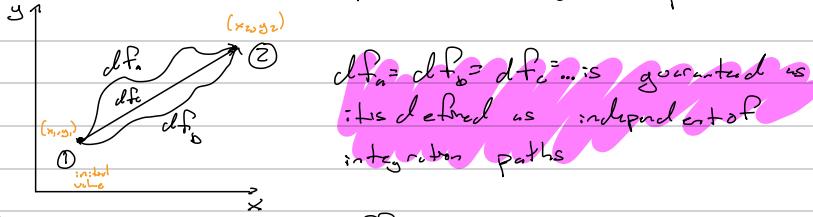




Math Review

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

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

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

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

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

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

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

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

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

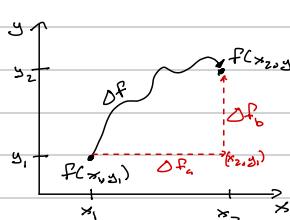
Summary - Common points of confusion!

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!

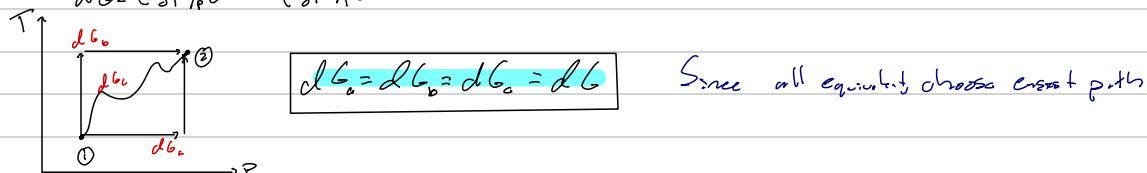
$\times \Rightarrow \int \frac{df}{dx}$ cannot be evaluated! Missing the differential!

\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.

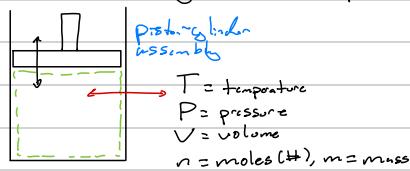
① System and Surroundings



• 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_{inf})	↑ extensive (C_{ext})	↑ int	↑ ext

density

We can convert ext quantities to int ones:

extensive = "specific property" ← lowercase with or w/o
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

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

③ "State" of a System

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

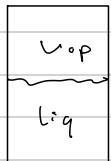
$$f = m - \pi + 2 \quad \text{for non-reactive systems}$$

↑
degrees of freedom
↑
of chemical species
↑
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



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

↓
molar of H_2O

↓
molar of H_2O

$$\sqrt{V_{\text{tot}}} = V = V_{m,L} n_L + V_{m,V} n_V \quad \rightarrow \text{divide by } \sqrt{n_{\text{tot}}} = \sqrt{n_L + n_V}$$

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

↑
average molar volume

↑
 $x \equiv \text{"quality"}$

$$\text{Also: } \beta = (1-x) \beta_L + x \beta_V$$

We can't for example specify V_m for each phase

→ constrains all other intensive properties

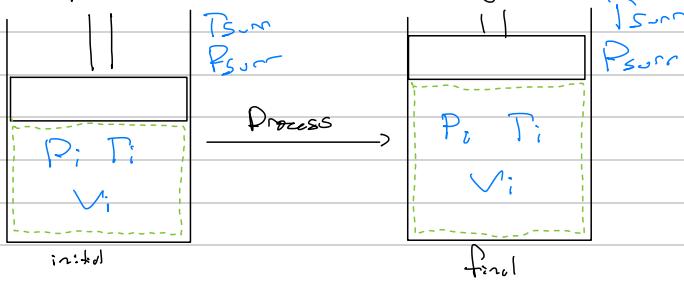
Also need to specify total amount in each phase

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

$Q_{\text{liq}} = \text{mole or mass fraction of } \text{H}_2\text{O}$
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 $\rightarrow P_{\text{sys}} = P_{\text{sum}}$

