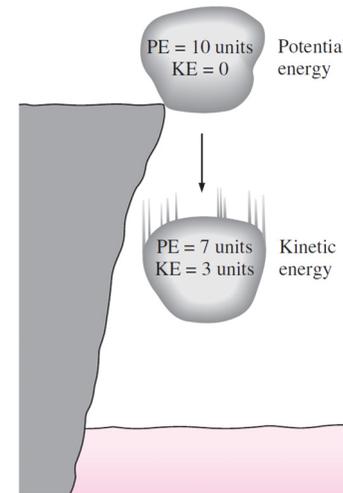


First Law of Thermodynamics and Energy Transfer Mechanisms

Physics Enhancement Programme
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Thermodynamics and Energy



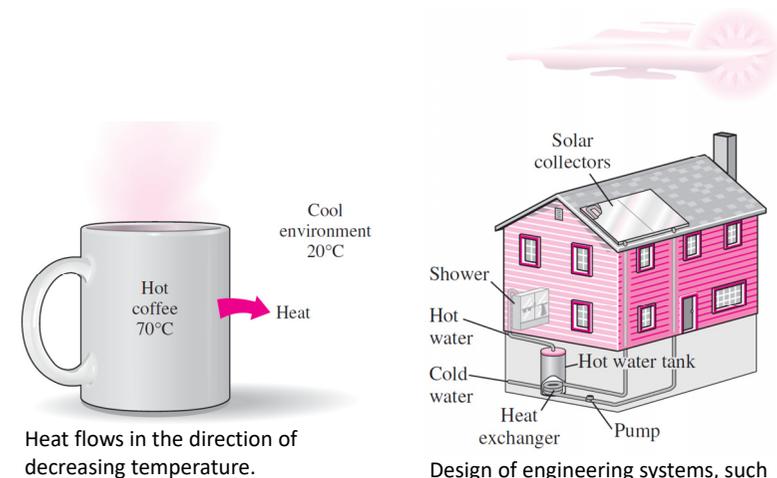
The most fundamental laws of nature:
Principle of Conservation of Energy

Energy cannot be created or destroyed, but during an interaction, energy can change from one form to another, and the total energy remains constant.

Historical Background of Thermodynamics

- Thermodynamics and mechanics were considered to be distinct branches of physics.
 - Until about 1850.
 - Experiments by James Joule and others showed a connection between them.
- Connection: Between the transfer of energy by heat in thermal processes and the transfer of energy by work in mechanical processes.
- The concept of energy was generalized to include internal energy.
- The Law of Conservation of Energy emerged as a universal law of nature.

Applications of Thermodynamics



Design of engineering systems, such as solar hot water system, involves thermodynamics.

Internal Energy

- Internal energy is all the energy of a system that is associated with its microscopic components.
 - These components are its atoms and molecules.
 - The system is viewed from a reference frame at rest with respect to the center of mass of the system.

Internal Energy and Other Energies

- Kinetic energy due to its motion through space is not included.
- Internal energy does include kinetic energies due to:
 - Random translational motion;
 - Rotational motion;
 - Vibrational motion.
- Internal energy also includes potential energy between molecules.

Heat

- Heat is defined as the transfer of energy across the boundary of a system due to a temperature difference between the system and its surroundings.
- The term heat will also be used to represent the amount of energy transferred by this method.

Changing Internal Energy

- Both heat and work can change the internal energy of a system.
- The internal energy can be changed even when no energy is transferred by heat, but just by work.
 - Example: compressing gas with a piston.
 - Energy is transferred by work.

Units of Heat

- Historically, the calorie was the unit used for heat.
 - One calorie is the amount of energy transfer necessary to raise the temperature of 1 g of water from 14.5°C to 15.5°C
 - The “Calorie” used for food is actually 1 kilocalorie.
- In the US Customary system, the unit is a BTU (British Thermal Unit)
 - One BTU is the amount of energy transfer necessary to raise the temperature of 1 lb of water from 63°F to 64°F
- SI Unit: Joules [J]

James Prescott Joule

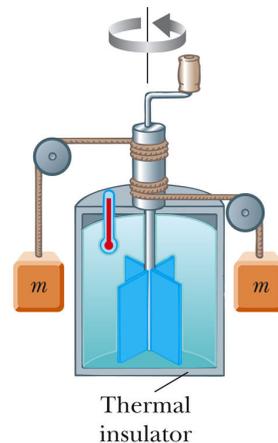
- 1818 – 1889.
- British Physicist.
- Research led to establishment of the principle of Conservation of Energy.
- Determined the amount of work needed to produce one unit of energy.



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Mechanical Equivalent of Heat

- Joule established the equivalence between mechanical energy and internal energy.
- The loss in potential energy associated with the blocks equals the work done by the paddle wheel on the water



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Mechanical Equivalent of Heat

- Joule found that it took approximately 4.18 J of mechanical energy to raise the water 1°C.
- Later, more precise, measurements determined the amount of mechanical energy needed to raise the temperature of water from 14.5°C to 15.5°C.
- Mechanical equivalent of heat:
 - 1 calorie = 4.186 J

Heat Capacity

- Heat capacity (C) is defined as the amount of energy needed to raise the temperature of a sample by 1°C .
- Energy (Q) to produce a change of temperature of ΔT :

$$Q = C \Delta T$$

Specific Heat

- Specific heat (c) is the heat capacity per unit mass.
- If energy Q transfers to a sample of a substance of mass m and the temperature changes by ΔT , then the specific heat is:

$$Q = mc \Delta T$$

Specific Heat

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

Quick Quiz

Imagine you have 1 kg each of iron, glass, and water, and that all three samples are at 10°C .

- Rank the samples from lowest to highest temperature after 100 J of energy is added to each sample.

Answer:

- Rank from least to greatest amount of energy transferred by heat if each sample increases in temperature by 20°C .

