

Thermodynamics I

Chapter 4

Lecture no.1

Reversible Processes (Non-Flow & Steady Flow)

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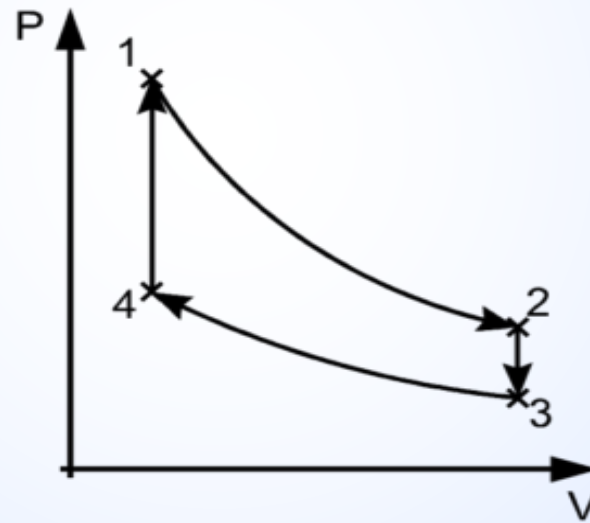
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Reversible Process

The reversible process is the ideal process which never occurs, while the irreversible process is the natural process that is commonly found in nature. When we tear a page from our notebooks, we cannot change this and 'un-tear'. This is an irreversible process. Whereas when water evaporates, it can also be condensed in the form of rains. This is a reversible process. Let us study more about them below.

A thermodynamic process is reversible if the process can return back in such a that both the system and the surroundings return to their original states, with no other change anywhere else in the universe. It means both system and surroundings are returned to their initial states at the end of the reverse process.



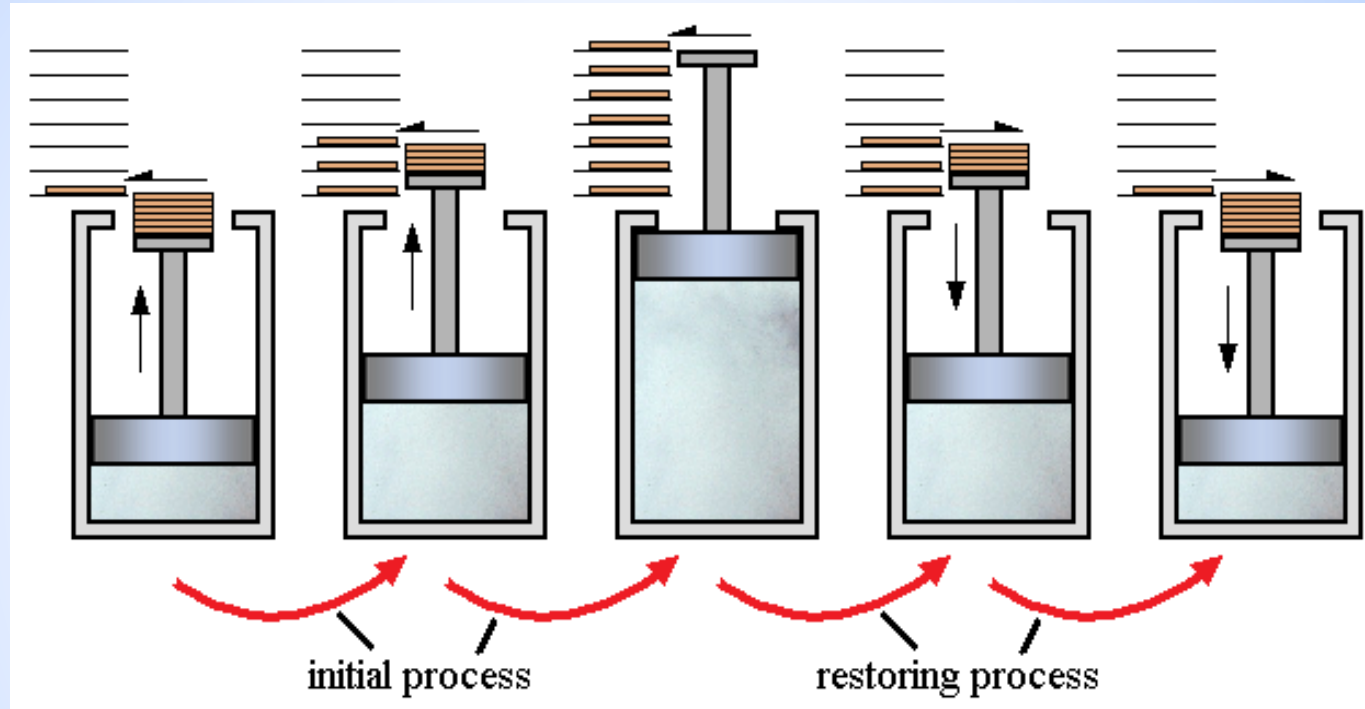
In the figure above, the system has undergone a change from state 1 to state 2. The reversible process can reverse completely and there is no trace left to show that the system had undergone thermodynamic change. During the reversible process, all the changes in state that occur in the system are in thermodynamic equilibrium with each other.



