**PHYS 1220, Engineering Physics, Chapter 19 – The First Law of Thermodynamics**

Instructor: TeYu Chien
Department of Physics and Astronomy
University of Wyoming

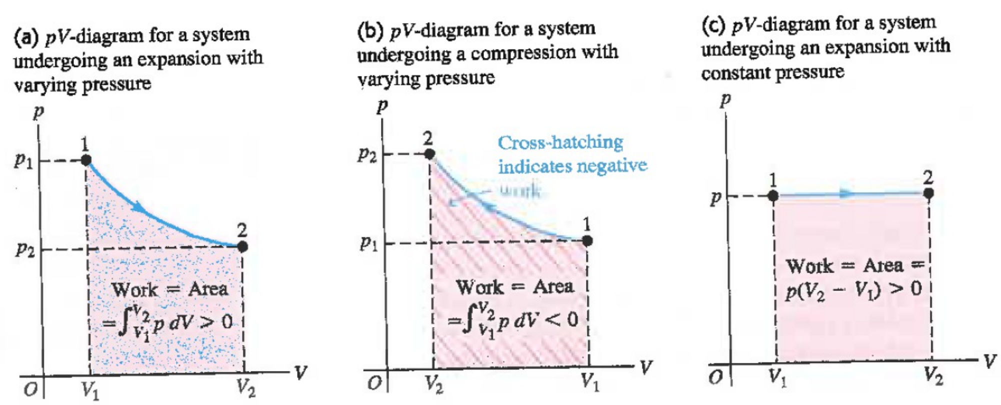
**Goal of this chapter is to learn the relationship between Heat, Work and Internal Energy of a system.**

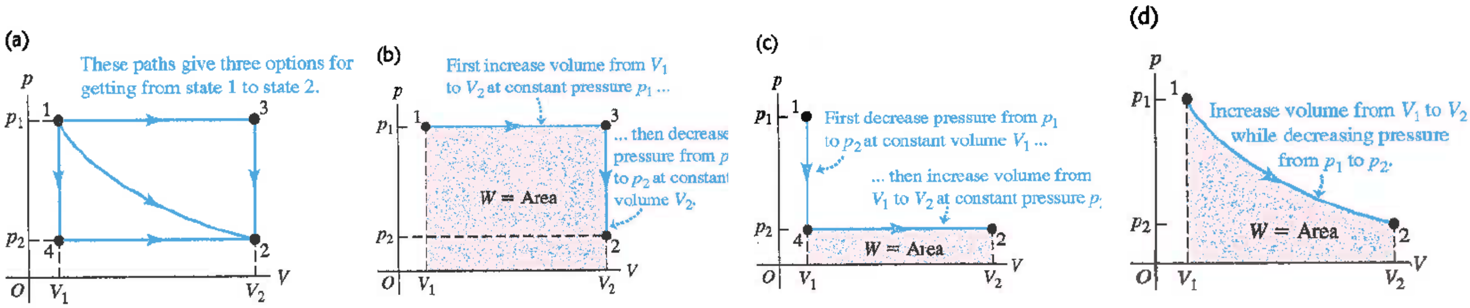
- What is a "system"?
  - A system is any **collection of objects** that is convenient to **regard as a unit**, and that may have the **potential to exchange energy with its surroundings**.

- What is a "status" of a system?
  - A status is to describe the states of a system. It **does not** depend on the **history** of the system, rather **it only depends** on its **thermodynamic state** (e.g. temperature, volume, pressure, number of molecules etc.)

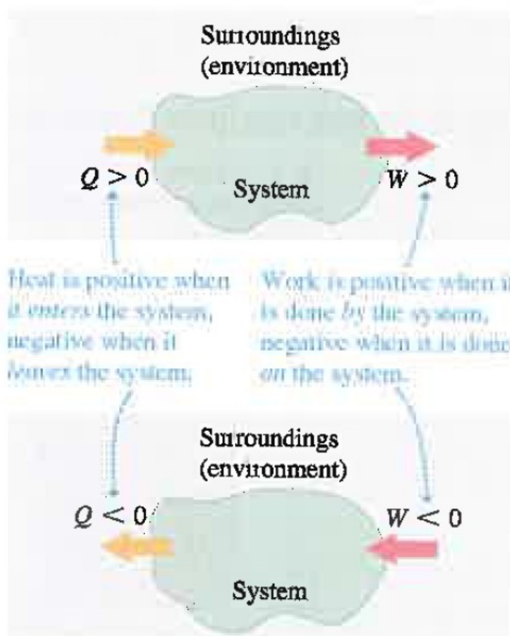
- What is a "process"?
  - A process is to describe how the system evolve from one state (initial state) to the other state (final state). Different process could have different evolution of **internal energy**, **heat** and **work** done by the system.

