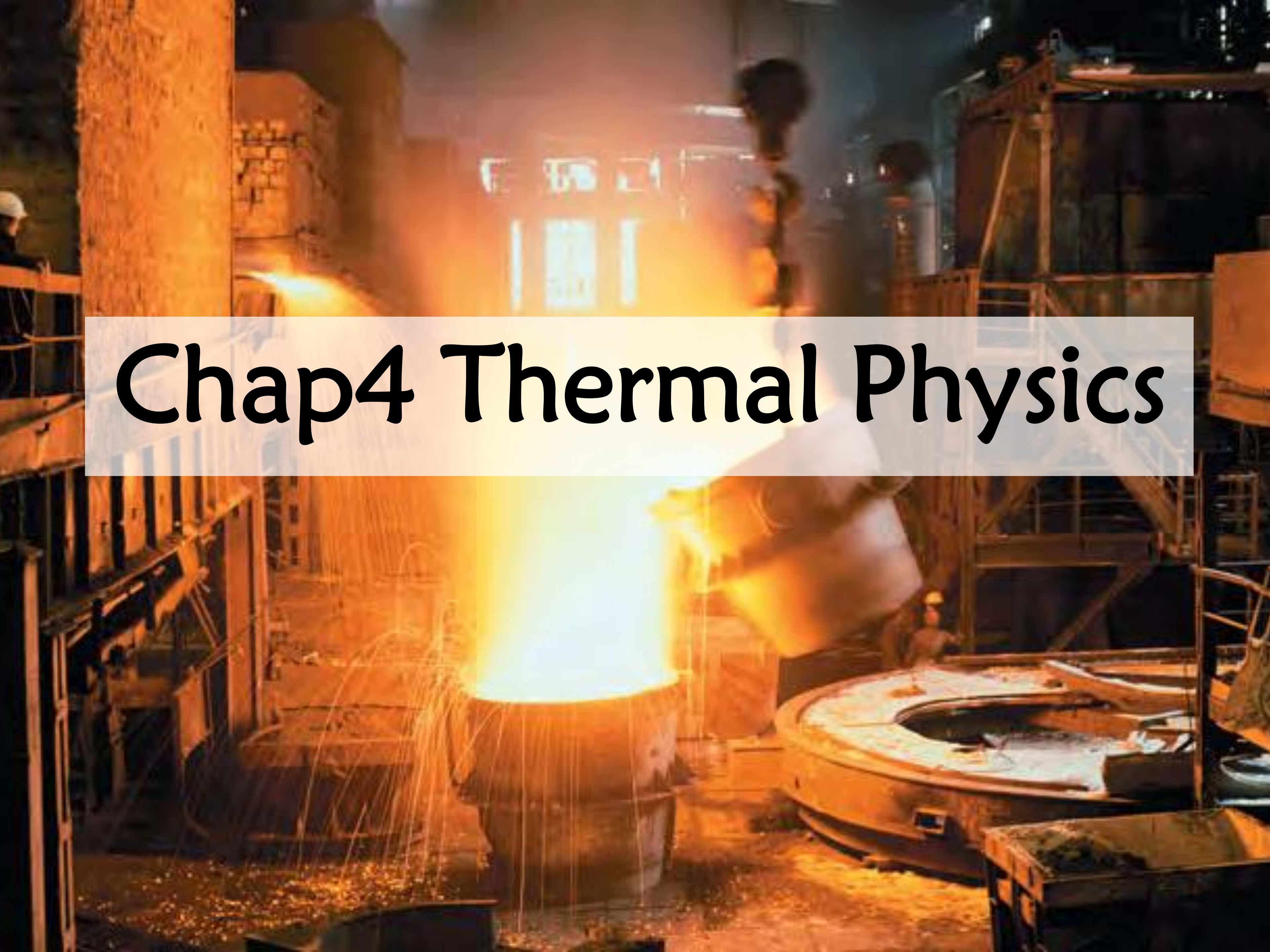


Chap4 Thermal Physics

Temperature Scales

Celsius scale °C

Fahrenheit scale °F

Absolute scale K
(*or* Kelvin scale)

$$T(^oF)=\frac{9}{5}T(^oC)+32$$

$$T(K)=T(^oC)+273.15$$

$$0K=-273.15^oC,$$

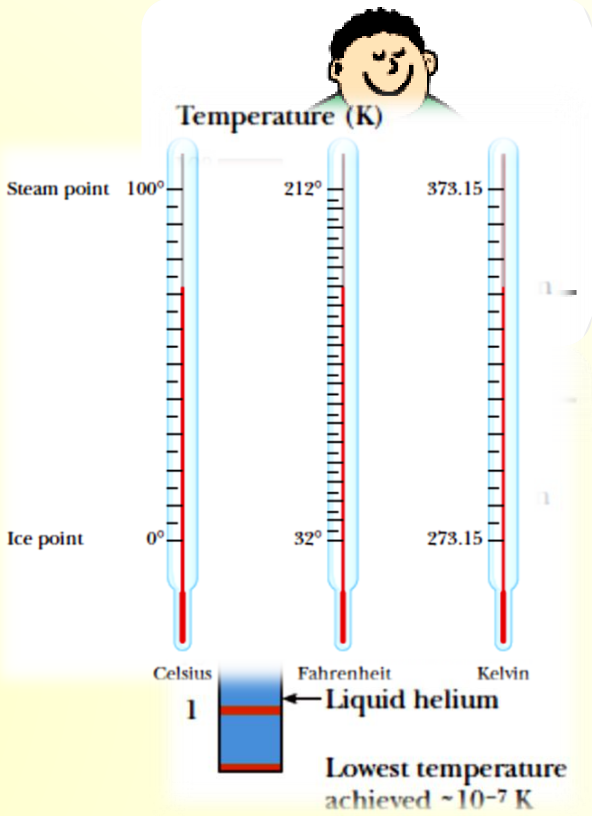
absolute zero

$$273.15K=0^oC,$$

$$373.15K=100^oC$$

$$0^oC=32^oF,\ 100^oC=212^oF$$

a change of 5 °C = a change of 9 °F



thermal equilibrium

When two objects are in *thermal contact*, energy can be exchanged between them; When there is no net exchange of energy, the two objects are in **thermal equilibrium**.

