Engineering Thermodynamics

Thermodynamics

- Energy can be viewed as ability to cause changes
- Thermodynamics can be defined as the science of energy.
- The name *thermodynamics* stems from the Greek words *therme* (heat) and *dynamis* (power), which is most descriptive of the early efforts to convert heat into power.
- Today the same name is broadly interpreted to include all aspects of energy and energy transformations including power generation, refrigeration, and relationships among the properties of matter.
- Thermodynamics is an axiomatic science which deals with the relation among heat, work and properties of system which are in equilibrium. It describes state and changes in state of physical systems.

Laws of Thermodynamics

- Thermodynamics basically entails four laws or axioms known as Zeroth, First, Second and Third law of thermodynamics.
- Zeroth Law: deals with thermal equilibrium and establishes a concept of temperature
- First Law:- an expression of the conservation of energy principle
- **Second Law:-** asserts that energy has *quality* as well as *quantity,* and actual processes occur in the direction of decreasing quality of energy.
- Third Law:- Defines the absolute zero of entropy.

- The **first law of thermodynamics** is simply an expression of the conservation of energy principle, and it asserts that *energy* is a thermodynamic property.
- The second law of thermodynamics asserts that energy has quality as well as quantity, and actual processes occur in the direction of decreasing quality of energy.
- For example, a cup of hot coffee left on a table eventually cools, but a cup of cool coffee in the same room never gets hot by itself (Fig. 1–3). The high-temperature energy of the coffee is degraded (transformed into a less useful form at a lower temperature) once it is transferred to the surrounding air



FIGURE 1–3

Heat flows in the direction of decreasing temperature.

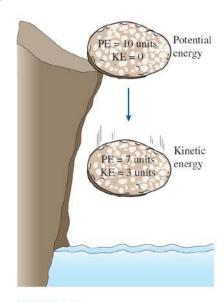


FIGURE 1–1
Energy cannot be created or destroyed; it can only change forms (the first law).

- It is well-known that a substance consists of a large number of particles called *molecules*. The properties of the substance naturally depend on the behavior of these particles.
- For example, the pressure of a gas in a container is the result of momentum transfer between the molecules and the walls of the container.
- However, one does not need to know the behavior of the gas particles to determine the pressure in the container. It would be sufficient to attach a pressure gage to the container.
- This macroscopic approach to the study of thermodynamics that does not require a knowledge of the behavior of individual particles is called classical thermodynamics. It provides a direct and easy way to the solution of engineering problems.
- A more elaborate approach, based on the average behavior of large groups of individual particles, is called **statistical thermodynamics**. This microscopic approach is rather involved and is used in this text only in the supporting role.

Application Areas of Thermodynamics

- All activities in nature involve some interaction between energy and matter; thus, it is hard to imagine an area that does not relate to thermodynamics in some manner.
- Generation of body heat and its rejection in the environment
- Heating and air-conditioning system,
- Humidifier
- Pressure cooker
- Automotive engines, rocket, jets and power plants
- Should have analysis that involves thermodynamics

System, Boundary and Surroundings

- •A thermodynamic system, or simply system, is defined as a quantity of matter or a region in space chosen for study.
- The region outside the system is called the surroundings.
- •The real or imaginary surface that separates the system from is called the boundary.
- The boundary of a system may be fixed or movable.

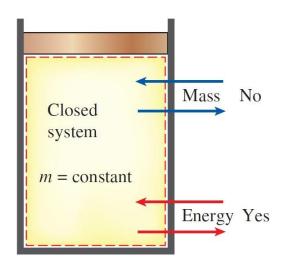
its surroundings
Surroundings

System

•Systems may be considered to be *closed* or *open*, depending on whether a fixed mass or a fixed volume in space is chosen for study.

Closed system

- A closed system (also known as a control mass or just system when the context makes it clear) consists of a fixed amount of mass, and no mass can cross its boundary.
- That is, no mass can enter or leave a closed system, as shown in. But energy, in the form of heat or work, can cross the boundary; and the volume of a closed system does not have to be fixed.
- If, as a special case, even energy is not allowed to cross the boundary, that system is called an isolated system.



Mass cannot cross the boundaries of a

closed system, but energy can.