Answer:

Sign Conventions

- If the temperature increases:
 - Q and ΔT are positive
 - Energy transfers into the system
- If the temperature decreases:
 - Q and ΔT are negative
 - Energy transfers out of the system

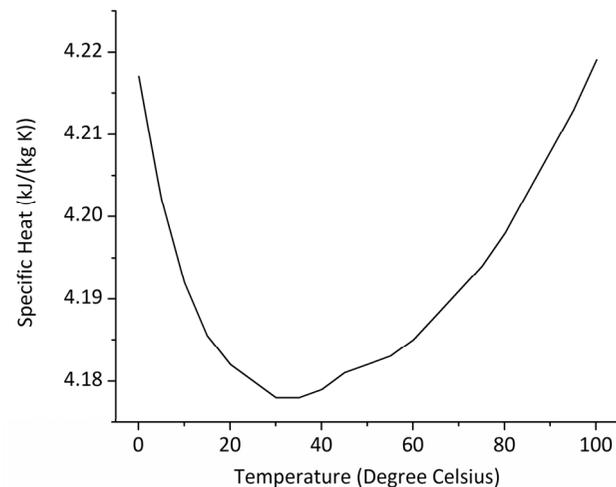
Temperature Dependent Specific Heat

- In real world situation, the specific heat is temperature dependent:

$$Q = m \int_{T_i}^{T_f} c(T) dT$$

- If the temperature intervals are not too large, the variation can be ignored and c can be treated as a constant.
 - For example, for water there is only about a 1% variation between 0° and 100°C .

Temperature Dependent Specific Heat of Water



On-Class Exercise

Temperature Dependent Specific Heat

The table lists the temperature dependent specific heat of water. Find the total energy to raise 1 kg of water from 0.01°C to 100°C ,

- if the specific heat is a constant. Hint: you may use $c = 4.2 \text{ kJ}/(\text{kg}\cdot\text{K})$; and
- if the specific heat is temperature dependent (please refer to table).
- What is the percentage error in calculating the total energy?

| Temperature (°C) | Specific Heat [kJ/(kg·K)] |
|------------------|---------------------------|
| 0.01 | 4.217 |
| 4 | 4.205 |
| 5 | 4.202 |
| 10 | 4.192 |
| 15 | 4.1855 |
| 20 | 4.182 |
| 25 | 4.18 |
| 30 | 4.178 |
| 35 | 4.178 |
| 40 | 4.179 |
| 45 | 4.181 |
| 50 | 4.182 |
| 55 | 4.183 |
| 60 | 4.185 |
| 65 | 4.188 |
| 70 | 4.191 |
| 75 | 4.194 |
| 80 | 4.198 |
| 85 | 4.203 |
| 90 | 4.208 |
| 95 | 4.213 |
| 100 | 4.219 |

Solution to Temperature Dependent Specific Heat



Specific Heat of Water

- Water has the highest specific heat of common materials.
- Responsible for many weather phenomena:
 - Moderate temperatures near large bodies of water;
 - Global wind systems;
 - Land and sea breezes.

Calorimetry

- Calorimetry is a technique to measure specific heat. Calorimeter is a instrument for measuring the quantity of heat exchanged.
- Procedure:
 - Heating a material;
 - Adding the material to a sample of water, and
 - Recording the final temperature.

Calorimetry

- The system of the sample and the water is isolated (without loss of energy).
- Conservation of energy requires that the amount of energy that leaves the sample equals the amount of energy that enters the water.

- Mathematically: $Q_{\text{Cold}} = -Q_{\text{Hot}}$
 $m_w c_w (T_f - T_w) = -m_x c_x (T_f - T_x)$
 $c_x = \frac{m_w c_w (T_f - T_w)}{m_x (T_x - T_f)}$

On-Class Exercise: Calorimetry

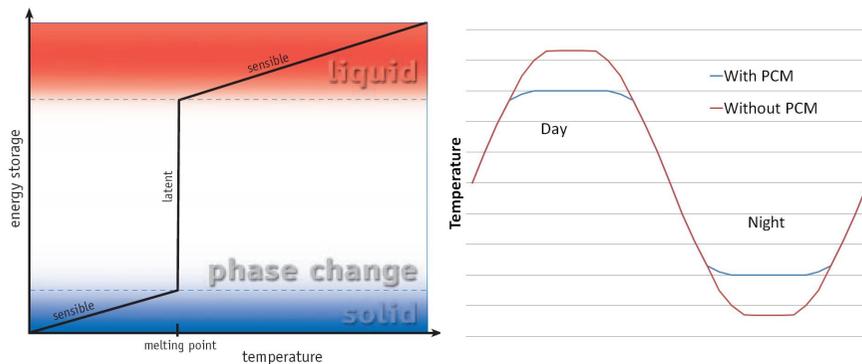
An ingot of metal (0.05 kg) is heated to 200°C and then dropped into a beaker of water (0.4 kg, initially at 20°C). The final equilibrium temperature of the mixture is 22.4°C. Find specific heat of the metal.



Phase Change

- Phase change: A substance changes from one form to another.
- Two common phase changes are
 - Solid to liquid or liquid to solid.
 - Liquid to gas or gas to liquid.
- During a phase change, there is no change in temperature.
 - For example, in boiling the increase in internal energy is represented by the breaking of the bonds between molecules, giving the molecules of the gas a higher intermolecular potential energy.

Application: Phase Change Materials for Building Energy Saving



Latent Heat

- Different substances react differently to the energy added or removed during a phase change.
 - Due to their different internal molecular arrangements.
- The amount of energy also depends on the mass of the sample.
- If an amount of energy Q is required to change the phase of a sample of mass m ,
 $L \equiv Q/m$.

Latent Heat

- Latent heat of fusion: phase change from solid to liquid.
- Latent heat of vaporization: phase change is from liquid to gas.
- The positive sign: the energy is transferred into the system.
 - This will result in melting or boiling.
- The negative sign: energy is transferred out of the system.
 - This will result in freezing or condensation.

Latent Heat Values

TABLE 20.2

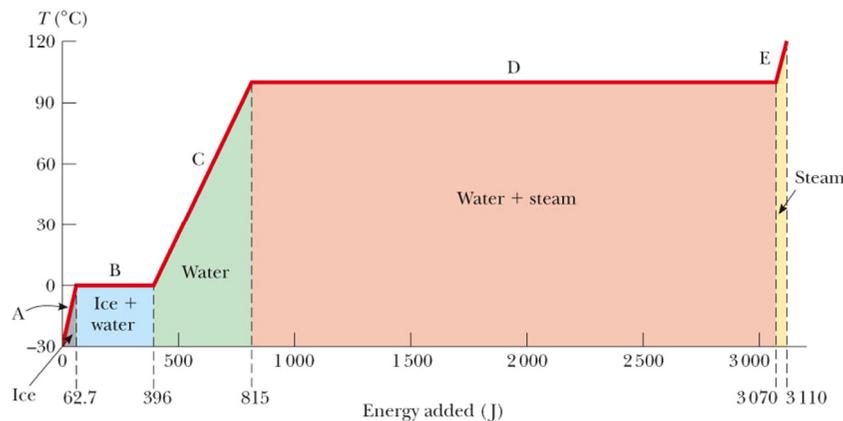
Latent Heats of Fusion and Vaporization

| 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 |

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Example: Ice to Steam

Consider the energy required to convert 1 g of ice at -30°C to steam:



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Ice Thermal Energy Storage

- Ice thermal energy storage system at the EMSD (Electrical and Mechanical Services Department) Headquarters.
- Five 110 m^3 ice storage tanks.
- Made by a 700 kW ammonia ice maker.

