Thermodynamics

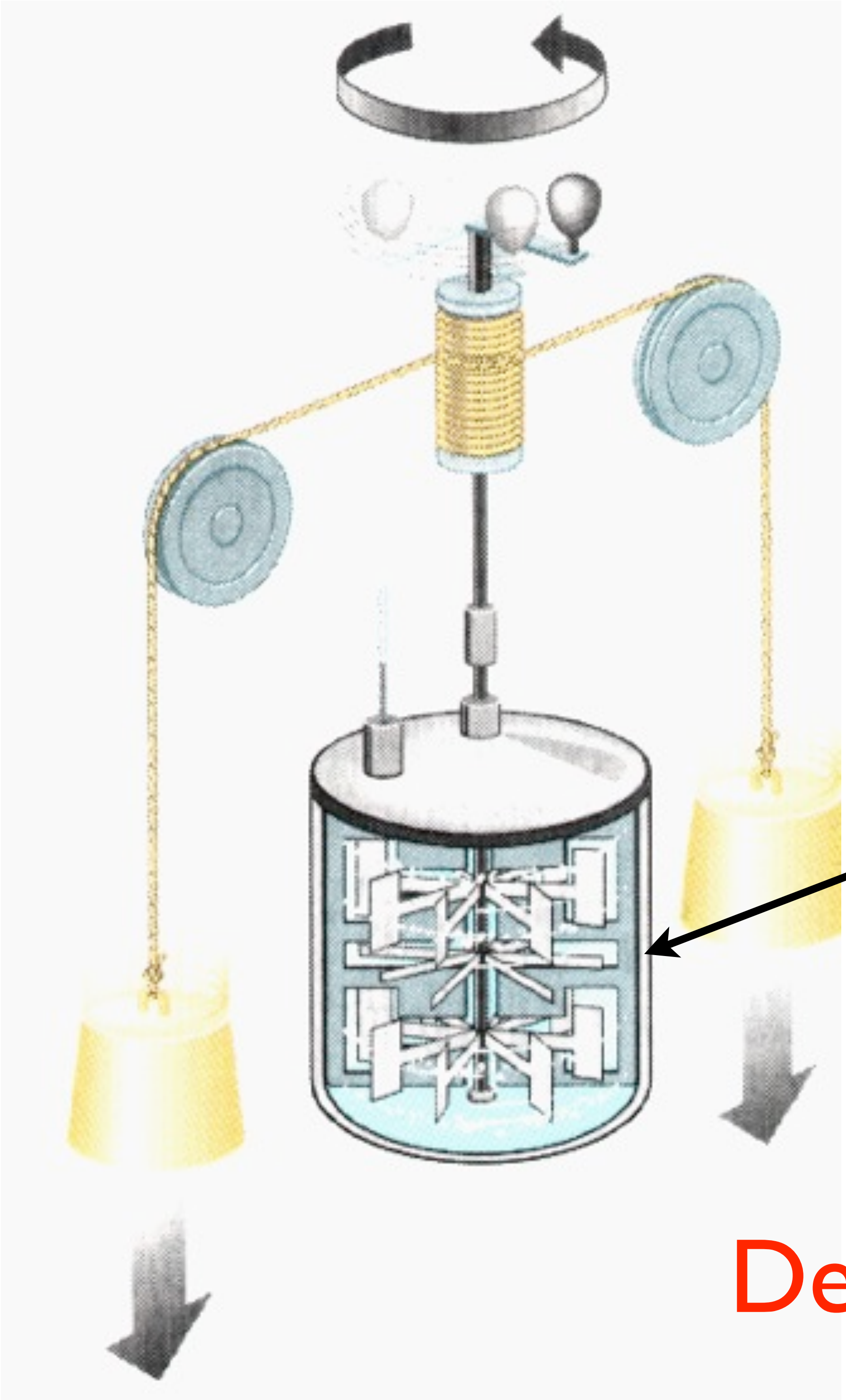
PHY 215
Thermodynamics and
Modern Physics

Fall 2025
MSU

Outline

- Mechanical equivalent of Heat
- Reversible vs. Irreversible
- Work done by gas system
- Thermodynamic Cycles
- “State Functions”
- 1st law
- C_p vs. C_v

Mechanical Equivalent of Heat



James Joule, 1843
(Julius von Mayer, 1854)

$\Delta Q/W$
constant!

$\Delta Q = Mc\Delta T$

$\Delta T \uparrow$

1 cal = 4.186 J

$W = mgh$

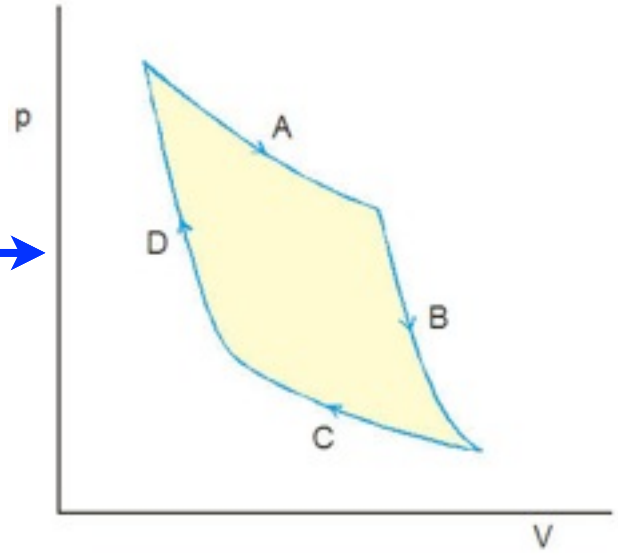
Work can be converted to heat, and vice versa!

