

Thermodynamics

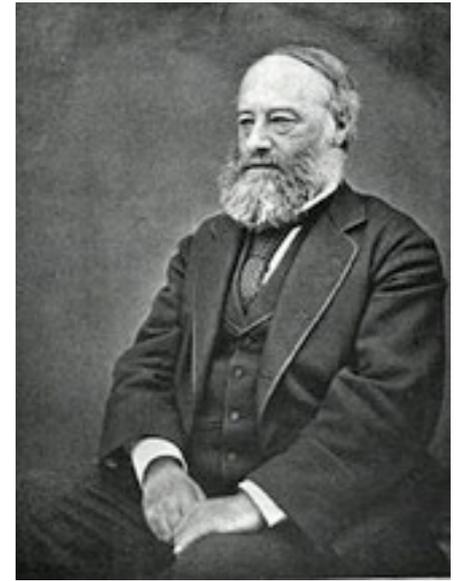
PHY 215
Thermodynamics and
Modern Physics

Spring 2026
MSU

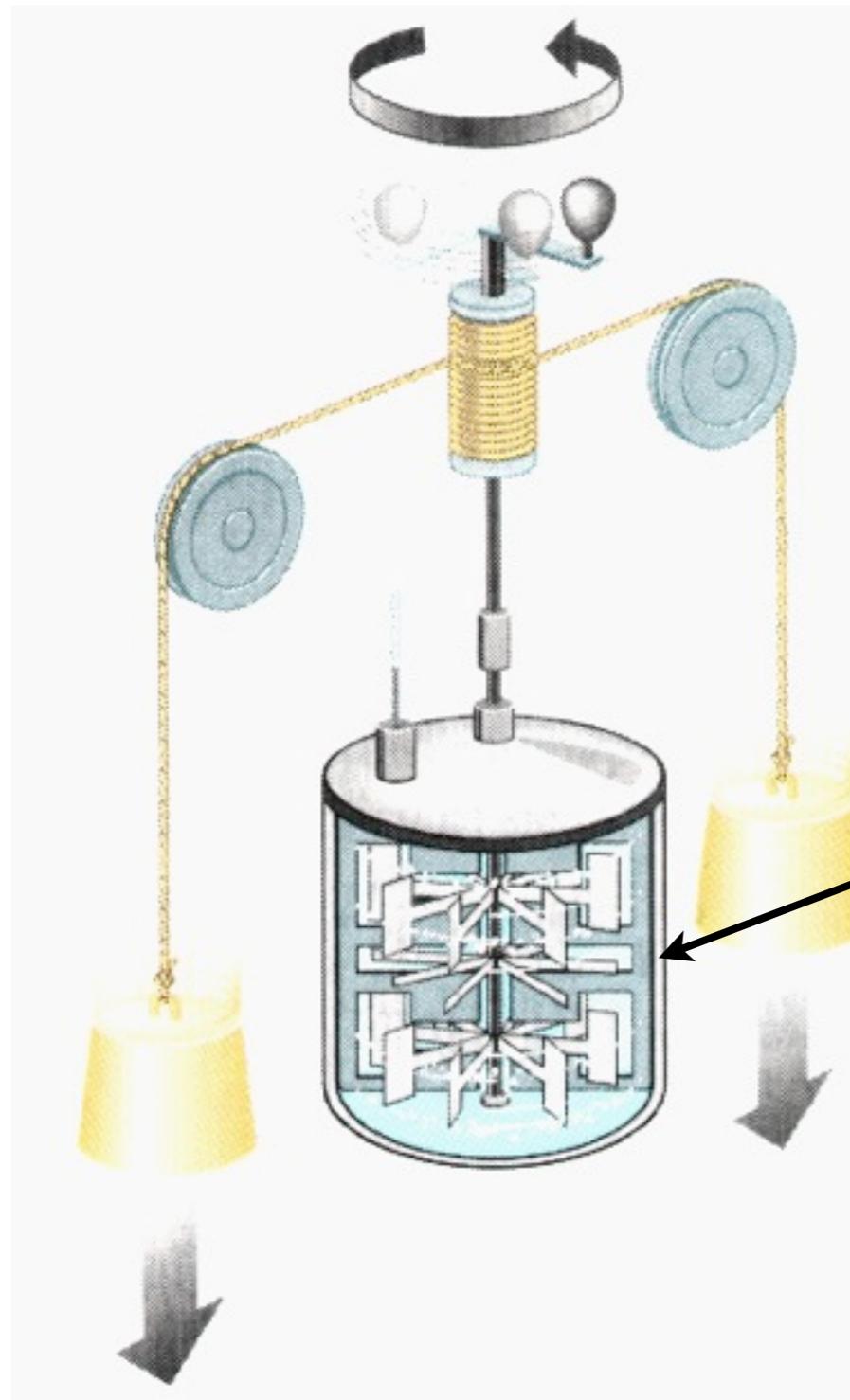
Outline

- 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!