Exercise:

How much the “ice energy” that is stored in the 5 ice storage tanks, in comparison to ambient temperature?



Supercooling

- Supercooling: If liquid water is held perfectly still, it is possible for the temperature to drop below 0°C without freezing.
- It arises because the water requires a disturbance of some sort for the molecules to move apart and start forming the open ice crystal structures.
 - This structure makes the density of ice less than that of water.
- If the supercooled water is disturbed, it immediately freezes and the energy released returns the temperature to 0°C .

Supercooling



Superheating

- Superheated water: liquid water under pressure at temperatures between the usual boiling point (100°C) and the critical temperature (374°C).



Source: Wikipedia

Work and Heat in Thermodynamic Processes

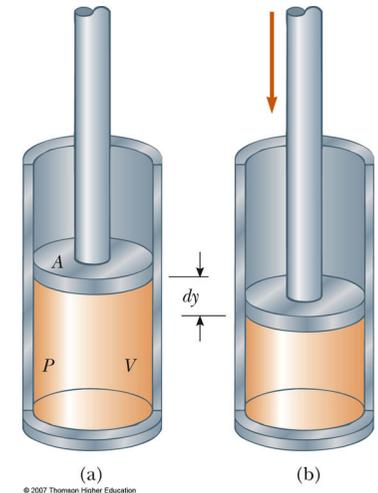
- State variables describe the state of a system.
- In the macroscopic approach to thermodynamics, variables are used to describe the state of the system.
 - Examples of state variables: Pressure, temperature, volume, internal energy.
- The macroscopic state of an isolated system can be specified only if the system is in thermal equilibrium internally.
 - For a gas in a container, this means every part of the gas must be at the same pressure and temperature.

Work and Heat in Thermodynamic Processes

- Transfer variables are zero unless a process occurs in which energy is transferred across the boundary of a system.
- Transfer variables are not associated with any given state of the system, only with changes in the state.
 - Heat and work are transfer variables.

Work in Thermodynamics

- Work can be done on a deformable system, such as a gas.
- Consider a cylinder with a moveable piston.
- A force is applied to slowly compress the gas.
 - The compression is slow enough for all the system to remain essentially in thermal equilibrium.
 - This is said to occur quasi-statically.



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Work in Thermodynamics

- The piston is pushed downward by a force (F) through a displacement of (dy):

$$\begin{aligned} dW &= \vec{F} \cdot d\vec{r} \\ &= -F\hat{j} \cdot dy\hat{j} \\ &= -Fdy \\ &= -PA dy \end{aligned}$$

- $A \cdot dy$ is the change in volume of the gas (dV).
- Therefore, the work done on the gas is:

$$dW = -PdV$$

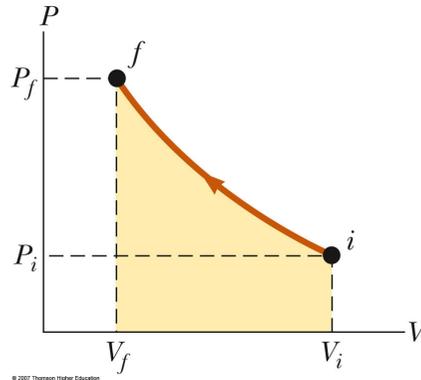
Work in Thermodynamics

- Interpretation of $dW = -PdV$
 - If the gas is compressed, dV is negative and the work done on the gas is positive.
 - If the gas expands, dV is positive and the work done on the gas is negative.
 - If the volume remains constant, the work done is zero.
- The total work done is:

$$W = - \int_{V_i}^{V_f} PdV$$

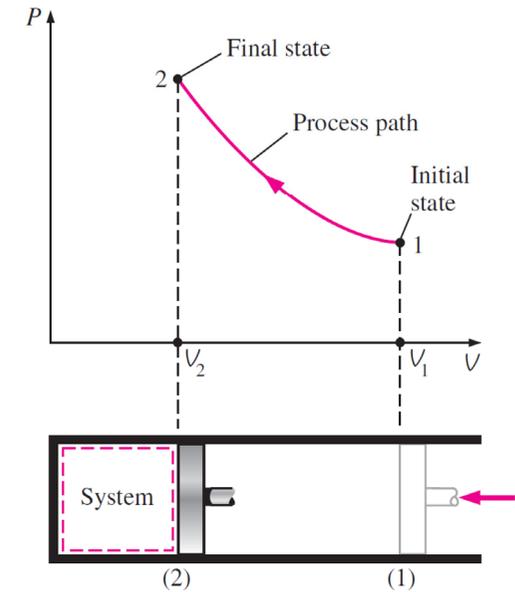
PV Diagrams

- *PV* diagram: The state of the gas.
- Allows us to visualize the process through which the gas is progressing.
- The curve is called the path.



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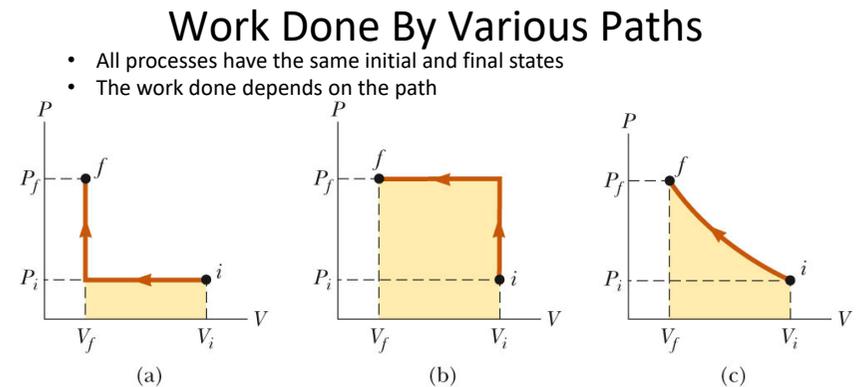
PV Diagrams



Example:
The P-V diagram of a
compression process.

PV Diagrams

- The work done on a gas in a quasi-static process that takes the gas from an initial state to a final state is the negative of the area under the curve on the *PV* diagram, evaluated between the initial and final states.
 - This is true whether or not the pressure stays constant.
 - The work done does depend on the path taken.



