

## Units

①

### Dimension

### English Engineering

### S.I. units

Length

foot (ft)

meter (m)

Force

lbf

N

Mass

(slug) / pound mass (lbm)

kg

Temp.

°F

Absolute temp.

°R (Rankine)

°C

Density ( $\rho = \frac{1}{v}$ )

v = specific vol.

lbm/ft<sup>3</sup> or slug/ft<sup>3</sup>

kg/m<sup>3</sup>

Specific weight ( $\gamma = \rho g$ )

lbf/ft<sup>3</sup>

N/m<sup>3</sup>

Pressure

lbf/ft<sup>2</sup>

N/m<sup>2</sup>

R = (Individual gas const.)

ft-lbf/lbm-°R

N.m/kg.K

$$[\text{for perfect gas}] P = \rho R T \Rightarrow R = \frac{P}{\rho T}$$

$$\text{Again, } R = \frac{P}{\rho T} = \frac{\text{lbf}/\text{ft}^2}{\frac{\text{lbm}}{\text{ft}^3} \cdot \text{°R}} = \frac{\text{N}/\text{m}^2}{\frac{\text{kg}}{\text{m}^3} \times \text{°K}} = \frac{\text{N} \cdot \text{m}}{\text{kg} \cdot \text{K}}$$

$$= \frac{\text{lbf} \cdot \text{ft}}{\text{lbm} \cdot \text{°R}} \quad R_{\text{air}} = 287 \text{ N.m/(kg.K)}$$

