Chapter 14: Heat

Two objects at different temperature are put in contact: heat flows from the hotter to colder one. Flow of heat to equalize the temperature.

When the temperature becomes equal = **thermal equilibrium** is reached.

Heat and temperature: different concepts

We used to think about heat as a material (caloric) that would flow from one object to another. Even though we now know *heat refers to transfer of energy*, a common unit still used is named after caloric:
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.

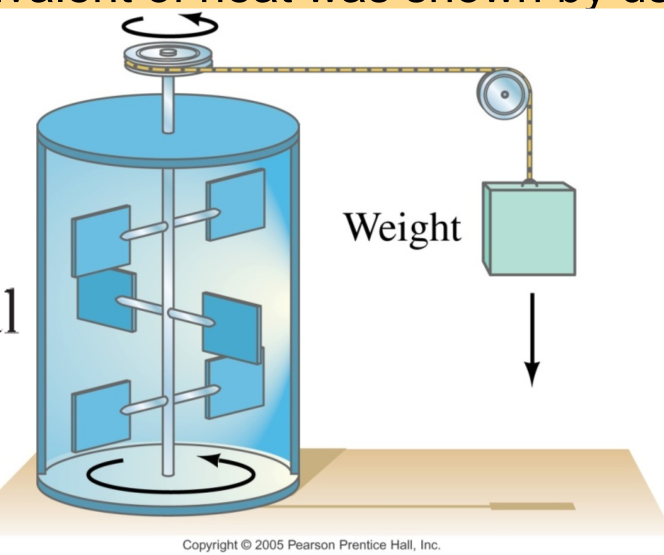
Heat as Energy Transfer

Joule and others performed experiments that established our present view that heat, as work, represents a transfer of energy.

He found that a given amount of work done was always equivalent to a particular amount of heat input

The mechanical equivalent of heat was shown by using the falling weight to heat the water:

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



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

Unit (SI): J

Ex. 14-1 Consider a 60 kg- person. How much total height should she climb to compensate for 500 Calories she ate? 3600m

Temperature, Heat and Internal Energy

The **temperature** of a gas is a measure of the *average* kinetic energy of its molecules.

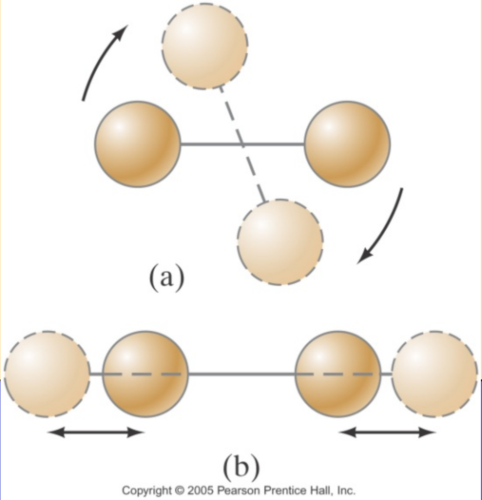
Internal Energy is the **total** energy of **all** the molecules in the object.

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

Internal energy of an ideal (atomic) gas:
(sum of the translational kinetic energies of all atoms)

$$U = N\left(\tfrac{1}{2} m\overline{v^2}\right)$$

$$\overline{KE} = m\overline{v^2}/2 = 3kT/2 \Rightarrow U = \tfrac{3}{2} nRT$$



<u>IDEAL GAS:</u> U depends only on n and T

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

<u>REAL GAS</u>: U depends on T, but also P and V

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

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

Specific Heat

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

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

The **specific heat**, *c*, is characteristic of the material.
UNIT (SI): J/(kg .°C)

Ex. 14-2 (a) How much heat input is needed to raise the temperature of an empty 20-kg vat made of iron from 10°C to 90°C? (b) What if the vat is filled with 20 kg of water?

(a) 720 kJ (b) (720 +6700) kJ

Ex. 14-3 An empty and very hot frying pan (200°C) is put in the bottom of a sink with a few inches of cool water. Will the water boil? Why?

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

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 across its boundaries.

