# Physics 262/266

George Mason University

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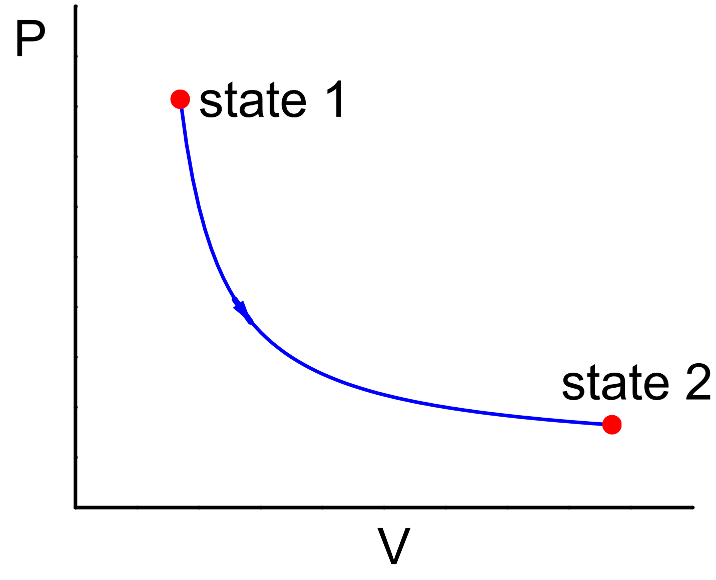
# Chapter 19: The 1st Law of Thermodynamics

- Heat, Work, and Thermodynamic Processes
- Internal Energy and the 1st Law of Thermodynamics
- Types of Thermodynamic Processes
- Adiabatic Processes
- More on Heat Capacities

# Thermodynamic Systems

- **A Thermodynamic System**: A collection of objects (considered as one "system") that may have potential to exchange energy (heat AND work) with each others and its surrounding.
- **System States & Thermodynamic Processes**:



  - **State** of a thermodynamic system is characterized by a set of marcoscopic variables (*P, V, T, n*) and, it can be visualized as a point ● in the PV diagram.
  - A thermodynamic system changes from one state (1) to another state (2) through a **thermodynamic process** indicated by the blue curve/path in the PV diagram.

# Thermodynamic Systems

- **Notes on thermodynamic processes:**
  - Different processes will have different paths on the PV graph
  - Thermodynamic *states* can only be specified if the system is in a *thermal equilibrium*!

    → Every (macroscopic) subparts of the system should have the same values of

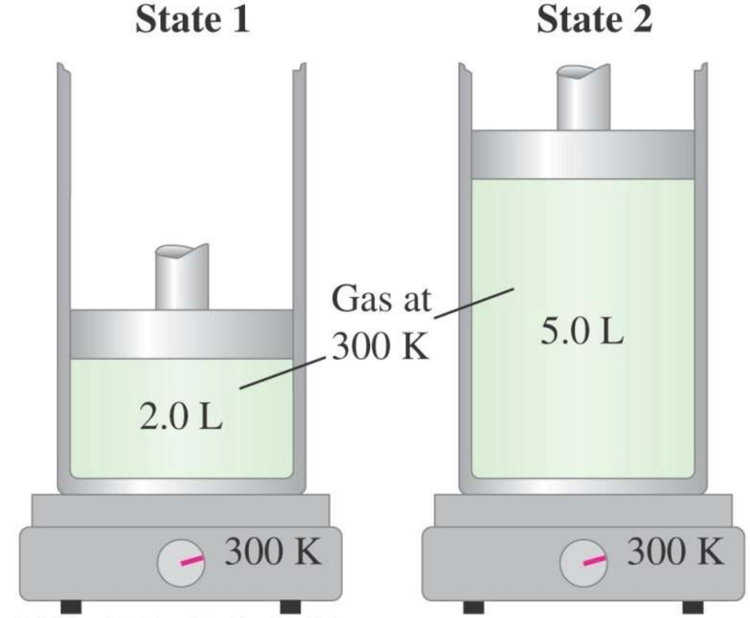
    $P, T, V, n$, etc.
  - For $n$ fixed in $PV = nRT$ in an Ideal Gas, we only need 2 out of 3 state variables $(P, V, T)$ to specify a state uniquely!

# Thermodynamic Systems

- **Notes on thermodynamic processes**:
  - A *path* for a thermodynamic process can only be represented in the PV diagram as a curve if the process is *reversible* (quasi-static).
  - A *quasi-static* process can be thought of as a sufficiently *slow* (still fast in macroscopic time) process such that the system is approximately *near equilibrium* at each step.

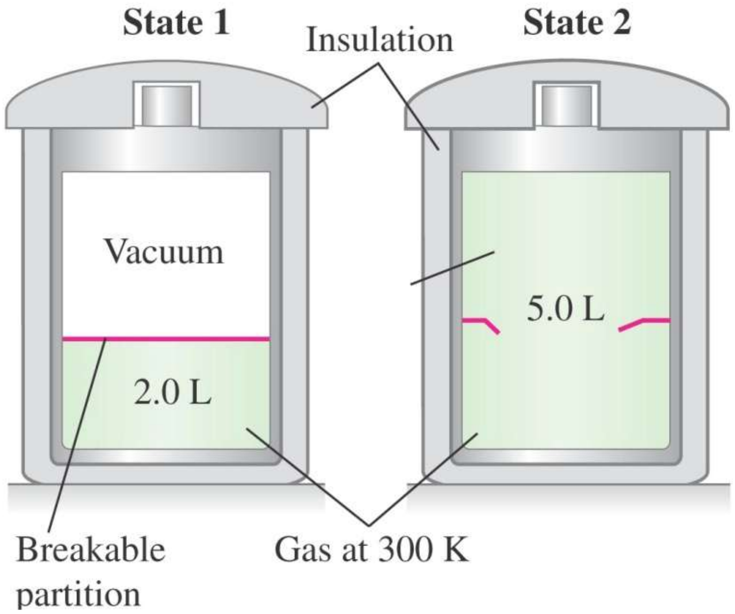
# Reversible vs. Non-Reversible Processes