$$= 53.3 \text{ lbf.ft/lbm.°R}$$

### Relation:

$$°R = \cancel{°F} + 459.67 \approx \underline{°F + 460.}$$

$$K = °C + 273$$

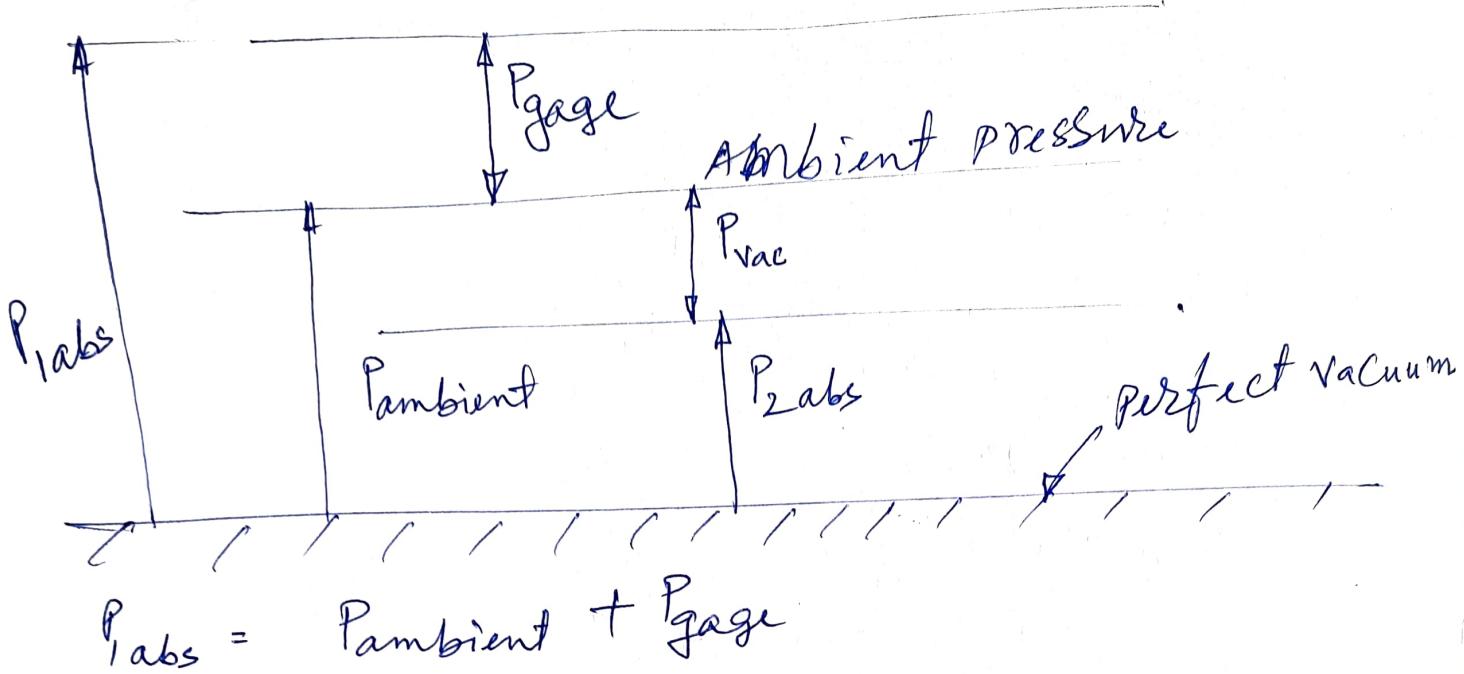
$$\frac{C}{5} = \frac{F-32}{9}$$

$R_u$  = Universal gas constant

$$R = \frac{R_u}{M_m}, M_m = \text{Molecular mass of gas.}$$

$$R_u = 8314 \text{ N.m/kg mole.K} = 1544 \cancel{\text{lbf.ft/(lbmole.°R)}}$$

Pressure:



$$P_{\text{2abs}} = P_{\text{ambient}} - P_{\text{vac}}$$

Units of pressure:  $\text{IN/m}^2$ ,  $\text{lbf/ft}^2$ ,  $\text{MPa}$  ( $= 10^6 \text{ N/m}^2$ )

$$1 \text{ bar} = 1 \times 10^5 \text{ N/m}^2, 1 \text{ atm} = 14.69 \text{ psi}$$

$$1 \text{ atm} = 0.1013 \text{ MPa} = 1.013 \times 10^5 \text{ N/m}^2$$

$$1 \text{ atm} = 1.013 \text{ bar}$$

Viscosity:-  $\eta = \frac{\text{Shear Stress}}{\text{Rate of angular deformation}}$

$$\eta_{\text{abs}} (\mu) \quad \text{Kinematic. } \left( \frac{\mu}{\rho} \right) \Rightarrow \text{lbf/ft}^2, \text{ N/m}^2.$$

$$\text{lbf}\cdot\text{sec}/\text{ft}^2 \text{ or. } \text{N.S}/\text{m}^2$$

To use this expansion the function must be continuous and possess continuous derivatives throughout the interval  $x_1$  to  $x_2$ . It should be noted that all derivatives in the expression above must be evaluated about the point of expansion  $x_1$ .

If the increment  $\Delta x = x_2 - x_1$  is small, only a few terms need be evaluated to obtain an accurate answer for  $f(x_2)$ . If  $\Delta x$  is allowed to approach zero, all higher-order terms may be dropped and

$$f(x_2) \approx f(x_1) + \left( \frac{df}{dx} \right)_{x=x_1} dx \quad \text{for } dx \rightarrow 0 \quad (1.26)$$

## 1.4 THERMODYNAMIC CONCEPTS FOR CONTROL MASS ANALYSIS

We apologize for the length of this section, but a good understanding of thermodynamic principles is essential to a study of gas dynamics.

### General Definitions

*Microscopic approach:* deals with individual molecules, and with their motion and behavior, on a statistical basis. It depends on our understanding of the structure and behavior of matter at the atomic level. Thus this view is being refined continually.

*Macroscopic approach:* deals directly with the average behavior of molecules through observable and measurable properties (temperature, pressure, etc.). This classical approach involves no assumptions regarding the molecular structure of matter; thus no modifications of the basic laws are necessary. The macroscopic approach is used in this book through the first 10 chapters.

*Control mass:* a fixed quantity of mass that is being analyzed. It is separated from its surroundings by a boundary. A control mass is also referred to as a *closed system*. Although no matter crosses the boundary, energy may enter or leave the system.

*Control volume:* a region of space that is being analyzed. The boundary separating it from its surroundings is called the *control surface*. Matter as well as energy may cross the control surface, and thus a control volume is also referred to as an *open system*. Analysis of a control volume is introduced in Chapters 2 and 3.

*Properties:* characteristics that describe the state of a system; any quantity that has a definite value for each definite state of a system (e.g., pressure, temperature, color, entropy).

State  
of  
process

*Intensive property:* depends only on the state of a system and is independent of its mass (e.g., temperature, pressure).

*Extensive property:* depends on the mass of a system (e.g., internal energy, volume).

