

THERMODYNAMICS, CONT.

- note: read thermal expansion sections on your own.

HEAT.

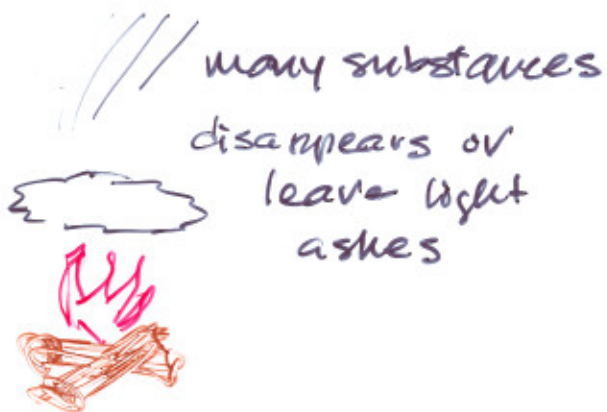
LONG... long... HISTORY.

- combustion
- frictional heating

} 18th Century theories
↓

wrong & shown to
be wrong--

Combustion:



metal gets heavier!

wood, for example... (steam & CO_2 take away
apparent mass)

George Stahl (1660-1734):

burning objects \rightarrow ~~releases~~ release PHLOGISTON

metals? confused... phlogiston has negative
weight?

Antoine Laurent Lavoisier (1743-1794)

showed combustion = oxygen + substance

\uparrow & necessary for
metabolism.

$\hookrightarrow \text{CO}_2$ & H_2O



evacuated...
diamond would not burn

Actually measured air, reaction

products... carefully

conservation
of
mass*

"The state needs no scientists."

* we know to be false
now, don't we!

Joseph Black (1728-1799)

How much heat required to raise temperature?

defined:

SPECIFIC HEAT

Quantity of heat required to raise
the temperature of 1g by 1°C.

substance-dependent.

Calorie/g.°C

$$Q = c \cdot m \cdot \Delta T$$

Annotations:

- ↑ Q : Cal
- ↑ c : specific heat, Cal/g/°C
- ↑ m : mass, g.
- ← ΔT : change in temperature, °C

<u>ΔQ</u>	<u>substance, 1gm</u>	<u>ΔT</u>	
1 Cal	water	1°C	← HIGH
1 Cal	mercury	30°C	
1 Cal	iron	9°C	

subtle point:

HEAT CAPACITY: just the proportionality:

$$Q = C \Delta T$$

depends on amount of stuff

so SPECIFIC HEATs are better units...

characteristic of material, not amount of it

SPECIFIC HEAT UNITS:

$$\text{cal/g}^\circ\text{C}$$

$$\text{cal/gK}$$

$$\text{J/kg}\cdot\text{K}$$

$$c = 1 \text{ cal/g}^\circ\text{C} = 1 \text{ cal/gK} = 1 \text{ BTU/16}^\circ\text{F} \\ = 4190 \text{ J/kgK}$$

From Black's work: WHY do some substances require more heat than others to raise their temps the same amount?

CALORIC... that's why.

A Fluid... that some substances had
a particular affinity for



ADD CALORIC



HEAT IT UP

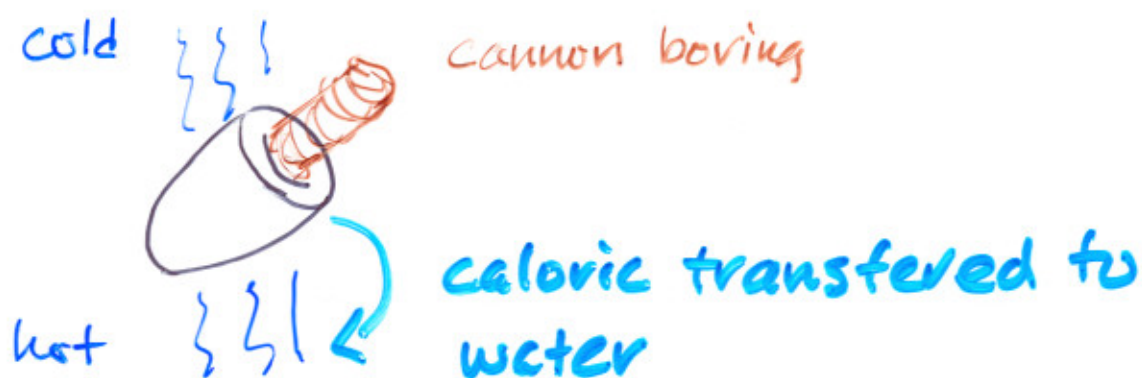
+ REVERSE

HEAT... A MATERIAL SUBSTANCE
CALORIC

TRAITOROUS THEORY -

Benjamin Thompson (1753-1814)

→ Count Rumford



PUT IT ALL UNDER WATER

boiled... and boiled...

and boiled...

