Lecture PowerPoints

Chapter 14
Physics: Principles with Applications, 7th edition
Giancoli

Chapter 14

Heat

Contents of Chapter 14

14-1 Heat As Energy Transfer

We often speak of heat as though it were a material that flows from one object to another; it is not. Rather, it is a form of energy.

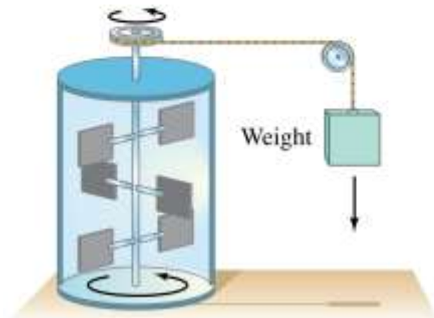
Unit of heat: calorie (cal)

1 cal is the amount of heat necessary to raise the temperature of 1 g of water by 1 Celsius degree.

Don't be fooled—the calories on our food labels are really kilocalories (kcal or Calories), the heat necessary to raise 1 kg of water by 1 Celsius degree.

14-1 Heat As Energy Transfer

If heat is a form of energy, it ought to be possible to equate it to other forms. The experiment below found the mechanical equivalent of heat by using the falling weight to heat the water:



4.186 J = 1 cal

4.186 kJ = 1 kcal

14-1 Heat As Energy Transfer

Definition of heat:

Heat is energy transferred from one object to another because of a difference in temperature.

- Remember that the temperature of a gas is a measure of the kinetic energy of its molecules.

14-2 Internal Energy

The sum total of all the energy of all the molecules in a substance is its internal (or thermal) energy.

Temperature: measures molecules' average kinetic energy

Internal energy: total energy of all molecules

Heat: transfer of energy due to difference in temperature

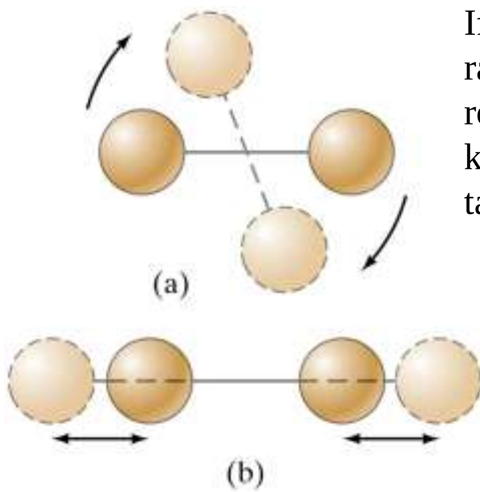
14-2 Internal Energy

Internal energy of an ideal (atomic) gas is equal to the average kinetic energy per molecule multiplied by the number of molecules.

But since we know the average kinetic energy in terms of the temperature, we can write:

$$U = \tfrac{3}{2}nRT, \qquad \left[\begin{array}{l}\text{internal energy of} \\ \text{ideal monatomic gas}\end{array}\right] \qquad (14\text{-}1)$$

14-2 Internal Energy



If the gas is molecular rather than atomic, rotational and vibrational kinetic energy needs to be taken into account as well.

14-3 Specific Heat

TABLE 14–1 Specific Heats (at 1 atm constant pressure and 20°C unless otherwise stated)

Substance	Specific Heat, c: J/kg · C°	Specific Heat, c: kcal/kg · C° (= cal/g · C°)
Aluminum	900	0.22
Alcohol (ethyl)	2400	0.58
Copper	390	0.093
Glass	840	0.20
Iron or steel	450	0.11
Lead	130	0.031
Marble	860	0.21
Mercury	140	0.033
Silver	230	0.056
Wood	1700	0.4
Water		
Ice (−5°C)	2100	0.50
Liquid (15°C)	4186	1.00
Steam (110°C)	2010	0.48
Human body (average)	3470	0.83
Protein	1700	0.4

The amount of heat required to change the temperature of a material is proportional to the mass and to the temperature change:

$$Q = mc\,\Delta T, \quad (14\text{-}2)$$

The specific heat, *c*, is characteristic of the material. Some values are listed at left.