#### Types of properties:

1. *Observable:* readily measured  
(pressure, temperature, velocity, mass, etc.)

*System:* - System is defined as an identifiable quantity  
 → Closed system/control mass ⇒ Mass cannot cross boundary of matter.  
 → Open system/control volume ⇒ Both mass & Energy can cross the boundary  
 → Isolated system: - Neither mass nor energy can cross the boundary

} Fluid  
as continuum.

2. *Mathematical*: defined from combinations of other properties (density, specific heats, enthalpy, etc.)
3. *Derived*: arrived at as the result of analysis
  - a. *Internal energy* (from the first law of thermodynamics)
  - b. *Entropy* (from the second law of thermodynamics)

*State change*: comes about as the result of a change in any property.

*Path or process*: represents a series of consecutive states that define a unique path from one state to another. Some special processes:

$$\begin{aligned} pV = \text{const} &= \text{Adiabatic} \rightarrow \text{no heat transfer} & \gamma = \frac{C_p}{C_v} = 1.4 \text{ for air} \\ pV = RT &= \text{Isothermal} \rightarrow T = \text{constant} \\ \leftarrow & \text{Isobaric} \rightarrow p = \text{constant} \rightarrow \text{Isobaric} \\ & \text{Isentropic} \rightarrow s = \text{constant} \\ & \text{Polytropic} \rightarrow pV^n = \text{const}. \end{aligned}$$

*Cycle*: a sequence of processes in which the system is returned to the original state.

*Point functions*: another way of saying *properties*, since they depend only on the state of the system and are independent of the history or process by which the state was obtained.

*Path functions*: quantities that are *not* functions of the state of the system but rather depend on the path taken to move from one state to another. *Heat* and *work* are path functions. They can be observed crossing the system's boundaries *during* a process.

## Laws of Classical Thermodynamics

- 0<sup>2</sup> Relation among properties
- 0 Thermal equilibrium
- 1 Conservation of energy
- 2 Degradation of energy (irreversibilities)

- Equation of state

The 0<sup>2</sup> law (sometimes called the 00 law) is seldom listed as a formal law of thermodynamics; however, one should realize that without such a statement our entire thermodynamic structure would collapse. This law states that we may assume the existence of a relation among the properties, that is, an equation of state. Such an equation might be extremely complicated or even undefined, but as long as we know that such a relation exists, we can continue our studies. The equation of state can also be given in the form of tabular or graphical information.

For a single component or pure substance only three *independent* properties are required to fix the state of the system. Care must be taken in the selection of these properties; for example, temperature and pressure are not independent if the substance exists in more than one phase (as in a liquid together with its vapor). When dealing with a unit mass, only two independent properties are required to fix the state. Thus

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one can express any property in terms of two other known independent properties with a relation such as

1)

$$P = f(x, y)$$

$$\cancel{PV=RT} \quad PXT = mRT$$

$$\cancel{PV=RT} \quad PXT = nRuT$$

- 2) If two systems are separated by a nonadiabatic wall (one that permits heat transfer), the state of each system will change until a new equilibrium state is reached for the combined system. The two systems are then said to be in *thermal equilibrium* with each other and will then have one property in common which we call the *temperature*.

The zeroth law states that two systems in thermal equilibrium with a third system are in thermal equilibrium with each other (and thus have the same temperature). Among other things, this allows the use of thermometers and their standardization.

3)

### First Law of Thermodynamics

The first law deals with conservation of energy, and it can be expressed in many equivalent ways. Heat and work are two extreme types of energy in transit. Heat is transferred from one system to another when an effect occurs solely as a result of a temperature difference between the two systems.

Heat is always transferred from the system at the higher temperature to the one at the lower temperature.

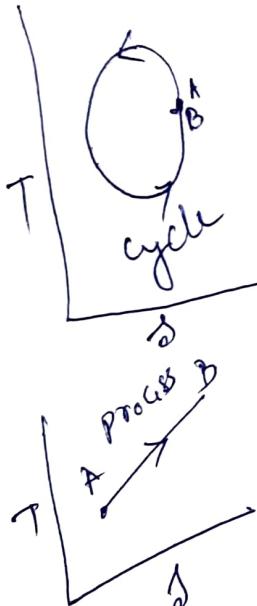
Work is transferred from a system if the total external effect can be reduced to the raising of a mass in a gravity field. For a closed system that executes a complete cycle,

$$\sum Q = \sum W \quad (1.27)$$

where

$Q$  = heat transferred *into* the system = +ve, heat going out system = -ve

$W$  = work transferred *from* the system : WORK done by the system is +ve.  
" " " on the " " " -ve.



