Thermal Physics
**Efficiency**

In this chapter, we’ll look at practical limits on energy transfers and transformations.

---

**Looking Back**

10.1 Basic energy concepts; forms of energy  
10.6 The law of conservation of energy

---

Both bulbs put out the same amount of light, but the one on the right uses 1/4 the electric power. Both bulbs perform the same transformation, but one is much more efficient.

---

**Energy in the Body**

All the energy that your body uses for all of the tasks you complete during the day comes from food. How efficient is your body at converting this energy? How much energy does your body actually use to run, climb, and move?

---

Climbing stairs requires a change in potential energy. How much energy does your body use to make this climb?

---

These fighters wear masks that allow a direct measurement of the energy used by their bodies as they exercise.
Thermal Energy and Temperature
An object’s temperature is related to its thermal energy, the energy of motion on an atomic scale.

We’ll use the ideal gas model to help us understand the nature of thermal energy and temperature.

Looking Back  
10.7 Thermal energy

Heat and Thermodynamics
Processes in which only thermal energy changes are the domain of thermodynamics.

The thermal energy of the kettle is increased by the heat from the burner, which is at a higher temperature.
Heat Engines and Heat Pumps

A **heat engine** can convert thermal energy into other forms of energy. A **heat pump** moves thermal energy from one place to another.

- This geothermal plant uses volcanic thermal energy to generate electricity. “Waste” heat warms the lagoon. Why must any energy be “wasted”?
- The inside of a refrigerator is cold because heat has been “pumped” out. What’s the energy cost to move this heat?

Entropy

Entropy is a measure of disorder at an atomic level that helps us explain some basic observations about the world.

When you stir a cup of coffee, the cream mixes in. This is irreversible; stirring backward won’t cause it to unmix. The concept of entropy explains why the future is different from the past and why there are theoretical limits on energy use.
SUMMARY

The goals of Chapter 11 have been to learn more about energy transformations and transfers, the laws of thermodynamics, and theoretical and practical limitations on energy use.

GENERAL PRINCIPLES

Energy and Efficiency

When energy is transformed from one form into another, some may be "lost" usually to thermal energy, due to practical or theoretical constraints. This limits the efficiency of processes. We define efficiency as:

\[ e = \frac{\text{what you get}}{\text{what you had to pay}} \]

The Laws of Thermodynamics

The first law of thermodynamics is a statement of conservation of energy for systems in which only thermal energy changes:

\[ \Delta E_a = W + Q \]

The second law of thermodynamics specifies the way that isolated systems can evolve:

The entropy of an isolated system always increases.

This law has practical consequences:
- Heat energy spontaneously flows only from hot to cold.
- A transformation of energy into thermal energy is irreversible.
- No heat engine can be 100% efficient.

Entropy and Irreversibility

Systems move toward more probable states. These states have higher entropy — more disorder. This change is irreversible. Changing other forms of energy to thermal energy is irreversible.

Increasing probability Increasing entropy

Heat is energy transferred between two objects at different temperatures. Energy will be transferred until thermal equilibrium is reached.

\[ T_i \quad Q \quad T_j \quad T_i < T_j \]

IMPORTANT CONCEPTS

Thermal energy

- For a gas, the thermal energy is the total kinetic energy of motion of the atoms.
- Thermal energy is random kinetic energy and so has entropy.

Temperature

- For a gas, temperature is proportional to the average kinetic energy of the motion of the atoms.
- Two systems are in thermal equilibrium if they are at the same temperature. No heat energy is transferred at thermal equilibrium.

\[ E_{\text{kin}} = \frac{3}{2} N k_B T \]

\[ T = \frac{2}{3} \frac{E_{\text{kin}}}{k_B} \]

A heat engine converts thermal energy from a hot reservoir into useful work. Some heat is exhausted into a cold reservoir, limiting efficiency.

\[ e_{\text{max}} = 1 - \frac{T_c}{T_H} \]

\[ \text{COP}_{\text{max}} = \frac{1}{\frac{T_H}{T_c} - 1} \]

A heat pump uses an energy input to transfer heat from a cold side to a hot side. The coefficient of performance is analogous to efficiency. For cooling, it is:

EFFICIENCIES

Energy in the body

Cells in the body metabolize chemical energy in food. Efficiency for most actions is about 25%.

\[ \text{Energy used by body} \quad 480 \text{ W} \]

\[ \text{Energy for forward propulsion at rate of} \quad 120 \text{ W} \]

\[ \text{Waste heat} \]

\[ \text{Chemical energy in} \]

\[ \text{Useful work out} \]

Power plants

A typical power plant converts about 1/3 of the energy input into useful work. The rest is exhausted as waste heat.

\[ T(K) = T(\text{C}) + 273 \]

Temperature scales

Zero on the Kelvin temperature scale is the temperature at which the kinetic energy of atoms is zero. This is absolute zero. The conversion from °C to K is:

\[ T(K) = T(\text{C}) + 273 \]

All temperatures in equations must be in kelvin.
Reading Quiz

1. A machine uses 1000 J of electric energy to raise a heavy mass, increasing its potential energy by 300 J. What is the efficiency of this process?

   A. 100%
   B. 85%
   C. 70%
   D. 35%
   E. 30%
Answer

1. A machine uses 1000 J of electric energy to raise a heavy mass, increasing its potential energy by 300 J. What is the efficiency of this process?

A. 100%
B. 85%
C. 70%
D. 35%
E. 30%
Reading Quiz

2. When the temperature of an ideal gas is increased, which of the following also increases? (1) The thermal energy of the gas; (2) the average kinetic energy of the gas; (3) the average potential energy of the gas; (4) the mass of the gas atoms; (5) the number of gas atoms.

A. 1, 2, and 3
B. 1 and 2
C. 4 and 5
D. 2 and 3
E. All of 1–5
2. When the temperature of an ideal gas is increased, which of the following also increases? (1) The thermal energy of the gas; (2) the average kinetic energy of the gas; (3) the average potential energy of the gas; (4) the mass of the gas atoms; (5) the number of gas atoms.

