Chapter 18
Temperature, Heat, and the First Law of Thermodynamics

Problems: 8, 11, 13, 17, 21, 27, 29, 37, 39, 41, 47, 51, 57
Thermodynamics – study and application of thermal energy

temperature – quantity that is related to the amount of thermal energy of a substance

Temperature is one of the 7 SI base quantities.
Temperature is measured in SI using the Kelvin scale.
Objects are in thermal equilibrium with each other, if they are at the same temperature.

Zeroth law of thermodynamics: If objects A and B are in thermal equilibrium with object C, then objects A and B are in thermal equilibrium with each other.

In the lab if we want to see if two objects are in thermal equilibrium we take their temperature. The third object is the thermometer.
Kelvin Scale

The Kelvin scale makes use of the triple point of water. At 273.16 K, water can coexist as a gas, liquid and a solid.

Thus the size of the kelvin is 1/273.16 of the difference between absolute zero and the triple point of water.
We will see later that temperature is related to kinetic energy. Because of this temperature does a lower limit. In Kelvin, the lower limit is 0 K. This temperature is known as absolute zero. This is when all motion is stopped. Record for lowest temperature produced in a lab is on the order of nK ($10^{-9}$ K)

The average temperature of the universe is about 3 K.
Celsius and Fahrenheit Scales

Celsius scale is a commonly used scale that is popular for scientific use. It is similar to the Kelvin scale except it is shifted by about $273^0$. 

$$T_c = T - 273.15 \quad (T \text{ is temperature in Kelvin})$$

Celsius scale is set up so that water freezes at $0^0C$ and boils at $100^0C$.

One degree change in Kelvin is the same as one degree change in Celsius.

If we are interested in changes in temperature the two give equivalent results.
Fahrenheit scale – used in the U.S.

$$T_F = \frac{9}{5}T_C + 32$$

Water freezes at 32°F and boils at 212°F.

At -40 degrees, the Celsius scale and Fahrenheit scales give the same temperature.
Thermal Expansion

As the temperature of a substance increases, the volume increases.

Atoms are separated from each other by some distance. As the temperature increases, this separation increases. Thus the whole object expands as temperature increases.

The object expands in all dimensions.

Important to consider when building structures such as bridges. Use thermal expansion joints to compensate for the changes in length.
Thermal Expansion

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

or

\[ L - L_0 = \alpha L_0 (T - T_0) \]

$L_0$ is the length when temperature is $T_0$

$\alpha =$ coefficient of linear expansion

Units of $\alpha$ are $\frac{1}{^\circ C}$
Example of a problem with thermal expansion

Pour hot water in a cold glass. \((\alpha=9\times10^{-6}/C^0)\)

The inside surface of the glass heats and expands.
The outside surface is cooler and expands less.
The glass may not withstand the difference in expansion and the glass breaks.

Pyrex glass \((\alpha=3.2\times10^{-6}/C^0)\) has a smaller coefficient of linear expansion. Thus the thermal stresses are reduced.
Volume expansion

As the temperature increases, the volume expands. If the coefficient of linear expansion is the same in all directions, then $\beta = 3\alpha$.

$\Delta V = \beta V_0 \Delta T$

$\beta = \text{coefficient of volume expansion}$

As global warming happens, the volume of water in the ocean increases. This contributes to rising sea levels.
Side Note about Water

Density of water does not consistently change with the temperature.

As the temperature decreases the volume decreases, UNTIL it drops to 4°C. The water expands as it is lowered to the freezing point. Thus ice is less dense that liquid water. Ice floats.
When two objects at different temperatures have different amounts of average thermal energy, they are not in thermal equilibrium. Thermal energy will be transferred from the warmer object to the colder object until they are in equilibrium. The transferred thermal energy is called heat.

Units of thermal energy (heat):
- calorie
- British-thermal unit
- joule
Thermal Energy

Thermal energy is related to the size/amount of an object.

Take a cool lake and a hot cup of coffee. The coffee has a higher temperature, since the coffee has a higher average thermal energy per molecule.

However, the lake can have more thermal energy, because it has so much more mass.

If you compared the cup of coffee to a single cup of lake water, the coffee would have more thermal energy.
The calorie is defined to be the amount of heat needed to raise the temperature of 1 gram of water from 14.5\textdegree C to 15.5\textdegree C.

1 cal = 4.1868 Joules

In nutrition, the calorie is the amount of energy that is in the food. Plus the calorie in nutrition is really a kilocalorie.
Heat Capacity

Heat capacity is the proportionality that relates the amount of heat needed to produce a change in temperature.

\[ Q = C \Delta T = C (T_f - T_i) \]

\( Q \) = heat
\( C \) = heat capacity

What we usually care about is the \textbf{specific heat capacity}. Heat capacity per unit mass.
Specific Heat

Remember the awkward definition for the calorie?

