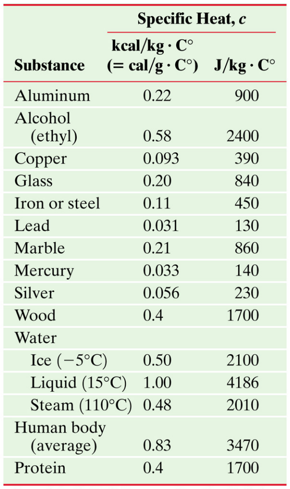
Thermodynamics

6 – Specific Heat



Substance	Specific Heat, c kcal/kg · C° (= cal/g · 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 (c): refers to the amount of heat required to raise the temperature of 1 kg of a substance by 1 K (or 1 °C).

But what if you had ***more*** (or ***less***!) than 1 kg…

$$\boxed{Q = mc\Delta T}$$

Where: Q = Required Heat (J)
m = mass (kg)
c = specific heat of substance
ΔT = change in temp (°C)

Example:

Copper vs. Water

$c_{water} = 4186$ J / kg °C
$c_{copper} = 389$ J / kg °C

Starting with 1 gram of copper at 0 °C and 1 gram of water at 0 °C. If you then raised the temperature of each by 1°C which substance required a larger heat input? Why?

WATER!

$c_{water} > c_{copper}$

Example:

Mr. Sandor, who currently weighs in at 85 kg's, just caught the flu (***oh no…*** ☹). His body temperature increased from 37.0 °C to 39.0 °C. How much energy was required to raise the body's temperature?

$$Q = mc_{body}\Delta T$$

$$Q = (85)(3470)(2) = \boxed{590\text{ KJ}}$$

Hot Object added to a Cold Liquid

ΔQ = 0 (if it's a closed system)
Q heat in = Q heat out
Q lost = Q gained
In this case….

$$-Q_{hot} = Q_{cold}$$

$$-m_{hot}c_{hot}(T_{final} - T_{hot}) = m_{cold}c_{cold}(T_{final} - T_{cold})$$

Example:

500. grams of 20.0° C water is added to 700. g of 85° C water. What is the final temperature of the mixture?

$$-Q_{hot} = Q_{cold}$$

$$-m_{hot}\cancel{c_{hot}}(T_f - T_{hot}) = m_{cold}\cancel{c_{cold}}(T_f - T_{cold})$$

$$-0.5T_f + 10 = 0.7T_f - 59.5$$

$$69.5 = 1.2T_f$$

$$\boxed{T_f = 57.9°C}$$

Example:

We wish to determine the specific heat of a new alloy. A 0.150 kg sample of the alloy is heated too 540 °C. It is then quickly placed in 400. g of water at 15.0°, which is contained in a 200. g aluminum calorimeter cup. (*Assume that the insulating jacket insulates well, so the temperature does not change significantly*). *The final temperature of the mixture is 30.5 °C.* ***Calculate the specific heat of the alloy***.

$$-Q_{hot} = Q_{cold}$$

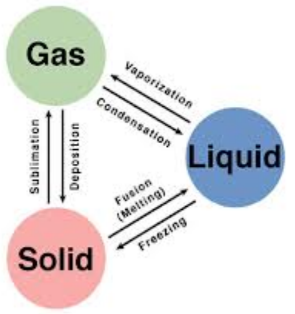
$$-m_s c_s \Delta T = m_w c_w \Delta T + m_{cal}c_{cal}\Delta T$$

