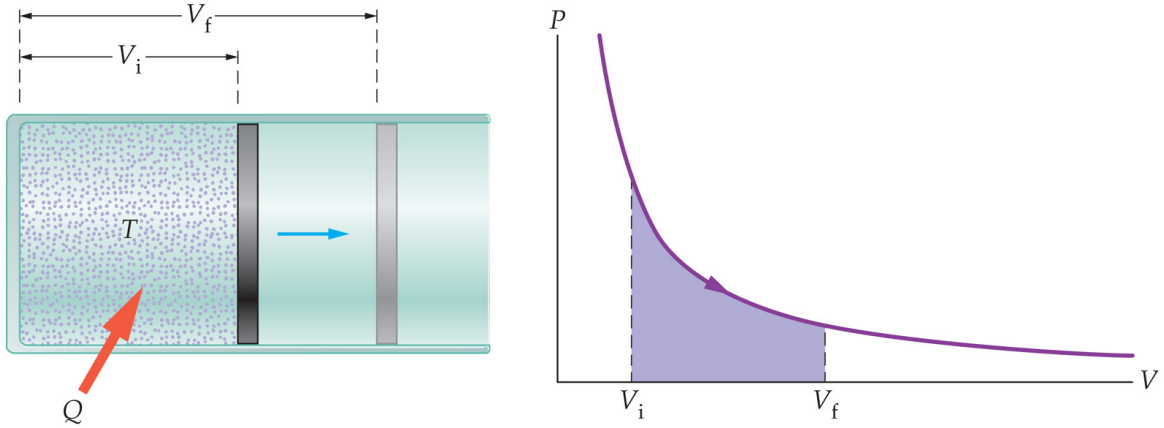
Physics 115
General Physics II

Session 12



Thermodynamic processes

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Lecture Schedule
(up to exam 2)

21-Apr	Mon	12	Specific Heats, Second Law	18.4-18.6
22-Apr	Tues	13	Entropy	18.7-18.10
24-Apr	Thurs	14	Charges	19.1-19.3
25-Apr	Fri	15	E field	19.4-19.5
28-Apr	Mon	16	Conductors	19.6
29-Apr	Tues	17	Gauss law	19.7
1-May	Thurs	18	Electrical potential	20.1-20.3
2-May	Fri	19	Potential, conductors	20.4
5-May	Mon	20	Capacitors	20.5-20.6
6-May	Tues	21	Current	21.1-21.2
8-May	Thurs	22	Power, Series & Parallel Circuits	21.3-21.4
9-May	**Fri**		**EXAM 2 - Ch. 18,19,20**	

Today

Announcements

- Exam 1 scores just came back from scan shop – grades will be posted on WebAssign gradebook later today
  - I will post on Catalyst Gradebook tomorrow, along with statistics (avg, standard deviation)

Solutions: see ex1-14-solns.pdf posted in class website *Slides* directory

Next topic: Laws of Thermodynamics

(Each of these 4 "laws" has many alternative versions)

We'll see what all these words mean later...

- 0^{th} Law: if objects are in thermal equilibrium, they have the same T, and no heat flows between them
  - Already discussed
- 1^{st} Law: Conservation of energy, including heat:

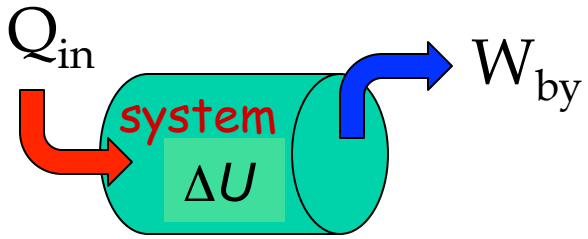
  Change in internal energy of system = Heat added – Work done
- 2^{nd} Law: when objects of different T are in contact, *spontaneous* heat flow is from higher T to lower T
- 3^{rd} Law: It is impossible to bring an object to T=0K in any finite sequence of processes

Internal energy and 1st Law

- **1st Law of Thermodynamics:**
  The change in internal energy of a system equals the heat transfer **into** the system* minus the work done **by** the system.

(Essentially: conservation of energy).

$$\Delta U = Q_{\text{in}} - W$$



Sign convention:
W > 0 → work done **by** the system
W < 0 → work done **on** the system

Minus sign in equation means:
U decreases if W is by system
U increases if W is on system