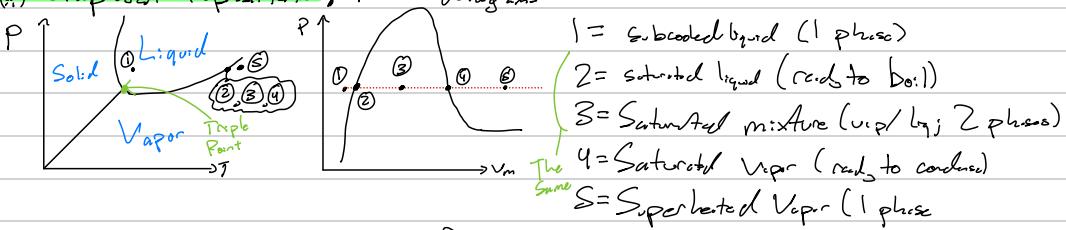
(iii) Phase Equilibrium \rightarrow Composition of phases is constant

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

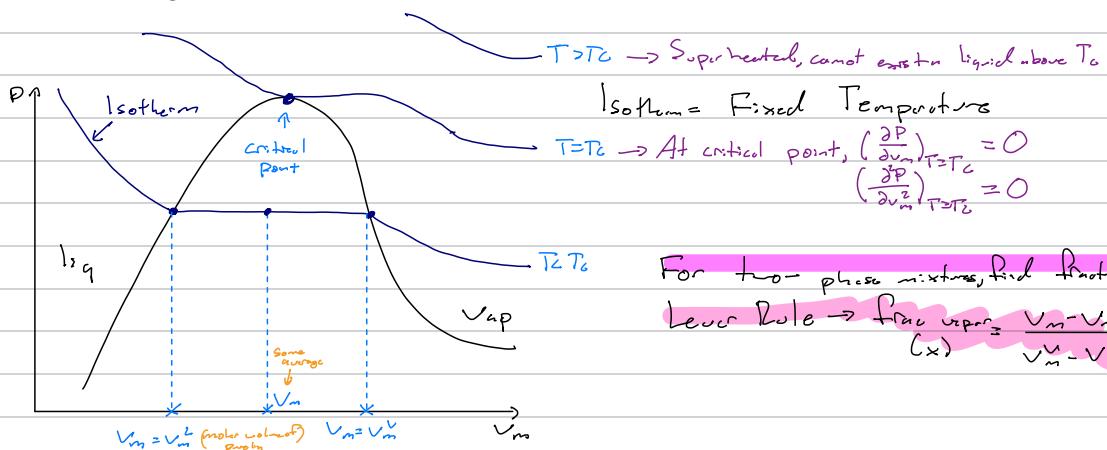
Thermodynamic Property Sources

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

(ii) Graphical Representation: PVT diagrams



• PVT diagrams are very useful!



For two-phase mixtures, find fraction of vapor using mass balance
Lever Rule \rightarrow frac vapor $= \frac{V_m - V_m^L}{V_m - V_m^V}$ (Derivation on Canvas)

(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 $\dot{V}_i \geq \dot{V}_{i,sat}$ you have pure vapor
- If $\dot{V}_{i,sat} > \dot{V} > \dot{V}_{i,l}$ you have a mixed phase
- If $\dot{V}_{i,l} \geq \dot{V}$ 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)