A. 1, 2, and 3
B. 1 and 2
C. 4 and 5
D. 2 and 3
E. All of 1–5
Reading Quiz

3. A refrigerator is an example of a

A. reversible process.
B. heat pump.
C. cold reservoir.
D. heat engine.
E. hot reservoir.
Answer

3. A refrigerator is an example of a

   A. reversible process.
   B. heat pump.
   C. cold reservoir.
   D. heat engine.
   E. hot reservoir.
Heat Energy is a flow of energy from hotter to colder because of a difference in temperature. Objects do not have heat. [Heat] = Joule

Heat Energy entering or leaving a system will cause either a Temperature Change: \( Q = mc\Delta T \) or a Phase Change: \( Q = mL \)
Any two systems placed in thermal contact will have an exchange of heat energy until they reach the same temperature.

If the systems are in thermal equilibrium then no changes take place.
• **Heat Energy** is a flow of energy from hotter to colder because of a difference in temperature. Objects do not have heat. \([\text{Heat}] = \text{Joule}\)

• **Internal Energy** of a system is a measure of the total Energy due to ALL random molecular motions INTERNAL of the system (Translations KE, Rotational KE, Vibrational KE) and internal POTENTIAL energies due to interactive forces (electromagnetic, strong, weak, gravitational) Objects have energy.

• **Mechanical Energy** is due to the kinetic and potential energies of the system itself in an external reference frame.

• **Mechanical Equivalent of Heat**: mechanical energy converted to heat energy by doing work on the system: \(1.000 \text{ kcal} = 4186\text{J}\)
Zeroeth Law

- Two systems individually in thermal equilibrium with a third system (such as a thermometer) are in thermal equilibrium with each other.
- That is, there is no flow of heat within a system in thermal equilibrium.
Checking Understanding

When you walk at a constant speed on level ground, what energy transformation is taking place?

A. $E_{chem} \rightarrow U_g$
B. $U_g \rightarrow E_{th}$
C. $E_{chem} \rightarrow K$
D. $E_{chem} \rightarrow E_{th}$
E. $K \rightarrow E_{th}$
Answer

When you walk at a constant speed on level ground, what energy transformation is taking place?

A. $E_{\text{chem}} \rightarrow U_g$
B. $U_g \rightarrow E_{\text{th}}$
C. $E_{\text{chem}} \rightarrow K$
D. $E_{\text{chem}} \rightarrow E_{\text{th}}$
E. $K \rightarrow E_{\text{th}}$
Heat flows spontaneously from a substance at a higher temperature to a substance at a lower temperature and does not flow spontaneously in the reverse direction.

Heat flows from hot to cold.
Heat is never spontaneously transferred from a colder object to a hotter object.
It is not possible to lower the temperature of any system to absolute zero.

Zero Temperature means zero volume ($PV=nRT$) and that is not possible since atoms take up space.
Hydrogen bomb
100,000,000 K

Center of the sun
20,000,000 K

Surface of a hot star
50,000 K

Plasma
20,000 K

Surface of the sun
6000 K

Carbon arc lamp
4300

Iron melts
1800 K

Tin melts
500 K

Water boils
400 K

Ice melts
273 K

Ammonia boils
300 K

Dry ice vaporizes
200 K

Oxygen boils
100 K

Helium boils
100 K

All molecules have broken up: no solids or liquids

-273°C  0 K
What is "room temperature" (68 degrees F) in Celsius and Kelvin?

\[ T(\degree C) = \frac{5}{9} \left( T(\degree F) - 32 \right) \]

\[ = \frac{5}{9} \left( 68 - 32 \right) = 20 \degree C \]

\[ T(K) = T(\degree C) + 273.15 \]

\[ = 293.15 K \]
Temperature

Celsius Chant

30 is HOT.
20 is NICE.
10 is CHILLY.
Zero is ICE!
• The change of internal energy of a system due to a temperature or phase change is given by:

  Temperature Change: \[ Q = mc\Delta T \]
  Phase Change: \[ Q = mL \]

• Q is positive when the system GAINS heat and negative when it LOSES heat.
First law of thermodynamics: For systems in which only the thermal energy changes, the change in thermal energy is equal to the energy transferred into or out of the system as work $W$, heat $Q$, or both:

$$\Delta E_{th} = W + Q$$

**Diagram:**
- **Environment**
- **System**
- Work on system ($W > 0$)
- Work by system ($W < 0$)
- Energy in ($Q > 0$)
- Energy out ($Q < 0$)
- Heat to system
- Heat from system
Consider your body as a system. Your body is “burning” energy in food, but staying at a constant temperature. This means that, for your body,

A. $Q > 0$.
B. $Q = 0$.
C. $Q < 0$.

**First law of thermodynamics** For systems in which only the thermal energy changes, the change in thermal energy is equal to the energy transferred into or out of the system as work $W$, heat $Q$, or both:

$$\Delta E_{th} = W + Q$$
Answer

Consider your body as a system. Your body is “burning” energy in food, but staying at a constant temperature. This means that, for your body,

A. \( Q > 0 \).
B. \( Q = 0 \).
C. \( Q < 0 \).

First law of thermodynamics For systems in which only the thermal energy changes, the change in thermal energy is equal to the energy transferred into or out of the system as work \( W \), heat \( Q \), or both:

\[ \Delta E_{\text{th}} = W + Q \]
Temperature

Temperature is measured by a thermometer. Kelvin is the *Absolute* Scale.

\[
T(\degree F) = \frac{9}{5} T(\degree C) + 32
\]

\[
T(\degree C) = \frac{5}{9} \left[ T(\degree F) - 32 \right]
\]

\[
T(K) = T(\degree C) + 273.15
\]
Rank the following temperatures, from highest to lowest.