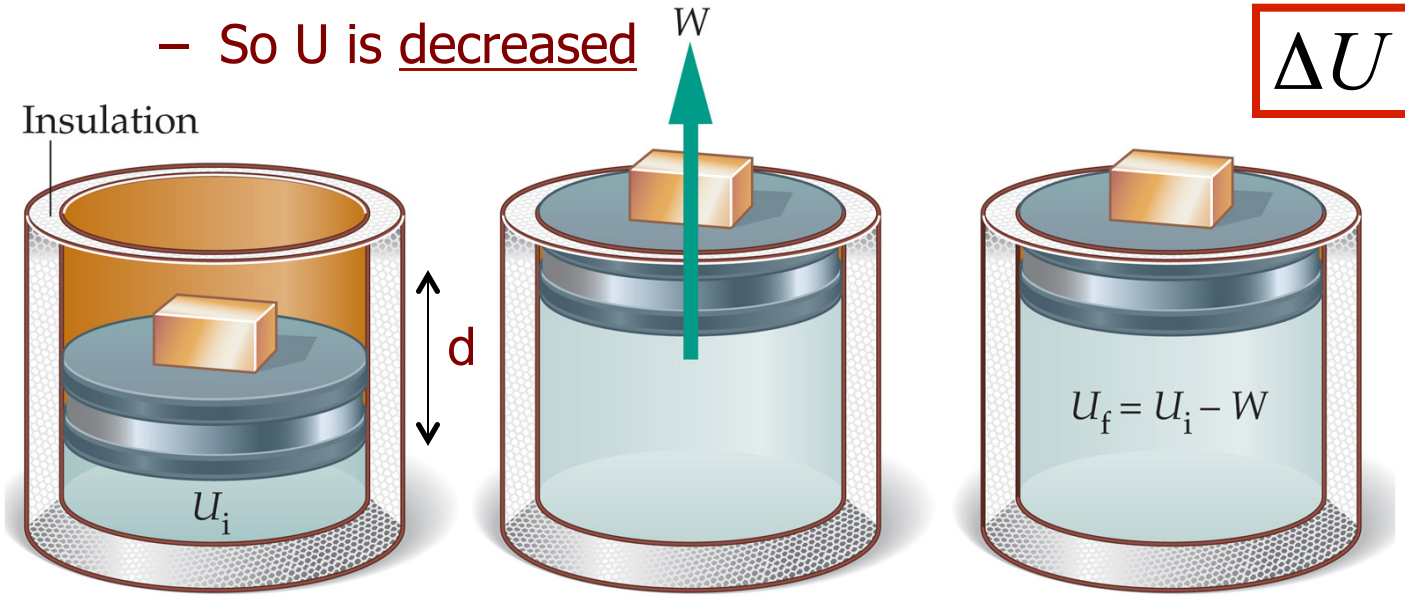
**System* could be *n* moles of ideal gas, for example...

Work done **by** the system
Example: expanding gas pushes a piston some distance
Work done **on** the system
Example: piston pushed by external force compresses gas

Example of sign convention

- Ideal gas in insulated container: no Q in or out: Q=0
- Gas expands, pushing piston up (F=mg, so W=mgd)
  - Work is done by system, so W is a positive number
  - So U is decreased

$$\Delta U = -W$$



- If instead: we add more weights to compress gas
  - Work done on gas → now W is a *negative* number
  - U is increased

$$\Delta U = -\left(-W\right) = +W$$

State of system, and state variables

- U is another thermodynamic property, like P, V, and T, used to describe the **state of the system**
  - They are connected by equations describing system behavior: for ideal gas, PV=NkT, and U=(3/2)NkT

    "equation of state"
- Q and W are not state variables: they describe changes to the state of the system
  - Adding or subtracting Q or W moves the system from one state to another: points in a {P,V,T} coordinate system
  - The system can be moved from one point to another via different sequences of intermediate states

