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Physics 1

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Lecture 6

- **Heat flow and the first law of thermodynamics.**
- **Kind of thermodynamic process. Adiabatic processes.**

Heat

- When the temperature of a thermal system in contact with a neighboring system changes, we say that there has been a **heat flow** into or out of the system.
- An energy unit related to thermal processes is the calorie (cal), which is defined as the amount of energy transfer necessary to raise the temperature of 1 gram of water by 1 degree (from 14.5°C to 15.5°C).

Mechanical equivalent of heat

- Mechanical energy is not conserserved in the presence of nonconservative forces. It transforms into internal energy. For example, friction produces heating

  1 cal = 4.186 J

Specific heat capacity

- The heat capacity *C* of a particular sample of a substance is defined as the amount of energy needed to raise the temperature of that sample by 1 °C.

$$C=Q/\Delta T$$

- The specific heat capacity *c* of a substance is the heat capacity per unit mass.

$$c=C/m=Q/(m\Delta T)$$

- Specific heat is essentially a measure of how thermally insensitive a substance is to the addition of energy. The greater a material’s specific heat, the more energy must be added to a given mass of the material to cause a particular temperature change.

Energy transfer and specific heat capacity

- From this definition, we can relate the energy ***Q*** transferred between a sample of mass ***m*** and specific heat capacity ***c*** of a material and its surroundings to a temperature change ΔT as

$$Q=mc\,\Delta T$$

Specific Heats of Some Substances at 25°C and Atmospheric Pressure

	Specific heat *c*	
Substance	**J/kg · °C**	**cal/g · °C**
Elemental solids		
Aluminum	900	0.215
Beryllium	1 830	0.436
Cadmium	230	0.055
Copper	387	0.092 4
Germanium	322	0.077
Gold	129	0.030 8
Iron	448	0.107
Lead	128	0.030 5
Silicon	703	0.168
Silver	234	0.056
Other solids		
Brass	380	0.092
Glass	837	0.200
Ice (− 5°C)	2 090	0.50
Marble	860	0.21
Wood	1 700	0.41
Liquids		
Alcohol (ethyl)	2 400	0.58
Mercury	140	0.033
Water (15°C)	4 186	1.00
Gas		
Steam (100°C)	2 010	0.48

Dependence of specific heat capacity on temperature

- Specific heat varies with temperature. For example, the specific heat of water varies by only about 1% from 0 c °C to 100 °C at atmospheric pressure. Usually such variations are negligible.

Dependence of specific heat capacity on volume and pressure

- Measured values of specific heats are found to depend on the conditions of the experiment. In general, measurements made in a constant pressure process are different from those made in a constant volume process. For solids and liquids, the difference between the two values is usually no greater than a few percent and is often neglected.

Phase transition

- It can be that transfer of energy does not result in a change in emperature. This is the case when the physical characteristics of the substance change from one form to another; such a change is called a phase change. Two common phase changes:
  - melting: from solid to liquid
  - boiling: from liquid to gas
  - change in the crystalline structure of a solid
- All such phase changes involve a change in internal energy but no change in temperature.
- The increase in internal energy in boiling, for example, is represented by the breaking of bonds between molecules in the liquid state; this bond breaking allows the molecules to move farther apart in the gaseous state, with a corresponding increase in intermolecular potential energy.

Latent heat

- Quantitative measure of phase transition is latent heat L:

$$Q=\pm mL$$

- Latent heat of fusion L_f is the term used when the phase change is from solid to liquid,
- Latent heat of vaporization L_v is the term used when the phase change is from liquid to gas (the liquid “vaporizes vaporizes”).