A. 300 °C > 300 K > 300 °F
B. 300 K > 300 °C > 300 °F
C. 300 °F > 300 °C > 300 K
D. 300 °C > 300 °F > 300 K

\[
T(\degree F) = \frac{9}{5} T(\degree C) + 32
\]

\[
T(\degree C) = \frac{5}{9} \left[ T(\degree F) - 32 \right]
\]

\[
T(K) = T(\degree C) + 273.15
\]

Temperature differences are the same on the Celsius and Kelvin scales. The temperature difference between the freezing point and boiling point of water is 100°C or 100 K.
Answer
Rank the following temperatures, from highest to lowest.

A. 300 °C > 300 K > 300 °F
B. 300 K > 300 °C > 300 °F
C. 300 °F > 300 °C > 300 K
D. 300 °C > 300 °F > 300 K

Temperature differences are the same on the Celsius and Kelvin scales. The temperature difference between the freezing point and boiling point of water is 100°C or 100 K.
Hot Question

Suppose you apply a flame to 1 liter of water for a certain time and its temperature rises by 10 degrees C. If you apply the same flame for the same time to 2 liters of water, by how much will its temperature rise?

a) 1 degree  b) 5 degrees  c) 10 degrees  d) zero degrees

Think Kinetic!

Hot stove
Efficiency

1. Electric energy runs a pump to push water uphill.
2. Some energy is “lost” to thermal energy.
3. Less than 100 J is stored.
4. The water flows downhill and runs a generator. Some energy is “lost” to thermal energy in this process.
5. At the end, half the original electric energy is recovered.

\[
e = \frac{\text{what you get}}{\text{what you had to pay}}
\]

General definition of efficiency
1. Burning coal produces 100 J of thermal energy.

2. Steam with 100 J of thermal energy enters the turbine.

3. The turbine turns a generator, producing 35 J of electric energy.

4. 65 J of thermal energy is exhausted into the environment.
Operation of a Heat Engine

1. Heat energy $Q_H$ is transferred from the hot reservoir to the system.

2. Part of the energy is used to do useful work $W_{out}$.

3. The remaining energy $Q_C = Q_H - W_{out}$ is exhausted to the cold reservoir as waste heat.

Hot reservoir $T_H$

Heat engine

Cold reservoir $T_C$
Operation of a Heat Pump

The amount of heat exhausted to the hot reservoir is larger than the amount of heat extracted from the cold reservoir.

External work is used to remove heat from a cold reservoir and exhaust heat to a hot reservoir.
Coefficient of Performance of a Heat Pump

$\text{COP}_{\text{max}} = \frac{T_C}{T_H - T_C}$

Theoretical maximum coefficient of performance of a heat pump used for cooling

$\text{COP}_{\text{max}} = \frac{T_H}{T_H - T_C}$

Theoretical maximum coefficient of performance of a heat pump used for heating
Checking Understanding: Increasing Efficiency of a Heat Pump

Which of the following changes would allow your refrigerator to use less energy to run? (1) Increasing the temperature inside the refrigerator; (2) increasing the temperature of the kitchen; (3) decreasing the temperature inside the refrigerator; (4) decreasing the temperature of the kitchen.

A. All of the above
B. 1 and 4
C. 2 and 3
Answer

Which of the following changes would allow your refrigerator to use less energy to run? (1) Increasing the temperature inside the refrigerator; (2) increasing the temperature of the kitchen; (3) decreasing the temperature inside the refrigerator; (4) decreasing the temperature of the kitchen.

A. All of the above
B. 1 and 4
C. 2 and 3
Additional Questions

The following pairs of temperatures represent the temperatures of hot and cold reservoirs for heat engines. Which heat engine has the highest possible efficiency?

A. 300°C 30°C
B. 250°C 30°C
C. 200°C 20°C
D. 100°C 10°C
E. 90°C 0°C
The following pairs of temperatures represent the temperatures of hot and cold reservoirs for heat engines. Which heat engine has the highest possible efficiency?

A. 300°C 30°C
B. 250°C 30°C
C. 200°C 20°C
D. 100°C 10°C
E. 90°C 0°C
Atomic Units

The Basics

• Atomic Number: # protons
• Atomic Mass: # atomic mass units (u)
• Atomic Mass Unit: 1/12 mass of C-12 atom
• amu = u = 1.66 x 10^{-27} kg
• Atomic Mass of C = 12.011u (1% is C-13)
• Mass of 1 C = (12.011u) (1.66 x 10^{-27} kg/u)
More on Moles

The mass / mol for any substance has the same numerical value as its atomic mass:

\[ \text{mass/mol C-12} = 12 \text{ g / mol} \]
\[ \text{mass/mol Li} = 6.941 \text{ g / mol} \]

\[ n = \frac{\text{mass}}{(\text{mass/mole})} = \frac{\text{mass}}{\text{atomic mass}} \]
Q: How many moles are in 1 kg of Sodium?

mass/mole = atomic mass
Na: 22.9898 g / mol
n = mass / (mass/mole)
  = 1000 g / (22.9898g/mol)
  = 435 moles

Q: How many atoms in 1 kg of Sodium?

# particles in a sample is:  \( N = n \, N_A \)

\[ N = (435\text{mol}) \times 6.022 \times 10^{23} \text{mol}^{-1} \]

= \( 2.62 \times 10^{25} \) atoms
Moles and Avogadro’s Number

\[ N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \]

• Mole (mol) = \# atoms or molecules (particles) as are in 12 grams of Carbon-12:

\[ 1 \text{ mole} = 6.022 \times 10^{23} \text{ particles} \]

• Avogadro’s Number: the number of particles in one mole:

\[ N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \]

• \# moles \( n \) contained in a sample of \( N \) particles:

\[ n = \frac{N}{N_A} \]

• \# particles in a sample is:

\[ N = n \times N_A \]
The Kinetic Theory

- Temperature ~ Average KE of each particle
- Particles have different speeds
- Gas Particles are in constant RANDOM motion
- Average KE of each particle is: \( \frac{3}{2} kT \)
- Pressure is due to momentum transfer