Specific Heat: [c]= J/(kg*K)
Molar Heat Capacity: [C]= J/(mol*K)

$$Q = m\,c\,\Delta T$$

m: kg; c: $\frac{J}{kg\,K}$; ΔT: K

$$= n\,(M c)\,\Delta T$$

n: moles; M: molar mass

$$C = cM$$

Specific Heat: [c]= J/(kg*K)
Molar Heat Capacity: [C]= J/(mol*K)

- Molar heat capacity for elemental solids ~ 25 J/(mol*K)

Table 17.3 Approximate Specific Heats and Molar Heat Capacities (Constant Pressure)

Substance	Specific Heat, c (J/kg · K)	Molar Mass, M (kg/mol)	Molar Heat Capacity, C (J/mol · K)
Aluminum	910	0.0270	24.6
Beryllium	1970	0.00901	17.7
Copper	390	0.0635	24.8
Ethanol	2428	0.0461	111.9
Ethylene glycol	2386	0.0620	148.0
Ice (near 0°C)	2100	0.0180	37.8
Iron	470	0.0559	26.3
Lead	130	0.207	26.9
Marble ($CaCO_3$)	879	0.100	87.9
Mercury	138	0.201	27.7
Salt (NaCl)	879	0.0585	51.4
Silver	234	0.108	25.3
Water (liquid)	4190	0.0180	75.4

Specific Heat: [c]= J/(kg*K)
Molar Heat Capacity: [C]= J/(mol*K)

- Rule of Dulong and Petit
  - The molar heat capacity for elemental solids is approximately constant 25 J/(mol*K)
  - number of atoms per mole is N_A, heat needed to raise temperature of one atom same.
  - Atomic mass does not matter, only *how many atoms* matters.
  - Why? 1 molecule's average KE = 3/2 kT
  - Total energy per atom = KE + PE = 3kT
    - SHM, PE like vibrating in 3D, equipartition of energy principle: each velocity component contributes 1/2kT of energy per molecule

Specific Heat: [c]= J/(kg*K)
Molar Heat Capacity: [C]= J/(mol*K)

- Rule of Dulong and Petit
  - The molar heat capacity for elemental solids is approximately constant 25 J/(mol*K)
  - Atomic mass does not matter, only *how many atoms* matters.
  - Why? 1 molecule's average KE = 3/2 kT
  - Total energy per atom = KE + PE = 3kT
  - total energy in one mole = $3N_AkT = 3RT$
  - Heat needed to raise temperature by 1K is

    $C_V = 3R \sim 24.9$ J/(mol*K)

  (volume constant so that all energy goes internally, not used up to do work of expanding the volume)

Test Your Understanding of Section 17.5 You wish to raise the temperature of each of the following samples from 20°C to 21°C. Rank these in order of the amount of heat needed to do this, from highest to lowest. (i) 1 kilogram of mercury; (ii) 1 kilogram of ethanol; (iii) 1 mole of mercury; (iv) 1 mole of ethanol.

14-3 Specific Heat

Heat in = Increase in Internal Energy – Work done on piston

Specific heats of gases are more complicated, and are generally measured at constant pressure (c_P) or constant volume (c_V).

Solids: usually measure c_P at constant atmospheric pressure

Gas: c_V usually easier to measure

c_V and c_P are different especially for gas

c_V <u>< = ></u> c_P

14-3 Specific Heat

Specific heats of gases are more complicated, and are generally measured at constant pressure (c_P) or constant volume (c_V).