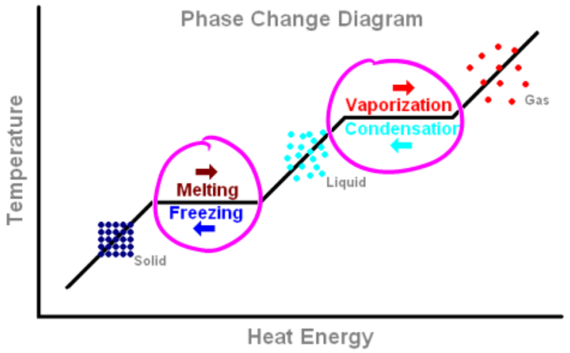
$$-(0.150)c_s(30.5 - 540) = (0.40)4186(305 - 15) + (0.20)(900)(30.5 - 15)$$

$$c_s = \frac{(0.40)(4186)(30.5 - 15) + (0.20)(900)(30.5 - 15)}{-(0.150)(30.5 - 540)}$$

$$\boxed{c_s = 376\,\frac{J}{kg\,°C}}$$

Phase Change: When a substance changes state (example: $S \rightarrow L$; OR $L \rightarrow G$)





Notice how at certain points energy is added to the system but temperature does not increase….

Energy is required to separate particles from each other and therefore does not increase their kinetic energy (***temperature***)

$$\boxed{Q = mL_f \text{ OR } Q = mL_v}$$

Where: Q = Latent Heat (J)
m = mass (kg)
L_f = heat of fusion (melt/freeze)
L_v = heat of vaporization (boil/condense)

Substance	Melting Point (°C)	Heat of Fusion		Boiling Point (°C)	Heat of Vaporization	
		kcal/kg[†]	kJ/kg		kcal/kg[†]	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.

Example:

How much energy does a refrigerator have to remove from 1.5 kg of water at 20.0 °C to make ice at -12 °C.

$$Q_{total} = Q_{20\rightarrow 0} + Q_0 + Q_{0\rightarrow -12}$$

$$Q_{total} = m_w c_w \Delta T + mL_f + m_{ice} c_{ice} \Delta T$$

$$Q_{total} = (1.5)(4186)(0-20) + (1.5)(-333{,}000) + (1.5)(2100)(-12-0)$$

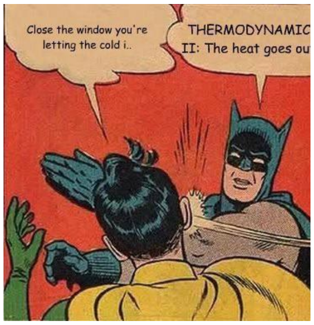
$$Q_{total} = \cancel{-5.67\times10^{5}\,J}\ \ -6.63\times10^{5}\,J$$

Fridge must remove ~~567 kJ~~ 6.63×10^{2} kJ of energy

Conduction: Thermal Energy can be transferred from one material to another in a process known as ***heat***.

The only requirement for heat transfers is a difference in temperature.

Ok… A couple *MORE* important details

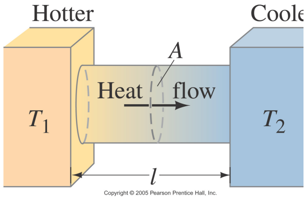


- Q increases if the temperatue difference (ΔT) is increased
- Q increases if the Area of the object increases
- Q increases if the Length of the object is decreased
- Some materials transfer heat better than others… (k = thermal conductivity)

Heat Transfer

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

Where: Q = Heat J
Δt = time (s)
k = thermal conductivity
A = cross-sectional area (m²)
ΔT = change in temp (°C)
L = Length (m)



Example:
If air has such a low thermal conductivity (0.22), why do we need to wear clothes? If we omit the obvious answer… *decency*!

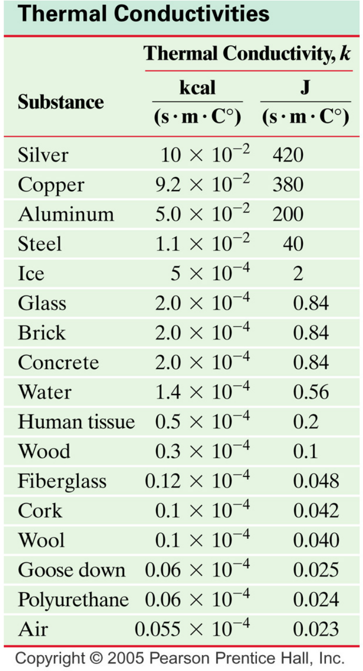
Clothes trap air against our body.
Energy is transferred from our bodies to the trapped air.
If air is constant replaced we must constantly warm NEW air.