"... evidently... inexhaustible..."

the metal chips?

identical to the cannon material.

→ THE HORSES & THE MOTION OF THE DRILL
THEY SUPPLIED.

SO... WHAT IS HEAT?

if not a substance?

IT TOOK SO MORE YEARS TO START
TO FORM AN ANSWER...

- James Prescott Joule (1818-1889)

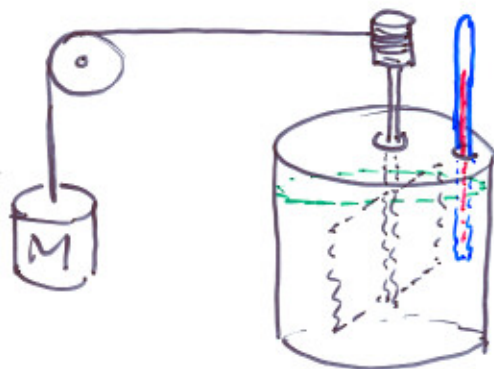
MANCHESTER BREWER

Became convinced:

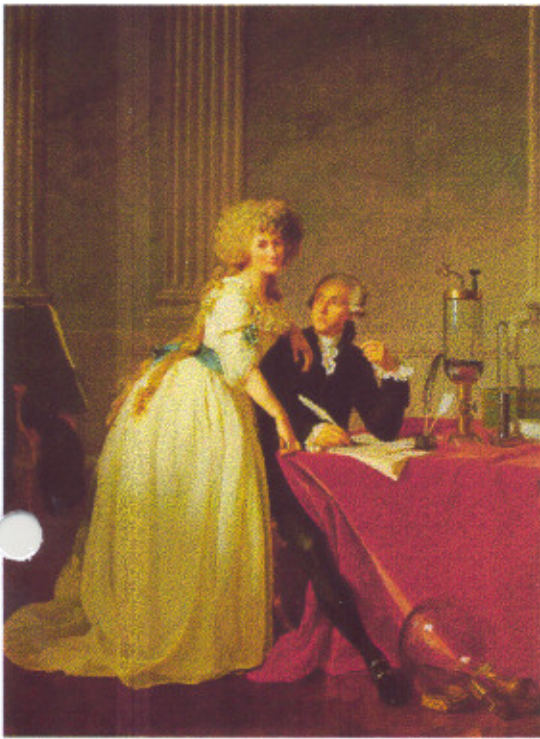
mechanical work \leftrightarrow heat

literally a numerical conversion...

1 Calorie \equiv 4.184 Joules



VERY PRECISE
EXPERIMENTS



PHY215

syllabus
Important Stuff
Links to biography & history
blog

fall2006

SO, WHERE'S IT GO?

HEAT IS ABSORBED ACCORDING TO ITS $\Delta T \uparrow$

SOMETIMES, A PHASE TRANSITION HAPPENS...

