

PHYS 104: THERMAL PROPERTIES OF MATTER

Lecturer: Evgeni Burovski
(B75 Physics building)

Lectures/Seminars/Workshops:

Tuesday	9am–10am Biology LT
	11am–12pm Bowland Main LT
Thursday	11am–12pm Biology LT
Friday	9am–10am George Fox LT 1
	11am–12am Biology LT

Official office hour: Tuesday 10am–11am
Open door policy.

PHYS 104: THERMAL PROPERTIES OF MATTER

ASSESSMENT

Assessment: 60% Exam (Summer term), 40% coursework (2 worksheets and the in-class test).

WORKSHEETS

will be posted to MLE on Mondays weeks 16 and 19, you have a week to complete a worksheet. For example, worksheet 1 will be posted to MLE on Monday week 16, and should be handed in by 11am on Monday week 17 in the relevant INBOX in the physics foyer. Work handed in later than this but before the 11am the following Thursday will incur a penalty of 1 grade reduction. Work handed in later than that will not count towards your continuous assessment.

PHYS 104: THERMAL PROPERTIES OF MATTER

	Tue 9am	Tue 11am	Thurs 11am	Fri 9am	Fri 11am
Week 16	Lect.	Lect.	Wkshp	Lect.	Lect.
Week 17	Lect.	Lect.	Lect.	Lect.	Sem.
Week 18	Lect.	Lect.			Test
Week 19	Lect.	Lect.	Wkshp	Lect.	Sem
Week 20	Lect.	Lect.		Lect.	Sem.

Extra “warm-up” and “challenge” questions will be on Moodle.

PHYS 104: THERMAL PROPERTIES OF MATTER

Textbook: Young and Freedman, “University physics” 13th ed.
In particular, Chapters 17–20.

Lecture notes: will be on MLE.

These have an observable positive impact on the exam score:

- Perusing the textbook regularly, and
- Taking notes during the lectures

NB: Lec notes are *not* self-contained. And they have a lot of white space for you to jot down your notes.

FROM THE HANDBOOK:

(available on the department's web site)

On completion of this module, students should be able to:

- appreciate the role of thermodynamics in describing macroscopic physical situations.
- display a familiarity with fundamental thermal phenomena.
- understand the basic concepts through which the phenomena are described, in particular those of temperature, work, heat, internal energy and entropy.
- apply their knowledge to modelling real phenomena and situations.

THERMAL PROPERTIES OF MATTER

Matter is made of atoms and molecules. Quite a lot of them, in fact: a typical number is $\sim 10^{24}$ particles.

Two complementary approaches:

- Statistical physics: look at the properties of individual particles and their interactions with each other. A *microscopic* approach.
- Thermodynamics: Express observable phenomena in terms of *phenomenological* parameters. A *macroscopic* approach.

TEMPERATURE I

WHAT IS TEMPERATURE?

Microscopically: it's a measure of an average kinetic energy of molecules of a material. (more on this later on)

Empirically: it's a way of quantifying the notions of “cold”, “warm”, “tepid”, “scalding” etc

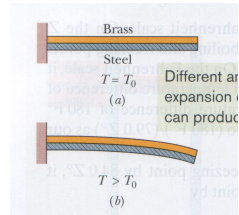
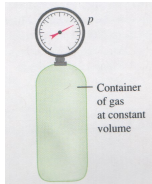
Quantify \iff measure: need a *thermometer*.

The main idea: use a measurable property of some system which varies with degree of “hotness”

TEMPERATURE II: THERMOMETERS

FOR EXAMPLE:

- Matter tends to expand when heated



- Electric resistance changes with temperature
- Matter emits electromagnetic waves, the amount of radiation depends on temperature

Q: are different ways of measuring temperature consistent with each other?

THERMAL EQUILIBRIUM



Take a cup of hot coffee, let it sit on a table for a while.

If we wait long enough, the “hotness” of the coffee no longer changes (and matches that of the air).

→ it reaches *thermal equilibrium* with the air in the room.