Speed ‘Distribution’ at CONSTANT Temperature is given by the Maxwell Speed Distribution
The Ideal Gas Model

\[ K_{\text{ave}} = \frac{3}{2} k_B T \]

\[ E_{\text{th}} = N K_{\text{ave}} = \frac{3}{2} N k_B T \]

Thermal energy of an ideal gas of \( N \) atoms

\[ T = \frac{2}{3} \frac{K_{\text{ave}}}{k_B} \]

3. The temperature also increases.

1. Heat is added to an ideal gas.

2. This heat increases the kinetic energy of the gas atoms.
The Kelvin Temperature of an ideal gas is a measure of the average translational kinetic energy per particle:

\[
\frac{3}{2} kT = KE = \frac{1}{2} m v_{rms}^2
\]

\( k = 1.38 \times 10^{-23} \text{ J/K} \) Boltzmann’s Constant

Root-mean-square speed:

\[
v_{rms} = \sqrt{v^2} = \sqrt{\frac{3kT}{m}}
\]
Kinetic Theory Problem

Calculate the RMS speed of an oxygen molecule in the air if the temperature is 5.00 °C. The mass of an oxygen molecule is 32.00 u 
(\( k = 1.38 \times 10^{-23} \text{ J/K}, \ u = 1.66 \times 10^{-27} \text{ kg} \))

\[ v_{rms} = \sqrt{\frac{3kT}{m}} \]

What is \( m \)?

\( m \) is the mass of one oxygen molecule in kg.

What is \( u \)?

How do we get the mass in kg?
Kinetic Theory Problem

Calculate the RMS speed of an oxygen molecule in the air if the temperature is 5.00 °C. The mass of an oxygen molecule is 32.00 u

\[ k = 1.38 \times 10^{-23} \text{ J/K, } u = 1.66 \times 10^{-27} \text{ kg} \]

\[
\nu_{rms} = \sqrt{\frac{3kT}{m}}
\]

\[
= \sqrt{\frac{3(1.38 \times 10^{-23} \text{ J/K})278K}{(32u)(1.66 \times 10^{-27} \text{ kg/u})}}
\]

\[ = 466 m/s \]

What is \( m \)?

\( m \) is the mass of one oxygen molecule.

Is this fast? \( \text{YES!} \)
Maxwell Speed Distribution

Most probable speed is near 400 m/s

Most probable speed is near 800 m/s

300 K

1200 K

Percentage of molecules per unit speed interval

Molecular speed, m/s
Checking Understanding

Two containers of the same gas (ideal) have these masses and temperatures:

• Which gas has atoms with the largest average thermal energy?
• Which container of gas has the largest thermal energy?

A. P, Q
B. P, P
C. Q, P
D. Q, Q

\[ E_{\text{th}} = N K_{\text{avg}} = \frac{3}{2} N k_B T \]

Thermal energy of an ideal gas of \( N \) atoms
Two containers of the same gas (ideal) have these masses and temperatures:

- Which gas has atoms with the largest average thermal energy?
- Which container of gas has the largest thermal energy?

A. P, Q
B. P, P
C. Q, P
D. Q, Q

Answer

\[ E_{th} = NK_{\text{avg}} = \frac{3}{2} Nk_B T \]

Thermal energy of an ideal gas of \( N \) atoms
More Kinetic Theory Problems

A gas molecule with a molecular mass of 32.0 u has a speed of 325 m/s. What is the temperature of the gas molecule?

A) 72.0 K  B) 136 K  C) 305 K  D) 459 K

E) A temperature cannot be assigned to a single molecule.

Temperature ~ Average KE of all particles
The absolute Pressure $P$ of an ideal gas is directly proportional to the absolute (Kelvin) temperature $T$ and the number of moles $n$ of the gas and inversely proportional to the volume $V$ of the gas:

$$PV = nRT$$

$n$ = # moles

$R = 8.31 \text{ J/(mol-K)}$ Universal Gas Constant
Ideal Gas Law

\[ PV = nRT \]

- \( n \) = # moles
- \( R = 8.31 \text{ J/(mol-K)} \) Universal Gas Constant

\[ PV = Nkt \]

- \( N \) = # particles
- \( k = 1.38 \times 10^{-23} \text{ J/K} \) Boltzmann’s Constant

Note: \( PV \) is units of Energy!
The only interaction between particles are elastic collisions (no sticky - no loss of KE)
This requires LOW DENSITY
Excellent Approximation for O, N, Ar, CO2 @ room temperature and pressures
“State” is described by the Ideal Gas Law
Non “Ideal” are Van der Waals gases
Ideal Gas Problem

An ideal gas with a fixed number of molecules is maintained at a constant pressure. At 30.0 °C, the volume of the gas is 1.50 m³. What is the volume of the gas when the temperature is increased to 75.0 °C?

\[
\frac{PV_1}{nRT_1} = \frac{V_1}{T_1} = \frac{PV_2}{nRT_2} = \frac{V_2}{T_2}
\]

\[
V_2 = V_1 \frac{T_2}{T_1} = 1.5m^3 \frac{348K}{303K} = 1.72m^3
\]
Thermal Expansion: Linear

\[ \Delta L = \alpha L_0 \Delta T \]

Coefficients determined experimentally!
Thermal Expansion: Volume

\[ \Delta V = \beta V_0 \Delta T \]

\[ \beta \sim 3\alpha \]
Most Solids and Liquids Expand upon Heating

The Exception: Water

1. Liquid water below 4°C is bloated with ice crystals.
2. Upon warming, the crystals collapse, resulting in a smaller volume for the liquid water.
3. Above 4°C, liquid water expands as it is heated because of greater molecular motion.
Thermal Expansion: Linear
Thermal Expansion: Linear

The coefficient of linear expansion of steel is $12 \times 10^{-6}/{^\circ C}$. A railroad track is made of individual rails of steel 1.0 km in length. By what length would these rails change between a cold day when the temperature is -10 °C and a hot day at 30 °C?