Other sign conventions are sometimes used but we shall adopt those above for this book.

For a closed system that executes a process,

$$Q = W + \Delta E \quad \text{for total} \quad (1.28)$$

where  $E$  represents the total energy of the system. On a unit mass basis, equation (1.28) is written as

$$q = w + \Delta e \quad \text{for unit mass} \quad (1.29)$$

The total energy may be broken down into (at least) three types:

$$e \equiv u + \frac{V^2}{2g_c} + \frac{g}{g_c} z \quad (1.30)$$

**System!** System is defined as an identifiable quantity of matter.

→ closed system : control mass  $\Rightarrow$  mass constant

→ open system : control volume  $\Rightarrow$  volume constant.

Isolated System :

where

$u$  = the intrinsic internal energy manifested by the motion of the molecules within the system

$\frac{v^2}{2g_e}$  = For Thermally perfect gas,  $u = u(T)$  only.  
 $\frac{v^2}{2g_e}$  = the kinetic energy represented by the movement of the system as a whole

$\frac{g_z}{g_e}$  = the potential energy caused by the position of the system in a field of gravity

It is sometimes necessary to include other types of energy (such as dissociation energy), but those mentioned above are the only ones that we are concerned with in this book.

For an infinitesimal process, one could write equation (1.29) as

$$\delta q = \delta w + de \quad (1.31)$$

$$de = d(u + \frac{v^2}{2} + gz)$$

Note that since heat and work are *path functions* (i.e., they are a function of how the system gets from one state point to another), infinitesimal amounts of these quantities are not exact differentials and thus are written as  $\delta q$  and  $\delta w$ . The infinitesimal change in internal energy is an exact differential since the internal energy is a point function or property. For a stationary system, equation (1.31) becomes

$$\delta q = \delta w + du \quad (1.32)$$

The reversible work done by pressure forces during a change of volume for a stationary system is

$$\delta w = p dv \quad (1.33)$$

Combination of the terms  $u$  and  $pv$  enters into many equations (particularly for open systems) and it is convenient to define the property *enthalpy*:

$$h \equiv u + pv \quad (1.34)$$

Enthalpy is a property since it is defined in terms of other properties. It is frequently used in differential form:

$$dh = du + d(pv) = du + p dv + v dp \quad (1.35)$$

Other examples of defined properties are the specific heats at constant pressure ( $c_p$ ) and constant volume ( $c_v$ ):

$$h = u + Pv$$

$$dh = du + Pdv + vdp$$

$$dh - vdp = du + pdv$$

$$\delta q = \delta w + du = pdv + du \quad \text{--- (1) Reversible}$$

$$\delta q = dh - vdp \quad \text{--- (2) Stationary system}$$

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8  $h = h(T, P)$

$$dh = \left(\frac{\partial h}{\partial T}\right)_P dT + \left(\frac{\partial h}{\partial P}\right)_T dp$$

$$\delta q = dh = C_p dT \quad \text{Const. Press. from equ^u(2).}$$

$$\delta q = pdv + du$$

(7)

$$u = u(T, v)$$

$$du = \left(\frac{\partial u}{\partial T}\right)_v dT + \left(\frac{\partial u}{\partial v}\right)_T dv$$

$$\delta q = du = C_v dT$$

(1.36)

$$C_p \equiv \left(\frac{\partial h}{\partial T}\right)_P$$

$$C_v \equiv \left(\frac{\partial u}{\partial T}\right)_v \quad (1.37)$$

$$= \cancel{C_p}$$

$$= C_v dT$$

Const. Volume from equ^u(1).

### Second Law of Thermodynamics

The *second law* has been expressed in many equivalent forms. Perhaps the most classic is the statement by Kelvin and Planck stating that it is impossible for an engine operating in a *cycle* to produce *net* work output when exchanging heat with only one temperature source. Although by itself this may not appear to be a profound statement, it leads the way to several corollaries and eventually to the establishment of a most important property (entropy).