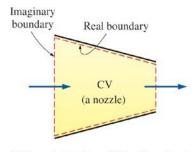
Moving boundary Gas 2 kg Gas $1.5 \, \mathrm{m}^3$ 2 kg 1 m^3 Fixed boundary FIGURE 1-20 A closed system with a moving boundary.

Closed System

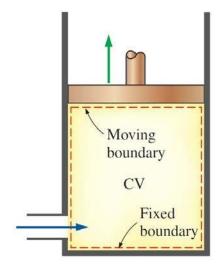
- •A closed system consists of a fixed amount of mass and no mass may cross the system boundary. The closed system boundary may move.
- •Examples of closed systems are sealed tanks and piston cylinder devices (note the volume does not have to be fixed). However, energy in the form of heat and work may cross the boundaries of a closed system.

Open System

- A large number of engineering problems involve mass flow in and out of a system and, therefore, are modeled as *control volumes*.
- An open system, or control volume, has mass as well as energy crossing the boundary, called a control surface.
- A water heater, a car radiator, a turbine, and a compressor all involve mass flow and should be analyzed as control volumes (open systems) instead of as control masses (closed systems).
- In general, any arbitrary region in space can be selected as a control volume.



(a) A control volume (CV) with real and imaginary boundaries



(b) A control volume (CV) with fixed and moving boundaries as well as real and imaginary boundaries

Isolated System Boundary Heat = 0Mass = 0Mass System Surr 3 Across Mass Isolated Surr 1 Boundary

Isolated System

An isolated system is a general system of fixed mass where no heat or work may cross the boundaries.

An isolated system is a closed system with no energy crossing the boundaries and is normally a collection of a main system and its surroundings that are exchanging mass and energy among themselves and no other system.

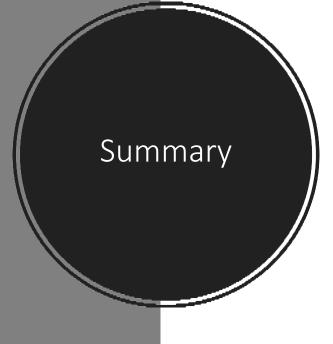
- Adiabatic system: An adiabatic system is one which is thermally insulated from its surroundings, it can however, exchanges work with its surroundings. If it does not, it becomes an isolated system.
 - An example of an adiabatic process is the vertical flow of air in the atmosphere; air expands and cools as it rises, and contracts and grows warmer as it descends. Another example is when an interstellar gas cloud expands or contracts.

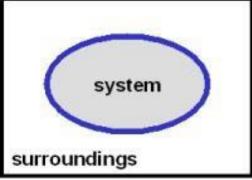
• Homogeneous System:

A homogeneous thermodynamic system is defined as the one whose chemical composition and physical properties are the same in all parts of the system, or change continuously from one point to another.

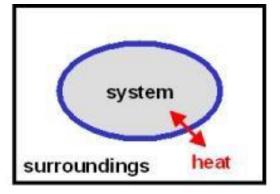
 A homogeneous system can be exemplified by imagining a column of atmospheric air, which is a mixture of a number of gases, mainly nitrogen and oxygen. In a system of this kind, acted upon by the force of gravity, both the composition of the system and its physical properties will continuously change from one point to another.

- **Heterogeneous System:** A heterogeneous system is denned as one consisting of two or more homogeneous bodies. There chemical and physical properties are not same throughout the system and will change abruptly while moving to different state.
- The homogeneous bodies of a heterogeneous system are referred to as phases. Each phase is separated from other phases by interfaces, or boundaries, and in passing over such a boundary the chemical composition of the substance or its physical properties abruptly change. An example of a heterogeneous system is water with ice floating in it. This system has two homogeneous bodies, water and ice. The chemical composition of the two phases is the same, but their physical properties differ drastically.
- Another example of a heterogeneous system is the content of a sealed steel tube containing liquid mercury, liquid ethyl alcohol, and a mixture of saturated vapors of the alcohol and mercury. This heterogeneous system comprises three phases. The first phase is the liquid mercury, the second is the liquid ethyl alcohol, and the third phase is represented by the mixture of saturated vapors. Here the chemical compositions and the physical properties of all phases are different.

