Chapter 10
Temperature and Heat
What are temperature and heat?

Are they the same?

What causes heat?
What Is Temperature?

How do we measure temperature?
What are we actually measuring?
Temperature and Its Measurement

- If two objects are in contact with one another long enough, the two objects have the same temperature.
- This begins to define *temperature*, by defining when two objects have the same temperature.
  - When the physical properties are no longer changing, the objects are said to be in *thermal equilibrium*.
  - Two or more objects in thermal equilibrium have the same temperature.
  - This is the *zeroth law of thermodynamics*.

*When the physical properties are no longer changing, the objects are said to be in thermal equilibrium.*

*This is the zeroth law of thermodynamics.*
The first widely used temperature scale was devised by Gabriel Fahrenheit.

Another widely used scale was devised by Anders Celsius.

The Celsius degree is larger than the Fahrenheit degree: the ratio of Fahrenheit degrees to Celsius degrees is 180/100, or 9/5.

\[
T_C = \frac{5}{9} (T_F - 32)
\]

\[
T_F = \frac{9}{5} T_C + 32
\]

They are both equal at -40°.
The zero point on the Fahrenheit scale was based on the temperature of a mixture of salt and ice in a saturated salt solution.

The zero point on the Celsius scale is the freezing point of water.

Both scales go below zero.

Is there such a thing as absolute zero?
What is absolute zero?

- If the volume of a gas is kept constant while the temperature is increased, the pressure will increase.
- This can be used as a means of measuring temperature.
- A constant-volume gas thermometer allows the pressure to change with temperature while the volume is held constant.
- The difference in height of the two mercury columns is proportional to the pressure.
- We can then plot the pressure of a gas as a function of the temperature.
- The curves for different gases or amounts are all straight lines.
- When these lines are extended backward to zero pressure, they all intersect at the same temperature, \(-273.2^\circ C\).
- Since negative pressure has no meaning, this suggests that the temperature can never get lower than \(-273.2^\circ C\), or \(0\) \(K\) (kelvin).

\[ T_K = T_C + 273.2 \]
- Can anything ever get **colder** than 0 K?
  - No.
- Can absolute zero ever be reached?
  - No.

\[ T_K = T_C + 273.2 \]
Heat and Specific Heat Capacity

- What happens when objects or fluids at different temperatures come in contact with one another?
  - The colder object gets hotter, and the hotter object gets colder, until they both reach the same temperature.
  - What is it that flows between the objects to account for this?

- We use the term *heat* for this quantity.
Heat and Specific Heat Capacity

- Heat flow is a form of energy transfer between objects.
- One-hundred grams of room-temperature water is more effective than 100 grams of room-temperature steel shot in cooling a hot cup of water.

- Steel has a lower specific heat capacity than water.
- The specific heat capacity of a material is the relative amount of heat needed to raise its temperature.
The **specific heat capacity** of a material is the quantity of heat needed to change a unit mass of the material by a unit amount in temperature.

- For example, to change 1 gram by 1 Celsius degree.
- It is a property of the material, determined by experiment.
- The specific heat capacity of water is 1 cal/g·C°: it takes 1 **calorie** of heat to raise the temperature of 1 gram of water by 1°C.

We can then calculate how much heat must be absorbed by a material to change its temperature by a given amount:

\[ Q = mc\Delta T \]

where

- \( Q \) = quantity of heat
- \( m \) = mass
- \( c \) = specific heat capacity
- \( \Delta T \) = change in temperature
- When two objects at different temperatures are placed in contact, heat will flow from the object with the higher temperature to the object with the lower temperature.
- Heat added increases temperature, and heat removed decreases temperature.
- Heat and temperature are not the same.
- Temperature is a quantity that tells us which direction the heat will flow.
When an object goes through a change of phase or state, heat is added or removed without changing the temperature. Instead, the state of matter changes: solid to liquid, for example.

The amount of heat needed per unit mass to produce a phase change is called the latent heat.

- The latent heat of fusion of water corresponds to the amount of heat needed to melt one gram of ice.
- The latent heat of vaporization of water corresponds to the amount of heat needed to turn one gram of water into steam.
If the specific heat capacity of ice is 0.5 cal/g·C°, how much heat would have to be added to 200 g of ice, initially at a temperature of -10°C, to raise the ice to the melting point?

\[ Q = mc\Delta T \]

\[ \begin{align*}
  m &= 200 \text{ g} \\
  c &= 0.5 \text{ cal/g·C°} \\
  T &= -10 \text{°C} \\
  \Delta T &= 10 \text{°C} \\
  Q &= (200 \text{ g})(0.5 \text{ cal/g·C°})(10 \text{°C}) \\
  &= 1,000 \text{ cal} \\
 \end{align*} \]

(heat required to raise the temperature)

a) 1,000 cal 
b) 2,000 cal 
c) 4,000 cal 
d) 0 cal
If the specific heat capacity of ice is 0.5 cal/g·°C, how much heat would have to be added to 200 g of ice, initially at a temperature of -10°C, to completely melt the ice?

\[ L_f = 80 \text{ cal/g} \quad Q = mL_f \]

\[ \begin{align*}
Q &= (200 \text{ g})(80 \text{ cal/g}) \\
 &= 16,000 \text{ cal}
\end{align*} \]

(heat required to melt the ice)

Total heat required to raise the ice to 0 °C and then to melt the ice is:

\[ 1,000 \text{ cal} + 16,000 \text{ cal} = 17,000 \text{ cal} = 17 \text{ kcal} \]
Rumford noticed that cannon barrels became hot during drilling.