A process can be reversible only when its satisfying two conditions

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- Dissipative force must be absent.
- The process should occur in infinite small time.



In simple words, the process which can reverse back completely is a reversible process. This means that the final properties of the system can perfectly reverse back to the original properties. The process can be perfectly reversible only if the changes in the process are infinitesimally small. In practical situations it is not possible to trace these extremely small changes in extremely small time, hence the reversible process is also an ideal process. The changes that occur during the reversible process are in equilibrium with each other, friction effects can be assumed to be negligible and heat must never be transferred to or from the system through a finite temperature difference .



Non-Flow Reversible Processes :

(Heating/Cooling and Expansion/Compression of Gases)

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Until now we have studied that thermodynamic system of a gas is used to convert heat energy into work energy or vice versa which is required in a number of practical applications.

Thermodynamic system is also of many types. Here we will study a closed system which can exchange energy with surroundings but not the mass. Thus mass does not flow in or out of the system and so the processes of heating/cooling/compression/expansion etc are undergone by this fix mass of gas confined in continuous closed boundary are called Non-flow processes. In these processes some property of the gas may change and some may not change based on which a particular process is characterized. By applying first law of thermodynamics, the various forms of energy exchange can be calculated considering the processes as reversible i.e. taking all internal/external losses due to friction etc as nil.

For this let us consider a fix quantity of an ideal gas filled in a metallic cylinder as shown in below figure, one side of which is covered or fixed by a solid end plate and other side is covered by a moving piston. The outer wall of piston matches with inner wall of cylinder such that it makes a leak proof sliding joint. This moving piston makes one of the boundaries of system as moving or flexible.

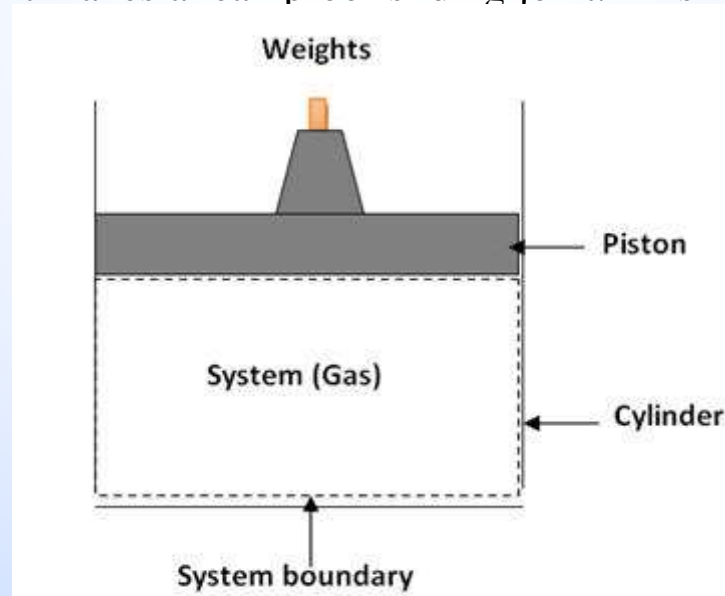


Fig. Closed system of gas filled in metallic cylinder



When infinitely small heat δQ is supplied to gas from outside through the wall of cylinder, the gas tends to expand and forces the piston weight F to move up. Let piston moves by a short distance, dl . Then the infinitesimal work done can be calculated as

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$$\begin{aligned}\delta W &= F \times dl \\ &= \frac{F}{A} \times dl \times A \\ &= P \cdot dV\end{aligned}$$