The *second law* also recognizes the degradation of energy quality by irreversible effects such as internal fluid friction, heat transfer through a finite temperature difference, lack of pressure equilibrium between a system and its surroundings, and so on. All real processes have some degree of irreversibility present. In some cases these effects are very small and we can envision an ideal limiting condition that has none of these effects and thus is reversible. A *reversible process* is one in which *both* the system and its surroundings can be restored to their original states.

By prudent application of the second law it can be shown that the integral of  $\delta Q/T$  for a reversible process is independent of the path. Thus this integral must represent the change of a *property*, which is called *entropy*:

$$\Delta S \equiv \int \frac{\delta Q_R}{T} \quad (1.38)$$

where the subscript *R* indicates that it must be applied to a reversible process. An alternative expression on a unit mass basis for a differential process is

$$ds \equiv \frac{\delta q_R}{T} \quad (1.39)$$

Although you have no doubt used entropy for many calculations, plots, and so on, you probably do not have a good feeling for this property. In Chapter 3 we divide entropy changes into two parts, and by using it in this fashion for the remainder of this book we hope that you will gain a better understanding of this elusive "creature."

### Property Relations

Some extremely important relations come from combinations of the first and second laws. Consider the first law for a stationary system that executes an infinitesimal process:

$$\delta q = \delta w + du \quad (1.32)$$

*Main = 28.8 kg/kg-mole.*  
 $\delta = 1.4$

(8)

If it is a reversible process,

$$\delta w = p dv \quad (1.33) \quad \text{and} \quad \delta q = T ds \quad (\text{from 1.39})$$

Substitution of these relations into the first law yields

$$T ds = du + p dv \quad (1.40)$$

Differentiating the enthalpy, we obtained

$$dh = du + p dv + v dp \quad (1.35)$$

Combining equations (1.35) and (1.40) produces

$$T ds = dh - v dp \quad (1.41)$$

Although the assumption of a reversible process was made to derive equations (1.40) and (1.41), the results are equations that *contain only properties and thus are valid relations to use between any end states*, whether reached reversibly or not. These are important equations that are used throughout the book.

$$\underline{T ds = du + p dv} \quad (1.40)$$

$$\underline{T ds = dh - v dp} \quad (1.41)$$

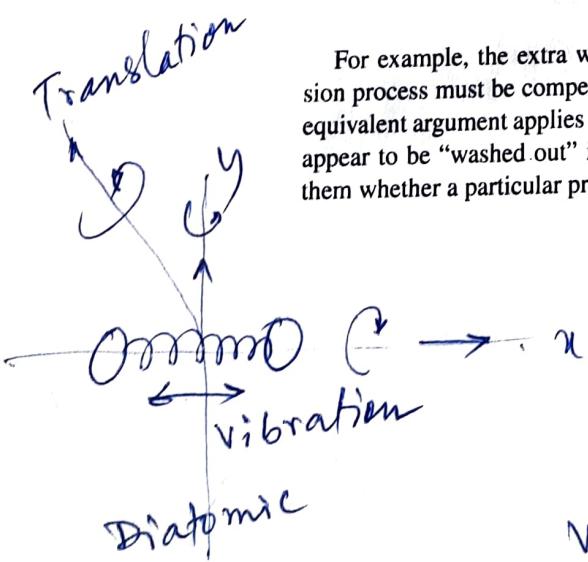
If you are uncomfortable with the foregoing technique (one of making special assumptions to derive a relation which is then generalized to be always valid since it involves only properties), perhaps the following comments might be helpful. First let's write the first law in an alternative form (as some authors do):

$$\delta q - \delta w = du \quad (1.32a)$$

Since the internal energy is a property, changes in  $u$  depend only on the end states of a process. Let's now substitute an irreversible process *between the same end points* as our reversible process. Then  $du$  must remain the same for both the reversible and irreversible cases, with the following result:

$$(\delta q - \delta w)_{\text{rev}} = du = (\delta q - \delta w)_{\text{irrev}}$$

For example, the extra work that would be involved in an irreversible compression process must be compensated by *exactly the same amount of heat released* (an equivalent argument applies to an expansion). In this fashion, irreversible effects will appear to be "washed out" in equations (1.40) and (1.41) and we cannot tell from them whether a particular process is reversible or irreversible.