NB: Here, the air serves as a *heat bath*. (it's big enough so that we can neglect the changes in its properties on approach to the thermal equilibrium)

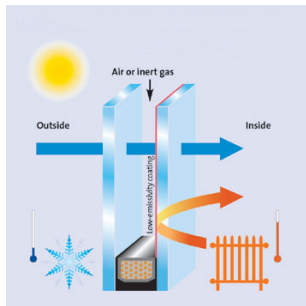
A CAVEAT: IF WE WAIT LONG ENOUGH...

Q: How long is long enough?

A: Well, it depends. Strictly speaking, it's infinitely long. Practically, it depends on the quality of thermal contact.

Good thermal conductors: e.g. metals

Thermal insulators: wood, fiberglass, vacuum.



Sometimes we just don't want to reach thermal equilibrium! e.g.: double glazed windows

THERMAL EQUILIBRIUM: ZEROth LAW OF THERMODYNAMICS

Take three systems, A , B and C , in thermal contact with each other.



A : air in the room

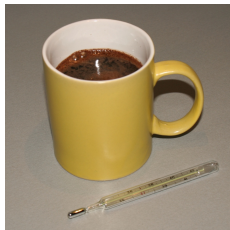
B : coffee

C : the spoon

Q: If A is in thermal equilibrium with B and B is in thermal equilibrium with C , are A and C in thermal equilibrium with each other?

A: Yes. This is known as a *zeroth law of thermodynamics*

ZEROth LAW OF THERMODYNAMICS



The reading of a thermometer is the same in both cases.

In other words: *If two systems are in thermal equilibrium with each other, they have the same temperature.*

Q: : What does a thermometer measure?

A: : Its own temperature.

TEMPERATURE SCALES: °C AND °F

To make a thermometer useful, we need to calibrate it.

CELSIUS SCALE

zero: Freezing of pure water

100: Boiling of pure water

Divide the distance between the two into a hundred equidistant intervals called *degrees*, °C.

FAHRENHEIT SCALE

Anecdotally: zero for the coldest night of 1708/09 in Danzig,
hundred for Mr Fahrenheit's body temperature

32: Freezing of water

212: Boiling of water

Which makes it 180 *degrees Fahrenheit*, °F, in between.

TEMPERATURE SCALES: °C AND °F

CONVERSION FROM °C TO °F:

$$T_F = \frac{9}{5} T_C + 32^\circ\text{C}$$

$$T_C = \frac{5}{9} (T_F - 32^\circ\text{F})$$

Why $\frac{5}{9}$? It's just $\frac{100}{180} = \frac{5}{9}$

If you only need a rough idea: “subtract thirty, divide by two”.

WE NEED A MORE ACCURATE TEMPERATURE SCALE

What we need is a reproducible, reliable and material-independent way of calibrating thermometers.

The definition of a temperature scale based on boiling and freezing of water is not accurate enough.

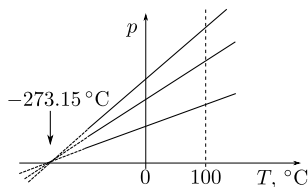
- Need pure water: chemical contamination shifts the freezing point.
- Distilled water can be superheated (heated above 100°C) or supercooled (cooled below 0°C).
- Both boiling and freezing points depend on the ambient pressure.

We need to look for an experimentally accessible law of Nature.

GAS THERMOMETERS

Take a fixed-volume vessel of a gas. At constant volume, the pressure, p , increases with temperature, T .

Measure the pressure: A mixture of ice and water gives 0°C , a mixture of water and steam gives 100°C . Draw a straight line between the two. Repeat for different quantities of different gases.



Experimentally: extrapolated to zero pressure, all curves cross at $T \approx -273.15^{\circ}\text{C}$.

Use this fact for establishing the *Kelvin temperature scale* a.k.a. *absolute temperature scale*.

ABSOLUTE TEMPERATURE

$$T_K = T_C + 273.15$$

So that $0\text{ K} = -273.15^\circ\text{C}$,
and temperature *difference* of $1^\circ\text{C} = 1\text{ K}$

To formally define the absolute temperature scale we need two points. One is zero Kelvin. The second one is chosen to be the *triple point* of water: a unique combination of p and T such that ice, water and vapor coexist (more on this later).

$$\text{water : } T_{\text{triple}} = 0.01^\circ\text{C} = 273.16\text{ K}$$

$$p_{\text{triple}} = 610\text{ Pa} \approx 6.1 \times 10^{-3}\text{ atm}$$

In thermal physics we *always* use the absolute temperature, T .

