

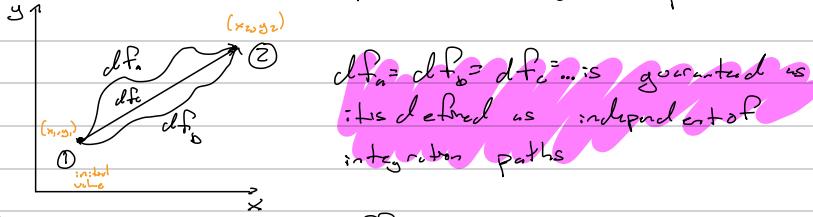


Midterm Exam 1



Math Review

① **Differentials** \equiv an infinitesimal change in a function or variable
 ex: If $f = f(x, y)$, then the total differential of f is $df = P(x, y)dx + Q(x, y)dy$
 and is exact and independent of integration paths



② **Partial Derivatives** \equiv the differential change in one variable with respect to a differential change in another variable, while holding other variables fixed/constant

ex: Let $f(x, y) = x^2y + xy + \sin x \sin y$.

$$\begin{aligned} \text{Diff. change in } f & \text{ (} \frac{\partial f}{\partial x} \text{)} @ \text{constant } y = \left(\frac{\partial f}{\partial x} \right) \\ \text{Diff. change in } x & \text{ (} \frac{\partial x}{\partial x} \text{)} \end{aligned}$$

then $\left(\frac{\partial f}{\partial x} \right) = 2xy + \sin y + \cos x \sin y \leftarrow P$

Similarly $\left(\frac{\partial f}{\partial y} \right) = x^2 + x \cos y + \sin x \cos y \leftarrow Q$

Now note that for our total differential, $df = P(x, y)dx + Q(x, y)dy$
 \rightarrow if y is held constant and hence $dy = 0$, $df = P(x, y)dx \Rightarrow \left(\frac{\partial f}{\partial x} \right)_y = P(x, y)$
 \rightarrow similarly for constant x , then $\left(\frac{\partial f}{\partial y} \right)_x = Q(x, y)$

③ **Integration** \equiv continuous sum over differential changes

$$\int f(x) dx$$

Integral \uparrow Differential

1. **Indefinite Integral:** No upper/lower bounds

$$\text{ex: } \int (A + BT + CT^2 + \frac{D}{T} + \frac{E}{T^2}) dT \quad \text{where } A, B, C, D, E \text{ are constants, } T \text{ is temperature}$$

$$= AT + \frac{1}{2}BT^2 + \frac{1}{3}CT^3 + DT - E + (\text{constant})$$

2. **Definite Integral:** Specified lower & upper bounds (no constant of integration)

$$\int_{T_0}^{T_1} (A + BT) dT = \left[AT + \frac{B}{2}T^2 \right]_{T=T_0}^{T=T_1} = A(T_1 - T_0) + \frac{B}{2}(T_1^2 - T_0^2)$$

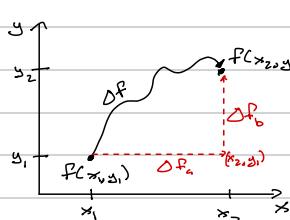
Summary - Common points of confusion!

① **Differential:** infinitesimal changes denoted df

② **Derivative:** change in one variable with respect to another (\approx differential / another differential)
 ex: $\frac{df}{dx}, \frac{df}{dt}, \frac{df}{dx}_y$

③ **Integral:** Continuous sum over differential changes
 ex: $f = f(x)$: $f(x_b) - f(x_a) = \int_{x_a}^{x_b} df$ problem: Need to integrate w.r.t. variable in our bounds of integration $= \int_{x_a}^{x_b} \left(\frac{df}{dx} \right) dx$ derivative!

ex: $f = f(x, y)$:



$$f(x_2, y_2) - f(x_1, y_1) = \Delta f = \Delta f_a + \Delta f_b$$

$$= \int_{x_1}^{x_2} \left(\frac{\partial f}{\partial x} \right)_{y=y_1} dx + \int_{y_1}^{y_2} \left(\frac{\partial f}{\partial y} \right)_{x=x_2} dy$$

We can do this bc the differential is independent of path

* Math no-no!

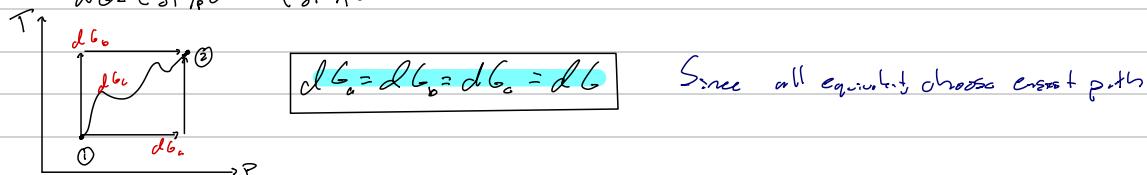
$\times \Rightarrow \int \frac{df}{dx}$ cannot be evaluated! Missing the differential!

\hookrightarrow Must integrate over differential changes

$\checkmark \Rightarrow \int \left(\frac{df}{dx} \right) dx$ is allowed!

Chapter 1 - Thermodynamic Properties

Thermodynamic Properties are state functions that do not depend on path and have exact differentials.
 ex: $G = G(T, P)$ Gibbs Free Energy
 $dG = \left(\frac{\partial G}{\partial T}\right)_P dT + \left(\frac{\partial G}{\partial P}\right)_T dP$



Inexact Differentials are path functions (i.e., they do depend on path)

ex: work (w) and heat (Q)

→ dW and dQ do depend on path

→ use notation δQ and δW to denote an inexact differential

What is Thermodynamics?

↳ A self-contained, logically consistent theory that can be expressed in two laws:

• First Law: Conservation of Energy

$\Delta E_{\text{univ}} = 0$ for any process

• Second Law: Entropy increases

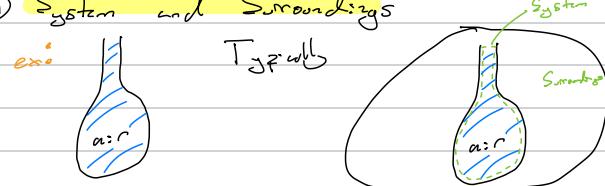
$\Delta S_{\text{univ}} \geq 0$ for any process

• Two benefits of the 2nd Law

(i) Is our process consistent with 2nd Law?

(ii) Entropy of system + surroundings should reach a maximum over time.

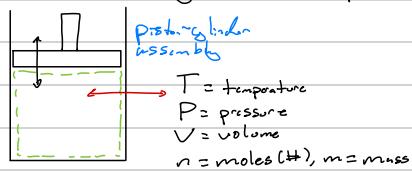
① System and Surroundings



• Three Types of Systems

| | "closed" | "open" | "isolated" |
|-----------------|----------|--------|------------|
| energy transfer | yes | yes | No |
| mass transfer | No | yes | No |

② Thermodynamic Properties of Fluids (mostly gases & liquids, only some solids)



Types of Properties

- (i) extensive = varies w/ size of system
- (ii) intensive = independent of system size

Q: Which of the following thermodynamic properties are intensive? (out of P, T, n, V, ρ)

$$\textcircled{1} P, T, n$$

$$\textcircled{3} V, n$$

$\textcircled{2} P, T, \rho$ None of these depend on how large the system is, or how much stuff is in it!

| | | | |
|------------------------------------|--|----------|----------|
| P | V | T | n |
| ↑ intensive (C_{∞}) | ↑ extensive (C_{ext}) | ↑ int | ↑ ext |

density

We can convert ext quantities to int ones:

extensive = "specific property" ↪ lowercase with no $\frac{1}{m}$
 mass → intensive quantity normalized by mass

$$\text{ext: } \frac{\text{Volume}}{\text{mass}} = \frac{V}{m} = \frac{1}{\rho} \quad \text{specific volume}$$

extensive = "molar property" ↪ lowercase w/ subscript m
 moles → intensive

$$\text{ext: } \frac{\text{Volume}}{\text{moles}} = V_m \quad \begin{array}{l} \text{relation for ext} \\ \text{readily & agreed} \\ \text{is distinct from textbook} \end{array}$$

③ "State" of a System

- The "state" of a system depends on its thermodynamic properties
 ↳ can be described by the Gibbs phase rule:

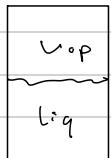
$$f = m - \pi + 2 \quad \text{for non-reactive systems}$$

\uparrow degrees of freedom
 \uparrow # of chemical species
 \uparrow # of phases

o Degrees of Freedom \Rightarrow # of intensive properties that must be specified to constrain properties of system

- To fully specify a state, also need to know amount of material

Example: Vapor-liquid mixture of H_2O



$$m=1 \quad \pi=2 \quad \Rightarrow f=1$$

mol/mol of liq mol/mol vap

$$V_{\text{tot}} = V = V_{m,\text{L}} n_{\text{L}} + V_{m,\text{V}} n_{\text{V}} \quad \Rightarrow \text{divide by } V_{\text{tot}} = n = n_{\text{L}} + n_{\text{V}}$$

$$\Rightarrow \frac{V}{n} = V_m = (1-x) V_{m,\text{L}} + x V_{m,\text{V}}$$

\uparrow average molar volume
 $x \equiv \text{"quality"}$

$$\text{Also: } \beta = (1-x)\beta_{\text{L}} + x\beta_{\text{V}}$$

We can't for example specify V_m for each phase

→ constrains all other intensive properties

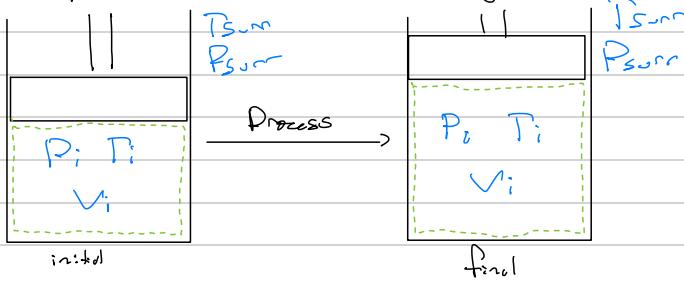
Also need to specify total amount in each phase

$$\rightarrow V_{\text{tot}} = V_{\text{liq}} + V_{\text{vap}}$$

$Q_{\text{liq}} = \text{mole or mass fraction of } \text{H}_2\text{O}$
 for saturated (vapor) mixture

④ Processes

- A process is a change in the system



4 Special Types of Processes

- Isothermal** $\rightarrow \text{const } T$
- Isobaric** $\rightarrow \text{const } P$
- Isochoric** $\rightarrow \text{const } V$
- Adiabatic** $\rightarrow \text{No heat transfer to/from sum}$
↳ can also be described as well-insulated

⑤ Equilibrium

- A system does not change @ equilibrium and there is no driving force for change

→ Examples:

(i) Thermal Equilibrium $\rightarrow T_{\text{sys}} = T_{\text{sum}}$ (also btwn phases)

(ii) Force/mechanical $\rightarrow P_{\text{sys}} = P_{\text{sum}}$

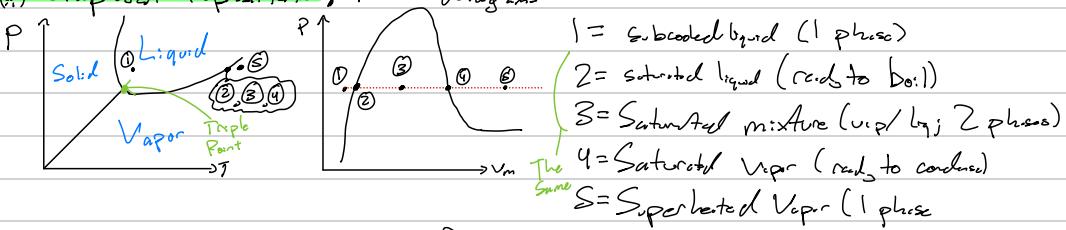
(iii) Phase Equilibrium \rightarrow Composition of phases is constant

(iv) Chemical Equilibrium \rightarrow No mass transfer into/out of a phase, or no net chemical reaction

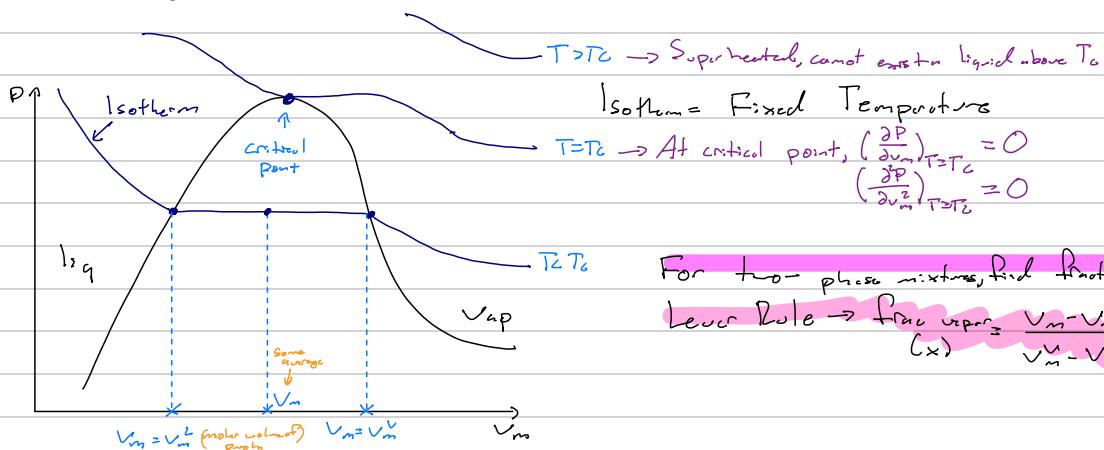
Thermodynamic Property Sources

(i) Experiment $\rightarrow P, T, V, \dots$

(ii) Graphical Representation: PVT diagrams



• PVT diagrams are very useful!