Some sample values:

TABLE 14–2
Specific Heats of Gases
(kcal/kg · C°)

Gas	c_P (constant pressure)	c_V (constant volume)
Steam (100°C)	0.482	0.350
Oxygen	0.218	0.155
Helium	1.15	0.75
Carbon dioxide	0.199	0.153
Nitrogen	0.248	0.177

14-4 Calorimetry—Solving Problems

Closed system: no mass enters or leaves, but energy may be exchanged

Open system: mass may transfer as well

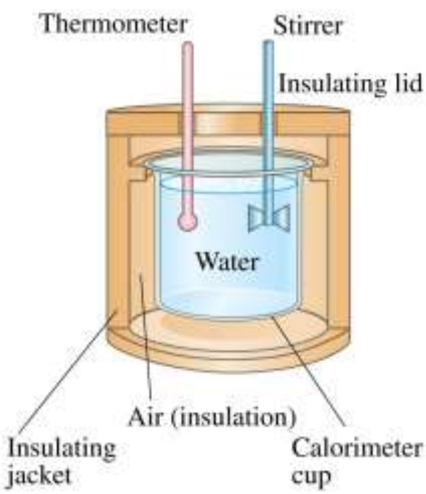
Isolated system: closed system where no energy in any form is transferred

For an isolated system,

Energy out of one part = energy into another part

Or: heat lost = heat gained

14-4 Calorimetry – Solving Problems



The instrument to the left is a calorimeter, which makes quantitative measurements of heat exchange. A sample is heated to a well-measured high temperature, plunged into the water, and the equilibrium temperature measured. This gives the specific heat of the sample.

14-4 Calorimetry—Solving Problems

Another type of calorimeter is called a bomb calorimeter; it measures the thermal energy released when a substance burns.

This is the way the caloric content of foods is measured.

Example 5

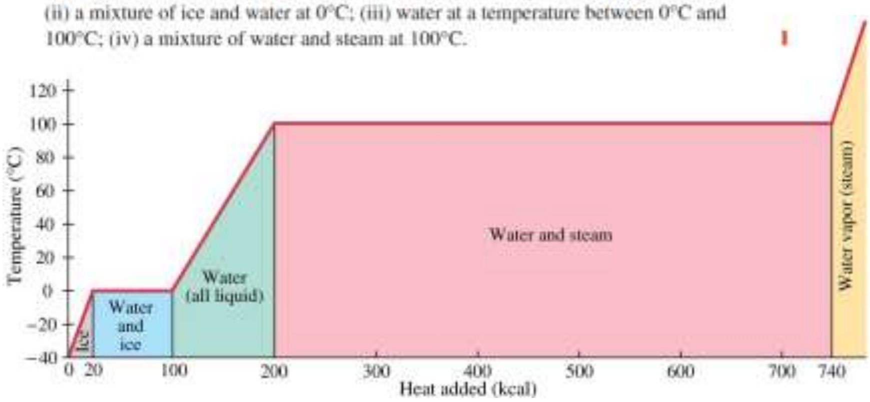
- 0.150 kg sample of an alloy is heated to 540°C.
- It is then quickly placed in 0.400 kg of water at 10.0°C
  - c_water = 4186 J/(kgK)
- which is contained in an aluminum cup 0.200 kg
  - c_Al = 900 J/(kgK)
- The final temperature of the system is 30.5°C
- What is the specific heat of the alloy?

17.20 The metal gallium, shown here melting in a person's hand, is one of the few elements that melt in the vicinity of room temperature. Its melting temperature is 29.8°C, and its heat of fusion is 8.04×10^4 J/kg.

14-5 Latent Heat

Energy is required for a material to change phase, even though its temperature is not changing.

Test Your Understanding of Section 17.6 You take a block of ice at 0°C and add heat to it at a steady rate. It takes a time t to completely convert the block of ice to steam at 100°C. What do you have at time $t/2$? (i) all ice at 0°C; (ii) a mixture of ice and water at 0°C; (iii) water at a temperature between 0°C and 100°C; (iv) a mixture of water and steam at 100°C.



