Temperature, Thermal Equilibrium and Thermometers

In everyday language, many of us use the terms temperature and heat interchangeably. But in physics, these terms have very different meanings. Think about this...

- When you touch a piece of metal and a piece of wood both resting in front of you, which feels warmer?
- When do you feel warmer when the air around you is 90°F and dry or when it is 90°F and very humid?
- In both cases the temperatures of what you are feeling is the same. Why then are you feeling a difference?

In this unit, we will learn about temperature, heat, and the laws of thermodynamics that relate heat, mechanical work and other forms of energy.
Thermometers and Thermal Equilibrium

To measure temperature of a substance, we need...

1. A measuring device (Thermometer) that changes visibly and is calibrated to a scale. Thermometers relate the change in a physical property of substance to temperature. Examples include:
   - The change of volume of a gas or liquid
   - The change in length of a metal strip or wire
   - The light or infrared radiation emitted by an object

2. To bring the Thermometer into contact with the substance
   - When the thermometer has settled on a value, we say that the thermometer and the substance are in Thermal Equilibrium.

Temperature Scales

Recall that temperature can be defined as...

- a measure of the warmth or coldness of an object or substance with reference to some standardized numerical measures or scales

Three common scales are:

- Fahrenheit (°F) - used mainly in the USA
- Celsius (°C) - used in most of the world
- Kelvin (K) - used in the physical sciences
  - also known as the Absolute Temperature Scale

Temperature Conversions

<table>
<thead>
<tr>
<th>Celsius # Fahrenheit</th>
</tr>
</thead>
<tbody>
<tr>
<td>( T_F = 1.8T_C + 32 )</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Celsius # Kelvin</th>
</tr>
</thead>
<tbody>
<tr>
<td>( T_K = T_C + 273.16 )</td>
</tr>
</tbody>
</table>

1. Which temperature scale does not have negative values?
   - Fahrenheit
   - Celsius
   - Kelvin
   - All the above
   - None of the above
1. Which temperature scale does not have negative values?
   - Fahrenheit
   - Celsius
   - Kelvin
   - All the above
   - None of the above

   Answer: C Kelvin

2. Water freezes at 32°F. What temperature would this be on the Celsius scale?
   - 32 °C
   - 0° C
   - 25° C
   - 212°C
   - 100° C

   Answer: B 0°C

3. Water boils at 100° C. What temperature would this be on the Fahrenheit scale?
   - 32 °F
   - 100° F
   - 0° F
   - 212° F
   - 180° F

   Answer: D 212°F

4. “Room temperature” is often taken to be 68 °F; what is this on the Celsius scale?
   - 34 ° C
   - 37.78° C
   - 5.78° C
   - 20° C
   - 52° C

   Answer:
4 "Room temperature" is often taken to be 68 °F; what is this on the Celsius scale?

- 34 °C
- 37.78 °C
- 5.78 °C
- 20 °C
- 52 °C

**D 20°C**

use \( T_C = \frac{(T_F - 32)}{1.8} \)

5 The coldest temperature recorded on earth was −89.2 °C at the Soviet Vostok Station in Antarctica, on July 21, 1983. What would a Fahrenheit scale thermometer have measured?

\[ T_F = 1.8T_C + 32 \]

\[ = 1.8(89.2) + 32 \]

\[ = -128.6 °F \]

6 The coldest temperature recorded on earth was −89.2 °C at the Soviet Vostok Station in Antarctica, on July 21, 1983. What would a Kelvin scale thermometer have measured?

\[ T_K = T_C + 273 \]

\[ = -89.2 + 273 \]

\[ = 183.8 K \]

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**Thermal Equilibrium: The Zeroth Law of Thermodynamics**

Two objects placed in thermal contact will eventually come to the same temperature. When they do, we say they are in thermal equilibrium.

The zeroth law of thermodynamics says that if two objects are each in equilibrium with a third object, they are also in thermal equilibrium with each other.

That is if \( T_1 = T_3 \) and \( T_2 = T_3 \) then \( T_1 = T_2 \)
Thermal Conductors and Insulators

**Conductors** - materials that allow heat to flow easily (metals)

**Insulators** - materials that slow or block heat flow (wood, plastic, fiberglass)

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7. Three objects A, B, and C initially have different temperatures \( T_A > T_B > T_C \). Objects A and B are separated by an insulating plate but they are in contact with object C through a conducting plate. Which of the following is true when objects A and B reach thermal equilibrium with object C?

- A. The temperatures of all three objects do not change
- B. Object A has a higher temperature than Object B and Object C
- C. Object C has a higher temperature than Object A and Object B
- D. Object B has a higher temperature than Object A and Object C
- E. All three objects have the same temperature

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Thermal Expansion

- Most materials expand when their temperatures increase.
- Liquids expand in a thermometer.
- A tight metal jar lid can be loosened by running it in hot water.
- These are examples of thermal expansion.

We consider two types of thermal expansion:
- Linear
- Volume

---

Linear Expansion

Suppose a rod composed of some substance has a length \( L_0 \) at an initial temperature of \( T_0 \).

If the temperature is changed by \( \Delta T \), the length changes by \( \Delta L \).

If \( \Delta T \) is not too large, \( \Delta L \) is directly proportional to \( \Delta T \).

The change in length is:

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

where \( \alpha \) is the coefficient of linear expansion.
**Linear Expansion**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Coefficient of Linear Expansion $\alpha$ ($\times 10^{-6}$ / °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>23.1</td>
</tr>
<tr>
<td>Diamond</td>
<td>1</td>
</tr>
<tr>
<td>Copper</td>
<td>17</td>
</tr>
<tr>
<td>Glass</td>
<td>8.5</td>
</tr>
<tr>
<td>Iron</td>
<td>11.8</td>
</tr>
<tr>
<td>Gold</td>
<td>14</td>
</tr>
<tr>
<td>Steel</td>
<td>13.2</td>
</tr>
<tr>
<td>Ice</td>
<td>51</td>
</tr>
</tbody>
</table>

8 A steel rod measures 10 meters at 0°C. Given that the coefficient of linear expansion of steel is $1.2 \times 10^{-5}$ per °C, what will the rod measure at 75°C?

\[
L = L_0 + \Delta L = L_0 (1 + \alpha \Delta T) = 100(1 + 0.000012 \times 75) = 100.09\text{ m}
\]

9 A simple pendulum is made of a steel string supporting a brass sphere. The temperature in a room with the pendulum is increased from 15°C to 30°C. Which of the following is true about the period of oscillations?