(iii) Thermodynamic Tables (ex. steam tables \rightarrow Appendix B of Kotz et al.)

• Always use steam tables for water

• Linear interpolation required for conditions between values in the table

(iv) Equations of State

\rightarrow ideal gas law, $PV_m = RT$

\rightarrow van der Waals...

(v) Statistical Mechanics (molecular calculations)

\hookrightarrow not focused on in this course

Final Notes about (Potentially) 2-phase systems

- If $V \geq V_{i,c}^{sat}$ you have pure vapor
- If $V_{i,c}^{sat} > V > V_{i,l}^{sat}$ you have a mixed phase
- If $V_{i,l}^{sat} \geq V$ you have pure liquid

Similar ideas work for any specific (or master) property: v_m, u_m, h_m, \dots

Example A rigid 100L container w/ saturated water @ 100°C. If the initial reaches the critical point.

a) Draw the process on a P-v diagram.

initial (i)

$$V_i = 100L$$

$$T_i = 100^\circ C$$

saturated

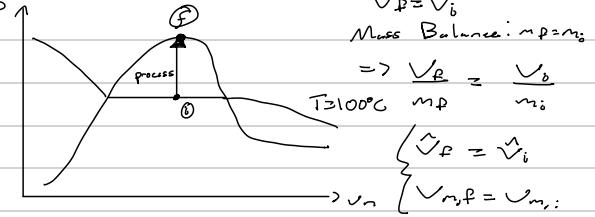
final (f)

$$V_f = 100L = V_i$$

$$T_f = T_c$$

$$P_f = P_c$$

process reaches
at critical point



b) Find the mass and initial quality.

(outcomes)

→ mass balance: $m_f = m_i$

rigid: $V_f = V_i = V$

$$\rightarrow V_f = V_i = V$$

Steam Tables

because P_c , T_c state is a crit point

$$V_p = V_i = V_c = 0.0033155 \frac{m^3}{kg}$$

$$V = 0.0033155 \frac{m^3}{kg}$$

bottom of tube

↳ of temp → $V_{i,g} = 0$

$$M = \frac{V}{V} = \frac{100 L}{0.0033155 \frac{m^3}{kg}} \times \frac{1 m^3}{1000 L} = 31.6 \text{ kg}$$

(continued)
Q. 1. b

$$V_i = V_{i,L}^{sat} (1-x) + V_{i,V}^{sat} x$$

$$x = \frac{V_i - V_{i,L}^{sat}}{V_{i,V}^{sat} - V_{i,L}^{sat}} = 0.0015 \quad (\text{nearly all liquid})$$

S. + Steam C 100°C

$$V_{i,L}^{sat} = 0.001047 \frac{m^3}{kg}$$

$$V_{i,V}^{sat} = 1.4194 \frac{m^3}{kg}$$

$$\rightarrow V_{i,L}^{sat} < V < V_{i,V}^{sat} \Rightarrow L/V mix$$

Chapter 2 - The First Law of Thermodynamics

- First Law: $\Delta E_{\text{univ}} = 0$ or $E_{\text{univ}} = \text{constant}$

Q: What is E_{univ} ?

1) Macroscopic Kinetic Energy ($E_k = \frac{1}{2} m (v)^2$)

2) Macroscopic Potential Energy ($E_p = mg$)

3) Internal Energy (U_m): all energy at molecular level
 $\rightarrow U_m$ includes:

- molecular kinetic energy $\rightarrow T$

- molecular potential energy (attractions/repulsions between molecules)

o Changes in U_m come from:

First part of the class - changes in T, P , or other properties

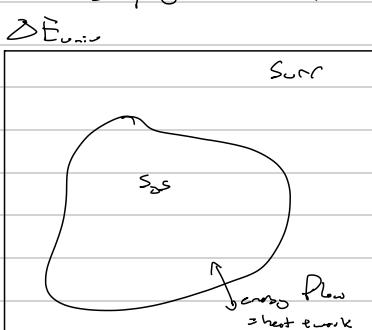
- phase changes

- chemical reactions (breaking/forming chemical bonds)

Note: An ideal gas by definition has no (0) potential energy between molecules

$U_m = U_m(T)$ only true for ideal gases

- We can simplify 1st Law (for closed systems):



$$\Delta E_{\text{sys}} = -\Delta E_{\text{surr}} = Q + W$$

heat
work

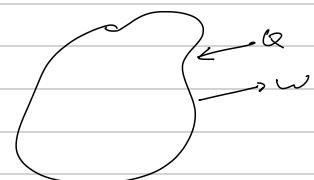
Q: Heat and work are flowing across the system boundary. What can we say about Q and W?

1) both $Q & W > 0$

2) $Q > 0, W < 0$

3) $Q < 0, W > 0$

4) both $Q & W < 0$



Recall $\Delta E = \text{internal} + \text{kinetic} + \text{potential}$

$$\Delta E_{\text{sys}} = \Delta U + \Delta E_k + \Delta E_p$$

Leave off "sys" for convenience

so, exterior: $\Delta U + \Delta E_k + \Delta E_p = Q + W$ First Law for closed systems

interior: $\Delta U + \Delta E_{\text{int}} + \Delta E_{\text{p,int}} = Q + W$

And for small processes, $\Delta u \rightarrow du$

$$du + dE_{\text{int}} + dE_{\text{p,int}} = \delta Q + \delta W \leftarrow Q \text{ and } W \text{ are moment differentials for path functions}$$

Accumulation

How do we calculate variables in the first law?

Not currently focused on
 in this course → ECHE360

① Heat (Q): energy transferred across system boundaries by Temperature gradients (conduction, radiation, etc.)

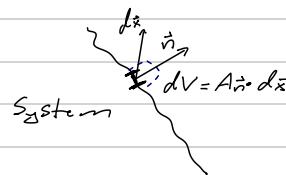
② Work (W): this class!

a) System boundary moves against an external force

$$\rightarrow \text{From physics, } W = \int_{\text{path}} \vec{F}_E \cdot d\vec{x}$$

where $\vec{F}_E = -P_E \vec{A}$

$$\text{so then, } W = - \int_{\text{path}} P_E \vec{A} \cdot d\vec{x} = - \int P_E dV \quad \text{"PV work"}$$



→ Calculating PV work requires us to know external pressure

* P_E does not always equal P_{ext}

b) Shaft Work (W_s)

example: propeller in fluid (macroscopic, not PV)

$$\therefore W = W_s - \int P_E dV$$



Work (and Heat) can be transferred either reversibly or irreversibly.

Reversible Processes - Characteristics & overview

- a) Slow → never more than differentially away from equilibrium
- b) Driven by small changes in forces (or pressures)
- c) Can be reversed at any point
- d) No friction or dissipation → highly idealized process

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Ex

Reversible work for ideal gas expansion under isothermal conditions.



- isothermal → T is constant
- ideal gas → $PV_m = RT$
- work → PV work → $W = - \int P_E dV$
- mole balance → $n = \text{constant}$ (closed sys)

• reversible → $P_E \approx P$

$$\text{Thus, } W = - \int P_E dV \stackrel{\text{rev.}}{\approx} - \int P dV \\ \stackrel{\text{IGL}}{=} - \int \frac{RT}{V_m} d(V_{m,f} - V_{m,i})$$

n is constant

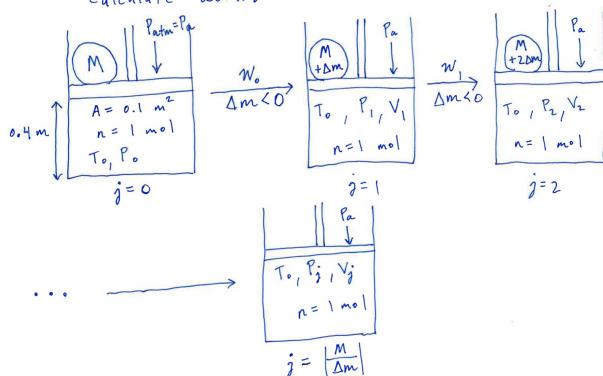
$$\frac{W}{n} = w_m = - RT_i \int_{V_{m,i}}^{V_{m,f}} \frac{dV_m}{V_m}$$

$$w_m = - RT_i \ln \left(\frac{V_{m,f}}{V_{m,i}} \right)$$

reversible work,
isothermal,
ideal gas

ex: ideal gas in a piston expands isothermally in increments of changes in external pressure.

calculate work.



• mole balance: $n = \text{const} = 1 \text{ mol}$

• final mass = 0

• ideal gas: $nRT_0 = P_0V_0 = P_1V_1 = P_2V_2 = \dots = P_jV_j$

• isothermal: $T = \text{const} = T_0$

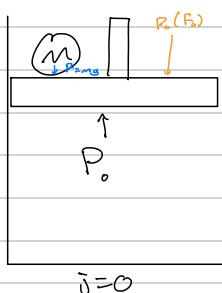
• PV work def'n → $W = - \int P_E dV$ external pressure

↳ what is P_E ?

• At eq. P_E should equal P_{eq}

• pressure / force balance:

Pressure / Force Balance



$$\text{initial } (j=0): A P_0 = A P_{E,0} \Rightarrow A \left(\frac{Mg}{A} + P_0 \right)$$

cross-section area of cylinder

$$\rightarrow P_0 = P_{atm} + \frac{Mg}{A}$$

$$\text{Similarly } (j=1): P_1 = P_{E,1} = \left(\frac{(M+\Delta m)g}{A} + P_0 \right)$$

$\Delta m < 0$ for this problem given

$$W_0 = - \int_{V_0}^{V_1} P_{E,1} dV = - P_1 \int_{V_0}^{V_1} dV = - P_1 (V_1 - V_0) = - P_1 \left(\frac{P_0 V_0}{P_1} - V_0 \right) = - V_0 (P_0 - P_1)$$

from $j=0$ to $j=1$

$P_{E,1} = P_1$

Subsidy law

$P_1 V_1 = P_0 V_0$

$$\text{Now, sub } P's \text{ from force balance} = - V_0 \left[\left(P_1 + \frac{Mg}{A} \right) - \left(P_0 + \frac{Mg}{A} + \frac{\Delta m g}{A} \right) \right] = + V_0 \cdot \frac{\Delta m g}{A} < 0$$

negative indicates S_d class work

$$\text{Similarly, } W_1 = - P_{E,2} \Delta V = - P_2 (V_2 - V_1) = - P_2 V_2 \left(1 - \frac{V_1}{V_2} \right) = - P_0 V_0 \left(1 - \frac{V_1}{V_2} \right) = - P_0 V_0 \left(1 - \frac{P_2}{P_1} \right)$$

$(\text{from } j=1 \text{ to } j=2)$

$$\rightarrow \text{use force balance} = - V_0 \left(P_{atm} + \frac{Mg}{A} \right) \left(\frac{(-\Delta m g)}{P_0 + \frac{Mg}{A} + \frac{\Delta m g}{A}} \right)$$

P_0

$$\Rightarrow W_1 = \frac{V_0 \Delta m g}{A} \cdot \frac{P_0 + \frac{Mg}{A}}{P_0 + \frac{Mg}{A} + \frac{\Delta m g}{A}} < 0$$

Expansion $\rightarrow S_d$ class work

Example Continued

$$W_2 = -P_3(V_3 - V_2) = \frac{V_0 \Delta mg}{A} \frac{\frac{mg}{A}}{P_a + \frac{mg}{A} + \frac{\Delta mg}{A}} \text{ etc...}$$

(from $j=2$
to $j=3$)

so, total work done is:

$$W_{\text{tot}} = \sum_{j=0}^{j=1 \dots 1-1} W_j = \sum_{j=0}^{j=1-1} \left[\frac{V_0 \Delta mg}{A} \cdot \frac{\frac{mg}{A}}{P_a + \frac{mg}{A} + j \frac{\Delta mg}{A}} \right]$$

What happens w/ different $\Delta m's$? Suppose $M = 1020 \text{ kg}$

| Final State | $\Delta m (\text{kg})$ | $W_{\text{tot}} (\text{J})$ |
|-------------|------------------------|-----------------------------|
| $j=1$ | -1020 | -4000 |
| $j=2$ | -510 | -4666 |
| $j=3$ | -340 | -4928 |
| : | : | : |

Thermodynamic properties (i.e., path independent quantities) like T, P, V do not change between initial and final states between scenarios.