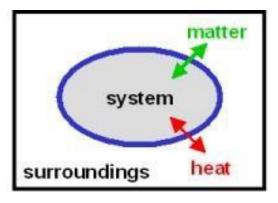




- "Isolated" system:
- no exchange of matter
- no exchange of heat



- "Closed" system:
- no exchange of matter
- can exchange heat energy



- "Open" system:
- · can exchange matter
- can exchange heat energy

14

Properties of a System

- Any characteristic of a system is called a property.
- •The property is independent of the path used to arrive at the system condition.
- •Some thermodynamic properties are pressure *P*, temperature *T*, volume *V*, and mass *m*.
- Properties may be intensive or extensive.
- •Extensive properties are those that vary directly with size--or extent--of the system.
- •An extensive property is a physical quantity whose value is proportional to the size of the <u>system</u> it describes, or to the quantity of matter in the system.

Extensive Properties

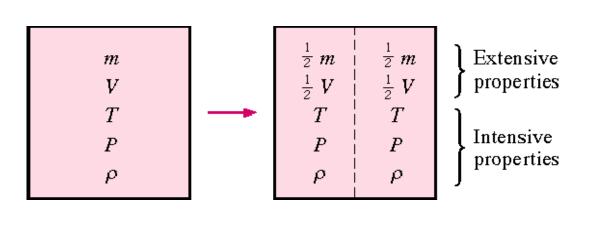
- •Extensive properties are those that vary directly with size--or extent--of the system.
- •An extensive property is a physical quantity whose value is proportional to the size of the system it describes, or to the quantity of matter in the system.
- Some Extensive Properties

```
a.Mass, m
b.Volume, V
c.total energy, E d.Amount
of substance, n e.Heat
capacity, (Cp in J/K)
f. mass dependent property
```

Intensive properties

- Intensive properties are those that are independent of size.
- An intensive property is a physical quantity whose value does not depend on the amount of the substance for which it is measured.
- Some Intensive Properties

```
a.temperature
b.pressure
c.age
d.Color
e.Specific heat capacity (Cp in J/kg.K)
f. any mass independent property
```



Extensive properties per unit mass are intensive properties.

For example, the specific volume *v*, defined as

$$v = \frac{Volume}{mass} = \frac{V}{m} \left(\frac{m^3}{kg} \right)$$

and density ρ , defined as

$$\rho = \frac{mass}{volume} = \frac{m}{V} \left(\frac{kg}{m^3} \right)$$

State, Equilibrium, Process, and Properties State

- **State** It is the condition of the system at an instant of time as described or measured by its properties. Or each unique condition of the system is called **state**.
- At a given state, all the properties of a system have fixed values. If the value of even one property changes, the state will change to a different one.

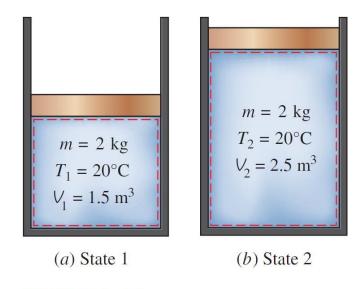
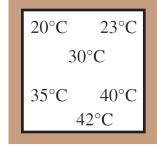
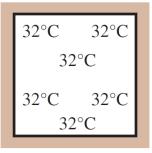


FIGURE 1–26
A system at two different states.





(a) Before

(b) After

FIGURE 1-27

A closed system reaching thermal equilibrium.

- Thermodynamics deals with *equilibrium* states. The word **equilibrium** implies a state of balance. In an equilibrium state there are no unbalanced potentials (or driving forces) within the system. A system in equilibrium experiences no changes when it is isolated from its surroundings.
- A system is said to be in thermodynamic equilibrium if it maintains thermal (uniform temperature), mechanical (uniform pressure), phase (the mass of two phases, e.g., ice and liquid water, in equilibrium) and chemical equilibrium.

A system is in thermal equilibrium

- if the temperature is the same throughout the entire system.
- That is, the system involves no temperature differential, which is the driving force for heat flow.

A system is in Mechanical equilibrium

- If here are no unbalanced forces within the system or between the surroundings.
- It is mainly related to pressure, and a system is in mechanical equilibrium if there is no change in pressure at any point of the system with time.
- If a system involves two phases, it is in phase equilibrium
 - when the mass of each phase reaches an equilibrium level and stays there.