Demo: Rub Hands

Image: http://atropos.as.arizona.edu

Reversible vs. Irreversible

- A reversible thermodynamic (heat transfer) process can be *reversed!*
  - Quasi-Static: system always, instantaneously, in thermal equilibrium. Use (p,V) diagram →
  - No dissipation.
  - No hysteresis: system proceeds backwards and forwards through same sequence - no memory!



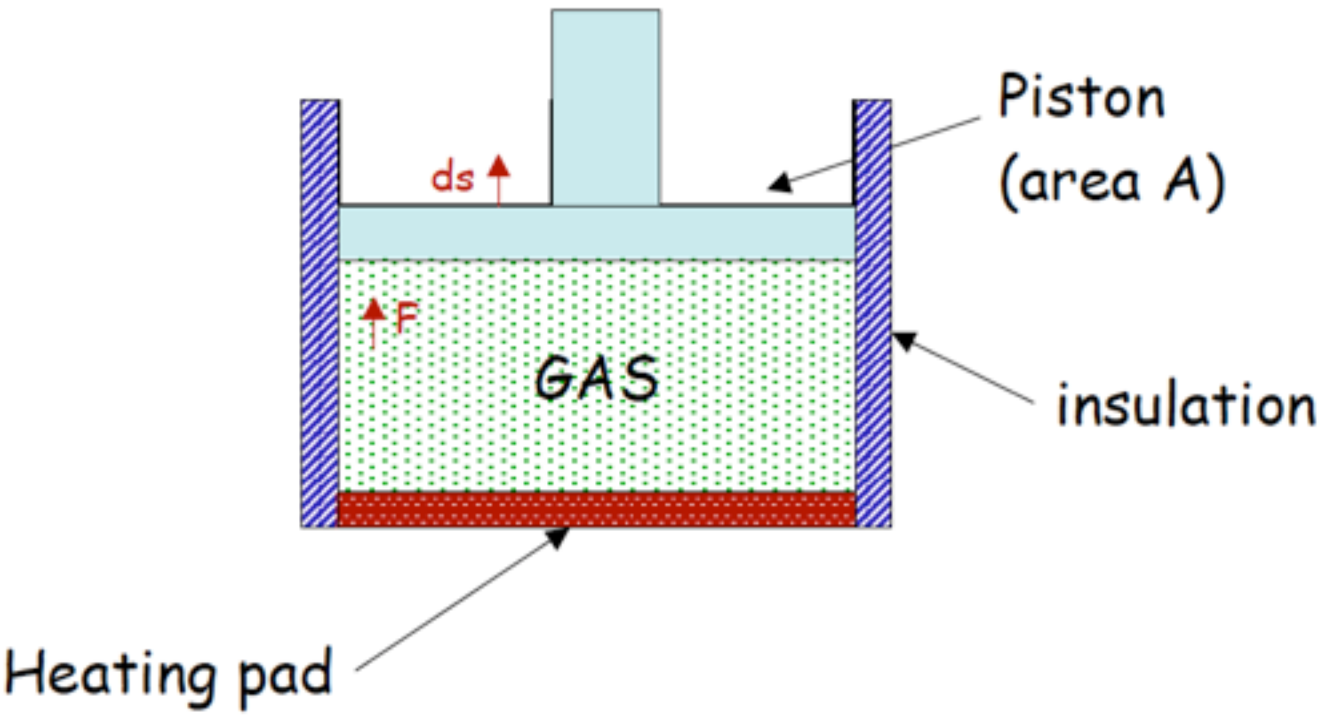
- An irreversible process is not reversible!

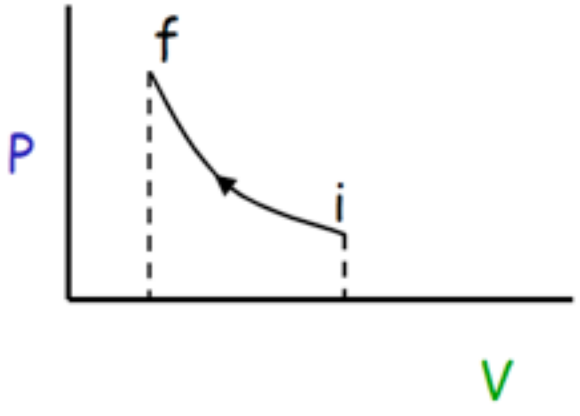
Demo: undo warm hands?

http://www.saburchill.com

Work done by Gas

Consider this system:





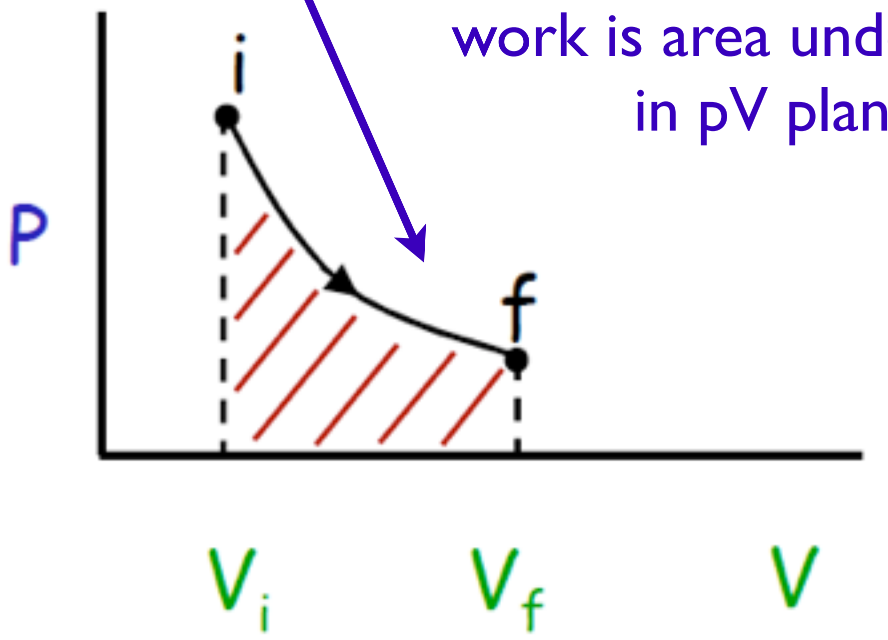
If volume decreases, the work done (by the gas) is negative.

If piston moves ds, then work done by the gas:

$$\boxed{dW = F\,ds = P\,A\,ds = P\,dV}$$

Total work done by the gas in moving from V_i to V_f :

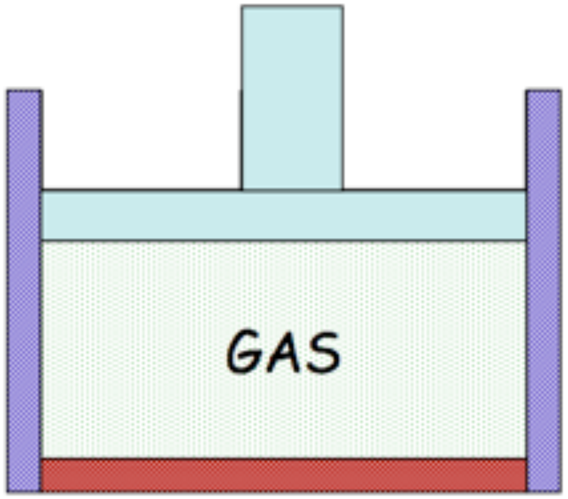
$$\Delta W = \int_{V_i}^{V_f} P\,dV$$



work is area under curve in pV plane

Example: Isothermal Expansion

$V_1 \rightarrow V_2$



T constant

$$
\begin{aligned}
W &= \int_{V_1}^{V_2} p\,dV \\
&= \int_{V_1}^{V_2} (nRT)\frac{dV}{V} = (nRT)\log\left(\frac{V_2}{V_1}\right)
\end{aligned}
$$