\[
\Delta L = \alpha L_0 \Delta T
\]

\[
\Delta L = (12 \times 10^{-6} \ /{^\circ C})(10^3 \ m)(30 \ ^\circ C - (-10 \ ^\circ C))
\]

\[
\Delta L = .48 \ m
\]
Thermal Expansion: Linear

\[ \Delta L = \alpha L_0 \Delta T \]

What change in temperature is needed to fill the gap, 1.3 x 10^{-3} m?

\[ \alpha_{\text{brass}} = 19 \times 10^{-6} \text{ } ^0\text{C}^{-1} \quad \alpha_{\text{Al}} = 23 \times 10^{-6} \text{ } ^0\text{C}^{-1} \]

\[ \Delta L_{\text{brass}} + \Delta L_{\text{Al}} = 1.3 \times 10^{-3} m \]

\[ \Delta T = \frac{1.3 \times 10^{-3} m}{\alpha_{\text{brass}} L_{\text{brass}} + \alpha_{\text{Al}} L_{\text{Al}}} = 11^0\text{C} \]
When the temperature of a metal ring increases, does the hole become larger? Smaller? Or stay same?
Circle Expansion

The coefficient of linear expansion of aluminum is $23 \times 10^{-6}/°C$. A circular hole in an aluminum plate is 2.725 cm in diameter at 0°C. What is the diameter of the hole if the temperature of the plate is raised to 100°C?

\[
\Delta L = \alpha L_0 \Delta T
\]

\[
= (23 \times 10^{-6} / °C)(2.725cm)100 °C
\]

\[
= 6.3 \times 10^{-3} \, cm \quad d = 2.731cm
\]
• **Heat Energy** is a flow of energy from hotter to colder because of a difference in temperature. Objects do not have heat. [Heat] = Joule

• Heat Energy entering or leaving a system will cause either a **Temperature Change:** \( Q = mc\Delta T \) or a **Phase Change:** \( Q = mL \)
First Law of Thermo

• The change of internal energy of a system due to a temperature or phase change is given by:

  Temperature Change: \[ Q = mc\Delta T \]
  Phase Change: \[ Q = mL \]

• \( Q \) is positive when the system GAINS heat and negative when it LOSES heat.
Specific Heat:  *Thermal Inertia*

The Specific Heat of a substance is the amount of Energy it requires to raise the temperature of 1 kg, 1 degree Celsius.

\[ Q = mc\Delta T \]

\[ c = \frac{Q}{m\Delta T} = \frac{J}{kg \cdot ^0C} \]

- The higher the specific heat, the more energy it takes and the longer it takes to heat up and to cool off.

- The lower the specific heat, the less energy it takes and the quicker it takes to heat up and cool off.

- Substances with HIGH specific heat STORE heat energy and make good thermal moderators. (Ex: Water, Oceans)
## Some Specific Heat Values

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat $c$ (J/kg (\cdot) °C)</th>
<th>Specific heat $c$ (cal/g (\cdot) °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Elemental solids</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aluminum</td>
<td>900</td>
<td>0.215</td>
</tr>
<tr>
<td>Beryllium</td>
<td>1830</td>
<td>0.436</td>
</tr>
<tr>
<td>Cadmium</td>
<td>230</td>
<td>0.055</td>
</tr>
<tr>
<td>Copper</td>
<td>387</td>
<td>0.0924</td>
</tr>
<tr>
<td>Germanium</td>
<td>322</td>
<td>0.077</td>
</tr>
<tr>
<td>Gold</td>
<td>129</td>
<td>0.0308</td>
</tr>
<tr>
<td>Iron</td>
<td>448</td>
<td>0.107</td>
</tr>
<tr>
<td>Lead</td>
<td>128</td>
<td>0.0305</td>
</tr>
<tr>
<td>Silicon</td>
<td>703</td>
<td>0.168</td>
</tr>
<tr>
<td>Silver</td>
<td>234</td>
<td>0.056</td>
</tr>
</tbody>
</table>
More Specific Heat Values

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat $c$</th>
<th>J/kg·°C</th>
<th>cal/g·°C</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Other solids</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Brass</td>
<td>380</td>
<td></td>
<td>0.092</td>
</tr>
<tr>
<td>Glass</td>
<td>837</td>
<td></td>
<td>0.200</td>
</tr>
<tr>
<td>Ice ($-5, ^\circ$C)</td>
<td>2090</td>
<td></td>
<td>0.50</td>
</tr>
<tr>
<td>Marble</td>
<td>860</td>
<td></td>
<td>0.21</td>
</tr>
<tr>
<td>Wood</td>
<td>1700</td>
<td></td>
<td>0.41</td>
</tr>
<tr>
<td><strong>Liquids</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Alcohol (ethyl)</td>
<td>2400</td>
<td></td>
<td>0.58</td>
</tr>
<tr>
<td>Mercury</td>
<td>140</td>
<td></td>
<td>0.033</td>
</tr>
<tr>
<td>Water ($15, ^\circ$C)</td>
<td>4186</td>
<td></td>
<td>1.00</td>
</tr>
<tr>
<td><strong>Gas</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Steam ($100, ^\circ$C)</td>
<td>2010</td>
<td></td>
<td>0.48</td>
</tr>
</tbody>
</table>
Specific Heat

\[ c_{\text{water}} = 4186 \frac{J}{kg \cdot ^0C} \]

\[ c_{\text{glycerin}} = 2410 \frac{J}{kg \cdot ^0C} \]

\[ c_{\text{iron}} = 452 \frac{J}{kg \cdot ^0C} \]

Why does water have such a high specific heat?