Work = Pressure x Change in volume

Total work during a non-flow process 1-2 i.e. compression or expansion of gases can be calculated as taking integral of $P \cdot dV$

$$W = \int_1^2 P \cdot dV$$

Thus, first law of thermodynamics can be written as :

$$\delta Q = P \cdot dV + dU$$

All these reversible processes of heat exchange (Heating or Cooling) and work exchange (compression or expansion) by a system of ideal gas with its surroundings can take place in various ways as discussed below:



1. Constant volume (Isochoric) heating/cooling process

Let the piston is fixed at one point in the walls of cylinder such that volume $V = \text{Constant}$

So the ideal gas law $PV/T = \text{Constant}$ will reduce to

$$\frac{P}{T} = \text{Constant}$$

Now let a small increment of heat δQ is supplied to the gas. According to the 1st law of thermodynamics,

$$\delta Q = \delta W + dU$$

$$\delta Q = P \cdot dV + dU$$

As $V = \text{Constant}$, $\delta W = P \cdot dV = 0$, then the equation was written

For any work substance (fluid)

$$Q = \Delta U$$

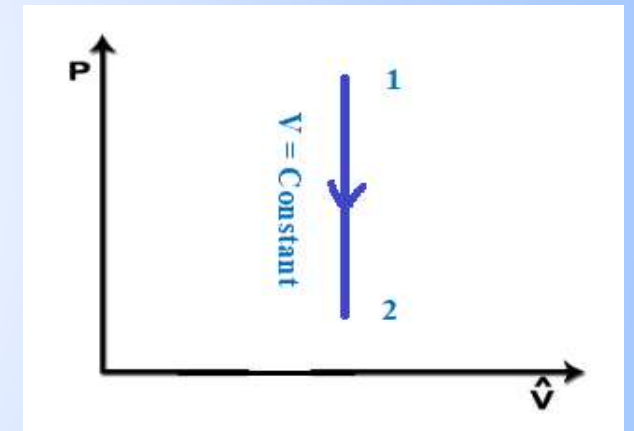
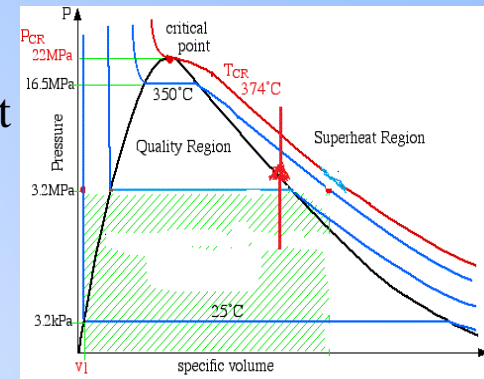
And for perfect gas, the equation can be written as :

$$\text{So } \delta Q = \Delta U = mC_v \Delta T$$

Thus knowing the values of δQ and C_v , ΔT (Increase in temperature) can be calculated and so ΔP can also be calculated from eq.

$$\frac{P}{T} = \text{Constant}$$

By measuring the small increment in temperature and pressure of system, the quantity δQ & ΔU can also be calculated.



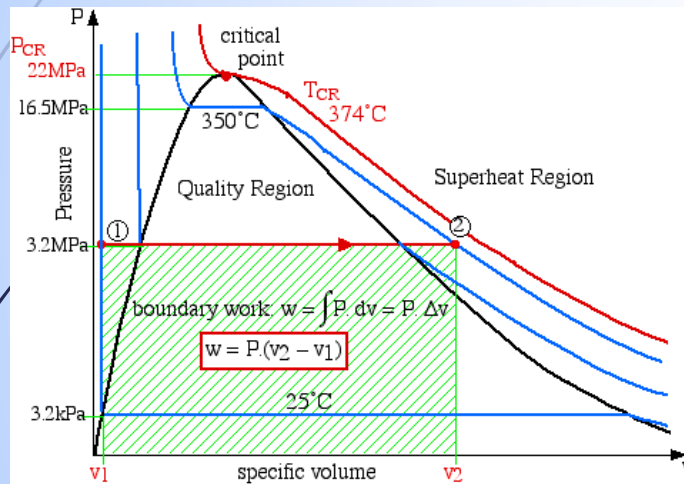
2. Constant pressure (Isobaric) process

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Let the piston is free to move in the cylinder and force, F or pressure, P on the piston remains constant. So now the boundaries of the system can move and so the gas can expand or can be compressed i.e. the system is able to exchange work.