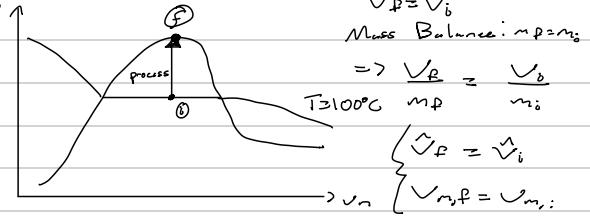
$$\begin{aligned} V_i &= 100L \\ T_i &= 100^\circ\text{C} \\ &\text{saturated} \end{aligned}$$

final (f)

$$\begin{aligned} V_f &= 100L = V_i \\ T_f &= T_c \\ P_f &= P_c \end{aligned}$$

process

→ 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

$$\dot{V}_f = \dot{V}_i = \dot{V}_c = 0.0033155 \frac{\text{m}^3}{\text{kg}}$$

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

bottom of tube

↳ of temp → $\dot{V}_{i,g} = 0$

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

(continued)

$$\dot{V}_i = \dot{V}_{i,L}^{sat} (1-x) + \dot{V}_{i,V}^{sat} x$$

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

S + Steam C 100°C

$$\dot{V}_{i,L}^{sat} = 0.001047 \frac{\text{m}^3}{\text{kg}}$$

$$\dot{V}_{i,V}^{sat} = 1.4194 \frac{\text{m}^3}{\text{kg}}$$

$$\rightarrow \dot{V}_{i,L}^{sat} < \dot{V} < \dot{V}_{i,V}^{sat} \Rightarrow L/V \text{ mixture}$$

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

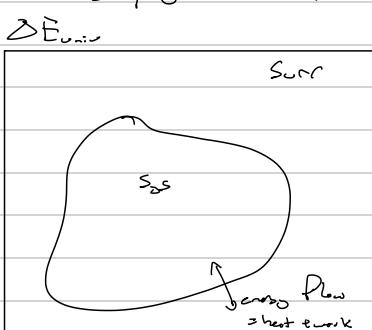
- phase changes

- 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

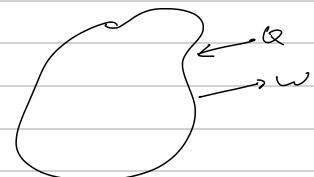
Q: Heat and work are flowing across the system boundary. What can we say about Q and W ?

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.)

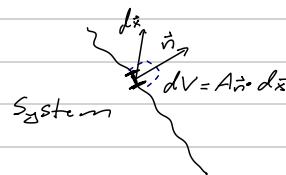
② Work (W): this class!

a) System boundary moves against an external force

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

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

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



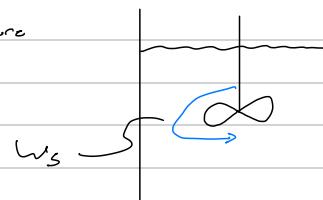
→ Calculating PV work requires us to know external pressure

* 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

New
Rec'd
T_{ip} Brovo
21
7

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

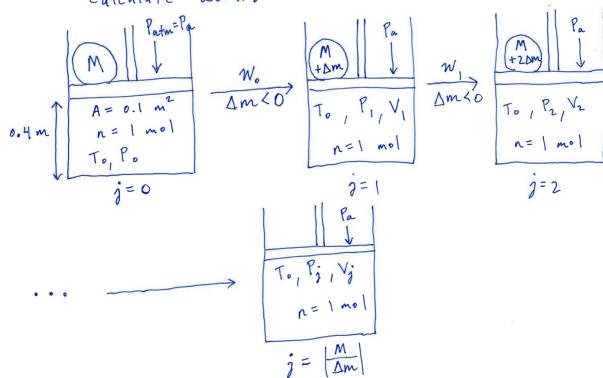
$$\frac{W}{n} = w_m = - RT_i \int_{V_{m,i}}^{V_{m,f}} \frac{dV_m}{V_m}$$

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

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

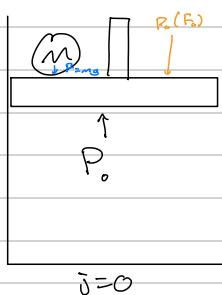
• PV work def'n → $W = - \int P_E dV$ external pressure

↳ what is P_E ?