## Reversible (Quasi-static)

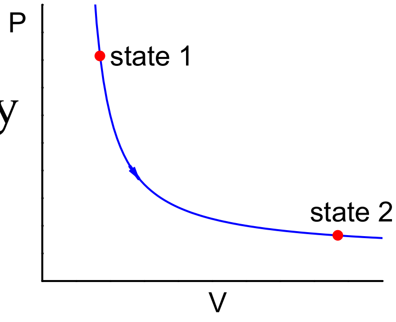


*same* initial & final states

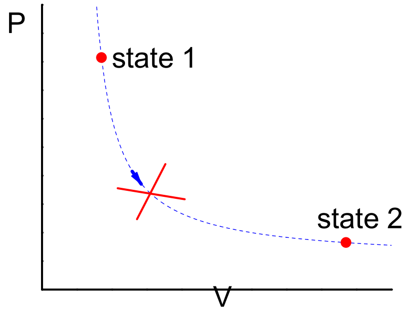
## Non-Reversible (Non-Quasi-static)



Moving the piston slightly with small heat exchange can increase/decrease the volume reversibly .



NO small changes can stop the gas in filling the container after the partition is broken.

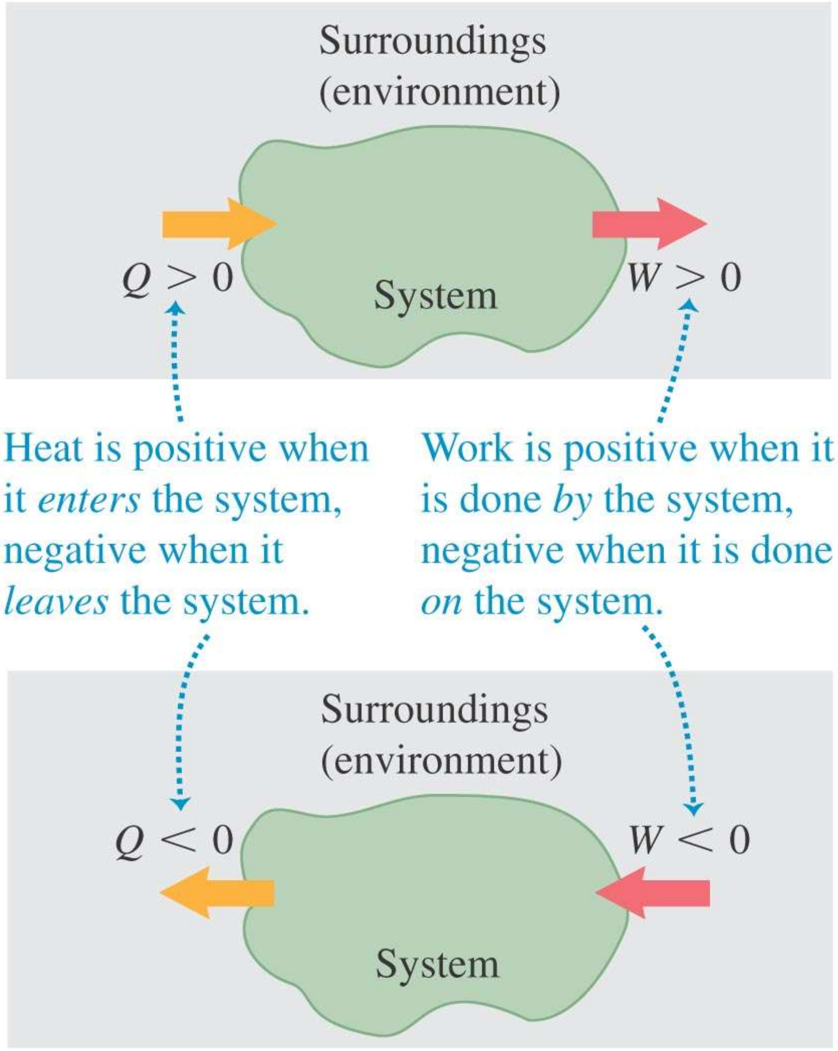


# Energy Transfer in a Thermodynamic System

**Heat** (*Q*): A thermodynamic system can *absorb* or *release* heat during a thermodynamic process.

**Work** (*W*): Work is either *done on* or *done by* a system during a thermodynamic process.

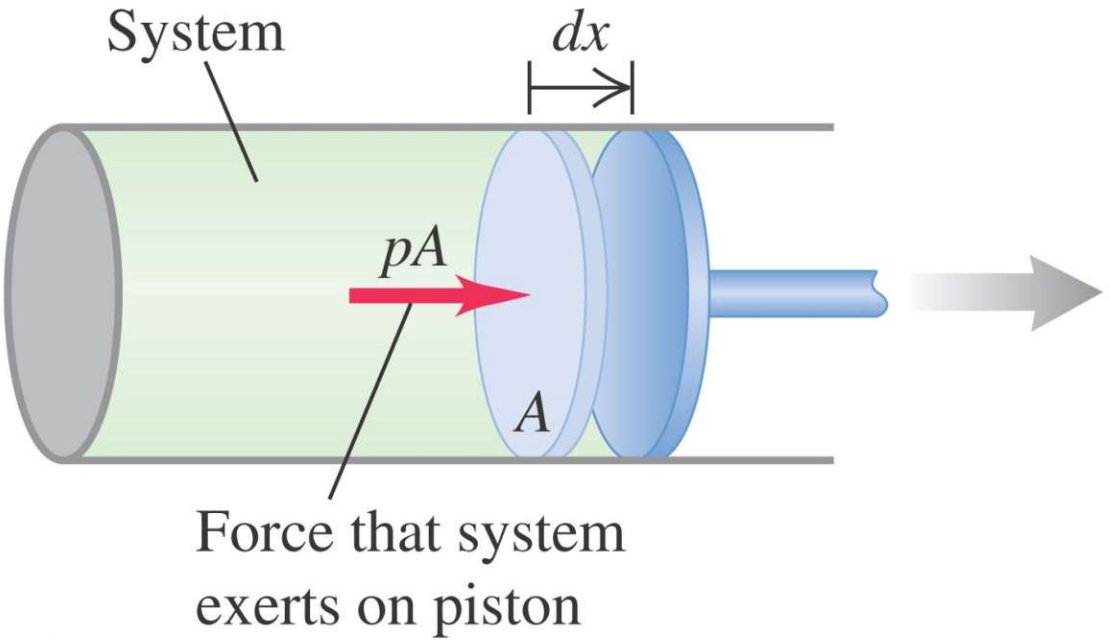
**Note**: Play attention to the sign conventions for *Q* and *W*!