$$\Delta Q = Mc\Delta T$$

$$1 \text{ cal} = 4.186 \text{ J}$$

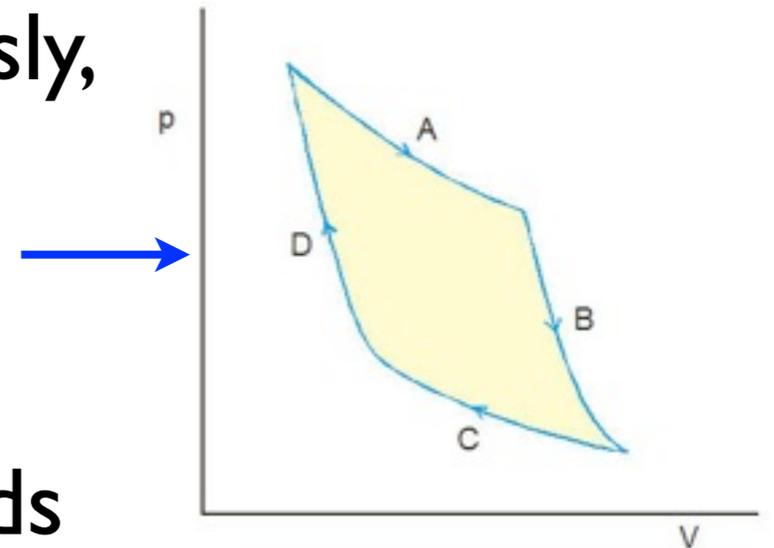
$$\Delta T \uparrow$$

$$W = mgh$$

Work can be converted to heat, and vice versa!