Thermal Expansion

This gap in the roadway of a bridge is called an **expansion joint**. Without the joint, the surfaces would buckle due to thermal expansion on very hot days or crack due to contraction on very cold days.

Extreme heat on a July day caused the buckling of these railroad tracks.

Thermal Expansion of Solids & Liquids

- If the thermal expansion of an object is sufficiently small compared with the object's initial dimensions, the change in any dimension is proportional to the temperature change.

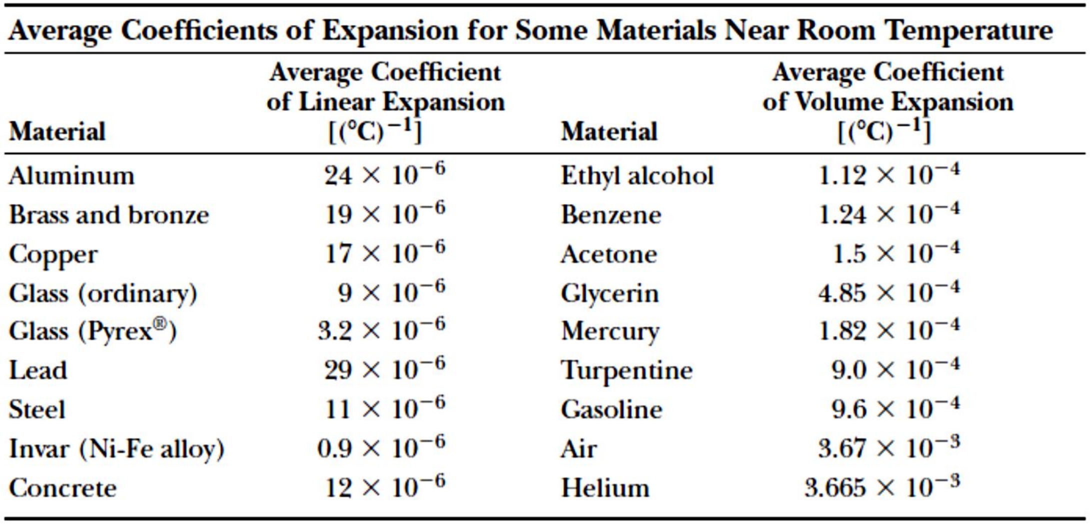
$$\boxed{\Delta L = \alpha L_0 \Delta T}$$

(*Note:* L_0 *is the initial length,* ΔT *is the temperature change, and* α *is the coefficient of linear expansion for a given material and has units of* ($^{o}C^{-1}$).)

- The increase in area of an object accompanying a change in temperature:

$$\boxed{\Delta A = \gamma A_0 \Delta T}$$

(*Note:* γ *is the coefficient of area expansion*. $\gamma = 2\alpha$ only if α of the object is the same in all directions.)



Average Coefficients of Expansion for Some Materials Near Room Temperature

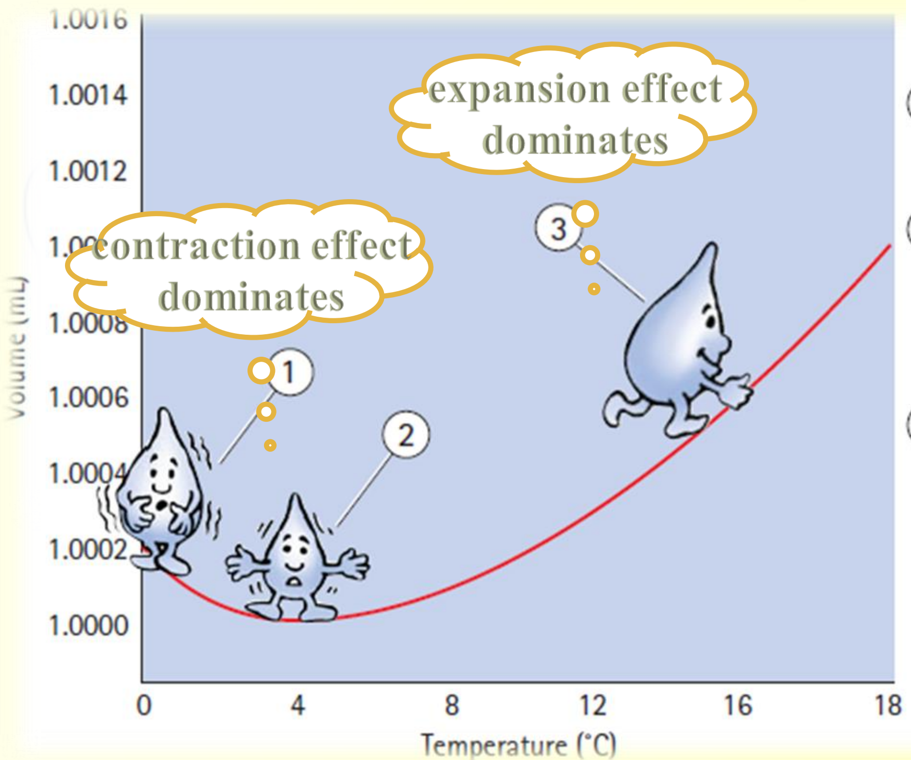
Material	Average Coefficient of Linear Expansion [$(^\circ C)^{-1}$]	Material	Average Coefficient of Volume Expansion [$(^\circ C)^{-1}$]
Aluminum	24×10^{-6}	Ethyl alcohol	1.12×10^{-4}
Brass and bronze	19×10^{-6}	Benzene	1.24×10^{-4}
Copper	17×10^{-6}	Acetone	1.5×10^{-4}
Glass (ordinary)	9×10^{-6}	Glycerin	4.85×10^{-4}
Glass (Pyrex®)	3.2×10^{-6}	Mercury	1.82×10^{-4}
Lead	29×10^{-6}	Turpentine	9.0×10^{-4}
Steel	11×10^{-6}	Gasoline	9.6×10^{-4}
Invar (Ni-Fe alloy)	0.9×10^{-6}	Air	3.67×10^{-3}
Concrete	12×10^{-6}	Helium	3.665×10^{-3}

- The increase in volume of object accompanying a change in temperature:

$$\boxed{\Delta V = \beta V_0 \Delta T}$$

(*Note:* β *is the coefficient of volume expansion*. $\beta = 3\alpha$ only if α of the object is the same in all directions.)