Heat is positive when it *enters* the system, negative when it *leaves* the system.

Work is positive when it is done *by* the system, negative when it is done *on* the system.

# Work Done by a Gas



The system (the gas) exerts a force ($F = pA$) on its surrounding (the piston) through a distance $dx$.

The infinitesimal work done by the system $dW$ is given by:

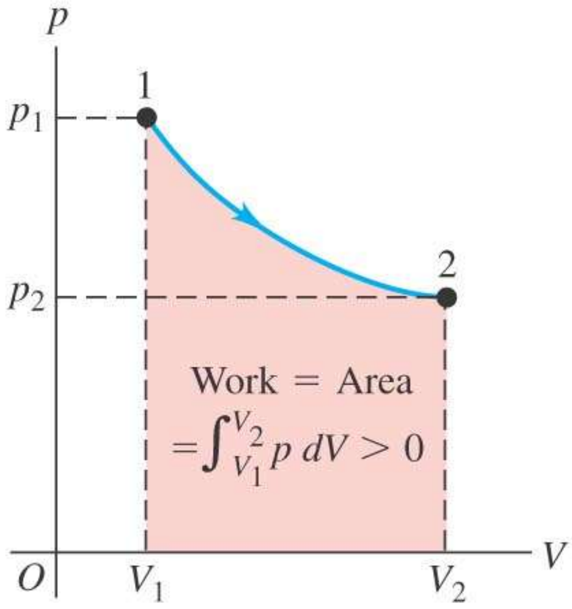
$$dW = F\,dx = PA\,dx = P\,dV$$

# Work Done by a Gas

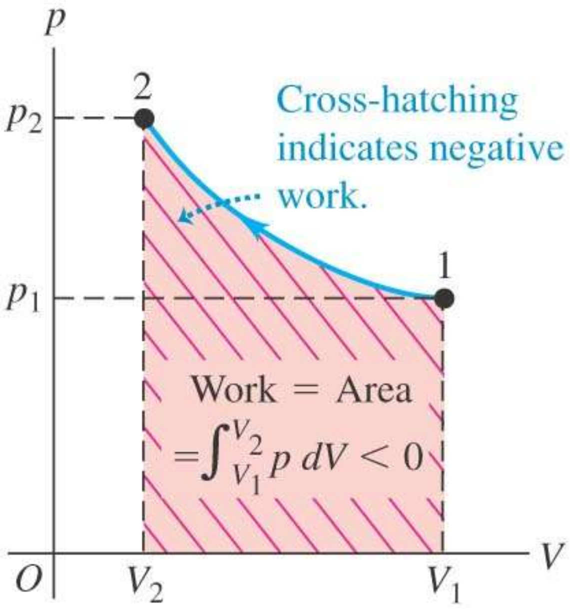
For a finite change in the volume from $V_1$ to $V_2$, the total work done is given by integrating the differential:

$$W = \int_{V_1}^{V_2} P dV$$

(a) $pV$-diagram for a system undergoing an expansion with varying pressure

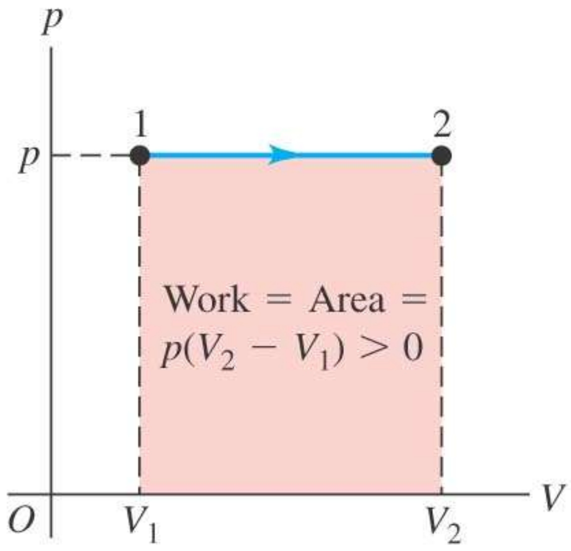


(b) $pV$-diagram for a system undergoing a compression with varying pressure



Cross-hatching indicates negative work.

(c) $pV$-diagram for a system undergoing an expansion with constant pressure



# Work Done by a Gas

Case 1: **isochoric** processes (*V* is constant, $dV=0$)

$$W = 0$$

Case 2: **isobaric** processes (*P* is constant)