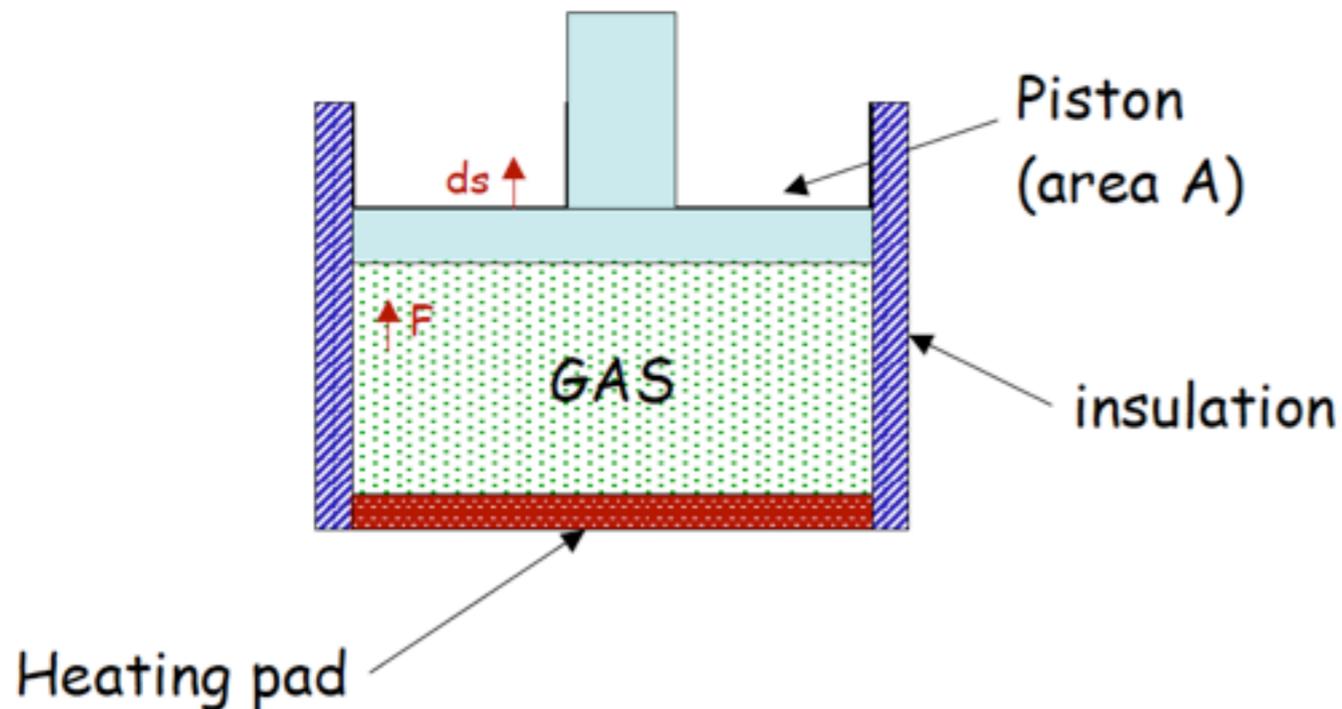
Reversible vs. Irreversible

- 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!



Work done by Gas

Consider this system:

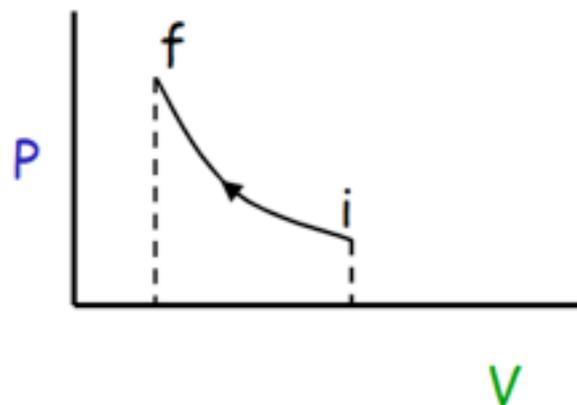
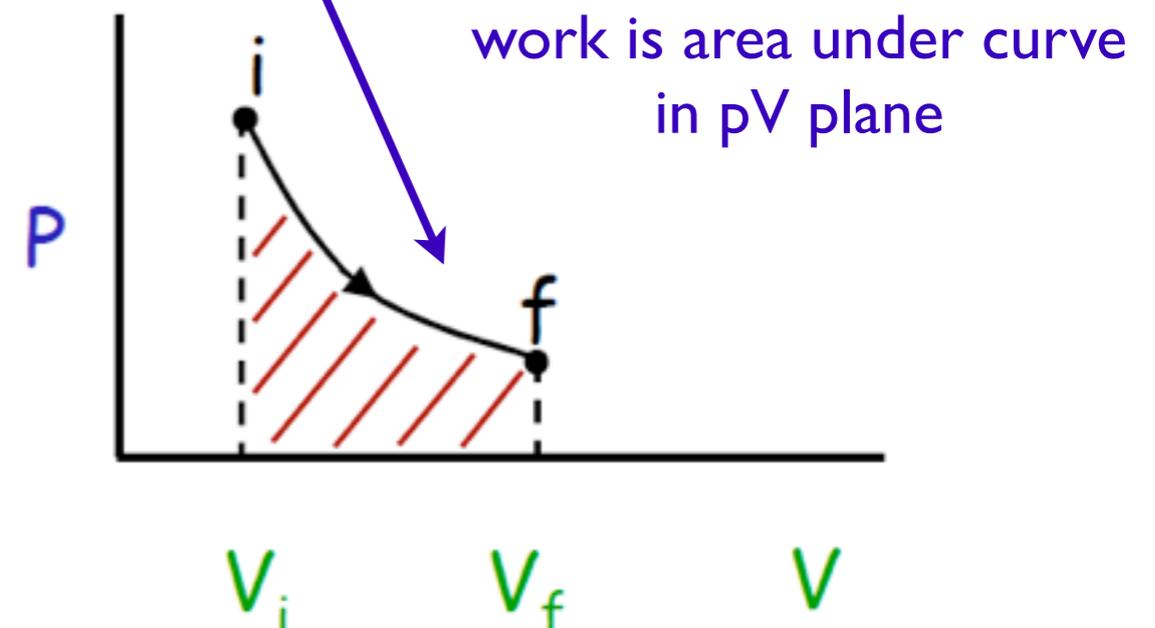


