Physics 262/266

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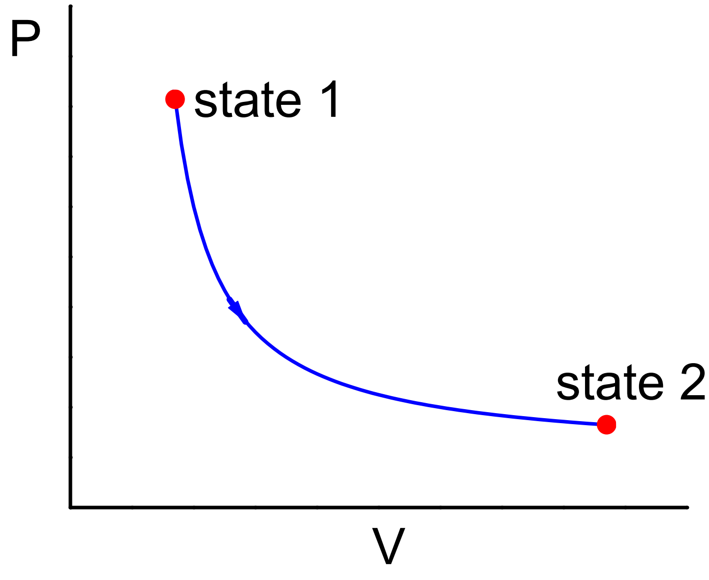
Chapter 19: The 1st Law of Thermodynamics

- 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

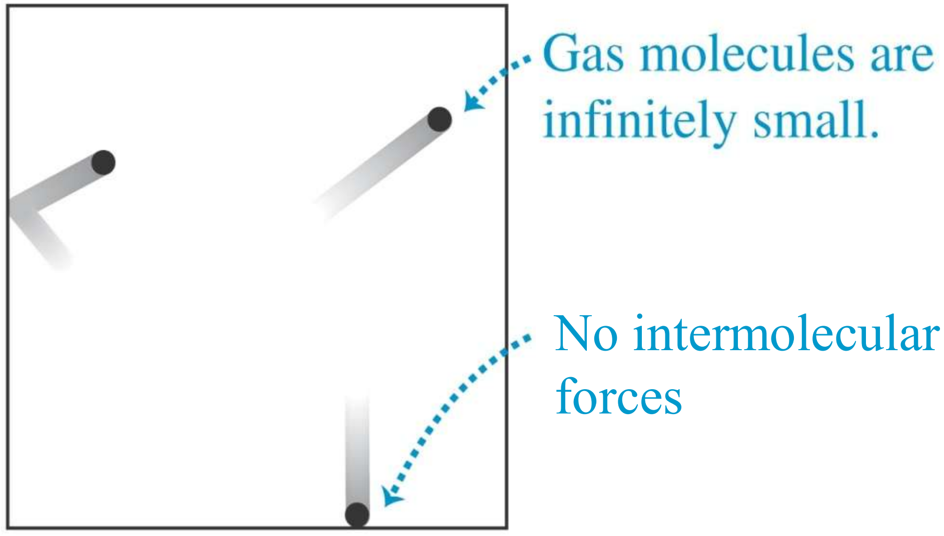
- **A Thermodynamic System**: A collection of objects (considered as one "system") that may have potential to exchange energy (heat AND work) with each others and its surrounding.
- **System States & Thermodynamic Processes**:



- **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.

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.



3 translational degrees of freedom only!

→ *N* molecules

→ each contributes ½ *kT*

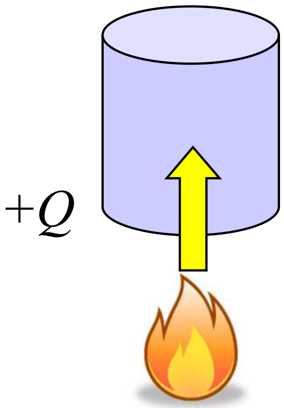
$$U = 3N\left(\frac{1}{2}kT\right) = \frac{3}{2}NkT$$