14-5 Latent Heat

Heat of fusion, L_F: heat required to change 1.0 kg of material from solid to liquid

Heat of vaporization, L_V: heat required to change 1.0 kg of material from liquid to vapor

$$Q = mL$$

TABLE 14–3 Latent Heats (at 1 atm)

Substance	Melting Point (°C)	Heat of Fusion kJ/kg	Heat of Fusion kcal/kg[†]	Boiling Point (°C)	Heat of Vaporization kJ/kg	Heat of Vaporization kcal/kg[†]
Oxygen	−218.8	14	3.3	−183	210	51
Nitrogen	−210.0	26	6.1	−195.8	200	48
Ethyl alcohol	−114	104	25	78	850	204
Ammonia	−77.8	33	8.0	−33.4	137	33
Water	0	333	79.7	100	2260	539
Lead	327	25	5.9	1750	870	208
Silver	961	88	21	2193	2300	558
Iron	1538	289	69.1	3023	6340	1520
Tungsten	3410	184	44	5900	4800	1150

[†]Numerical values in kcal/kg are the same in cal/g.

Water: fusion 333, vaporization 2260.

Makes sense?

Evaporation as a cooling process?

Example 6

- 1.5 kg water at 20^{o}C
- Turn into ice, -12^{o}C
- Energy needed?

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Example 7

- 0.50 kg ice at -10^{o}C
- 3.0 kg tea (water) at 20^{o}C
- a) Will the final mixture be solid or liquid?
- b) What is the final temperature of the mixture?

14-5 Latent Heat

The total heat required for a phase change depends on the total mass and the latent heat:

$$Q = mL, \quad (14\text{-}4)$$

Problem Solving: Calorimetry

1. Is the system isolated? Are all significant sources of energy transfer known or calculable?
2. Apply conservation of energy.
3. If no phase changes occur, the heat transferred will depend on the mass, specific heat, and temperature change.

14-5 Latent Heat

4. If there are, or may be, phase changes, terms that depend on the mass and the latent heat may also be present. Determine or estimate what phase the final system will be in.
5. Make sure that each term is in the right place and that all the temperature changes are positive.
6. There is only one final temperature when the system reaches equilibrium.
7. Solve.

14-5 Latent Heat

The latent heat of vaporization is relevant for evaporation as well as boiling. The heat of vaporization of water rises slightly as the temperature decreases.

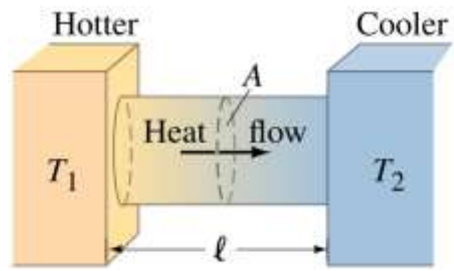
On a molecular level, the heat added during a change of state does not go to increasing the kinetic energy of individual molecules, but rather to break the close bonds between them so the next phase can occur.

14-6 Heat Transfer: Conduction

Heat conduction can be visualized as occurring through molecular collisions.

The heat flow per unit time is given by:

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{\ell} \quad (14\text{-}5)$$



Protective tile (ceramic) used in space shuttle

- thermal conductivity: low or high?
- heat capacity: low or high?

14-6 Heat Transfer: Conduction

TABLE 14–4
Thermal Conductivities

Substance	Thermal Conductivity, k	
	J / (s · m · C°)	kcal / (s · m · C°)
Silver	420	10×10^{-2}
Copper	380	9.2×10^{-2}
Aluminum	200	5.0×10^{-2}
Steel	40	1.1×10^{-2}
Ice	2	5×10^{-4}
Glass	0.84	2.0×10^{-4}
Brick	0.84	2.0×10^{-4}
Concrete	0.84	2.0×10^{-4}
Water	0.56	1.4×10^{-4}
Human tissue	0.2	0.5×10^{-4}
Wood	0.1	0.3×10^{-4}
Fiberglass	0.048	0.12×10^{-4}
Cork	0.042	0.10×10^{-4}
Wool	0.040	0.10×10^{-4}
Goose down	0.025	0.060×10^{-4}
Polyurethane	0.024	0.057×10^{-4}
Air	0.023	0.055×10^{-4}

The constant k is called the thermal conductivity.