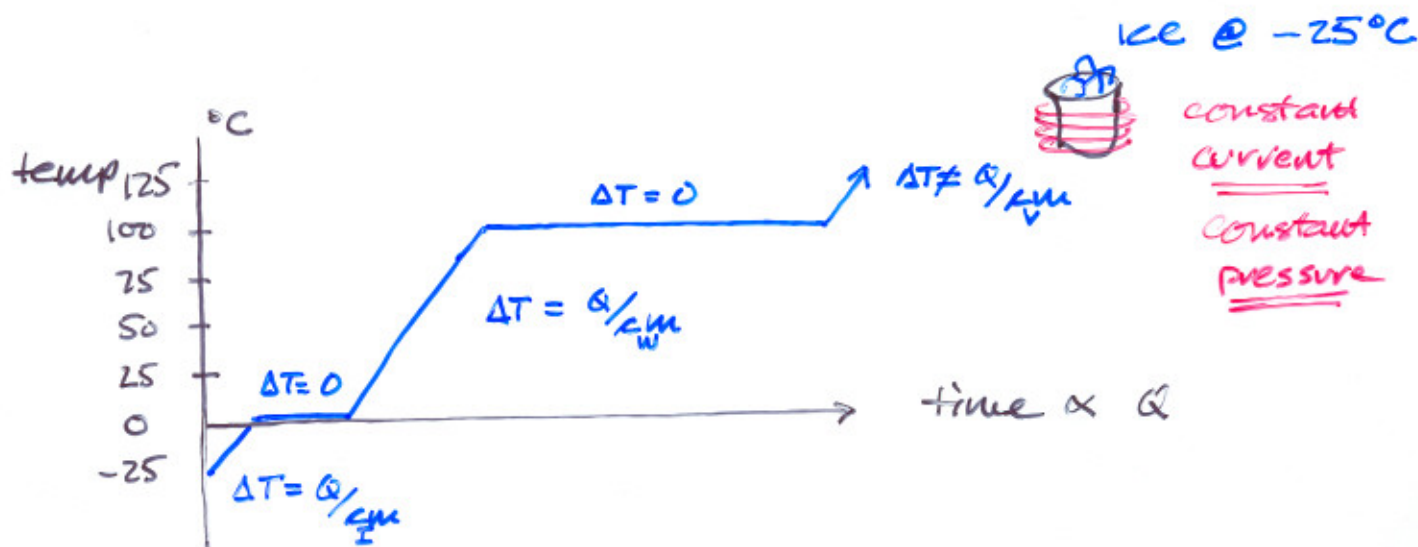
$\Delta T = 0$

according to the material's LATENT HEAT

$Q = mL$

(heat of transformation)

- | | | |
|-------|---|-------------------------|
| L_V | latent heat of vaporization
(heat of vaporization) | } Boiling
Condensing |
| L_F | latent heat of fusion
(heat of fusion) | |





**WATER IS A FAMILIAR (BUT UNUSUAL)
SUBSTANCE**

$$L_F = 80 \text{ Cal/g} = 333 \text{ kJ/kg}$$

$$L_V = 540 \text{ Cal/g} = 2256 \text{ kJ/g}$$

subtle points...

NOT ALWAYS grams, kilograms, etc.

SOMETIMES moles



$$1 \text{ mole} = 6.02 \times 10^{23} \text{ "units"}$$



Avogadro's Number... N_A .

$$1 \text{ mole of Aluminum} = 6.02 \times 10^{23} \text{ atoms}$$

$$\text{hydrogen} = 6.02 \times 10^{23} \text{ molecules}$$

So, if masses are replaced by moles -

molar specific heats

molar heats of fusion / vaporization

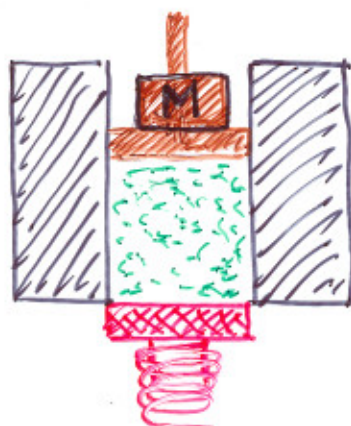
HEAT & WORK.

Always interested in relations among:

- temperature
- work
- heat
- pressure
- volume

THE
"SYSTEM"

↓
THE
GAS



PISTON → can measure force applied
can measure volume

WELL-INSULATED
SHELL

HEAT SOURCE — can measure heat applied
or absorbed

$$\text{PRESSURE}_G = \frac{F_G}{\text{area}} = \frac{F}{A}_G$$

move piston a bit... $d\vec{s}$ -- by gas 

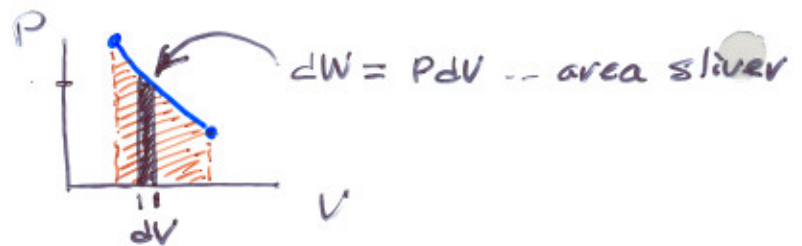
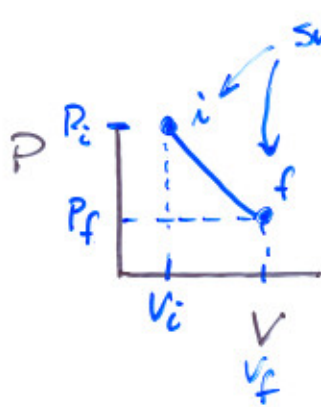
$$dW(\text{gas}) = \vec{F}_G \cdot d\vec{s} = P_G \vec{A} \cdot d\vec{s} = P_G dV$$