3 - Translation mode

2 - rotational mode

$$\text{about } x\text{-axis is rotation energy} = \frac{1}{2} I \omega^2 \rightarrow 0$$

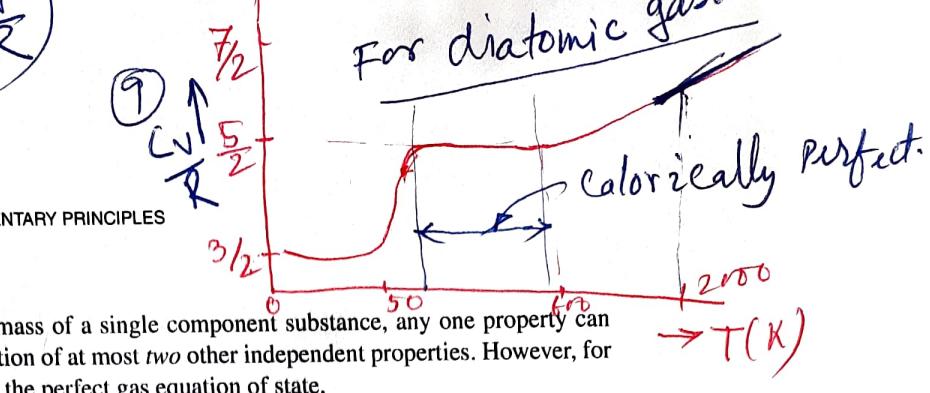
$I \ll \text{low}$   
Vibrational mode excites beyond  $2000^\circ K$

$$\frac{C_p}{R} = 1 + \frac{C_v}{R}$$

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### Perfect Gases

Recall that for a unit mass of a single component substance, any one property can be expressed as a function of at most *two* other independent properties. However, for substances that follow the perfect gas equation of state,



$$p = \rho RT \quad (1.13)$$

it can be shown (see p. 173 of Ref. 4) that *the internal energy and the enthalpy are functions of temperature only*. These are extremely important results, as they permit us to make many useful simplifications for such gases.

Consider the specific heat at constant volume:

$$c_v \equiv \left( \frac{\partial u}{\partial T} \right)_v \quad (1.37)$$

If  $u = f(T)$  only, it does not matter whether the volume is held constant when computing  $c_v$ ; thus the partial derivative becomes an ordinary derivative. Thus

$$c_v = \frac{du}{dT} \quad (1.42)$$

or

$$du = c_v dT \quad (1.43)$$

Similarly, for the specific heat at constant pressure, we can write for a perfect gas:

$$h = u + Pv = u + RT = f(T) \quad (1.44)$$

It is important to realize that equations (1.43) and (1.44) are applicable to *any and all processes* (as long as the gas behaves as a perfect gas). If the specific heats remain reasonably constant (normally good over limited temperature ranges), one can easily integrate equations (1.43) and (1.44):

$50K \text{ to } 600K \Rightarrow C_p \text{ & } C_v \text{ constant} \Rightarrow \text{calorically perfect.}$

$$\Delta u = c_v \Delta T \quad (1.45)$$

$$\Delta h = c_p \Delta T \quad (1.46)$$

In gas dynamics one simplifies calculations by introducing an arbitrary base for internal energy. We let  $u = 0$  when  $T = 0$  absolute. Then from the definition of enthalpy,  $h$  also equals zero when  $T = 0$ . Equations (1.45) and (1.46) can now be rewritten as

$$u = c_v T \quad (1.47)$$

$$h = c_p T \quad (1.48)$$

$$\textcircled{1} \quad C_p - C_v = R \quad C_p$$

$$\frac{C_p}{R} - \frac{C_v}{R} = 1$$

$$\frac{C_p}{R} = 1 + \frac{C_v}{R}$$

Typical values of the specific heats for air at normal temperature and pressure are  $c_p = 0.240$  and  $c_v = 0.171 \text{ Btu/lbm} \cdot ^\circ\text{R}$ . Learn these numbers (or their SI equivalents)! You will use them often.