Materials with large k are called conductors; those with small k are called insulators.

Air is a good insulator!

Wool sweater, fur

Application Fur Versus Blubber

The fur of an arctic fox is a good thermal insulator because it traps air, which has a low thermal conductivity k. (The value $k = 0.04$ W/m·K for fur is higher than for air, $k = 0.024$ W/m·K, because fur also includes solid hairs.) The layer of fat beneath a bowhead whale's skin, called blubber, has six times the thermal conductivity of fur ($k = 0.24$ W/m·K). So a 6-cm thickness of blubber ($L = 6$ cm) is required to give the same insulation as 1 cm of fur.

Conduction

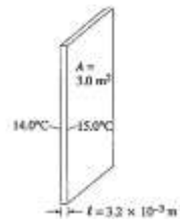
EXAMPLE 14–8 **Heat loss through windows.** A major source of heat loss from a house in cold weather is through the windows. Calculate the rate of heat flow through a glass window 2.0 m × 1.5 m in area and 3.2 mm thick, if the temperatures at the inner and outer surfaces are 15.0°C and 14.0°C, respectively (Fig. 14–7).

APPROACH Heat flows by conduction through the 3.2-mm thickness of glass from the higher inside temperature to the lower outside temperature. We use the heat conduction equation, Eq. 14–5.

SOLUTION Here $A = (2.0\,\text{m})(1.5\,\text{m}) = 3.0\,\text{m}^2$ and $\ell = 3.2 \times 10^{-3}\,\text{m}$. Using Table 14–4 to get k, we have

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{\ell} = \frac{(0.84\,\text{J/s}\cdot\text{m}\cdot\text{C}^\circ)(3.0\,\text{m}^2)(15.0^\circ\text{C} - 14.0^\circ\text{C})}{(3.2 \times 10^{-3}\,\text{m})} = 790\,\text{J/s}.$$

NOTE This rate of heat flow is equivalent to $(790\,\text{J/s})/(4.19 \times 10^3\,\text{J/kcal}) = 0.19\,\text{kcal/s}$, or $(0.19\,\text{kcal/s}) \times (3600\,\text{s/h}) = 680\,\text{kcal/h}$.



14-6 Heat Transfer: Conduction

Building materials are measured using *R*–values rather than thermal conductivity:

$$R = \frac{\ell}{k}.$$

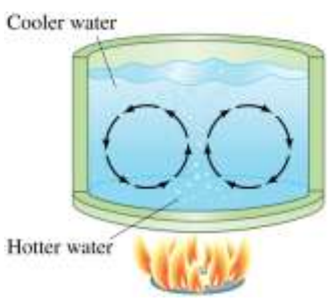
Here, *l* is the thickness of the material.

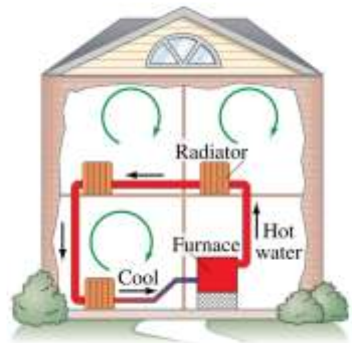
TABLE 14–5 *R*-values

Material	Thickness	*R*-value ($\text{ft}^2\cdot\text{h}\cdot\text{F}^\circ/\text{Btu}$)
Glass	$\frac{1}{8}$ inch	1
Brick	$3\frac{1}{2}$ inches	0.6–1
Plywood	$\frac{1}{2}$ inch	0.6
Fiberglass insulation	4 inches	12

14-7 Heat Transfer: Convection

Convection occurs when heat flows by the mass movement of molecules from one place to another. It may be natural or forced; both these examples are natural convection.