Now let a small quantity of heat, δQ is supplied to the gas through the walls of cylinder.

On absorbing this heat, the temperature of gas will tend to increase and simultaneously the gas will tend to expand against force F on the piston. Applying 1st law of thermodynamic to this process.



$$\delta Q = \delta W + dU$$

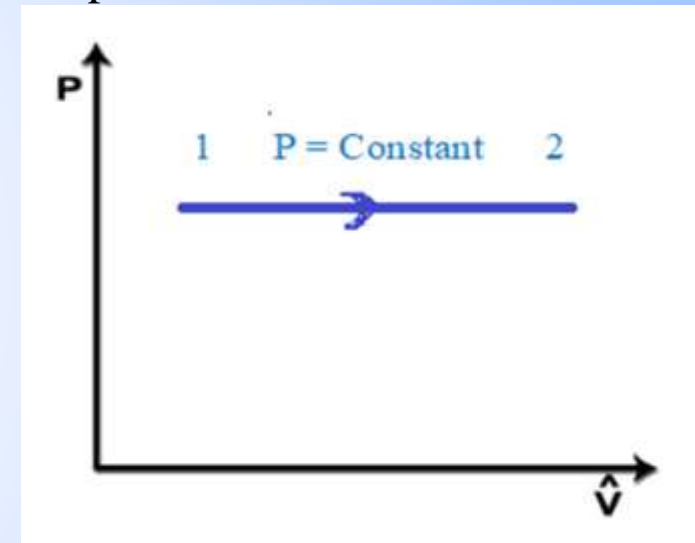
$$\text{Or } \delta Q = P.dV + m C_v dT$$

$$Q = P (V_2 - V_1) + (U_2 - U_1)$$

$$Q = (U_2 + P V_2) - (U_1 + P V_1)$$

For any working substance (fluid)

$$Q = H_2 - H_1$$



And for perfect gas, the equation can be written as :

$$Q = C_P T_2 - C_P T_1$$

$$Q = C_P (T_2 - T_1)$$

$$Q = C_P \Delta T = C_P dT$$

$$Q = P(V_2 - V_1) + mC_v(T_2 - T_1)$$



Also, work can be calculated for a perfect gas by equation :

$$W = P \int dV$$

$$W = P (V_2 - V_1) = PV_2 - PV_1 = m RT_2 - mRT_1 = mR (T_2 - T_1) = m \int R dT$$

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Thus the increase in volume of gas due to expansion and increase in temperature are interrelated and if one can be measured the other can be found and the quantities W , ΔU and so Q can be calculated from equation as By ideal gas law

As pressure, $P = \text{Constant}$, the gas law reduce to $V/T = \text{Constant}$

$$\frac{PV}{T} = \text{Constant} = mR$$

And the first law of thermodynamics

$$\delta Q = \delta W + dU$$

$$\delta Q = P.dV + m C_v dT$$

$$\text{or } mC_p dT = mRdT + mC_v dT \quad \left[\because \frac{P.dV}{dT} = mR \right]$$

$$\text{or } C_p = R + C_v$$

$$\text{or } C_p - C_v = R$$

Thus characteristic gas constant of an ideal gas is the difference between specific heat at constant pressure (C_p) and at constant volume (C_v).

The ratio of specific heats C_p / C_v is denoted as γ .

The physical meaning of R or γ can be taken as the characteristic of an ideal gas to expand under the influence of heat or we can say the increase in product of pressure and volume PV with increase in T .



3 . Constant temperature (Isothermal) Process

a. Steam

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It can be seen, that the isothermal process for a wet steam lies at the constant pressure line also. Thus the heat calculated by equation :