• A system is in chemical equilibrium

• if its chemical composition does not change with time, that is, no chemical reactions occur. A system will not be in equilibrium unless all the relevant equilibrium criteria are satisfied.

Process

- A process occurs when the system undergoes a change in a state or an energy transfer at a steady state.
- Any change that a system undergoes from one equilibrium state to another is called a process, and the series of states through which a system passes during a process is called the path of the process.
- To describe a process completely, one should specify the initial and final states of the process, as well as the path it follows, and the interactions with thesurroundings.

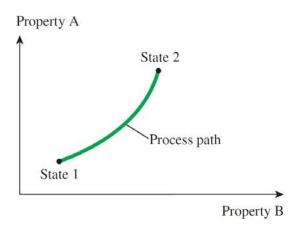
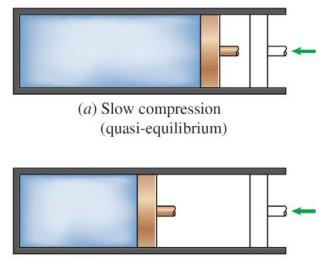


FIGURE 1–29
A process between states 1 and 2 and the process path.

- When a process proceeds in such a manner that the system remains infinitesimally close to an equilibrium state at all times, it is called a quasistatic, or quasi-equilibrium, process.
- A quasi-equilibrium process can be viewed as a sufficiently slow process that allows the system to adjust itself internally so that properties in one part of the system do not change any faster than those at other parts.



(b) Very fast compression (nonquasi-equilibrium)

FIGURE 1-30

Quasi-equilibrium and nonquasiequilibrium compression processes.

- It should be pointed out that a quasi-equilibrium process is an idealized process and is not a true representation of an actual process. But many actual processes closely approximate it, and they can be modeled as quasi-equilibrium with negligible error.
- Engineers are interested in quasi equilibrium processes for two reasons. First, they are easy to analyze; second, work-producing devices deliver the most work when they operate on quasi-equilibrium processes. Therefore, quasi-equilibrium processes serve as standards to which actual processes can be compared.

- Process diagrams plotted by employing thermodynamic properties as coordinates are very useful in visualizing the processes. Some common properties that are used as coordinates are temperature *T*, pressure *P*, and volume *V* (or specific volume *v*). Figure 1–31 shows the *P-V* diagram of a compression process of a gas.
- Note that the process path indicates a series of equilibrium states through which the system passes during a process and has significance for quasi-equilibrium processes only. For non quasi-equilibrium processes, we are not able to characterize the entire system by a single state, and thus we cannot speak of a process path for a system as a whole. A non quasi-equilibrium process is denoted by a dashed line between the initial and final states instead of a solid line.

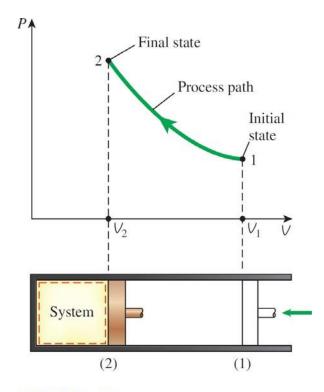


FIGURE 1–31
The *P-V* diagram of a compression process.

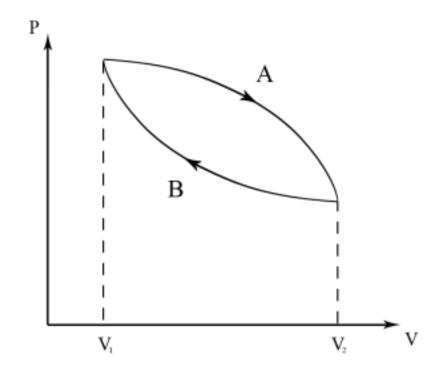
Processes maintaining constant property

- •Prefix iso- is used to designate a process for which a particular property remains constant.
- •In most of the processes that we will study, one thermodynamic property is held constant. Some of these processes are

Process	Property held constant
isobaric	pressure
isothermal	temperature
Isochoric/ isometric	volume
isentropic	entropy

CYCLE

- A system is said to have undergone a **cycle** if it returns to its initial state at the end of the process. That is, for a cycle the initial and final states are identical.
- Cycle: Any process or series of processes whose end states are identical i.e. the system returns to its original thermodynamic state
- Figure shows a cycle in which a system commencing at condition '1' changes in pressure and volume through a path and return to its initial condition '1'.



