Lecture 2 – Definitions

Purdue ME 200, Thermodynamics I

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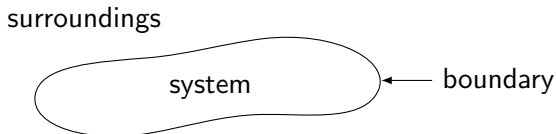
Outline

System, surroundings and boundary

- a **system** is a portion of space or matter we want to study
- the **surroundings** are everything but the system
- the **boundary** separates the system from the surroundings

Types of system

- an **open system** or **control volume** is a region of space
- a **closed system** or **control mass** is a collection of matter
- an **isolated system** does not interact with the surroundings

	open	closed	isolated
energy crosses boundary?	yes	yes	no
matter crosses boundary?	yes	no	no

Outline

Extensive, intensive and specific properties

- system **properties** are characteristics that can be quantified with no knowledge of the system's history
- **extensive** properties are additive when systems combine (clone the system and they double)
- **intensive** properties are independent of system size (clone the system and they stay the same)
- a **specific** property is an extensive property per unit mass
- e.g., **specific volume** is $v = V/m = 1/\rho$ (in m^3/kg)
- is specific volume extensive? intensive?

Pressure

- **pressure** is force per unit area
- its units are Pascals ($1 \text{ Pa} = 1 \text{ N/m}^2$)
- common assumption:
  - $\diamond$ the pressure at any point is the same in all directions
- this is true for fluids at rest
- it's a good approximation for most purposes in this class
- this assumption lets us calculate pressure by
  1. finding the weight pressing down on a horizontal surface
 2. dividing by the surface area
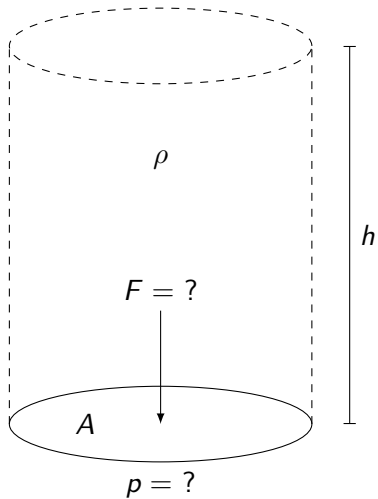
Absolute, gage and vacuum pressures

- this class almost always uses **absolute pressure**
  - ⋄ absolute pressure is measured relative to a vacuum
  - ⋄ a system's absolute pressure is denoted by p
- **atmospheric pressure** is $p_{atm} = 101.3$ kPa
- the quantity $|p - p_{atm}|$ is called
  - ⋄ **gage pressure** for systems above atmospheric pressure
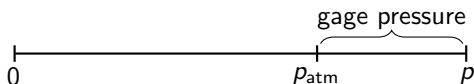
  - ⋄ **vacuum pressure** for systems below atmospheric pressure
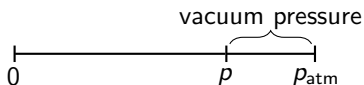
Temperature

- **temperature** is a surprisingly subtle concept
- it's not quite what we sense as 'hot' or 'cold'
- it's closely related to
  - ⋄ a pointer in the direction of natural heat transfer
  - ⋄ an averaged kinetic energy of molecules, atoms or particles

Thermal equilibrium

- imagine a block that feels hot and another that feels cold
- bring them together and periodically touch them
- the hot block will start to feel cooler; the cold block, warmer
- eventually, their feelings of warmth will stop changing
- and both blocks will feel equally warm

- at that point, the blocks are in **thermal equilibrium**
- and we postulate that their temperatures are equal

0th law of thermodynamics

- **0th law**: thermal equilibrium is transitive
    - ⋄ if A is in thermal equilibrium with B
    - ⋄ and B is in thermal equilibrium with C
    - ⋄ then A is in thermal equilibrium with C
- the 0th law underpins all thermometers
- to measure the temperature of thing C,
    - ⋄ in the lab, calibrate thermometer B to well-understood A
   - ⋄ in the field, bring B into thermal equilibrium with C

Defining temperature scales

- pick temperatures T_1 and T_2 at conditions 1 and 2
- a degree is defined as the fraction $1/(T_2 - T_1)$
- Anders Celsius picked
  - $\diamond$ $T_1 = 0$ °C at water's freezing point at atmospheric pressure
  - $\diamond$ $T_2 = 100$ °C at water boil
  - $\diamond$ so 1 °C is 1/100th of the ΔT between water freeze and boil
- Daniel Fahrenheit picked
  - $\diamond$ $T_1 = 0$ °F on a very cold day (later 32 °F at water freeze)
  - $\diamond$ $T_2 = 100$ °F at body temperature (later 212 °F at water boil)
  - $\diamond$ so 1 °F is 1/180th of the ΔT between water freeze and boil

Absolute temperature scales

- experiments show that there's a lower limit on temperature
- we call this temperature **absolute zero**
- **absolute temperature scales** have $T = 0$ at absolute zero
- Kelvin and Rankine defined absolute temperature scales
  - ⋄ Kelvin (William Thomson) matched Celsius' degree definition
    (1 K is 1/100th of the ΔT between water freeze and boil)
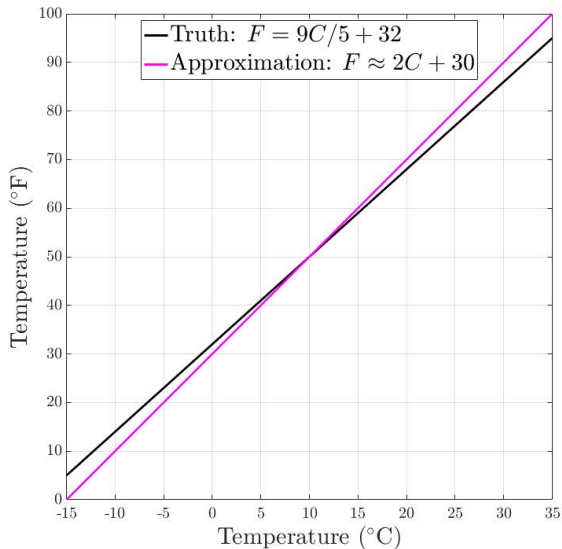 - ⋄ William Rankine matched Fahrenheit's definition (1/180th)