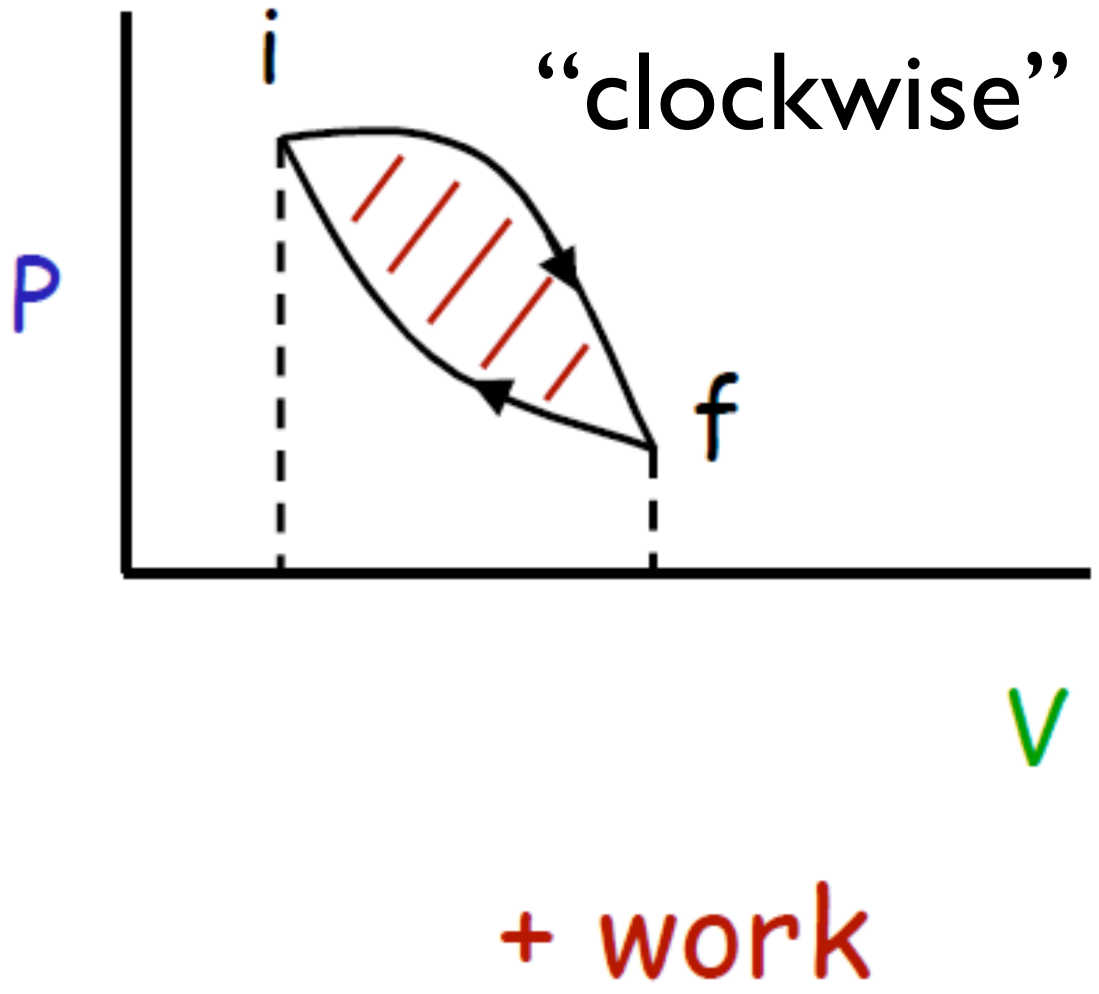
$\boxed{pV=nRT}$

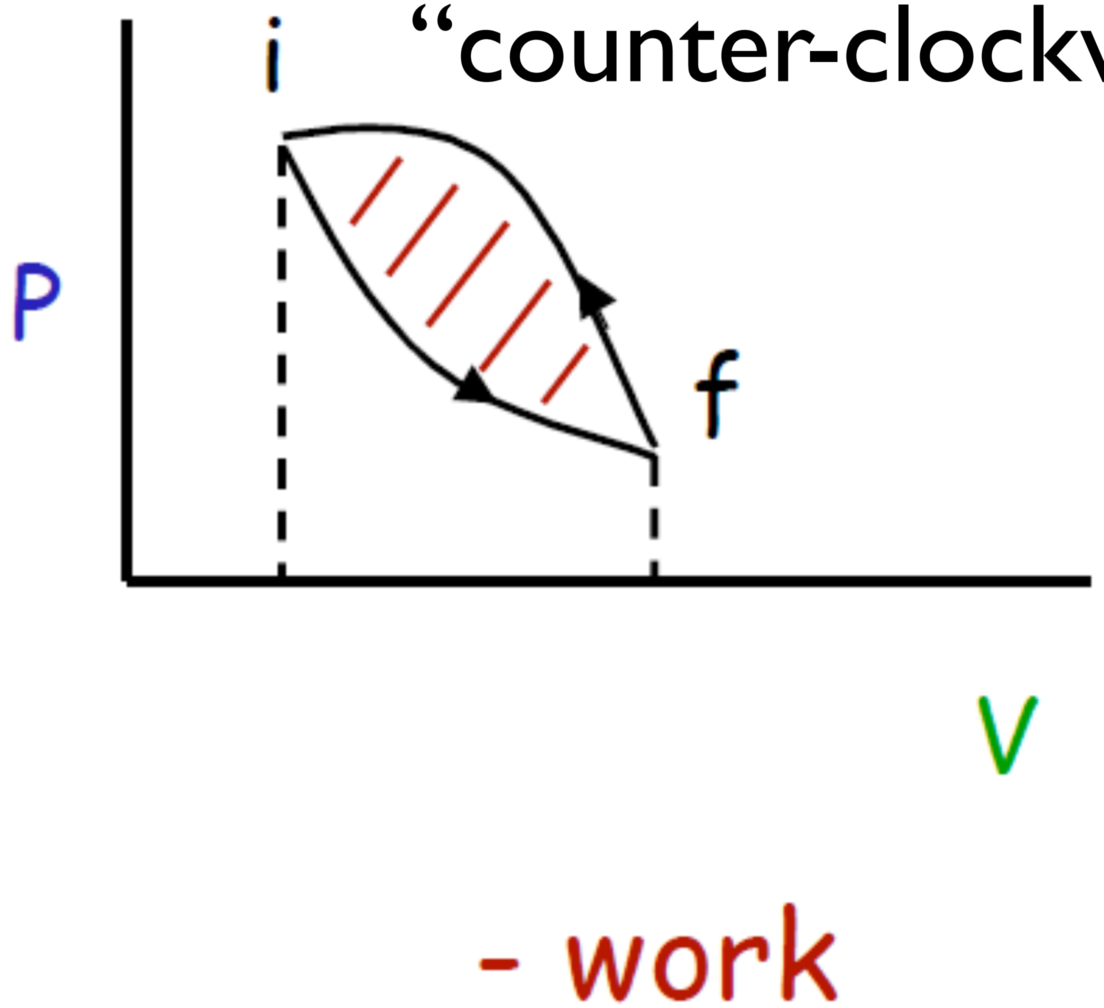
W>0: work done by gas

Gas at 40°C (=313°K) expands by 2x, does 3J work.
How many moles were present? (NB: R=8.3 J/mol/°K)