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

Energy Exchanges in a Thermal System

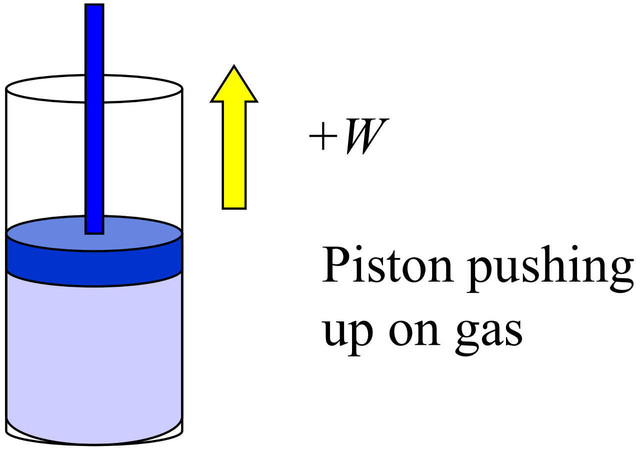
Both heat and work can change the internal energy of the system. In particular,

Heat



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

Work



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,

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

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…

→ The **total** internal energy of the system *U* is a **state variable** and,

→ A change in U, i.e., *ΔU, is independent of path/process*. U depends on the system's state only. Specifically, for an Ideal Gas, *U* is a function of *T* only.

Simple Processes

1. Isolated Systems (no interactions with surrounding)

   Since $Q = W = 0$,

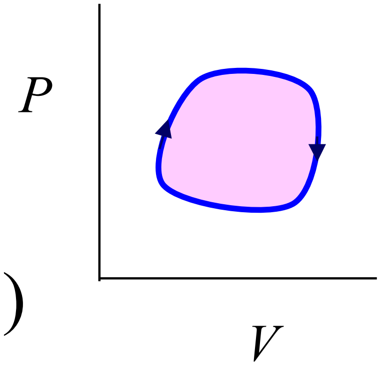
$$\Delta U = 0$$

U remains a constant in an isolated system.

2. Cyclic Processes (starting state = end state)

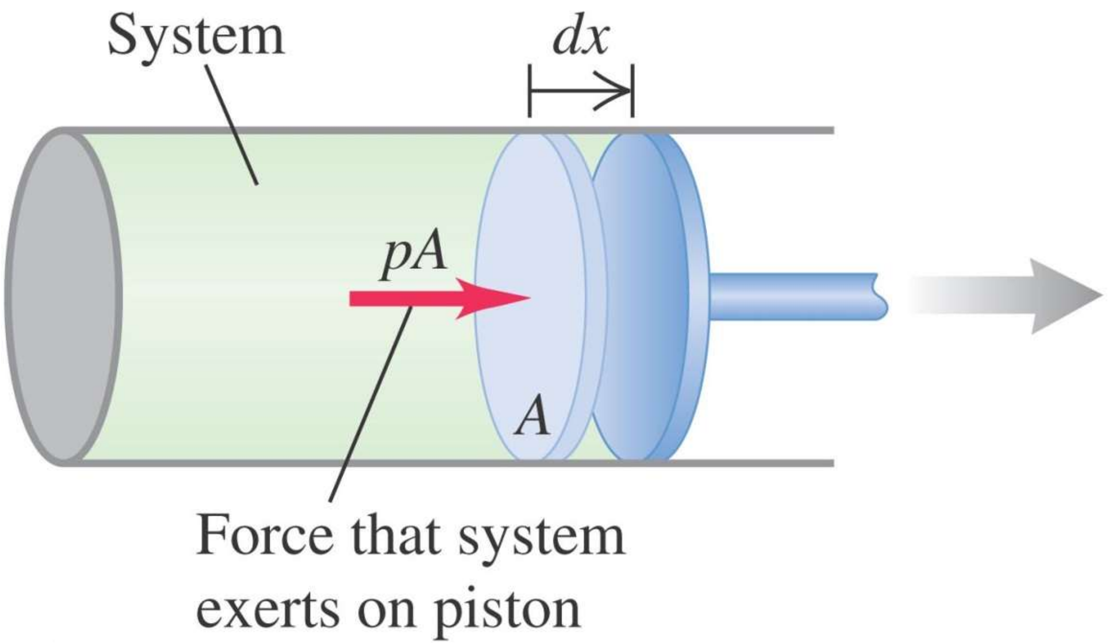
   Since *U* is a *state variable*, $\Delta U = 0$.

