Basic Definitions

1 Definitions (Alphabetic Order)

1.1 Absolute Quantity

Both pressure and temperature could be defined on absolute basis. If there are no molecules present in a volume, then there is no momentum flux, then $p_{abs} = 0$. When molecules in a substance can have zero (internal) energy, $T_{abs} = 0$. Notice here may not be vacuum, if the temperature reaches the lowest point, all molecules could not have motion, then $T_{abs} = 0$.

1.2 Adiabatic

System has no heat transfer across its boundaries.

1.3 Closed System

System where no **mass** can cross system boundary (control surface), also called **control mass**.

1.4 Control Surface

The imaginary boundary encloses the system.

1.5 Extensive Property

Depend on mass or extent of system.

1.6 Fluid

When a force is applied to the surface of a **solid**, it will experience a **finite** deformation, while **fluid** will experience a **continuously increasing deformation**.

1.7 Heat Transfer

Heat transfer is the transfer of thermal energy between physical systems depending on the temperature difference between them. It is a **process** that system is undergoing. There are 3 main types of heat transfer:

- 1. Heat Conduction: This is the transfer of heat within a body or between bodies in direct contact, due to the motion and interaction of molecules and free electrons. A common example is when one end of a metal rod is heated, the other end also gets hot after some time because heat is conducted from the hot end to the cooler end.
- 2. Heat Convection: This is the transfer of heat in a fluid (which can be a liquid or a gas) by the movement of heated particles to another region. This can occur naturally (due to the density differences caused by heating the fluid) or it can be forced (by a fan or a pump). A common example of convection is the heat circulation in a pot of boiling water.
- 3. Heat Radiation: This is the transfer of heat in the form of electromagnetic waves, typically infrared radiation. Unlike conduction and convection, radiation can occur in a vacuum, which is how the Sun's heat reaches Earth.

More details could be found here.

1.8 Homogeneous System

Has only one phase.

1.9 Inhomogeneous System

Has more than one phase.

1.10 Intensive Property

Do not depend on mass or extent of system. Notice that intensive property could be obtained from extensive property by dividing by mass. For example:

$$u = U/m \tag{1}$$

Here, u is the internal energy per mass, which is an intensive property.

1.11 Internal Energy (U)

Molecules have energy due to their internal structure, which is called **internal energy**. It includes:

- 1. Translational KE (of center-of-mass (CM) of molecule)
- 2. Rotational KE (of nuclei abour CM)
- 3. Vibrational KE+PE (of nuclei relative to CM)
- 4. Electronic KE+PE (of electrons relative to nuclei)
- 5. Chemical PE (between nuclei)
- 6. Nuclear (within nucleus)

1.12 Isolated System

No energy and mass pass through control surface of the system.

1.13 Kinetic Energy (KE)

The energy that a molecule possesses due to its motion:

$$KE = \frac{1}{2}mv^2\tag{2}$$

1.14 Knudsen Number (Kn)

For liquid, we can easily treat them as continuous. However, for the gas, this may be hard. We use **Knudsen Number** to justify the continuity:

$$Kn = \frac{\lambda}{L} \tag{3}$$

Here λ is the mean free path, L is the dimension of any body/surface that can also interact with the gas molecules. If $Kn \leq 0.01$, we can say the flow is **continuum flow**. In contrast, if Kn > 0.01, we call the flow as **rarified flow**.

1.15 Macroscopic Matter

- 1. Ignore the internal structure of masses
- 2. Treat the matter as continuous (no empty spaces)
- 3. Combine (sum, average, ...) effects caused by the presence and interactions of the many **microscopic** particles within a mass into **macroscopic** properties
- 4. This approximation may not always work if there are too few particles to average.
- 5. Also does not work if many properties of matter depend on internal structure

1.16 Mean Free Path (λ)

Distance a molecule travels on average before colliding with another molecule.

1.17 Microscopic Matter

Molecules, atoms, electrons, protons, neutrons, quarks, leptons.

1.18 Open System

Matter can pass through boundary, also called **control volume**.

1.19 Path

The specifics of how the state of a TD system changes from initial state to final state. **Path function** refers the quantity that depends on the path taken, such as the amount of work and heat transfer involved in the process. **TD properties mathematically described as point functions.**

1.20 Phases of Matter

Normally we divide the phases of matter into: solid, liquid, gas and plasma.

Gas	Liquid	Solid
Assumes shape and volume of its container	Assumes shape of the part of container it occupies	Retains a fixed volume and shape
Compressible – lots of free space between particles	Not easily compressible – little free space between particles	Not easily compressible – little free space between particles
Flows easily – particles move past one another	Flows – particles can move past one another but often colliding	Rigid – most particles can't move past one another

Figure 1: Phases of Matter

1.21 Plasma

A gas where charged particle interactions are dominant.

1.22 Potential Energy (PE)

The energy that a molecule possesses due to its location in field. Normal types:

- 1. Spring
- 2. Gravitational
- 3. Electric

1.23 Power

Work rate (per time) is power.

1.24 Pressure

Normal force per unit area, or called stress:

$$P = \frac{dF}{dA} \tag{4}$$

From Newton's Law, we know force is due to a change in momentum. Therefore, we can say **pressure exists due to rate of change of momentum as molecules cross any imaginary plane**.

1.25 Process

What actually happens and how a system evolves as it undergoes a change

1.26 Property

A measurable quantity or characteristic of system.

1.27 Pure Substance

Has same chemical structure in all states, but may have several distinguishable phases.

1.28 Quasi-Equilibrium Process

If we assume TD properties remain well-defined, system remains in equilibrium along each infinitesimal step in process, we call this process as **quasi-equilibrium process**, which is **internally reversible**

1.29 State

A unique combination of all the TD properties of a system. If we know values for **all** the TD properties of a system, its state **is defined**. Notice that we can only define the state of a system in **thermodynamic equilibrium**, because many TD properties could only be defined in TD equilibrium.

1.30 Surrounding

The rest of the world outside the system.

1.31 System

A region enclosed by an imaginary boundary for analysis.

1.32 Thermodynamic Equilibrium

When system has no **internal tendency** to change in time. It includes:

- 1. Mechanical Equilibrium: Mechanical forces are in balance throughout system.
- 2. Thermal Equilibrium: Temperature is uniform in system
- 3. Chemical and Phase Equilibria (composition): Relative amount of each chemical substance and its distribution among various phases is same as it would achieve if that matter was isolated and at same T and P.

Notice that TD Equilibrium is not the same as steady-state d/dt = 0. We can externally "force" system causing it to change $d/dt \neq 0$ but stay in equilibrium.

1.33 Thermodynamic Property

This refers the property of matter or energy that are independent of any external coordinate system. Notice that we can make new TD property by combining other TD properties, for example enthalpy:

$$H = U + pV \tag{5}$$

By adding kinetic energy, we can get the stagnation enthalpy:

$$H_0 = H + KE = H + \frac{1}{2}mv^2 \tag{6}$$

Notice that **this is not a TD property!** It depends on velocity, which **depends** on external reference frame

1.34 Total Energy

In macroscopic view, the total energy includes:

$$KE(\text{macroscopic}) + PE(\text{macroscopic}) + U(\text{microscopic})$$
 (7)

1.35 Work

Work is the energy transferred to or from a system by means other than heat transfer. For a system to do work, it must exert a force over a displacement.