TOTAL WORK DONE BY GAS...

$$W = \int_{V_i}^{V_f} P dV$$

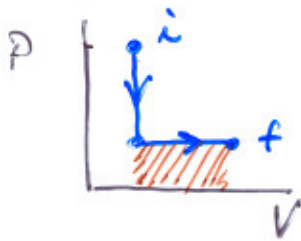
{ raise piston $\rightarrow +W_G$
 lower piston $\rightarrow -W_G$

P-V DIAGRAMS... tell the work story:

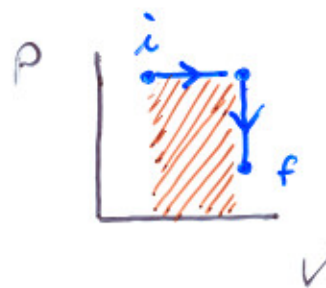


$W = \int P dV$ \rightarrow total area under P-V curve

lots of ways...



less



more

work

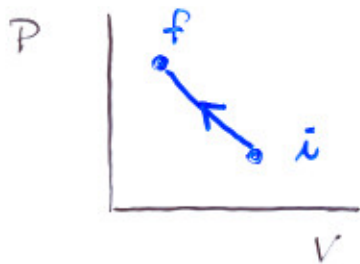
each increase V , so piston raised, so

$+W_G$

MUCH OF MECHANICAL ENGINEERING:

managing Thermodynamic Cycles...

go back:



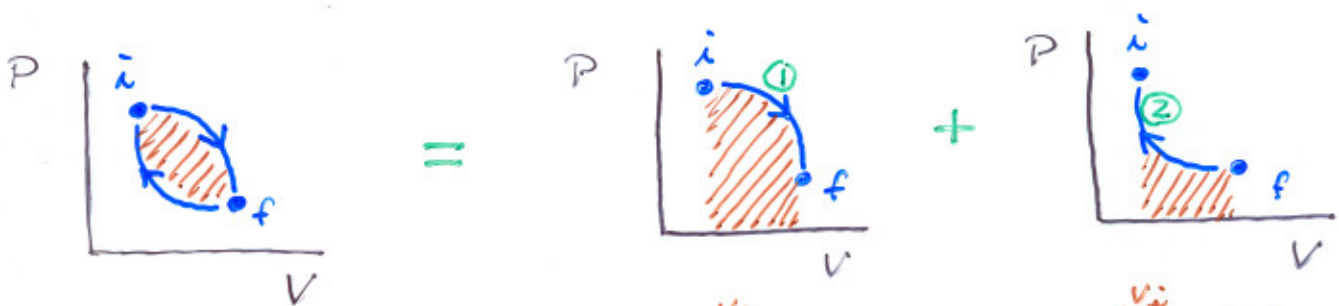
$i \rightarrow f$ here:

- reduces V of gas
- lowers piston
- $-W$ done by gas/system
($\Rightarrow +W$ done by piston)

back \neq forth \rightarrow path dependence for P vs V .



necessary to
get net work



A CYCLE

$W(\text{net})$

$=$

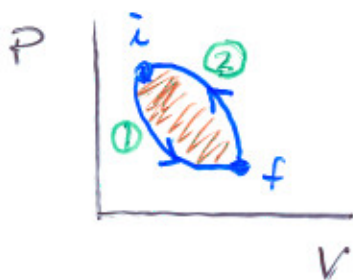
$W_{if}^{(1)}$

$+$

$W_{fi}^{(2)}$

> 0

negative work:



$$W_2 = \int_f^i P_2 dV$$

volume gets smaller

< 0

$$W_1 = \int_i^f P_1 dV$$

volume gets larger

BUT: $W_1 < W_2$

SO: $W(\text{net}) < 0 = W_1 - W_2$

ADD HEAT...

- CAN RAISE TEMPERATURE \rightarrow "increase" internal energy"
- CAN DO WORK

BUT

SOMETHING STAYS THE SAME:

\leftarrow experimental observation

$Q - W$

\hookrightarrow so, this is an intrinsic property
of a system.