14-7 Heat Transfer: Convection

Many home heating systems are forced hot-air systems; these have a fan that blows the air out of registers, rather than relying completely on natural convection.

Our body temperature is regulated by the blood; it runs close to the surface of the skin and transfers heat. Once it reaches the surface of the skin, the heat is released through convection, evaporation, and radiation.

14-8 Heat Transfer: Radiation

The most familiar example of radiation is our own Sun, which radiates at a temperature of almost 6000 K.

14-8 Heat Transfer: Radiation

The energy radiated has been found to be proportional to the fourth power of the temperature:

$$\frac{Q}{t} = \epsilon\sigma A T^4. \quad (14\text{-}6)$$

The constant σ is called the Stefan-Boltzmann constant:

$$\sigma = 5.67 \times 10^{-8}\ \mathrm{W/m^2{\cdot}K^4}$$

The emissivity e is a number between zero and one characterizing the surface; black objects have an emissivity near one, while shiny ones have an emissivity near zero.

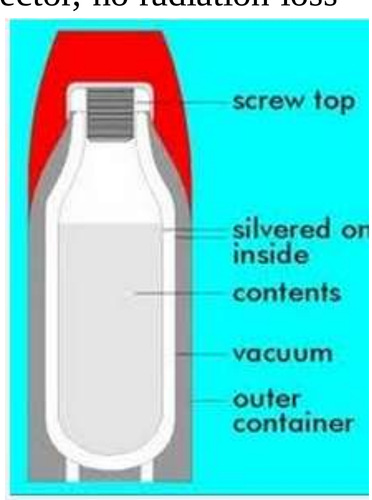
Radiation

- A good absorber is also a good emitter
- blackbody e =1

- A reflector e = 0 (shiny, light colored objects)
  - is a bad absorber and bad emitter

Dewar Flask

- vacuum – no conduction, convection
- silver – great reflector, no radiation loss



Radiation

EXAMPLE 14–9 | ESTIMATE | Cooling by radiation. An athlete is sitting unclothed in a locker room whose dark walls are at a temperature of 15°C. Estimate the body's rate of heat loss by radiation, assuming a skin temperature of 34°C and $\epsilon = 0.70$. Take the surface area of the body not in contact with the chair to be 1.5 m².

APPROACH We use Eq. 14–7, which requires Kelvin temperatures.

SOLUTION We have

$$\frac{Q}{t} = \epsilon\sigma A(T_1^4 - T_2^4)$$
$$= (0.70)(5.67 \times 10^{-8}\,\mathrm{W/m^2 \cdot K^4})(1.5\,\mathrm{m^2})[(307\,\mathrm{K})^4 - (288\,\mathrm{K})^4] = 120\,\mathrm{W}.$$

NOTE This person's "output" is a bit more than what a 100-W bulb uses.

NOTE Avoid a common error: $(T_1^4 - T_2^4) \neq (T_1 - T_2)^4$.

Radiation

- resting person internal metabolism 100W
- Example, 120W loss
- Lose more energy than you make!
- Uncomfortable even if air is 25°C
- radiation ~ 50% heat loss in normal room!
- Comfortable: Floors/walls warmer, even if air is not so warm. Incubators

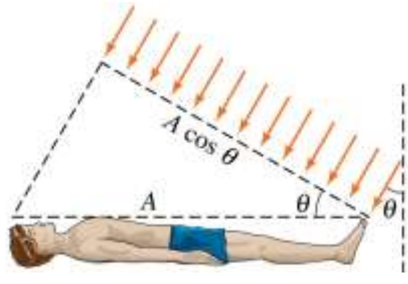
14-8 Heat Transfer: Radiation

If you are sitting in a place that is too cold, your body radiates more heat than it can produce. You will start shivering and your metabolic rate will increase unless you put on warmer clothing.