$$W = P(V_2 - V_1)$$

Case 3: **isothermal** processes (*T* is constant) Example 19.1

Starting with $W = \int_{V_1}^{V_2} P\,dV$

Using the Ideal Gas law: $PV=nRT$

# Work Done by an Isothermal Process

We have, $W = \int_{V_1}^{V_2} \frac{nRT}{V} dV$

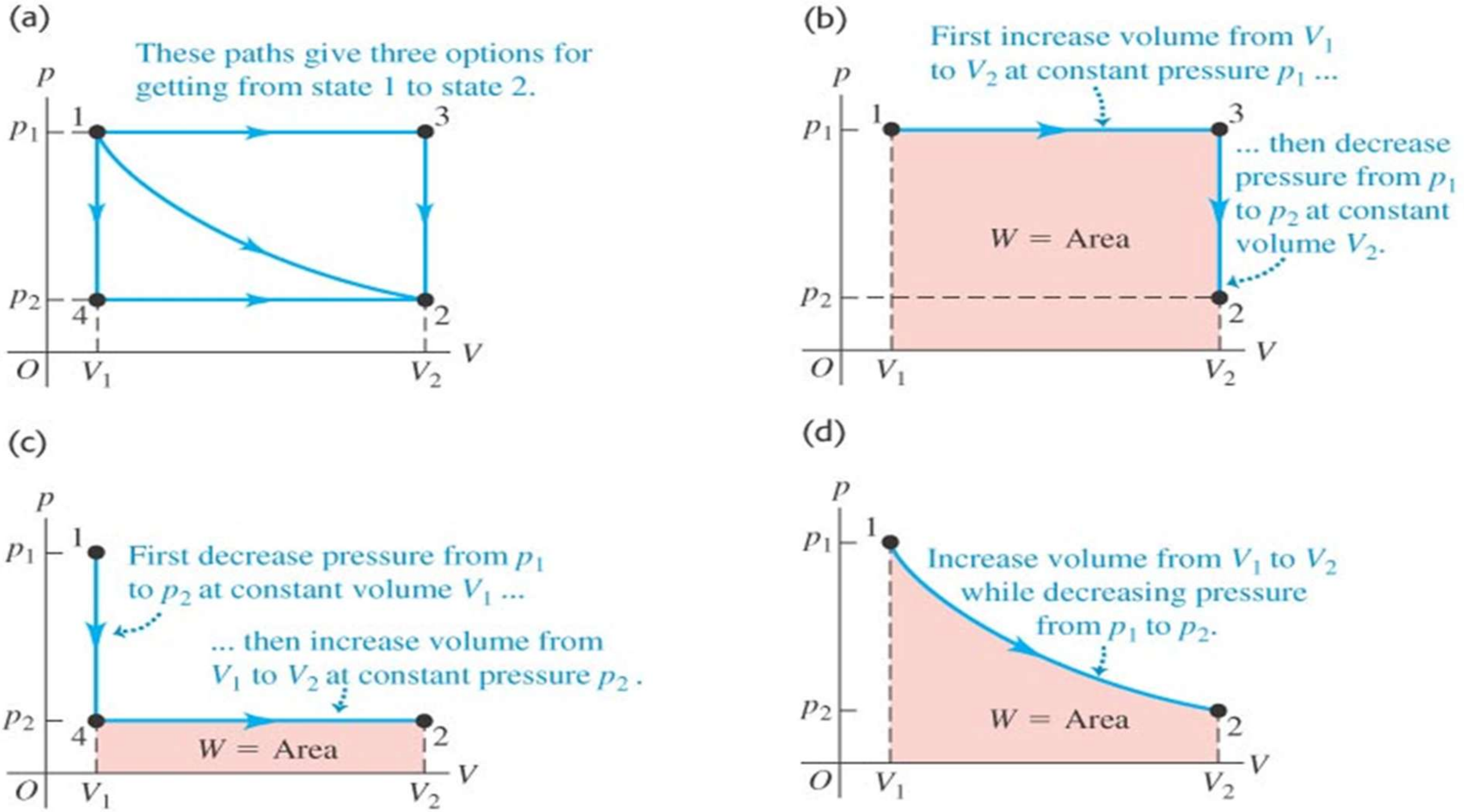
*T* is constant and we can pull it out of the integral together with the other two constants *nR* (fixed amount of gas),

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

For *T* constant, we also have $P_1 V_1 = P_2 V_2 \;\; or \;\; \frac{V_2}{V_1} = \frac{P_1}{P_2}$

So, we can also write, $W = nRT \ln\left(\frac{P_1}{P_2}\right)$

# Work Done is Path Dependent



(a) These paths give three options for getting from state 1 to state 2.

(b) First increase volume from $V_1$ to $V_2$ at constant pressure $p_1$ ... ... then decrease pressure from $p_1$ to $p_2$ at constant volume $V_2$.

(c) First decrease pressure from $p_1$ to $p_2$ at constant volume $V_1$ ... ... then increase volume from $V_1$ to $V_2$ at constant pressure $p_2$.

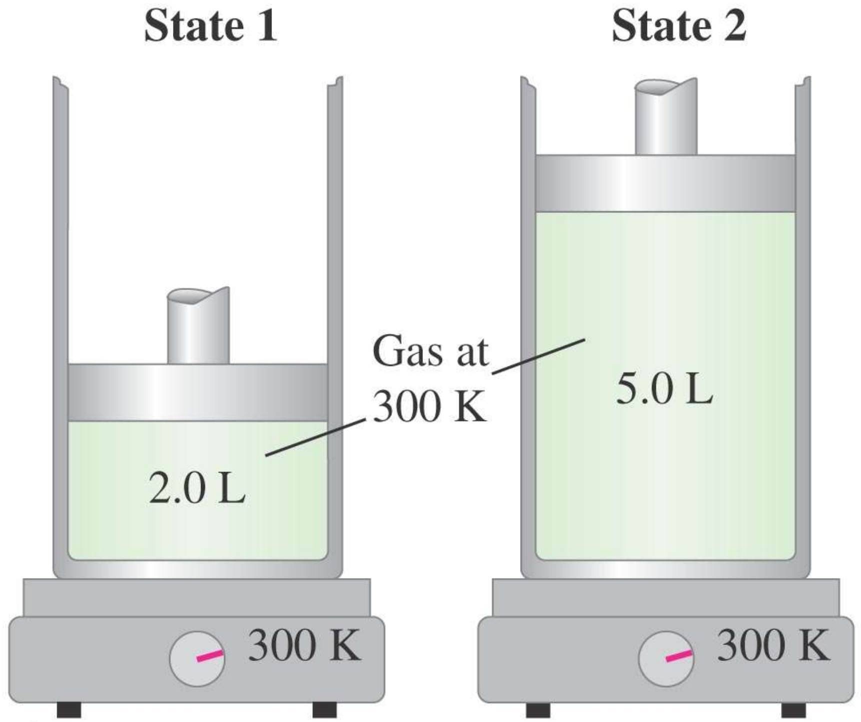
(d) Increase volume from $V_1$ to $V_2$ while decreasing pressure from $p_1$ to $p_2$.