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a) Work Done, $W_a = -P_i(V_f - V_i)$

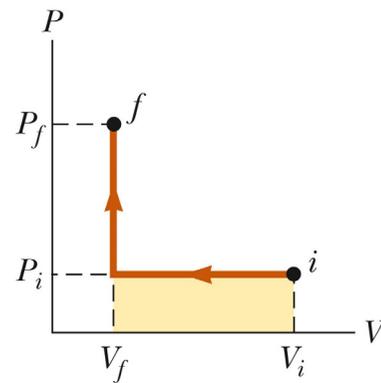
b) Work Done, $W_b = -P_f(V_f - V_i)$

c) Work Done, $W_c = -P(V) \cdot (V_f - V_i)$

Work done depends on the path between the initial and final states: $W_b > W_c > W_a$

Work From a PV Diagram: Example 1

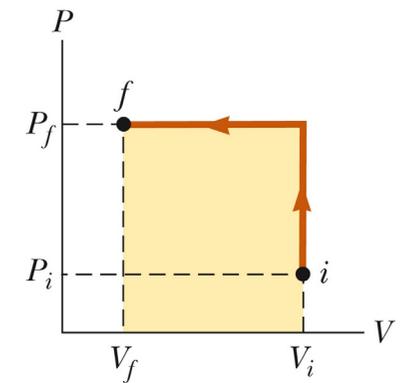
- The volume of the gas is first reduced from V_i to V_f at constant pressure P_i .
- Next, the pressure increases from P_i to P_f by heating at constant volume V_f .
- $W = -P_i(V_f - V_i)$



(a)

Work From a PV Diagram: Example 2

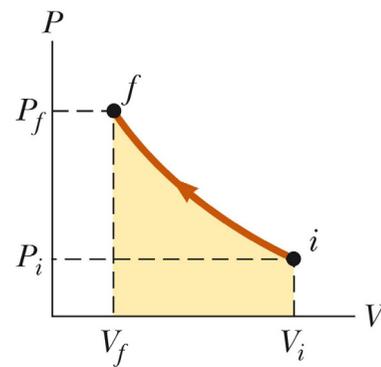
- The pressure of the gas is increased from P_i to P_f at a constant volume.
- The volume is decreased from V_i to V_f .
- $W = -P_f(V_f - V_i)$



(b)

Work From a PV Diagram: Example 3

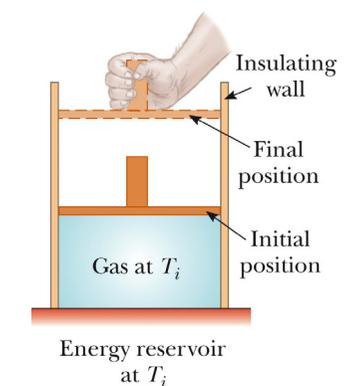
- The pressure and the volume continually change.
- The work is between $-P_f(V_f - V_i)$ and $-P_i(V_f - V_i)$.
- To evaluate the actual amount of work, the function $P(V)$ must be known.



(c)

Heat Transfer: Example 1

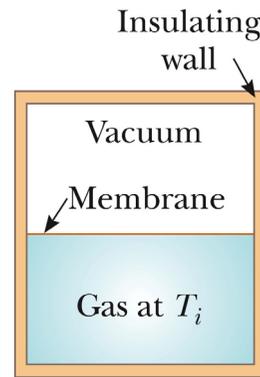
- The energy transfer, Q , into or out of a system also depends on the process.
- The energy reservoir is a source of energy that is considered to be so great that a finite transfer of energy does not change its temperature.
- The piston is pushed upward, the gas is doing work on the piston.



(a)

Heat Transfer: Example 2

- This gas has the same initial volume, temperature and pressure as the previous example.
- The final states are also identical.
- No energy is transferred by heat through the insulating wall.
- No work is done by the gas expanding into the vacuum.



(b)

Summary: Energy Transfer

- Energy transfers by heat, like the work done, depend on the initial, final, and intermediate states of the system.
- Both work and heat depend on the path taken.
- Neither can be determined solely by the end points of a thermodynamic process.

The First Law of Thermodynamics

- The First Law of Thermodynamics is a special case of the Law of Conservation of Energy.
 - the only type of energy that changes in a system is internal energy and the energy transfers are by heat and work.

- The First Law of Thermodynamics states that:

$$\Delta E_{int} = Q + W.$$

- For infinitesimal changes in a system: $dE_{int} = dQ + dW$.

Isolated Systems

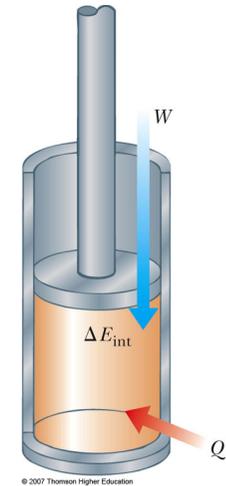
- An isolated system is one that does not interact with its surroundings.
 - No energy transfer by heat takes place;
 - Work done on the system is zero;
 - $Q = W = 0 \quad \Leftrightarrow \quad \Delta E_{int} = 0$.
- The internal energy of an isolated system remains constant.

Cyclic Processes

- A cyclic process is one that starts and ends in the same state.
 - This process would not be isolated.
 - On a PV diagram, a cyclic process appears as a closed curve.
- The internal energy must be zero since it is a state variable.
- If $\Delta E_{\text{int}} = 0$, $Q = -W$.
- In a cyclic process, the net work done on the system per cycle equals the area enclosed by the path representing the process on a PV diagram.

Adiabatic Process

- An adiabatic process is one during which no energy enters or leaves the system by heat.
 - $Q = 0$.
 - This is achieved by:
 - Thermally insulating the walls of the system.
 - Having the process proceed so quickly that no heat can be exchanged.



Adiabatic Process

- Since $Q = 0$, $\Delta E_{\text{int}} = W$.
- If the gas is compressed adiabatically, W is positive so ΔE_{int} is positive and the temperature of the gas increases.
- If the gas expands adiabatically, the temperature of the gas decreases.

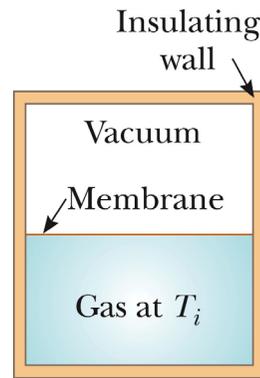
Adiabatic Process

Examples of adiabatic processes:

- The expansion of hot gases in an internal combustion engine.
- The liquefaction of gases in a cooling system.
- The compression stroke in a diesel engine.