INTERNAL ENERGY, U

a "State Function"

$$\Delta U = U_f - U_i$$

(= " E_{int} " in handouts)

$$\Delta U = \Delta Q - \Delta W$$



change in
internal
energy

↑
change
in heat
input

↑
work
done by
gas/system



**TEMPERATURE
MEASURES U**

FIRST "LAW" OF THERMODYNAMICS:

$$\Delta Q = \Delta U + \Delta W$$

$$dQ = dU + dW$$

dX is an "inexact" differential... just a little change
not representative of some function $X(P, V)$

CONSERVATION OF ENERGY

1ST LAW IS A LINKING OF HEAT ENERGY

&

MECHANICAL ENERGY

ASSOCIATED WITH

MAYER

JOULE

HELMHOLTZ



controversy, intrigue,
pathos, victory...

A SCIENTIFIC SOAP OPERA

MECHANICAL
ENERGY CHANGES FROM ONE FORM TO ANOTHER

often with the loss through heating

HEAT
ENERGY + MECHANICAL ENERGY IS CONSTANT

nothing comes from nothing---

STATE FUNCTION...

path - INDEPENDENT quantity

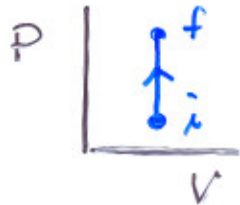
DOESN'T MATTER HOW YOU "GET THERE"

... just a point on the P-V diagram

U, T, P, V -- all state functions

W -- not so much.

DIFFERENT PATHS... DIFFERENT CONSEQUENCES

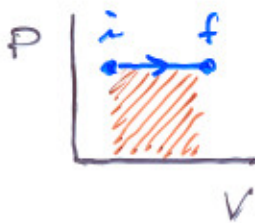


constant volume: **ISOCORIC**

$$\Delta W = \int P dV = 0 \Rightarrow \Delta Q = \Delta U + \Delta W$$

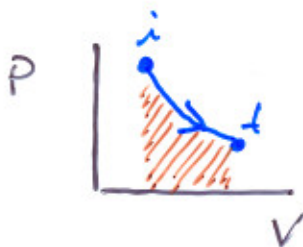
$$\Delta Q = \Delta U$$

heat all goes to increasing temperature



constant pressure: **ISOBARIC**

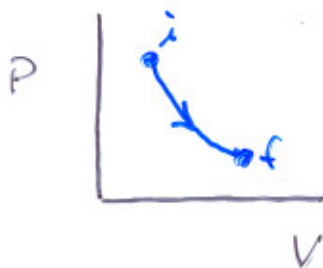
$$\Delta W = \int P dV > 0 \text{ here } \Delta Q = \Delta U + \Delta W$$



constant temperature: **ISOTHERMAL**

$$\Delta U = 0 \Rightarrow \Delta Q = \Delta W$$

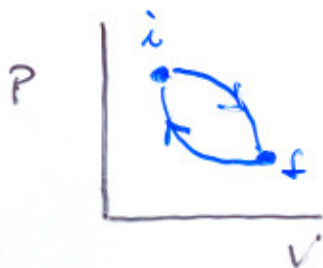
heat all goes to doing work



constant heat: **ADIABATIC**

$$\Delta Q = 0 \Rightarrow \Delta U = -\Delta W$$

work done by reducing internal energy



a cycle .. $i \rightarrow f \rightarrow i$

$$\Delta U = 0 \Rightarrow \Delta Q = \Delta W$$

all heat goes to do work

PROBLEM 73E. 200 J of work is done on a system
& 70.0 cal of heat is extracted from the system.
What are ΔW , ΔQ , ΔU ?

work done ON system \Rightarrow volume of system gets smaller

$$\Rightarrow \Delta W < 0$$

$$\Delta W = -200 \text{ J}$$

heat taken from system

$$\Rightarrow \Delta Q < 0$$

$$\Delta Q = -70 \text{ cal}$$

$$= -(70 \text{ cal})(4.186 \text{ J/cal})$$

$$\Delta Q = -293 \text{ J}$$

internal energy from 1st law:

$$\Delta Q = \Delta U + \Delta W$$

$$\Delta U = -\Delta W + \Delta Q \quad !$$

$$\Delta U = 200 \text{ J} - 293 \text{ J} = -93 \text{ J}$$

internal energy goes down, no temperature reduced.