• At eq. P_E should equal P_{eq}

• pressure / force balance:

Pressure / Force Balance



$$\text{initial } (j=0): A P_0 = A P_{E,0} \Rightarrow P_0 = P_{E,0} = A \left(\frac{M_0}{A} + P_0 \right)$$

cross-section area of piston

$$\rightarrow P_0 = P_0 + \frac{M_0}{A}$$

$\Delta m < 0$ for this problem given

$$\text{Similarly } (j=1): P_1 = P_{E,1} = \left(\frac{(M+M_0)_0}{A} + P_0 \right)$$

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

$P_{E,1} = P_1$

Substitution law

$P_1 V_1 = P_0 V_0$

$$\text{Now, sub } P's \text{ from force balance} = -V_0 \left[(P_0 + \frac{M_0}{A}) - (P_0 + \frac{M_0 + \Delta m_0}{A}) \right] = +V_0 \cdot \frac{\Delta m_0}{A} < 0$$

negative indicates S_{down} class work

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

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

$$\rightarrow \text{use force balance} = -V_0 \left(P_0 + \frac{M_0}{A} \right) \left(\frac{(-\Delta m_0)}{P_0 + \frac{M_0}{A} + \frac{\Delta m_0}{A}} \right)$$

P_0

$$\Rightarrow W_1 = \frac{V_0 \Delta m_0}{A} \cdot \frac{P_0 + M_0}{P_0 + \frac{M_0}{A} + \frac{\Delta m_0}{A}} < 0$$

Expansion → S_{down} 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)

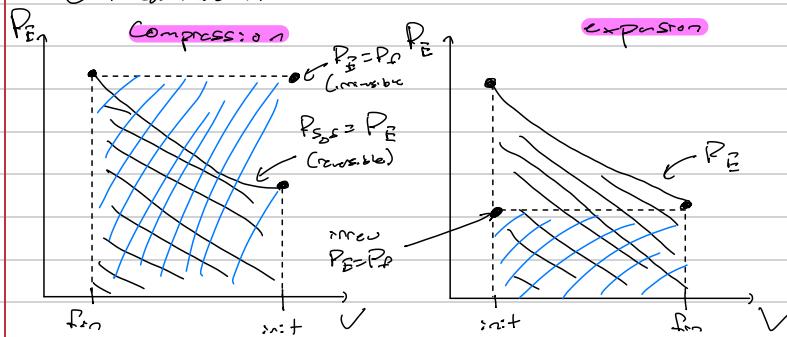
This is a reason why assuming a reversible process can be convenient. $W_{\text{re}} = -5545 \text{ J}$

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 (Combination 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: $dU_m = dq_m - Pdv_m$ ($P_E = P$)

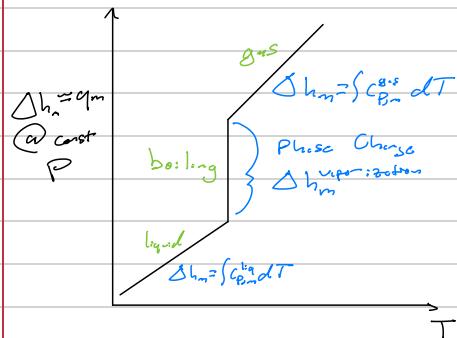
$dU_m = dq_m - d(PV_m)$ can make this more only bc Pressure P is constant

$$dU_m + d(PV_m) = dq_m$$

$$\underbrace{d(U_m + PV_m)}_{h_m} = dH_m = C_{p,m}dT = dq_m \quad \leftarrow \text{measure } q_m \text{ @ const } P$$

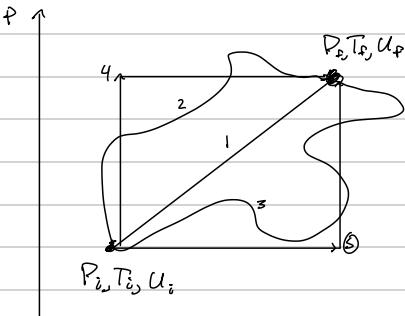
\leftarrow measure T to determine h_m & $C_{p,m}$