- the period doubles
- the period does not change
- the period slightly increases
- the period slightly decreases
- the period increases by $\sqrt{2}$

**Volume Expansion**

Suppose a volume of some substance (gas, liquid or solid) has a length $V_0$ at an initial temperature of $T_0$.

If the temperature is changed by $\Delta T$, the volume changes by $\Delta V$.

If $\Delta T$ is not too large, $\Delta V$ is directly proportional to $\Delta T$.

The change in volume is

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

where $\beta$ is the coefficient of volume expansion.
### Volume Expansion

<table>
<thead>
<tr>
<th>Substance</th>
<th>Coefficient of Volume Expansion $\beta$ ($\times 10^{-6}/^{\circ}\text{C}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gasoline</td>
<td>950</td>
</tr>
<tr>
<td>Glycerine</td>
<td>485</td>
</tr>
<tr>
<td>Water</td>
<td>207</td>
</tr>
<tr>
<td>Aluminum</td>
<td>69</td>
</tr>
<tr>
<td>Diamond</td>
<td>3</td>
</tr>
<tr>
<td>Glass</td>
<td>9.9</td>
</tr>
<tr>
<td>Steel</td>
<td>32.4</td>
</tr>
<tr>
<td>Gold</td>
<td>42</td>
</tr>
</tbody>
</table>

### Volume Expansion of Water

Water above 4°C water expands when heated. BUT in the temperature range from 0°C to 4°C, the volume decreases as temperature increases. This means that 4°C water is more dense than 0°C water. Hence water has its greatest density at 4°C. 4°C water will sink to the bottom of a lake, so ice will form in the 0°C water floating on top. This protects life in bodies of freshwater since the water on the bottom will be at worst 4°C, not freezing!

### Heat and Temperature Change

When you pour hot water into a cold cup, the water cools down and the cup warms up as they approach thermal equilibrium. The reason for these temperature changes is that...

- Energy flows from the the higher temperature object to the lower temperature object. This flow of energy is called HEAT.

$Q$ is the symbol for HEAT

- Heat and Temperature are Different
  - Temperature is a quantitative measure of an object's hotness or coldness.
  - Heat is the energy that moves from one object to another because of a temperature difference.

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Units of Heat

Because heat is a form of energy...

the SI unit for heat is the **joule (J)**

Units of Heat

Other common units are:

- **calorie (cal)**
  - the amount of heat required to raise the temperature of 1 gram of water by 1°C.
  - 1 cal = 4.186 J

- **kcal or Calorie (food calorie)**
  - 1 kcal = 1000 cal = 4186 J

- **BTU (British Thermal Unit)**
  - the quantity of heat required to raise the temperature of 1 pound of water by 1°F
  - 1 BTU = 252 cal = 1055 J

Specific Heat

Consider the cup of water shown below.

If it takes an amount of heat Q to raise the temperature of the water to 20°C (a change of temperature $\Delta T = 10°C$)

It will take 2Q to raise its temperature 30°C (by $\Delta T = 20°C$)

Heat required is directly proportional to the change in temperature

$$Q \sim \Delta T$$

Specific Heat

Consider the cups below: #1 contains half the mass of water as #2.

If it takes an amount of heat Q to raise the temperature of the Cup #1 by $\Delta T$.

It will take 2Q to raise the temperature of Cup #2 by the same $\Delta T$.

Heat required is directly proportional to the change in mass m as well as $\Delta T$

$$Q \sim m\Delta T$$

Specific Heat

Consider the objects below: a cup containing a mass m of water and a piece of copper of mass m

Experiment tells us that it takes less heat to raise the temperature of the copper than the water by the same amount

Heat required depends on the nature of each substance. The quantity that adjusts for substance is call the **Specific Heat** (c).

So...

$$Q = mc\Delta T$$

Units for Specific Heat

$$Q = mc\Delta T$$

rearrange this for c:

$$c = \frac{Q}{m\Delta T}$$

so it follows that the units for specific heat are

$$J \left/ kg \degree C \right.$$
### Specific Heat

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat (J/kg°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>water (H₂O)</td>
<td>4186</td>
</tr>
<tr>
<td>ethylene glycol (anti-freeze)</td>
<td>2386</td>
</tr>
<tr>
<td>ice (H₂O)</td>
<td>2093</td>
</tr>
<tr>
<td>aluminum (Al)</td>
<td>837</td>
</tr>
<tr>
<td>copper (Cu)</td>
<td>419</td>
</tr>
<tr>
<td>gold (Au)</td>
<td>126</td>
</tr>
</tbody>
</table>

Notice that water has the highest heat capacity in the table.

**Challenge:** Can you find a material with a higher heat capacity?  
**Note:** Metals have much lower heat capacities than water, so...

> It takes less heat to raise the temperature of a mass of metal by ΔT than it takes to raise the same mass of water by ΔT.

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### Question 11

How much heat is required to raise the temperature of 0.5 kg of aluminum (c = 837 J/kg°C) from 15°C to 40°C?

**Answer**

\[ Q = mc\Delta T \]
\[ = (0.5)(837)(40-15) \]
\[ = 10,462 \text{ J} \]

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### Question 12

It takes 2 minutes to raise the temperature of 1 liter of water by 50°C with a hot plate. How much time would it take to raise the temperature of 2 liters of water 50°C using the same hot plate?