Heat goes into other modes of energy so that temperature changes slowly.
$$Q = mc\Delta T$$

How much heat is required to raise the temperature of a 0.750kg aluminum pot containing 2.50kg of water at 30°C to the boiling point?

$$Q = m_{Al}c_{Al}\Delta T + m_{w}c_{w}\Delta T$$

$$= (m_{Al}c_{Al} + m_{w}c_{w})\Delta T$$

$$= \left[.75\text{kg}(900\text{J/kg}^\circ\text{C}) + 2.5\text{kg}(4186\text{J/kg}^\circ\text{C}) \right](70^\circ\text{C})$$

$$Q = 7.798 \times 10^5 \text{ J}$$
Phase Change $Q = mL$

- A change from one phase to another
- A phase change always occurs with an exchange of energy!
- A phase change always occurs at constant temperature!
Sample Latent Heat Values

\[ Q = mL \]

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting Point (°C)</th>
<th>Latent Heat of Fusion (J/kg)</th>
<th>Boiling Point (°C)</th>
<th>Latent Heat of Vaporization (J/kg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td>-269.65</td>
<td>(5.23 \times 10^3)</td>
<td>-268.93</td>
<td>(2.09 \times 10^4)</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>-209.97</td>
<td>(2.55 \times 10^4)</td>
<td>-195.81</td>
<td>(2.01 \times 10^5)</td>
</tr>
<tr>
<td>Oxygen</td>
<td>-218.79</td>
<td>(1.38 \times 10^4)</td>
<td>-182.97</td>
<td>(2.13 \times 10^5)</td>
</tr>
<tr>
<td>Ethyl alcohol</td>
<td>-114</td>
<td>(1.04 \times 10^5)</td>
<td>78</td>
<td>(8.54 \times 10^5)</td>
</tr>
<tr>
<td>Water</td>
<td>0.00</td>
<td>(3.33 \times 10^5)</td>
<td>100.00</td>
<td>(2.26 \times 10^6)</td>
</tr>
<tr>
<td>Sulfur</td>
<td>119</td>
<td>(3.81 \times 10^4)</td>
<td>444.60</td>
<td>(3.26 \times 10^5)</td>
</tr>
<tr>
<td>Lead</td>
<td>327.3</td>
<td>(2.45 \times 10^4)</td>
<td>1750</td>
<td>(8.70 \times 10^5)</td>
</tr>
<tr>
<td>Aluminum</td>
<td>660</td>
<td>(3.97 \times 10^5)</td>
<td>2450</td>
<td>(1.14 \times 10^7)</td>
</tr>
<tr>
<td>Silver</td>
<td>960.80</td>
<td>(8.82 \times 10^4)</td>
<td>2193</td>
<td>(2.33 \times 10^6)</td>
</tr>
<tr>
<td>Gold</td>
<td>1063.00</td>
<td>(6.44 \times 10^4)</td>
<td>2660</td>
<td>(1.58 \times 10^6)</td>
</tr>
<tr>
<td>Copper</td>
<td>1083</td>
<td>(1.34 \times 10^5)</td>
<td>1187</td>
<td>(5.06 \times 10^6)</td>
</tr>
</tbody>
</table>
Phase Change

Energy goes into the system and breaks molecular bonds.

Energy is given up by the system by forming molecular bonds.
Phase Change: Melting & Freezing

Melting: Energy goes into the system and breaks molecular bonds.

Freezing: Energy is given up by the system by forming molecular bonds.
Phase Change: Melting & Freezing

(a) Ice

Liquid water
Phase Change: Melting & Freezing

- Melting: Solid to Liquid @ the melting temperature
- Melting is a *cooling* process
- Freezing: Liquid to Solid @ the melting temperature
- Freezing is a *warming* process.
Why do farmers spray peaches with water to save them from frost?

Freezing is a warming process!
If you were in an igloo on a freezing night. You would be warmed more by

a) a bucket of ice melting.

b) a bucket of water freezing

c) the same either way.

d) neither - are you nuts?
Phase Change: Evaporation

• Takes place at the surface of a liquid due to escaping molecules.
• Occurs at all temperatures
• Evaporation occurs when water vapor pressure in the liquid exceeds the pressure of water vapor in the surrounding air.
• Evaporation is a cooling process.
Evaporation is a Cooling Process

1. Liquid water molecule having sufficient kinetic energy to overcome surface hydrogen bonding approaches liquid surface.

2. Liquid water cooled as it loses this high-speed water molecule.
Phase Change: Boiling

• Boiling is evaporation under the surface of the liquid.
• Liquid boils at the temperature for which its vapor pressure exceeds the external pressure (mostly atmospheric pressure.)
• Boiling point depends on temperature AND pressure:
  • @ 1 atm, bp of water is 100ºC, @ 5 atm, bp of water is 374 ºC
• Boiling is a cooling process.
• At low pressures, liquids are boiled (‘freeze-dried’) into solids.
Phase Change: Condensation

- Gas molecules condense to form a liquid.
- Condensation is a *warming* process.
- Why is a rainy day warmer than a cloudy or clear day in winter?
- Why do we feel uncomfortable on a muggy day?
Condensation is a Warming Process

Fast-moving water vapor molecule bounces off surface

Slow-moving water vapor molecule sticks to liquid surface

Gas warmed by removal of slower molecule

Liquid warmed by formation of hydrogen bonds
Phase Change: Humidity

• Vapor is the gas phase of a substance below its boiling temperature.
• Air can ‘hold’ only so much water vapor before it becomes saturated and condensation occurs. Humidity is a measure of vapor density.
• Warm air can hold more water vapor. More condensation occurs at cooler temperatures because the molecules are moving slower.

Slow moving water molecules coalesce upon collision.
Windward: *Wet*
Leeward: *Dry*

Cools and condenses at Top

Warm
Humid
Air
Pushed
Up

Warm
Dry
Air
Falls
Down
Stormy Weather

When warm air rises, it expands and cools. The water vapor in the air soon condenses into water droplets, which form clouds and eventually these droplets fall from the sky as rain.
Phase Change: Sublimation

The conversion of a solid directly to a gas & visa versa
Examples: snowflakes, Moth Balls, dry ice
Phase Change: Triple Point

A temperature and pressure at which all three phases exist in equilibrium.