Latent Heats of Fusion and Vaporization

Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium	−269.65	5.23×10^3	−268.93	2.09×10^4
Nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
Oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
Ethyl alcohol	−114	1.04×10^5	78	8.54×10^5
Water	0.00	3.33×10^5	100.00	2.26×10^6
Sulfur	119	3.81×10^4	444.60	3.26×10^5
Lead	327.3	2.45×10^4	1 750	8.70×10^5
Aluminum	660	3.97×10^5	2 450	1.14×10^7
Silver	960.80	8.82×10^4	2 193	2.33×10^6
Gold	1 063.00	6.44×10^4	2 660	1.58×10^6
Copper	1 083	1.34×10^5	1 187	5.06×10^6

State variables - Thermodynamic process - Thermal equilibrium

- We describe the *state* of a system using such variables as pressure, volume, temperature, and internal energy. These quantities are called state variables. Macroscopic state of a system can be specified only if the system is in thermal equilibrium. When we regard a thermodynamic process we imply that all its state variables change quasi-statically, that is, slowly enough to allow the system to remain essentially in thermal equilibrium at all times.

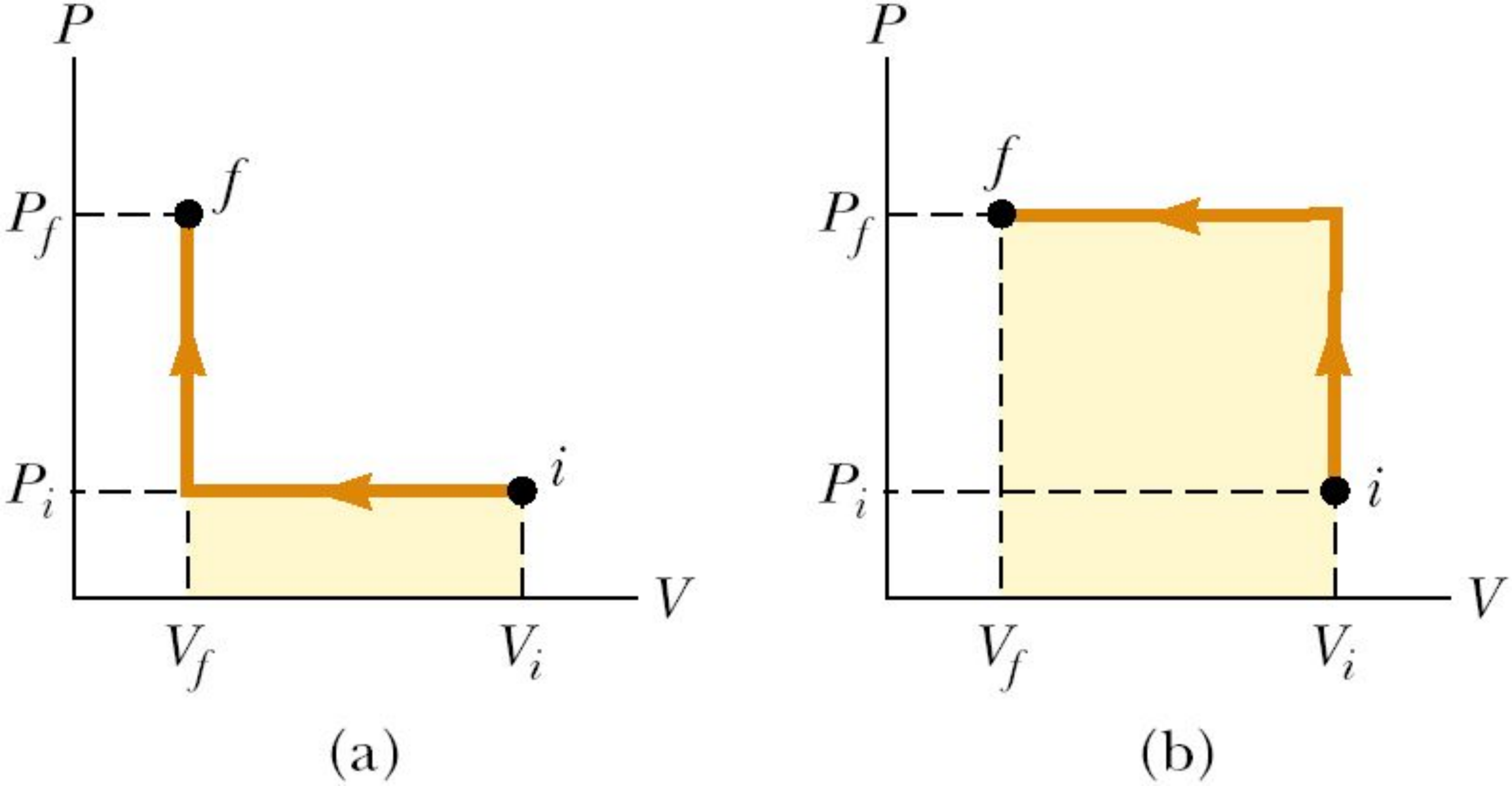
Work and heat in thermodynamic process

- The total work done by the gas as its volume changes from V_i to V_f is

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

- The work done by a gas in a quasi-static process equals the area under the curve on a *PV* diagram, evaluated between the initial and final states. It depends on the path between the initial and final states.

Work depends on the path:



(a): $W_a = P_i(V_f - V_i)$

(b): $W_b = P_f(V_f - V_i)$

1) $W_a < W_b$ as $P_f < P_i$

2) $W_a < W_b$ as the coloured area in (b) case is large then the area in (a) case

Two ways of energy transfer