For an isolated system, (internal *conservation of energy*)

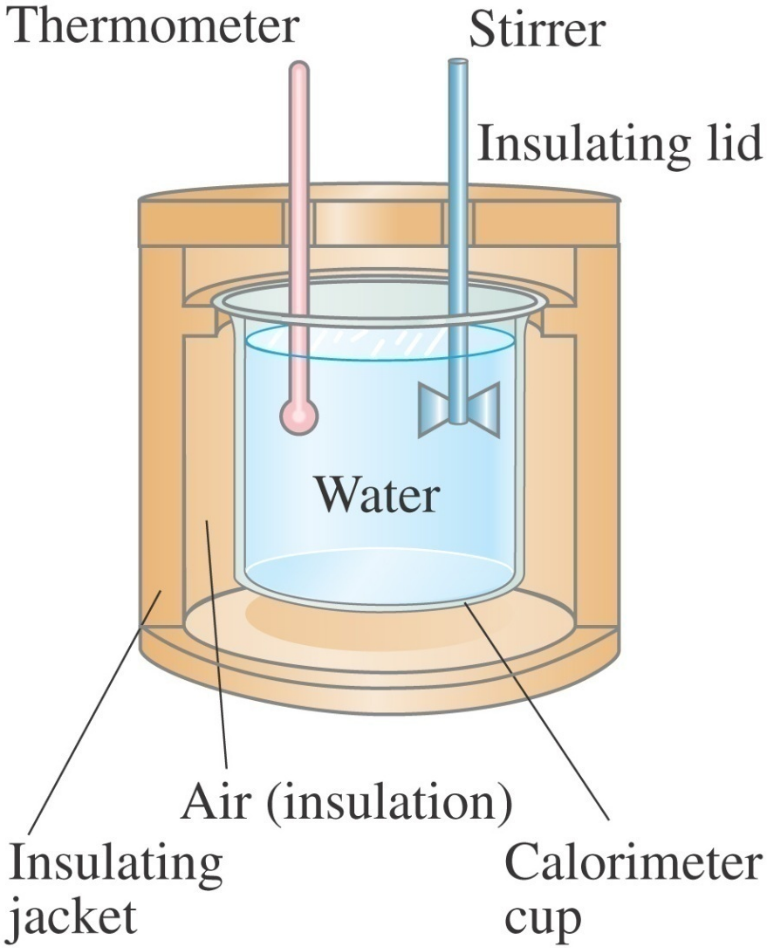
Energy out of one part = energy into another part

heat lost by one part = heat gained by another part

Ex. 14-4 If 200 cm^3 of tea at 95 ^oC is poured into a 150-g glass cup initially at 25^oC, what will be the common final temperature of the tea and cup when equilibrium is reached, assuming no heat flows to the surroundings?

86 ^oC

Calorimetry



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.

Ex. 14-5 An engineer wishes to determine the specific heat of a new metal alloy. A 0.150-kg sample of the alloy is heated to 540°C. It is then quickly placed in 400 g of water at 10.0°C, which is contained in a 200-g aluminum calorimeter cup. The final temperature of the system is 30.5°C. Calculate the specific heat of the alloy

500J/kg. °C

Calorimetry

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,

Ex. 14-6 Determine the energy content of 100 g of cookies from the following measurement. A 10-g sample of a cookie is allowed to dry before putting it in a bomb calorimeter. The aluminum bomb has a mass of 0.615 kg and is placed in 2.00 kg of water contained in an aluminum calorimeter cup of mass 0.524 kg. The initial temperature of the system is 15.0^{o}·C, and its temperature after ignition is 36.0^{o}·C.