$$Q = H_2 - H_1$$

or

$$Q = U_2 - U_1 + W$$

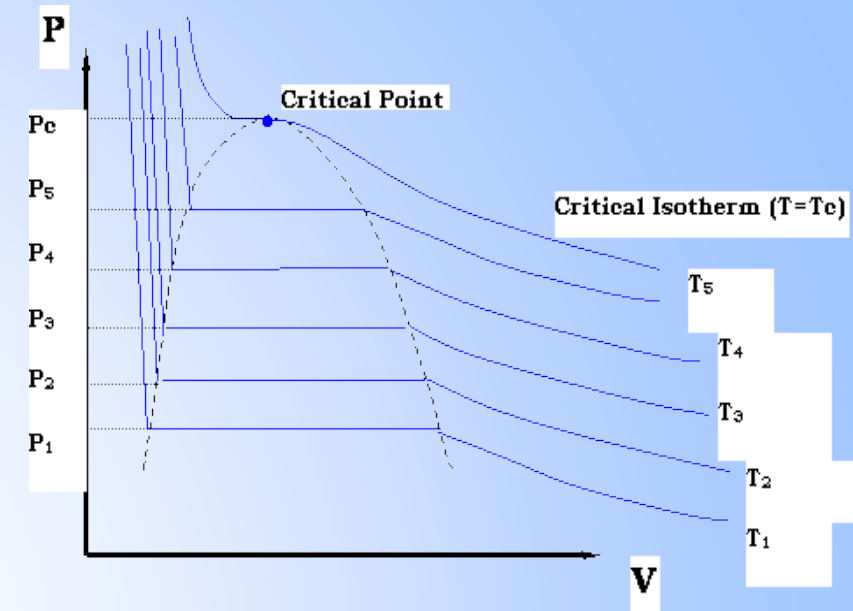
b. Ideal gas (perfect gas)

The beauty of isothermal expansion or compression process (**when the working substance is ideal gas**) is that the internal energy remains constant and so whole of the heat exchange by system is converted to work and vice versa. However the process is difficult to visualize while thinking that when the system (gas filled behind piston in the cylinder) absorbs heat, its temperature should always increase. But what happens in the isothermal process that while absorbing heat, simultaneously the gas expands thus decrease in its pressure and temperature takes place and the net change in temperature is zero. While expanding, the gas gives positive work equal to the heat supplied. Conversely also if the gas is compressed by doing extra work on the gas, its pressure and temperature tends to increase, but simultaneously if the gas is cooled in such proportion that the net change in temperature remains zero, the compression process becomes isothermal. In this isothermal compression process, again the work supplied to the gas is given away by the system in the form of heat. Applying first law of thermodynamics to this isothermal process in which

$$du = 0, \text{ because } dt = 0$$

$$\delta Q = \delta W = P \cdot dV$$

$$Q = W = \int_1^2 P \cdot dV$$



From ideal gas law

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At $T = \text{constant}$

Gas law reduces to $PV = \text{constant}$ or

Putting this in equation

$$\frac{PV}{T} = \text{constant}$$

$$P_1 V_1 = P_2 V_2 = PV$$

$$Q = W = \int_1^2 \frac{P_1 V_1}{v} \cdot dV = P_1 V_1 \int_1^2 \frac{dV}{V}$$

$$Q = W = P_1 V_1 \ln \frac{v_2}{v_1}$$

By using this equation we can calculate the work exchange or heat exchange during isothermal process if we know the change in volume.

Also in case of Isothermal process we know that

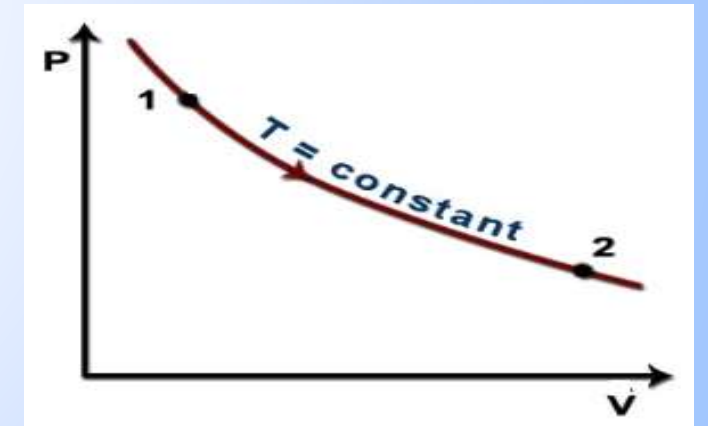
$$P_1 V_1 = P_2 V_2$$

$$\frac{V_2}{V_1} = \frac{P_1}{P_2}$$

Putting in eq.

$$Q = W = P_1 V_1 \ln \frac{P_1}{P_2}$$

By using this equation, we can calculate the work exchange or heat exchange if we know the change in pressure.