There exist two ways in which energy can be transferred between a system and its surroundings:

- One way is work done by the system, which requires that there be a macroscopic displacement of the point of application of a force.
- The other is heat, which occurs on a molecular level whenever a temperature difference exists across the boundary of the system.

Both mechanisms result in a change in the internal energy of the system and therefore usually result in measurable changes in the macroscopic variables of the system, such as the pressure, temperature, and volume of a gas.

The First Law of Thermodynamics

- The change in internal energy ***ΔU*** of the system is equal to the heat ***Q*** put into a system minus the work ***W*** done by the system.

$$\Delta U = Q - W$$

- Note: here W is with the minus sign as the work is done by the system.

- The first law of thermodynamics is a special case of the law of conservation of energy that encompasses changes in internal energy and energy transfer by heat and work. It provides a connection between the microscopic and macroscopic approaches.

Ideal Gas Processes

- Here W is work done by the system,
- ΔQ - heat flow into the system.
- Isobaric (constant pressure):

$$W = P\Delta V$$

$$dQ = C_p dT$$

- Isochoric (constant volume):

$$\Delta W = 0$$

$$\Delta Q = \Delta U$$

$$dQ = C_V dT$$

C_p, C_V are specific heat capacities, $C_p = C_V + nR$, n is the number of moles.

- Isothermal (constant temperature):

$$\Delta U = 0$$

$$\Delta Q = \Delta W$$

- Adiabatic (no heat flow, Q=0):

$$\Delta W = -\Delta U$$

The curve of adiabatic process is described by formula:

$$PV^{\gamma} = const$$

$$TV^{\gamma-1} = const$$

$$\gamma = C_P/C_V$$

Polytropic processes

PV^{γ} = const, γ=const.