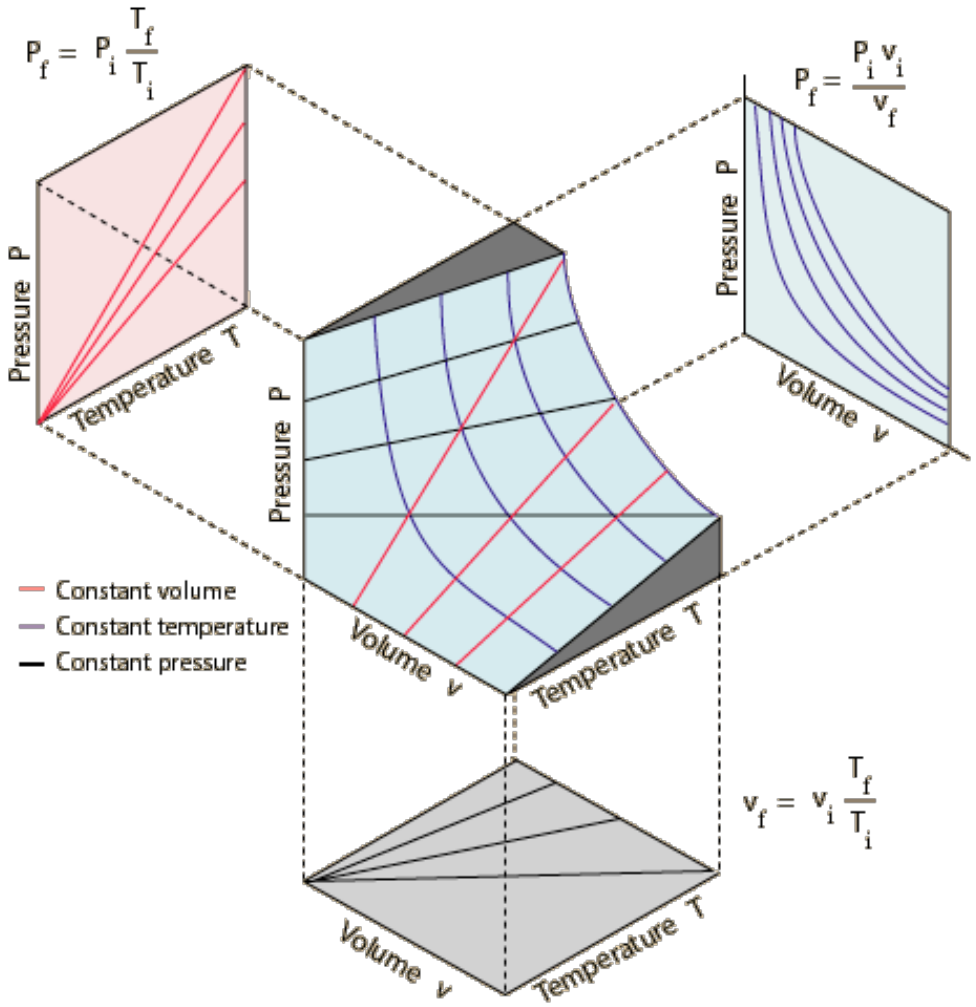
    = different paths in PVT space

    = different sequences of adding/subtracting W and Q

    = different thermodynamic processes

3D model of PVT surface

- Ideal gas law PV=NkT constrains state variables P,V,T to lie on the curved surface shown here
- Every point on the surface is a possible state of the system
- Points off the surface cannot be valid combinations of P,V,T, for an ideal gas



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Thermodynamic processes

- For ideal gas, we can describe processes that are
  - Isothermal (T=const)
  - Constant P
  - Constant V
  - Adiabatic (Q=0)
- Quasi-static processes : very slow changes
  - System is approximately in equilibrium throughout

  Example: push a piston in very small steps

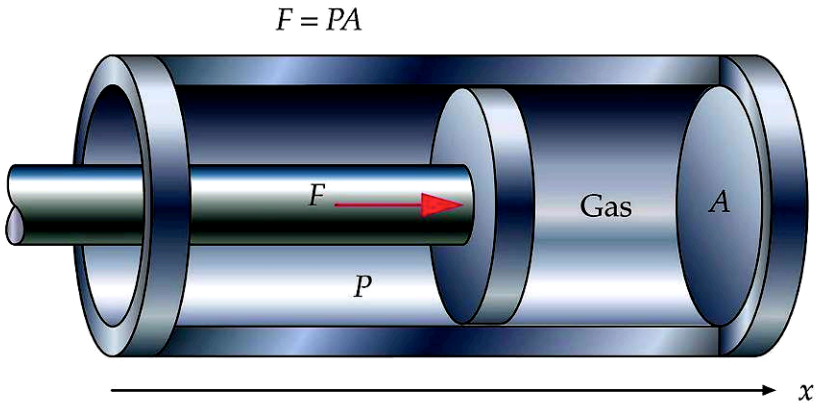
  At each step, wait to let system regain equilibrium

  "Neglect friction"

Such processes are **reversible**

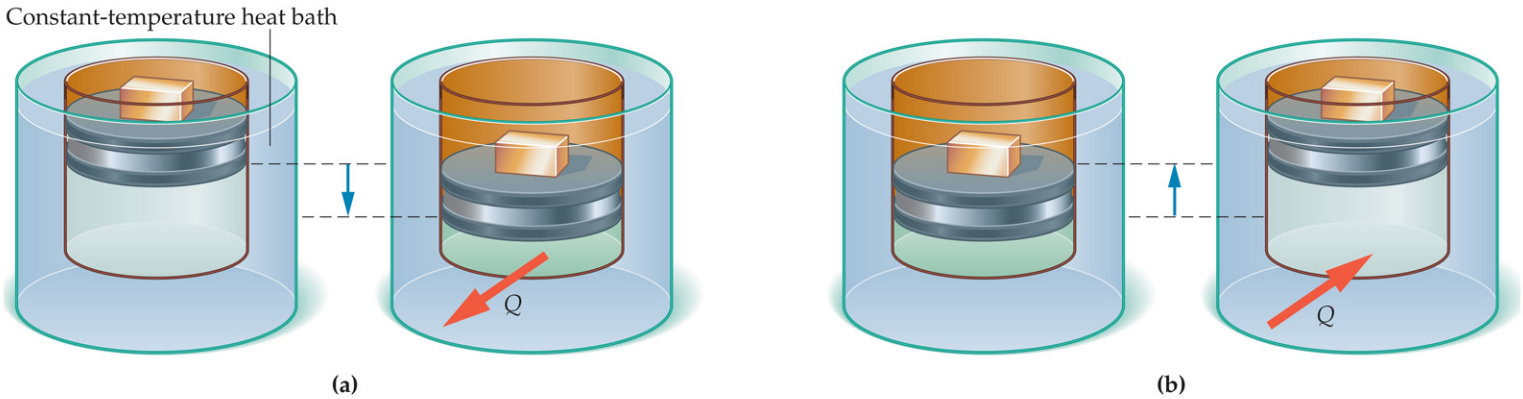
- Could run process backwards, and return to initial (P,V,T) state