If piston moves ds , then work done by the gas:

$$dW = F ds = P A ds = P dV$$

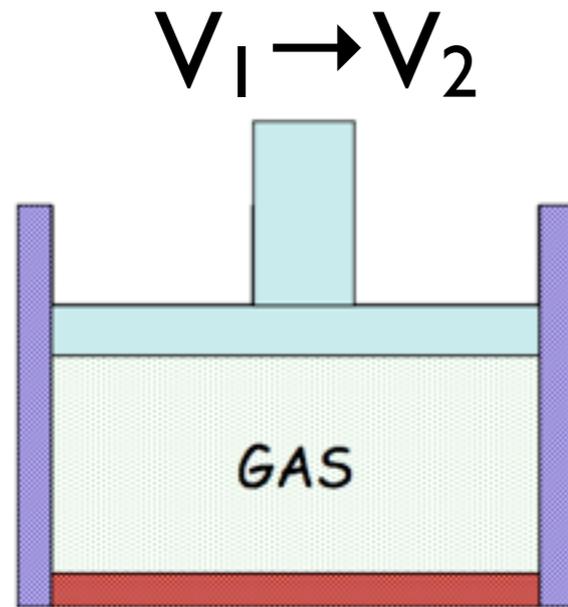
Total work done by the gas in moving from V_i to V_f :

