Thermodynamics
Phase Changes
Work Done on a Gas

Lana Sheridan

De Anza College

April 27, 2020

Last time

- applying the ideal gas equation

- thermal energy

- heat capacity

- phase changes

- latent heat

Overview

- more about phase changes

- work, heat, and the first law of thermodynamics

Practice

The specific heat capacity of ice is about 0.5 cal/g$^\circ$C. 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.

[1]Hewitt, Problem 2, page 314.

Practice

The specific heat capacity of ice is about 0.5 cal/g°C. 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.

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:

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

---

[1]Hewitt, Problem 2, page 314.

Practice

The specific heat capacity of ice is about 0.5 cal/g°C. 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.

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:

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

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

---

[1]Hewitt, Problem 2, page 314.

Practice

The specific heat capacity of ice is about 0.5 cal/g°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?

[1]Hewitt, Problem 2, page 314.

Practice

The specific heat capacity of ice is about 0.5 cal/g°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?

boiling:

$$Q_4 = mL_v = (1 \text{ g}) \left(\frac{2.26 \times 10^6 \text{ J/kg}}{4.186 \text{ J/cal}} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = 540 \text{ cal}$$

---

[1]Hewitt, Problem 2, page 314.

Practice

The specific heat capacity of ice is about 0.5 cal/g°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?

boiling:

$$Q_4 = mL_v = (1 \text{ g}) \left(\frac{2.26 \times 10^6 \text{ J/kg}}{4.186 \text{ J/cal}} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = 540 \text{ cal}$$

The energy required to transform the water to steam is much bigger than the energy required to heat the ice, convert it to water, and continue heating up to $100°$C.

---

[1]Hewitt, Problem 2, page 314.

Question

Suppose the same process of adding energy to the ice cube is performed as discussed in the last question, but instead we graph the internal energy of the system as a function of energy input, $Q + W$. What would this graph look like?

Phase Change paths
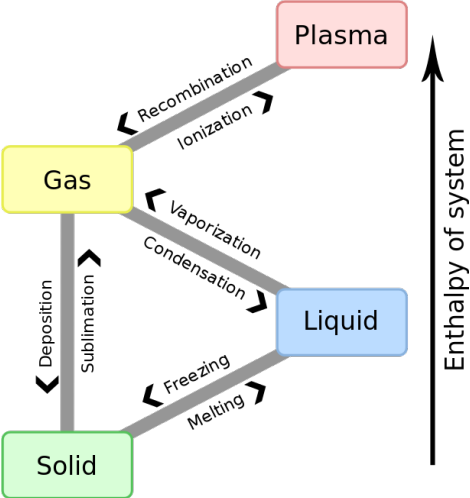
Evaporation

evaporation

the process by which a liquid changes to a gas at the liquid surface

Since changing from a liquid to a gas requires heat, when a liquid evaporates it takes heat from its surroundings.

This is why humans sweat in hot weather, pigs wallow puddles, and dogs pant. All are trying to use evaporation of water to reduce body temperature.

Evaporation

Ben Franklin noticed that a wet shirt kept him feeling cool on a hot day.

He decided to experiment to see if the temperature of objects could be lowered by this process.

In 1758 he and John Hadley took a mercury thermometer and repeatedly wet the bulb with ether while using bellows to keep air moving over it.

Despite it being a warm day, they recorded temperatures as low as 7° F (-14°C) at the bulb of the thermometer.

This is the basic idea behind refrigeration!

Latent Heat Practice

The heat of vaporization of ethyl alcohol is about 200 cal/g. If 2 kg of this fluid were allowed to vaporize in a refrigerator, show that 5 kg of ice (at $0°C$) would be formed from $0°C$ water.

---

[1]Hewitt, Problem 8, page 314.

Latent Heat Practice

The heat of vaporization of ethyl alcohol is about 200 cal/g. If 2 kg of this fluid were allowed to vaporize in a refrigerator, show that 5 kg of ice (at 0°C) would be formed from 0°C water.

Hint: in the last problem we melted 1 g of ice and found it required 80 cal.

[1]Hewitt, Problem 8, page 314.

Latent Heat Practice

The heat of vaporization of ethyl alcohol is about 200 cal/g. If 2 kg of this fluid were allowed to vaporize in a refrigerator, show that 5 kg of ice (at $0°$C) would be formed from $0°$C water.

Hint: in the last problem we melted 1 g of ice and found it required 80 cal.

energy needed for vaporization:

$$Q = mL_{v,ea} = (2 \text{ kg})(200 \text{ cal/g}) = 4.0 \times 10^5 \text{ cal}$$

assuming this same amount of energy was taken from the water:

$$m = \frac{Q}{L_f} = \frac{4.0 \times 10^5 \text{ cal}}{80 \text{ cal/g}} = 5000 \text{ g} = 5 \text{ kg} \checkmark$$

---

[1]Hewitt, Problem 8, page 314.