$$
\begin{aligned}
n &= \frac{W}{RT\log\left(\frac{V_2}{V_1}\right)} = \frac{3\,\mathrm{J}}{\left(8.3\,\frac{\mathrm{J}}{\mathrm{mol}\,{}^\circ K}\right)(313\,{}^\circ K)\log(2)} \\
&= 0.0017\,\mathrm{mol}
\end{aligned}
$$

Thermodynamic Cycles


“clockwise”


“counter-clockwise”

“State Functions”

- “State Functions” depend only on the parameters which define an equilibrium state, e.g. p, V, T, n, ...
- Example: U(n,T)=(3/2)nRT, monoatomic gas

$$\int_i^f dU = U_f - U_i \quad \& \quad \oint dU = 0$$

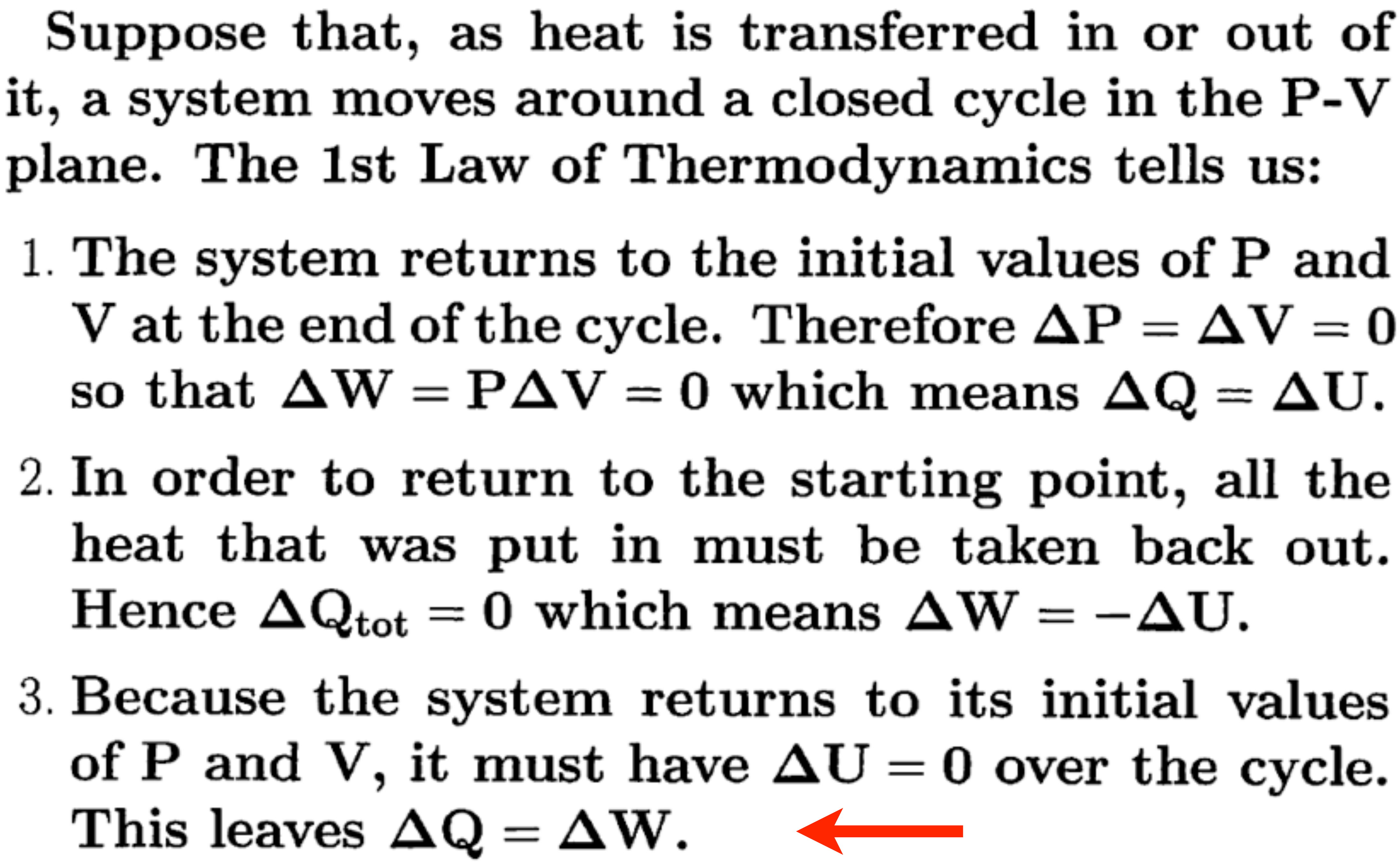
- W and Q are *not* state functions: depend on process which produced state.

$$\oint dW, dQ \neq 0$$

1st Law

- Heat (dQ) added to a system could
  - increase the internal energy (dU)
  - or go into work done by system (dW)
- dQ=dU+dW (Energy conservation!)
  - True for finite change: $\Delta Q=\Delta U+\Delta W$
  - True for reversible & irreversible processes
  - ALWAYS TRUE! (must include all possible types of work - electrical, elastic, etc.)