→ But work is not a thermodynamic property.
It is path dependent

As we increase the number of steps w/ smaller changes per step, magnitude of work done by the system on the surroundings increases

The Limit? $\lim_{\Delta m \rightarrow 0}$ = reversible process w/ max work done by system on surroundings

($P_E \approx P_{\text{sys}}$ for all steps j)

$$\text{Reversible: } W = - \int_{V_0}^{V_F} P_B dV = - \int_{V_0}^{V_F} P dV$$

$P_B = P$ for a reversible process
changes happening very slowly, so
 $P \approx P_E$ (small changes)

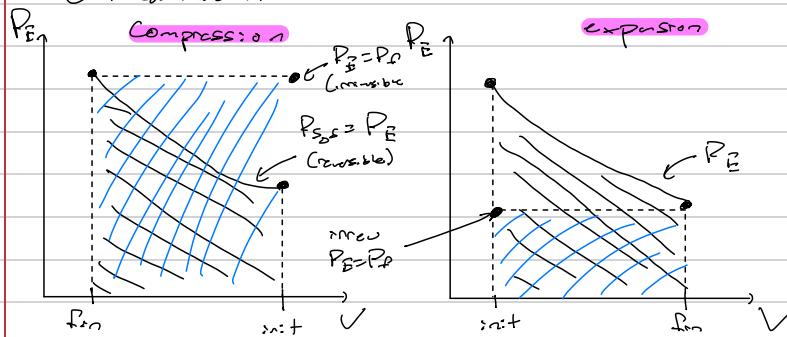
This is a reason why assuming a reversible process can be convenient. $W_{\text{re}} = -5545 \text{ J}$

We've shown that, for closed systems that undergo an isothermal reversible change in Volume, that:

$$W_{\text{isoth}} = -RT \ln \left(\frac{V_{\text{final}}}{V_{\text{initial}}} \right)$$

→ If a process were reversible and $V_F < V_i$:

- ① $|W_{\text{min,rev}}| > |W_{\text{max,rev}}|$ This is compression $P_E > P$
- ② $|W_{\text{min,rev}}| < |W_{\text{max,rev}}|$ True for expansion
- ③ $|W_{\text{min,rev}}| = |W_{\text{max,rev}}|$
- ④ Need more information



$$|W_{\text{min,rev}}| > |W_{\text{max,rev}}| \quad |W_{\text{min,rev}}| < |W_{\text{max,rev}}|$$

$$W = - \int P_E dV$$

Internal Energy

• How do we calculate ΔU (dU)?

→ Several Options:

(i) Property tables (NIST, steam tables)

(ii) Specific (molar) heat capacities

a) Ideal Gas: $U_m = U_m(T)$ only (Combining of Temperature)
 $\rightarrow \text{Thus, } C_{v,m} = \frac{dU_m(T)}{dT}$ for ideal gas only

Typically correlations are used: $C_{v,m}(T) = A + BT + CT^2 + \frac{D}{T^2} + ET^3$ where $A-E$ are constants for the gas

Reminder: $C_{v,m} = \left(\frac{\partial U_m}{\partial T} \right)_{V_m} = \frac{dU_m(T)}{dT}$
 \uparrow constant volume
 \uparrow generally true
 \uparrow ideal gas only

Similarly,
 $C_{p,m} = \left(\frac{\partial h_m}{\partial T} \right)_P = \frac{dh_m(T)}{dT}$
 \uparrow constant pressure
 \uparrow generally true
 \uparrow ideal gas only

Enthalpy (h_m)

• Mathematically: $h_m = U_m + PV_m$ (general defn)

→ For ideal gas, $PV_m = RT$

$$\therefore h_m = U_m(T) + RT \Rightarrow h_m = h_m(T) \text{ only for an ideal gas}$$

Note that heat capacities are defined for a single phase.

How to Measure $C_{p,m}$

→ closed system, calorimeter, system @ const P

const P → $P_E = P$ (force balance)

1st Law: $dU_m = dq_m - Pdv_m$ ($P_E = P$)

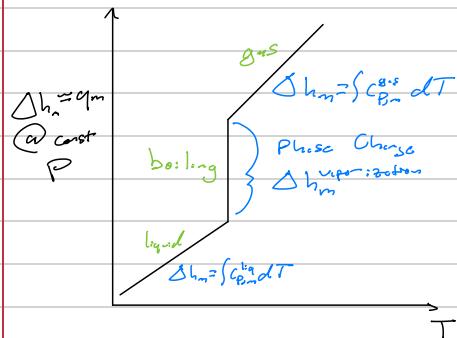
$dU_m = dq_m - d(PV_m)$ can measure more only bc Pressure P is constant

$dU_m + d(PV_m) = dq_m$

$d(h_m + PV_m) = dh_m = C_{p,m}dT = dq_m$ ← measure q_m @ const P

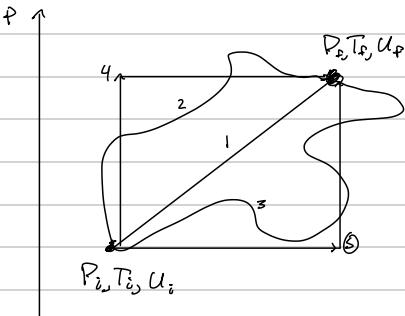
← measure T to determine h_m & $C_{p,m}$

③ Latent Heat \equiv enthalpy (or internal energy) of a phase change



Calculation Paths

Used to calculate changes in thermodynamic properties as these do not depend on path ($T, P, V_m, U_m, h_m, \dots$)

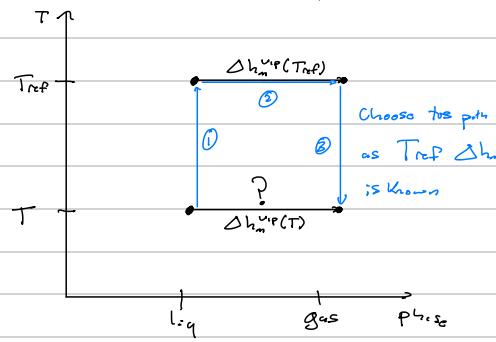


$$\Delta U = U_f - U_i$$

$$= \Delta U_1 = \Delta U_2 = \Delta U_f = \Delta U_i$$

True for any process!

We can use calculation path to find $\Delta h_m^{vp}(T)$



* T_{ref} is usually 298K

Solution:

$$\Delta h_m^{vp}(T) = \Delta h_m^v + \Delta h_m^g + \Delta h_m^s$$

$$\textcircled{1} \text{ Sink phase heating/cooling } \Delta h_{m,1} = \int_{T_{ref}}^{T_f} C_{P,m}^{l,g} dT$$

$$\textcircled{2} \text{ Phase change @ } T_{ref}; \Delta h_{m,2} = \Delta h_m^{vp}(T_{ref})$$

$$\textcircled{3} \text{ Sink phase cooling/heating } \Delta h_{m,3} = \int_{T_{ref}}^T C_{P,m}^{s,s}(T) dT$$

$$\therefore \Delta h_m^{vp}(T) = \Delta h_m^{vp}(T_{ref}) + \int_{T_{ref}}^T [C_{P,m}^{s,s}(T) - C_{P,m}^{l,g}(T)] dT$$

Example 1st Law w/ Property Tables

A rigid container has sat'd water @ 2 bar and quality of $x=0.42$. The water is heated to 540°C. Find the final pressure and the $\Delta \hat{U}, \hat{q}$, and \dot{m} for the process.

| initial | final |
|--|---|
| $\text{Sat'd } H_2O$ $x = 0.42$ $P = 2 \text{ bar}$ V_i | $T_f = 540^\circ\text{C}$ $\xrightarrow{\text{Process: Closed System, Constant volume}}$ $V_f = V_i (\text{rigid})$ |

- initial state is sat'd water, $x = 0.42$
- rigid $\rightarrow V_f = V_i$, and $\dot{V}_f = \dot{V}_i = 0$
- closed system $\rightarrow m_i = m_f$ From mass balance
- water is heated $\rightarrow q \neq 0$
- $\dot{W}_s = 0$ no moving parts or mixing
- $\dot{W} = - \int_{V_i}^{V_f} P_B dV = 0$ (no volume change)
- $\rightarrow KE, PE / mg$ small and negligible

$$1^{\text{st}} \text{ Law: } \Delta \hat{U} = \hat{q} + \dot{W}$$

$$\Delta \hat{U}: \text{initial} \rightarrow \text{sat'd water @ } P_{s,t} = 2 \text{ bar (0.2 MPa)}$$

| | |
|--|---|
| $\text{Table B.2: } \begin{cases} \hat{u}_i = 504.47 \frac{\text{kJ}}{\text{kg}} \\ \hat{v}_i = 0.001061 \frac{\text{m}^3}{\text{kg}} \end{cases}$ | $\hat{u}_f = 2529.5 \frac{\text{kJ}}{\text{kg}}$ $\hat{v}_f = 0.8857 \frac{\text{m}^3}{\text{kg}}$ |
|--|---|

$$\text{Recall: } \hat{u}_i = x \hat{u}_w + (1-x) \hat{u}_l$$

orange

$$\hat{v}_i = x \hat{v}_w + (1-x) \hat{v}_l$$

$$\Rightarrow \begin{cases} \hat{u}_i = 1385 \frac{\text{kJ}}{\text{kg}} \\ \hat{v}_i = 0.373 \frac{\text{m}^3}{\text{kg}} \end{cases}$$

$$\text{Final: Known } T_f = 540^\circ\text{C}, V_f = 0.373 \frac{\text{m}^3}{\text{kg}} \text{ (rigid)}$$

Table B.2 shows $T_f > T_c \rightarrow$ superheated steam

- Interpolate super heated water vapor tables (B4)
- $\rightarrow P_s = 10 \text{ bar (1 MPa)}$

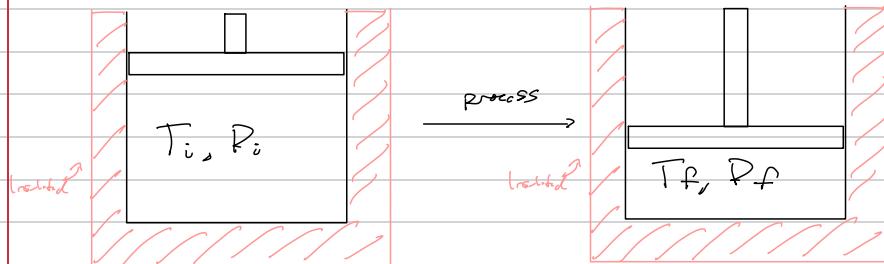
$$\hat{u}_f = 3194 \frac{\text{kJ}}{\text{kg}}$$

$$1^{\text{st}} \text{ Law: } \Delta \hat{U} = \hat{q}$$

$$\Delta \hat{U} = \hat{u}_f - \hat{u}_i = 1839 \frac{\text{kJ}}{\text{kg}}$$

| |
|--|
| $P_f = 10 \text{ bar}$ $\Delta \hat{U} = \hat{q} = 1839 \frac{\text{kJ}}{\text{kg}}$ $\dot{m} = 0$ |
|--|