- A ½ minute
- B 1 minute
- C 2 minutes
- D 4 minutes

**Answer**

\[ Q_{\text{original}} = m c \Delta T \]
\[ Q_{\text{new}} = 2m c \Delta T \]
\[ Q_{\text{new}} = 2Q_{\text{original}} \]

same hotplate ⇨ same power (energy/time)

2Q takes twice the time

---

### Question 13

The ocean temperature doesn't change drastically because of

- A Water is a good heat conductor
- B Water has a high boiling point
- C Water has a high specific heat
- D Water has a low melting temperature
13. The ocean temperature doesn't change drastically because of
   A. Water is a good heat conductor
   B. Water has a high boiling point
   C. Water has a high specific heat
   D. Water has a low melting temperature

   **Answer:** C
   Water has a high specific heat

14. A solid copper ball, a solid silver ball and a solid aluminum ball, all having the same mass and at room temperature, are placed in a 300°C oven at the same time. Which of the three will increase in temperature fastest? (hint: look up the specific heats in the table provided earlier)

   - the aluminum ball
   - the copper ball
   - the silver ball
   - they all increase temperature at the same rate

   **Answer:** C
   Silver has the lowest specific heat, so it takes the least amount of energy of the three materials to increase its temperature

15. A solid copper ball, a solid silver ball and a solid aluminum ball, all having the same mass and a temperature of 200°C are placed on a huge block of ice at 0°C. Which ball will melt the most ice?

   - The silver ball
   - The copper ball
   - The aluminum ball
   - They all melt the same amount of ice

   **Answer:** C
   Aluminum has the greatest heat capacity. Therefore has the most heat energy to give up per degree

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**Thermal Equilibrium: Heat Calculations**
Thermal Equilibrium and Conservation of Energy

When two objects (isolated from their surroundings) are brought in contact with one another, we know that heat will flow from the hotter object to the colder object until the objects reach thermal equilibrium.

Because there is no where else for the heat to go, and because heat is energy and energy is conserved...

Heat Lost from hotter object + Heat Gained by colder object = ZERO

\[ \sum Q = 0 \]

Heat Calculations

Example 1: Two equal masses of water at different temperatures are mixed together. What is the equilibrium temperature \( T_F \)?

\[ Q = mc\Delta T \]

\[ \Delta T = T_F - T_0 \]

Example 1: Two equal masses of water at different temperatures are mixed together. What is the equilibrium temperature \( T_F \)?

\[ \Sigma Q = 0 \]

\[ Q = mc\Delta T \]

\[ \Delta T = T_F - T_0 \]

16 A styrofoam cup containing 400 g of 60°C water is poured into another styrofoam cup containing 900 g of 15°C water. What is the temperature of the combination?

- 75.0°C
- 45.0°C
- 37.5°C
- 30.0°C
- 15.0°C

17 Two vials of mercury sit on a chemist's desk. One vial contains 80 g of mercury at 20°C while the other vial contains 20 g of mercury at 60°C. She then pours one vial into the other. What will the final temperature of the mercury be? (The specific heat of mercury is 0.14 J/g)

- 40 °C
- 67 °C
- 21 °C
- 28 °C
17 Two vials of mercury sit on a chemist's desk. One vial contains 80 g of mercury at 20°C while the other vial contains 20 g of mercury at 60°C. She then pours one vial into the other. What will the final temperature of the mercury be? (The specific heat of mercury is 0.14 J/g)

A 40 °C  
B 67 °C  
C 21 °C  
D 28 °C

\[ Q_1 + Q_2 = 0 \]
\[ Q = mc_1 \Delta T \]
\[ m_1(T_f - T_1) + m_2(T_f - T_2) = 0 \]
\[ 80(T_f - 20) + 20(T_f - 60) = 0 \]
\[ 100T_f - 2800 = 0 \]
\[ T_f = 28 \, ^\circ C \]

Calorimetry

Often, a calorimeter (heat measuring apparatus) is used to find the initial temperature, specific heat or other thermal properties of a substance.

A simple Calorimeter is composed of a container that is insulated from the outside environment (so practically no heat can enter or leave).

For our simple “experiments”, the calorimeter's inner cup may be made of metal (Aluminum). The calorimeter will be filled with a quantity of water. We will drop our sample substance from the top before inserting the stopper.

Example 2: A 0.4 kg Aluminum calorimeter is filled with 0.8 kg of water. The calorimeter and water come to thermal equilibrium at a temperature of \( T_0 = 20°C \). 0.1 kg of a material (U) with a specific heat of 628 J/(kg°C) and an initial temperature of 300°C is dropped into the calorimeter. Find the equilibrium temperature of the combination.

A simple Calorimeter is composed of a container that is insulated from the outside environment (so practically no heat can enter or leave).

For our simple “experiments”, the calorimeter's inner cup may be made of metal (Aluminum). The calorimeter will be filled with a quantity of water. We will drop our sample substance from the top before inserting the stopper.

\[ \sum Q = 0 \]
\[ Q_1 + Q_2 + Q_3 = 0 \]

substitute \( Q = mc \Delta T \) for each component:
\[ (mc\Delta T)_U + (mc\Delta T)_W + (mc\Delta T)_Al = 0 \]

substitute \( \Delta T = T_f - T_0 \) for each component:
\[ 0.1(628)(T_f - 300) + 0.8(4186)(T_f - 20) + 0.4(837)(T_f - 20) = 0 \]

solve for \( T_f \):
\[ 3746.4T_f = 92,511 \quad \text{or} \quad T_f = 24.69°C \]

Phase Transitions
**Phase Transitions**

The term **phase** refers to a specific state of matter, such as:
- solid
- liquid
- gas

When a substance undergoes a change from one state to another, it is called a **phase transition**.

**Phase transitions** (at a given pressure) take place at a **constant temperature**, are accompanied by **addition or removal of heat**, and may involve a change in volume (density). Examples include:
- water freezing
- ice melting
- water vaporizing
- steam condensing

**Latent Heat of Fusion**

We will focus on phase transitions in water.

Ice melts at 0°C. For 1 kg of 0°C ice to change completely to water at 0°C requires the addition of 334 kJ of heat.

Similarly to freeze 1 kg of 0°C water into 0°C ice requires the removal of 334 kJ of heat.

The heat added to change a solid to liquid -or- removed to change a liquid to a solid is called the **latent heat of fusion** \( L_f \).

For water: \( L_f = 334 \text{ kJ/kg} \)

**Latent Heat of Vaporization**

We will focus on phase transitions in water.