Expansion of Water

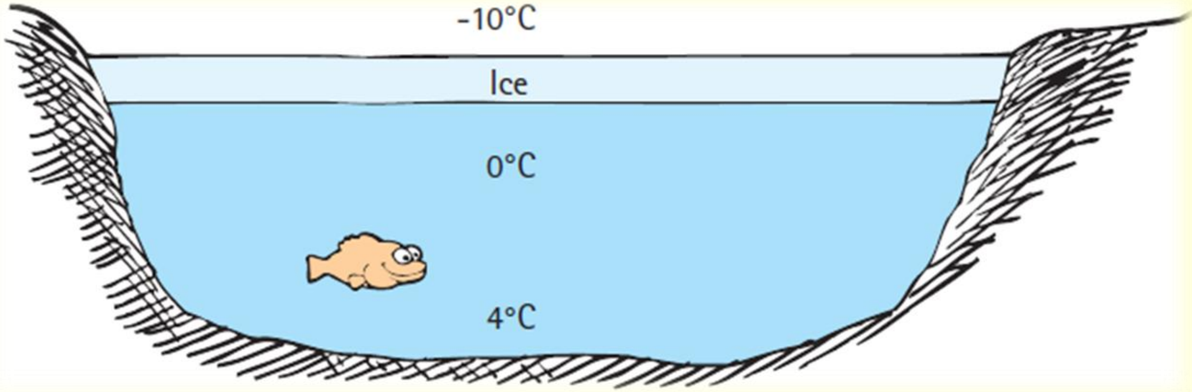


① Liquid water below 4˚C is bloated with ic crystals.

② Upon warming, the crystals collapse, resulting in a smaller volume for the liquid water.

③ Above 4˚C, liquid water expands as it is heated because of greater molecular motion.

Between 0°C and 4°C, the volume of liquid water *decreases* as temperature *increases*. Above 4°C, thermal expansion exceeds contraction effect and volume *increases* as temperature *increases*.



Because water is most dense at 4°C, colder water rises and freezes on the surface. This means that the ice forms an insulating layer that slows heat loss from the underlying water, offering thermal protection for marine life!

Ideal Gas

- An ideal gas is a collection of atoms or molecules that move randomly and exert *no long-range* forces on each other. Each particle of the ideal gas is individually point-like, occupying a negligible volume.
- Most gases at room temperature and atmospheric pressure behave approximately as ideal gases.
- One mole (mol) of any substance is the amount of the substance that contains as many particles (atoms, molecules, or other particles) as there are atoms in 12 g of the isotope carbon-12.

Avogadro’s number: $N_A = 6.02 \times 10^{23}\ \text{particles/mol}$

The number of moles: $n = \dfrac{m}{molar\ mass}$ (unit: g/mol)

The number of molecules: $N = n \cdot N_A$

Ideal Gas Law (Equation of state for ideal gases)

♣ $$\boxed{PV = nRT}$$

the number of moles

universal gas constant

$$R = 8.31\ \mathrm{J/(mol\cdot K)}$$

- The volume occupied by 1 mol of *any* ideal gas at atmospheric pressure and at 0 ^{o}C is 22.4 L.

♣ $$\boxed{PV = Nk_BT}$$

the number of molecules

Boltzmann's constant

$$k_B = R/N_A = 1.38\times10^{-23}\ \mathrm{J/K}$$

Kinetic Theory of Gases Model

Assumption of kinetic theory

- ♠ The number of molecules in the gas is large, and the average separation between them is large compared with their dimensions.
- ♠ The molecules obey Newton’s laws of motion, but as a whole they move randomly.
- ♠ The molecules make elastic collisions with the walls.
- ♠ The molecules interact only through short-range forces during elastic collisions.
- ♠ All molecules in the gas are identical.
- ♠ *U* associated with weak attractive forces among molecules is small compared to *K*, and we ignore it.

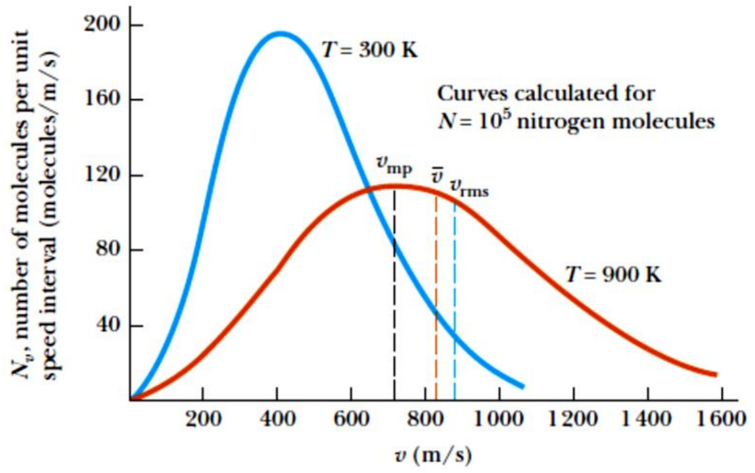
Molecular Kinetic Theory

- Average kinetic energy per molecule:

$$\boxed{K = \frac{1}{2}m\overline{v^2} = \frac{3}{2}k_BT}$$

Boltzmann's constant

Maxwell speed distribution



- Total translational kinetic energy of *N* molecules of gas:

$$\boxed{K_{tot} = N\left(\frac{1}{2}m\overline{v^2}\right) = \frac{3}{2}Nk_BT = \frac{3}{2}nRT}$$

Universal gas constant

  - The total translational kinetic energy of a system of molecules is proportional to the absolute temperature of the system.