Adiabatic Free Expansion

- The process is adiabatic because it takes place in an insulated container.
- Because the gas expands into a vacuum, it does not apply a force on a piston and $W = 0$.
- Since $Q = 0$ and $W = 0$, $\Delta E_{\text{int}} = 0$ and the initial and final states are the same.
 - No change in temperature is expected.



(b)

Isobaric Process

- Isobaric process: Constant pressure.
- The values of the heat and the work are generally both nonzero.
- The work done is $W = -P(V_f - V_i)$ where P is the constant pressure.

Isovolumetric Process

- Isovolumetric process: No change in the volume.
- Since the volume does not change, $W = 0$.
- From the first law, $\Delta E_{\text{int}} = Q$.
- If energy is added by heat to a system kept at constant volume, all of the transferred energy remains in the system as an increase in its internal energy.

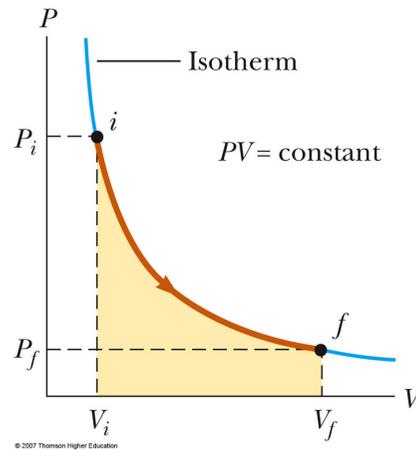
Isothermal Process

- Isothermal process: Constant temperature.
- Since there is no change in temperature, $\Delta E_{\text{int}} = 0$.
- Therefore, $Q = -W$.
- Any energy that enters the system by heat must leave the system by work.

Isothermal Process

Isotherm in PV diagram:

- Isothermal process; and
- Ideal gas yields a hyperbolic curve.



Quick Quiz

Characterize the paths (isobaric, isovolumetric, isothermal, or adiabatic). For path B, $Q = 0$.

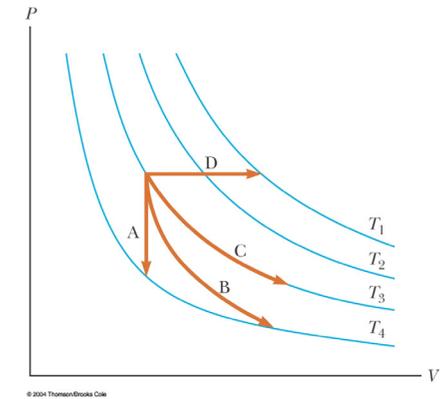
Answers:

Path A:

Path B:

Path C:

Path D:

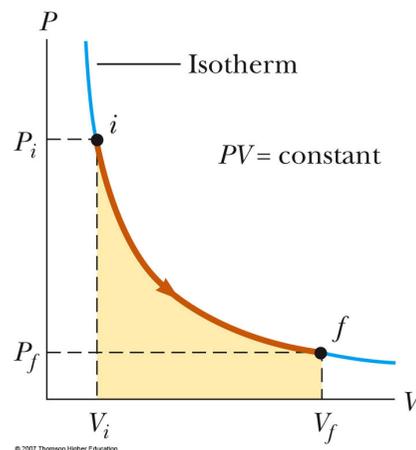


Exercise: Isothermal Expansion

The PV diagram exhibits $PV = \text{constant}$. Find the work done on gas in the process of isothermal expansion from state i to state f .

Remarks:

- If gas expands, $V_f > V_i$, work done on gas is negative.
- If gas is compressed, $V_f < V_i$, work done on gas is positive.



Summary on Special Processes

- **Adiabatic**
 - No heat exchanged.
 - $Q = 0$ and $\Delta E_{int} = W$.
- **Isobaric**
 - Constant pressure.
 - $W = P(V_f - V_i)$, and $\Delta E_{int} = Q + W$.
- **Isothermal**
 - Constant temperature.
 - $\Delta E_{int} = 0$ and $Q = -W$.
- **Isovolumetric**
 - Constant volume.
 - No work is done in a process.
 - $\Delta E_{int} = Q$.

Mechanisms of Energy Transfer by Heat

- Major mechanisms for energy transfer:
 - Conduction
 - Convection
 - Radiation

Conduction

- Atomic scale:
 - Exchange of kinetic energy between microscopic particles by collisions.
 - Microscopic particles: atoms, molecules, or free electrons.
 - Less energetic particles gain energy during collisions with more energetic particles.
- Rate of conduction depends upon the characteristics of the substance.

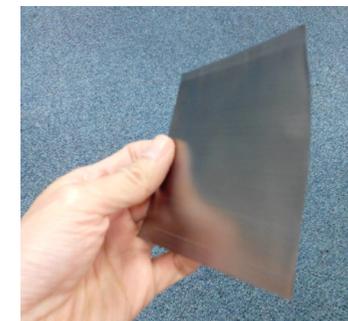
Conduction

- In general, metals are good thermal conductors.
 - Contain large numbers of electrons that are relatively free to move through the metal.
 - Transport energy from one region to another
- Poor conductors include paper and gases.
- Conduction can occur only if there is a difference in temperature between two parts of the conducting medium.

Applications: Materials with Good/Poor Thermal Conductivity



Since air is a good thermal insulator (thermal conductivity of $0.025 \text{ Wm}^{-1}\text{K}^{-1}$), double glazed windows (or double layer glass windows) with still air in between are applied to minimize the heat flow (and noise) between indoors and outdoors.



Graphite sheet is a very good thermal conduction material (thermal conductivity of $700 \text{ Wm}^{-1}\text{K}^{-1}$), which acts as heat sink for heat dissipation in integrated circuits.