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



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

Example: Isothermal Expansion



T constant

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

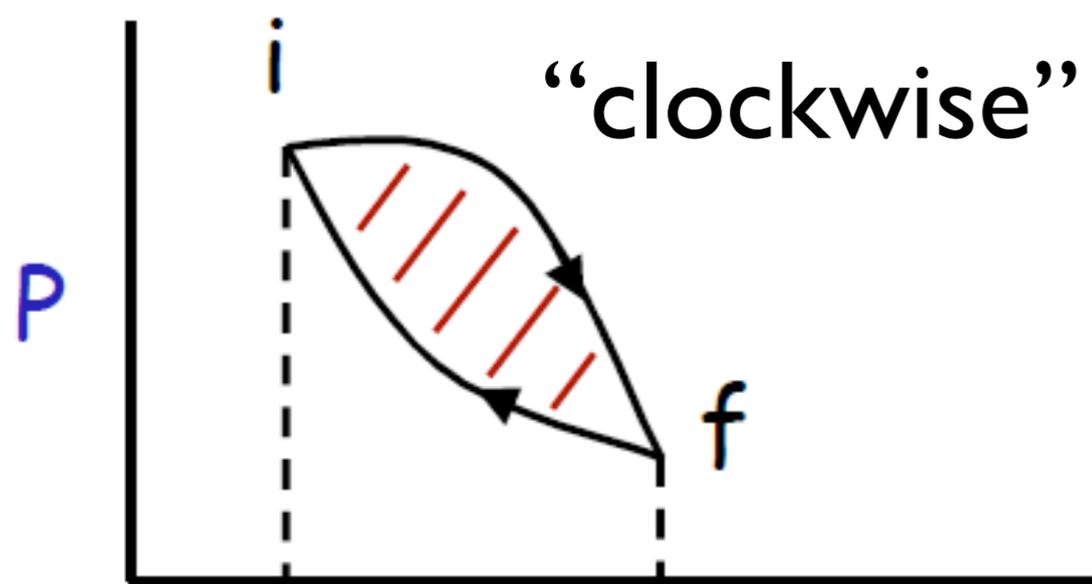
$$pV = nRT$$

$W > 0$: work done by gas

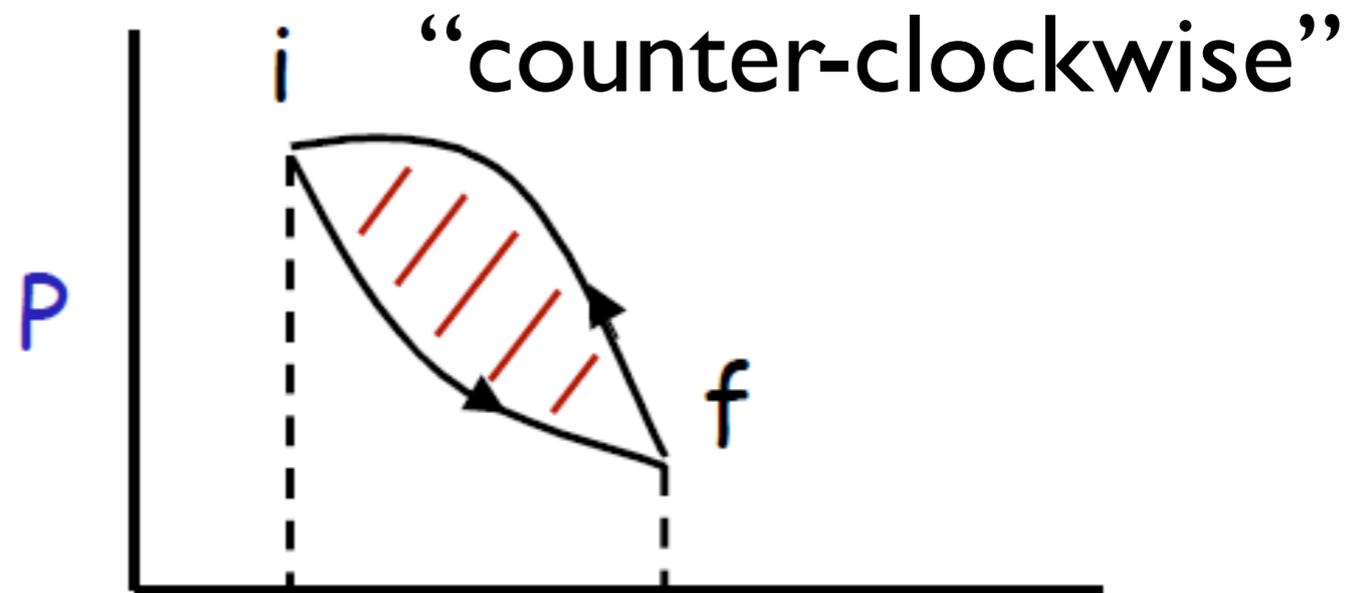
Gas at 40°C ($=313^\circ\text{K}$) expands by 2x, does 3J work.
How many moles were present? (NB: $R=8.3 \text{ J/mol/}^\circ\text{K}$)

$$n = \frac{W}{RT \log \left(\frac{V_2}{V_1} \right)} = \frac{3 \text{ J}}{\left(8.3 \frac{\text{J}}{\text{mol}^\circ\text{K}} \right) (313^\circ\text{K}) \log(2)}$$
$$= 0.0017 \text{ mol}$$

Thermodynamic Cycles



+ work



- work

“State Functions”

- “State Functions” depend only on the parameters which define an equilibrium state, e.g. p, V, T, n, \dots
- 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.)