ABSOLUTE ZERO: $T = 0\text{ K}$

Q: Is it correct to say that at $T = 0\text{ K}$ the motion of atoms, molecules etc stops?

A: No. There is the so-called *zero-point* motion, as required by quantum mechanics.

Q: Can a physical system reach absolute zero?

A: No. That's called the *third law of thermodynamics*. We can only approach absolute zero. The lowest temperatures attained in labs are about 50 nK.

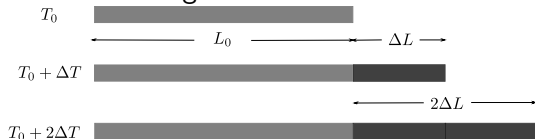
Q: Are there negative temperatures?

A: No. In some cases we can define a temperature-like quantity, an “effective T ”. The latter can be negative. If interested, google for “population inversion” or “lasers”.

THERMAL EXPANSION

Most materials expand(contract) with heating(cooling).

Consider a long thin rod:



Experimentally (as long as $\Delta T \ll T_0$ and $\Delta L \ll L_0$):

- the change in length, $\Delta L \propto \Delta T$.
- the change in length, $\Delta L \propto L_0$.

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

Here α is *linear expansion coefficient*.

THERMAL EXPANSION II

- The linear expansion coefficient is measured in K^{-1} .
- The numerical values of α are different for different materials (cf Y&F, Chap. 17.4).
- They also depend on the initial temperature. For solids this dependence is typically weak and can often be neglected.

For example, for brass at around room temperature $\alpha = 2 \times 10^{-5} \text{K}^{-1}$. Raising the temperature by 10 degrees, we have

$$\begin{aligned}\frac{\Delta L}{L} &= 2 \times 10^{-5} \text{K}^{-1} \times 10 \text{K} = 2 \times 10^{-4} \\ &= 0.0002 = 0.02\%\end{aligned}$$

THERMAL EXPANSION III: RAIL TRACKS

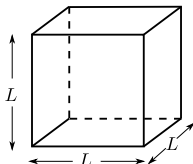


VOLUME EXPANSION

Volume expansion:

$$\frac{\Delta V}{V_0} = \beta \Delta T, \quad \Delta T \ll T_0,$$

where β is the *volume expansion coefficient*.



If T increases by dT , then

$$dV = \frac{dV}{dL} dL = 3L^2 dL$$

using $dL = \alpha L dT$, we obtain

$$dV = 3L^2 \alpha L dT = 3\alpha (L^3) dT$$

Or, equivalently,

$$\beta = 3\alpha$$

(assuming that the material is isotropic)

VOLUME EXPANSION OF LIQUIDS

For liquids β is typically larger than that for solids.

At around room temperature



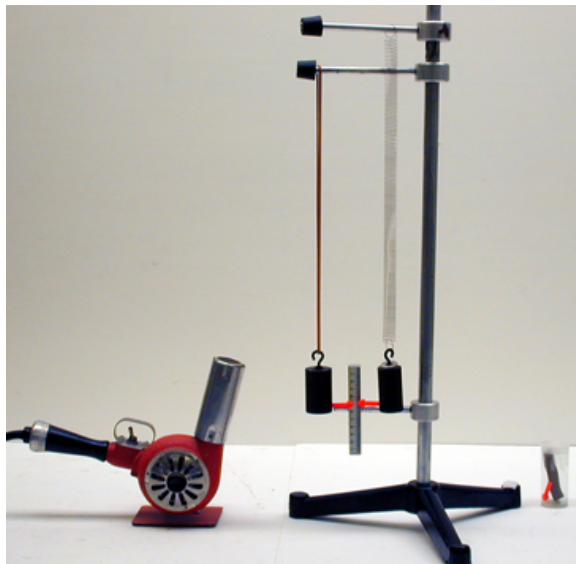
glass: $\beta \approx 1\text{--}3 \times 10^{-5} \text{K}^{-1}$

mercury: $\beta \approx 18 \times 10^{-5} \text{K}^{-1}$

For most materials $\beta > 0$. Loosely speaking, “matter tends to expand when heated”.