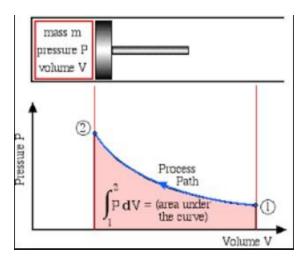
Point Function

when two properties locate a point on the graph (co-ordinate axes) then those properties are called as **point function.** A variable (as the temperature of the air) each value of which is associated with and determined by the position of some point in space.

$$\int_{1}^{2} dV = V_2 - V_1$$

• **Path functions** There are certain quantities which cannot be located on a graph by but are given by the area or so, on that graph. In that case, the area on the graph, to the particular process, *is a function of the path of the process*. Such quantities are

called **path functions**.



Thus
$$\int_{1}^{2} \delta Q \neq Q_2 - Q_1$$
 and is shown as ${}_{1}Q_2$ or Q_{1-2}

Similarly
$$\int_{1}^{2} \delta W \neq W_2 - W_1$$
, and is shown as $_1W_2$ or W_{1-2}

N.B. The operator δ is used to denote inexact differentials and operator d is used to denote exact differentials.

Plot process on P vs V axes: $W = P_0 \Delta V = \text{area under path of process in the P-V plot this applies}$ to any path, not just for constant-P: Work done = area under path on a P vs V plot

- Path functions have inexact differentials designated by the symbol d. Therefore, a differential amount of heat or work is represented by d*Q* or d*W*, respectively, instead of d*Q* or d*W*.
- Properties, however, are **point functions** (i.e., they depend on the state only, and not on how a system reaches that state), and they have **exact differentials** designated by the symbol *d*.
- A small change in volume, for example, is represented by dV, and the total volume change during a process between states 1 and 2 is

$$\int_{1}^{2} dV = V_2 - V_1 = \Delta V$$

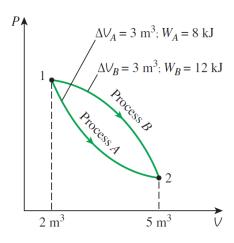


FIGURE 2–22
Properties are point functions;
but heat and work are path
functions (their magnitudes
depend on the path followed).

• That is, the volume change during process 1–2 is always the volume at state 2 minus the volume at state 1, regardless of the path followed (Fig. 2–22). The total work done during process 1–2, however, is

$$\int_{1}^{2} \delta W = W_{12} \quad (not \ \Delta W)$$

 That is, the total work is obtained by following the process path and adding the differential amounts of work (dW) done along the way. The integral of dW is not W2 - W1 (i.e., the work at state 2 minus work at state 1), which is meaningless since work is not a property and systems do not possess work at a state.

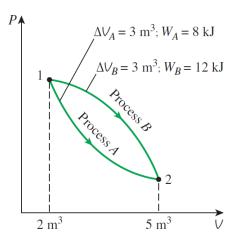
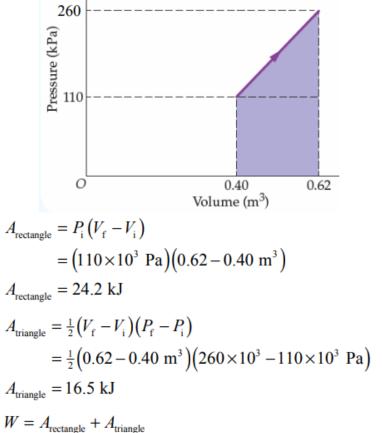


FIGURE 2–22
Properties are point functions;
but heat and work are path
functions (their magnitudes
depend on the path followed).

Example:

• Ideal gas expands from $V_{initial} = 0.40$ m₃ to $V_{final} = 0.62$ m₃ while its pressure increases linearly from $P_i = 110$ kPa to $P_f = 230$ kPa. Calculate the work done.