Real processes are **irreversible** (due to friction, etc)



Quasi-static compression: reversible

- Ideal gas in piston, within heat reservoir
  - Compress gas slowly; heat must go out to keep T=const
    - Temperature reservoir can absorb heat without changing its T
- Reverse the process: expansion
  - Let gas pressure push piston up slowly; reservoir must supply Q to keep T=const
  - System (gas) and reservoir ("surroundings") are back to their original states
  - Reversible process for ideal gas, ideal reservoir, no friction
    - But: No real process is truly reversible (friction in piston, etc)



Process diagrams

- Example: constant P expansion of ideal gas
  - Ideal gas with $P=P_0$, at x_i
  - Frictionless piston moves to x_f
  - Volume increases from V_i to V_f
- Gas has done work on piston

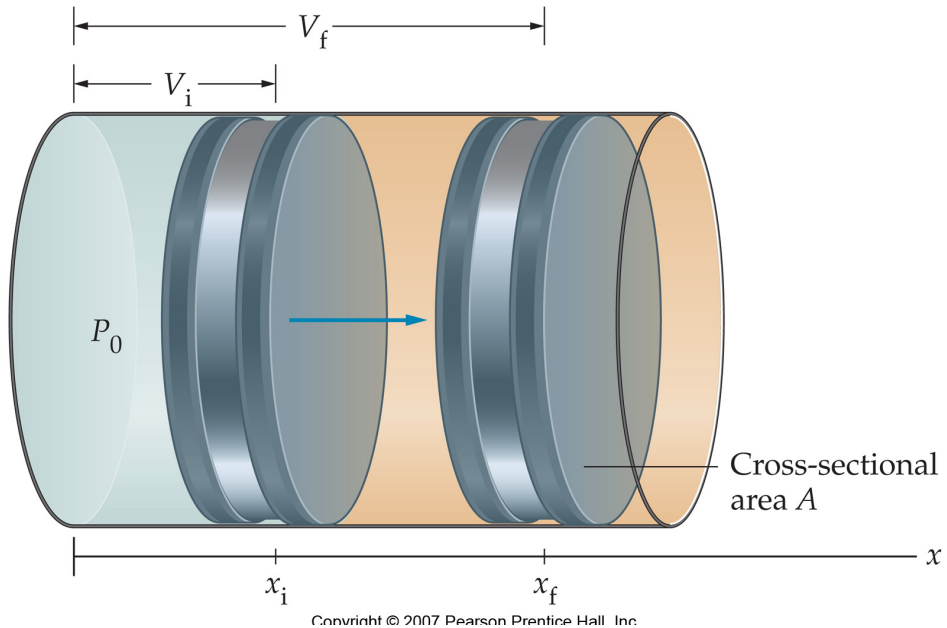
  $W = F\Delta x = (P_0\ A)\ \Delta x = P_0\ \Delta V$

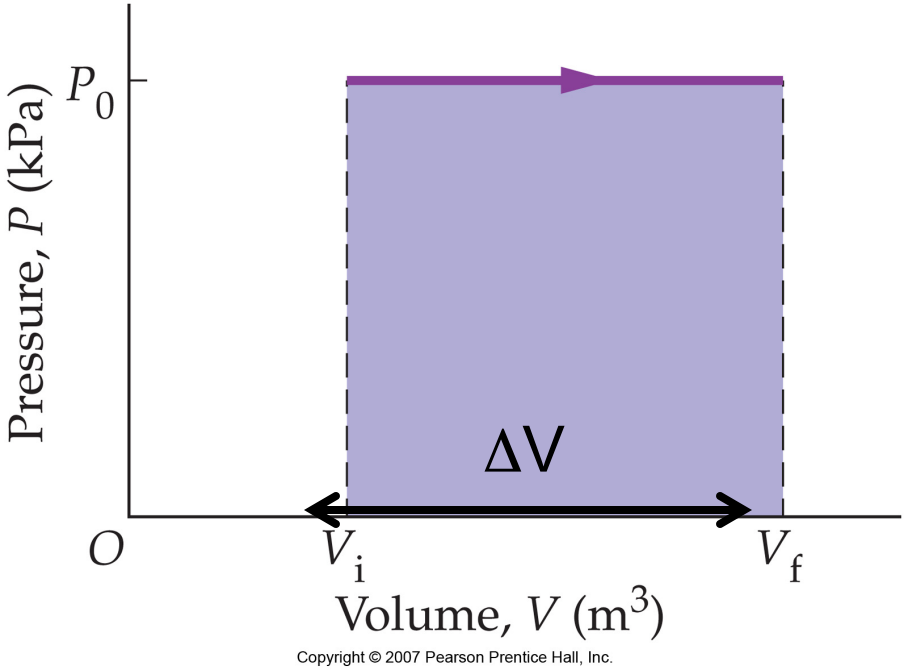
  Plot process on P vs V axes:

  Notice: $W = P_0\ \Delta V$ = **area under path** of process in the P-V plot

  This applies to any path, not just for constant-P:

  Work done = area under path on a P vs V plot





Example

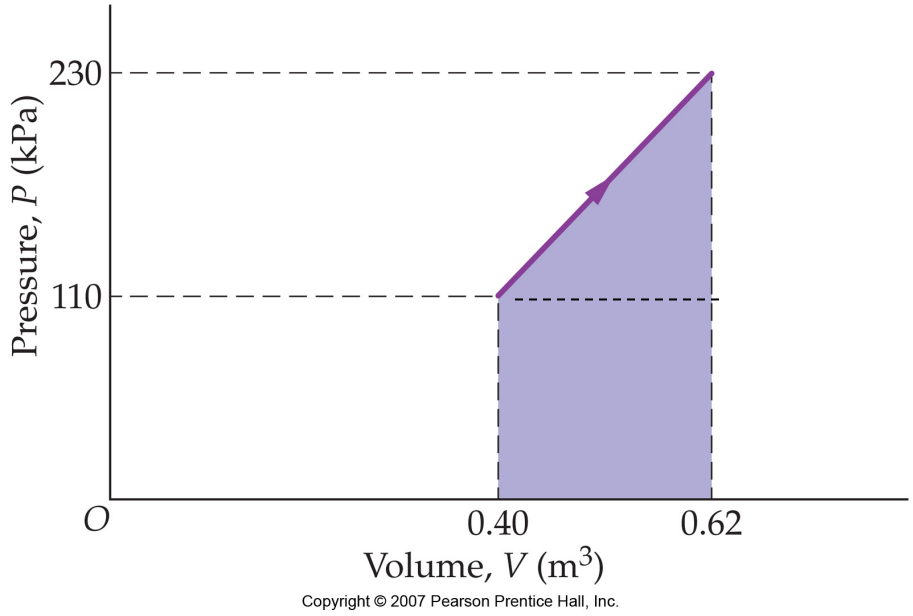
- Ideal gas expands from $V_{initial}$ =0.40 m^3 to V_{final} =0.62 m^3 while its pressure increases linearly from P_i = 110kPa to P_f = 230kPa
- Work done = area under path

Without calculus, we can calculate as Pythagoras would have done:

Area = rectangle + triangle = (110kPa)(0.22m^3)+½(120kPa)(0.22m^3)