4. Adiabatic Process

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- The system does not exchange heat with the surrounding
- System is completely insulated from surrounding

From non – flow equation (first law of thermodynamics)

$$Q = U_2 - U_1 + W$$

$$Q = 0$$

Thus, for any adiabatic process whether it is reversible or not, and for a vapour undergoing a reversible adiabatic process the work done can be found by :

$$U_1 - U_2 = W$$

And for an ideal gas, where $Q = 0$, The non flow equation is :

$$dq = du + dw = 0$$

$$dq = du + p dv = 0$$

$$du + p dv = 0$$

$$P V = R T$$

For unit mass



$$P = \frac{RT}{V}$$

$$du + \frac{RT}{V} \frac{dV}{V} = 0$$

$$C_v dT + \frac{RT}{V} \frac{dV}{V} = 0$$

$$C_v \frac{dT}{T} + R \frac{dV}{V} = 0$$

$$C_v \ln T + R \ln V = \text{Constant}$$

Where $T = \frac{PV}{R}$

[$C_v \ln \frac{PV}{R} + R \ln V = \text{Constant}$] divided by C_v

$$\ln \frac{PV}{R} + \frac{R}{C_v} \ln V = \text{Constant}$$

Where $C_v = \frac{R}{\gamma - 1}$, $\gamma - 1 = \frac{R}{C_v}$

$$\ln \frac{PV}{R} + (\gamma - 1) \ln V = \text{Constant}$$

$$\ln \frac{PV}{R} + \ln V^{(\gamma-1)} = \text{Constant}$$

$$\ln \frac{PV^\gamma}{R} = \text{Constant} \quad \text{Then} \quad \frac{PV^\gamma}{R} = e^{\text{constant}}$$

$$PV^\gamma = \text{Constant}$$



Thus perfect gas expand / compress according to the equation :

$$PV^\gamma = \text{Constant}$$

$$P = \frac{C}{V^\gamma}$$

Then the work found by :

$$W = U_1 - U_2 = \int_1^2 P dV$$

$$W = C_v dT = C_v (T_2 - T_1)$$

or
$$W = \int_1^2 \frac{C}{V^\gamma} dV = C \int_1^2 \frac{dV}{V^\gamma}$$

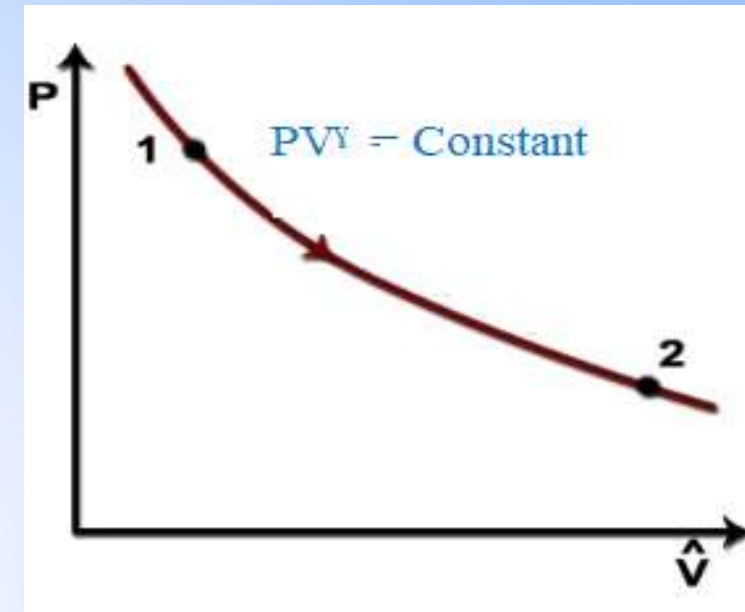
$$W = C \left[\frac{V^{1-\gamma}}{1-\gamma} \right]_1^2 = PV^\gamma \left[\frac{V^{1-\gamma}}{1-\gamma} \right]_1^2 = \frac{P_2V_2 - P_1V_1}{1-\gamma}$$

$$W = \frac{P_1V_1 - P_2V_2}{\gamma - 1}$$

By equations $PV^\gamma = \text{Constant}$, and $\frac{PV}{T} = \text{Constant}$, can found

$$T_2 = T_1 \left(\frac{V_1}{V_2} \right)^{\gamma-1}$$

$$T_2 = T_1 \left(\frac{P_2}{P_1} \right)^{(x-1)/\gamma}$$



5. Polytropic Process

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It is found that many processes in practice approximate to reversible law of the form $PV^n = \text{Constant}$, where n is a constant. The index n depends only on the heat and work quantities during the process. Both vapours and perfect gases obey this type of closely in many non-flow processes. Such processes are internally reversible. Thus perfect gas expand / compress according to the law :