Water vaporizes (boils) at 100°C. For 1 kg of 100°C water to change completely to steam at 100°C requires the addition of 2260 kJ of heat.

Similarly to condense 1 kg of 100°C steam into 100°C water requires the removal of 2260 kJ of heat.

The heat added to change a liquid to gas -or- removed to change a gas to a liquid is called the **latent heat of vaporization** \( L_v \).

For water: \( L_v = 2260 \text{ kJ/kg} \)

**Phase Transitions of Water**

- H₂O transforms from ice to liquid water to steam as heat is added
- During phase transitions the T-Q graph is horizontal - T is constant while heat is added
- While in a single phase (ice or water or steam), the temperature rises as heat is added

**Phase Transitions and Internal Energy**

Notice that when a phase transition is taking place, the temperature remains constant.

Where is the energy going to or coming from?

The answer is that it is going into making or breaking bonds between the atoms (or molecules) of the material.

For instance when ice melts, the bonds that hold the water molecules in place are broken with the addition of energy.

When that energy is removed from water at 0°C, the molecules bond again, crystalizing into ice.

19. As water vapor condenses
   - A. The temperature increases
   - B. The temperature decreases
   - C. Energy is absorbed
   - D. Energy is released
   - E. None from the above
19. As water vapor condenses:
A. The temperature increases
B. The temperature decreases
C. Energy is absorbed
D. Energy is released
E. None from the above

Answer: D. Energy is released (The temperature remains constant)

20. Ice is placed in an empty container in a room where the air temperature is 20°C and allow to melt. While the ice is melting the temperature of the water is...

- less than 0°C
- 0°C
- room temperature
- greater than 0°C but less than room temperature

Answer: B. 0°C during phase transition, the temperature remains constant

21. How much heat must be provided to melt a 0.3 kg chunk of ice, then raise the temperature of the melt water to 40°C?

\[ Q_{\text{total}} = Q_{\text{to melt ice}} + Q_{\text{to heat water}} \]
\[ Q_{\text{total}} = mL_f + mc\Delta T \]
\[ = 0.3(334,000) + 0.3(4186)(40-0) \]
\[ = 150,432 \text{ J} \]

Calorimetry

**Example 3:** A 0.5 kg Aluminum calorimeter is filled with 1 kg of water. The calorimeter and water come to thermal equilibrium at a temperature of \( T_0 = 80°C \). A 0.1 kg ice cube (at 0°C) is dropped into the calorimeter. Find the equilibrium temperature of the combination.

\begin{align*}
\text{Ice (I)} & \quad \text{Aluminum can (Al)} & \quad \text{Water (w)} \\
m & = 0.1 \text{ kg} & m & = 0.5 \text{ kg} & m & = 1.3 \text{ kg} \\
L_f & = 334,000 \text{ J/kg} & T_i & = 80°C & T_i & = 80°C \\
c & = 837 \text{ J/(kg°C)} & c & = 4186 \text{ J/(kg°C)} & c & = 4186 \text{ J/(kg°C)}
\end{align*}

The ice melts and then the melt water absorbs heat, The calorimeter and the water it contains lose heat
\[ \sum Q = 0 \]
\[ Q_{\text{ice melting}} + Q_{\text{water heat}} + Q_{\text{water}} + Q_{\text{Al}} = 0 \]
\[ mL_f + (mc\Delta T)_w + (mc\Delta T)_w + (mc\Delta T)_w = 0 \]
\[ 0.1\cdot334,000 + 0.1\cdot(4186)(T_w-0) + 1\cdot4186(T_w-80) + 0.5(837)(T_w-80) = 0 \]
\[ 5023.1\cdot T_w = 334,960 \quad \text{or} \quad T_w = 66.68°C \]
Heat Transfer

Earlier we defined thermal insulators and conductors. Now we are going to examine the mechanisms of heat transfer and the rates of heat transfer.

Mechanisms of Heat Transfer
There are three mechanisms of heat transfer:
- conduction
- convection
- radiation

Conduction
As indicated earlier, temperature is directly related to the average kinetic energy of the molecules in a substance. In a warmer region of an object ($T_2$), the molecules have a higher average kinetic energy than in a cooler region ($T_1$).

Conduction is the flow of heat due to the transfer of kinetic energy in molecular collisions within the object.

The rate of heat transfer ($\Delta Q/\Delta t$) depends:
- Directly on the relative temperature at both ends ($T_2 - T_1$)
- If no difference - no heat transfer
- Directly on the cross-sectional area (A) available for collisions to occur
- Inversely as the length (L) along which the heat must pass
- Directly on the conductivity (k) - ease of heat transfer - of the material

$$\frac{\Delta Q}{\Delta t} = \frac{kA}{L} (T_2 - T_1)$$

22 An iron tube with a length of 1 m and a radius of 10 cm is heated from the bottom by a lighter. If the bottom of the tube has a temperature of 35 degrees Celsius and the top of the tube has a temperature of 20 degrees Celsius, what is the rate of heat transfer through the tube? ($k=80$)

- $37.69 \text{ W}$
- $376,991 \text{ W}$
- $84.26 \text{ W}$
- $842,661 \text{ W}$
23 Two objects of different temperatures are separated by a wall. If the thickness (L) of the wall is doubled, the rate of heat transfer due will be...

- A Doubled
- B Quadrupled
- C Unchanged
- D Cut to one-half
- E Cut to one-fourth

\[
\frac{\Delta Q}{\Delta t} = \frac{kA}{L} (T_f - T_i)
\]

If L is doubled, \(\Delta Q/\Delta t\) is halved.

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24 Convection can occur

- A Only in solids
- B Only in liquids
- C Only in gasses
- D Only in liquids and gasses
- E In solids, liquids, and gasses

Convection can only occur where the molecules of a substance are free to move translationally (gasses and liquids).

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Radiation

Radiation is energy transfer by electromagnetic waves. You have directly experienced it as the warming you feel from the sun or sitting close to a fire.