Using that definition, the specific heat capacity of water is 1 cal/(g\(^0\)C)

\[ Q = mc \Delta T = m \cdot c \cdot (T_f - T_i) \]
Water

• Water has a relatively large specific heat.
  (see table on 485)

• $c_w$ is almost 5 times as large as $c_{Al}$

• It takes almost 5 times as much energy to change the temperature of a mass of water than to change an equal mass of aluminum by the same temperature difference.
Calorimetry

Used to measure the unknown specific heat of a material by placing it in thermal equilibrium with a material of known specific heat and measuring the temperature changes.

This works since the heat that leaves one material, goes into the other material.

For example: mixing hot aluminum with colder water. By finding the temperature changes, the specific heat of aluminum can be found.
Example

200 g of iron \( (c_{Fe} = 0.107 \text{ cal/(g}^0\text{C}) \) is dropped into 100 g of water. The iron is initially at 80 degrees Celsius, while the water starts at 20 degrees Celsius. What will be the final temperature?

Heat leaving iron = Heat entering water

\[ m_{Fe} c_{Fe} \Delta T_{Fe} = m_{w} c_{w} \Delta T_{w} \]
\[ m_{Fe} c_{Fe} \Delta T_{Fe} = m_{w} c_{w} \Delta T_{w} \]

\[ \Delta T_{FE} = 80^0C - T_f \]
\[ \Delta T_{w} = T_f - 20^0C \]

We want to find \( T_f \).

\[
(200 \text{ g})(0.107 \frac{\text{cal}}{\text{g} \cdot \text{C}})(80^0C - T_f) = (100 \text{ g})(1 \frac{\text{cal}}{\text{g} \cdot \text{C}})(T_f - 20^0C)
\]

Do algebra and solve for \( T_f \)

\[ T_f = 30.6^0C \]
Heat of Transformation

Many substances exist in more than one phase of matter.

Water can be a liquid, solid, or gas.

Even a metal such as lead can be a liquid or a gas.

Sometimes the heat added to a substance changes the phase of matter of the substance.
Most common examples: melting ice, boiling water
Vaporization

Vaporization is when a substance changes from a liquid to a gas.

The reverse of this process is condensation.

When a substance goes from liquid to gas, it must absorb heat.
When it condenses, it must release heat.
Fusion

Fusion occurs when a liquid freezes. The substance goes from liquid to solid.

Melting is the opposite process.

In order to melt ice you need to add heat. To freeze water, the water must lose heat.
Amount of heat required for fusion or vaporization.
\[ Q = Lm \]

L is the latent heat of fusion/vaporization.
   - L has units of energy/mass
m is the mass.

Q is positive if the substance is melted or vaporized.
Q is negative if the substance is fused (frozen) or condensed.

See table on pg. 487 for values of \( L_f \) and \( L_v \) for different materials and the temperatures at which the phase change occurs.
Water is used as a common example. Melting point 273 K or 0°C
$L_f$ 333 kJ/kg or 79.7 cal/gram

Boiling point 373 K of 100°C
$L_v$ 2256 kJ/kg or 540 cal/gram

From the table on pg. 487 you can see that ‘gases’ (hydrogen, oxygen) have boiling points well below room temperature.

‘Solids’ such as lead, silver can exist as liquids and gases but only at very high temperatures.
How about finding the energy to turn cold ice (T<0°C) to steam where T>100°C)
You have to break the problem into 5 parts.

1\(^{st}\) raise the temperature of the ice to 0°C.
2\(^{nd}\) melt ice
3\(^{rd}\) raise temperature of the water to 100°C
4\(^{th}\) change the liquid water to steam (vaporize)
5\(^{th}\) heat the steam to the final temperature

By adding the 5 heats required to do these steps you can find the total energy needed to do the complete process.
Heat and Work

Look at what happens when we compress gases.

We have a gas in a cylinder and can push down with a piston. $ds$ is the distance the piston moves.

\[ dW = F \cdot ds \]
\[ dW = (pA) \, ds \]
\[ dW = p \, (A \, ds) \]
\[ dW = p \, dV \]

\[ W = \int_{V_i}^{V_f} pdV \]
P-V diagram

The P-V diagram is a graph that shows the relationship between the pressure and the volume of the gas.

\[ W = \int_{v_i}^{v_f} p\,dv \]

The work is the area under the curve.

If the gas expands the work done by the gas is positive. The gas does work on its surroundings.