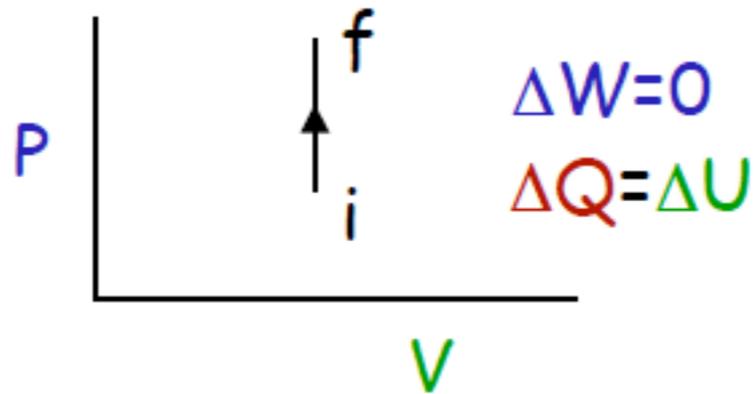
Concept Test

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:

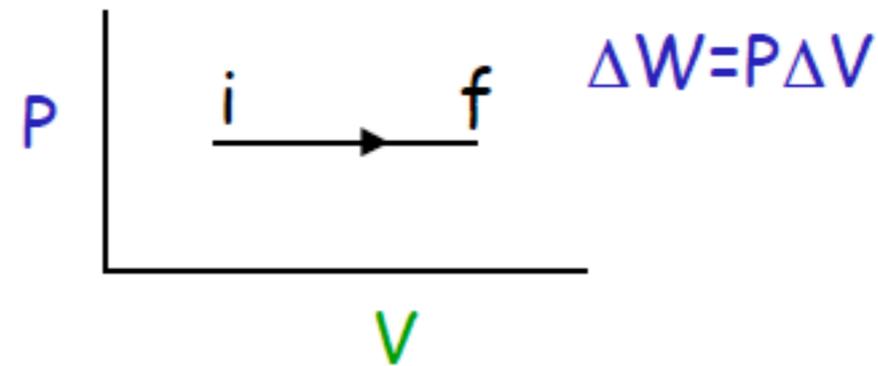
1. The system returns to the initial values of P and V at the end of the cycle. Therefore $\Delta P = \Delta V = 0$ so that $\Delta W = P\Delta V = 0$ which means $\Delta Q = \Delta U$.
2. In order to return to the starting point, all the heat that was put in must be taken back out. Hence $\Delta Q_{\text{tot}}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

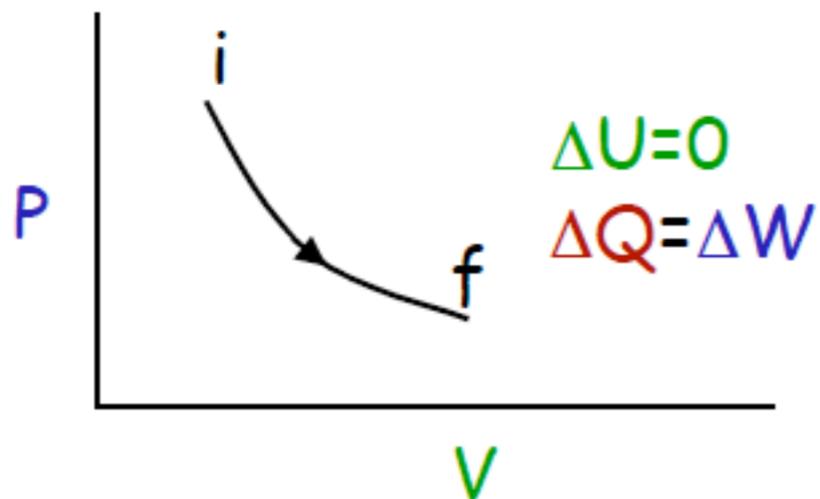
Constant Volume
(isochoric)