Example:
Suppose you sit down on a 8.0 °C concrete bench. You are only wearing a thin layer of clothing that provides negligible insulation and therefore your core temperature (37 °C) is only protected by 1.2 cm thick layer of fat (*on your bum!*) that touches the bench. Let's estimate the area of contact between your self and the bench to be 0.10 m². What is the heat loss due to conduction?

$$\frac{Q}{\Delta t} = \frac{kA\Delta T}{L} = \frac{(0.2)(0.10)(37-8.0)}{0.012}$$

$$\boxed{\frac{Q}{\Delta t} = 48.3\ J/s}$$

Bum → Bench!

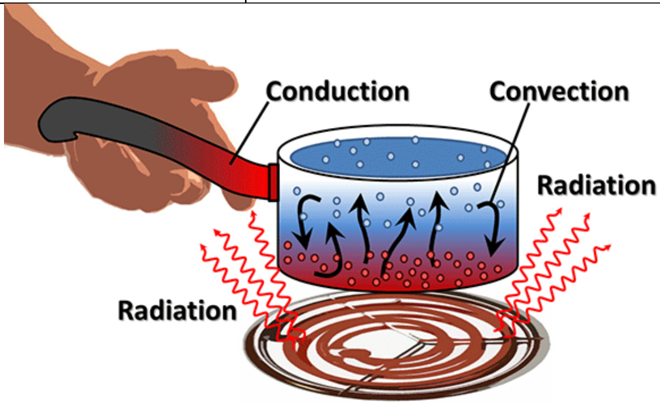


Thermal Conductivities

Substance	Thermal Conductivity, *k* kcal / (s · m · C°)	 J / (s · m · C°)
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

FYI… There are two other means to transfer thermal energy…

Convection: Transfer of thermal energy by motion in fluids.	Radiation: transfer of thermal energy through electromagnetic waves.



Putting it all together!

Example:

When the weather gets cold, the air we breathe in is heated when it comes in contact with warm lung tissue. The energy to heat the air comes from our body. The inhaled air warms to nearly the temperature of the interior of a human body 37 °C. When humans exhale, some heat is retained by the body, but most if lost. We will assume the exhaled air is about 30 °C.

263.15 K

A typical person takes 12 breaths each minute, with each breath taking in about 0.50 L of outside air. If the air outside is -10 °C and is heated to 30 °C… → 303.15 K

a. What is the volume of exhaled air with each breath?

n, P, R are Constant!

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1} = (0.50 L)\left(\frac{303.15 k}{263.15 k}\right) = \boxed{0.58 L}$$

b. Find the heat and the heat power required to warm one breath of air (MM = 29 g/mol and specific heat = 1.01 kJ/(kg °C)). (*the gases are exchanged as you breath – Oxygen to Carbon Dioxide – but to a good approximation the number of atoms, stay the same*).
Note: consider only the energy required to warm the air, not the energy lost to evaporation from the tissues of the lungs.

$$PV = nRT$$

$$n = \frac{RT}{PV} = \frac{(1.01\times10^5 Pa)(5.0\times10^{-4} m^3)}{(8.31 J/mol\cdot K)(263.15 K)} = 0.023 mol$$

12 breaths per min ⟶ 1 breath every 5 s

$$\frac{Q}{\Delta t} = \frac{m_{air} c_{air} \Delta T}{\Delta t} = \frac{n_{air} MM_{air} c_{air} \Delta T}{\Delta t} = \frac{(0.023 mol)(29 g/mol)(1.01)(30 - (-10))}{5 s}$$

$$\frac{Q}{\Delta t} = 5.4 J/s = \boxed{5.4 W}$$