- Work, Heat and Internal Energy
  - Work ($W$) done by a system:
    - $W = \int_{V_1}^{V_2} p\, dV$ ("+" means work done **by** the system (**out**); "**-**" means work done **to** the system (**into**))






- Heat ($Q$) flow into/out of the system:
  - "+" means heat flow **into** the system (**in**); "-" means heat flow **out** of the system (**out**)



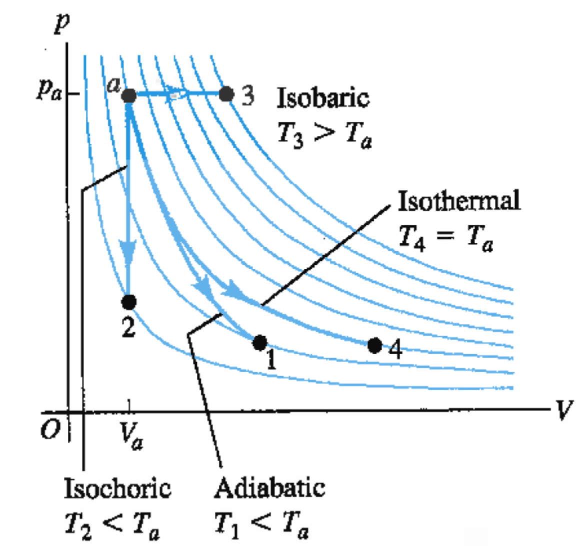

- Internal Energy ($U$) of a system:
  - It only depends on the thermodynamic states of the system (more detail when discuss Internal Energy of ideal gas)
- **First Law of Thermodynamics**
  - $U_2 - U_1 = \Delta U = Q - W$ (of a process)
  - or $dU = dQ - dW$ (infinitesimal process)

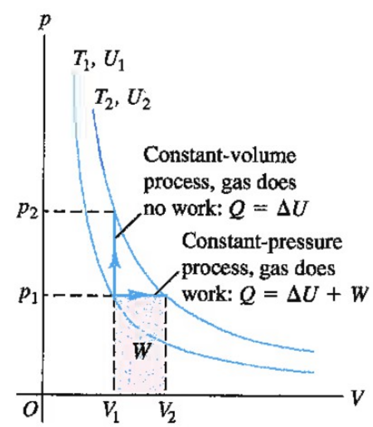
- Kinds of Thermodynamic Processes
- Adiabatic Process (no heat exchange, $\Delta Q = 0$ )
  - $\Delta U = -W$
- Isochoric Process (constant volume, $\Delta V = 0$ )

- No volume change means no work done, $\Delta W = 0$
- $\Delta U = Q$

- Isobaric Process (constant pressure, $\Delta p = 0$ )
  - constant pressure makes the calculation of the work each:
    $$W = \int_{V_1}^{V_2} p\,dV = p \int_{V_1}^{V_2} dV = p(V_2 - V_1)$$
  - $\Delta U = Q - p(V_2 - V_1)$

- Isothermal Process (constant temperature, $\Delta T = 0$ )
  - To have this process, at each infinitesimal process, the system needs to reach thermal equilibrium at all time.
  - **For systems which internal energy only depends on temperature**, $\Delta U = 0$ . Thus, $Q = W$




- Ideal Gas
  - Internal Energy ($U$) of an ideal gas depends **only on its temperature**, not on its pressure or volume.
  - $K_{tr} = \frac{dof}{2} nRT = U = nC_V T$ (remember: $C_V = \frac{dof}{2} R$ . *dof*: degrees of freedom)

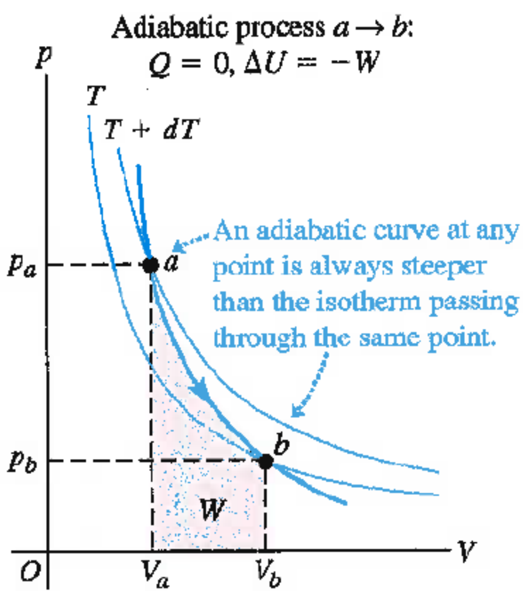
- Heat capacities ($C_V$, $C_p$) of an ideal gas
  - By definition of $C_V$: $dQ = nC_V\,dT$
  - For an ideal gas undergoes a isochoric process: $\Delta U = Q$ . $dU = nC_V\,dT$

- By definition of $C_p$: $dQ = nC_p dT$
- For an ideal gas undergoes a isobaric process: $dW = p\,dV = nRdT$ . $dW = nRdT$ . Thus, $dU = dQ + dW$ (first law) becomes $dQ = dU + nRdT = nC_p dT$

- Combine the results from constant volume and constant pressure: $nC_V dT + nRdT = nC_p dT$ . Thus: $C_V + R = C_p$ (for ideal gas)

- Ratio of heat capacities could tell how far from ideal gas
  - $\gamma = \frac{C_p}{C_V}$ ( $\gamma = 1.67$ for ideal monatomic gas; $\gamma = 1.40$ for ideal diatomic gas)

- Equation of state for Ideal gas undergoes an Adiabatic process
  - $dQ = 0$ (Adiabatic process), thus first law becomes: $dU = -dW$
  - $nC_V dT = -pdV$
  - $pV^{\gamma} = constant$ (for ideal gas undergoes an adiabatic process)




Math Preview for Chapter 20:
- Integration

Questions to think about for Chapter 20:

- Can heat flow from low to high temperature?  If not, why not? If yes, how and what are the limitations?