Regelation

This is a phenomenon seen in water because its density is lower in its solid state.
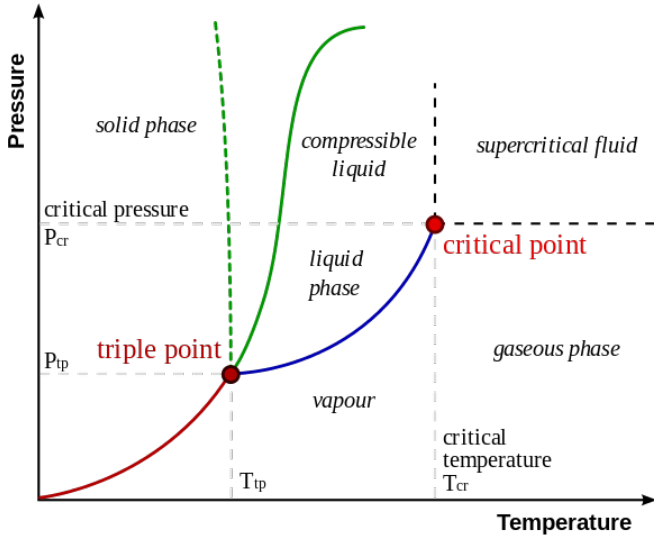
High pressure applied to solid water causes it to melt, even at low temperatures.

This makes ice skating work.

Phase Change paths
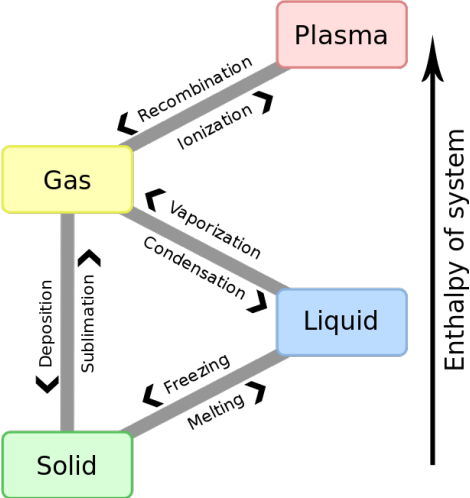
Phase Diagrams

---

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

Liquids below their melting point & above their boiling point

Phase changes require the bond structure of the substance to change.

Liquids below their melting point & above their boiling point

Phase changes require the bond structure of the substance to change.

As a liquid is cooled, it will generally start to freeze at one point before another. (Symmetry breaking.) For example, ice cubes in a freezer freeze from the top down.

As water freezes solid crystals form and spread throughout.

If the water is very pure and cooled without shaking, it can be cooled below $0°C$ without freezing.

This is called **supercooling** and can happen in some other liquids also under the right conditions.

Liquids below their melting point & above their boiling point

The same thing can happen when water (or other liquids) are heated just above their boiling points.

As the liquid starts to boil, bubbles of vapor form inside it. This happens most easily at defects in the edges of the container.

If the container is very smooth and not shaken, the only way the liquid can vaporize is from the middle. Surface tension opposes this.
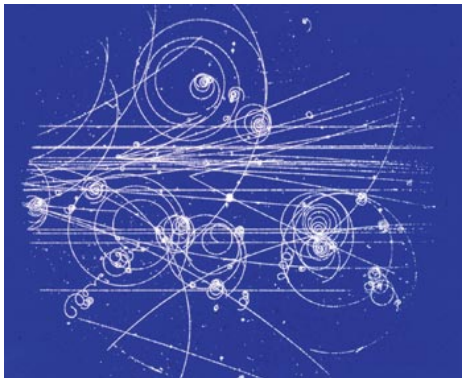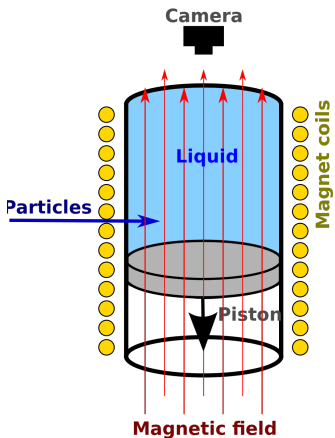
The liquid can be heated a bit above its boiling point.

This is called **superheating**.

It is used in **bubble chambers** to observe charged particles.

Liquids below their melting point & above their boiling point

Bubble chambers:

Heat and Work

We now take a closer look at the first law of thermodynamics.

To do this, we will take a deeper look at work and heat.

We also need to consider our system more carefully.