- Root-mean-square (rms) speed of the molecules:

$$\boxed{v_{\text{rms}} = \sqrt{\overline{v^2}} = \sqrt{\frac{3k_BT}{m}} = \sqrt{\frac{3RT}{M}}}$$

Molar mass

Heat

- ➢ Heat is the thermal energy transferred from one location to another due to a temperature difference.
- ➢ According to this definition, matter contains thermal energy --- *not* heat.
- ➢ Units of Heat:
  - SI Unit: Joule (J)
  - Common Units: calorie (cal), kilocalorie (kcal)
  - The calorie (cal) is defined as the energy necessary to raise the temperature of 1 g of water from 14.5 ^{o}C to 15.5 ^{o}C.

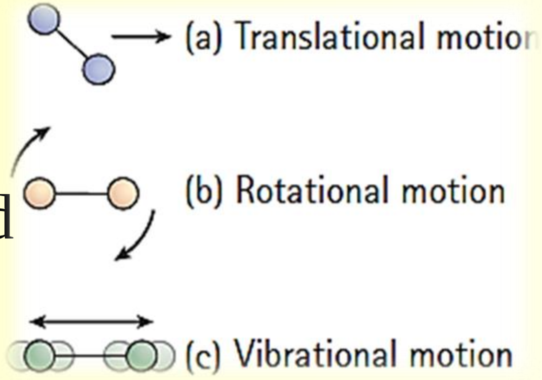
Food unit “Calorie”

⤢

$$1\text{cal} = 4.186\text{J};\ 1\text{kcal} = 4.186\text{k J}$$

Internal Energy *U*

- *U* is the energy associated with the microscopic components of a system --- the atoms and molecules of the system. The internal energy includes kinetic and potential energy associated with the random *translational*, *rotational*, and *vibrational* motion of the particles that make up the system, and any potential energy bonding the particles together.



- As for a *monatomic* ideal gas, *U* is associated with the translational motion of its atoms. In this special case, *U* is the total translational kinetic energy of the atoms.

$$\boxed{U = K_{tot} = \frac{3}{2}Nk_BT = \frac{3}{2}nRT}$$

Specific Heat (Capacity)

The filling of hot apple pie may be too hot to eat, even though the crust is not.

- The specific heat capacity of any substance is defined as quantity of heat required to change the temperature of a unit mass of the substance by 1°C
- Specific heat capacity is like ***thermal inertia*** because it signifies the resistance of a substance to a change in temperature.
- Units of Specific Heat:
  - SI Unit: Joule per kilogram-degree Celsius (J/kg·°C)
  - Common Unit: kilocalorie per kilogram-degree Celsius (kcal/kg·°C)
- Water has a much higher specific heat capacity than most other common materials.

specific heat

$$Q = cm\Delta T$$

Latent Heat (*L*) & Phase Change

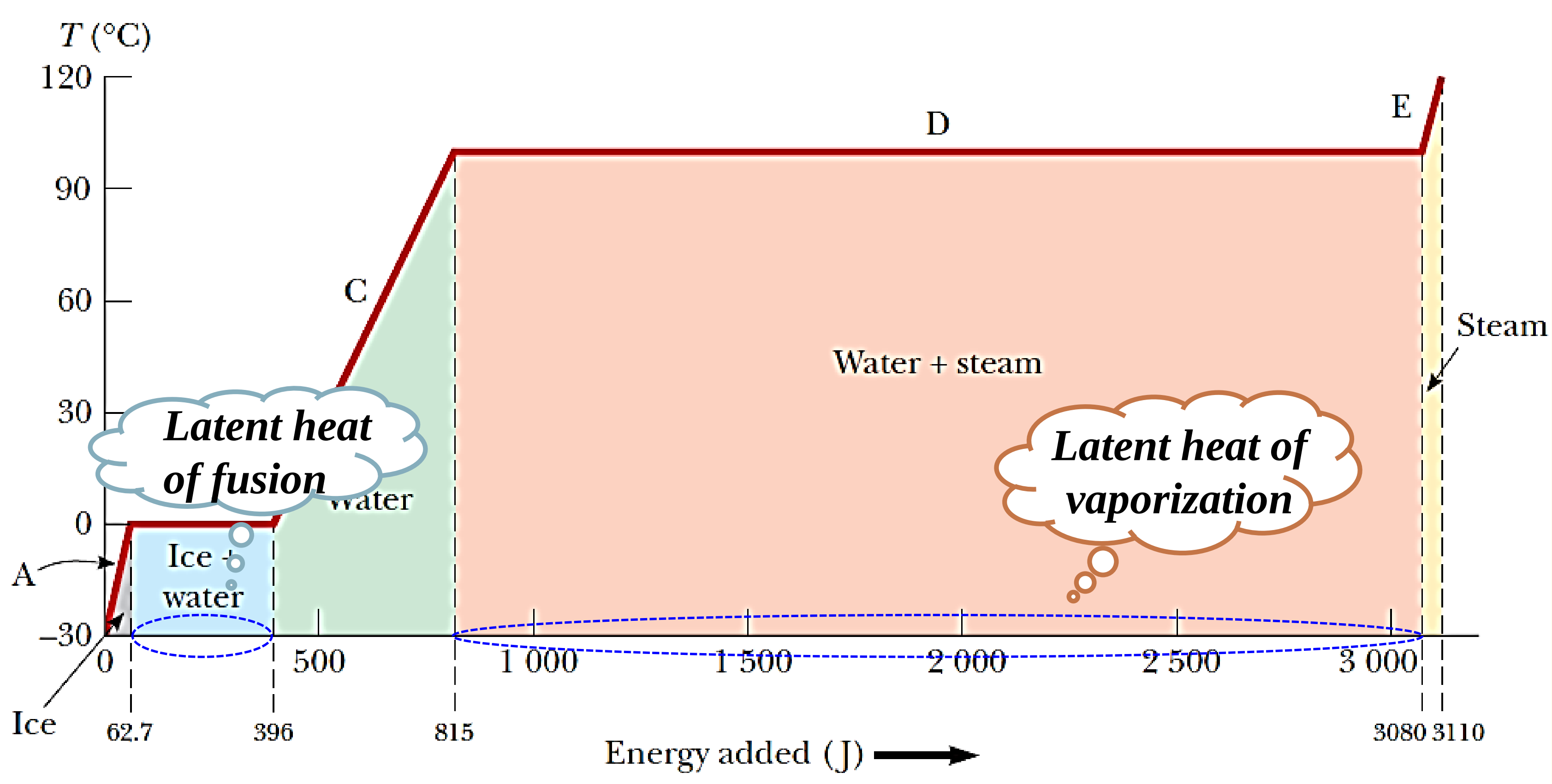
- The energy Q needed to change the phase of a given pure substance is $Q = mL$, where L, called the **latent heat** of the substance, depends on the nature of the phase change as well as on the substance.