Work done by a gas is *different* depending on path taken even though the starting and ending states are the same for all these processes.

# Heat Absorbed/Release is also Process Dependent

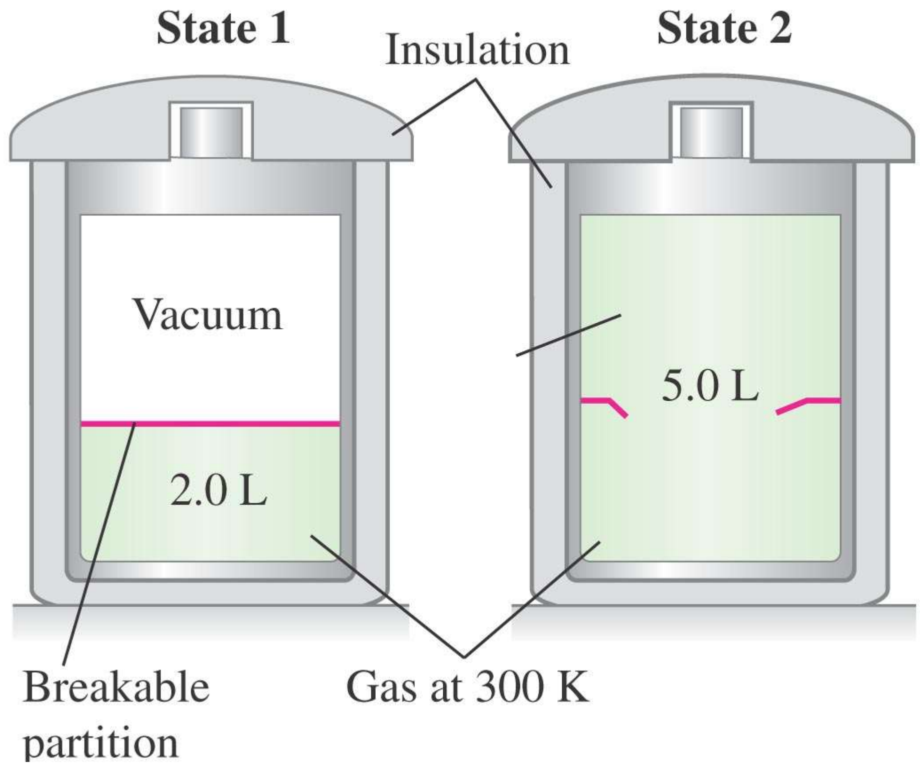
## Process 1 (isothermal expansion)

System does work on piston; hot plate adds heat to system ($W > 0$ and $Q > 0$).



## Process 2 (free expansion)

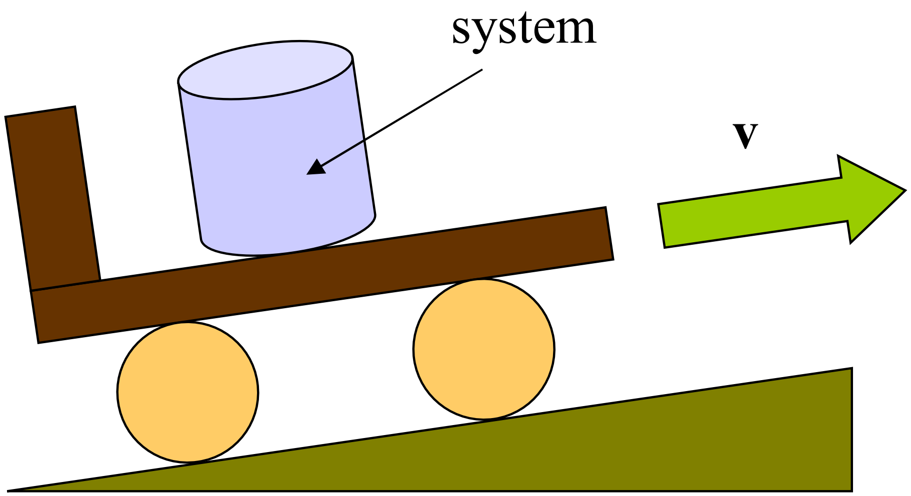
System does no work; no heat enters or leaves system ($W = 0$ and $Q = 0$).



Starting and ending states are the same for *both* processes!

# Internal Energy ($U$)

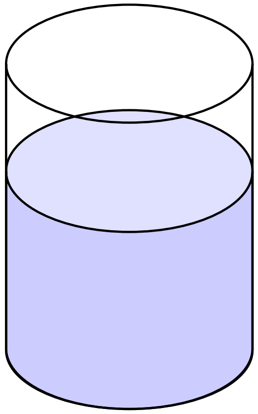
**Internal Energy** ($U$): the total amount of energy (KE + PE) *intrinsic* to the system associated with all its microscopic components when viewed in a reference frame at rest with the object.



The mechanical KE and PE associated with the motion of the *entire* system are NOT parts of the internal energy $U$ of the system.

# Internal Energy (example)

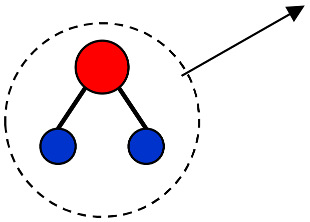
A bucket of water



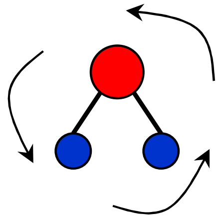
@ rest on the table

KEs:

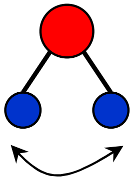
translational:



rotational:

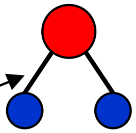


vibrational:



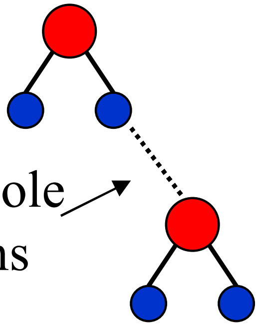
PEs:

within molecules:



molecular
bonds

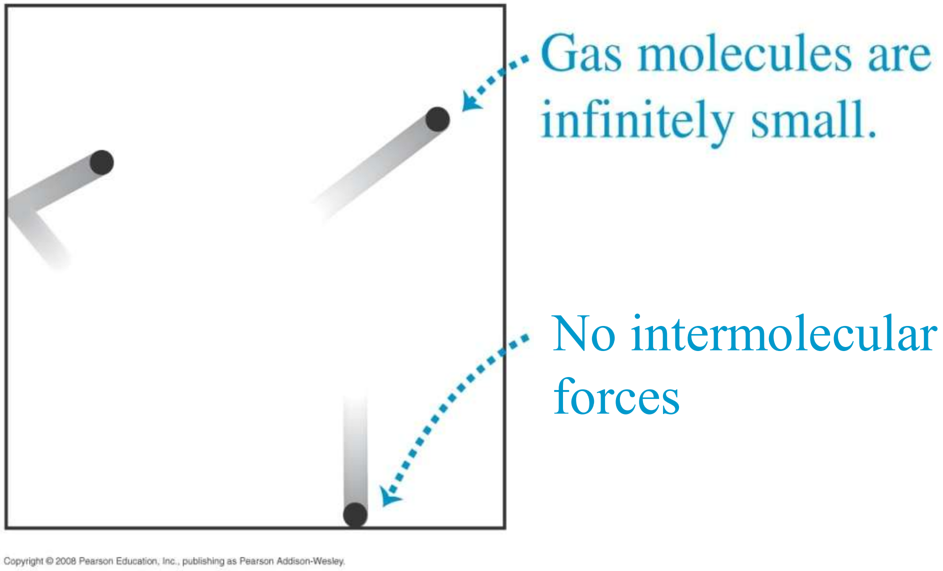
between molecules:



dipole-dipole
interactions

# Internal Energy for an Ideal Gas (monatomic)

Recall in an Ideal Gas…



We can explicitly calculated its Internal Energy *U*:

3 translational degrees of freedom only!

→ *N* molecules

→ each contributes ½ *kT*

$$U = 3N\left(\frac{1}{2}kT\right) = \frac{3}{2}NkT$$

(Ideal Gas monatomic)

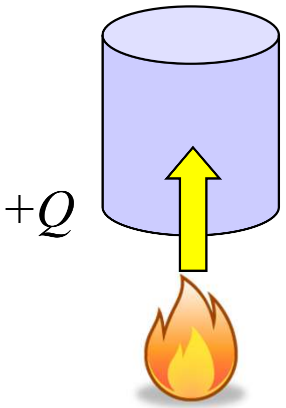
The Internal Energy for an Ideal Gas *U* is a function of *T* ONLY!

(Experimentally, this is also shown to be the case for most diluted real gases!)

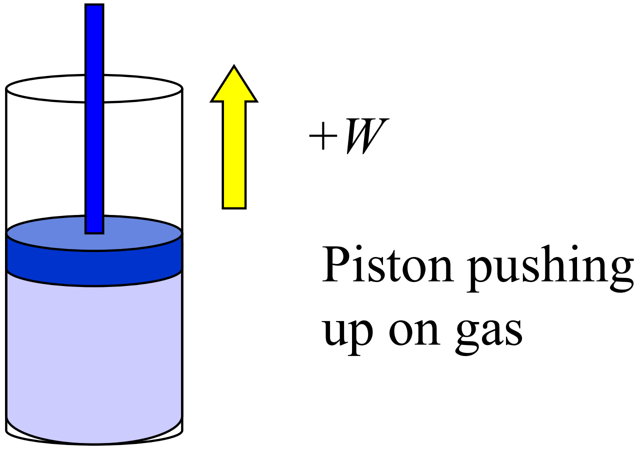
# Energy Exchanges in a Thermal System

We have seen that both heat and work can change the internal energy of the system. In particular,

Heat



Work



Piston pushing up on gas

*Heat* enters the system,
$U$ increases ($\Delta U > 0$)

*Work* done by the system,
$U$ decreases ($\Delta U < 0$)

# The 1st Law of Thermodynamics

Putting these two mechanisms together, we can combine them into one mathematical statement for the *change* in the **internal energy *U*** (total energy) in any processes,

$$\Delta U = Q - W \qquad \text{(1}^{\text{st}}\text{ Law of Therm.)}$$

This is a generalization of the principle of conservation of energy to include energy transfer through heat as well as through mechanical means (work).

# The $1^{st}$ Law of Thermodynamics

## Important Property of *U*

Although both *Q* and *W* are path/process dependent…

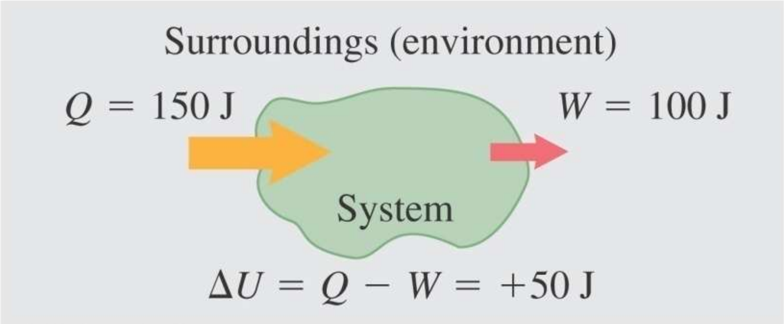
→ *U* is the **total** internal energy of the system. It is a **state variable** and it depends on the initial and final states only. Specially, for an Ideal Gas, *U* is a function of *T* only.

→ $\Delta U$ is *independent of path* since $\Delta U$ represents the change in the total internal energy of the system

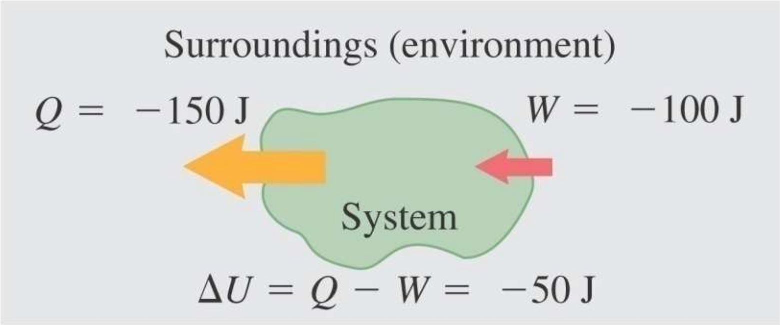
# 1st Law of Thermodynamics

$\Delta U$ can be positive, negative, and zero.