If the gas is compressed, the work done by the gas is negative. The surroundings do work on the gas.

see figs. 18-14
1\textsuperscript{st} Law of Thermodynamics

The 1\textsuperscript{st} Law of Thermodynamics states that the change in internal energy of a gas is equal to the difference between the heat added and the work done by the system.

\[ \Delta E_{\text{int}} = Q - W \]
\[ dE_{\text{int}} = dQ - dW \]

The internal energy increases if heat is added and it decreases if the gas does work.

This is really just a fancy way of saying conservation of energy.
Special cases of the 1\textsuperscript{st} Law of TD

Adiabatic process – process where there is not heat transfer.

This can occur is the process is so fast, or the system is insulated so that there is no heat transfer between system and environment.

\[ \Delta E_{\text{int}} = Q - W = 0 - W \]
\[ \Delta E_{\text{int}} = -W \]

If you allowed a gas that is insulated to expand, that is an adiabatic process.
Constant volume process (Isometric or isochoric)

\[ \Delta V = \text{zero, so no work is done} \]
\[ \Delta E_{\text{int}} = Q \]

This can happen if the gas is in a rigid container.
Cyclic process

If we look at a P-V diagram, we see the pressure and volume of the gas. Knowing these two variables describes the state of the gas. This includes the amount of internal energy. The change in internal energy of a process only depends on the endpoints.

If a process occurs that the initial and final positions on the P-V diagram are the same, that is a cyclic process. An example is what’s going on the cylinder of a car engine.
Cyclic process

Since the initial and final pressures and volumes are the same, $\Delta E_{\text{int}} = 0$

Therefore $0 = Q - W$

$Q = W$

If you go around any closed loop on a P-V diagram, the $\Delta E_{\text{int}} = 0$

The work is the enclosed area. $W > 0$ if clockwise.
The work is negative if counterclockwise.
(draw pictures)
Free expansion.
Free expansion is what happens when a gas is allowed to expands very quickly.
Imagine a gas that is confined to the left hand side of the insulated chamber.

If the divider in between the gas and the vacuum region breaks, the gas expands quickly to fill the entire chamber.
It is insulated to there is no heat transfer \((Q = 0)\).
The gas expands into a vacuum so there is no pressure to resist so the gas does no work \((W = 0)\).
Therefore \(\Delta E_{\text{int}} = 0\).
Types of Energy Transfer

3 ways for thermal energy transfer to occur.

**Thermal conduction**, when two objects at different temperature are in physical contact with each other, energy will be transferred from the hotter object to the cooler object.

Rate of conduction depends on the thermal conductivity of the materials involved. For example, a pot holder has poor thermal conductivity, so it helps keep your hand from being burnt by a hot pot.
Conduction

Conduction rate = $P_{\text{cond}}$

$$P_{\text{cond}} = \frac{Q}{t} = kA\frac{T_H - T_C}{L}$$

A = area

L = thickness

k = thermal conductivity  (see table 18-6)

Units of $\frac{W}{mK}$

R-value thermal resistance

R = $L/k$
Convection

• Convection is the transfer of energy by the movement of a substance.

• Examples:
  – Air in a convection oven circulates, carrying the heat to different locations in the oven.
  – Water cooling an engine by forcing cold water to pass by hot engine parts.
  – Currents in the ocean.
  – Atmospheric convection
Energy from the Sun

• The Earth is not in contact with the sun, so it can’t be heated by conduction.
• There is no substance such as air in between to circulate the energy from the sun to the Earth, so the Earth is not heated by convection.
• How is the Earth heated?
Radiation

**Radiation** - objects radiate energy in the form of electromagnetic waves due to the thermal vibrations of their molecules.

Those electromagnetic waves travel through space and deliver the energy.

Another example is the heat you feel when you put your hands near a hot light bulb. Atoms on the bulb are vibrating, which produces the waves that transit the energy.
Radiation

\[ P_{\text{rad}} = \sigma \varepsilon AT^4 \]

A = surface area
Stefan-Boltzmann constant
\( \varepsilon \) – emissivity goes from 0 – 1
\[ \sigma = 5.6704 \times 10^{-8} \frac{W}{m^2 K^4} \]

Emissivity has to do with how well the object absorbs thermal radiation.
\( \varepsilon = 1 \) is a blackbody radiation which is an upper limit. (perfect radiator)

Use equation with Kelvin scale so that an object at absolute zero radiates no energy.
Use similar equation to find rate that object absorbs energy via thermal radiation from its environment.

\[ P_{abs} = \sigma \varepsilon A T^4_{env} \]

A perfect black-body radiator will absorb all of the radiated energy it intercepts. Thus it is an idealization.

Objects radiate heat while they absorb heat so:

\[ P_{net} = P_{abs} - P_{rad} = \sigma \varepsilon A (T^4_{env} - T^4) \]
Problems 10, 26, 38, 48