- SI Unit: Joule per kilogram (J/kg)

- **Latent heat of fusion L_f** is used when a phase change occurs during melting or freezing.

- **Latent heat of vaporization L_v** is used when a phase change occurs during boiling or condensing.

- In general, the latent heat of vaporization is ***much greater*** than the latent heat of fusion.



Example: a plot of temperature vs. energy added when 1.00 g of ice, initially at -30.0℃, is converted to steam at 120℃.

Heat Transfer

- Heat transfer is defined as transfer of energy between objects that have different temperatures. The direction of heat flow depends on temperature, and from hot object (more thermal energy) to cold object (less thermal energy).

- Three modes: Conduction, Convection, Radiation

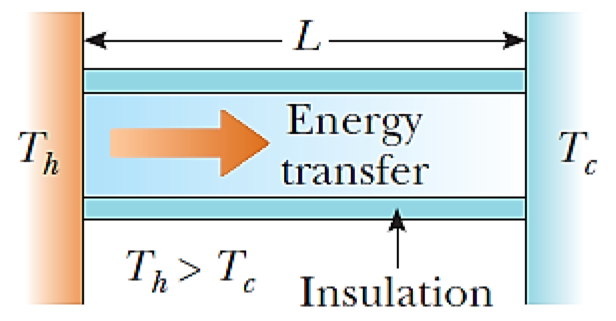
Conduction

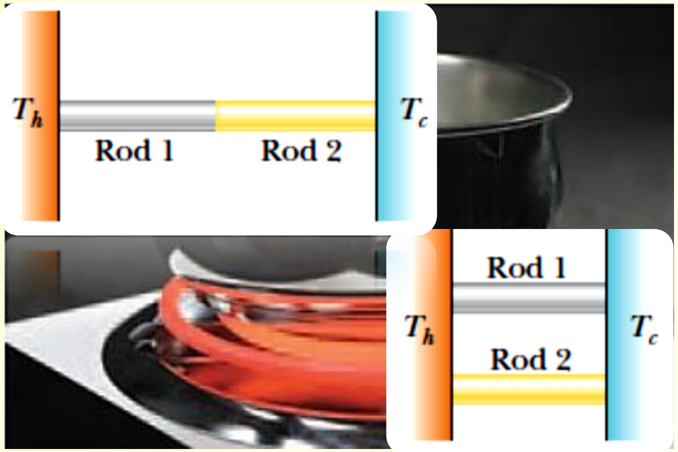
(mainly in solid)

♠ **the rate of heat transfer:**

$$\frac{Q}{\Delta t} = kA\frac{\Delta T}{L}$$

thermal conductivity

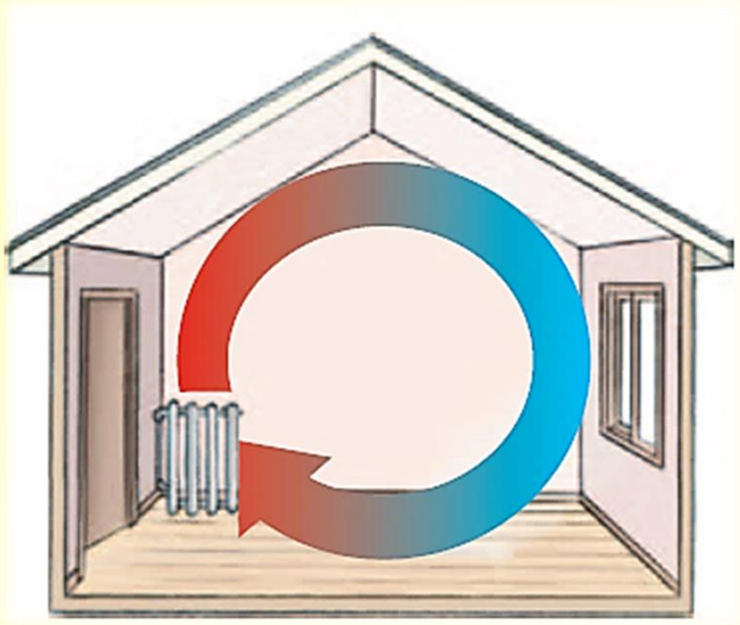




Convection (mainly in liquid or gas)

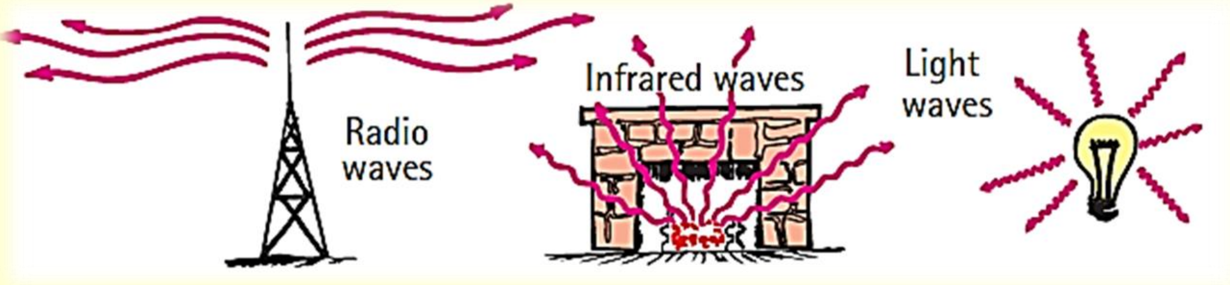
- The transfer of energy by the movement of a substance is called convection.

Warming a hand by convection.



Convection currents are set up in a room warmed by a radiator.