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



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

Work Done by a Gas

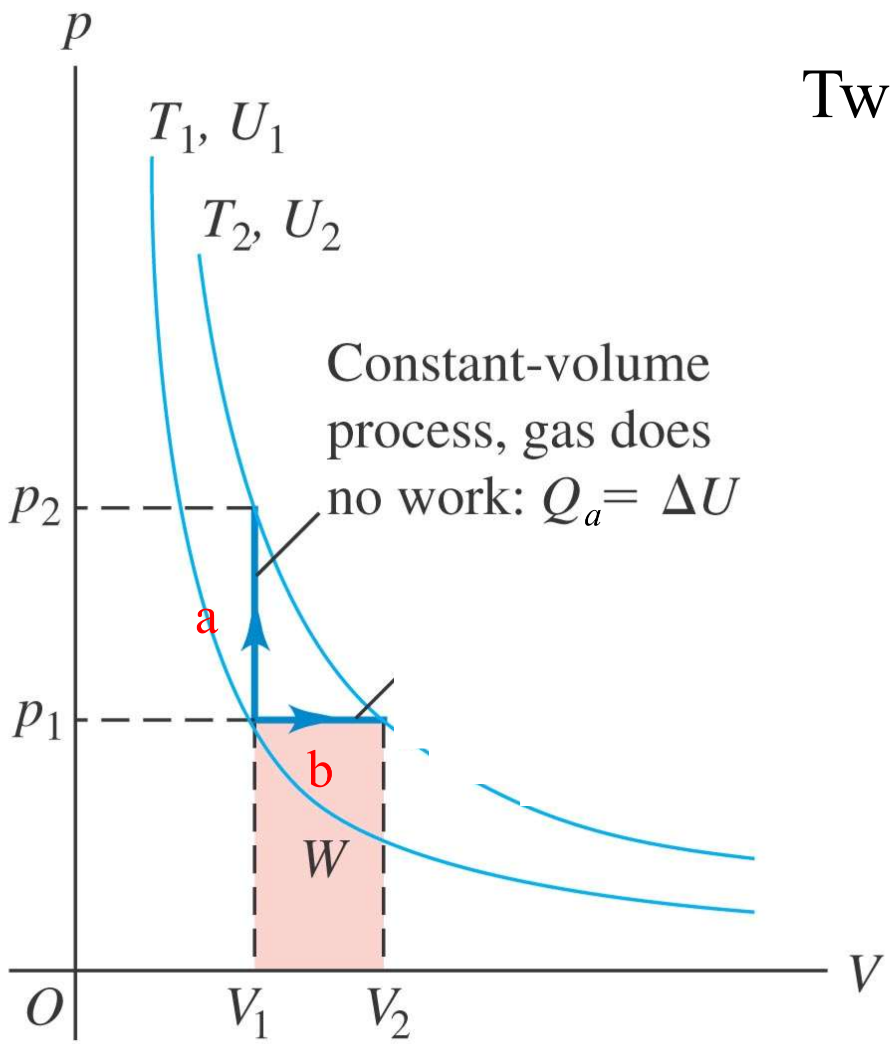


The system (the gas) exerts a force ($F = pA$) on its surrounding (the piston) through a distance dx.