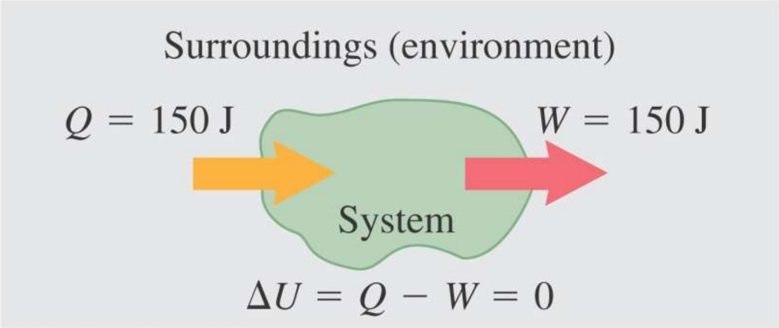
More heat is added to system than system does work: Internal energy of system increases.



More heat flows out of system than work is done: Internal energy of system decreases.



Heat added to system equals work done by system: Internal energy of system unchanged.



# Special Processes

1. Isolated Systems (no interactions with surrounding)

   Since $Q = W = 0$,

$$\Delta U = 0$$

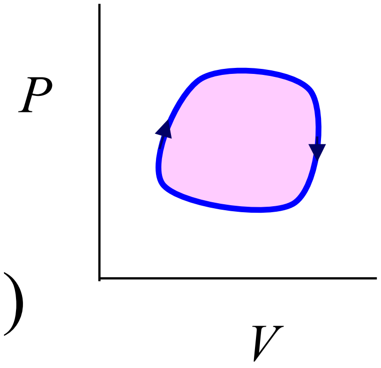
$U$ remains a constant in an isolated system.

2. Cyclic Processes (starting state = end state)

   Since $U$ is a *state variable*, $\Delta U = 0$.

   1st Law gives, $\Delta U = Q - W = 0$

$$Q = W$$ (= area encircled by curve)



# Special Processes

3. Isochoric (constant *V*):

NO volume change → NO work (*W*=0)

$$\Delta U = Q - W \xrightarrow{W=0} \boxed{\Delta U = Q}$$

4. Isobaric (constant *P*)

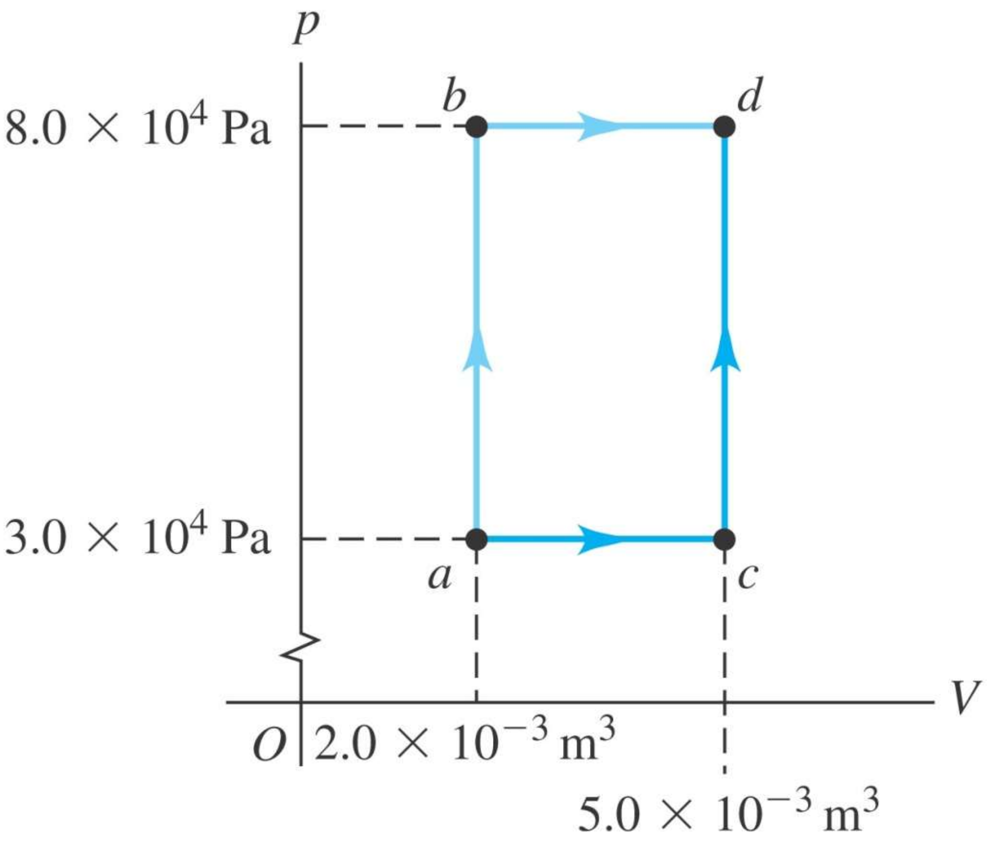
Since *P* is a constant, $dW = PdV$ can be integrated easily and gives,

$$W = P(V_2 - V_1)$$

$$\boxed{\Delta U = Q - P\left(V_2 - V_1\right)}$$

# Example: A Thermodynamical Cycle

- Example 19.4(2)



Given:

$Q_{ab} = 150$ J

$Q_{bd} = 600$ J

note

Find:

$\Delta U_{ab} = ?$

$\Delta U_{abd} = ?$

$Q_{acd} = ?$

http://complex.gmu.edu/www-phys/phys262/soln/ex19.4.pdf

# Special Processes

5. Isothermal (constant *T*)

For an Ideal Gas, *U*(*T*) depends only on *T* → *ΔU = 0 !*

$$\Delta U = Q - W \xrightarrow{\Delta U = 0} \boxed{Q = W}$$

Recall, work done by an isothermal process in an Ideal Gas:

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

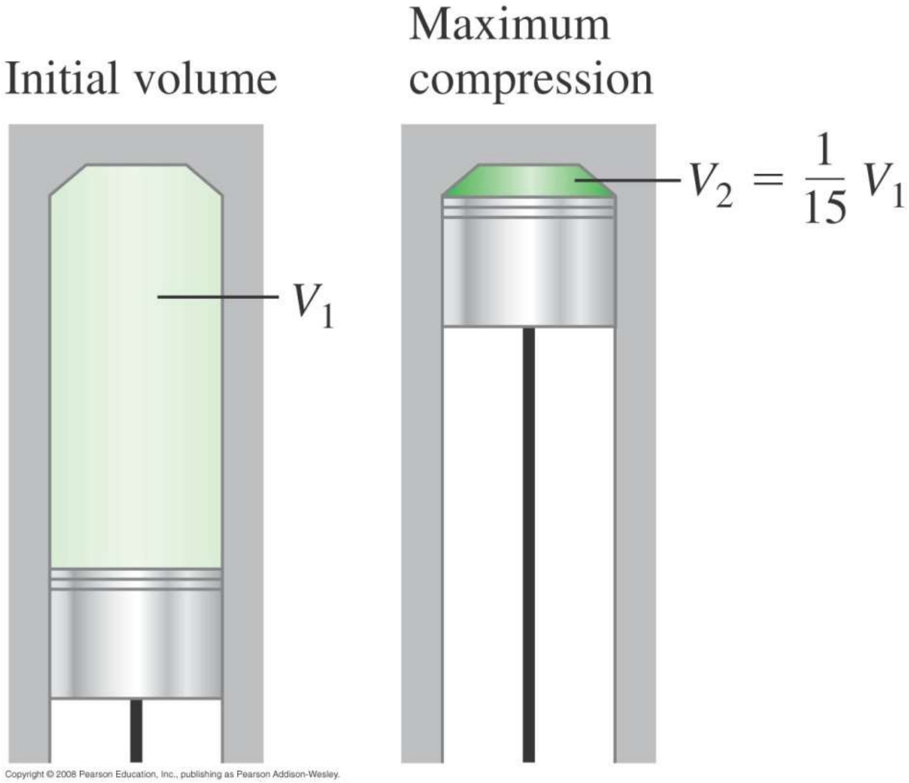
# Special Processes

6. Adiabatic (no heat exchange, $Q = 0$):

$$\Delta U = Q - W \xrightarrow{Q=0} \boxed{\Delta U = -W}$$

Note: The compression stroke in an internal combustion engine is quick and it can be well approximated by an *adiabatic* process.

(will come back to this example later)

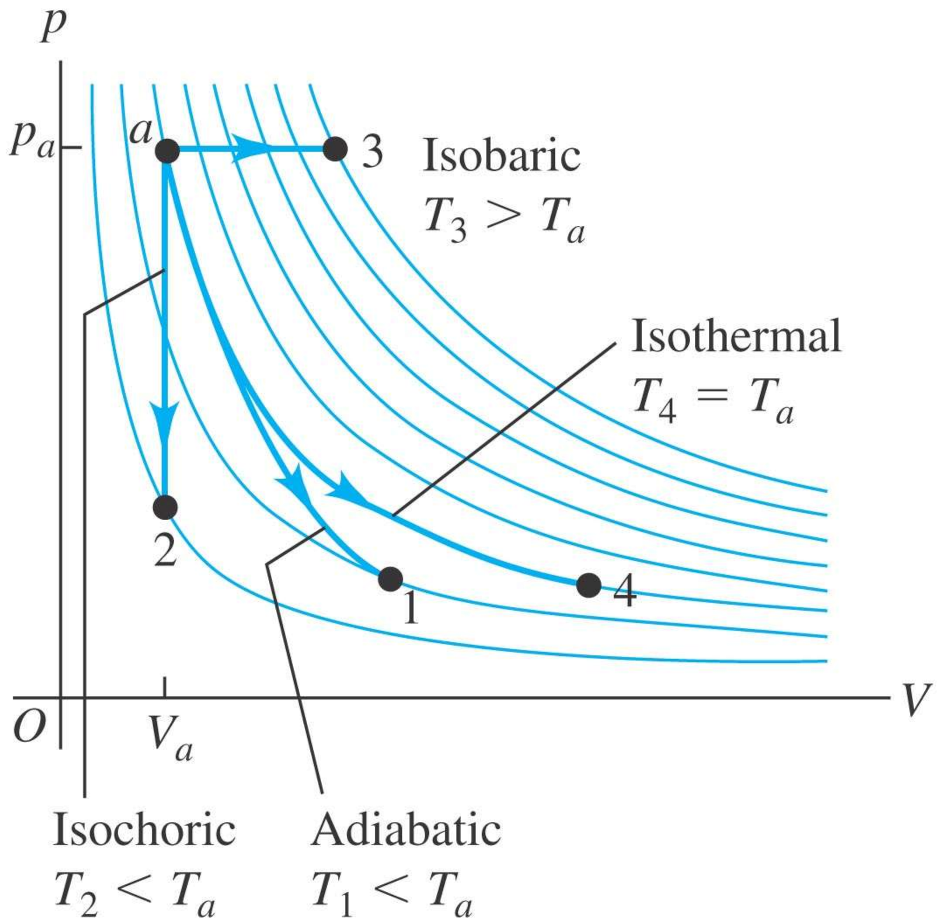


D fire piston

# Summary

Four Reversible (Quasi-static) processes:

1. Adiabatic ($Q$=0)
2. Isochroic ($\Delta V$=0)
3. Isobaric ($\Delta P$=0)
4. Isothermal ($\Delta T$=0)



1st Law for infinitesimal changes: $dU = dQ - dW = dQ - PdV$