 $= 24.2 + 16.5 \text{ kJ} = 40.7 \text{ kJ} = |4.1 \times 10^4 \text{ J}|$

• The terms *steady* and *uniform* are used frequently in engineering, and thus it is important to have a clear understanding of their meanings. The term *steady* implies *no change with time*. The opposite of steady is unsteady, or transient. The term uniform, however, implies no change with location over a specified region. These meanings are consistent with their everyday use (uniform properties, etc.).

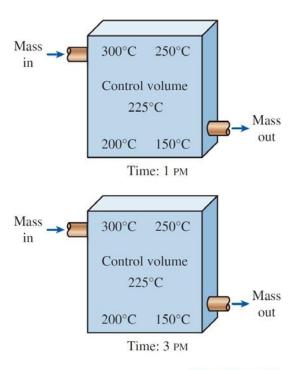


FIGURE 1-32

During a steady-flow process, fluid properties within the control volume may change with position but not with time.

- A large number of engineering devices operate for long periods of time under the same conditions, and they are classified as *steady-flow devices*.
- Processes involving such devices can be represented reasonably well by a somewhat idealized process, called the **steady-flow process**, which can be defined as a *process during which a fluid flows through a control volume steadily*

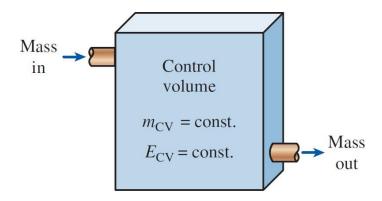


FIGURE 1-33

Under steady-flow conditions, the mass and energy contents of a control volume remain constant.

Pressure

- Pressure is defined as a normal force exerted by a fluid per unit area.
- Normally, we speak of pressure when we deal with a gas or a liquid. The counterpart of pressure in solids is *normal stress*. Note, however, that pressure is a scaler quantity while stress is a tensor. Since pressure is defined as force per unit area, it has the unit of newtons per square meter (N/m2), which is called a **pascal** (Pa).

That is, 1 Pa = $1 N/m^2$

Temperature

- The temperature is a thermal state of a body which distinguishes a hot body from a cold body. The temperature of a body is proportional to the stored molecular energy. i.e., the average molecular kinetic energy of the molecules in a system.
- A particular molecule does not have a temperature, it has energy. The gas as a system has temperature.
- Instruments for measuring ordinary temperatures are known as **thermometers** and those for measuring high temperatures are known as **pyrometers**.
 - **Absolute zero temperature:** It has been found that a gas will not occupy any volume at a certain temperature. This temperature is known as **absolute zero- temperature.**
 - Absolute temperature= Thermometer reading in -273.15 C or O K
 - Absolute temperature can also be represented in degrees kelvin, denoted by K (SI unit).

- They are related to absolute temperature scales by
 - T(K) = T(C) + 273.15
 - T(R) = T(F) 1 459.67
- The magnitudes of each division of 1 K and 1°C are identical, and so are the magnitudes of each division of 1 R and

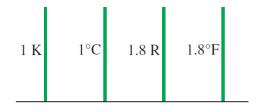
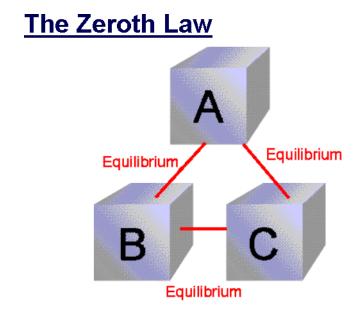


FIGURE 1–38
Comparison of magnitudes of various temperature units.

Zeroth law of thermodynamics

- The zeroth law of thermodynamics states that if two bodies are in thermal equilibrium with a third body, they are also in thermal equilibrium with each other.
- it serves as a basis for the validity of temperature measurement. By replacing the third body with a thermometer, the zeroth law can be restated as *two bodies are in thermal equilibrium if both have the same temperature reading even if they are not in contact*.
- According to this law if two systems named as A and B is in equilibrium with the third system named as C then A and C are also in thermal equilibrium



Laws of thermodynamics

- First Law
- Second Law
- Third Law

- The *first law of thermodynamics*, also known as *the conservation of energy principle*, provides a sound basis for studying the relationships among the various forms of energy and energy interactions.
- Based on experimental observations, the first law of thermodynamics states that *energy* can be neither created nor destroyed during a process; it can only change forms. Therefore, every bit of energy should be accounted for during a process.