Concept Test

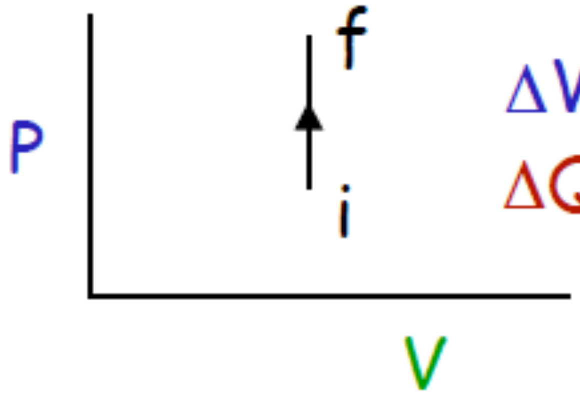


Suppose that, as heat is transferred in or out of it, a system moves around a closed cycle in the P-V plane. The 1st Law of Thermodynamics tells us:

1. The system returns to the initial values of P and V at the end of the cycle. Therefore $\Delta P = \Delta V = 0$ so that $\Delta W = P\Delta V = 0$ which means $\Delta Q = \Delta U$.
2. In order to return to the starting point, all the heat that was put in must be taken back out. Hence $\Delta Q_{\mathrm{tot}} = 0$ which means $\Delta W = -\Delta U$.
3. Because the system returns to its initial values of P and V, it must have $\Delta U = 0$ over the cycle. This leaves $\Delta Q = \Delta W$.

Thermodynamic Processes

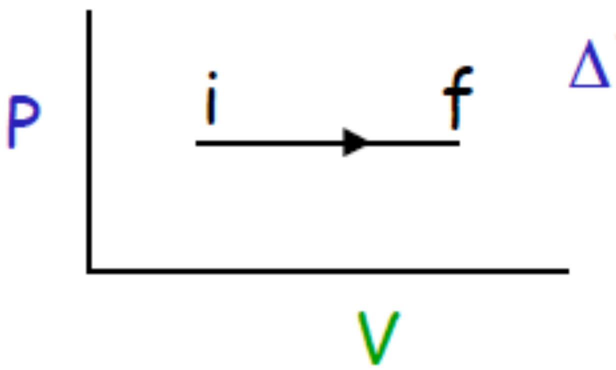
Constant Volume
(isochoric)



$\Delta W=0$

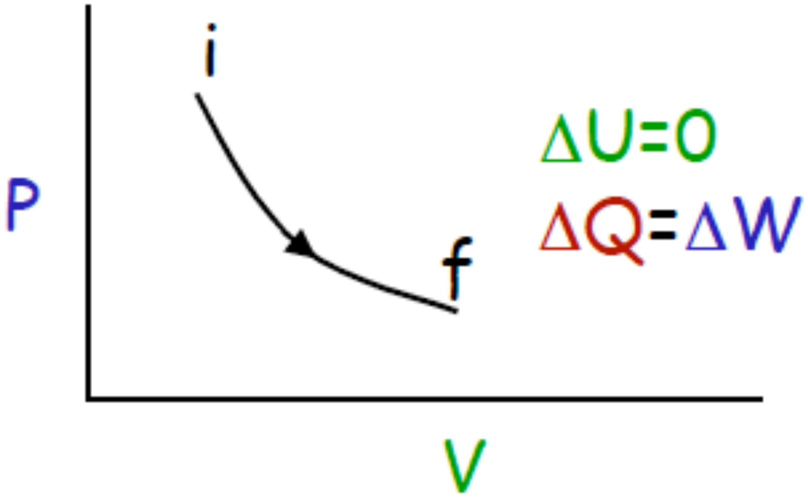
$\Delta Q=\Delta U$

Constant Pressure
(isobaric)



$\Delta W=P\Delta V$

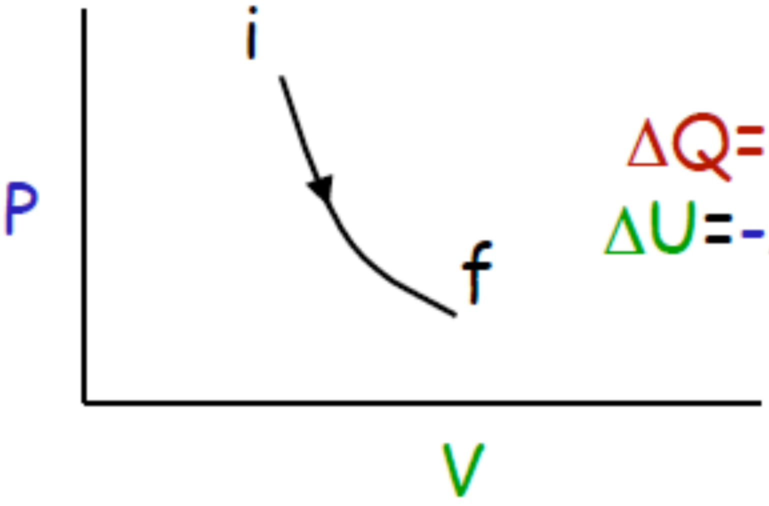
Constant Temp
(isothermal)