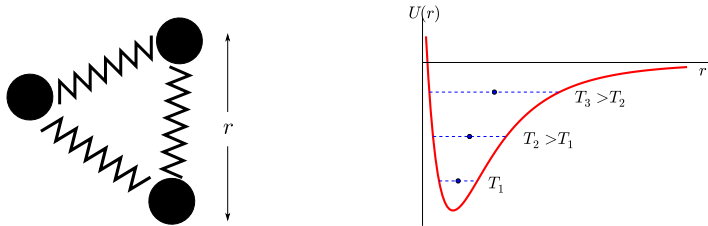
The so-called *anomalous liquids* have $\beta < 0$ in a certain temperature range. For example, water (more on this later.)

DID I SAY THINGS EXPAND WHEN HEATED?



<http://faraday.physics.uiowa.edu/heat/4A30.80.htm>

MICROSCOPIC ORIGIN OF THERMAL EXPANSION



The potential energy of interaction between molecules in a solid or a liquid can be approximated by a Lennard-Jones potential:

$$U(r) = \varepsilon \left[\left(\frac{r_0}{r} \right)^{12} - 2 \left(\frac{r_0}{r} \right)^6 \right],$$

which is *not symmetric*.

As the temperature increases, the average energy of the oscillations in the LJ potential increases, and the equilibrium distance between molecules shifts towards larger values.

HEAT, THE QUANTITY OF



As the coffee cools down to the room temperature, there is *energy transfer* the coffee to the air.

The amount of energy transferred because of the difference in T -s is called *heat*.

Loosely speaking,

heat \approx energy *in transit*.

TEMPERATURE VS HEAT

Imagine spilling 1/2 of coffee. The remaining liquid has the same temperature. But to raise its temperature by 1 degree, we need to supply *half* as much heat.

IN WHICH UNITS IS THE HEAT MEASURED?

... essentially, these are units of energy.

In the SI system, it's joule.

Historically, the unit was a *calorie*: 1 calorie is the amount of heat required to raise the temperature of 1 gram of water from 14.5 °C to 15.5 °C.

$$1 \text{ cal} = 4.186 \text{ J}$$

Sometimes a “food Calorie” is used (abbreviated as Cal):

$$1 \text{ Cal} = 1 \text{ kcal} = 10^3 \text{ cal.}$$

RECAP: TEMPERATURE, ENERGY AND HEAT

Temperature and *energy* relate to an object:

“The coffee in this cup is at 373 K”

“The energy value of this candy bar is about 42 kcal” (better: about 176 kJ).

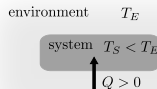
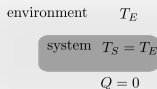
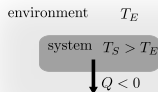
Heat relates to an interaction between objects:

“During the *heat exchange* between A and B, 42 J was transferred.”

But it's never “An object contains 42 J of heat”.

The latter statement went out of fashion by the end of the 18th century. If curious, google for “phlogiston”.

RECAP: TEMPERATURE, ENERGY AND HEAT



Consider a system at temperature T_S , interacting with the environment at T_E .

$T_S > T_E$: the system loses energy as heat, $Q < 0$

$T_S < T_E$: the system gains energy as heat, $Q > 0$

$T_S = T_E$: no heat exchange, $Q = 0$

SPECIFIC HEAT A.K.A. HEAT CAPACITY

Q: How much heat does it take to increase the temperature of a body by 1 K?

$$Q = cm\Delta T,$$

Where:

- ΔT is the temperature change, $\Delta T = T_{\text{final}} - T_{\text{initial}}$
- Q is the quantity of heat (notice that $Q > 0$ if $\Delta T > 0$)
- m is the mass of a sample
- c is the *specific heat* or *heat capacity*

The units of specific heat are $\text{J kg}^{-1}\text{K}^{-1}$.

Specific heat is different for different materials. It also depends on the temperature, but the dependence is weak and can often be ignored.