$$PV^n = \text{Constant}, \quad \text{then } P = \frac{C}{V^n}$$

Then the work found by :

$$W = \int_1^2 P dV$$

$$W = \int_1^2 \frac{C}{V^n} dV = C \int_1^2 \frac{dV}{V^n}$$

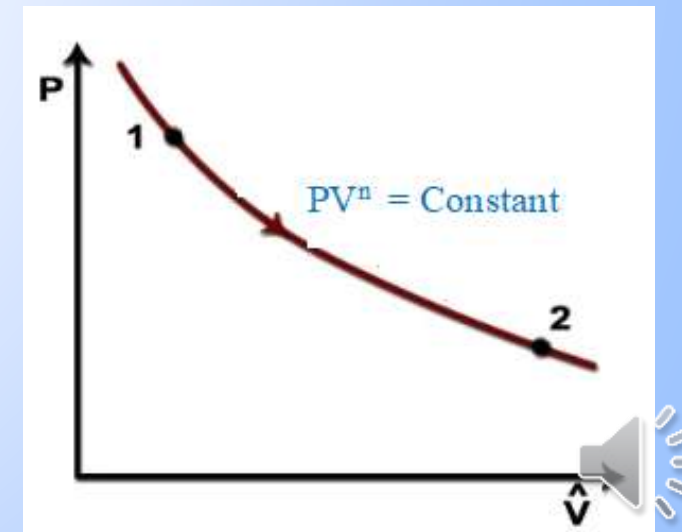
$$W = C \left[\frac{V^{1-n}}{1-n} \right]_1^2 = PV^n \left[\frac{V^{1-n}}{1-n} \right]_1^2 = \frac{P_2 V_2 - P_1 V_1}{1-n}$$

$$W = \frac{P_1 V_1 - P_2 V_2}{n-1}$$

By equations $PV^n = \text{Constant}$, and $\frac{PV}{T} = \text{Constant}$, can found

$$T_2 = T_1 \left(\frac{V_1}{V_2} \right)^{n-1}$$

$$T_2 = T_1 \left(\frac{P_2}{P_1} \right)^{(n-1)/n}$$



For unit mass , the work is

$$W = \frac{P_1 V_1 - P_2 V_2}{n-1}$$

$$w = \frac{R (T_1 - T_2)}{n-1}$$

While the total work is

$$W = \frac{m R (T_1 - T_2)}{n-1}$$

$$Q = U_2 - U_1 + W = C_v (T_1 - T_2) + \frac{R (T_1 - T_2)}{n-1}$$

Where $C_v = \frac{R}{\gamma - 1}$,

$$Q = \frac{R (T_1 - T_2)}{n-1} - \frac{R}{\gamma - 1} (T_1 - T_2)$$

$$Q = R (T_1 - T_2) \left[\frac{1}{n-1} - \frac{1}{\gamma - 1} \right]$$

$$Q = \left[\frac{\gamma - n}{\gamma - 1} \right] \frac{R (T_1 - T_2)}{n-1} = \left[\frac{\gamma - n}{\gamma - 1} \right] W$$



There are 2-types of specific heats:

1. Specific Heat at Constant Volume C_v :

It is the amount of heat required to rise the temperature of unit mass of a gas through one degree when the volume is constant. If a unit mass of a gas is taken in a closed vessel and is heated, the volume of the gas remains constant, but the temperature increases. As the volume remains constant, there is no external work done by the gas and as temperature of the gas increases, there is increase in internal energy of the gas.

Therefore heat supplied to the gas is completely utilized in increasing the I.E. of the gas,

2. Specific Heat at Constant Pressure C_p :

It is the amount of heat required to rise the temperature of unit mass of a gas through one degree when the pressure is kept constant.

Consider a unit mass of a gas in a cylinder fitted with a frictionless piston. When the gas is heated, the piston moves up, maintaining the same pressure. But the volume and temperature of the gas increases during heating.

As there is increase in volume, there is external work done by the gas and there is increase in temperature, there is increase in I.E.

Thus heat supplied when the pressure is constant, is utilized for two purposes:

- (a) To do some external work.
- (b) To increase the I.E of the gas.

Whereas in case of constant volume heating, the heat supplied is completely utilized for increasing the I.E.