Radiation



- If an object is at a temperature T, and its surroundings are at a temperature T_0, the net energy lost or gained each second by the object through radiation is:

$$H_{net} = \sigma A e (T^4 - T_0^4)$$

emissivity of the object

Stefan-Boltzmann constant

surface area of the object

- For an ideal absorber, $e = 1$
- For an ideal reflector, $e = 0$

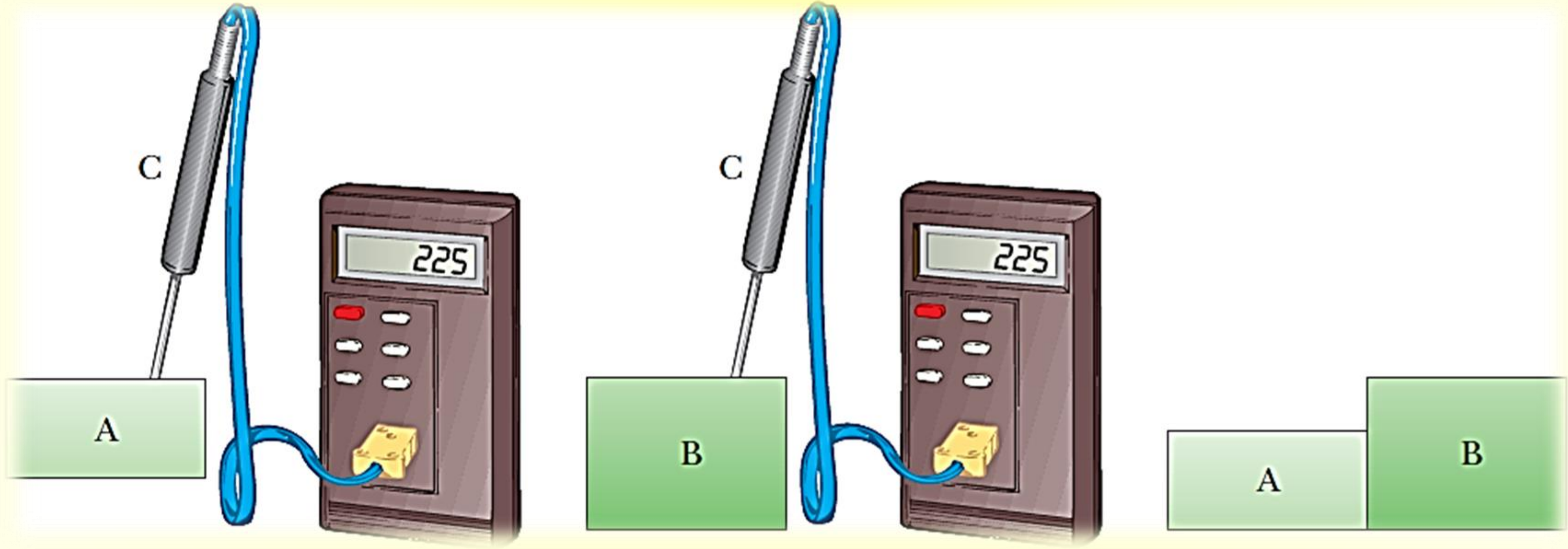
This thermogram of a house, made during cold weather, shows colors ranging from white and yellow (*areas of greatest energy loss*) *to blue and purple* (*areas of least energy loss*).

A radiation thermometer measures a person's temperature by monitoring the intensity of infrared radiation leaving the ear.

The Laws of Thermodynamics

The zeroth law of thermodynamics

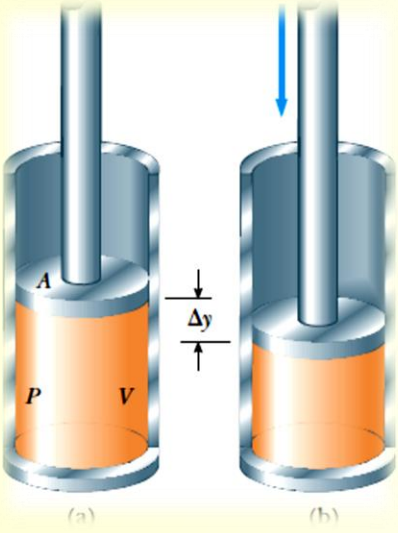
If objects A and B are separately in thermal equilibrium with a third object C, then A and B are in thermal equilibrium with each other.



The first law of thermodynamics

If a system undergoes a change from an initial state to a final state, where *Q* is the energy transferred to the system by heat and *W* is the work done on the system, the change in the internal energy △*U* of the system:

$$\boxed{\Delta U = Q + W}$$

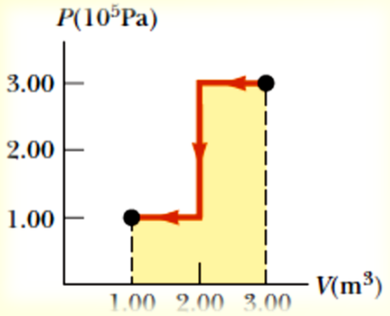


- The work *W* done on a gas at constant pressure:

$$\boxed{W = -P\Delta V}$$

(*Note: P is the pressure throughout the gas, △V is the change in volume of the gas during the process.*)

- The area under the graph in a *PV* diagram is *equal to* in magnitude to the work done on the gas.



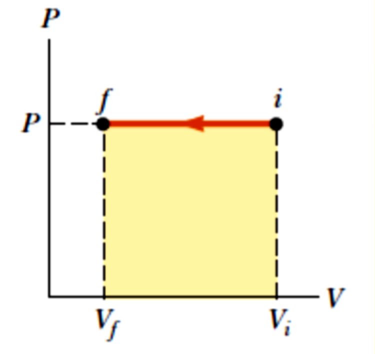
- For a monatomic ideal gas, change in internal energy:

$$\boxed{\Delta U = \frac{3}{2} nR\Delta T = nC_v\Delta T}$$

(*Note: C_v is the molar specific heat at constant volume of the gas,* $C_v = 3R/2$.)

Isobaric Processes:

- In an isobaric process, the pressure remains constant as the gas expands or is compressed.