Other frequently used relations in connection with perfect gases are

$$\gamma \equiv \frac{c_p}{c_v} \Rightarrow c_p = \gamma c_v \quad (1.49)$$

$$c_p - c_v = \frac{R}{J} \Rightarrow c_p - c_v = R \quad (1.50)$$

Notice that the conversion factor  $\Rightarrow c_v =$

$$J = 778 \text{ ft-lbf/Btu} \quad (1.51)$$

$\alpha_1 \gamma c_v - c_v = \frac{R}{J}$   
 $c_v = \frac{\gamma R}{\gamma - 1}$   
 $c_p = \frac{\gamma R}{\gamma - 1}$   
 in general

has been introduced in (1.50) since the specific heats are normally given in units of  $\text{Btu/lbm} \cdot ^\circ\text{R}$ . This factor will be omitted in future equations and it will be left for you to consider when it is required. It is hoped that by this procedure you will develop careful habits of checking units in all your work. What units are used for specific heat and  $R$  in the SI system? (See the table on gas properties in Appendix B.) Would this require a  $J$  factor in equation (1.50)?

### Entropy Changes

The change in entropy between any two states can be obtained by integrating equation (1.39) along any reversible path or combination of reversible paths connecting the points, with the following results for perfect gases:

$$ds = c_p \frac{dT}{T} - R \frac{dp}{P} \Rightarrow \Delta s_{1-2} = c_p \ln \frac{v_2}{v_1} + c_v \ln \frac{p_2}{p_1} \quad (1.52)$$

$$\Delta s_{1-2} = c_p \ln \frac{T_2}{T_1} - R \ln \frac{p_2}{p_1} \quad (1.53)$$

$$ds = c_v \frac{dT}{T} + R \frac{dv}{V} \Rightarrow \Delta s_{1-2} = c_v \ln \frac{T_2}{T_1} + R \ln \frac{v_2}{v_1} \quad (1.54)$$

Remember, absolute values of pressures and temperatures must be used in these equations; volumes may be either total or specific, but both volumes must be of the same type. Watch the units on  $c_p$ ,  $c_v$ , and  $R$ .

### Process Diagrams

Many processes in the gaseous region can be represented as a *polytropic process*, that is, one that follows the relation

$$T ds = P dv + c_v dT \quad p v^n = \text{const} = C_1 \quad (1.55)$$

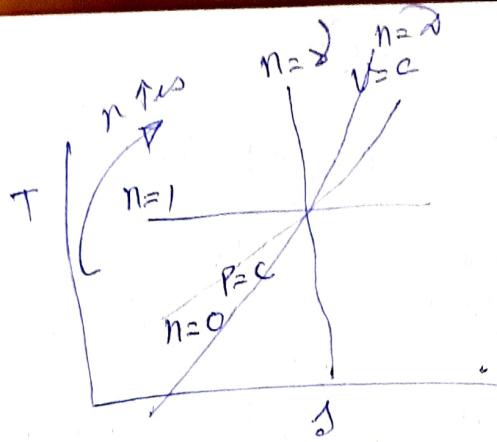
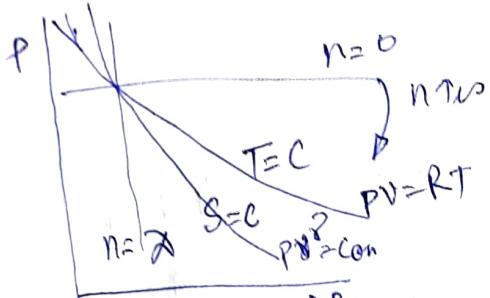
$$\begin{aligned} ds &= \frac{P}{T} dv + \frac{c_v}{T} dT \\ &= R \frac{dv}{V} + \frac{c_v}{RT} (P dv + V dp) \\ &= R \frac{dv}{V} + c_v \left( \frac{P dv + V dp}{RT} \right) \\ &= (R + c_v) \frac{dv}{V} + c_v \frac{dp}{P} \\ &= c_p \frac{dv}{V} + c_v \frac{dp}{P} \end{aligned}$$

From,

$$\begin{aligned} P &= PRT \\ PV &= RT \\ P dv + V dp &= R dT \\ \therefore dT &= \frac{P dv + V dp}{R} \end{aligned}$$

$$\frac{P}{T} = \frac{R}{V}, \quad RT = PV$$

(11)



$$\left(\frac{dp}{dv}\right)_{T=C} = -\frac{P}{v}$$

$$\therefore \left(\frac{dp}{dv}\right)_{S=C} = -\frac{\delta P}{v}$$

$$\therefore \left(\frac{dp}{dv}\right)_{S=C} > \left(\frac{dp}{dv}\right)_{T=C}$$

