



# Conceptual Physics

## Phase Changes

## Thermodynamics

Lana Sheridan

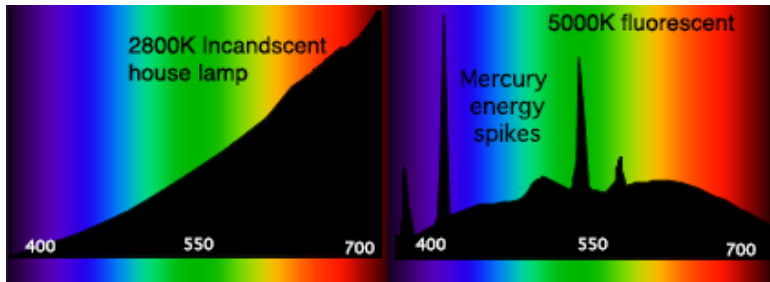
De Anza College

July 31, 2017

## Last time

- heat capacity
- thermal expansion
- heat transfer mechanisms
- Newton's law of cooling
- the greenhouse effect

# Fluorescent Lamps



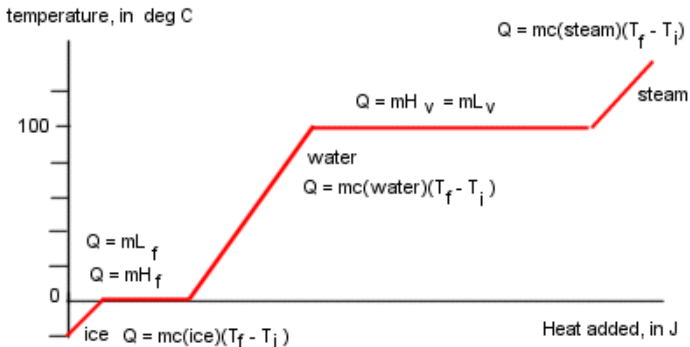
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<sup>1</sup>Figure, Wikipedia, uploaded by user Dkroll2.

# Overview

- phases of matter and phase change
- the laws of thermodynamics
- entropy
- heat engines

# Phase Changes



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

# Latent Heat

## latent heat of fusion, $L_f$

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

$$\text{heat, } Q = mL_f$$

## latent heat of vaporization, $L_v$

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

$$Q = mL_v$$

## Practice

Problem 2, page 314.

The specific heat capacity of ice is about  $0.5 \text{ cal}/(\text{g}^\circ\text{C})$ . 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 ( $-273^\circ\text{C}$ ) to 1 g of boiling water. How does this compare to the number of calories required to change the same gram of  $100^\circ\text{C}$  boiling water to  $100^\circ\text{C}$  steam?

Reminder: 1 cal is the heat required to raise the temperature of 1 g of water by  $1^\circ\text{C}$ .

From the textbook: the latent heat of fusion of water is  $80 \text{ cal/g}$ , the latent heat of vaporization is  $540 \text{ cal/g}$ .

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<sup>1</sup>Hewitt, Problem 2, page 314, and see page 309.

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warming ice:

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

melting:

$$Q_2 = mL_f = (1 \text{ g})(80 \text{ cal}/\text{g}) = 80 \text{ cal}$$

warming water:

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

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Total  $Q_1 + Q_2 + Q_3 = 320 \text{ cal}$ .

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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  $100^\circ\text{C}$ .

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<sup>1</sup>Hewitt, Problem 2, page 314.

## Practice

Problem 8, page 314.

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.

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energy needed for vaporization:

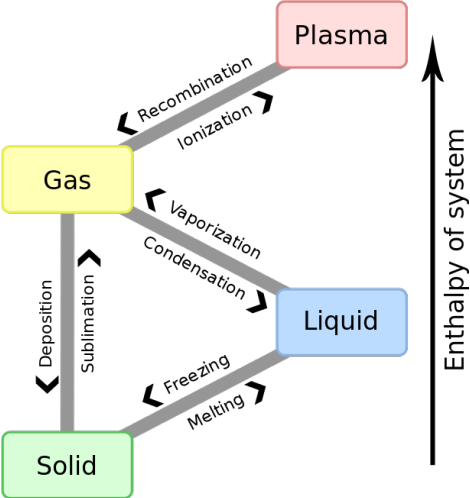
$$Q = mL_{v,ea} = (2000 \text{ g})(200 \text{ cal/g}) = 4.0 \times 10^5 \text{ cal}$$

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

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



# Phase Change paths



# Melting

## melting

the change of a solid to a liquid

As a substance reaches the temperature at which it melts, additional heat energy goes into the internal energy of the atoms / molecules making up the substance.