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 ($T, P, V_m, U_m, h_m, \dots$)

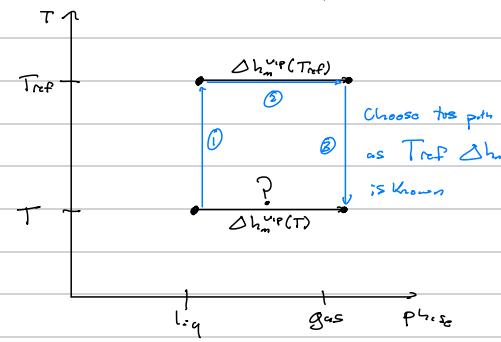


$$\Delta U = U_f - U_i$$

$$= \Delta U_1 = \Delta U_2 = \Delta U_3 = \Delta U_f$$

True for any process!

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



* T_{ref} is usually 298K

Solution:

$$\Delta h_m^{vp}(T) = \Delta h_m^{vp} + \Delta h_{m2} + \Delta h_{m3}$$

$$\textcircled{1} \text{ Sink phase heating/cooling} \quad \Delta h_{m1} = \int_{T_{ref}}^T C_{P,m}^{lq} dT$$

$$\textcircled{2} \text{ Phase change at } T_{ref}; \Delta h_{m2} = \Delta h_m^{vp}(T_{ref})$$

$$\textcircled{3} \text{ Sink phase cooling/heating} \quad \Delta h_{m3} = \int_{T_{ref}}^T C_{P,m}^{ss}(T) dT$$

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

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	$T_f = 540^\circ\text{C}$ $V_f = V_i (\text{rigid})$

Process:
closed system
constant volume

- 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
- $\dot{W} = - \int_{V_i}^{V_f} P_E dV = 0$ (no volume change)
- $\rightarrow KE, PE / mg$ small and negligible

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

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

Table B.2	$\hat{u}_i = 504.47 \frac{\text{kJ}}{\text{kg}}$	$\hat{u}_f = 2529.5 \frac{\text{kJ}}{\text{kg}}$
$B.2$	$\dot{V}_i = 0.001061 \frac{\text{m}^3}{\text{kg}}$	$C_v = 0.8857 \frac{\text{m}^2}{\text{kg}}$

$$\Rightarrow \dot{V}_i = \dot{V}_f = 0.373 \frac{\text{m}^3}{\text{kg}}$$

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

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

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

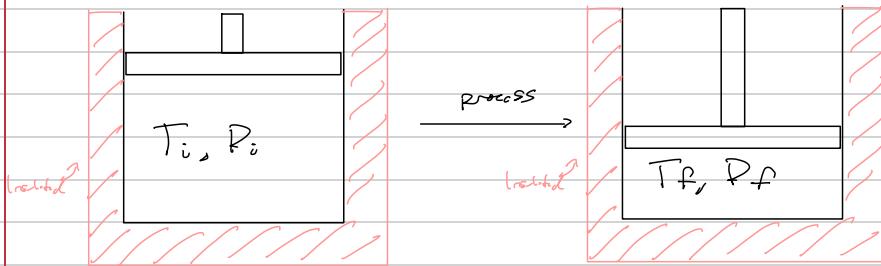
$$\hat{u}_f = 3194 \frac{\text{kJ}}{\text{kg}}$$

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

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

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

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)
- $\delta W = -P_B dV$ (definition of P_B)
- $V_f = V_i$ (ideal gas $\propto C_{v,m}dT$)
- $P_f = P_i$ (Reversible)

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

$$n C_{v,m} dT = -P dV$$

$$n C_{v,m} dT = -\frac{P_B T}{V} dV$$

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

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

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

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

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

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