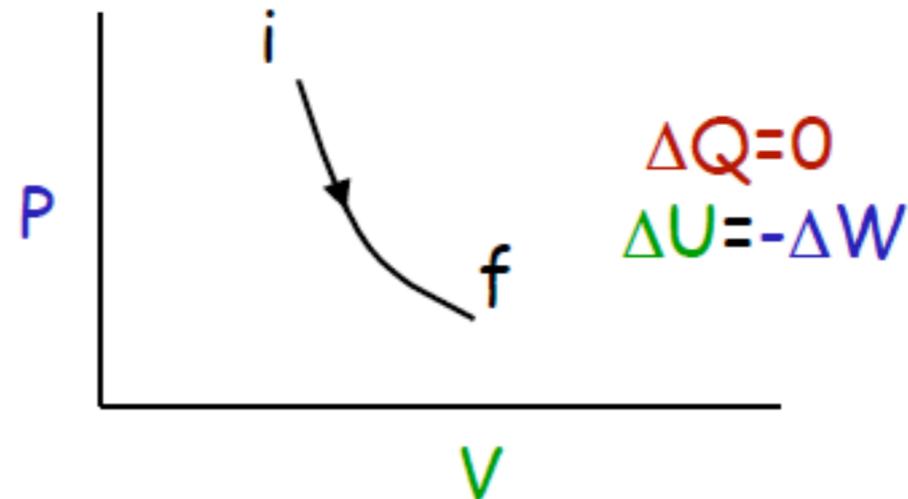
Constant Pressure
(isobaric)



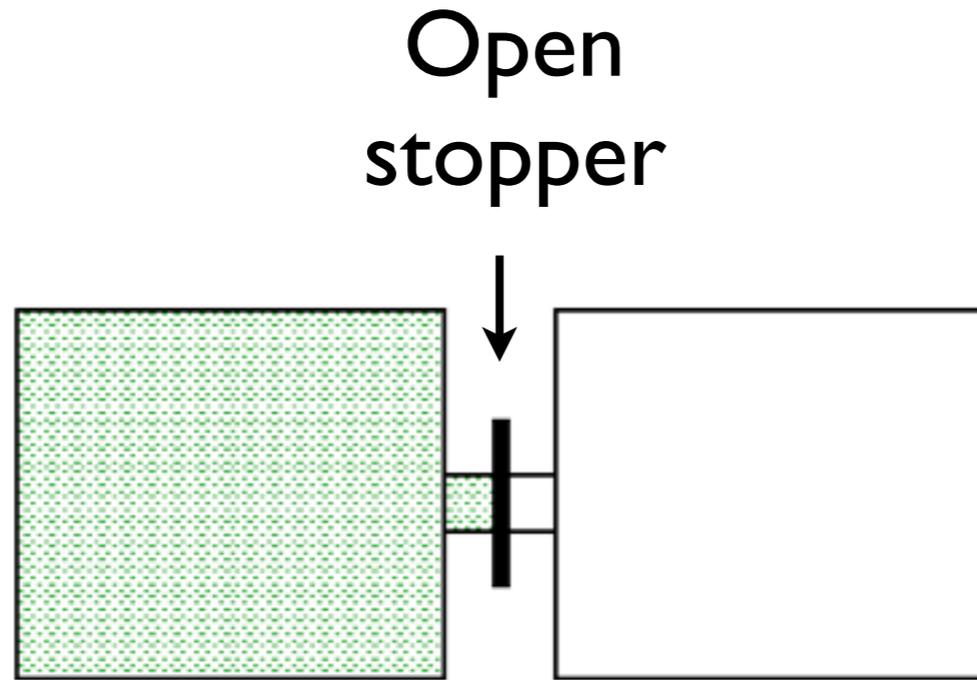
Constant Temp
(isothermal)



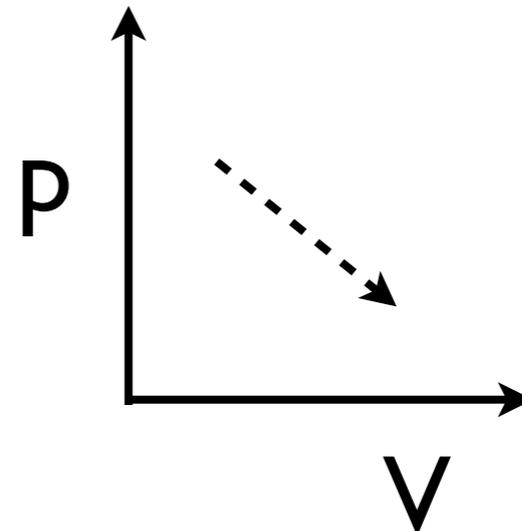
Constant Heat
(adiabatic)