$\Delta U=0$

$\Delta Q=\Delta W$

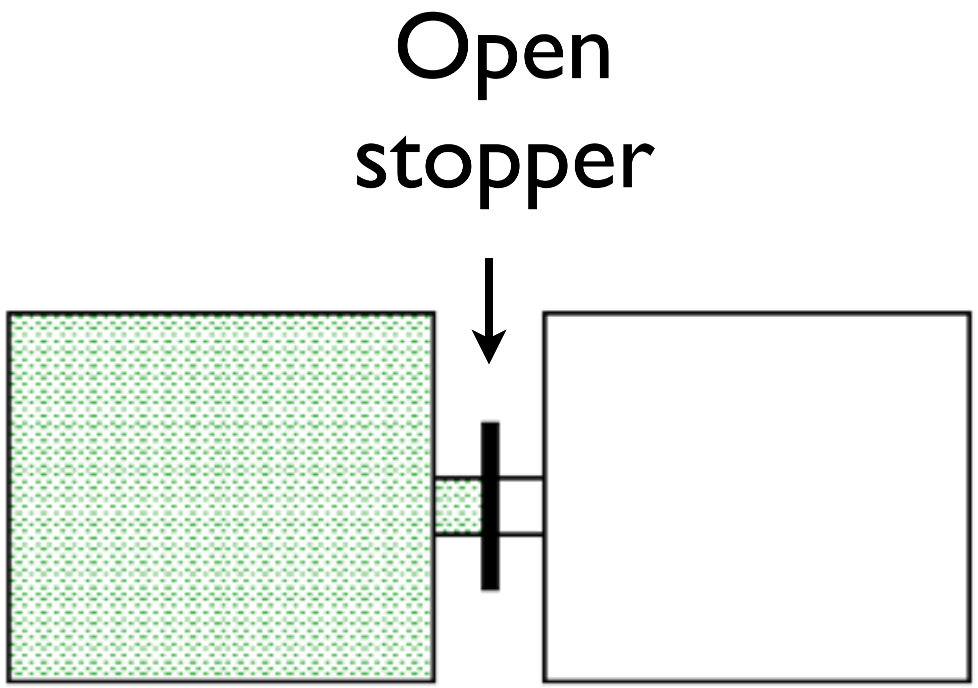
Constant Heat
(adiabatic)



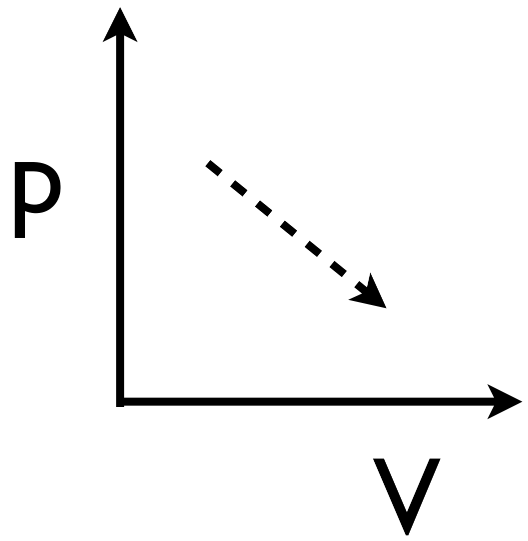
$\Delta Q=0$

$\Delta U=-\Delta W$

Free Expansion



Not quasi-static



$\Delta Q = 0$ (adiabatic)

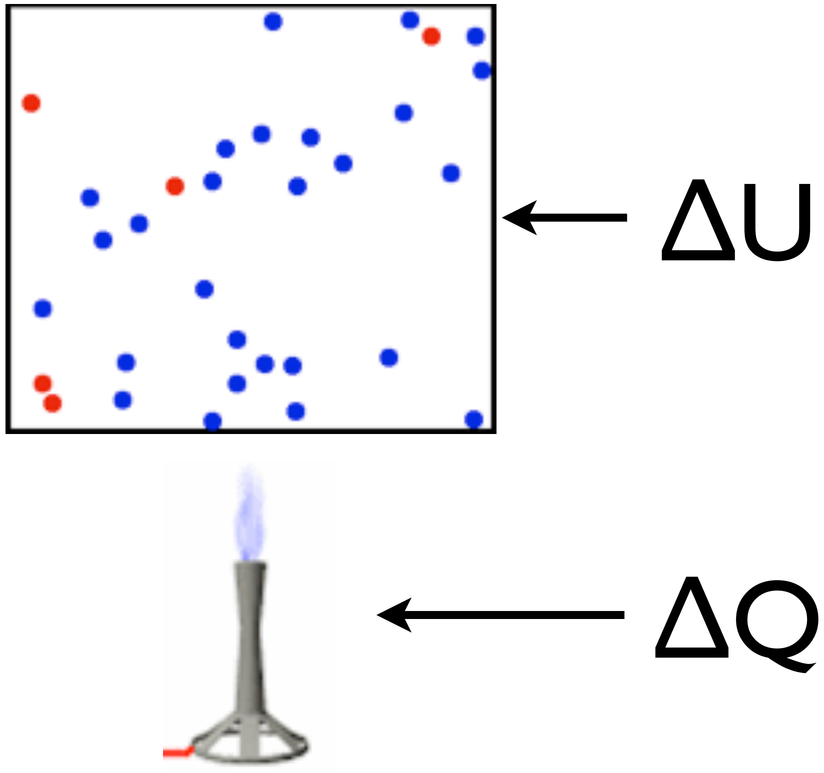
$\Delta W = 0$ (nothing to work against)