14-8 Heat Transfer: Radiation

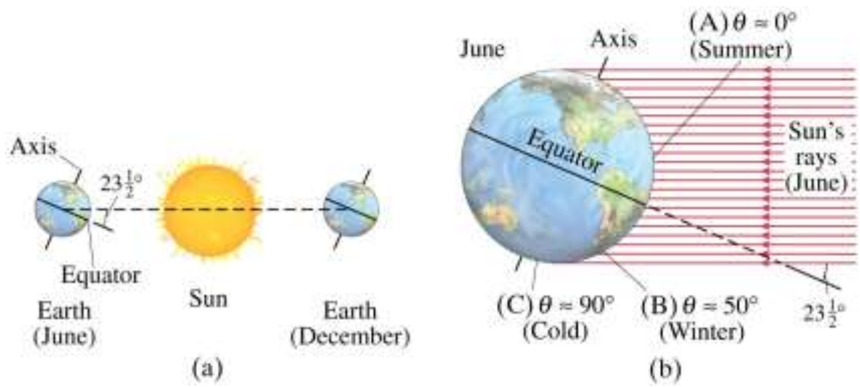
If you are in the sunlight, the Sun's radiation will warm you. In general, you will not be perfectly perpendicular to the Sun's rays, and will absorb energy at the rate:

$$\frac{Q}{t} = (1000\ \mathrm{W/m^2})\epsilon A\cos\theta, \quad (14\text{-}8)$$



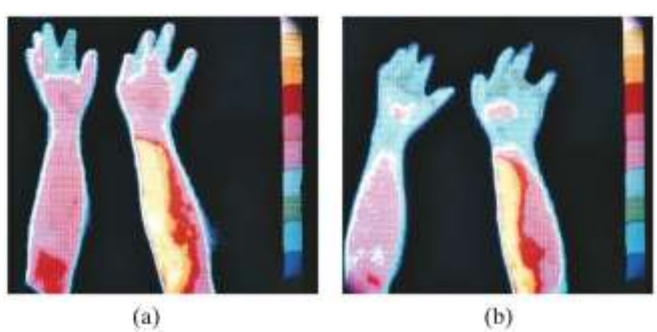
14-8 Heat Transfer: Radiation

This cos θ effect is also responsible for the seasons.



14-8 Heat Transfer: Radiation

Thermography—the detailed measurement of radiation from the body—can be used in medical imaging. Warmer areas may be a sign of tumors or infection; cooler areas on the skin may be a sign of poor circulation.



Summary of Chapter 14

- Internal energy U refers to the total energy of all molecules in an object. For an ideal monatomic gas,

$$U = \tfrac{3}{2}nRT, \qquad \left[\begin{array}{l}\text{internal energy of}\\ \text{ideal monatomic gas}\end{array}\right] \qquad (14\text{-}1)$$

- Heat is the transfer of energy from one object to another due to a temperature difference. Heat can be measured in joules or in calories.
- Specific heat of a substance is the energy required to change the temperature of a fixed amount of matter by 1° C.

Summary of Chapter 14

- In an isolated system, heat gained by one part of the system must be lost by another.
- Calorimetry measures heat exchange quantitatively.
- Phase changes require energy even though the temperature does not change.
- Heat of fusion: amount of energy required to melt 1 kg of material.
- Heat of vaporization: amount of energy required to change 1 kg of material from liquid to vapor.

Summary of Chapter 14

- Heat transfer takes place by conduction, convection, and radiation.
- In conduction, energy is transferred through the collisions of molecules in the substance.
- In convection, bulk quantities of the substance flow to areas of different temperature.
- Radiation is the transfer of energy by electromagnetic waves.