Bonded atoms / molecules are in a state that has low potential energy.

In a liquid the rigid bonds between molecules are broken. Breaking these bonds corresponds to increasing the potential (internal) energy of the molecules.

# Freezing

## freezing

the change of a liquid to a solid

Here heat is lost from the substance, causing molecules to lose potential (internal) energy.

The solid material bonds reform.

# 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.

# Condensation

## condensation

the process by which a gas changes to a back into a liquid

A glass of ice water in will collect water on the outside of it in hot, humid weather.

The drink is heated by the gas in the air, but as it takes heat from the air, water vapor condenses out onto the glass.

This is also the process that causes dew to collect overnight.

# Boiling

## boiling

the change of a liquid to a gas that occurs not only at the surface but throughout the liquid

In a rolling boil, bubbles can be seen to be forming within the middle of water in a pot.

These bubbles are steam - gaseous water. It forms randomly in pockets throughout the whole pot.

# Evaporation

## evaporation

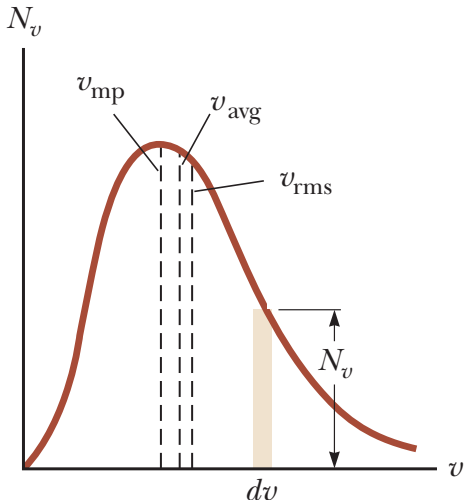
the process by which a liquid changes to a gas at the liquid surface

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 water, pigs wallow puddles, and dogs pant. All are trying to use evaporation of water to reduce body temperature.

# Evaporation and Distribution of Molecular Speeds

Distribution of speeds of molecules at a given temperature:



There are always some molecules with more kinetic energy than most molecules: those can vaporize at a lower temperature.



# Evaporation

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

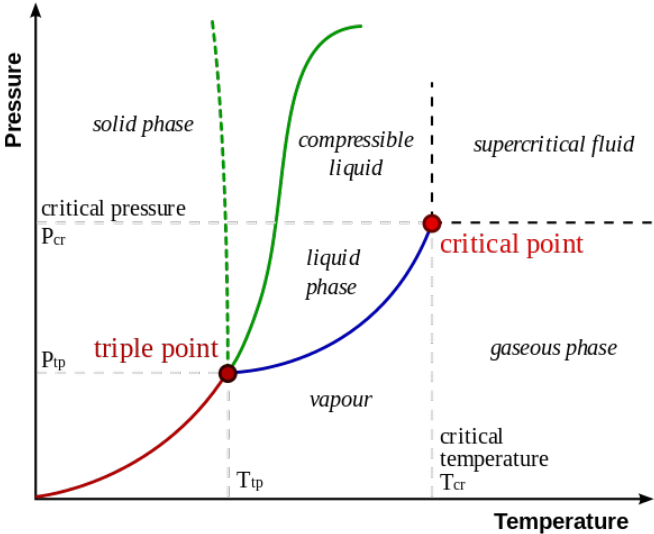
He decided to experiment to see if the temperature of objects could be lowered by this process.

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  $7^{\circ}\text{F}$  ( $-14^{\circ}\text{C}$ ) at the bulb of the thermometer.

This is the basic idea behind refrigeration!

# Phase Diagrams



# Thermodynamics

The name of the field literally means “movement of heat”.

Thermodynamics is the study of the relation between heat and work, and volume, pressure, temperature, and entropy.

All of the observable variables are macroscopic: they only refer to bulk properties of the fluids under study.

When thermodynamics was first developed, there was no understanding at all of atoms or microscopic interactions.

# Thermodynamics Motivation

Thermodynamics was developed in the 1800 because of a need to understand the relation between work and heat.

Steam engines were first experimented with 2000 years ago, and demonstrated practically in the 1600s.