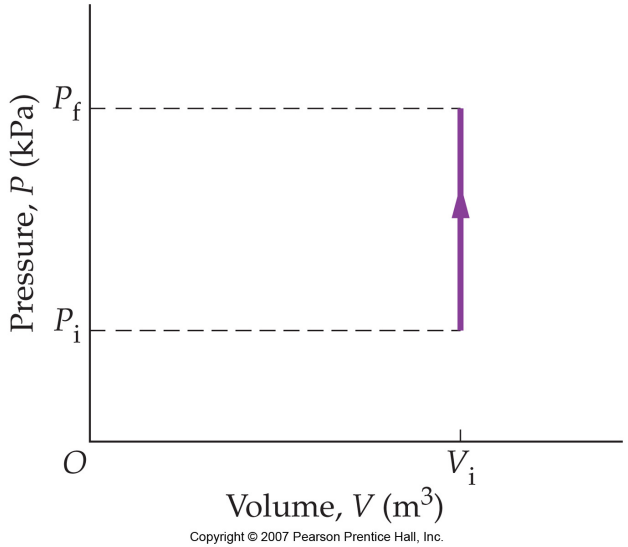
Area =W = 3.7 x 10^4 J

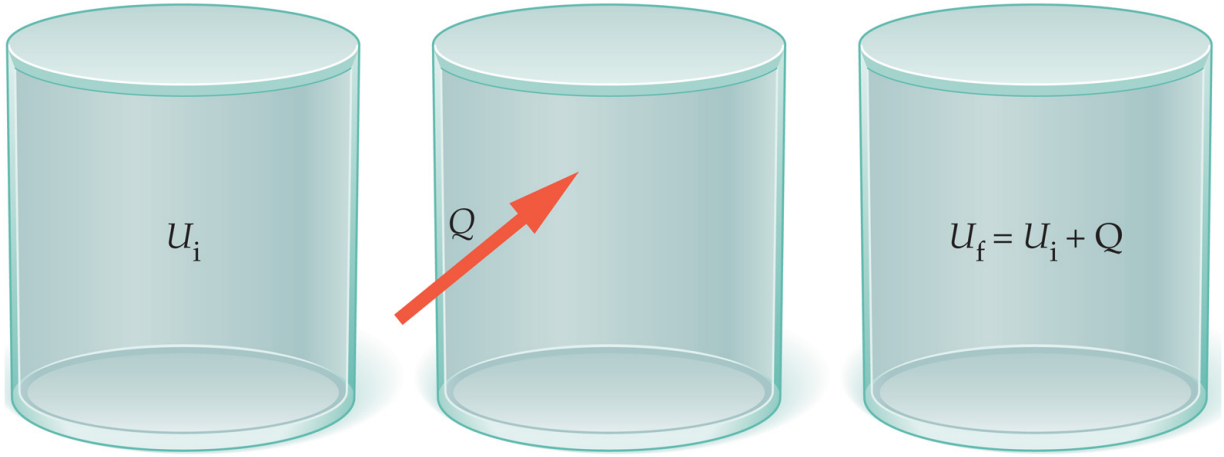
This work was done BY the expanding gas, so its internal energy is reduced



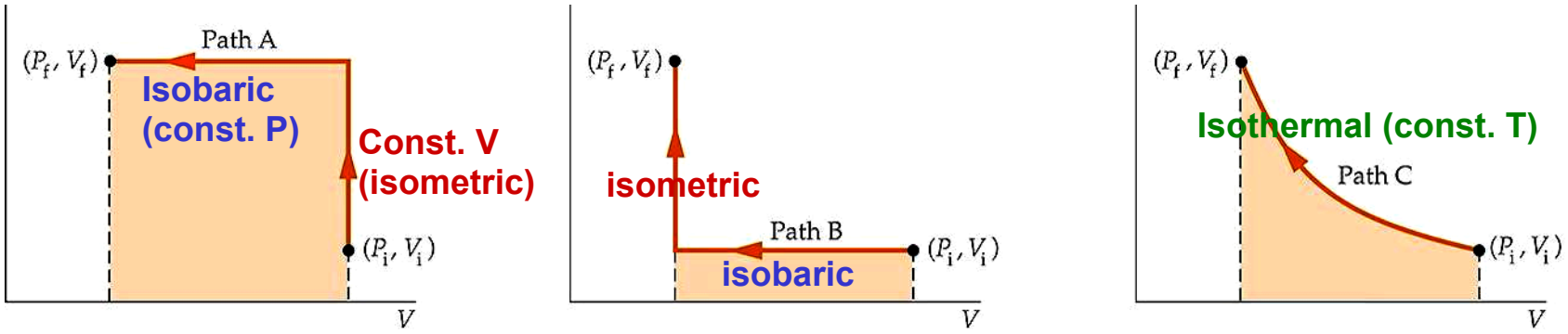
Constant volume processes

- For V=constant, P=(nR/V) T
- No work done (area=0)
- T must increase to change P
- 1st Law: $\Delta U = Q - W \Rightarrow \Delta U = Q$





Process paths in P,V plots



- Many possible paths on a P vs V diagram, for a gas moved from state (P_i , V_i) to (P_f , V_f).
- Suppose the final state has unchanged T, then

  $P_iV_i = nRT = P_fV_f$
- U depends only on T, so the initial and final internal energies U also must be equal

  Note: To keep constant T ($\Delta U=0$), heat transfer must occur, with Q = W
- Therefore, since area under {P,V} path = work done,

$$W_{\text{Path A}} > W_{\text{Path C}} > W_{\text{Path B}} \quad \text{(Is work } by \text{ or } on \text{ the gas?)}$$

Isothermal process

- If ideal gas expands but T remains constant, P must drop: P = nRT/V
  - Family of curves for each T
  - Shape is always ~ 1 / V
  - W by gas= area under plot
    - Add up slices of width ΔV