Thermodynamic Equilibrium States

We will study **thermodynamic systems**.

These systems are in **thermodynamic equilibrium internally**.

> **Thermodynamic equilibrium state**
>
> a state of a system in which every part of the system will be at the same temperature, T, and if the system is a gas, at the same pressure, P.

In classical thermodynamics, it is a postulate that any system left isolated will come to a thermal equilibrium state given enough time, and then remain in that state.

Thermodynamic Equilibrium States

We will study **thermodynamic systems**.

These systems are in **thermodynamic equilibrium internally**.

> **Thermodynamic equilibrium state**
>
> a state of a system in which every part of the system will be at the same temperature, T, and if the system is a gas, at the same pressure, P.

In classical thermodynamics, it is a postulate that any system left isolated will come to a thermal equilibrium state given enough time, and then remain in that state.

In particular, for our present discussion we will be considering an **ideal gas**. (Can use $PV = nRT$.)

Ideal Gases

Ideal gas assumptions (kinetic theory model).

We make the following assumptions in the ideal gas model:

- the volume of the gas particles is negligible compared to the total gas volume
- molecules are identical hard spheres (will relax this later)
- collisions between molecules are elastic
- there are no intermolecular forces (aside from hard-sphere collisions)
- there are no long-range forces from the environment (can be relaxed)

A real gas is behaves as an idea gas when it is at high temperature and low density (far from condensation).

Variables

The variables we will use can be broken into types:

state variables – describe system's state / properties – T, P, V, and E_{int}.

transfer variables – describing energy transferred into our out of the system – Q, W

Variables

The variables we will use can be broken into types:

state variables – describe system's state / properties – T, P, V, and E_{int}.

transfer variables – describing energy transferred into our out of the system – Q, W

intensive variables – variables that don't change value when the system is doubled in size – P, T, ρ, c

extensive variables – variables that double their value when the system is doubled in size – V, E_{int}, m, C

Work done <u>on</u> a gas

Imagine compressing or expanding a gas in a piston **quasi-statically** (meaning slowly enough so the gas remains in thermal equilibrium).
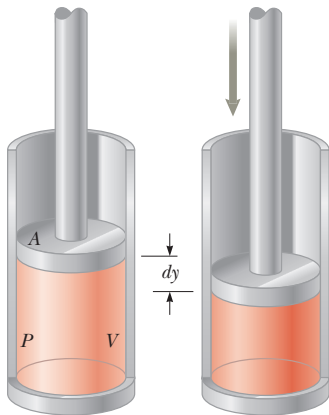
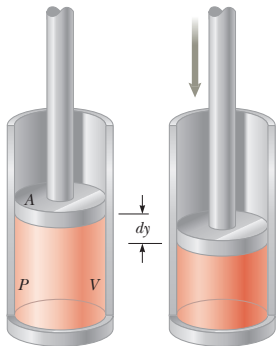How much work is done **on** the gas?

Work done <u>on</u> a gas

Definition of work:

$$W = \int \mathbf{F} \cdot d\mathbf{r}$$
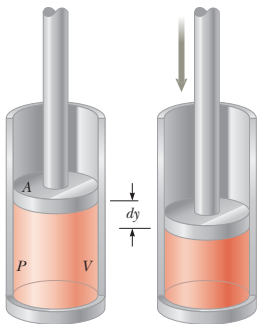
For this system:

$$
\begin{aligned}
W &= \int (-F\mathbf{j}) \cdot dy\,\mathbf{j} \\
&= -\int F\,dy \\
&= -\int P A\,dy \\
&= -\int P\,dV
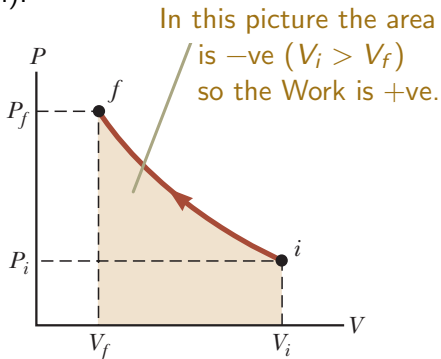\end{aligned}
$$

$$W = -\int_{V_i}^{V_f} P\,dV$$

Work done on a gas

Here, the volume decreases, so the work done on the gas is positive.

$$W = -\int_{V_i}^{V_f} P \, dV$$

The work done is the area under the *P-V* curve (with the appropriate sign).

In this picture the area is −ve ($V_i > V_f$) so the Work is +ve.

Summary

- more about phase changes
- work, heat, and the first law of thermodynamics

Next Test Tuesday, May 5 (? TBC), on Ch19, 20.

Homework Serway & Jewett:

- Read chapter 20 and look at the examples.