$\Rightarrow \Delta U = 0$

- Not represented by a smooth path in (p,V) plane
- Experiment: p drops, v increases, T constant

Molar Specific Heat

at constant volume



Monatomic Gas

$$\Delta Q = \Delta U = (3/2)\, nR\Delta T$$

$$\Delta Q = n\, C_V \Delta T$$

molar specific heat

$$C_V = (3/2)\, R = 12.5\ \mathrm{J/(mol \cdot K)}$$

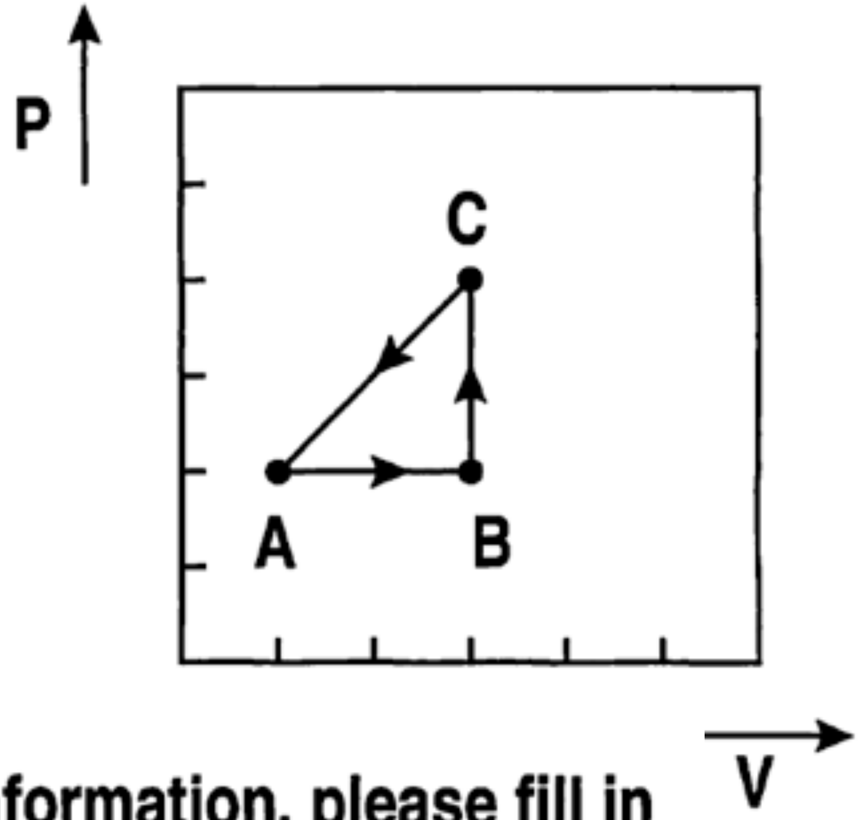
Images: http://en.wikipedia.org, http://www.sciencetime.co.uk

C_p vs. C_V

- C_V: Molar specific heat at constant volume
- Constant pressure: gas heats up and expands, doing work $\Delta W = p\Delta V$.
- Ideal gas law: $p\Delta V = nR\Delta T$, for constant p
- $\Delta Q = nC_p\Delta T = \Delta U + p\Delta V = \Delta U + (nR\Delta T)$
- $\Delta U = nC_V\Delta T$
- Therefore: $C_p = C_V + R$

In Class Exercise

A system is taken along the indicated path in the P-V plane: A -> B -> C -> A.



Given the following information, please fill in the remaining spaces in the table with either "0" or "+" or "-" as appropriate. Be prepared to explain your answers!

	dQ	dW	dU
A -> B	+	+	+
B -> C	+	0	+
C -> A	-	-	-

- Direction
- 1st Law
- U a state function

Summary

- Heat is a form of energy; 1 cal = 4.186 J.
- Thermodynamics processes can be *reversible* or *irreversible*.
- Thermodynamic Cycles in (p,V) plane
- Reversible processes are quasi-static, and have no hysteresis.
- Work done by gas system: $dW=pdV$
- 1st Law: $dQ=dU+dW$
- Q and W are not "state functions"!
- $C_P=C_V+R$