List of Thermal Conductivities

| Substance | Thermal Conductivity (W/m · °C) |
|---------------------------------------|---------------------------------|
| <i>Metals (at 25°C)</i> | |
| Aluminum | 238 |
| Copper | 397 |
| Gold | 314 |
| Iron | 79.5 |
| Lead | 34.7 |
| Silver | 427 |
| <i>Nonmetals (approximate values)</i> | |
| Asbestos | 0.08 |
| Concrete | 0.8 |
| Diamond | 2 300 |
| Glass | 0.8 |
| Ice | 2 |
| Rubber | 0.2 |
| Water | 0.6 |
| Wood | 0.08 |
| <i>Cases (at 20°C)</i> | |
| Air | 0.023 4 |
| Helium | 0.138 |
| Hydrogen | 0.172 |
| Nitrogen | 0.023 4 |
| Oxygen | 0.023 8 |

Reference:
https://en.wikipedia.org/wiki/Thermoelectric_materials



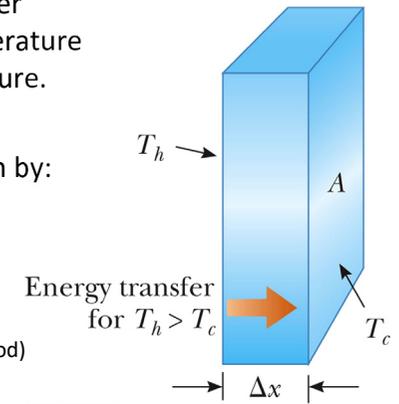
Conduction

- The slab allows energy to transfer from the region of higher temperature to the region of lower temperature.

- The rate of heat transfer is given by:

$$\frac{dQ}{dt} = -kA \frac{dT}{dx}$$

A = Cross-sectional area.
 Δx = Thickness of the slab (or length of a rod)
 Q is in Joules and t is in seconds
 k is the thermal conductivity

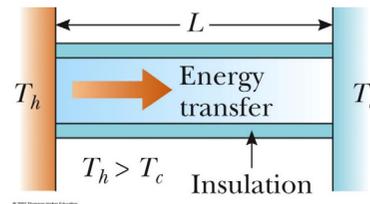


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Temperature Gradient

- The quantity $|dT/dx|$ is called the temperature gradient.
 - It measures the rate at which temperature varies with position.
- For a rod, the temperature gradient can be expressed as:

$$\left| \frac{dT}{dx} \right| = \frac{T_h - T_c}{L}$$



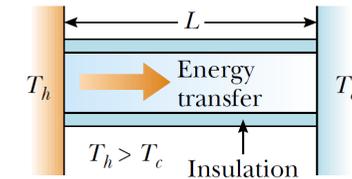
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Thermal Resistance

- Thermal Resistance R:

$$R = \frac{L}{k}$$

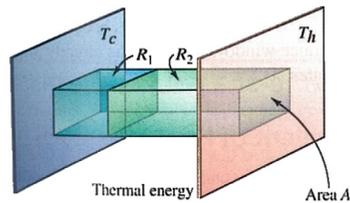
$$\frac{\Delta Q}{\Delta t} = \frac{1}{R} A \Delta T$$



Thermal Resistance in Series and in Parallel

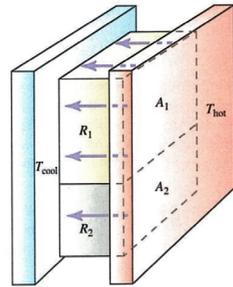
- Thermal Resistance in Series:

$$R_{\text{eff}} = R_1 + R_2$$



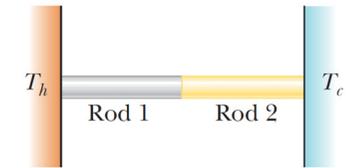
- Thermal Resistance in Parallel:

$$\frac{1}{R_{\text{eff}}} = \frac{1}{A_1 + A_2} \left(\frac{A_1}{R_1} + \frac{A_2}{R_2} \right)$$

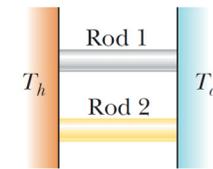


Quick Quiz on Thermal Conductivity

Two rods of the same length and diameter with different materials are used to connect two regions of different temperature such that energy transfer through the rods by heat. They can be connected in series or in parallel. In which case is the rate of energy transfer by heat larger?



(a)

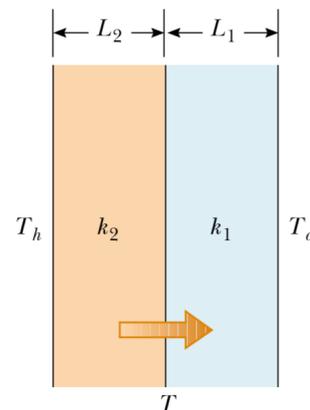


(b)

Answer:

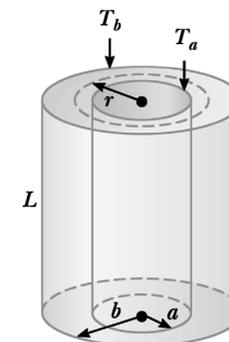
Exercise on Thermal Conductivity I

Two slabs, thickness L_1 and L_2 and thermal conductivities k_1 and k_2 , are in thermal contact with each other. The temperatures of their outer surfaces are T_c and T_h , respectively ($T_h > T_c$). Determine the temperature, T , at the interface and the rate of energy transfer under steady-state condition.



Exercise on Thermal Conductivity II

The following hollow cylinder is filled with water and maintained at temperature T_a , whereas the outer surface is T_b ($T_a > T_b$). The wall of the cylinder has a thermal conductivity k . Find the rate of energy transfer from the inner wall to the outer surface.



Exercise on Thermal Conductivity and Latent Heat

- A pond of water at 0°C is covered with a layer of ice 4 cm thick. If the air temperature stays constant at -10°C, how long does it take for the ice thickness to increase to 8 cm?

Hints:

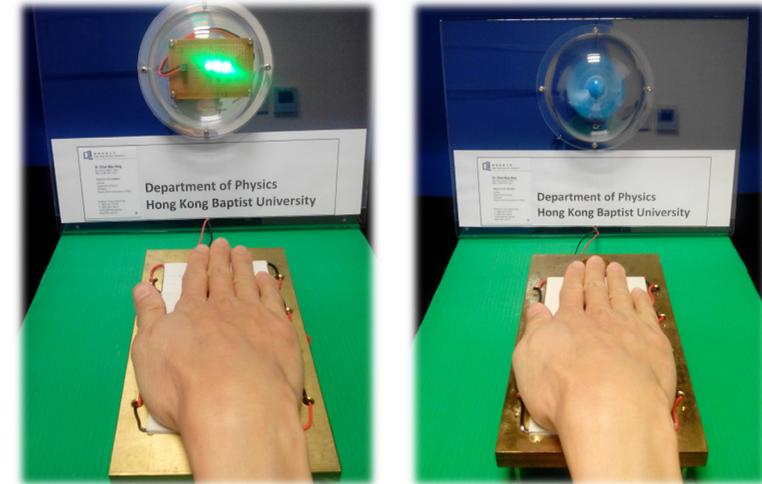
Latent Heat fusion of water: 334 kJ/kg

Latent heat vaporization of water: 2260 kJ/kg

Density of ice: 917 kg/m³

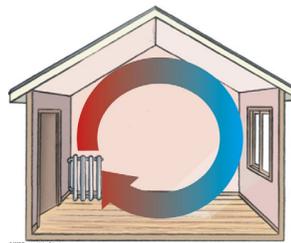
Thermal conductivity of ice: 2 W/mK

Applications and Demonstrations: Thermal Energy and Thermal Conductivity



Convection

- Energy transferred by the movement of a substance.
 - Natural convection: movement results from differences in density.
 - Forced convection: movement is forced by a fan or a pump.
- Example of Convection:
 - Air directly above the radiator is warmed and expands.
 - Air density decreases, and it rises.
 - A continuous air current is established.