Joule performed a series of experiments showing that mechanical work could raise the temperature of a system.

In one such experiment, a falling mass turns a paddle in an insulated beaker of water, producing an increase in temperature.
Joule’s experiments led to Kelvin’s statement of the first law of thermodynamics.

- Both work and heat represent transfers of energy into or out of a system.
- If energy is added to a system either as work or heat, the internal energy of the system increases accordingly.

The increase in the internal energy of a system is equal to the amount of heat added to a system minus the amount of work done by the system. \( \Delta U = Q - W \)
Joule’s Experiment and the First Law of Thermodynamics

- This introduced the concept of the *internal energy* of a system.
  - An increase in internal energy may show up as an increase in temperature, or as a change in phase, or any other increase in the kinetic and/or potential energy of the atoms or molecules making up the system.
  - Internal energy is a property of the system uniquely determined by the state of the system.
- The internal energy of the system is the sum of the kinetic and potential energies of the atoms and molecules making up the system.
A hot plate is used to transfer 400 cal of heat to a beaker containing ice and water; 500 J of work are also done on the contents of the beaker by stirring. What is the increase in internal energy of the ice-water mixture?

\[ W = -500 \text{ J} \]
\[ Q = 400 \text{ cal} = (400 \text{ cal})(4.19 \text{ J/cal}) = 1680 \text{ J} \]

\[ \Delta U = Q - W = 1680 \text{ J} - (-500 \text{ J}) = 2180 \text{ J} \]

- a) 900 J
- b) 1180 J
- c) 1680 J
- d) 2180 J
A hot plate is used to transfer 400 cal of heat to a beaker containing ice and water; 500 J of work are also done on the contents of the beaker by stirring. How much ice melts in this process?

\[ L_f = 80 \text{ cal/g} \]
\[ = (80 \text{ cal/g})(4.19 \text{ J/cal}) \]
\[ = 335 \text{ J/g} \]

\[ \Delta U = mL_f \]
\[ m = \Delta U / L_f \]
\[ = (2180 \text{ J}) / (335 \text{ J/g}) \]
\[ = 6.5 \text{ g} \]
Gas Behavior and The First Law

Consider a gas in a cylinder with a movable piston. If the piston is pushed inward by an external force, work is done on the gas, adding energy to the system.

- The force exerted on the piston by the gas equals the pressure of the gas times the area of the piston: \( F = PA \)
- The work done equals the force exerted by the piston times the distance the piston moves: \( W = Fd = (PA)d = P\Delta V \)
If the gas is being compressed, the change in volume is negative, and the work done is negative.
- Work done on the system is negative.
- Negative work increases the energy of the system.

If the gas is expanding, positive work is done by the gas on its surroundings, and the internal energy of the gas decreases.
An **ideal gas** is a gas for which the forces between atoms are small enough to be ignored.

- For an ideal gas, absolute temperature is directly related to the average kinetic energy of the molecules of the system.
- Most gases behave approximately as ideal gases.

If the process is **adiabatic**, no heat flows into or out of the gas.

Even though no heat is added, the temperature of a gas will increase in an adiabatic compression, since the internal energy increases.
In an *isothermal* process, the temperature does not change.

- The internal energy must be constant.
- The change in internal energy, $\Delta U$, is zero.
- If an amount of heat $Q$ is added to the gas, an equal amount of work $W$ will be done by the gas on its surroundings, from $\Delta U = Q - W$.

In an *isobaric* process, the pressure of the gas remains constant.

- The internal energy increases as the gas is heated, and so does the temperature.
- The gas also expands, removing some of the internal energy.
- Experiments determined that the pressure, volume, and absolute temperature of an ideal gas are related by the *equation of state*:

$$PV = NkT$$

where $N$ is the number of molecules and $k$ is Boltzmann’s constant.
What process makes a hot-air balloon rise?

- When gas is heated in a hot-air balloon, the pressure, not the temperature, remains constant.
- The gas undergoes an isobaric expansion.
- Since the gas has expanded, the density has decreased.
- The balloon experiences a buoyant force because the gas inside the balloon is less dense than the surrounding atmosphere.
The Flow of Heat

- There are three basic processes for heat flow:
  - Conduction
  - Convection
  - Radiation
In conduction, heat flows through a material when objects at different temperatures are placed in contact with one another.
(Conduction Contd.)

- The rate of heat flow depends on the temperature difference between the objects.
- It also depends on the *thermal conductivity* of the materials, a measure of how well the materials conduct heat.
- For example, a metal block at room temperature will feel colder than a wood block of the exact same temperature.
- The metal block is a better thermal conductor, so heat flows more readily from your hand into the metal.
- Since contact with the metal cools your hand more rapidly, the metal feels colder.
In convection, heat is transferred by the motion of a fluid containing thermal energy.

- Convection is the main method of heating a house.
- It is also the main method heat is lost from buildings.
In **radiation**, heat energy is transferred by electromagnetic waves.

- The electromagnetic waves involved in the transfer of heat lie primarily in the infrared portion of the spectrum.
- Unlike conduction and convection, which both require a medium to travel through, radiation can take place across a vacuum.
- For example, the evacuated space in a thermos bottle.
- The radiation is reduced to a minimum by silvering the facing walls of the evacuated space.
What process makes a car’s interior heat up when parked in the sun?
Why are houses insulated with material in the walls instead of just empty space?
Why is this insulated material often foil-backed?
Is a light-colored roof or a dark-colored roof more energy efficient?
What heat-flow processes are involved in solar collectors?
What heat-flow processes are involved in the greenhouse effect?