Free Expansion



Not quasi-static



$$\Delta Q = 0 \quad (\text{adiabatic})$$

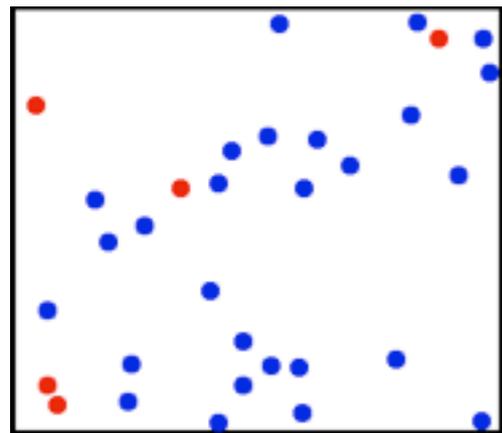
$$\Delta W = 0 \quad (\text{nothing to work against})$$

$$\Rightarrow \Delta U = 0$$

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

Molar Specific Heat

at constant volume



← ΔU



← ΔQ

Monatomic Gas

$$\Delta Q = \Delta U = (3/2) n R \Delta T$$

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

← molar specific heat

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

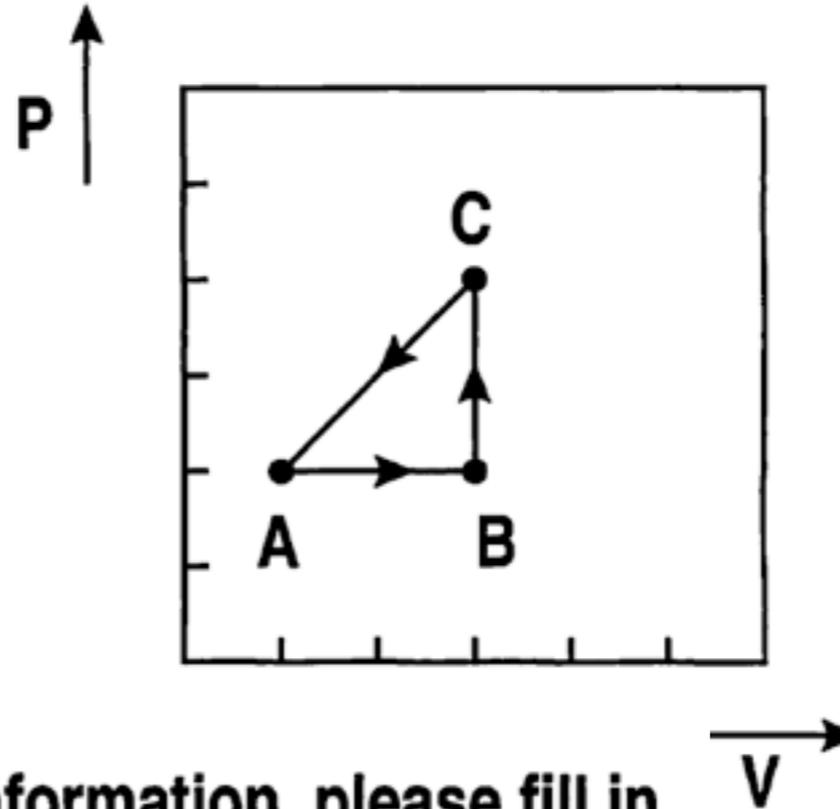
C_p vs. C_v

- 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: $C_p = C_v + R$

$C_v = \frac{f}{2}R$, where
of degree of freedom
 $f = 3$ for monatomic gas
 $f = 5$ for diatomic gas

In Class Exercise

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



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 \rightarrow B	+	+	+
B \rightarrow C	+	0	+
C \rightarrow A	-	-	-

- Direction
- 1st Law
- U a state function

Summary

- Heat is a form of energy; $1 \text{ cal} = 4.186 \text{ J}$.
- 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”!
- $C_p = C_v + R$