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

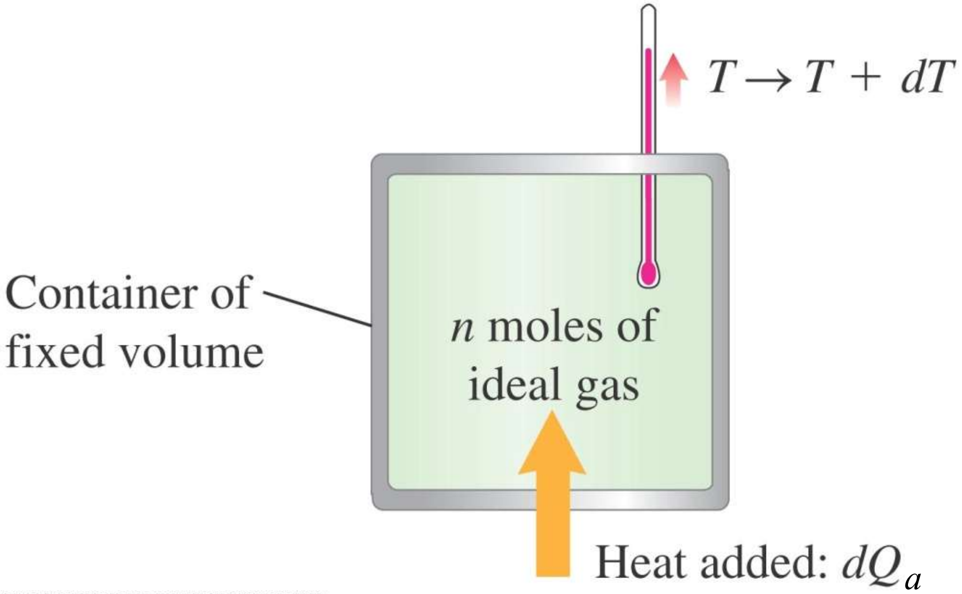
$$W = \int_{V_1}^{V_2} P\,dV$$

Heat Exchanges during Constant *P* & Constant *V* Processes



Two different ways to change $dT = T_2 - T_1$:

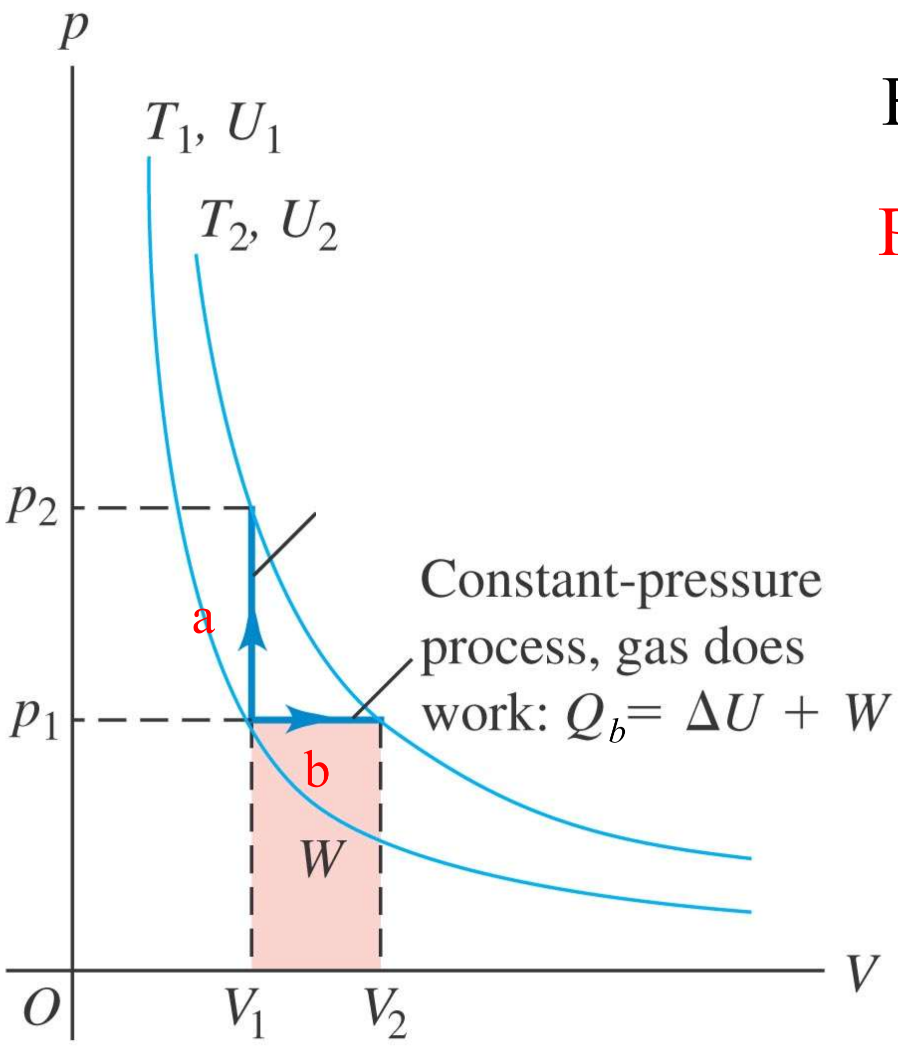
Process a: Constant *V*



$$dQ_a = nC_v dT$$

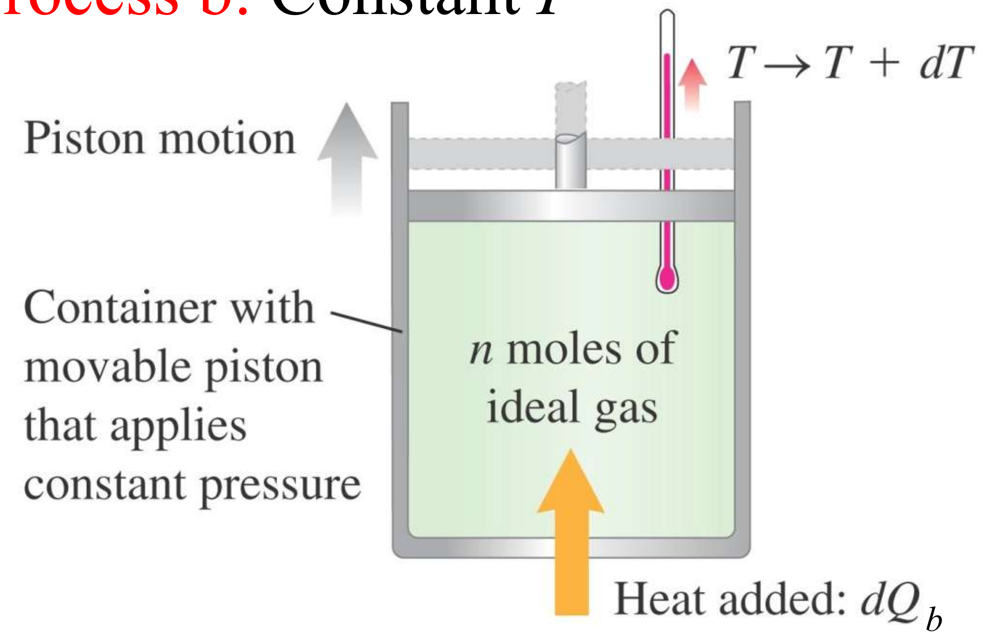
Molar Specific Heat with Constant *V*: C_v

Heat Exchanges during Constant *P* & Constant *V* Processes



For the same dT :