Example: Reversible adiabatic expansion/compression of an ideal gas with $C_{v,m} = \text{constant}$.
Relate P_i, P_f, V_i, V_f



- closed, mass/mol fixed: $nF = n$:
- 1st Law: $dU = \delta Q + \delta W$
- $\delta Q = 0$ (adiabatic)
- $\delta W = -P_{ext}dV$ (definition of P_{ext})
- Ideal gas: $nC_{v,m}dT$
- $-P_{ext}dV = nC_{v,m}dT$
- $P_B = P$ (Reversible)

$$\rightarrow nC_{v,m}dT = -P_B dV$$

$$nC_{v,m}dT = -P dV$$

$$\cancel{nC_{v,m}dT} = -\cancel{\frac{RT}{V}} dV$$

$$\int_{T_0}^{T_f} \frac{C_{v,m}}{T} dT = \int_{V_i}^{V_f} -\frac{R}{V} dV$$

$$\text{If } C_{v,m} = \text{constant} \Rightarrow C_{v,m} \ln\left(\frac{T_f}{T_i}\right) = -R \ln\left(\frac{V_f}{V_i}\right)$$

$$\text{take exp: } \left(\frac{T_f}{T_i}\right)^{C_{v,m}} = \left(\frac{V_f}{V_i}\right)^{-R}$$

$$\Rightarrow C_{p,m} = C_{v,m} + R \quad (\text{notes to follow})$$

$$P_f V_f^{\frac{C_{p,m}}{C_{v,m}}} = P_i V_i^{\frac{C_{p,m}}{C_{v,m}}}$$

ideal gas, constant $C_{v,m}$, reversible
and adiabatic

Final Notes on Ideal Gas Heat Capacities

$$\Delta h_m = \int_{h_{m,1}}^{h_{m,2}} dh_m = \int_{T_1}^{T_2} C_{p,m} dT \quad (\text{always true for ideal gas})$$

$$\text{Because } h_m \equiv U_m + PV_m = U_m + RT$$

$$\frac{dh_m}{dT} = \frac{du_m}{dT} + R$$

$$\uparrow \qquad \uparrow$$

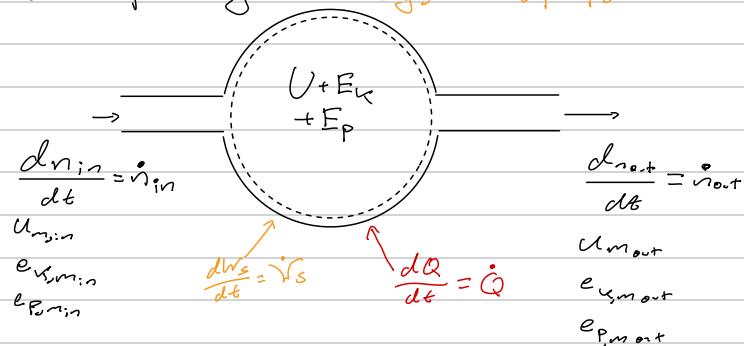
$$C_{p,m} = C_{v,m} + R$$

↑ from u_m ↑ from PV_m

from T only from T only

(for ideal gas)

Skip §2.8.2.1 1st Law Open Systems e.g. turbines, pumps, reactors...



* Important Calculus Note:

$$\dot{\nu}_{in} = \frac{d\dot{m}_{in}}{dt} \neq \dot{m}_{in}$$

discrete ↑ differential ↑

Form of PV Work?

→ Flow Work

Rate of flow work:

$$\frac{dW_{flow}}{dt} = \dot{W}_{flow} = -P_E \frac{dV_{in}}{dt}$$

$$= -P_{in} \frac{d(V_{in} m_{in})}{dt} = -P_{in} \frac{d(m_{in} V_{in})}{dt}$$

$$= P_{in} \frac{d(m_{in} V_{in})}{dt} = P_{in} V_{in} \frac{dm_{in}}{dt} = P_{in} V_{in} \dot{m}_{in} = \dot{W}_{flow} \because \text{molar rate} = \frac{dm}{dt} = (P_{in} V_{in})_{in}$$

= flow work on sys

→ $P > 0, V_m > 0$, so flow work $\dot{W}_{flow} > 0$ always

⇒ Add flow work to energy balance

$$\frac{d}{dt} (U + E_k + E_p) = \dot{Q} + \dot{W}_s + \dot{\nu}_{in} (U_{in} + e_{kin,in} + e_{pot,in}) + \dot{\nu}_{in} (P_{in} V_{in})_{in}$$

accumulation

$$- \dot{\nu}_{out} (U_{out} + e_{kin,out} + e_{pot,out}) - \dot{\nu}_{out} (P_{out} V_{out})_{out}$$

$$\text{Real: } h_m \equiv U_m + P_m$$

$$\frac{d}{dt} (U + E_k + E_p) = \dot{Q} + \dot{W}_s + \dot{\nu}_{in} (h_m + e_{kin,in} + e_{pot,in}) - \dot{\nu}_{out} (h_m + e_{kin,out} + e_{pot,out})$$

First Law, Open Systems

Can also write on a mass basis:

IF open system is at steady-state

- no mass/moles accumulating in sys
- mass balance, $\frac{d\dot{m}}{dt} = 0 = \dot{m}_{in} - \dot{m}_{out} \Rightarrow \dot{m}_{in} = \dot{m}_{out} = \dot{m}$

- no change in energy of system

$$\frac{d}{dt} (U + E_k + E_p) = 0$$

- no change in properties with time

IF NOT steady-state

✓ mole accumulation

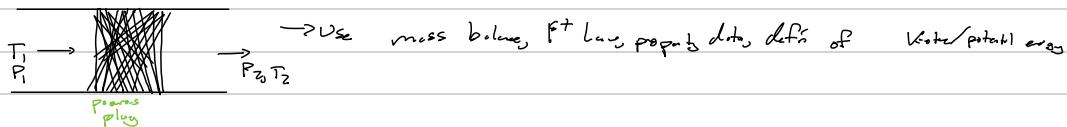
$$\rightarrow \text{mole balance: } \frac{d\dot{m}}{dt} = \dot{m}_{in} - \dot{m}_{out}$$

- Total mass/mole accumulation:

$$\Delta n = n_{fin} - n_{ini} = \int_{n_{ini}}^{n_{fin}} dn$$

Example 6: Throttling Process

- Simple obstruction to reduce pressure in a pipe (constant-insulated)



Steady state: $\frac{d}{dt}(u + E_k + E_p) = 0; \frac{du}{dt} = 0 \quad \dot{m}_{in} = \dot{m}_{out} = \dot{m}$

$\dot{m}_s = 0$ (no moving parts)

$\dot{Q} = 0$ (well-insulated)

$e_{kin} \approx 0$ (smaller than u^2) ← good approximation if velocity is much lower than the speed of sound

$e_{p,m} = 0$ (no height difference)

does change but the change is negligible with respect to other quantities

→ Simplified F+ Law

$0 = \dot{m}_{in} h_{m,in} - \dot{m}_{out} h_{m,out}$

so

$$\begin{cases} 0 = h_{m,in} - h_{m,out} \\ 0 = \hat{h}_{in} - \hat{h}_{out} \end{cases}$$

Mole Balance ($\dot{m}_{in} = \dot{m}_{out}$)

mole basis } independent of fluid
mass basis } of fluid

a) Let's assume fluid is water:

$T_1 = 350^\circ\text{C}$

$P_1 = 10 \text{ MPa}$

$\rightarrow T_2 ?$
 $P_2 = 1 \text{ bar}$

Find T_2 !

$\hat{h}_1 = \hat{h}(350^\circ\text{C}, 10 \text{ MPa}) = 2923.4 \frac{\text{kJ}}{\text{kg}}$
by law $\hat{h}_2 = \hat{h}_1$ since $\Delta \hat{h} = 0$

Find T when $\hat{h}_2 = 2923.4 \frac{\text{kJ}}{\text{kg}}$ at $P_2 = 1 \text{ bar} = 0.1 \text{ MPa}$

\rightarrow At 1 bar, $\hat{h}(200^\circ\text{C}) = 2875.3$
 $\hat{h}(250^\circ\text{C}) = 2974.3$

$\left(\frac{2974.3 - 2875.3}{250 - 200} \right) (T_2 - 200) + 2875.3 = 2923.4$

$\rightarrow \boxed{T_2 = 224.293^\circ\text{C}}$

b) What will happen if fluid is ideal gas?

1) $T_2 < T_1$

3) $T_2 = T_1$

2) $T_2 > T_1$

4) real more $\Rightarrow P$

Why?

\rightarrow Still: $h_{m,2} - h_{m,1} = 0$ (same first law)

$\Delta h_m = 0$

$\Delta h_m = \int_{T_1}^{T_2} dh_m = 0$ (Change in properties is constant of diffn.)

\rightarrow Ideal G.s.: $h_m = h_m(T)$ only

$dh_m = C_{pm} dT$ For ideal g.s.

$\Delta h_m = \int_{T_1}^{T_2} C_{pm} dT = 0$

$\rightarrow C_{pm}(T_2 - T_1) = 0 \Rightarrow T_2 = T_1 = 350^\circ\text{C}$ also true if $C_{pm} = C_{pm}(T)$ [i.e., not constant]

Note: Compare H_2O + ideal gas case of throttle

- ideal gas: $h_m = h_m(T)$ only
 $dh_m = C_p dT$

- water (real fluid): $\hat{h} \neq \hat{h}(T)$ only

$$S_o \quad \hat{h}_2 = \hat{h}_1 \quad \text{but } T_2 \neq T_1$$

Q: What is the significance of enthalpy \hat{h} , the first law for open systems?

1) A more fundamental molecular quantity for flow systems

2) Sum of internal energy + flow work between system & surroundings

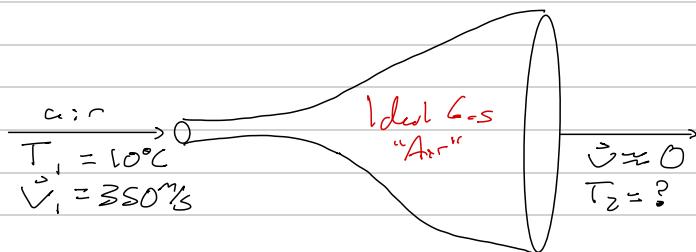
3) Flows in open systems have enthalpy \hat{h} , but not internal energy. Of course \hat{h} has internal energy!

4) Amount of heat transferred at constant pressure

5) All of the above will do! closed systems

Ex: Nozzle or Diffuser

→ These relates to accelerate (or slow) flow streams



Process occurs @ constant pressure, $P_2 = P_1$

• Open system

• Steady state? yes!

$$\rightarrow \text{MB: } \frac{dm}{dt} = 0 = m_{in} - m_{out} \rightarrow m_{in} = m_{out}$$

• 1st Law, accumulations = 0:

$$\frac{d}{dt}(U + E_k + E_p) = 0$$

• $\dot{W}_s = 0$ no moving parts, assume adiabatic $\dot{Q} = 0$

$$\cdot e_{p,m,in} \approx e_{p,m,out} \approx 0$$

$$\cdot e_{k,m,out} = \frac{1}{2} m v^2 = \frac{1}{2} m (0)^2 = 0, e_{k,m,in} \approx 0$$

$$\cdot \text{Simpl. form 1st Law: } \hat{h}_2 = \hat{h}_1 + \hat{e}_{k,1} \quad (\text{Solving on basis of basis is bc of kinetic energy})$$

$$\rightarrow \hat{e}_{k,1} = \frac{1}{2} m | \vec{v} |^2 \\ \hat{e}_{k,1} = \frac{1}{2} | \vec{v} |^2 = \frac{1}{2} (350 \frac{\text{m}}{\text{s}})^2 \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 61.25 \frac{\text{kJ}}{\text{kg}}$$

$$1^{\text{st}} \text{ Law} \rightarrow \hat{h}_2 - \hat{h}_1 = \hat{e}_{k,1}$$

$$\Delta \hat{h} = \int_{in}^{out} d\hat{h}_m = \int_{in}^{out} \hat{C}_p dT = \int_{T_1}^{T_2} \hat{C}_p dT$$

ideal g.s.