By the mid 1800s, these engines were becoming economical. They became the cornerstone of transport and production in the industrial revolution.

Better understanding how to design them was crucial.

## Some quantities

We already introduced internal energy. Reminder:

**internal energy,  $E_{\text{int}}$  or  $U_{\text{int}}$**

The energy that an object has as a result of its temperature and all other molecular motions, effects, and configurations.

Later we will need another quantity also: entropy.

# First Law of Thermodynamics

## 1st Law

The change in the internal energy of a system is equal to the sum of the heat added to the system and the work done on the system.

$$\Delta U_{\text{int}} = W + Q$$

This is just the conservation of energy written in a different way!

It takes into account that heat is energy.

# Second Law

## 2nd Law

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This is obvious from experience, but it's not obvious *why* this should happen.

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It also indicates there are processes in the physical world that seem not to be time-reversible.



# Entropy

**Entropy is a measure of disorder in a system.**

It can also be used as a measure of information content.

Intriguingly, entropy was introduced separately in physics and then later in information theory. The fact that these two measures were the same was observed by John von Neumann.

# Entropy

According to Claude Shannon, who developed Shannon entropy, or information entropy:

“I thought of calling it ‘information’, but the word was overly used, so I decided to call it ‘uncertainty’. [...] Von Neumann told me, ‘You should call it entropy, for two reasons. In the first place your uncertainty function has been used in statistical mechanics under that name, so it already has a name. In the second place, and more important, nobody knows what entropy really is, so in a debate you will always have the advantage.’ ”

## So what is entropy?

Consider the Yo. app (valued at \$5-10 million in 2014).

You can only use it to send the message “yo.”

If you get a message on the app, you can guess what it will say:  
there was only 1 possible state of the message.

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You can only use it to send the message “yo.”

If you get a message on the app, you can guess what it will say: there was only 1 possible state of the message.

The message has no information content, and it is perfectly ordered, there is no uncertainty.

The entropy of the message is zero.

## So what is entropy?

But what if the message could be “yo” or “no”?

“If you get the yo, let’s meet for drinks, if no, I’m still in a meeting and can’t join you.”

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But what if the message could be “yo” or “no”?

“If you get the yo, let’s meet for drinks, if no, I’m still in a meeting and can’t join you.”

Now you learn something when you get the message: there were two possible states of the message.

The entropy of the message is greater than zero.

# Entropy in Thermodynamics

Consider the atmosphere, it is mostly Oxygen and Nitrogen.

Have you ever walked into a room and been unable to breathe because all of the oxygen is on the other side of the room?

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Have you ever walked into a room and been unable to breathe because all of the oxygen is on the other side of the room?



As more oxygen molecules are added, the probability that there is oxygen is on both sides increases.



# Macrostates and Microstates

A **macrostate** is something we can observe on a large scale.

The macrostates here could be:

- all oxygen on the left
- all oxygen on the right
- oxygen mixed throughout the room.



# Macrostates and Microstates

A **microstate** is a state too small / complex to easily observe, but represents one way a macrostate can be achieved.

We want to consider the number of microstates for each macrostate.

The macrostates here could be:

- all oxygen on the left — 1 microstate
- all oxygen on the right — 1 microstate
- oxygen mixed throughout the room — 6 microstates



## Entropy and disorder

Entropy,  $S$ , depends on the number of microstates that an observed macrostate has. The more microstates, the more entropy.

A macrostate that has very many microstates can be thought of as a disordered state.

If all the oxygen is on the left of the room, all the nitrogen on the right, the room is organized, or ordered.

But this is very unlikely!

Even if a room starts out ordered, you would expect it to become disordered right away, because a disordered room is more probable.

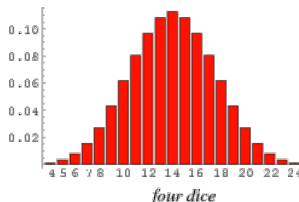
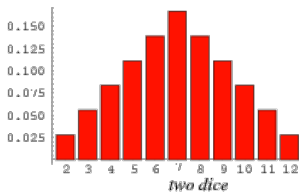
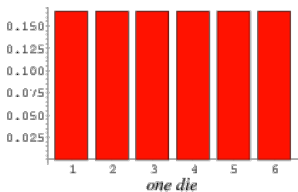
## Entropy example

Problem 8, page 331.