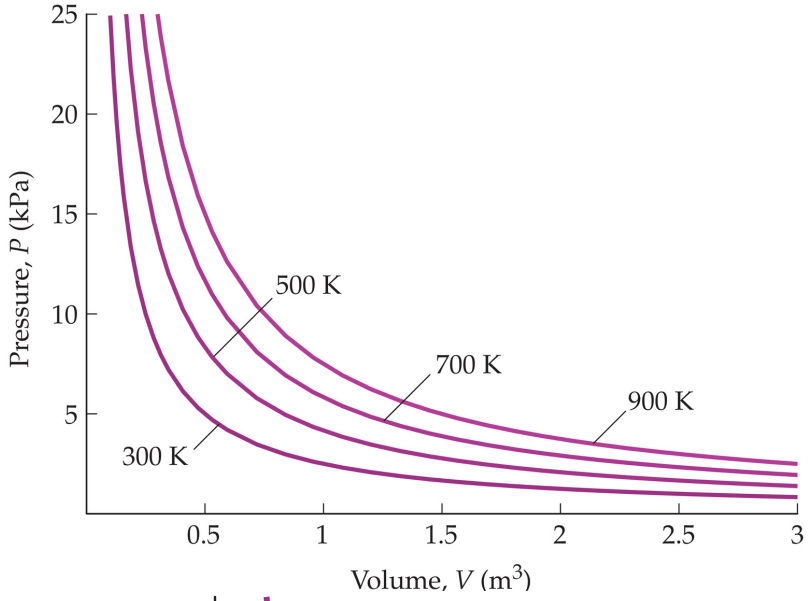
$$W \approx \sum_{V_i}^{V_f} \left(\frac{nRT}{V} \right) \Delta V$$

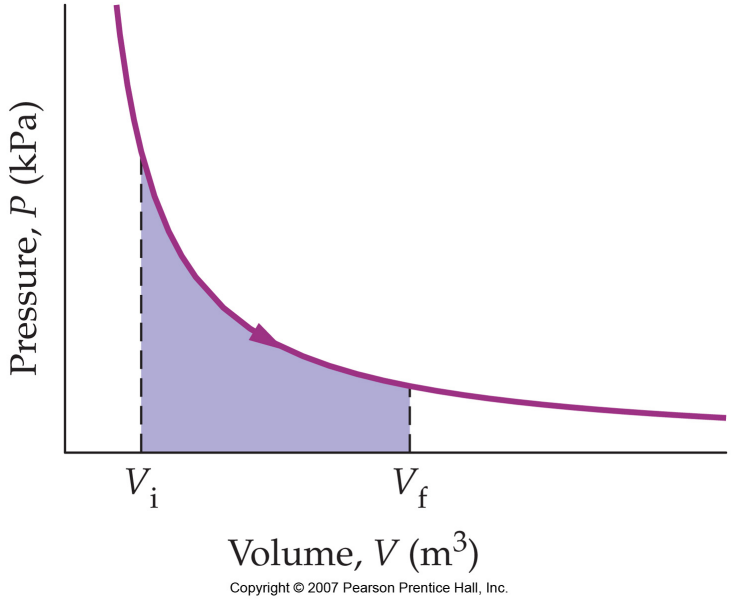
    - Use calculus, *integrate* to find area:

$$\Rightarrow \int_{V_i}^{V_f} \frac{nRT}{V} dV = nRT \ln\left(\frac{V_f}{V_i} \right)$$

"ln" = natural (base e) logarithm

$$e = \frac{1}{1} + \frac{1}{2!} + \frac{1}{3!} + \cdots = 2.72\ldots$$





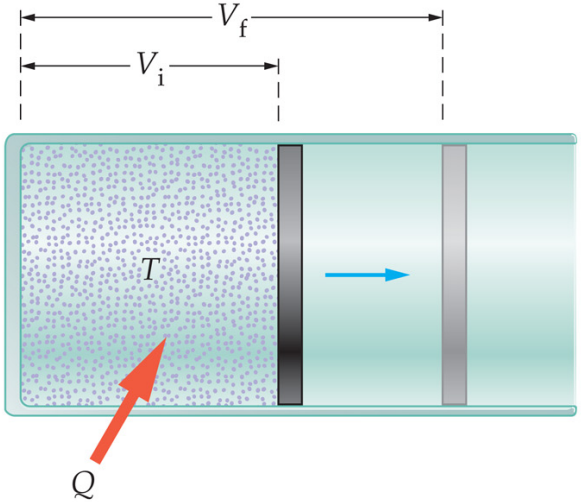
Isothermal example

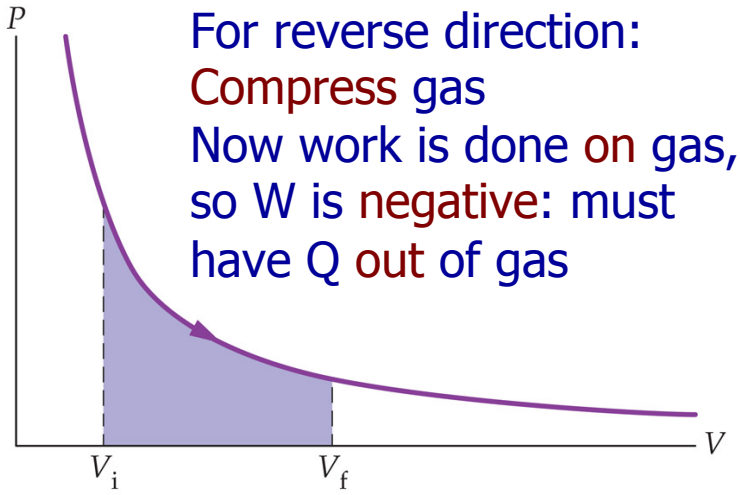
- For T = constant, U must be constant: U =(3/2) nRT
  - So Q – W = 0; W is done by gas, so Q=W
  - +Q means into system
  - We must add heat Q=W to keep T constant