As you will learn later in the Electromagnetic Waves Unit,

- radiation includes radio waves, microwaves, visible light and more, not just the infrared radiation you feel
- does not matter to transfer heat - it can travel through a vacuum
Radiation

The rate of heat transfer by radiation is given by the Stefan-Boltzmann equation.

\[
\frac{\Delta Q}{\Delta t} = e \sigma A T^4
\]

where
- \( e \) - emissivity, is a number between 0 and 1 that increases with the darkness of the surface of an object
- \( \sigma \) - the Stefan-Boltzmann constant \( \sigma = 5.67 \times 10^{-8} \frac{W}{m^2 \cdot K^4} \)
- \( A \) - the surface area of the object

The rate at which an object radiates energy is proportional to the fourth power of the absolute temperature.

25. A block of ice has a temperature of 0 °C and a surface area of 1 m². What is the rate of heat transfer by radiation? (Ice has an emissivity of 0.97)

- A 0 W
- B 413 W
- C 305 W
- D 612 W

\[
\Delta Q / \Delta t = 0.97 (5.67 \times 10^{-8})(1)(273)^4
\]

\[
= 305 \text{ W}
\]

26. When the temperature of a heater is doubled, by what factor does the radiating power change?

- A 2
- B 4
- C 8
- D 16
- E 32

if \( T \) is doubled, \( \Delta Q / \Delta t \) will increase by \( 2^4 = 16 \)

27. Which of the following is responsible for raising the temperature of water in a pot placed on a hot stove?

- A Conduction
- B Convection
- C Radiation
- D Vaporization
- E Condensation
27 Which of the following is responsible for raising the temperature of water in a pot placed on a hot stove?

- A Conduction
- B Convection
- C Radiation
- D Vaporization
- E Condensation

Answer: A and B

The water in contact with the pot is heated by conduction. This heat is transmitted to the rest of the water in the pot by convection.

Gas Laws

Physical properties of gases include pressure (P), volume (V) and temperature (T).

The relationships between these quantities were studied by a progression of scientists:

Boyle (17th century)  Charles (18th century)

Boyle's Law

Robert Boyle performed a set of experiments measuring the volume (V) of gases as the pressure (P) was changed.

Boyle's Law...

"The pressure exerted on an ideal gas is inversely proportional to the volume it occupies (if the temperature and amount of gas remain unchanged within a closed system)"

\[ P \alpha \frac{1}{V} \]

or

\[ PV = \text{constant} \]

or

\[ P_1 V_1 = P_2 V_2 \]

Note Boyle's Law assumes an isothermal (constant T) process.

Different temperatures give different P-V diagrams.

28 Which of the following graphs represents the isothermal process?
28. Which of the following graphs represents the isothermal process?

- A
- B
- C
- D
- E

**Answer:** E

PV = nRT
so P ~ 1/V

29. A container filled with an ideal gas at pressure P is compressed to one-fourth of its volume while the temperature is kept constant. What is the new pressure in the gas relative to its original pressure P?

- A 2P
- B 4P
- C P
- D 1/2P
- E 1/4P

**Answer:** B

PV = nRT
so P ~ 1/V

30. Which of the following graphs represents the isobaric process?

- A
- B
- C
- D
- E

**Charles's Law**

Jacques Charles performed a set of experiments measuring the volume (V) of gases as the temperature (T) was changed.

Charles’s Law…

“When the pressure of an ideal gas is kept constant, the volume is directly proportional to the absolute (Kelvin) temperature”

Note this is an isobaric (constant P) process.

\[ \frac{V_1}{V_2} = \frac{T_1}{T_2} \]

Note Charles’s Law assumes an isobaric (constant P) process.

Different pressures give different V-T diagrams.
30. Which of the following graphs represents the isobaric process?

- **A**
- **B**
- **C**
- **D**
- **E**

**Answer:** **B**

isobaric means $P = \text{constant}$

31. An ideal gas is taken from one state at temperature $T_1 = 273\, \text{K}$ to another state at temperature $T_2 = 546\, \text{K}$ isobarically. What happens to the volume of the ideal gas?

- **A** It quadruples
- **B** It is cut to one-fourth
- **C** It doubles
- **D** It is cut to a half
- **E** It doesn’t change during the isobaric process

**Answer:** **C**

$PV = nRT$ \text{isobaric}\ P = \text{constant} \text{ so } V \sim T$ 

$T$ doubles, $V$ doubles

**Gay-Lussac's Law**

Joseph Louis Gay-Lussac formulated the last three of the gas laws.

“When the volume of an ideal gas is kept constant, the pressure is directly proportional to the absolute (Kelvin) temperature”

\[
\frac{P_1}{P_2} = \frac{T_1}{T_2}
\]

Note Gay-Lussac’s Law assumes an \textit{isochoric} (constant $V$) process.

Different pressures give different $P$-$T$ diagrams.

32. Which of the following graphs represents the isochoric process?

**Answer:** **C**

isochoric means $V = \text{constant}$
33 A sample of an ideal gas is enclosed in a container with rigid walls. The temperature of the gas is raised from 20°C to 60°C. What happens to the pressure of the gas?

- A It doubles
- B It quadruples
- C It triples
- D It is cut to one-third
- E It is slightly increased

Answer: E

PV=nRT
rigid walls ➔ V = constant so P~T
T increases from 293K to 333K
so P increases slightly

The Ideal (or Combined) Gas Law

The Boyle, Charles and Gay-Lussac Laws can be combined into a single general relationship among the pressure, volume, and temperature a fixed quantity of gas.

\[ PV \propto T \]

-or-

\[ PV = nRT \]

The Ideal Gas Law

where
- \( n \) is the number of moles of gas
- \( R = \frac{8.31}{mol \cdot K} \) is the universal gas constant

34 The number of moles of an ideal gas is doubled while the temperature and container volume remain the same. What happens to the pressure of the gas?

- A It doubles
- B It quadruples
- C It remains the same
- D It is decreased to one-half
- E It is decreased to one-fourth

Answer: A

PV=nRT
T, V are constant so P~n
n doubles, so P doubles

35 An ideal gas is taken through a closed cycle A⇒B⇒C⇒A. As shown on the diagram. Which point is associated with the highest temperature?