MOLAR HEAT CAPACITY

Heat capacity, c , measures how much heat it takes to raise the temperature of 1 kg of material by 1 K. It is sometimes more convenient to talk about the quantity of matter in terms of the number of *moles*. Recall that a *mole* of pure substance is defined to contain $N_A \approx 6.022 \times 10^{23}$ molecules.

The *molar mass*, μ , is a mass of a mole of material (for pure substances: atomic mass in the periodic table gives μ in g/mol).

For ν moles of a material of the molar mass μ , the mass is $m = \mu\nu$

$$Q = c\mu\nu\Delta T \equiv C\nu\Delta T$$

where $C = c\mu$ is the so-called *molar heat capacity* (or *molar specific heat*.)

The units of the molar specific heat are $\text{J mol}^{-1}\text{K}^{-1}$.

CALORIMETRY CALCULATIONS: AN EXAMPLE

A brick of brass of m_b is heated to a temperature $T_b = 70^\circ\text{C}$ and dropped into a bucket of water. The mass of water is m_w and its initial temperature is $T_w = 15^\circ\text{C}$. Assuming that the bucket itself has negligible heat capacity and that the system is kept isolated from the surroundings, find the final temperature of the system at thermal equilibrium.

Solution: Let the (unknown) final temperature is T .

Heat transfer for brass: $Q_1 = c_b m_b (T - T_b) < 0$

Heat transfer for water: $Q_2 = c_w m_w (T - T_w) > 0$

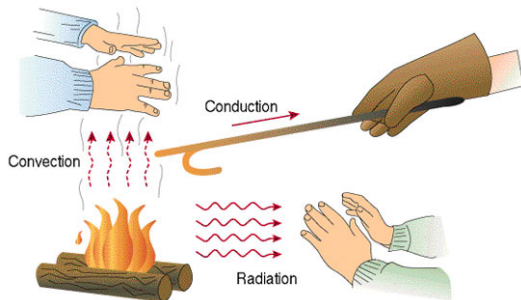
Conservation of energy: $0 = Q_1 + Q_2$

Therefore: $c_b m_b (T - T_b) + c_w m_w (T - T_w) = 0$

And finally,

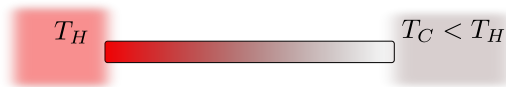
$$T = \frac{c_b m_b T_b + c_w m_w T_w}{c_b m_b + c_w m_w}$$

MECHANISMS OF HEAT TRANSFER



- Conduction
- Convection
- Radiation

HEAT CONDUCTION



Imagine a thin rod connecting a hot reservoir at temperature T_H and a cold reservoir at temperature $T_C < T_H$. The heat flows from the hot reservoir to the cold one.

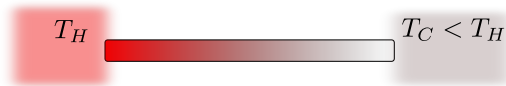
If δQ of heat flows along the rod in time δt , then $H = \delta Q / \delta t$ is the *heat current* (units: Watt).

$$H = kA \frac{T_H - T_C}{L}$$

where A is the area of the cross-section of the rod, L is the rod's length, and

k is the *thermal conductivity* ($\text{Wm}^{-1} \text{K}^{-1}$).

HEAT CONDUCTION II



To find the distribution of temperature along the rod (or if interested in the time evolution of the temperature), use the generalization of the last formula:

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

Here x is the coordinate along the rod, and dT/dx is the temperature *gradient*.

Notice the minus sign: *heat flows from hotter regions to colder regions*.

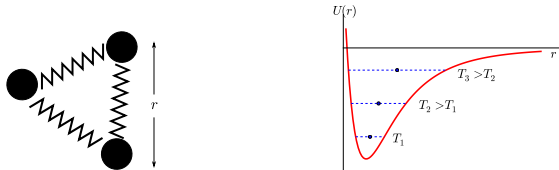
THERMAL CONDUCTIVITY OF VARIOUS SUBSTANCES

Representative values at room temperature and atmospheric pressure.