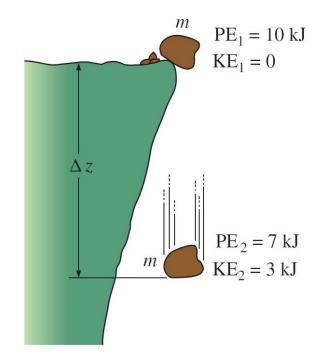
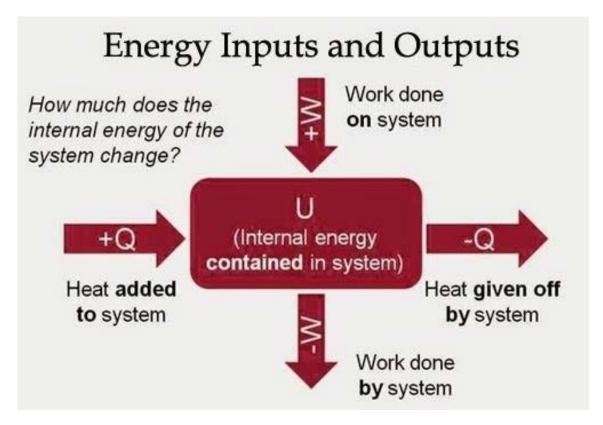


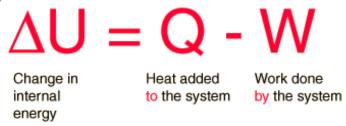
FIGURE 2-39

Energy cannot be created or destroyed; it can only change forms.



Source: https://me-mechanicalengineering.com/first-law-of-thermodynamics/

The change in internal energy of a system is equal to the heat added to the system minus the work done by the system.



Source: https://www.wikilectures.eu/w/1st law of thermodynamics

$$\begin{pmatrix}
\text{Total energy} \\
\text{entering the system}
\end{pmatrix} - \begin{pmatrix}
\text{Total energy} \\
\text{leaving the system}
\end{pmatrix} = \begin{pmatrix}
\text{Change in the total} \\
\text{energy of the system}
\end{pmatrix}$$

or

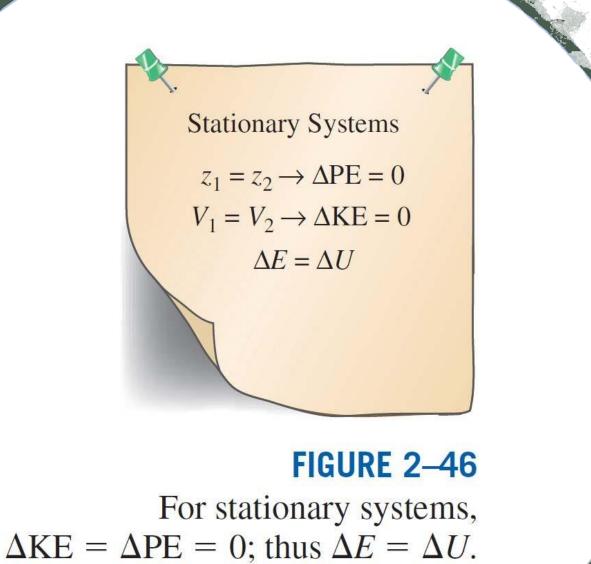
$$E_{\rm in} - E_{\rm out} = \Delta E_{\rm system}$$

$$\Delta E = \Delta U + \Delta KE + \Delta PE$$

$$\Delta U = m(u_2 - u_1)$$

$$\Delta KE = \frac{1}{2}m(V_2^2 - V_1^2)$$

$$\Delta PE = mg(z_2 - z_1)$$



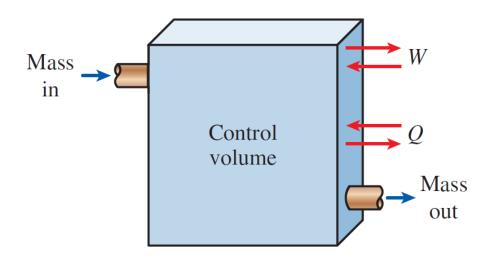


FIGURE 2-47

The energy content of a control volume can be changed by mass flow as well as heat and work interactions.

