kinetic energy; it is independent of mass.

Thermal Equilibrium

Two bodies are said to be at thermal equilibrium if they are at the same temperature. This means there is no net exchange of thermal energy between the two bodies. The top pair of objects are in contact, but since they are at different temps, they are not in thermal equilibrium, and energy is flowing from the hot side to the cold side.



No net heat flow

The two purple objects are at the same temp and, therefore are in thermal equilibrium. There is no net flow of heat energy here.

Thermodynamic Systems

1. Consider these scenarios and state whether work is done by the system on the environment (SE) or by the environment on the system (ES):

(a) opening a carbonated beverage;

(b) filling a flat tire;

(c) a sealed empty gas can expands on a hot day, bowing out the walls.

- [3.2 Work, Heat, and Internal Energy]()
- **2.** Is it possible to determine whether a change in internal energy is caused by heat transferred, by work performed, or by a combination of the two?
- **3.** When a liquid is vaporized, its change in internal energy is not equal to the heat added. Why?
- **4.** Why does a bicycle pump feel warm as you inflate your tire?
- **5.** Is it possible for the temperature of a system to remain constant when heat flows into or out of it? If so, give examples.

3.3 First Law of Thermodynamics

- **6.** What does the first law of thermodynamics tell us about the energy of the universe?
- **7.** Does adding heat to a system always increase its internal energy?
- **3.4 Thermodynamic Processes**
- **9.** When a gas expands isothermally, it does work. What is the source of energy needed to do this work?
- **11.** It is unlikely that a process can be isothermal unless it is a very slow process. Explain why. Is the same true for isobaric and isochoric processes? Explain your answer. **13.** Most materials expand when heated. One notable exception is water between **0°C** and **4°C**, which actually decreases in volume with the increase in temperature. Which is greater for water in this temperature region, C_p or C_V

- **3.6 Adiabatic Processes for an Ideal Gas**
- **15.** Is it possible for **γ** to be smaller than unity?
- **16.** Would you expect **γ** to be larger for a gas or a solid? Explain.
- **17.** There is no change in the internal energy of an ideal gas undergoing an isothermal process since the internal energy depends only on the temperature. Is it therefore correct to say that an isothermal process is the same as an adiabatic process for an ideal gas? Explain your answer.
- **18.** Does a gas do any work when it expands adiabatically? If so, what is the source of the energy needed to do this work?