Thermodynamics
Heat Capacity
Phase Changes

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Last time

- finish applying the ideal gas equation
- thermal energy
- introduced heat capacity

Warm Up Question

Quick Quiz 20.1[1] Imagine you have 1 kg each of iron, glass, and water, and all three samples are at 10°C.

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

Heat capacities: glass – 837 J kg^{-1} K^{-1}
iron – 448 J kg^{-1} K^{-1}

(A) iron, glass, water ⟵

(B) water, iron, glass

(C) water, glass, iron

(D) glass, iron, water

[1]Serway & Jewett, pg 579.

Warm Up Question

Quick Quiz 20.1[1] Imagine you have 1 kg each of iron, glass, and water, and all three samples are at 10°C.

(b) Rank the samples from greatest to least amount of energy transferred by heat if each sample increases in temperature by 20°C.

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Overview

- heat capacity
- phase changes
- latent heat

Heat Capacity

Heat Capacity, C

of a sample of substance is the quantity of heat required to change the temperature **of that sample** by 1 degree C (or K).

$$Q = C\,\Delta T$$

where ΔT is the change in temperature and Q is the heat.

Specific Heat Capacity, c

of a substance is the quantity of heat required to change the temperature of a unit mass of that substance by 1 degree C (or K).

$$Q = cm\,\Delta T$$

m is the mass of the object.

Heat and Temperature Change

Energy that causes a change in temperature does not have to enter our system as heat.

It can be a different form of energy transfer.

Examples:

- in a microwave, energy T_{ER} enters the food as **electromagnetic waves**
- **work** can cause a temperature change in two surfaces rubbed together, or as a bicycle pump pressurizes air in the bike tires, the air's temperature rises

These energy transfers to our system will increase the internal energy of the system, E_{int}.

Units of Internal Energy and Heat: Calories

The units of both internal energy and heat are Joules, J.

However, heat was not always understood to be an amount of energy, so other units have been defined for it, and are still sometimes used.

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The "calories" listed on food labels are sometimes called "Calories" (capital C) because they are in fact kilocalories.

1 Calorie = 1 kilocalorie = the heat required to raise the temperature of 1 kilogram of water by 1 degree Celsius.

1 calorie = 4.18 Joules.

Calorimetry

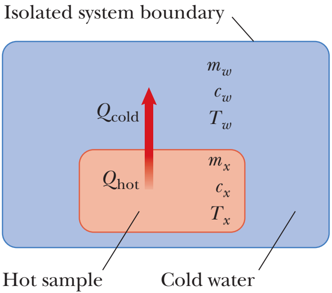
Calorimetry

a technique for determining the specific heat capacity of a sample by heating it to a known temperature, then transferring it to a known quantity of water and observing the temperature change in the water.

Steps:

1. sample of known mass m_x is heated to temperature T_x
2. sample is moved to an isolated container of water, containing mass m_w of water at temperature $T_w < T_x$
3. the sample and the water are allowed to reach thermal equilibrium
4. the final temperature of the water, T_f, is measured

Calorimetry



Since the heat transferred to the cold water is equal to the heat transferred from the hot sample:

$$
\begin{aligned}
Q_c &= -Q_h \\
m_w c_w (T_f - T_w) &= -m_x c_x (T_f - T_x) \\
c_x &= \frac{m_w c_w (T_f - T_w)}{m_x (T_x - T_f)}
\end{aligned}
$$

[1]Figure from Serway & Jewett, page 595.

Phase Changes

The processes by with matter changes from one state to another.

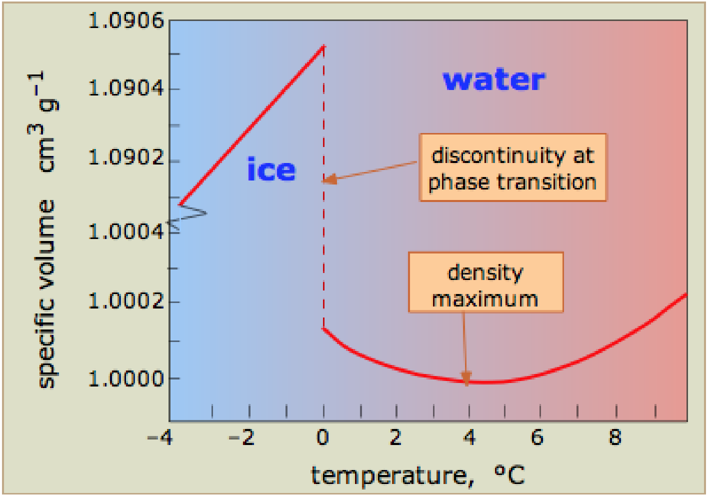
The different states of matter: solid, liquid, gas, plasma, are also called *phases* of matter.

Phase Changes

Phase changes tend to be dramatic.

If sudden, obvious changes in the properties and behaviors of a substance did not occur as we vary the temperature, we would probably have no need to refer to different states of matter!

Phase Changes



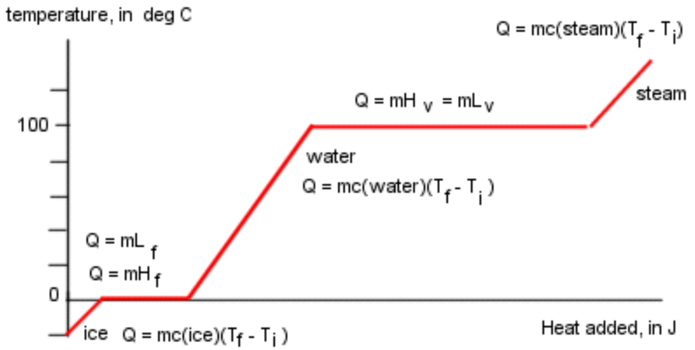
Notice the discontinuity!