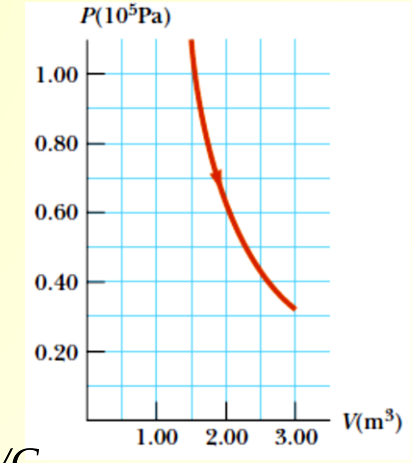


$$\left.\begin{aligned} Q = \Delta U - W &= \Delta U + P\Delta V \\ \Delta U &= nC_v\Delta T \\ P\Delta V &= nR\Delta T \\ Q &= nC_p\Delta T \end{aligned}\right\} \Rightarrow \boxed{C_p = C_v + R}$$

(Note: C_p is the ***molar specific heat*** *at constant pressure of the gas. As for a monoatomic ideal gas, $C_p = 5R/2$.)*

Adiabatic Processes:

- In an adiabatic process, no energy enters or leaves the system by heat, so $Q = 0$.
- A sufficiently rapid process can be considered approximately adiabatic because heat does not have time to flow in or out of the system.



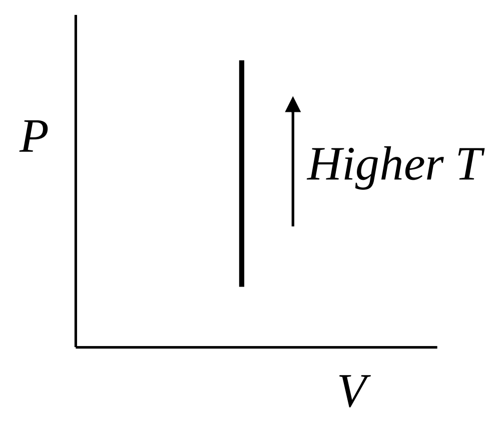
$$\boxed{PV^\gamma = Const.}$$

(Note: γ is the adiabatic index of the gas, $\gamma = C_p/C_v$. As for a monoatomic ideal gas, $\gamma = 5/3$.)

♠ Isovolumetric (isochoric) Processes:

- An isovolumetric process proceeds at constant volume, corresponding to vertical lines in a *PV* diagram.
- In an isovolumetric process, $W = 0$.

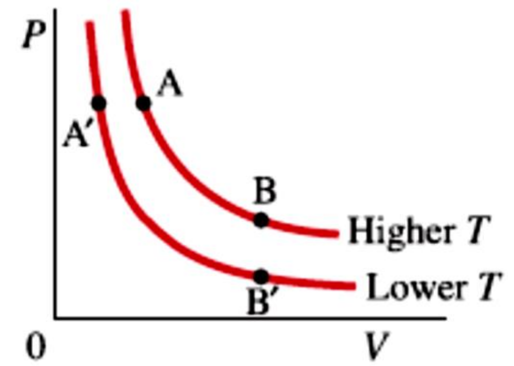
$$\boxed{Q = \Delta U = nC_v \Delta T}$$

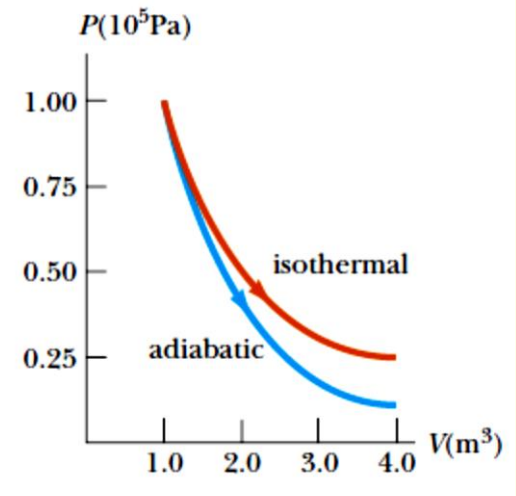


♠ Isothermal Processes:

- During an isothermal process, the temperature of a system does not change, so $\triangle U = 0$.
- The work done on the gas system during an isothermal compression process:

$$\boxed{W = nRT \ln\left(\frac{V_i}{V_f}\right)}$$

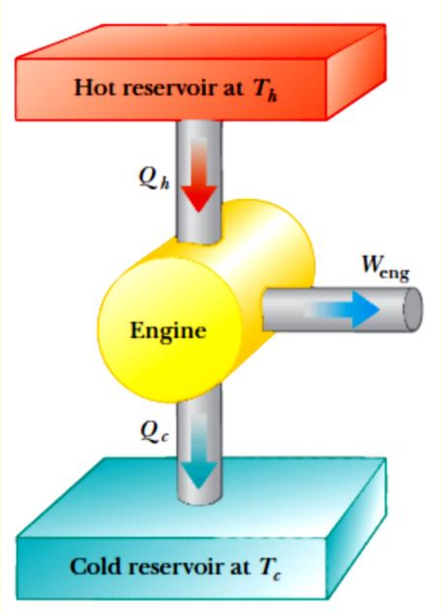




Heat Engines

- A **heat engine** takes in energy by heat and partially converts it to other forms, such as electrical and mechanical energy.

- A heat engine carries some working substance through a *cyclic process* during which (1) energy is transferred by heat from a source at a high temperature, (2) work is done by the engine, and (3) energy is expelled by the engine by heat to a source at lower temperature.



$$\Delta U = 0 \Rightarrow Q_{\text{net}} = -W = W_{\text{eng}} \Rightarrow W_{\text{eng}} = |Q_h| - |Q_c|$$

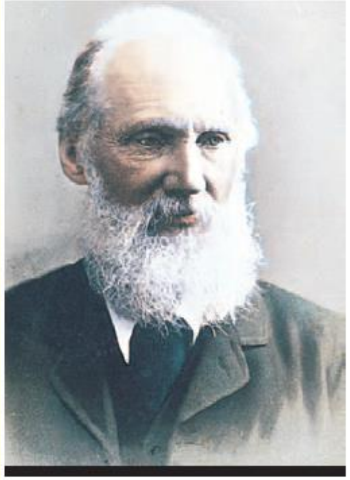
$$\Downarrow$$

- Thermal efficiency of a heat engine: $\boxed{e \equiv \frac{W_{\text{eng}}}{|Q_h|} = 1 - \frac{|Q_c|}{|Q_h|}}$

The second law of thermodynamics

♠ Kelvin statement:

- No heat engine operating in a cycle can absorb energy from a reservoir and use it entirely for the performance of an equal amount of work.