47 kcal in 10 g
470 kcal in 100 g

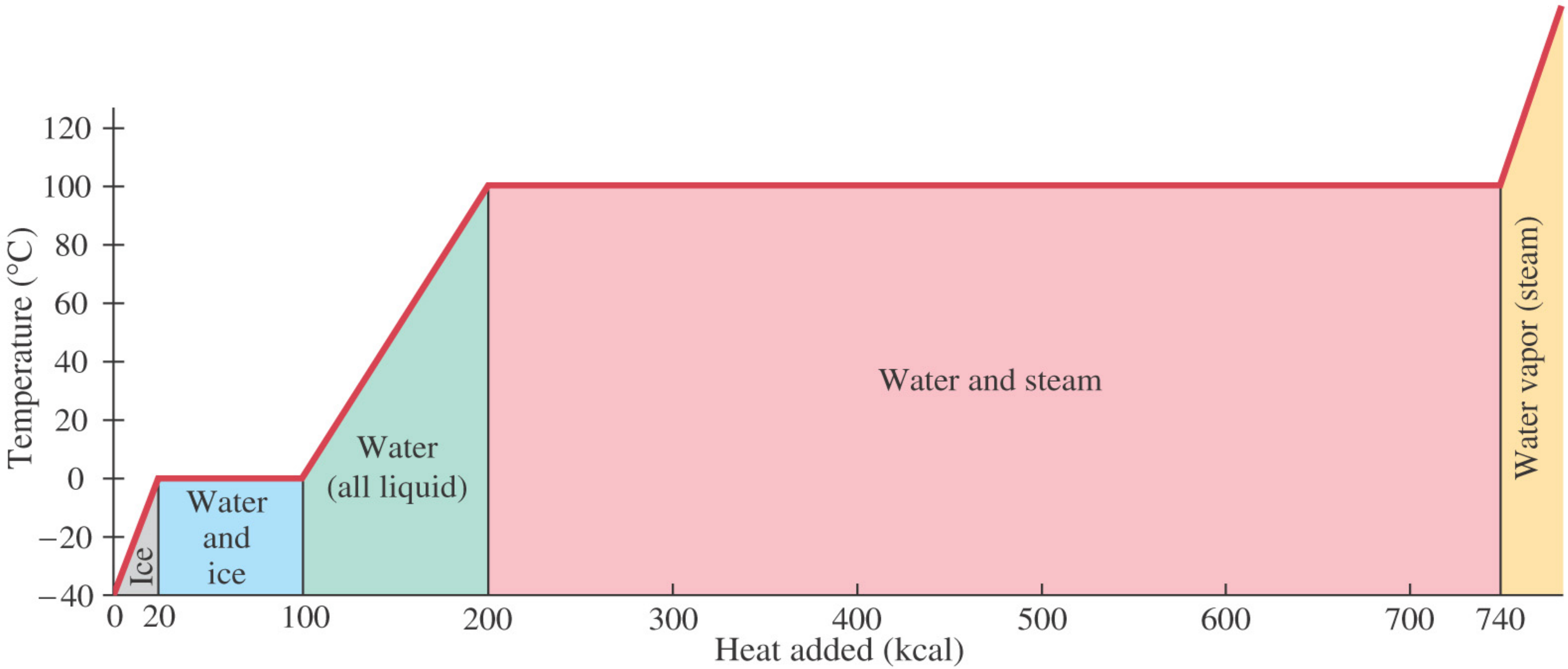
Latent Heat

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

Latent heats:

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



Latent Heat - exercises

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

$$Q = mL$$

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.

Ex. 14-7 How much energy does a freezer have to remove from 1.5 kg of water at 20^oC to make ice at -12^oC

660kJ

Ex. 14-8 At a reception, a 0.50-kg chunk of ice at -10^oC is placed in 3.0 kg of “iced” tea at 20^oC. At what temperature and in what phase will the final mixture be? The tea can be considered as water. Ignore any heat flow to the surroundings, including the container

5.0^oC

Latent Heat - table

TABLE 14–3 Latent Heats (at 1 atm)

Substance	Melting Point (°C)	Heat of Fusion kcal/kg[†]	Heat of Fusion kJ/kg	Boiling Point (°C)	Heat of Vaporization kcal/kg[†]	Heat of Vaporization kJ/kg
Oxygen	−218.8	3.3	14	−183	51	210
Nitrogen	−210.0	6.1	26	−195.8	48	200
Ethyl alcohol	−114	25	104	78	204	850
Ammonia	−77.8	8.0	33	−33.4	33	137
Water	0	79.7	333	100	539	2260
Lead	327	5.9	25	1750	208	870
Silver	961	21	88	2193	558	2300
Iron	1808	69.1	289	3023	1520	6340
Tungsten	3410	44	184	5900	1150	4800

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

Problem Solving

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. 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.

Ex. 14-9 The specific heat of a liquid mercury is 140 J/kg^{o}C. When 1.0 kg of a solid mercury at its melting point of -39^{o}C is placed in a 0.50-kg aluminum calorimeter filled with 1.2 kg of water at 20.0^{o}C, the final temperature of the combination is found to be 16.5^{o}C. What is the heat fusion of mercury in J/kg?

11 kJ/kg

Heat Transfer: Conduction

Heat conduction can be visualized as occurring through molecular collisions. Experimentally it is found that the heat flow per unit time is given by:

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{l}$$

Constant k is called the **thermal conductivity**.