	$k, \text{W m}^{-1} \text{K}^{-1}$				
Copper	400	Concrete	0.1–2	Dry air	0.024
Lead	35	Wood	0.1–0.4	Argon	0.016
Mercury	8	Marble	2–3	Helium	0.14

- Metals are better heat conductors than insulators
- Gases are poor heat conductors (think double glazed windows and wool sweaters!)

MICROSCOPIC ORIGIN OF THERMAL CONDUCTIVITY



Temperature is a measure of the average kinetic energy atoms/molecules have. Atoms collide with each other and exchange energy.

Notice that there is a flow of heat even though there is not net flow of atoms!

Metals transmit heat mainly via *conduction* electrons which are nearly free to roam around the sample. (these are the same electrons that conduct electricity)

CONVECTION

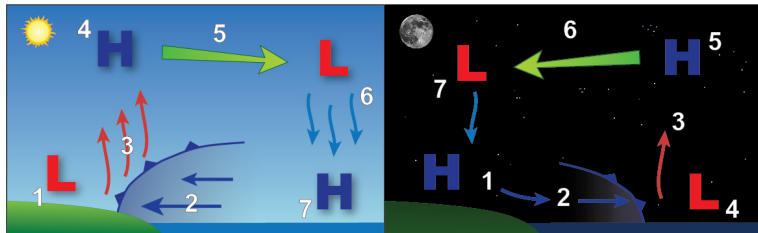
Convection is the heat transfer via mass motion of a fluid from one region of space to another.

- *Forced convection*: blood flow in blood vessels where the heart works as pump
- *Natural convection*: due to thermal expansion, hotter fluid is less dense than colder fluid.

Qualitatively: Convection is due to an interplay between thermal expansion and the Archimedes force.

CONVECTION: ON- AND OFF-SHORE BREEZE

The heat capacity of water is larger than that of the earth.



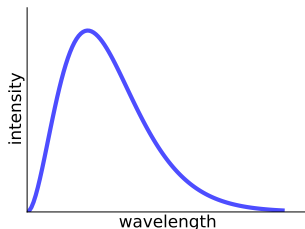
During the day, the sea is slightly colder than the land, hence the air above the sea is slightly denser. The hot air above the land raises upwards \implies there is an onshore wind.

Image: NOAA National Weather Service

RADIATION

Radiation is the transfer of heat via electromagnetic waves.

This is the so-called *black body radiation*.
In fact, every body emits EM radiation.
The spectrum (and overall intensity)
depends on the temperature.



For the Sun ($T \sim 6000 \text{ K}$) the maximum of the intensity is in the visible range.

For bodies at about room temperature ($T \sim 300 \text{ K}$) the maximum is in the infrared.

BLACK BODY RADIATION

The heat current H ($=dQ/dt$) emitted by a body of an area A kept at temperature T the *Stefan-Boltzmann law*:

$$H = e\sigma AT^4$$

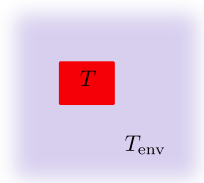
where

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

is the *Stefan-Boltzmann constant*, and e is the *emissivity* of a body (which depends on the material and on the quality of the surface.)

Maximum emissivity equals 1 (*ideal black body*), minimum emissivity is zero. It's typically larger for dark surfaces than for lighter ones.

BLACK BODY RADIATION OF THE ENVIRONMENT



A body at temperature T emits EM radiation. And so does its environment, which is at temperature T_{env} .

If the body is in thermal equilibrium with the environment, $T = T_{\text{env}}$, the rates of emission and *absorption* must be equal to each other.

The former is given by $H = e\sigma AT^4$, hence the latter must be $e\sigma AT_{\text{env}}^4$.

Therefore, the *net emission rate* (emission *and* absorption) is

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

PHASES OF MATTER AND PHASE TRANSITIONS

Phases of matter (= states of matter): liquid, solid, gaseous.

In fact, there are more: plasma, superfluids etc etc. Which we are not going to consider just now.

Each phase is stable in a certain range of temperatures and pressures. *E.g.* at atmospheric pressure, liquid water is stable between 0°C and 100°C .

Once the boundary of the region of stability is reached \rightarrow *phase transition*.