"Air": assuming ideal gas

$$\rightarrow \text{From Koretz } K_{20} \hat{C}_{p,m} \approx 3.355 \text{ kJ/kgK}$$

$$\text{constant } C_p \quad \frac{K}{2.053} \hat{C}_p = \frac{C_p}{M w_{air}} \leftarrow (-28 \times 10^{-3} \text{ kJ/mol})$$

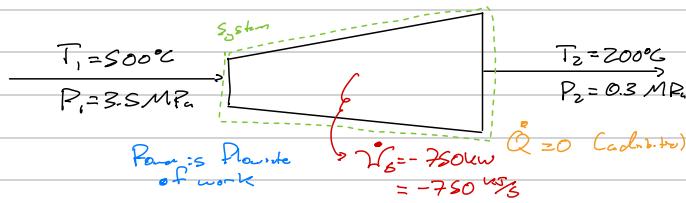
$$\Delta \hat{h} = \int_{T_1}^{T_2} \hat{C}_p dT \approx \hat{C}_p (T_2 - T_1) = \hat{e}_{k,1} \rightarrow T_2 = \frac{\hat{e}_{k,1}}{\hat{C}_p} + T_1 \quad \text{Solve: } T_2 = 73^\circ\text{C}$$

* Note: could be more accurate w/ T-dependent $\hat{C}_p(T)$

: (macroscopic) kinetic energy "contributes" to T (macroscopic)

Example: Turbo extracts work (power) from gas.

Adiabatic turbine: input steam @ $T_1 = 500^\circ\text{C}$, $P_1 = 3.5 \text{ MPa}$ expanded to $T_2 = 200^\circ\text{C}$ and $P_2 = 0.3 \text{ MPa}$. Turbine generates 750 kW of power. What is the required mass flow rate of steam?



- Open system
- Steady-state? Yes!
- Mass balance?

$$\frac{dm}{dt} = 0 = \dot{m}_{in} - \dot{m}_{out} \Leftrightarrow \dot{m}_{in} = \dot{m}_{out} = \dot{m}$$

• 1st Law: $\frac{d}{dt} (U + E_k + E_p) = 0$ (steady-state)

$$\Delta \dot{E}_p \approx 0$$

$$\Delta \dot{E}_k \approx 0$$

1st Law: $\dot{Q} = \dot{m}_{in} h_{in} - \dot{m}_{out} h_{out} + \dot{W}_s$

$$MB: \dot{Q} = \dot{m}(h_{in} - h_{out}) + \dot{W}_s$$

$$\dot{m} = \frac{\dot{W}_s}{h_{out} - h_{in}}$$

steam tables

Supersaturated steam tables (guts h_{in} , h_{out})

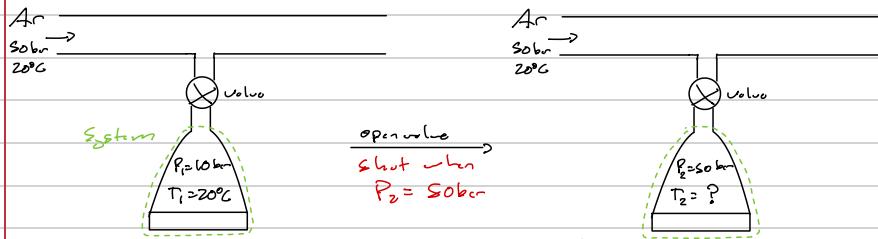
$$\dot{m} = \frac{-750 \text{ kW}}{2865.9 - 3451 \frac{\text{kJ}}{\text{kg}}} = 1.28 \frac{\text{kg}}{\text{s}}$$

How to find V ?

$V = \dot{m} V_{in}$ we like to control inlet conditions
so use V_{in} to convert!

Kortesley Problem 2.47

Argon gas, $C_{p,m} \approx \frac{5}{2}$ ($C_{p,m} = \frac{5}{2} R$). MW = 40 kg/kmol, treat as an ideal gas



a) T_2 right after valve is closed? ($\dot{Q} \approx 0$)

• Open system \rightarrow define system as tank

• Steady-state? No! transient

• Mass balance?

$$\frac{dn}{dt} = \dot{m}_{in} - \dot{m}_{out}$$

gas only flows into tank

• 1st Law open system

$$\frac{dU}{dt} = \dot{m}_{in}(h_{in} + e_{kin} + e_{pot}) - \dot{m}_{out}(h_{out} + e_{kin} + e_{pot})_{out} + \dot{W}_s$$

ignores PE

no mass出入

assumption

$$\frac{dU}{dt} = \dot{m}_{in} h_{min} = \dot{m}_{in} \frac{dn}{dt}$$

Gas flows into tank

$$\int_{U_1}^{U_2} dU = \int_{h_{min}}^{h_{max}} dn$$

Gas insulation

$$\Rightarrow h_{max} = h_m(\text{Ar}, 20^\circ\text{C})$$

$$\rightarrow U_2 - U_1 = h_{max}(n_2 - n_1) \quad \left. \begin{array}{l} \text{mols & } T \text{ change + same } \\ \text{extensive } \Delta U \text{ for sys(tank)} \end{array} \right\} \text{constant } T \text{ for Ar!}$$

$$n_2 U_{max} - n_1 U_{min} = h_{max}(n_2 - n_1)$$

$$\rightarrow n_1 (h_{min} \rightarrow U_{min}) = n_2 (h_{max} - U_{max})$$

defn: $h_m = U_m + P_m$ $\rightarrow h_{min} = U_{min} + P_{min}$

$$h_m = U_m + P T \quad (\text{for ideal gas})$$

$$\rightarrow n_1 [(h_{min} - h_{max}) + RT_1] = n_2 [(h_{max} - h_{min}) + RT_2]$$

ideal gas: $\Delta h_m = \int_{T_1}^{T_2} C_{p,m} dT = 0$

$$\Delta h_m = \int_{T_1}^{T_2} C_{p,m} dT$$

bc $T_1 = T_2$

$$\left. \begin{array}{l} n_1 RT_1 = n_2 RT_2 + n_2 C_{p,m} \int_{T_2}^{T_1} dT \\ n_1 RT_1 = n_2 RT_2 + n_2 C_{p,m} (T_1 - T_2) \end{array} \right\} \begin{array}{l} n_2 T_2 \text{ unknown} \\ (\text{eq, 2 unknowns}) \end{array}$$

but $V_2 = V_1$ (rigid tank)

$$\frac{n_1 RT_2}{P_2} = \frac{n_1 RT_1}{P_1} \rightarrow n_2 = \left(\frac{n_1 T_1}{P_1} \right) \left(\frac{P_2}{T_2} \right)$$

* not given n_1 , but it cancels out

→ sub n_2 in \rightarrow

$$T_2 = \frac{\frac{n_1}{P_1} T_1}{\frac{n_2}{P_2} + \frac{3}{2}} = 187.9^\circ\text{C}$$

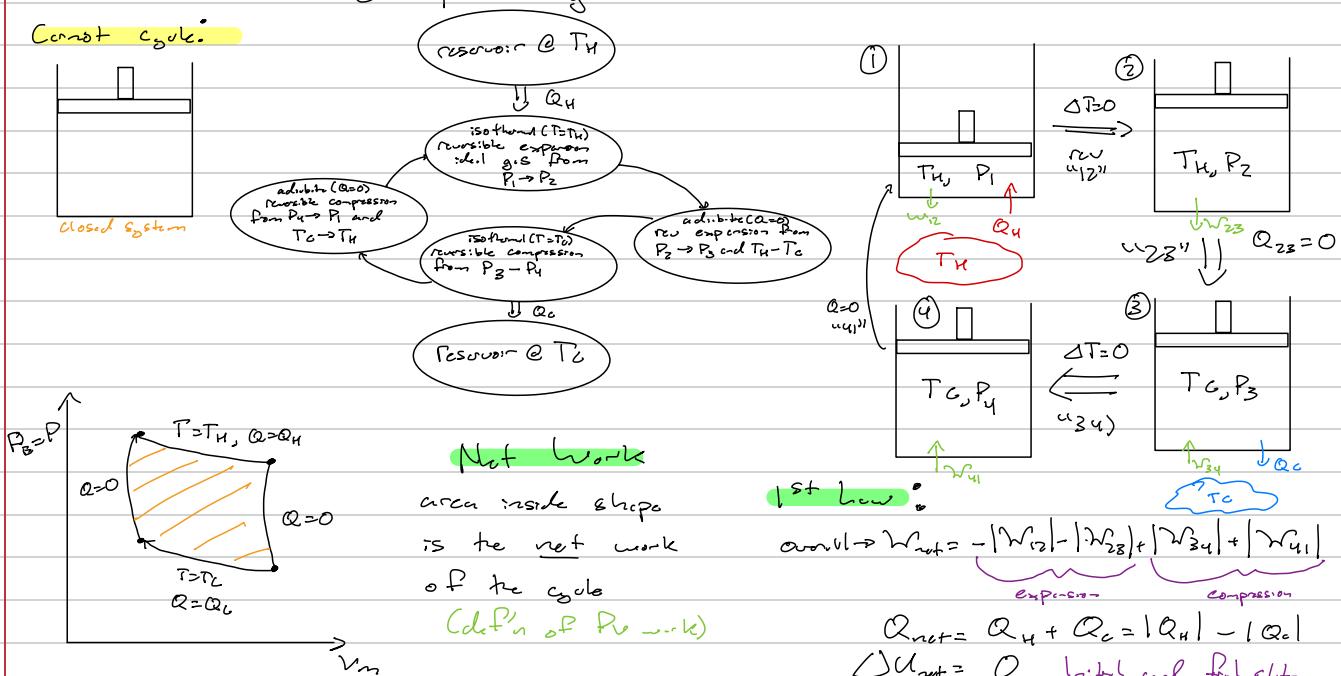
$(C_{p,m} = \frac{5}{2} R)$

Chapter 3 - Entropy and the Second Law

Recall: Any thermodynamic property (T, P, V, s, u, h, \dots) is independent of path
 → If we find a new property (S) that is also independent of path, it must also be a
 thermodynamic property.
 → ΔS for a cycle (same initial/ final state) is zero.

Consider a Carnot cycle (Ch 2 § 2.9): ideal g.s. reversible process
 → Practical use: Converts temperature gradient to useful work

Carnot cycle:



Process 12: Isothermal ($T=T_H$)

$$\rightarrow 1\text{st Law: } \Delta U_{12} = Q_{12} + W_{12} \quad W_{12} = -Q_{12}$$

$$\Delta U_{12} = n \int_{P_1}^{P_2} C_v dT \Rightarrow T_1 = T_2 = T_H \text{ so } \Delta U_{12} = 0$$

$$\begin{cases} W_{12} = nRT_H \ln \left(\frac{P_2}{P_1} \right) < 0 & \text{expansion} \\ P_2 < P_1 \end{cases}$$

$$\begin{aligned} W_{12} &= - \int P_2 dV = -n \int P_2 dV \\ \text{reverse} &= -n \int P_2 dV \\ &= -n \int \frac{RT_H}{V} dV \\ &= -n \int_{V_{12}}^{V_{21}} \frac{RT_H}{V} dV \\ &\rightarrow W_{12} = -nRT_H \ln \left(\frac{V_{12}}{V_{21}} \right) \end{aligned}$$

Process 23: Adiabatic ($Q_{23}=0$)

adiabatic, reversible, constant C_p, m

$$\rightarrow PV^k = \text{constant}$$

$$W_{23} = - \int P dV = - \int \frac{P_2 V_2^k}{V^k} dV \quad \text{Too many unknowns:}$$

→ solve P_3 and R_1 in terms of P_1, P_2 & P_3

"23" is adiabatic $\rightarrow P_2 V_2^k = P_3 V_3^k$

"41" is adiabatic $\rightarrow P_4 V_4^k = P_1 V_1^k$

Substitute

$$\frac{P_3}{P_2} = \frac{P_4}{P_1}$$

To Summary

$$W_{12} = nRT_H \ln \left(\frac{P_2}{P_1} \right), \quad Q_{12} = Q_H = -nRT_H \ln \left(\frac{P_2}{P_1} \right)$$

* Ideal g.s. constant C_p, m , reversible

$$W_{23} = \frac{nR}{k-1} (T_H - T_C), \quad Q_{23} = 0$$

$$W_{34} = -nRT_C \ln \left(\frac{P_2}{P_1} \right), \quad Q_{34} = Q_C = -nRT_C \ln \left(\frac{P_2}{P_1} \right)$$

$$W_{41} = \frac{nR}{k-1} (T_C - T_H), \quad Q_{41} = 0$$

Discuss Overall Cycle

$$\textcircled{1} \quad \eta = \text{Carnot efficiency} = \frac{W_{\text{out}}}{Q_H} = \frac{\text{work out}}{Q_H} = \frac{T_H - T_C}{T_H} \quad (\text{all rev})$$

$\rightarrow \eta$ is never 1. Heat is never converted entirely to work, so not all processes are reversible!

$$\textcircled{2} \quad \text{Consider: } \sum_{\text{cycle}} \frac{Q}{T} = \frac{Q_{12}}{T_H} + \frac{Q_{23}}{T} + \frac{Q_{34}}{T_C} + \frac{Q_{41}}{T}$$

$$= \frac{Q_H}{T_H} + \frac{Q_C}{T_C} = -nR \ln\left(\frac{P_2}{P_1}\right) + nR \ln\left(\frac{P_1}{P_2}\right) \quad \text{Thus, } \sum_{\text{cycle}} \frac{Q}{T} = 0 \quad \text{for a full cycle}$$

idea: $\frac{Q}{T}$ is a thermodynamic property! Call this entropy (S).

$$\Delta S = \frac{Q_{\text{rev}}}{T} \quad \text{or} \quad dS = \frac{dQ_{\text{rev}}}{T}$$

Would calculate a state function (S) from path function (Q).