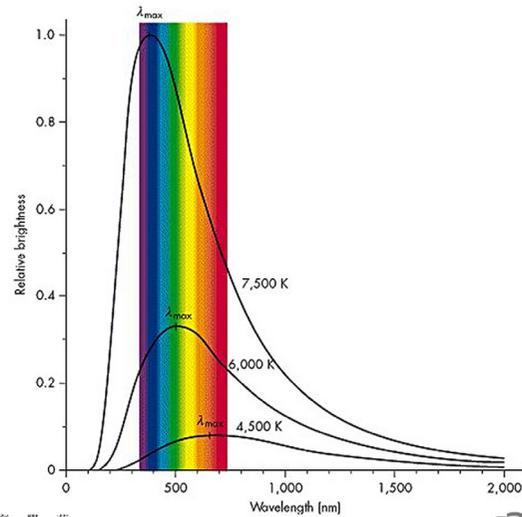
Blackbody Radiation

- Planck's law: the radiance L (in $\text{Wm}^{-2}\text{m}^{-1}\text{steradian}^{-1}$, Watt per unit area per unit wavelength per solid angle) at a temperature T is given by:

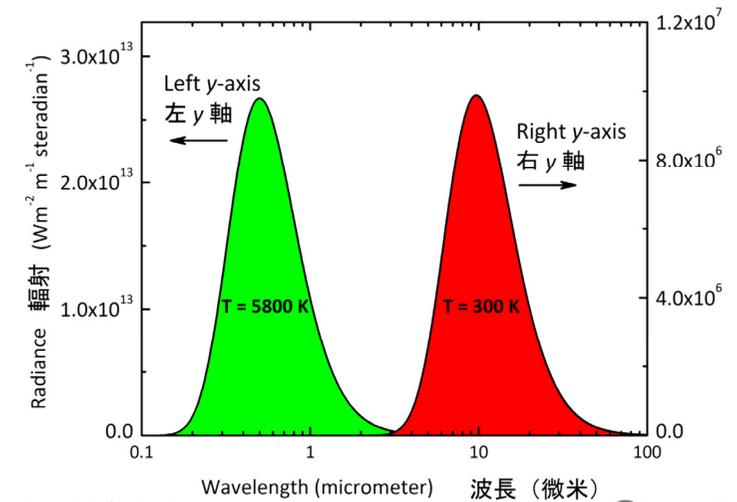
$$L(T) = \frac{2hc^2}{\lambda^5 [\exp(hc/k\lambda T) - 1]}$$

where h is the Planck's constant ($=6.63 \times 10^{-34}$ Js), c is the speed of light in vacuum ($=3 \times 10^8$ ms⁻¹), and k denotes the Boltzmann's constant (1.38×10^{-23} J/K).

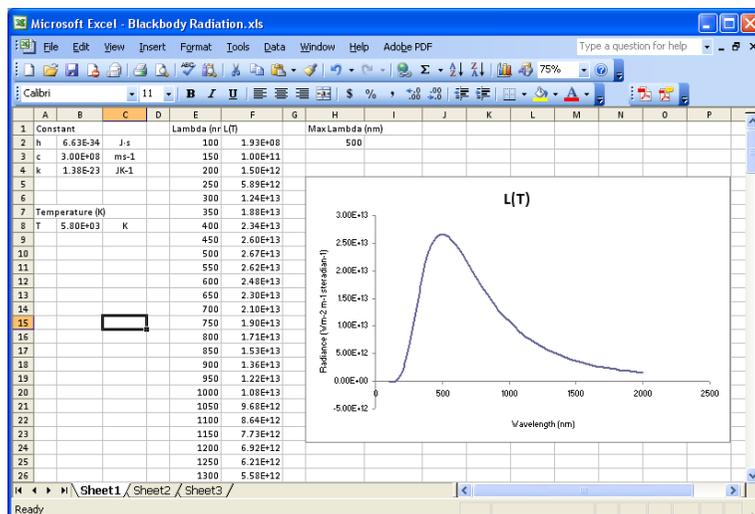
Blackbody Radiation



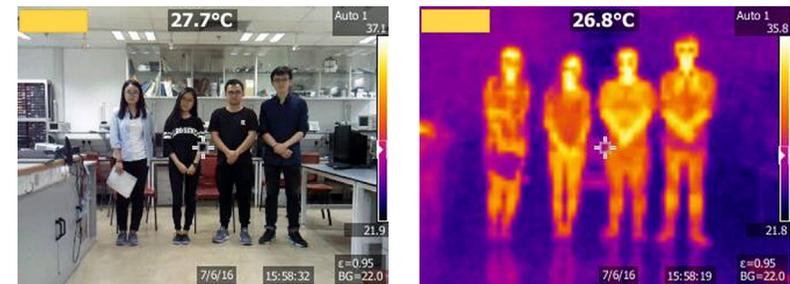
Blackbody Radiation



Blackbody Radiation



Blackbody: IR Image



Visible image (0.4 – 0.76 μm) and thermal infrared image (9 – 14 μm) of students at the same physics laboratory.

Radiation and Stefan's Law

- In isotropic radiation, the irradiance $E(T)$ (in Wm^{-2}) at a temperature T [in K] is given by:

$$E(T) = \int_0^{\infty} \pi \cdot L(T) d\lambda$$

$$= \sigma_{bb} T^4$$

where the Stefan-Boltzmann constant $\sigma_{bb} = 2\pi^5 k^4 / 15c^2 h^3$ is given by $5.67 \times 10^{-8} \text{ Wm}^{-2} \text{K}^{-4}$.