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

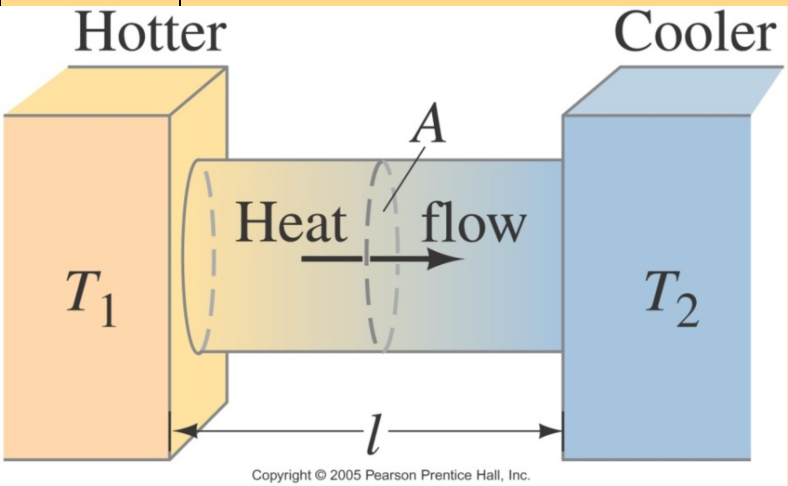


TABLE 14–4
Thermal Conductivities

Substance	Thermal Conductivity, k: $\frac{\text{kcal}}{(\text{s}\cdot\text{m}\cdot\text{C}^\circ)}$	$\frac{\text{J}}{(\text{s}\cdot\text{m}\cdot\text{C}^\circ)}$
Silver	10×10^{-2}	420
Copper	9.2×10^{-2}	380
Aluminum	5.0×10^{-2}	200
Steel	1.1×10^{-2}	40
Ice	5×10^{-4}	2
Glass	2.0×10^{-4}	0.84
Brick	2.0×10^{-4}	0.84
Concrete	2.0×10^{-4}	0.84
Water	1.4×10^{-4}	0.56
Human tissue	0.5×10^{-4}	0.2
Wood	0.3×10^{-4}	0.1
Fiberglass	0.12×10^{-4}	0.048
Cork	0.1×10^{-4}	0.042
Wool	0.1×10^{-4}	0.040
Goose down	0.06×10^{-4}	0.025
Polyurethane	0.06×10^{-4}	0.024
Air	0.055×10^{-4}	0.023

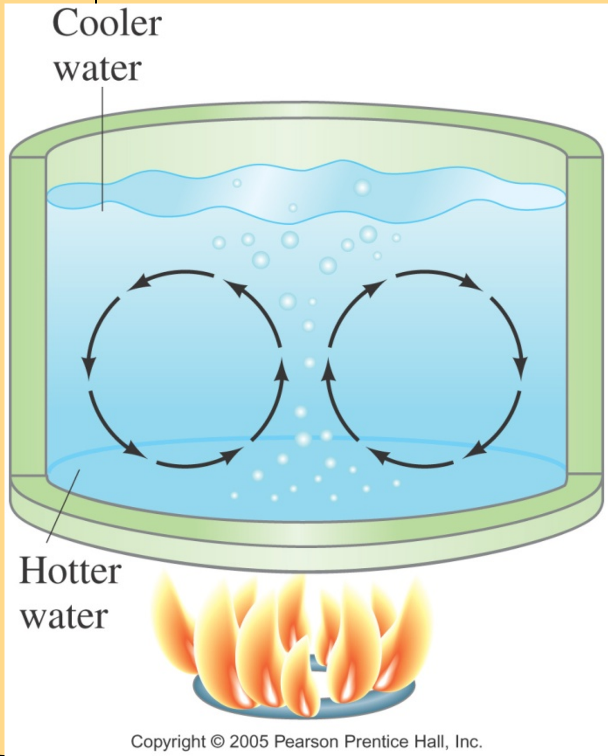
Ex. 14-10 A major source of heat loss from a house is through the windows. Calculate the rate of heat flow through a glass window 2.0mx1.5m in area and 3.2 mm thick, if the temperatures at the inner and outer surfaces are 15.0°C and 14.0°C, respectively