Entropy

\rightarrow words associated with entropy?

- randomness
- disorder
- chaos
- number of states

Related to molecular probability and statistics

\rightarrow associated w/ # of molecular configurations a system can adopt

Q: Arrange the following : terms of the magnitude of their entropy, S_m .

- 1) solid \Rightarrow liquid \Rightarrow gas
- 2) $\text{gas} \Rightarrow \text{liquid} \Rightarrow \text{solid}$
- 3) solid \Rightarrow liquid \Rightarrow gas
- 4) need more info

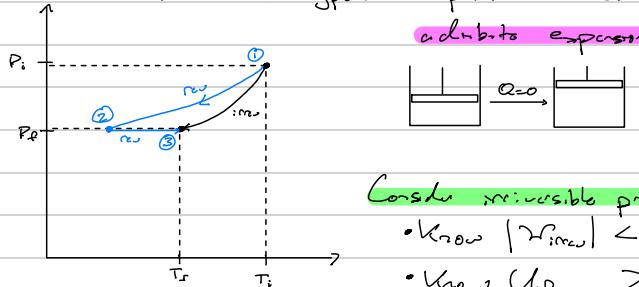
Q: Which process has the most positive change of the fluid P ΔS_m ?

- 1) evaporate $\text{H}_2\text{O(l)} \rightarrow \text{gas}$ $\Delta S_m > 0$
- 2) condense $\text{H}_2\text{O(g)} \rightarrow \text{liquid}$ $\Delta S_m < 0$
- 3) cool $\text{H}_2\text{O(g)}$ by 100°C $\Delta S_m < 0$
- 4) heat $\text{H}_2\text{O(l)}$ by 100°C $\Delta S_m > 0$

Carnot Cycle $\rightarrow \Delta S_{\text{cycle}} = 0$ for reversible processes

\rightarrow How about calculating ΔS for irreversible processes

- Construct a hypothetical path consisting of only reversible processes



Process '12'

$$\begin{aligned} &\text{reversible adiabatic expansion} \\ &1^{\text{st Law: }} \Delta U_{12} = Q_{12, \text{rev}} + W_{12, \text{rev}} \\ &\Delta S_{12} = \int \frac{dQ_{12, \text{rev}}}{T} = 0 \end{aligned}$$

Consider irreversible process (1 \rightarrow 3)

- Know $|W_{12, \text{rev}}| < |W_{12, \text{irr}}|$ For expansion (H_2O)
- Know $U_{12, \text{rev}} > U_{12, \text{irr}}$
- Since S is a state function,

$$\Delta S_{13} = \Delta S_{12} + \Delta S_{23} \quad \begin{matrix} \text{(defining this to} \\ \text{be a reversible} \\ \text{process at constant} \\ \text{pressure)} \end{matrix}$$

(not adiabatic)

$$\Delta S_{23} = \int_{T_2}^{T_3} \frac{dQ_{23, \text{irr}}}{T}$$

$$1^{\text{st Law: }} dU = dQ + PdV$$

$$dQ = dU + PdV$$

$$dQ = dU + PdV$$

$$dQ = dU + d(PV)$$

$$dQ = dU + dH$$

$$dQ = dH$$

Specific to adiabatic expansion: Increase in entropy S for system for irreversible processes. $\Delta S_{13} > 0$

How to get from
SR to others

but!

SR to others

const P
(adiabatic)

$$\Rightarrow \Delta S_{23} = \int_{T_2}^{T_3} \frac{dU_{23}}{T} = \int_{T_2}^{T_3} \frac{C_{p,m}}{T} dT > 0$$

$$(T_3 > T_2)$$

$$(U_3 > U_2)$$

II
reversible & irreversible processes have very different entropy changes!

$$\Delta S_{13} = \Delta S_{12} + \Delta S_{23} > 0 \quad \text{increase in system } S \text{ for irreversible processes}$$

ΔS for Surroundings

$$\delta S_{\text{surroundings}} = \frac{\delta Q_{\text{surv}}}{T_{\text{sur}}} \quad \text{real heat transfer to the surroundings}$$

To calculate δS_{surv} , we can use the real heat transfer to the surroundings (does not have to be reversible)
 why? Surroundings are large \rightarrow constant T_{sur}

$$\delta Q_{\text{surv}} = -\delta Q \quad \begin{matrix} \text{real heat} \\ \text{transfer} \\ \text{not law} \end{matrix} \quad \rightarrow \text{From perspective of surroundings, my change in } S_{\text{surv}} \text{ is tiny} \\ \therefore \text{always seems like reversible } Q_{\text{surv}}$$

In our previous example...

$$\delta S_{\text{surv}} = \frac{\delta Q_{\text{surv}}}{T_{\text{sur}}} = -\frac{\delta Q}{T_{\text{sur}}} = 0 \quad \text{because process is adiabatic}$$

ΔS_{surv} is also 0 for adiabatic processes.

$$\Delta S_{13} > 0, \Delta S_{\text{surv}} = 0 \quad ; \text{missed}$$

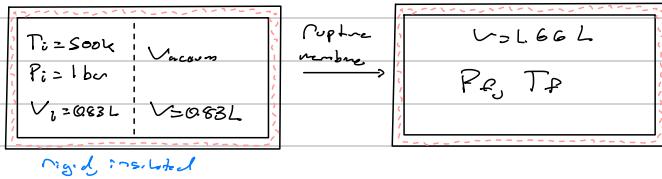
$$\Delta S_{12} = 0, \Delta S_{\text{surv}} = 0 \quad \text{reversible}$$

Conclusion: $\Delta S_{\text{univ}} = \Delta S + \Delta S_{\text{surv}}$

$$\rightarrow \boxed{\Delta S_{\text{univ}} > 0 : \text{missed}} \quad \text{True for } \xrightarrow{\text{process}}$$

↑ Second Law of Thermodynamics

Example: Ideal gas Closed Sys



Final T_f

closed system: $n_i = n_f = n$

• rupture \rightarrow not reversible due to sudden expansion

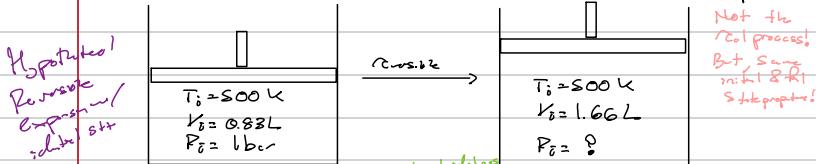
• 1st Law: $\Delta U = Q + W \xrightarrow{\text{adiabatic}} \text{no } Q, \text{ no } W \rightarrow \Delta U = 0$ (constant volume system)

$$\Delta U_{\text{int}} = 0 = \int_{T_i}^{T_f} C_v dT \rightarrow T_i = T_f = 500 \text{ K}$$

What about ΔS ?

→ use a convenient reversible calculation path! (ΔS is a state function / thermodynamic property)

→ must define a reversible calculation path, $\delta S = \frac{\delta Q_{\text{rev}}}{T}$



$$\Rightarrow \Delta S_{\text{calc}} = \Delta S_{\text{real process}}$$

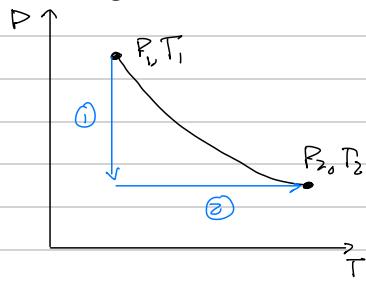
$$\Delta S = \int \frac{\delta Q_{\text{rev}}}{T} = \int \frac{dU_{\text{real}} - dW_{\text{real}}}{T} = - \int \frac{dW_{\text{real}}}{T} = - \int \frac{-P_{\text{real}} dV}{T} = \int \frac{P_{\text{real}} dV}{T} \xrightarrow{\substack{\text{isothermal} \\ \text{expansion}}} = n \int_{V_i}^{V_f} \frac{P}{T} dV = n R \ln \left(\frac{V_f}{V_i} \right)$$

$$\rightarrow \Delta S_m = \frac{\Delta S}{n} = R \ln \left(\frac{V_f}{V_i} \right) = +5.76 \text{ J/mole K}$$

$$\text{Surroundings: } \delta S_{\text{surv}} = \frac{\delta Q_{\text{surv}}}{T_{\text{sur}}} = -\frac{\delta Q_{\text{real}}}{T_{\text{sur}}} = 0 \quad \text{adiabatic}$$

$$\therefore \Delta S_{\text{univ}} = \Delta S_m + \Delta S_{\text{surv}} = +5.76 \text{ J/mole K} > 0 \quad \checkmark \text{ consistent w/ second law}$$

Entropy changes for ideal G.S.



"T cycle": hypothetical, reversible paths to get $\Delta S_{m,sys}$

Process ①: reversible, isothermal (ideal g.s.)

$$\Delta S_1 = \int \frac{S Q_{rev}}{T} = \int_{\text{const } P} \frac{dU_{rev} - SW_{rev}}{T} = \int \frac{P}{T} dV \quad (\text{lost } c_v)$$

$$\therefore \Delta S_1 = nR \ln\left(\frac{V_2}{V_1}\right) = -nR \ln\left(\frac{P_2}{P_1}\right)$$

Process ②: reversible, const P ($c_v + c_p$)

$$S Q_{rev} = dU_{rev} + P dV = d(C_v + PV) = dH$$

$$\Delta S_2 = \int \frac{S Q_{rev}}{T} = \int \frac{dH}{T} = n \int \frac{C_p dT}{T}$$

$$\text{If } C_{p,m} \text{ is const: } \Delta S_2 = n C_{p,m} \ln\left(\frac{T_2}{T_1}\right)$$

$$\Rightarrow \Delta S = n C_{p,m} \ln\left(\frac{T_2}{T_1}\right) - n R \ln\left(\frac{P_2}{P_1}\right) \quad \text{if } C_{p,m} \text{ is constant}$$

$$\text{general: } dS_m = \frac{C_{p,m}}{T} dT - \frac{R}{P} dP$$

ideal gas
entropy change

Q: For which quantities is it possible to calculate changes using an arbitrary hypothetical calculation path?

- 1) Q ← path function, property
- 2) U } Thermodynamic property
- 3) S }
- 4) W ← path function

Q: How do we calculate ΔS_{univ} for any process?

- ① ΔS_m from reversible, hypothetical path; $\Delta S_{m,sum}$ from real heat → See previous example
- ② ΔS_m from real heat; $\Delta S_{m,sum}$ from real heat
- ③ ΔS_m from reversible path; $\Delta S_{m,sum}$ from real heat
- ④ ΔS_m from real heat; $\Delta S_{m,sum}$ from real heat

Example: 1 mol liquid H₂O @ P=1 bar. T increases from 28°C to 50°C. Calculate ΔS , ΔS_{sum}

ΔS_{univ} at $T_{sum} = 700^\circ\text{C}$ and 100°C.