Freezing-Melting

Evaporation-Condensation

Sublimation

Lines of equilibrium
Phase Change \[ Q = mL \]

Phase change occurs at a *Constant Temperature*! Latent Heats of: Fusion & Evaporation \( L_f, L_v \)

Water:

\[ L_f = 334 \text{ kJ/kg} \quad (\text{solid-liquid}) \]
\[ L_v = 2256 \text{ kJ/kg} \quad (\text{liquid-gas}) \]
Phase Change: Water \( Q = mL \)

How much steam @ 100 °C does it take to melt 1kg of ice at -30 °C?

- How much energy is needed to raise the ices to 0 °C
- How much energy is needed to melt 1kg of ice?
- How much energy is given up by the steam?
- What happens to the steam that is melting the ice?

\[
L_f = 334 \text{ kJ/kg} \\
L_v = 2256 \text{ kJ/kg} \\
c_{\text{ice}} = 2090 \text{ J/kg} \cdot ^\circ C \\
c_{\text{water}} = 4186 \text{ J/kg} \cdot ^\circ C
\]
Phase Change: Water \[ Q = mL \]

How much steam @ 100 °C does it take to melt 1kg of ice at -30 °C?

How much energy is needed to raise the ices to 0 °C

\[ Q_1 = 1kg(2090 \text{ J/kg } \cdot 0°C)(30 \text{ °C}) \]

\[ = 62700J \]

\[ L_f = 334 \text{ kJ/kg} \]
\[ L_v = 2256 \text{ kJ/kg} \]
\[ c_{ice} = 2090 \text{ J/kg } \cdot 0°C \]
\[ c_{water} = 4186 \text{ J/kg } \cdot 0°C \]
Phase Change: Water \( Q = mL \)

How much steam @ 100 °C does it take to melt 1kg of ice at -30 °C?

How much energy is needed to melt 1kg of ice?

\[
Q_2 = mL = 1\text{kg}(334 \text{ kJ/kg})
\]

\[
Q_2 = 334 \text{kJ}
\]

\[
L_f = 334 \text{ kJ/kg}
\]

\[
L_v = 2256 \text{ kJ/kg}
\]

\[
c_{ice} = 2090 \text{ J/kg \cdot } ^\circ\text{C}
\]

\[
c_{water} = 4186 \text{ J/kg \cdot } ^\circ\text{C}
\]

\[
Q_1 = 62700 \text{J}
\]

\[
Q_2 = 334 \text{kJ}
\]
Phase Change: Water \( Q = mL \)

How much steam @ 100 °C does it take to melt 1kg of ice at -30 °C?

• How much energy is given up by the steam?
• What happens to the steam that is melting the ice?

\[
L_f = 334 \text{ kJ/kg} \\
L_v = 2256 \text{ kJ/kg} \\
c_{\text{ice}} = 2090 \text{ J/kg} \cdot ^0\text{C} \\
c_{\text{water}} = 4186 \text{ J/kg} \cdot ^0\text{C} \\
Q_1 = 62700 \text{ J} \\
Q_2 = 334 \text{ kJ} \\
Q_{\text{total}} = 397 \text{ kJ}
\]
The First Law of Thermodynamics

\[ \Delta E_{\text{int}} = Q + W \]

• The First Law of Thermodynamics is a special case of the Law of Conservation of Energy
  – It takes into account changes in internal energy and energy transfers by heat and work

• Although \( Q \) and \( W \) each are dependent on the path, \( Q + W \) is independent of the path
Heat Transfer

- Heat flows from HOT to COLD
- Conduction (solids)
- Convection (liquids & gases)
- Radiation (solids, gases, plasma)
Conduction, Convection & Radiation

- Conduction (hot air)
- Convection around windows and doors (cold air)
- Radiation

©2001 Brooks/Cole - Thomson Learning
Heat Conduction

Energy transferred via molecular collision
Conduction

Heat energy is transferred in solids by collisions between free electrons and vibrating atoms.

- Good Conductors: Most Metals (free electrons!)
- Bad Conductors: Organic & Inert Materials
- Good Insulators: Air, Water, Wood
- Good Conductors are BAD Insulators
- & Visa Versa
The heat $Q$ conducted during a time $t$ through a material with a thermal conductivity $k$. $dT/dx$ is the Temperature Gradient.

$$P = kA \frac{dT}{dx}$$
Some Thermal Conductivities

<table>
<thead>
<tr>
<th>Substance</th>
<th>Thermal Conductivity (W/m \cdot ^\circ C)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Metals (at 25^\circ C)</strong></td>
<td></td>
</tr>
<tr>
<td>Aluminum</td>
<td>238</td>
</tr>
<tr>
<td>Copper</td>
<td>397</td>
</tr>
<tr>
<td>Gold</td>
<td>314</td>
</tr>
<tr>
<td>Iron</td>
<td>79.5</td>
</tr>
<tr>
<td>Lead</td>
<td>34.7</td>
</tr>
<tr>
<td>Silver</td>
<td>427</td>
</tr>
<tr>
<td><strong>Nonmetals (approximate values)</strong></td>
<td></td>
</tr>
<tr>
<td>Asbestos</td>
<td>0.08</td>
</tr>
<tr>
<td>Concrete</td>
<td>0.8</td>
</tr>
<tr>
<td>Diamond</td>
<td>2300</td>
</tr>
<tr>
<td>Glass</td>
<td>0.8</td>
</tr>
<tr>
<td>Ice</td>
<td>2</td>
</tr>
<tr>
<td>Rubber</td>
<td>0.2</td>
</tr>
<tr>
<td>Water</td>
<td>0.6</td>
</tr>
<tr>
<td>Wood</td>
<td>0.08</td>
</tr>
<tr>
<td><strong>Gases (at 20^\circ C)</strong></td>
<td></td>
</tr>
<tr>
<td>Air</td>
<td>0.0234</td>
</tr>
<tr>
<td>Helium</td>
<td>0.138</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>0.172</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>0.0234</td>
</tr>
<tr>
<td>Oxygen</td>
<td>0.0238</td>
</tr>
</tbody>
</table>
Temperature Gradient

The quantity \( |dT / dx| \) is called the temperature gradient.
Conduction Problem

A bar of gold is in thermal contact with a bar of silver of the same length and area as shown. One end of the compound bar is maintained at 80.0\(^\circ\)C while the opposite end is at 30.0\(^\circ\)C. When the energy transfer reaches steady state, what is the temperature at the junction? Ignore thermal expansion of the metals.

\[ \phi = kA \left( \frac{T_h - T_c}{L} \right) \]
In the same room, at the same temperature, the tile floor feels cooler than wood floor. How can they be the same temperature?
Hot Air rises, expands and cools, and then sinks back down causing convection currents that transport heat energy.