- A A
- B B
- C C
- D The temperature is the same at A, B and C
- E More information is required
An ideal gas is taken through a closed cycle A → B → C → A. As shown on the diagram. Which point is associated with the highest temperature?

- A
- B
- C
- D
- E

T is greatest when P x V is greatest.

PV = nRT

35

Kinetic Theory

Atoms and the Properties of Matter

The idea that all familiar matter is made up of atoms goes back to the ancient Greeks. Democritus (460-370 BC) proposed that if one was to cut a piece of iron into smaller and smaller portions, there would be a portion which could not be divided further. This smallest piece he called an atom (indivisible).

Another of the scientists working with gases, John Dalton (1766-1844) reprised the theory of atoms. His experiments led him to propose the foundational postulates of modern chemistry. The first of these was that the elements (fundamental materials) are made of atoms, and that all atoms of a given element have the same physical properties (mass, size...).

Kinetic Theory

From Dalton's principles, it followed that properties of materials are determined by the properties of the atoms in the material, the motion of those atoms (or molecules), and their interactions.

In fact, the properties of matter we have studied so far: thermal expansion, melting, boiling, cooling, heating... can be explained based on the concept that matter is made up of atoms.

Kinetic theory is the first of the theories of materials. It relates the motion of the atoms (or molecules) which comprise a gas to the thermodynamic properties of the gas.

Kinetic theory provides us with a better understanding of:
- temperature
- heat
- internal energy
- pressure

Kinetic Theory of Ideal Gases

The study of a real gas can be mathematically complicated.

We would have to consider:
- the motion of the molecules in the gas
- the motion of the atoms inside each molecule
- whether the molecules repel or attract and even stick together
- whether they will combine to form different molecules.

In our application of kinetic theory we will be using a set of simplifying assumptions called The Ideal Gas Model.

The Ideal Gas Model

Assumptions:
1. The gas consists of a very large number of particles (atoms, molecules) in a container.
2. The particles behave as point particles; their size is small in comparison to the average distance between particles and to the size of the container.
3. The particles are in constant motion; they obey Newton's Laws of motion. Each particle collides occasionally with a wall of the container. These collisions are perfectly elastic.
4. The walls of the container are rigid and very massive.
36 Which of the following is not included into the assumptions of the ideal gas?

- A The number of molecules in a container is very large
- B The molecules interact when they collide with each other
- C The molecules interact all the time during their motion because of intermolecular forces
- D The collisions between molecules are perfectly elastic
- E The size of molecules can be ignored

---

Pressure

The first physical quantity we will model is with kinetic theory is pressure. Pressure is just the force per unit area on the walls of the container:

\[ P = \frac{F}{A} \]

Recalling the impulse-momentum change theorem, we can find the force from the change of momentum:

\[ F = \frac{\Delta p}{\Delta t} \]

So we'll begin our derivation by finding \( \Delta p \) and \( \Delta t \) for a single molecule colliding with one of the very massive walls of its perfectly rigid container.

Pressure

These collisions occur every time interval \( \Delta t \), which is the time it takes for the molecule to travel from one side of the box to the other and back, a distance of \( 2L \).

\[ 2L = v_x \Delta t \]

or

\[ \Delta t = \frac{2L}{v_x} \]
Pressure

To calculate the force due to all the molecules in the box, we have to add the force contributions from each molecule.

\[ F_x = \frac{m}{L} (v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2) \]

Defining the average value of the square of the \( x \) component of velocity:

\[ \overline{v_x^2} = \frac{(v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2)}{N} \]

We find that the force on a wall due to all the molecules is:

\[ F_x = \frac{m}{L} \overline{v_x^2} \]

Pressure

The velocity in one direction can be related to the speed using the Pythagorean Theorem:

\[ \overline{v^2} = \overline{v_x^2} + \overline{v_y^2} + \overline{v_z^2} \]

Since all molecules move in random directions and there is no preference between \( x \), \( y \), and \( z \), we can write the following:

\[ \overline{v_x^2} = \overline{v_y^2} = \overline{v_z^2} = \frac{1}{3} \overline{v^2} \]

The force on a wall is then:

\[ P = \frac{m}{3L}N\overline{v^2} \]

Recalling that \( P = F/A \) and that volume \( V = LA \), we get:

\[ P = \frac{1}{3} \frac{mN\overline{v^2}}{V} \]

The Pressure in an ideal gas is directly proportional to the average square speed of the molecules.

\[ \frac{P}{V} = C \]

The Ideal Gas Law and Average Kinetic Energy

Since \( KE = \frac{m\overline{v^2}}{2} \)

we can rewrite our expression for pressure as:

\[ PV = \frac{2}{3} N(KE) \]

Comparing this to the ideal gas law that emerged from experiments:

\[ PV = nRT \quad \text{or} \quad PV = NkT \]

Where \( k = 1.38 \times 10^{-23} \text{ J/K} \), the Boltzmann constant.

We find that:

\[ \frac{KE}{2} = \frac{3}{2} kT \]

- The average kinetic energy of molecules in a gas is directly proportional to the absolute temperature.
- This is the most important result of kinetic theory.
- The higher the temperature, the faster molecules move on the average.

The Ideal Gas Law and RMS Velocity

Comparing \( KE = \frac{m\overline{v^2}}{2} \) and \( KE = \frac{3}{2} kT \)

We can find the "root-mean-square" velocity:

\[ \overline{v_{rms}} = \sqrt{\frac{3kT}{m}} \quad \text{· The higher the temperature, the faster molecules move on average.} \]

Kinetic Theory Summary

- The pressure in the ideal gas is directly proportional to the average square of the velocity of molecules.

\[ P = \frac{1}{3} \frac{mN\overline{v^2}}{V} \]

- Temperature was explained on the microscopic level.
- The average kinetic energy of molecules in a gas is directly proportional to the absolute temperature.
- The temperature can’t be negative and can only reach absolute zero when the average kinetic energy of molecules is zero.

\[ KE = \frac{3}{2} kT \]

- The average velocity of molecules depends on absolute temperature and molecular mass.
- Increasing temperature causes molecules to move faster.
- The lighter the molecule, the faster it moves.