• Develop an expression for ΔS

→ system: rev calc path (constant P)

$$\Delta S = \int dS = n \int \frac{S Q_{rev}}{T} = n \int \frac{dH}{T} = n \int_{T_i}^{T_f} \frac{C_{p,m}}{T} dT$$

(lost)
const P

$$C_{p,m} = 25.4 \frac{\text{J}}{\text{molK}} \approx \text{constant} \rightarrow \Delta S = n C_{p,m} \ln\left(\frac{T_f}{T_i}\right) = 6.07 \frac{\text{J}}{\text{K}}$$

• ΔS_{sum} : use real heat transfer

$$\Delta S_{sum} = \int \frac{S Q_{real,sum}}{T_{sum}} = \frac{Q_{real,sum}}{T_{sum}} = -\frac{Q}{T_{sum}}$$

$$\rightarrow Q = \int S Q = \int dH = n \int_{T_i}^{T_f} C_{p,m} dT = (1\text{ mol})(25.4 \frac{\text{J}}{\text{molK}})(25\text{ K}) \quad [=] \text{ J}$$

$$\therefore T_{sum} = 700^\circ\text{C} \therefore \Delta S_{sum} = -\frac{25.4 \times 25}{700 + 273} = -1.44 \frac{\text{J}}{\text{K}}$$

$$T_{sum} = 100^\circ\text{C} \therefore \Delta S_{sum} = -\frac{25.4 \times 25}{100 + 273} = -5.05 \frac{\text{J}}{\text{K}}$$

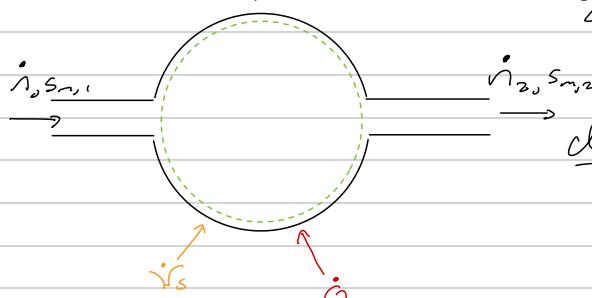
$$T_{sum} \mid \Delta S_{univ} = \Delta S + \Delta S_{sum}$$

$$700^\circ\text{C} \quad 3.87 \frac{\text{J}}{\text{K}}$$

$$100^\circ\text{C} \quad 1.02 \frac{\text{J}}{\text{K}}$$

Both ΔS_{univ} are positive, so it is thermodynamically allowed!

Second Law Open Systems



$$2^{\text{nd}} \text{ Law: } \frac{dS_{\text{univ}}}{dt} = \frac{dS_{\text{sys}}}{dt} + \frac{dS_{\text{env}}}{dt} \geq 0$$

accumulation = 0 : f. at steady state

(pol heat transfer)

$$\frac{dS_{\text{univ}}}{dt} = i_2 S_{m,2} - i_1 S_{m,1} + \frac{\dot{Q}_{\text{univ}}}{T_{\text{univ}}} \quad \text{Also, } \dot{Q}_{\text{univ}} = -\dot{Q}$$

$$2^{\text{nd}} \text{ Law: } \frac{dS_{\text{univ}}}{dt} = \frac{dS_{\text{sys}}}{dt} + i_2 S_{m,2} - i_1 S_{m,1} - \frac{\dot{Q}}{T_{\text{univ}}} \geq 0$$

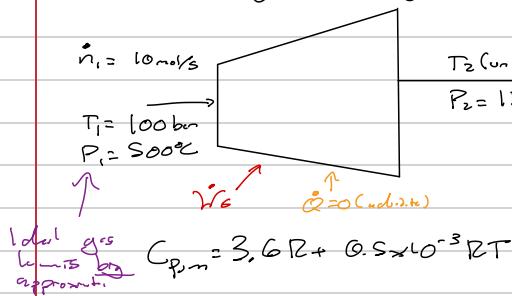
\hookrightarrow steady \Rightarrow static $\rightarrow \frac{dS_{\text{sys}}}{dt} = 0 \Rightarrow \frac{dS_{\text{univ}}}{dt} = i_2 S_{m,2} - i_1 S_{m,1} - \frac{\dot{Q}}{T_{\text{univ}}}$

What is Second Law useful?

- ① Another state function, and equation/inequality
- ② Consistency check for calculations, $\Delta S_{\text{univ}} = 0$
- ③ Limiting cases for evaluating maximum/minimum amount of work (reversibility)
- ④ Can be used for equilibrium analysis (will be done later)

Next example

Example: Ideal gas entering adiabatic turbine



ideal gas law approximation: $C_{p,m} = 3.6 R + 0.5 \times 10^{-3} RT$

Want to know:

a) Maximum amount of power that can be generated?
→ This is for a reversible process

recall: $P = P_2$ for closed system

- infinitesimal changes
- max work for expansion,
- min work for compression

For open systems: $\Delta S_{\text{univ}} = 0$

mole balance: $\frac{dn}{dt} = 0 = \dot{n}_{\text{in}} - \dot{n}_{\text{out}}$ (steady-state)

no accumulation $\rightarrow \dot{i}_1 = \dot{i}_2 = i$

1st Law: $0 = \dot{i}_1 h_{m,1} - \dot{i}_2 h_{m,2} + \dot{W}_S + \dot{Q}$ $\xrightarrow{\text{adiabatic}} \dot{W}_S = \dot{i}(h_{m,2} - h_{m,1})$

$\dot{W}_S = \dot{i} \int_{T_1}^{T_2} C_{p,m} dT$ $\xrightarrow[\text{Lat.}]{} \text{temperature dependent}$ Be careful! don't skip steps

\rightarrow Eqn (1st Law), 2 unknowns (\dot{W}_S, T_2)

Need more equations $\rightarrow 2^{\text{nd}} \text{ Law}$!!

2nd Law to get T_2 :

$$\frac{dS_{\text{univ}}}{dt} = 0 = \frac{dS_{\text{sys}}}{dt} + \dot{i}(S_{m,2} - S_{m,1}) - \frac{\dot{Q}}{T_{\text{univ}}} \rightarrow S_{m,2} - S_{m,1} = 0 \quad (\text{isentropic process})$$

(conserv.) (steady state) note: not static (adiabatic)

$$\Delta S_{\text{univ}} = \int_{i_1}^{i_2} dS_m = \int_{T_1}^{T_2} \frac{C_{p,m}}{T} dT - \int_{P_1}^{P_2} \frac{R}{P} dP \rightarrow 0 = \int_{T_1}^{T_2} \left(\frac{3.6}{T} + 0.5 \times 10^{-3} \right) dT - R \ln \frac{P_2}{P_1}$$

prop. of $\int dS_m = \frac{C_{p,m}}{T} dT - \frac{R}{P} dP$

ideal gas \rightarrow (ideal gas assumption only)

\rightarrow Back to 1st Law

$$\dot{W}_{\text{turb}} = \dot{i}_1 \int_{T_1=273K}^{T_2=282K} (3.6 + 0.5 \times 10^{-3} T) dT = -173.2 \frac{\text{kJ}}{\text{s}} \xrightarrow{\text{max power}} \dot{W}_{\text{turb}} = -173.2 \text{ kW}$$

LO so work done on surroundings

Example Continued

b) The efficiency is *80%, what is $\dot{W}_{s,\text{real}}$?

$$\dot{W}_{s,\text{real}} = 0.8 \cdot \dot{W}_{s,\text{rev}} = -138.6 \frac{\text{kJ}}{\text{s}}$$

denote \dot{W}_s

EPP

$$\eta = \begin{cases} \dot{W}_{s,\text{rev}} / \dot{W}_{s,\text{real}} & \text{for expansion} \\ \dot{W}_{s,\text{rev}} / \dot{W}_{s,\text{real}} & \text{for compression} \end{cases}$$

* This efficiency is the isentropic efficiency, not actual efficiency

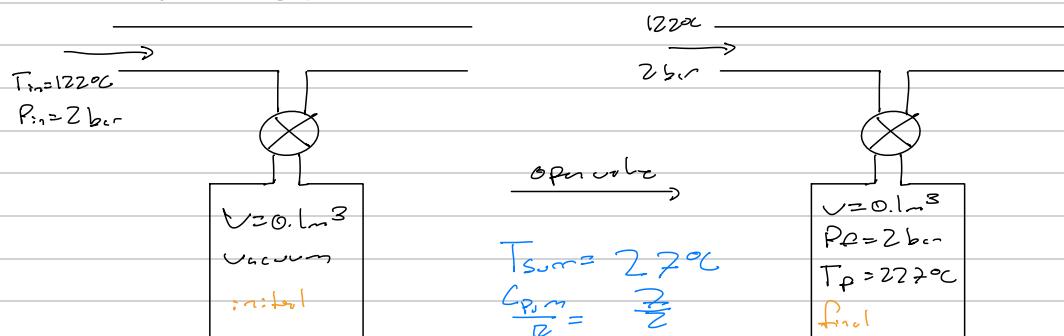
c) What is T_2 ? (80% efficiency)

$$\text{1st Law: } \dot{W}_s = 10 \int_{T_1}^{T_2} C_p dT$$

Know \dot{W}_s , $T_1 \rightarrow$ Solve to find $T_2 = 343 \text{ K}$

Example: Filling a tank w/ ideal gas, 6000J of heat are lost from system.

Find dS_{univ}



• Open system

$$\text{• mole balance: } \frac{dn}{dt} = \dot{n}_{in} - \dot{n}_{out} \Rightarrow \frac{dn}{dt} = \dot{n}_{in} = \frac{dn_{in}}{dt} \quad (\text{unsteady state})$$

$$\Rightarrow dn = dn_{in} \Rightarrow \int dn = \int dn_{in} \Rightarrow n_f = n_i$$

• 1st Law: skip (do not do this :)) \dot{Q} is not needed here given $\dot{Q} \& T_f$

$$\text{• 2nd Law: } \frac{dS_{\text{univ}}}{dt} = \frac{ds}{dt} + \dot{n}_{out} s_{out} - \dot{n}_{in} s_{in} - \frac{\dot{Q}}{T_{\text{sum}}} = \frac{ds}{dt} - \dot{n}_{in} s_{in} - \frac{\dot{Q}}{T_{\text{sum}}} \leftarrow \frac{dQ}{dt}$$

$$\Rightarrow \frac{dS_{\text{univ}}}{dt} = \frac{ds}{dt} + \frac{dn}{dt} s_{in} - \frac{\dot{Q}}{T_{\text{sum}}}$$

$$\text{different form: } dS_{\text{univ}} = ds + d_{n,i} s_{in} - T_{\text{sum}} \text{ constant} \quad \begin{array}{l} \text{entropy is state property} \\ \text{and int for prop are} \\ \text{constant T \& P} \end{array}$$

$$\text{so, } \int dS_{\text{univ}} = \Delta S_{\text{univ}} = \int ds + s_{in} d_{n,i} - \frac{1}{T_{\text{sum}}} \int \dot{Q}$$

$$\Delta S_{\text{univ}} = n_f s_{in} - n_i s_{in} - \frac{\dot{Q}}{T_{\text{sum}}} \\ = n_f (s_{in} - s_{in}) - \frac{\dot{Q}}{T_{\text{sum}}}$$

→ What about $(s_{in,f} - s_{in,i})$

- For ideal gas: integrate ds_m

$$\Delta S_m = s_{in,f} - s_{in,i} = \int ds_m = \int_{T_i}^{T_f} \frac{C_p}{T} dT - \int_{P_i}^{P_f} \frac{R}{P} dP$$

$$\Rightarrow \Delta S_m = \int_{T_i}^{T_f} \frac{C_p}{T} dT = \frac{3}{2} R \ln \left(\frac{T_f}{T_i} \right)$$

$$\Delta S_{\text{univ}} = n_f \Delta S_m - \frac{\dot{Q}}{T_{\text{sum}}}$$

$$\Delta S_{\text{univ}} = \frac{7}{2} R \ln \left(\frac{T_f}{T_i} \right) \left(\frac{P_f V}{R T_f} \right) - \frac{-6000 \text{ J}}{223 + 27} > 0$$

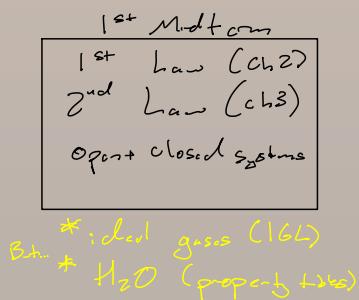
(complaining for remote answer)

✓ consistent w/
Second law

Midterm Exam 2



Previously...