Temperature scales
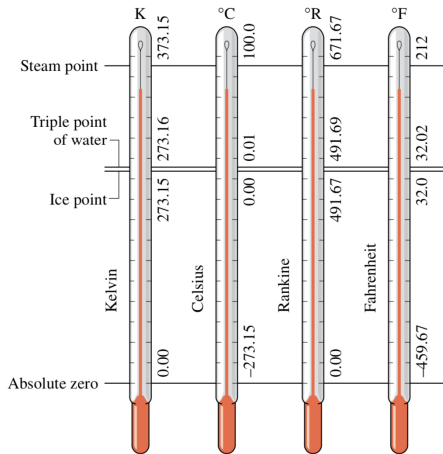
$$T[\text{K}] = T[^\circ\text{C}] + 273.2 = 5\,T[^\circ\text{R}]/9 = 5(\,T[^\circ\text{F}] + 459.7)/9$$

Moran et al. (2018): *Fundamentals of Engineering Thermodynamics*.

Mental conversion between °C and °F

Outline

States and equilibrium

- a **state** is a list of properties that fully characterize a system
- a system in **equilibrium** does not change state when isolated
  (loosely, properties are spatially uniform within the system)
- in equilibrium, the state is usually a *short* list of properties
- out of equilibrium, the list may be long and complicated
- an **equation of state** relates properties to one another
- different systems have different equations of state

Ideal gas example

- in an **ideal gas**, particles do not interact with one another
  - $\diamond$ no real gas behaves exactly like an ideal gas
  - $\diamond$ but it's often a good approximation (e.g. for air at p_{atm})
- a system of ideal gas has equation of state $pv = RT$
  - $\diamond$ $R = \bar{R}/M$ (in J/kg/K) is a constant for any particular gas
  - $\diamond$ $\bar{R} = 8.31$ J/mol/K is the universal gas constant
  - $\diamond$ $M = m/n$ is the molecular weight of the gas (in kg/mol)
  - $\diamond$ m is the mass of gas in the system
  - $\diamond$ n is the number of moles of gas in the system
- any two of (p, v, T) fully characterize a system of ideal gas
  - $\diamond$ given (p, v), temperature is determined by $T = pv/R$
  - $\diamond$ given (v, T), pressure is determined by $p = RT/v$
  - $\diamond$ given (p, T), specific volume is determined by $v = RT/p$

Quick exercise

Convert the ideal gas law from the form $pv = RT$ to $pV = n\bar{R}T$.

Hint: Use $v = V/m$, $R = \bar{R}/M$ and $M = m/n$.

Solution

$$pv = RT$$
$$p(V/m) = (\bar{R}/M)T$$
$$pV = \frac{m}{M}\bar{R}T$$
$$pV = \frac{m}{(m/n)}\bar{R}T$$
$$pV = n\bar{R}T$$

Processes

- a **process** is a transition between two equilibrium states
  (a process can pass through non-equilibrium states)
- a **path** is a sequence of states that a process passes through
- a **cycle** is a process that begins and ends in the same state
- a **quasi-equilibrium process** stays close to equilibrium states
- quasi-equilibrium approximation is inexact, but often useful

Non-equilibrium process
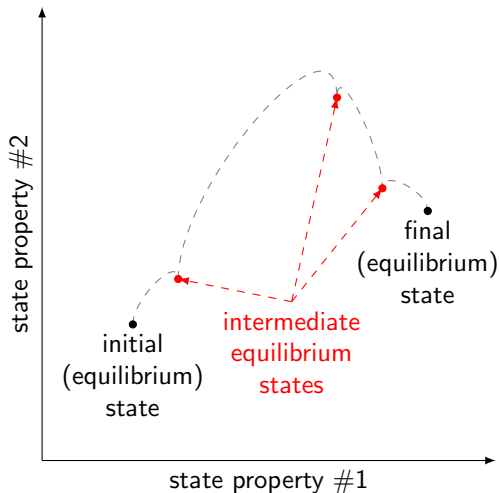
- system passes through weird, high-dimensional states
- path can't be visualized in two dimensions

Quasi-equilibrium process
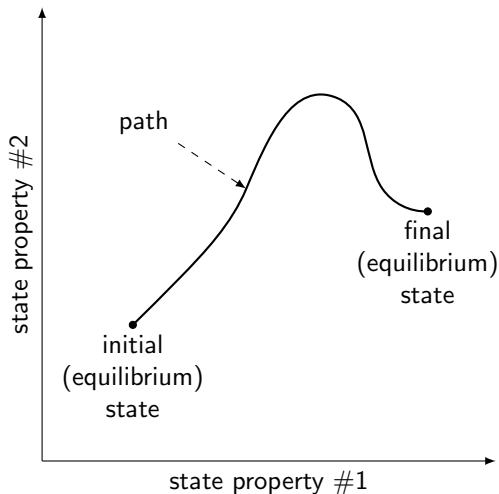
- system stays very close to equilibrium states
- path can be visualized in two dimensions