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

37 If the average kinetic energy of molecules is increased while the number of moles is kept constant, what happens to the pressure of an ideal gas?

- A It increases
- B It decreased
- C It remains constant
- D It decreases and then increases
- E None from the above
37 If the average kinetic energy of molecules is increased while the number of moles is kept constant, what happens to the pressure of an ideal gas?

- A It increases
- B It decreases
- C It remains constant
- D It decreases and then increases
- E None from the above

\[ PV = \frac{2}{3} N(KE) \]

if \( N/V \) is kept constant
but \( KE \) increases
then \( P \) increases

38 The average kinetic energy of molecules can be increased by increasing which of the following?

- A Pressure
- B Volume
- C Temperature
- D Number of moles
- E All of the above

\[ KE = \frac{3}{2} kT \]

if \( T \) is raised, \( KE \) is raised

39 If the temperature of an ideal gas is increased from 300 K to 600 K, what happens to the average kinetic energy of the molecules?

- A It doubles
- B It quadruples
- C It reduced to 1/2
- D It is reduced to 1/4

\[ KE = \frac{3}{2} kT \]

if the absolute \( T \) is doubled
\( KE \) is doubled

40 If the temperature of an ideal gas is increased from 25 C to 50 C, what happens to the average kinetic energy of the molecules?

- A It doubles
- B It quadruples
- C It is cut to one-half
- D It is cut to one-fourth
- E It slightly increases

\[ KE = \frac{3}{2} kT \]
40. If the temperature of an ideal gas is increased from 25°C to 50°C, what happens to the average kinetic energy of the molecules?

- A. It doubles
- B. It quadruples
- C. It is cut to one-half
- D. It is cut to one-fourth
- E. It slightly increases

If T is increased from 298K to 323K, KE is raised slightly.

41. If the absolute temperature of an ideal gas is doubled, what happens to the average speed of the molecules?

- A. It doubles
- B. It quadruples
- C. It increases by √2
- D. It decreases by √2
- E. It remains unchanged

If absolute T is doubled, \( v_{\text{rms}} \) is increased \( \sqrt{2} \).

Internal Energy

Internal Energy of an Ideal Gas

The internal energy of an ideal gas depends on temperature and the number of moles \( n \) of gas.

\[
U = \frac{3}{2} nRT
\]

An increase in temperature causes an increase in internal energy.

* A mole is a quantity of matter. It is defined as the 6.022 \times 10^{23} molecules of a substance. (Defined as the number of atoms of carbon-12 that would have a mass of 12 grams).
42. The temperature of a monatomic ideal gas is increased from 35°C to 70°C. How does it change its internal energy?

- A. It doubles
- B. It quadruples
- C. It is slightly increased
- D. It is decreased to one-half
- E. It is decreased to one-fourth

C

43. The state of an ideal gas is changed through the closed path 1 → 2 → 3 → 1. What happens to the internal energy of the gas between point 2 and point 3?

- A. It increases
- B. It decreases
- C. It remains constant
- D. It decreases and then increases
- E. It increases and then decreases

2 → 3 is isothermal
U is constant

44. Which of the following is true about the melting process?

- A. The energy is required to increase the average kinetic energy of molecules
- B. The energy is required to decrease the average kinetic energy of molecules
- C. The energy is required to increase the potential energy between the molecules
- D. The energy is required to decrease the potential energy between the molecules
- E. No energy is required for this process; it happens spontaneously
44. Which of the following is true about the melting process?

A. The energy is required to increase the average kinetic energy of molecules
B. The energy is required to decrease the average kinetic energy of molecules
C. The energy is required to increase the potential energy of molecules (moving them further apart)
D. The energy is required to decrease the potential energy of molecules
E. No energy is required for this process, it happens spontaneously

C. The energy is required to increase the potential energy between the molecules (moving them further apart)

Work in Thermodynamics

Internal Energy, Heat and Work

The state of any thermodynamic system can be described with the internal energy. The internal energy (U) of a thermodynamic system can be changed in two different ways: adding heat (Q) to the system or doing work (W) on the system.

In 1845, James Joule presented his paper "On the mechanical equivalent of heat". He was the first to discover that doing work on a system increases its temperature.

Work in Thermodynamics

A simple example of a thermodynamic system is a quantity of gas enclosed in a cylinder with a movable piston. Consider the work done by the gas during its expansion.

Note: The work done will be positive since the pressure (force) and the motion of the piston are in the same direction.

Work in Thermodynamics

Suppose that the cylinder has a cross-sectional area A and the pressure exerted by the gas is $P_{gas}$. The force exerted by the gas on the piston is $F = P_{gas}A$. When the piston moves up a distance $\Delta x$ and the pressure $P$ is constant, the work $W$ is

$$W = F \Delta x = (P_{gas}A) \Delta x$$

or

$$W = P \Delta V$$

Work in Thermodynamics

When the piston moves down, the gas volume decreases, and so the work $W$ done by the gas is negative.

$$W = -P_{gas} \Delta V$$

During the compression of the gas the work $W'$ done by the external force $F_{ext}$ is positive.

$$W' = P_{ext} \Delta V$$

The relationship between work done by the gas and work done on the gas can be presented by following:

$$W' = -W$$
Work in Thermodynamics

This relationship can be represented as a graph of $P$ as a function of $V$ on a PV-diagram.

The work done equals the area under the curve on a PV-diagram.

In an expansion, the work done by the gas is positive.

$$W = area = P \Delta V = \text{Work}$$

$W = area = P(V_B - V_A) > 0$

In a compression, the work done by the gas is negative.

$$W = area = P(V_B - V_A) < 0$$

45 The state of an ideal gas is changed in a closed path $1 \rightarrow 2 \rightarrow 3 \rightarrow 1$. Which of the following is true about work done by the gas between point 1 and point 2?

- A Work done by the gas is positive
- B Work done by the gas is negative
- C Work done by the gas is zero
- D Work done by the gas is greater than work done on the gas
- E Work done by the gas is less than work done on the gas

46 The state of an ideal gas is changed in a closed path $1 \rightarrow 2 \rightarrow 3 \rightarrow 1$. Which of the following is true about work done by the gas between point 2 and point 3?