Hot air rises because fast moving molecules tend to migrate towards regions of least obstruction - UP - into regions of lesser density!

Rising air cools because a decrease in density reduces number of collisions & speeds decrease.

As the air cools, it becomes denser, sinking down, producing a convection current.
Uneven heating on the earth and over water cause convection currents in the atmosphere, resulting in *WINDS*.

Global wind patterns (Trade Winds, Jet Streams) are due to convection current from warmer regions (equator) to cooler regions (poles) plus rotation of Earth.

Convection Currents in the Ocean (Gulf Stream) transport energy throughout the oceans.

Air & Ocean Convection causes the WEATHER.
Convection between water and land causes the Winds.
Sea Breeze

Land heats up more quickly than water

Cold air begins to push inland creating a breeze off the ocean

Rising warm air cools and moves over the ocean to replace cold air that moved inland

Local time: 12:00 p.m.
Land 85°F, Ocean 65°F
High Pressure
Dry Warm Weather

Low Pressure
Stormy Weather

1) Descending Air

2) Clouds
Rising Air

High Pressure
Converging Winds
Low Pressure
Converging Winds
High Pressure
Global Circulation
Ocean Convection

Great ocean conveyor belt

Sea-to-air heat transfer

Gulf Stream

Atlantic Ocean

Indian Ocean

Warm shallow current

Cold and salty deep current

Electromagnetic Radiation is emitted and absorbed via atomic excitations. All objects absorb and emit EM waves.
Electromagnetic Radiation is emitted and absorbed via atomic excitations. All objects absorb and emit EM waves.
When an object is heated it will glow first in the infrared, then the visible. Most solid materials break down before they emit UV and higher frequency EM waves.

Frequency ~ Temperature

Long
- (a) Cool
- (b) Medium
- (c) Hot

Short

EM radiation intensity

6000 K (white hot)
4000 K
3000 K (red hot)

JV  Visible range (violet → red)

©2001 Brooks/Cole - Thomson Learning
Stefan’s Law

• \( P = \sigma A e T^4 \)
  
  – \( P \) is the rate of energy transfer, in Watts
  
  – \( \sigma = 5.6696 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4 \)
  
  – \( A \) is the surface area of the object
  
  – \( e \) is a constant called the emissivity
    • \( e \) varies from 0 to 1
    • The emissivity is also equal to the absorptivity
  
  – \( T \) is the temperature in Kelvins
A good absorber reflects little and appears Black
A good absorber is also a good emitter.
Radiant heat makes it impossible to stand close to a hot lava flow. Calculate the rate of heat loss by radiation from 1.00 m² of 1200°C fresh lava into 30.0°C surroundings, assuming lava’s emissivity is 1.

The net heat transfer by radiation is:

\[ P = e\sigma A(T_2^4 - T_1^4) \]

\[ P = e\sigma A(T_2^4 - T_1^4) = 1(5.67 \times 10^{-8} \text{ J/smK}^4)1\text{m}^2((303.15\text{K})^4 - (1473.15\text{K})^4) \]

\[ P = -266 \text{ kW} \]
Insulators

How do fur coats keep you warm?

Fur is filled with air. Convection currents are slow because the convection loops are so small.
Any two systems placed in thermal contact will have an exchange of heat energy until they reach the same temperature.

If the systems are in thermal equilibrium then no net changes take place.
Why is winter cold and summer hot?
Intensity: The Radiation Power, $P$, passing through an area, $A$.

$$ I = \frac{P}{4\pi r^2} \left[ \frac{W}{m^2} \right] $$
Why are cloudy nights warmer than cold nights?
The heating effect of a medium such as glass or the Earth’s atmosphere that is transparent to short wavelengths but opaque to longer wavelengths: Short get in, longer are trapped!
Solar Radiation
Global Warming
CO$_2$ & Temperature Change
The Arctic Climate Impact Assessment, a study commissioned four years ago by the United States and the seven other countries with Arctic territory, projects that rising global concentrations of heat-trapping emissions will drive up temperatures particularly quickly at high latitudes.
CO$_2$ & Temperature Change
RISING SEAS One of the most important consequences of Arctic warming will be increased flows of meltwater and icebergs from glaciers and ice sheets, and thus an accelerated rise in sea levels.
Caught between rising seas on one side and expanding shrub-filled zones to the south, tundra ecosystems around the Arctic will likely shrink to their smallest extent in at least 21,000 years, the scientists concluded. This could reduce breeding areas for many tundra-dwelling bird species and grazing lands for caribou and other mammals.
1 Meter Rise In Florida
"Thirty years ago, there was no river here. If you come back here in another 30 years, one thing is for sure: There will definitely be no more ice here."

-Dr. Yao Tandong,
Institute of Tibetan Plateau Research
Global Glacial Ice Melting

On Kilimanjaro in Kenya, an 11,700-year-old ice cap that measured 4.3 square miles in 1912 had shrunk to 0.94 square miles in 2000, and is projected to disappear altogether in about 15 years. Melting of glaciers in Patagonia has doubled in recent years.
In Peru, the Quelccaya ice cap retreated a rate of more than 600 feet a year from 2000 to 2002 - up from just 15 feet a year in the 1960's and 70's - leaving a vast 80-foot-deep lake where none had existed when his studies began.
How will a decrease in Salinity change the ocean circulation?
Unless we change our direction, we are likely to end up where we are headed.