- $P = \sigma A e T^4$
 - P [in W] is the rate of energy transfer;
 - e is the emissivity (varies from 0 to 1);
 - The emissivity is also equal to the absorptivity

Wien Displacement Law

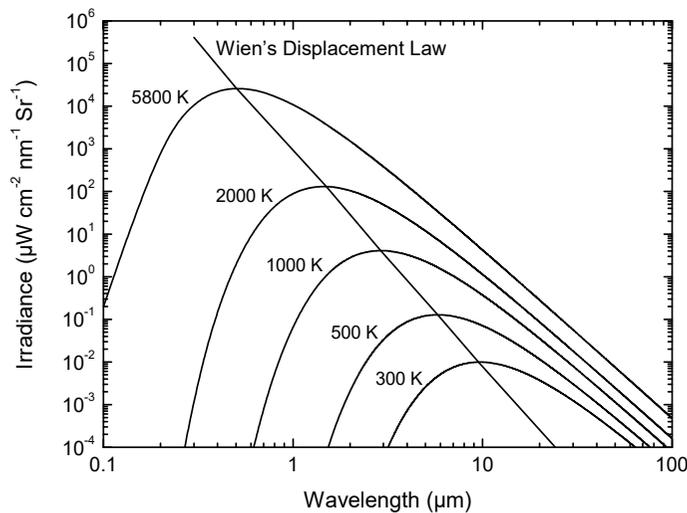
- Wien Displacement Law – The wavelength λ_{max} at which the maximum intensity emitted by a blackbody.
- λ_{max} can be derived by differentiating the radiance $L(T)$ with respect to wavelength under the condition:

$$dL(T)/d\lambda = 0$$

- The solution to the differentiation is:

$$\lambda_{\text{max}} \approx 2898/T \quad [\mu\text{m}\cdot\text{K}]$$

Wien Displacement Law



Exercise:

Estimation of Thermal Power Loss by Human

Body Surface Area (BSA)

- Du Bois Formula:

$$BSA = 0.007184 \times W^{0.425} \times H^{0.725}$$

- Mosteller Formula

$$BSA = \sqrt{\frac{W \times H}{3600}}$$

| | BSA |
|------------------|---------------------|
| Newborn | 0.25 m ² |
| Child of 2 years | 0.5 m ² |
| 9 years | 1.07 m ² |
| 10 years | 1.14 m ² |
| 12–13 years | 1.33 m ² |
| Women | 1.6 m ² |
| Men | 1.9 m ² |

Energy Absorption and Radiation Emission

- Considering an object (temperature T) in the surroundings (temperature T_o), the net rate at which the object radiates is:
 - $P_{\text{net}} = \sigma A e (T^4 - T_o^4)$.
 - When an object is in equilibrium with its surroundings, it radiates and absorbs at the same rate (temperature will not change).

Ideal Absorbers

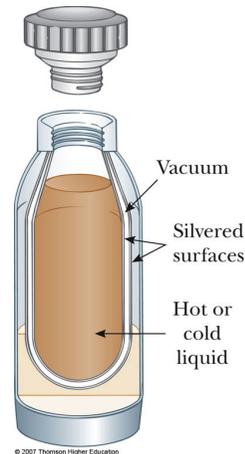
- An ideal absorber is defined as an object that absorbs all of the energy incident on it ($e = 1$).
- This type of object is called a black body.
- An ideal absorber is also an ideal radiator of energy.

Ideal Reflector

- An ideal reflector absorbs none of the energy incident on it ($e = 0$).

The Dewar Flask

- A Dewar flask is a container designed to minimize the energy losses by conduction, convection, and radiation.
 - Invented by Sir James Dewar (1842 – 1923)
- It is used to store either cold or hot liquids for long periods of time.
- The space between the walls is a vacuum to minimize energy transfer by conduction and convection.
- The silvered surface minimizes energy transfers by radiation (silver is a good reflector).
- The size of the neck is reduced to further minimize energy losses.



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Example: The Expected Surface Temperature of the Earth

- The expected surface temperature of the Earth can be estimated by means of radiative equilibrium.
- Radiative Equilibrium: Energy gained from the absorption of short wave radiation is equal to the emitted long wave radiation.
- Starting from the solar constant (the irradiance at the top of the Earth's atmosphere), $S_o = 1370 \text{ Wm}^{-2}$, the incident energy on the Earth is:

$$\text{Incident Energy} = S_o \times (\pi \times R_{\text{Earth}}^2)$$

Example: The Expected Surface Temperature of the Earth

- Not all the incident solar energy is absorbed. A portion is reflected back to space, as determined by the average reflectivity, or albedo A , of the Earth/atmosphere system.

Typical values of earth's albedo, (After Riehl, 1978)

| Surface type | Albedo (%) |
|---|------------|
| Forests | 3–10 |
| Fields, green | 3–15 |
| Fields, dry, plowed | 20–25 |
| Grass | 15–30 |
| Bare ground | 7–20 |
| Sand | 15–25 |
| Snow, fresh | 80 |
| Snow (old)/ice | 50–70 |
| Water, solar elevation $> 40^\circ$ | 2–4 |
| Water, solar elevation $30\text{--}5^\circ$ | 6–40 |

Example: The Expected Surface Temperature of the Earth

- Considering the albedo ($A \approx 0.3$),
the absorbed energy = $S_o \times (\pi \times R_{\text{Earth}}^2) \times (1 - A)$
- The Earth also emits long-wave (thermal) radiation back to the space,
the emitted energy = $(4\pi R_{\text{Earth}}^2) \times \sigma \times T^4$.
where the Stefan-Boltzmann constant $\sigma = 5.67 \times 10^{-8} \text{ Wm}^{-2}\text{K}^{-4}$.

Example: The Expected Surface Temperature of the Earth

- Under the condition of radiative equilibrium,

$$S_o \times (\pi \times R_{\text{Earth}}^2) \times (1 - A) = (4\pi R_{\text{Earth}}^2) \times \sigma \times T^4$$

$$S_o (1 - A) = 4 \sigma T^4$$

$$T = 255 \text{ K (or } -18^\circ\text{C)}$$

- However, the actual surface temperature is about 16°C .
- The inconsistency is due to the atmospheric greenhouse effect, where the greenhouse gases absorb the short-wave radiation and re-emit the long-wave radiation back to the Earth's surface.

Example: The Expected Surface Temperature of the Earth

- To illustrate the greenhouse effect on the Earth's surface temperature, the radiative balance equation can be modified by a greenhouse factor x (ranges from 0 to 1).

$$S_o (1 - A) = 4 \sigma T^4 (1 - x)$$

- For $x = 0$, there is no greenhouse effect.
- For $x \approx 0.4$, the equilibrium surface temperature is 17°C .