790 J/s

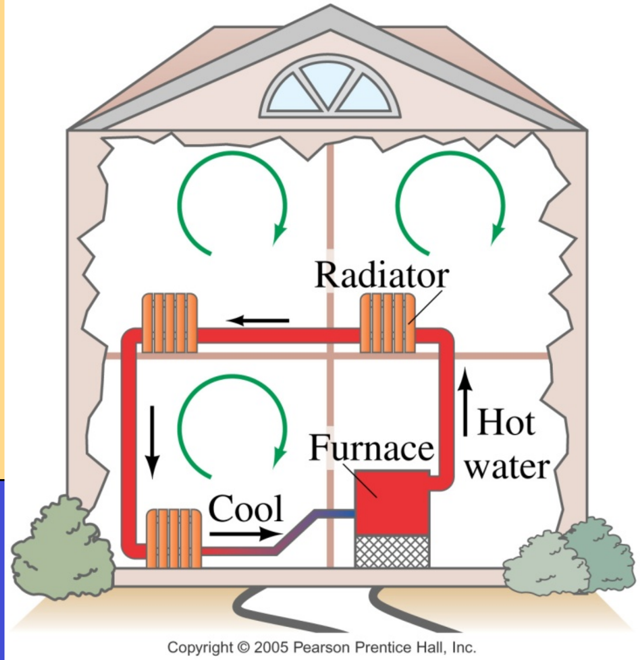
Heat Transfer: Convection

Gases and liquids are generally not good conductors, but they can transfer heat rapidly by **convection**. Convection occurs when heat flows by the mass movement of molecules from one place to another.
CONTRAST: Conduction: molecules move over small distances and collide.
Convection: movement of large numbers of molecules over large distances.

Convection may be natural or forced (ex: furnace+fan)



Hot air rises: Air expands as heated and hence its density decreases. Its density is less than the surrounding cool air and so it rises.
Examples: ocean currents, wind, blood, pot of water, hot water radiator



Heat Transfer: Radiation

Convection and conduction require the presence of matter as a medium to carry heat from the hotter to the colder region. But another type of heat transfer - **radiation** - occurs without any medium.

The most familiar example of radiation is the Sun, which radiates at a temperature of almost 6000 K (it consists of many wavelengths).

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

Temperature in KELVIN!!!

$$\frac{\Delta Q}{\Delta t} = e\sigma A T^4$$

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.

A good absorber is also a good emitter.

Net flow rate of radiation:

$$\frac{\Delta Q}{\Delta t} = e\sigma A(T_1^4 - T_2^4)$$

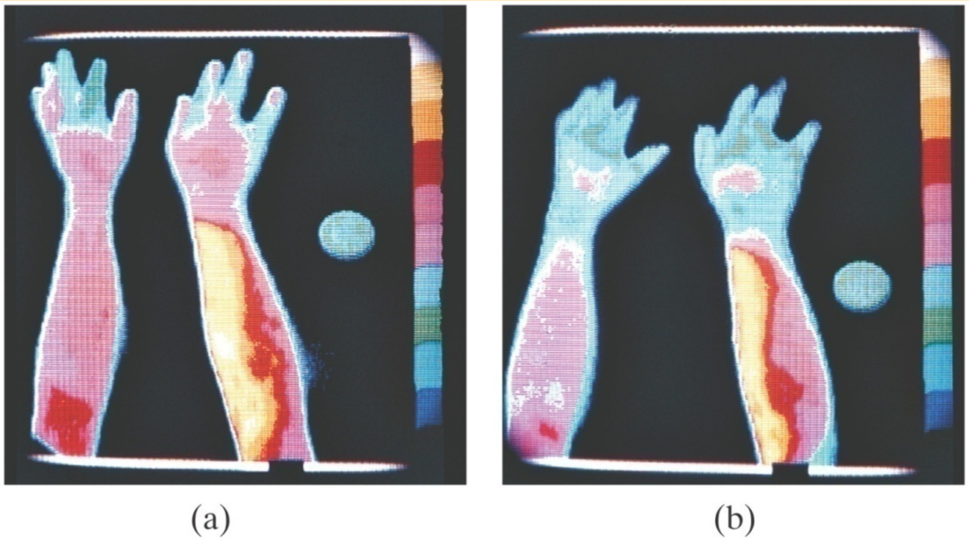
Exercise and Thermography

Ex. 14-11 An athlete is sitting unclothed in a locker room whose dark walls are at a temperature of 15°C. Estimate the rate of heat loss by radiation, assuming a skin temperature of 34 °C and e=0.70. Take the surface area of the body not in contact with the chair to be 1.5 m^2

120 W

If you are sitting in a room that is too cold, your body radiates more heat than it can produce. You will start shivering to increase the metabolic rate.

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.



Exercise

Ex. 14-12 A ceramic teapot (e=0.70) and a shiny one (e=0.10) each hold 0.75 L of tea at 95°C. (a) Consider that the area of the teapot is 0.05 m^2 and estimate the rate of heat loss due to radiation from each. (b) Estimate the temperature drop after 30 min for each. Consider only radiation, and assume the surrounding are at 20°C

a) 20 W for the ceramic pot and 3 W for the shiny one

b) 12°C for the ceramic pot and 2°C for the shiny one