Take a block of ice at 0°C and start adding heat slowly. The temperature does not change: ice melts instead. Now water and ice are in *phase coexistence*.

PHASE TRANSITIONS

Notice that ice melts when heat is added (not the other way around).

To completely melt a mass m of material, we need to add

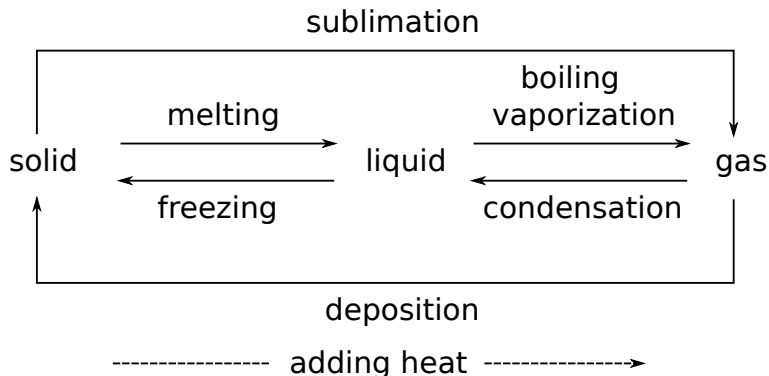
$$Q = mL_f$$

of heat. Here L_f is the *latent heat of fusion* (measured in J/kg.)

Q: How much heat do we have to remove in order to completely freeze m kg of liquid?

A: The same amount, $Q = mL_f$, — the process is reversible (if done slowly so that at all times the mixture is in thermal equilibrium)

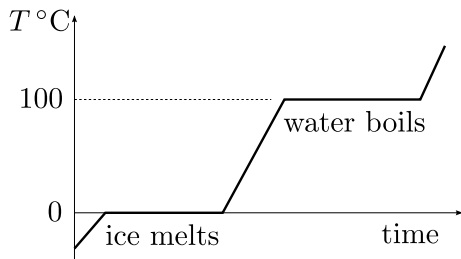
PHASE TRANSITIONS II



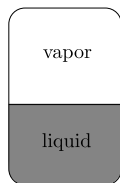
melting \longrightarrow the latent heat of fusion, L_f .
boiling \longrightarrow the latent heat of vaporization, L_v
sublimation \longrightarrow the latent heat of sublimation, L_s

PHASE TRANSITIONS III

Imagine taking a block of ice at $T < 0^\circ\text{C}$ at atmospheric pressure and slowly adding heat at a constant rate. How does the temperature depend on time?



DYNAMICS OF PHASE TRANSITIONS: BOILING



Imagine a hermetically sealed vessel with a mixture of liquid and its vapor (ignore the presence of air for now).

The vapor in equilibrium with the liquid is called *saturated vapor*. Its pressure (the so-called *vapor pressure*), p_s , depends on temperature.

Little bubbles of vapor are spontaneously created in the liquid all the time (it's called *fluctuations*).

At low temperature, these bubbles shrink and disappear.

Increasing the temperature, the vapor pressure increases. Once it reaches the ambient pressure, the bubbles do not shrink any more: they float upwards and escape — the liquid boils.

DYNAMICS OF PHASE TRANSITIONS: BOILING II

Liquid boils as soon as the vapor pressure equals the ambient pressure:

$$p_s(T) = p_{\text{ambient}}$$

p_s increases with temperature, therefore— the lower the ambient pressure, the lower the boiling temperature.

		height, m	$T_{\text{boiling}}, ^\circ\text{C}$
For example (water):	Ben Nevis	1344	95
	Mont Blanc	4260	85

DYNAMICS OF PHASE TRANSITIONS III

Bubble formation is aided by *nucleation centers*: dust, container wall roughness, shaking etc.

Clean liquids can be *overheated* (or *overcooled*). These are called *metastable* states: a slightest disturbance leads to an immediate boiling.

A WILSON CHAMBER:



A vessel with an overcooled vapor: an elementary particle creates little droplets of liquid along its track. (Here: a gamma ray created an electron and a positron pair).

There is much more— we'll come back to discussing phase transitions later on in this course.