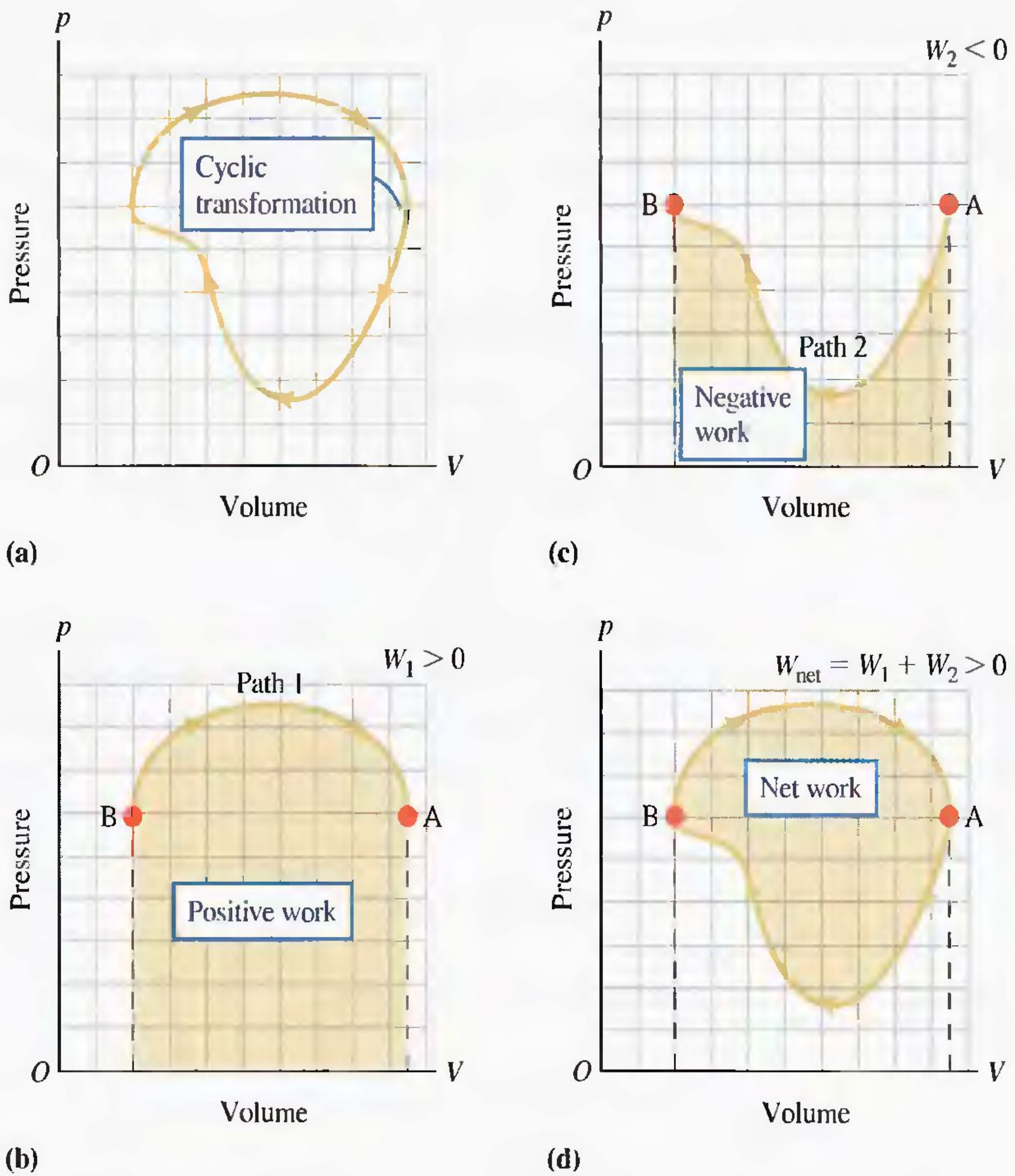
- Isobaric $\gamma=0$
- Isotermic $\gamma=1$
- Adiabatic $\gamma= C_P/C_V$
- Isochoric $\gamma=\infty$

Cyclic Processes

- If a nonisolated system is performing a cyclic process, the change in the internal energy must be zero. Therefore the energy *Q* added to the system must equal the negative of the work *W* done by the system during the cycle:

$$\Delta U = 0,$$
$$Q = W$$

- On a *PV* diagram, a cyclic process appears as a closed curve. In a cyclic process, the net work done by the system per cycle, equals the area enclosed by the path representing the process on a PV diagram.



In a closed cycle, the work done by a gas on its surroundings is the area enclosed by the curve of p versus V. To show this, the full cycle is broken into two paths – the upper and the lower.