Construct a table of all the possible combinations of numbers that can come up when you throw 2 dice.

The score of the dice will be the sum of the two numbers on the dice. Why is 7 the most likely score?

# Entropy example



The macroscopic world is like a game in which there are  $\sim 10^{23}$  dice, but you only ever see the (approximate) score, not the result on each die.

<sup>1</sup>Figure Weistein, Eric W. "Dice." From MathWorld—A Wolfram Web Resource. <http://mathworld.wolfram.com/Dice.html>

# Second Law of Thermodynamics

This gives us another way to state the second law:

## 2nd Law

In an isolated system, entropy does not decrease.

equivalently:

$$\frac{\Delta S}{\Delta t} \geq 0$$

Heat must flow from hotter to colder because there are more ways to distribute heat evenly than to keep it in one place only.

As time goes by, things tend to disorder.

# Third Law of Thermodynamics

## 3rd Law

As the temperature of a material approaches zero, the entropy approaches a constant value.

The constant value the entropy takes is very small. It is actually zero if the lowest energy state of the material is unique.

Another way to express the third law:

## 3rd Law - alternate

It is impossible to reach absolute zero using any procedure and only a finite number of steps.

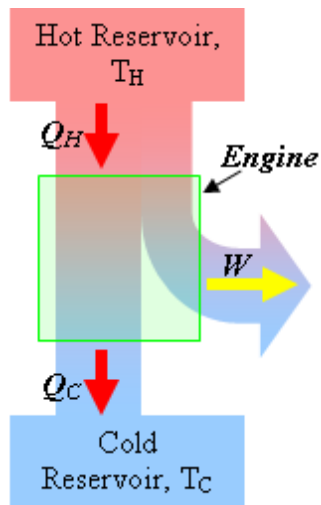
# Heat Engines

Steam engines and later incarnations of the engine run on a very simple principle: heat is transferred from a hot object to a colder object and mechanical work is done in the process.

Usually, some chemical energy (burning fuel) is used to raise the temperature of one object, and the colder object remains at the ambient temperature.



# Heat Engines



<sup>1</sup>Diagram from <http://www2.ignatius.edu/faculty/decarlo/>

## Efficiency of a Heat Engine

Efficiency:

$$e = \frac{Q_H - Q_C}{Q_H} = \frac{W}{Q_H}$$

An ideal engine, one that has the highest possible efficiency (a Carnot engine), has efficiency:

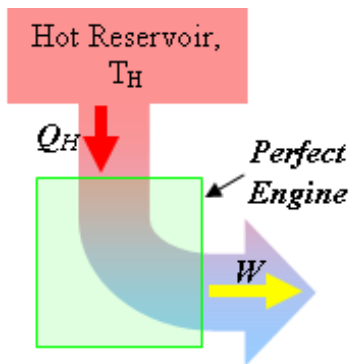
$$e = \frac{T_H - T_C}{T_H}$$

(T is measured in Kelvin!)

It is not possible for any engine to have efficiency higher than this without violating the 2nd law.

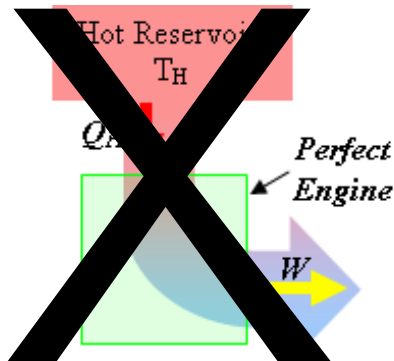
## “Perfect” but Impossible Engine

It would be nice if all heat energy  $Q_H$  could be converted to work.



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But this is not possible.

That would require  $T_C = 0$ , so that  $W = Q_H$ . Cannot happen.

# Second Law and Heat Engines

We can state the second law also as a fundamental limitation on heat engines.

## Second Law of Thermodynamics (Heat Engine version)

It is impossible to construct a heat engine that, operating in a cycle, produces no effect other than the input of energy by heat from a reservoir and the performance of an equal amount of work.

# Summary

- phases of matter and phase change
- Laws of thermodynamics
- entropy
- absolute zero
- heat engines
- practical engines

## Talks, Essay

## Homework Hewitt,

- Ch 17, onward from page 312. Exercises: 1, 11; Problems: 1
- Ch 18, onward from page 329. Exercises: 3, 7, 11, 17, 25, 29. Problems 1, 5, 7.