$$Example: \quad n = 0.5mol, T = 310K, V_i = 0.31m^3, V_f = 0.50m^3$$

$$W = nRT\ln\left(\frac{V_f}{V_i}\right) = \left(.5mol\right)\left(8.31J/mol/K\right)\left(310K\right)\ln\left(\frac{0.45m^3}{0.31m^3}\right)$$

$$\ln\left(1.45\right) = 0.373 \Rightarrow W = 480J = Q_{IN}$$



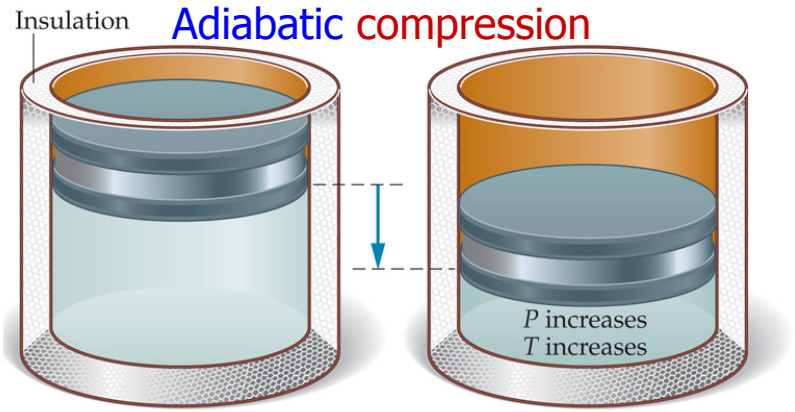


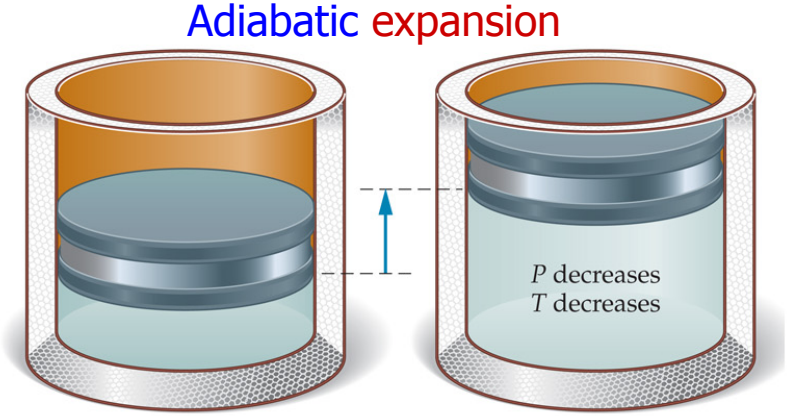
For reverse direction:
Compress gas
Now work is done on gas, so W is negative: must have Q out of gas

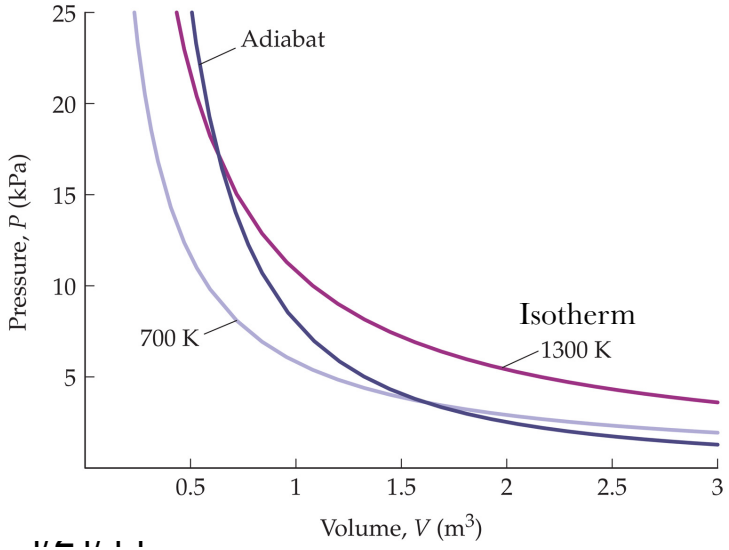
Adiabatic (Q=0) processes

- If ideal gas expands but no heat flows, work is done by gas (W>0) so U must drop: $\Delta U = Q - W \Rightarrow \Delta U = -W$

  U ~ T, so T must drop also







Adiabatic expansion path on P vs V must be steeper than isotherm path:
Temperature must drop →
State must move to a lower isotherm