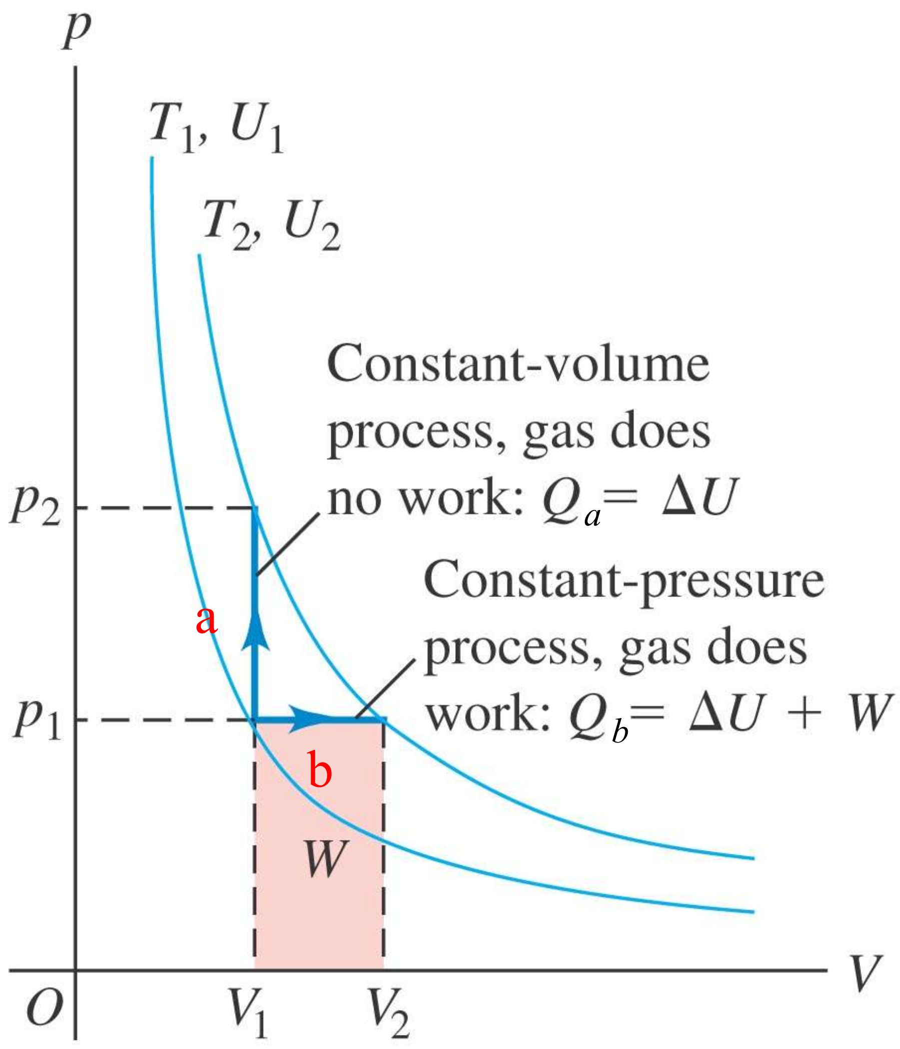
Process b: Constant *P*



$$dQ_b = nC_P dT$$

Molar Specific Heat with Constant *P*: C_P

Heat Exchanges during Constant *P* & Constant *V* Processes



For the same dT:

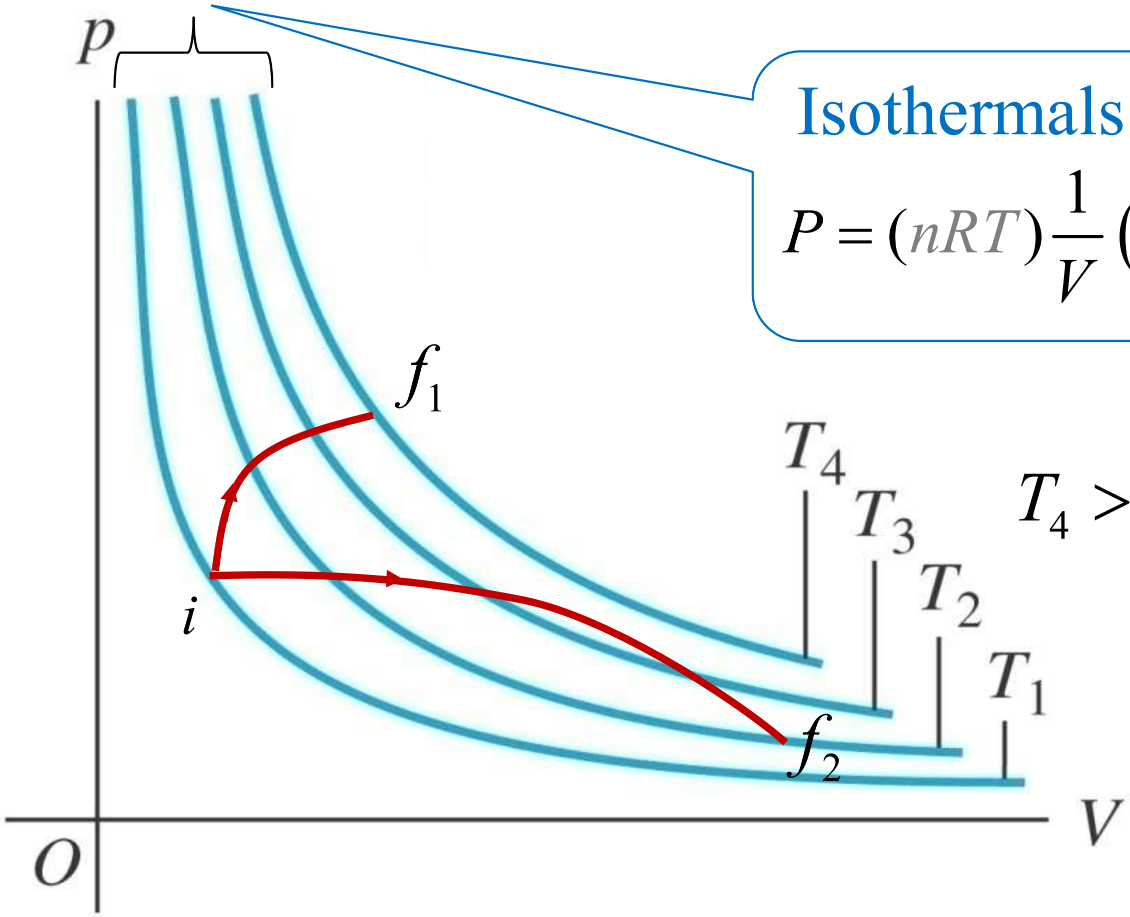
Process a: Constant *V*

Process b: Constant *P*

$$dQ_v \neq dQ_P$$

$$C_p = C_v + R$$

PV Diagrams



Isothermals

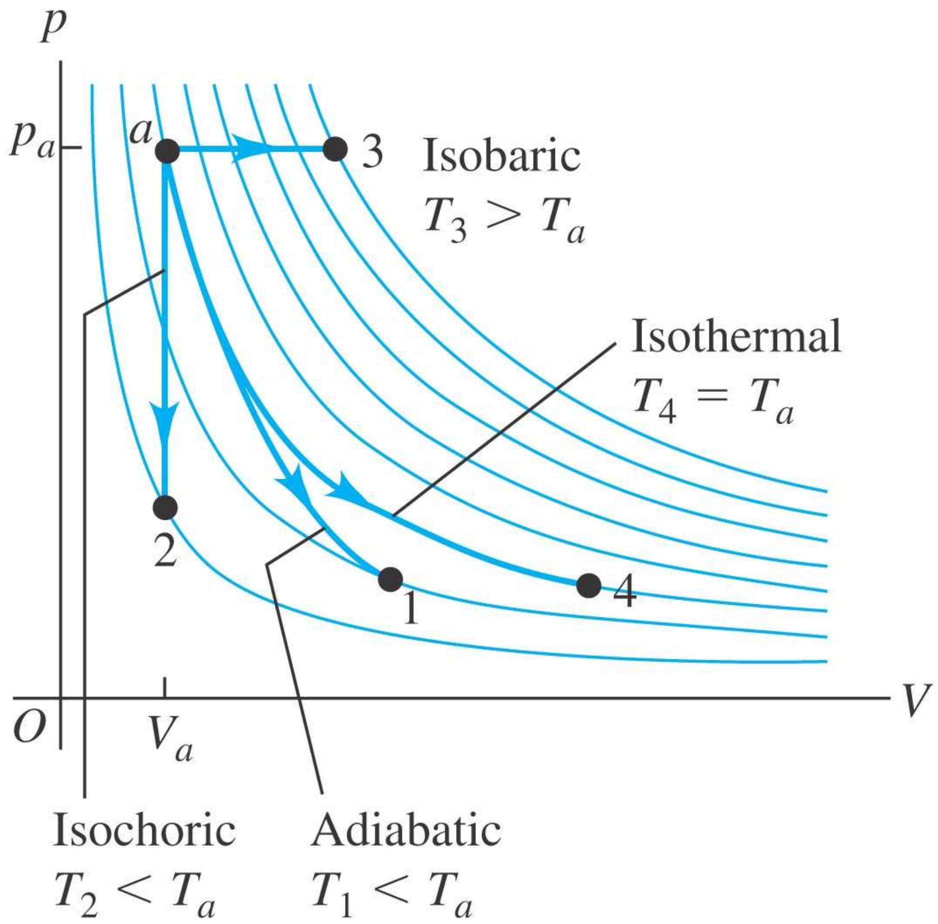
$$P = (nRT)\frac{1}{V}\left(\text{w/ cont } T\right)$$

$$T_4 > T_3 > T_2 > T_1$$

Special Processes – Non-Trivial

Four Reversible (Quasi-static) processes:

1. Adiabatic (Q=0)
2. Isochroic (ΔV=0)
3. Isobaric (ΔP=0)
4. Isothermal (ΔT=0)



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

Special Processes

3. Isochoric (constant *V*):

NO volume change → NO work $\boxed{W=0}$

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

4. Isobaric (constant *P*)

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

$$\boxed{W = P(V_2 - V_1)}$$

$$\Delta U = Q - W \rightarrow \boxed{\Delta U = Q - P\left(V_2 - V_1\right)} \qquad \boxed{Q = nC_P\Delta T}$$

Special Processes

5. Isothermal (constant *T*)

For an Ideal Gas, *U*(*T*) depends only on *T* → $\Delta U = 0$

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

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, $Q = 0$):

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

$$W = -nC_v(T_2 - T_1)$$

Additionally, for adiabatic process, we have the following relationship for the state variables:

$$TV^{\gamma-1} = \text{constant} \qquad PV^{\gamma} = \text{constant}$$

(alternate form)

where $\gamma = C_p / C_v$