$$E_{\rm in} - E_{\rm out} = (Q_{\rm in} - Q_{\rm out}) + (W_{\rm in} - W_{\rm out}) + (E_{\rm mass,in} - E_{\rm mass,out}) = \Delta E_{\rm system}$$

Some special cases of the First Law of Thermodynamics

Adiabatic processes: system insulated, no heat transfer

$$Q = 0$$
 therefore $\Delta E_{int} = Q - W = -W$

Constant-volume process: V is fixed

$$dW = pdV = 0$$
, $W = 0$ therefore $\Delta E_{int} = Q - W = Q$

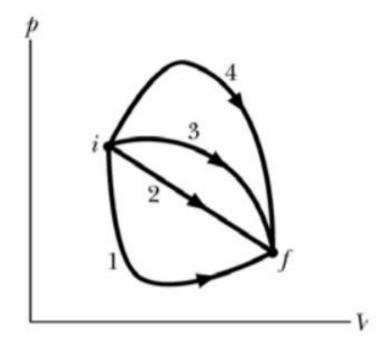
Cyclic processes: System goes back to the initial state

$$\Delta E_{int} = 0$$
 therefore $Q = W$

Numerical Interpretation

- The figure here shows four paths on a p-V diagram which a gas can be taken from state i to state f. Rank the paths according to the following parameters, greatest first.
 - A) the change Δ Eint
 - B) the greatest work W done by the gas
 - C) the magnitude of the energy transferred as heat Q.

https://www.doubtnut.com/question-answer-physics/the-figure-here-shows-paths-on-a-p-v-diagram-along-which-a-gas-can-be-taken-from-state-i-to-state-f--482963816

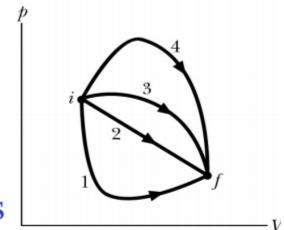


Numerical Interpretation

The figure here shows four paths on a p-V diagram which a gas can be taken from state i to state f. Rank the paths according to the following parameters, greatest first.

A) the change ΔE_{int}

All paths start at i and end at f, therefore all paths have the same change in internal energy, ΔE_{int}



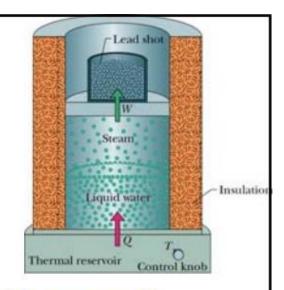
B) the greatest work W done by the gas

$$W = \int dW = \int_{V_i}^{V_f} p \, dV$$
 Path 4 has the maximum area under the p-V curve.

C) the magnitude of the energy transferred as heat Q.

$$Q = \Delta E_{int} + W$$
 W is greatest for path 4 => Q_4 is greatest.

Let 1.00 kg of liquid water at 100°C be converted to steam at 100°C by boiling at standard atmospheric pressure (1 atm) as shown. The volume of that water changes from an initial value of 1.00 x10⁻³m³ as a liquid to 1.671 m³ as steam.

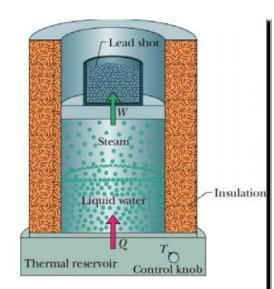


A) How much work is done by the system during this process?

B) How much energy is transferred as heat during the process?

C) What is the change in the system's internal energy during the process?

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A) How much work is done by the system during this process?

$$W = \int dW = \int_{V_i}^{V_f} p \, dV = (1.01 \times 10^5)(1.671 - 0.001)J = 169.0 \text{ kJ}$$

B) How much energy is transferred as heat during the process?

$$Q = m L_V = (1.00)(2256) = 2256 \text{ kJ}$$

C) What is the change in the system's internal energy during the process?

$$\Delta E_{int} = Q - W = 2256 \text{ kJ} - 169 \text{ kJ} = 2087 \text{ kJ}$$