LORD KELVIN, British Physicist and Mathematician (1824–1907)

☐ Reversible & Irreversible Processes:

- Most natural processes are known to be irreversible—the reversible process is an idealization.
- If a real process occurs so slowly that the system is virtually always in equilibrium, the process can be considered reversible.

☐ The Carnot Engine:

- A Carnot engine is a theoretical engine operating in an ideal, reversible cycle—now called a Carnot cycle — between two energy reservoirs.

J-L Charmet/SPL/Photo Researchers, Inc.

SADI CARNOT, French Engineer (1796–1832)

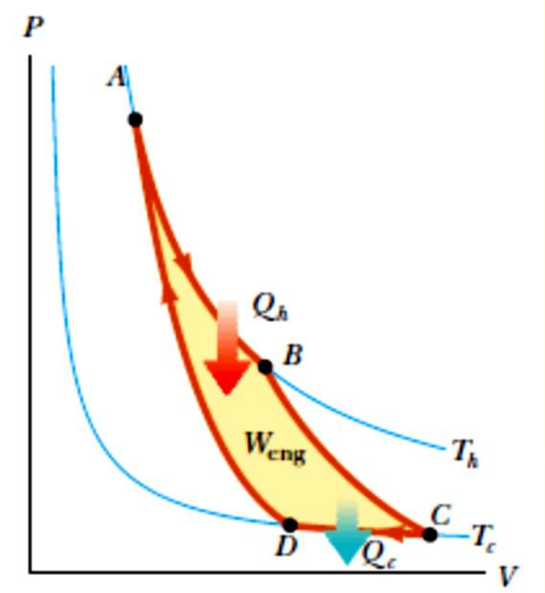
Carnot's Theorem:

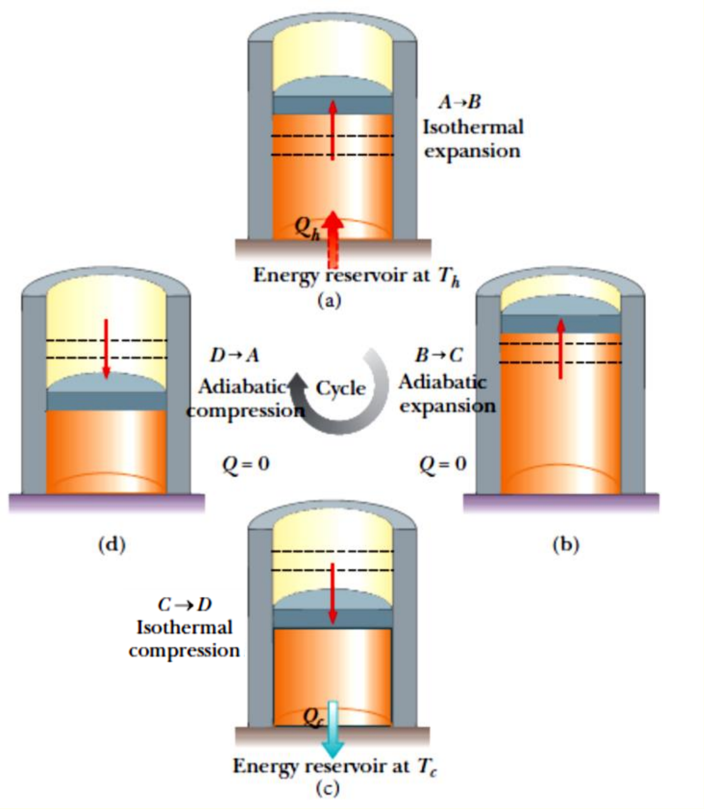
- No real engine operating between two energy reservoirs can be more efficient than a Carnot engine operating between the same two reservoirs.

Four stages of Carnot cycle:

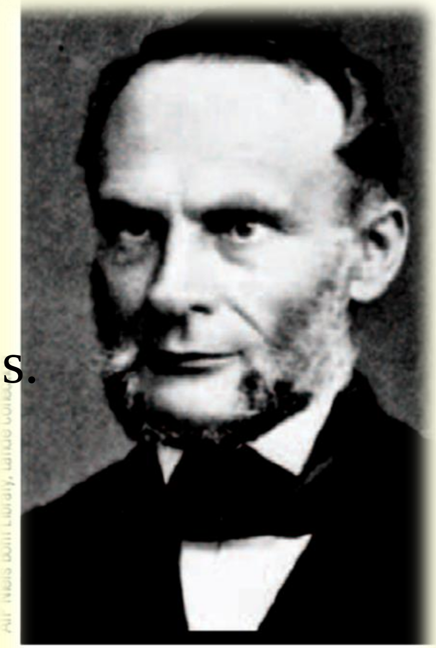
- Thermal efficiency of a Carnot engine:

$$\boxed{e_c = 1 - \frac{T_c}{T_h}}$$





☐ Entropy S:


AIP Niels Bohr Library, Lande Collection

RUDOLF CLAUSIUS, German Physicist (1822–1888)

- Temperature and internal energy, associated with the zeroth and first laws of thermodynamics, respectively, are both *state variables*. A state variable called the entropy S is related to the second law of thermodynamics.
- Entropy is a measure of the *disorder* of the system.

♠ Entropy increase statement for the 2nd law of thermodynamics:

- Natural processes spontaneously move toward a state of greater disorder, so the entropy of the isolated system increases in all natural processes.

The third law of thermodynamics

- It is impossible to lower the temperature of a system to absolute zero in a finite number of steps. That is, no system can reach absolute zero.

Maxwell's demon