∴ Specific heat at constant pressure is greater than specific heat at constant volume.

i.e., $C_p > C_v$ and, $C_p = \left[\frac{\delta q}{dT} \right]_p$

Also $C_p = \left[\frac{dh}{dT} \right]_p$

Therefore, C_p is also defined as, the rate of change of specific enthalpy with respect to temperature when the pressure is kept constant.



6. Adiabatic Index:

It is the ratio of specific heat at constant pressure to the specific heat at constant volume and is given by,

$$K \text{ or } \gamma = \frac{C_p}{C_v}$$

Note: For air, $C_p = 1.005 \frac{\text{kJ}}{\text{kg-K}}$ and $C_v = 0.718 \frac{\text{kJ}}{\text{kg-K}}$ and $\gamma = 1.4$

Internal Energy of an Ideal Gas

We will show that the internal energy of an ideal gas is a function of temperature only. This makes physical sense because there is an assumption in ideal gas behavior that there is no interaction between the molecules when we write

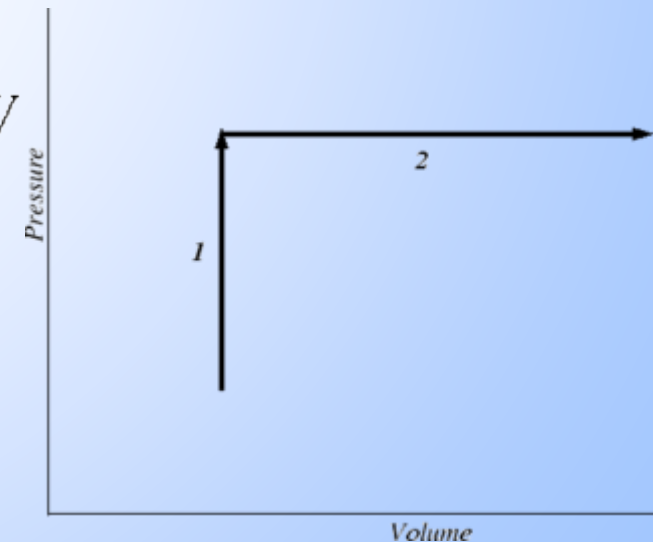
Start with a reversible process for an ideal gas:

$$P\bar{V} = RT$$

$$dU = dq + dw = dq - PdV$$

Consider two processes: one occurring at constant volume, the other occurring at constant pressure.

Figure : Two consecutive processes, constant volume followed by constant pressure.



For process 1: At constant V ($PdV = 0$)

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$$dU = C_V dT + 0$$

i.e. U is $f(T)$ only

$$dU = C_V dT$$

This can be integrated because T is the only thing that is changing on the right hand side (C_V is assumed to be independent of T and V).

For process 2:

$$dU = C_P dT - PdV$$

P is constant (i.e., not a function of T or V) so it can be integrated directly. Using the ideal gas law:

$$PV = nRT$$

For one mole of gas (i.e. $n = 1$)

$$PdV + VdP = RdT$$

At constant P ($VdP = 0$)

$$PdV = RdT$$



Then

$$dU = C_p dT - R dT$$

$$dU = (C_p - R) dT$$

So $C_v = C_p - R$

Then $dU = C_v dT$ or $dU = \frac{C_p}{\gamma} dT$, U is $f(T)$ only.

Note: that $\gamma = \frac{C_p}{C_v}$, $C_v = \frac{R}{\gamma - 1}$ and $C_p = \frac{\gamma R}{\gamma - 1}$

Summary

A perfect gases obey this type of closely in many non-flow processes. Such processes are internally reversible according to the law: $PV^{\text{index}} = \text{Constant}$. And $\frac{PV}{T} = \text{Constant}$.

1. If index = ∞ , the process is to be called Isochoric process i.e $V = \text{Constant}$ (P)^(1/ ∞) = Constant. If the gas is an ideal gas. Then, $P/T = \text{Constant}$
2. If index = 1, the process is to be called Isothermal process i.e $T = \text{Constant}$. Then, $PV = \text{Constant}$
3. If index = 0, the process is to be called Isobaric process i.e $P = \text{Constant}$. Then $\frac{V}{T} = \text{Constant}$.
4. If index = γ , the process is to be called adiabatic process. Then, $PV^\gamma = \text{Constant}$
5. If index = n , the process is to be called polytropic process. Then, $PV^n = \text{Constant}$