Equations of
State (Ch 4)

Next!

1st & 2nd Law Problems
w/ Real Fluids

Need a mathematical framework to
describe properties of real fluids

Chapter 4 - Equations of State

Equations of state for pure fluids (1 phase)

- Gibbs phase rule: For a pure fluid, need 2 intensive properties to specify all intensive properties

$$\hookrightarrow P = f(T, V_m) \quad \text{2 intensive properties}$$

• Compressibility Factor (Z):

$$Z = \frac{P_{\text{m}}}{RT} = g(T, P) \quad \left. \begin{array}{l} \text{2 intensive props} \\ \hline \end{array} \right\} Z=1 \text{ for ideal gas}$$

\hookrightarrow Tells us how "real" or "non-ideal" a gas is

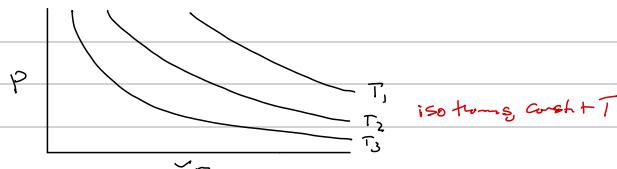
Examples of EOS's

(i) Ideal Gas Law:

$$\rightarrow PV = nRT$$

$$\rightarrow P V_m = RT$$

$$\rightarrow Z = \frac{PV_m}{RT} = 1$$



Assumptions: (a) Molecules do not interact with each other (No potential energy)

\hookrightarrow (b) No volume of molecules

\hookrightarrow We can improve our description of real gases by relaxing these assumptions

(ii) Clausius EOS

$$\rightarrow P = \frac{RT}{V_m - b}$$

where b : parameter for molecular volume ($b > 0$)

Consequence: less volume for the other gas molecules!

\hookrightarrow No intermolecular forces just molecular volume

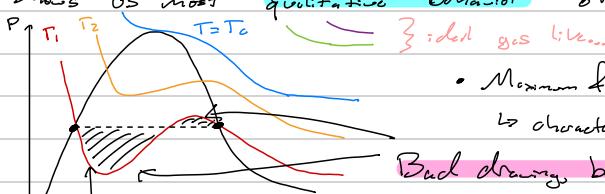
b is a const vol for a fluid, independent of the fluid's T & P .

(iii) Van der Waals EOS

$$\rightarrow P = \frac{RT}{V_m - b} - \frac{a}{V_m^2} \quad \text{where } a: \text{Attractions between molecules: } \text{Forces} \sim r^{-6} \sim V_m^{-2}$$

improving our description of intermolecular forces

• Shows us most qualitative behavior of real fluids



\hookrightarrow Maximum & inflection point @ $T=T_c$

\hookrightarrow characteristic of real fluids

Bad drawing, but these two regions should have equal areas

\hookrightarrow Points that gives "equal area" shows us VL regions

Isotherms do not give accurate pressure inside the VL dome

Observation: $\left(\frac{\partial P}{\partial V_m} \right)_{T=T_c} = 0$ at critical point

Universal Truths: $\left(\frac{\partial^2 P}{\partial V_m^2} \right)_{T=T_c} = 0$ at critical point (inflection point)

For VDW EOS

$$\left(\frac{\partial P}{\partial V_m} \right)_{T=T_c} = \frac{\partial}{\partial V_m} \left[\frac{RT}{V_m - b} - \frac{a}{V_m^2} \right]_{T=T_c} = -\frac{RT_c}{(V_{m,c}-b)^2} + \frac{2a}{V_{m,c}^3} = 0$$

$$\left(\frac{\partial^2 P}{\partial V_m^2} \right)_{T=T_c} = \dots = 0 \quad \text{not expected here}$$

Set equal and solve to get these as a, b

2 eqns,

2 unknowns (a, b)

\Rightarrow Solve for a, b in terms of $T_c, V_{m,c}, P_c$!!

$$a = \frac{27}{64} \cdot \frac{(RT_c)^2}{R_c} \quad \text{and} \quad b = \frac{V_{m,c}}{3} \cdot \frac{RT_c}{8P_c}$$

If we know $T_c, V_{m,c}, P_c$ we can calculate $a + b$ for VDW EOS

expressions valid only for VDW EOS

$$Z_c = \frac{P_c V_{m,c}}{R T_c} = \frac{3}{8} \quad @ \text{critical point, for any fluid, for VDW EOS}$$

$\hookrightarrow Z_c$ does not depend on $a+b$

∴ All fluids that follow VDW EOS have same deviation from ideality at critical point
True for more complex EOS's too!

Principle of Corresponding States (Hypothesis)

- all fluids have similar behavior (deviation from ideality) at critical point
- \Rightarrow all fluids have similar behavior for the same "reduced conditions"

$$P_r = \frac{P}{P_c}, \quad T_r = \frac{T}{T_c}, \quad V_{m,r} = \frac{V_m}{V_{m,c}}$$

• VDW EOS predicts $Z_c = \frac{3}{8} = 0.375$

• Experimentally, $Z_c \approx 0.29 \rightarrow$ need better EOS for real applications

Back to 'Examples of EOS'

(i) Redlich-Kwong EOS

$$\rightarrow P = \frac{RT}{Vm - b} - \frac{a}{\sqrt{Vm(Vm+b)}} \quad \begin{matrix} a & \leftarrow \text{attraction terms} \\ \text{less attrac. @ low } T \end{matrix}$$

$$\begin{cases} a = \frac{0.42748 P_c^{5/2} T_c}{R} \\ b = \frac{0.08664 R T_c}{P_c} \end{cases} \quad \begin{matrix} \text{Not the same as} \\ a \text{ and } b \text{ in VDW} \\ \text{EOS} \end{matrix}$$

$$\Rightarrow Z_c = \frac{1}{3} \approx 0.33 \quad \text{But for the VDW!}$$

(ii) Peng-Robinson EOS

$$\rightarrow P = \frac{RT}{Vm - b} - \frac{a \alpha(T)}{Vm(Vm+b) - b(Vm-b)}$$

Common in industry

$$\rightarrow Z_c = 0.307 \text{ pretty good!}$$

α is a factor (\propto shape of molecules)

(::)- (✓) acc. w/ cubic EOS's

Q: What is the about corresponding states

(✓) All fluids have same T_c, P_c

(::) Z_c is about the same for all fluids

(::) Properties not the same reduced variables are the same

(✗) None of these

✓ 1 "2" and "3"

Compressibility Charts

$$Z = Z^{(0)} + \omega Z^{(1)} = f(T, P) \quad \begin{matrix} \text{kind of boring & only not really total} \\ \text{in this case} \end{matrix}$$

\uparrow Simple molecules Correction for "non-sphericity"

Liquids and Solids

- Most EoS's are not very accurate for liquids
- Very strong intermolecular interactions; cannot be described by 1 parameter " α "
- **Defn:** $B = \frac{1}{V_m} \left(\frac{\partial V_m}{\partial T} \right)_P$ coefficient of thermal expansion
 $K = -\frac{1}{V_m} \left(\frac{\partial V_m}{\partial P} \right)_T$ coefficient of isothermal compressibility
- In liquids (not solids), B and K are often approximately constant
- Gibbs phase rule: Solid phase plus fluid, $f=2$ independent interior variables

$$\rightarrow V_m = V_m(T, P)$$

$$- \text{Total dF}_P \text{contd: } \frac{\partial V_m}{V_m} = \left(\frac{\partial V_m}{\partial T} \right)_P dT + \left(\frac{\partial V_m}{\partial P} \right)_T dP$$

$$\Rightarrow \frac{\partial V_m}{V_m} = B dT - K dP$$

useful eqn: P, B and K
are approximately constant

Example: Calculate volume occupied by 10 kg of benzene at 50 bar and 60°C using VdW EoS

→ How do we solve w/o given $a+b$?

⇒ can we solve if we have property data: T_c, P_c

algebra

↑ given or look up

Calculate T_r, P_r :

$$\begin{cases} P_r = \frac{P}{P_c} = \frac{50 \text{ bar}}{329 \text{ bar}} = 1.52 \\ T_r = \frac{T}{T_c} = \frac{333 \text{ K}}{425 \text{ K}} = 0.78 \end{cases}$$

D. difficult to tabulate

From $a+b$ expressions, rewrite VdW EoS as:

$$P_r = \frac{8 T_r}{3 V_{m,r} - 1} - \frac{3}{V_{m,r}^2} \quad (\text{a+b dependence cancels out})$$

↓ want to know
Solve numerically

$$V_{m,r} = 0.476$$

$$V_{m,r} = \sqrt{\frac{V_m}{V_{m,c}}} \rightarrow V_m = V_{m,r} \cdot V_{m,c}$$

$$Z_L = \frac{3}{8} = \frac{P_c V_{m,c}}{R T_c} \longrightarrow V_{m,c} = 0.000208 \frac{\text{m}^3}{\text{mol}}$$

$$S_o, V_m = 4.9 \times 10^{-5} \frac{\text{m}^3}{\text{mol}}$$

Convert V_m to V :

$$V = \text{mol} \cdot V_m$$

$$\text{mol} = 10 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{60 \text{ g}} = 0.01666 \text{ mol}$$

$$\checkmark = 0.017 \text{ m}^3$$

Compare w/ ideal gas:

$$V_{\text{ideal}} = 0.017 \text{ m}^3 \quad 5 \text{ times larger volume w/ 1.6L}$$

↑ because we are not account for attractive forces

Chapter 5 - Thermodynamic Web, and Calculus of Thermo

Goal: Make property predictions for real gases. Apply these predictions to first & second law problems.

→ Calculate U_m, h_m, s_m in terms of P, T, V_m, C_{pm}, C_u for real gases
(P, T given or EoS)

In introduce new Thermodynamic Relationships

First 3 Fundamental equations of thermodynamics (FE)

- Recall for closed system, 1st Law:

$$dU_m = \delta q_m + \delta w_m$$

- For reversible processes,

$$\rightarrow \delta w_m = -P dV_m \quad (P = P_B : \text{P reversible})$$

$$\rightarrow \delta q_m, \text{rev} = T dS_m$$

$$\therefore dU_m = T dS_m - P dV_m \quad \text{Fundamental Equations of Thermo}$$

- Recall: $h_m \equiv U_m + P V_m$

$$\rightarrow dh_m = d(U_m + P V_m) \stackrel{\text{chain rule}}{=} dU_m + P dV_m + V_m dP$$

$$(FE) dh_m = T dS_m - P dV_m + P dV_m + V_m dP$$

$$dh_m = T dS_m + V_m dP \quad \text{Fundamental Equations of Thermo (equivalent, just different form)}$$

Also define:

(just definitions)

$$\rightarrow \text{Gibbs Free energy: } G = H - TS$$

$$g_m = h_m - TS_m$$

$$\rightarrow \text{Helmholtz Free energy: } A = U - TS$$

$$a_m = u_m - TS_m$$

$$\Rightarrow dg_m = d(h_m - TS_m) = dh_m - T dS_m - S_m dT$$

$$dg_m = T dS_m + V_m dP - T dS_m - S_m dT$$

$$dg_m = -S_m dT + V_m dP \quad \text{Fundamental Equations of Thermo (Gibbs')}$$

$$\text{Similarly } da_m = -S_m dT - P dU_m \quad \text{Fundamental Equations of Thermo (Helmholtz')}$$

∴ Summary, Forms of FB:

Intensive

$$dU_m = T dS_m - P dV_m$$

$$dh_m = T dS_m + V_m dP$$

$$dg_m = -S_m dT + V_m dP$$

$$da_m = -S_m dT - P dU_m$$

Extensive

$$dU = T dS - P dV$$

$$dH = T dS + V dP$$

$$dG = -S dT + V dP$$

$$dA = -S dT - P dU$$

Q: The Fundamental Equations (any form) is valid for:

- ✓ 1) reversible processes
- ✓ 2) irreversible processes
- ✓ 3) open systems
- ✗ 4) none
- ✓ 5) all

1st part to drugs in thermodynamic properties!
State of system is only factor!