Phase Changes

We know that as we heat a solid it will eventually melt to form a liquid and if we keep heating the liquid will boil off as a gas.

But how does the temperature change during these processes?

Phase Changes



During a phase change, temperature doesn't change, even when heat is added!

Phase Changes

Why does this happen?

Where is the energy going?

It isn't increasing the translational speed of the atoms, that would relate to an increase in temperature.

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It isn't increasing the translational speed of the atoms, that would relate to an increase in temperature.

$\rightarrow$ It goes into breaking bonds.

Latent Heat

latent heat of fusion, L_f

The amount of energy (heat) per unit mass required to change a solid to a liquid.

$$Q = mL_f$$

where m is the mass of solid that is transformed into a liquid.

latent heat of vaporization, L_v

The amount of energy (heat) per unit mass required to change a liquid to a gas.

$$Q = mL_v$$

where m is the mass of liquid that is transformed into a gas.

[1] "Latent" from *latere*, "to lie hidden".

Latent Heat

Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium[a]	−272.2	5.23×10^3	−268.93	2.09×10^4
Oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
Nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
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

[a]Helium does not solidify at atmospheric pressure. The melting point given here corresponds to a pressure of 2.5 MPa.

[1]Table from Serway & Jewett, page 598; values at atmospheric pressure.

Latent Heat

Latent heat is the energy required for the to change to the higher energy phase per unit mass of the substance.

This includes the energy for both of two components[2]:

1. the energy required to overcome intermolecular forces / break the bonds
2. the work required to push aside gas at ambient pressure to allow for any increased volume of the new phase

Because of this latent heat of fusion/vaporization is also called *enthalpy of fusion/vaporization*.
Enthalpy, $H = E_{\text{int}} + PV$

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The latent heat depends on the temperature and pressure of the phase change.

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Latent Heat

Latent heat is the energy required for the to change to the higher energy phase per unit mass of the substance, including both ΔE_{int} and the work that must be done against the environment at constant pressure to change the volume of the system.

Practice

The specific heat capacity of ice is about 0.5 cal/g$^\circ$C. Supposing that it remains at that value all the way to absolute zero, calculate the number of calories it would take to change a 1 g ice cube at absolute zero (-273°C) to 1 g of boiling water. How does this number of calories required to change the same gram of 100°C boiling water to 100°C steam?

Reminder: 1 cal is the heat required to raise the temperature of 1 g of water by 1°C.

[1] Hewitt, Problem 2, page 314.

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warming ice:

$$Q_1 = mc_{\text{ice}}\,\Delta T = (1\text{ g})(0.5\text{ cal/g}^\circ\text{C})(273^\circ\text{C}) = 136.5\text{ cal}$$

melting:

$$Q_2 = mL_f = (1\text{ g})\left(\frac{3.33\times 10^5\text{ J/kg}}{4.186\text{ J/cal}}\right)\left(\frac{1\text{ kg}}{1000\text{ g}}\right) = 79.55\text{ cal}$$

warming water:

$$Q_3 = mc_{\text{water}}\,\Delta T = (1\text{ g})(1.0\text{ cal/g}^\circ\text{C})(100^\circ\text{C}) = 100\text{ cal}$$

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Total $Q_1 + Q_2 + Q_3 = 320$ cal.

[1]Hewitt, Problem 2, page 314.

Comment on Prob 23 on WebAssign HW

(Problem 23 is the 5th problem in that set.)

In an insulated vessel, 250 g of ice at 0°C is added to 600 g of water at 18.0°C (a) What is the final temperature of the system? (b) How much ice remains when the system reaches equilibrium?

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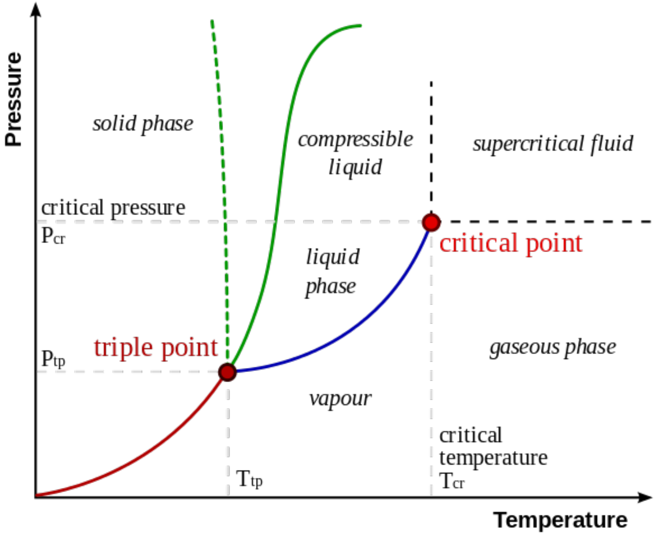
In an insulated vessel, 250 g of ice at 0°C is added to 600 g of water at 18.0°C (a) What is the final temperature of the system? (b) **How much ice remains** when the system reaches equilibrium?

Can you guess the answer to part (a) right now, without any calculation?

Phase Diagrams



[1]A typical phase diagram. The dashed green line shows the unusual behavior of water. Diagram by Matthieumarechal, Wikipedia.

Summary

- heat capacity
- phase changes
- latent heat

Homework (not collected, some recommended problems here are not in WebAssign)
Serway & Jewett:

- Look at examples 20.1–4.
- Ch 20. Probs: 19, 71