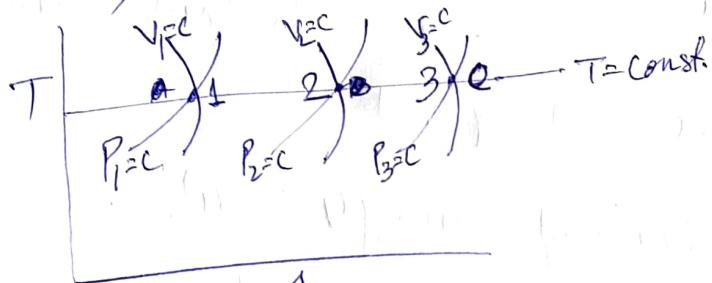
i.e. Isentropic Curves are Steeper than Isothermal

$$\left(\frac{dT}{dS}\right)_{V=C} = \frac{T}{C_V}$$

$$\left(\frac{dT}{dS}\right)_{P=C} = \frac{T}{C_P}$$

$$\left(\frac{dT}{dS}\right)_{V=C} > \left(\frac{dT}{dS}\right)_{P=C} [\because C_P > C_V]$$

For  $V=C$  curves,  $V$  goes rightward &  $P=C$  curves,  $P$  goes Leftward



For ~~T=C~~ For  $T=C$

$$ds = \frac{dT}{T} + R \frac{dv}{v}$$

$$ds = R \frac{dv}{v}$$

$$\int dv = \int \frac{ds}{R}$$

$$\therefore \ln \frac{v_2}{v_1} = \frac{s_2 - s_1}{R} = +ve$$

$$\therefore v_2/v_1 = e^{(s_2 - s_1)/R} > 1$$

$v$  goes rightward

$$\text{Similarly } ds = C_V \frac{dT}{T} - R \frac{dp}{p}$$

$$\int \frac{dp}{p} = - \int \frac{ds}{R}$$

$$\ln \frac{P_2}{P_1} = - \frac{\Delta S}{R}$$

$$\frac{P_2}{P_1} = e^{-\Delta S/R} < 1$$

$$\therefore P_1 > P_2$$

$P$  goes leftward.

# prove that  $n=1$  is isentropic process

$$ds = C_V \frac{dT}{T} + R \frac{dv}{v}$$

$$= C_V \left(1 - \frac{1}{n}\right) \frac{dv}{v} + R \frac{du}{u}$$

$$= \frac{dv}{v} \left(R + u(1 - \frac{1}{n})\right) \xrightarrow{ds=0} u = 0 \Rightarrow S=C.$$

$$\text{Now } PV^n = C_1 \Rightarrow \frac{RT}{V} \cdot V^n = C_1$$

$$\therefore T V^{n-1} = C_2$$

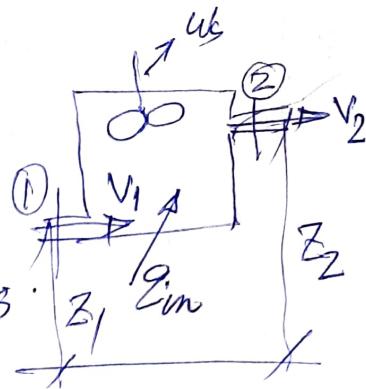
$$\frac{dT}{T} = (1-n) \frac{du}{u}$$

(12)

## # 1st law of thermodynamics

$$\frac{\delta Q}{\delta t} = \frac{\delta W_s}{\delta t} + \frac{\delta W_{pass}}{\delta t} + \frac{dE}{dT}$$

$$h_1 + \frac{V_1^2}{2} + gZ_1 + Z_{in} = h_2 + \frac{V_2^2}{2} + gZ_2 + W_s$$



## # 2nd law of thermodynamics :-

$$dS \geq \frac{\delta Q}{T} \Rightarrow dS \geq \frac{\delta Q}{T} \text{ per unit mass}$$

$$dS_R = \frac{\delta Q}{T} \text{ for Reversible}$$

$$dS_i \geq \frac{\delta Q}{T} \text{ u for irreversible.}$$

$$\# \text{ momentum eqn} : - \sum F_x = m(V_{out} - V_{in}).$$

$$\# \text{ Equat}^n \text{ of state} : P = \rho R T$$

$$\# \text{ Differential form of Energy eqn} : -$$

$$\delta Q = \delta W_s + dh_T \text{ when } h_1 = h + \frac{V^2}{2} + gZ.$$