- A Work done by the gas is positive
- B Work done by the gas is negative
- C Work done by the gas is zero
- D Work done by the gas is greater than work done on the gas
- E Work done by the gas is less than work done on the gas
The state of an ideal gas is changed along a closed path $X \Rightarrow B \Rightarrow Y \Rightarrow A \Rightarrow X$. What is the total amount of work done on the gas?

A $2PV$
B $-2PV$
C $PV$
D $-PV$

The work done along (area under) $X \Rightarrow B$ is $2PV$.
The work done along (area under) $Y \Rightarrow A$ is $-PV$.
The net work done is $PV$.

First Law of Thermodynamics

In previous sections of this chapter we defined the internal energy, heat, and work in thermodynamics.

Now we will combine them in one formula expressing conservation of energy in thermal processes.

We have learned that there are two ways to increase the Internal Energy ($\Delta U$) of a thermodynamic system:

- add Heat ($Q$) to the system
- do Work ($W = P_{ext}\Delta V$) on the system

$\Delta U = Q + W$  First Law of Thermodynamics

150 J of heat is added to a system and 100 J of work done on the system. What is the change in the internal energy of the system?

A 250 J  B 150 J  C 100 J  D 50 J  E 0 J
48 150 J of heat is added to a system and 100 J of work done on the
system. What is the change in the internal energy of the system?

A 250 J
B 150 J
C 100 J
D 50 J
E 0 J

\[\Delta U = Q + W'\]
\[= 150 + 100\]
\[= 250 \text{ J}\]

49 250 J of heat is added to a system and the system does 100 J of
work on surroundings. What is the change in the internal energy of
the system?

A 250 J
B 150 J
C 100 J
D 50 J
E 0 J

\[\Delta U = Q + W'\]
\[= 250 - 100\]
\[= 150 \text{ J}\]

Thermodynamic Processes

We will next look at 4 different types of thermodynamic processes
and their effect on the internal energy of a thermodynamic system.

Each process takes place in a situation where a particular
thermodynamic quantity is fixed.
- isothermal - constant temperature (\(\Delta T = 0\))
- isobaric - constant pressure (\(P = \text{constant}\))
- isochoric - constant volume (\(\Delta V = 0\))
- adiabatic - no heat added or removed (\(Q = 0\))

In an isothermal process, the
temperature is constant (\(\Delta T = 0\)).
For an ideal gas, since the internal
energy \(U\) depends on \(T\),
- the internal energy is constant
- \(\Delta U = 0\)

Since \(\Delta U = Q + W'\), or \(Q = -W'\)
- all heat added to the system is
  converted to work by the system

Isothermal Processes

"Isothermal processes" by Neothermal96 - Own work. Licensed under CC BY
https://commons.wikimedia.org/wiki/File:Isothermal_processes.svg
Isobaric Processes

In an isobaric process, the pressure is constant. Since $W' = P_{ext} \Delta V$
- work is done as soon as the volume as $V$ changes
- $\Delta U \neq 0$
- $\Delta U = Q + W'$ can be rearranged as $Q = \Delta U - W'$
- heat added to the system causes the system to do work and will increase its internal energy (and therefore $T$)

Isochoric Processes

In an isochoric process, the volume is constant ($\Delta V = 0$).
- no work is done
- $\Delta U = Q$
- heat added to the system increases the system's internal energy

Adiabatic Processes

In an adiabatic process, no heat enters or leaves the system ($Q = 0$).
- any work done on the system increases its internal energy
- any work by the system decreases its internal energy
- $\Delta U = W'$

Thermodynamic Processes

<table>
<thead>
<tr>
<th>Process</th>
<th>Condition</th>
<th>First Law</th>
</tr>
</thead>
<tbody>
<tr>
<td>isothermal</td>
<td>$\Delta T = 0$ ($\Delta U = 0$)</td>
<td>$0 = Q + W'$</td>
</tr>
<tr>
<td>isobaric</td>
<td>$\Delta P = 0$</td>
<td>$\Delta U = Q + W'$</td>
</tr>
<tr>
<td>isochoric</td>
<td>$\Delta V = 0$ ($W' = 0$)</td>
<td>$\Delta U = Q$</td>
</tr>
<tr>
<td>adiabatic</td>
<td>$Q = 0$</td>
<td>$\Delta U = W'$</td>
</tr>
</tbody>
</table>

50 A sample of an ideal gas is taken through a closed cycle. Which of the following is true about the change in internal energy and work done on the gas between point 2 and point 3?

- A $\Delta U = 0$, $W' > 0$
- B $\Delta U = 0$, $W' = 0$
- C $\Delta U = 0$, $W' < 0$
- D $\Delta U > 0$, $W' > 0$
- E $\Delta U < 0$, $W' < 0$

50 A sample of an ideal gas is taken through a closed cycle. Which of the following is true about the change in internal energy and work done on the gas between point 2 and point 3?

- A $\Delta U = 0$
- B $\Delta U = 0$
- C $\Delta U = 0$ since 2 $\rightarrow$ 3 is an isotherm
- D $\Delta U > 0$ since $W' < 0$ the gas is doing work
- E $\Delta U < 0$, $W' < 0$
A sample of an ideal gas is taken through a closed cycle. Which of the following is true about the change in internal energy and work done on the gas between point 1 and point 2?

A $\Delta U = 0$, $W' > 0$
B $\Delta U > 0$, $W' = 0$
C $\Delta U = 0$, $W' < 0$
D $\Delta U > 0$, $W' > 0$
E $\Delta U < 0$, $W' < 0$

Answer: B $\Delta U > 0$ since $1 \rightarrow 2$ PV=nRT is increasing
$W' = 0$ since V is constant

Second Law of Thermodynamics

- Many thermal processes proceed naturally in one direction but not the opposite. These are called irreversible.
- For example, heat by itself always flows from a hot object to a cooler object, never the reverse. The reverse process would not violate the first law of thermodynamics; energy would be conserved.
- In order to properly account for irreversible processes, the second law of thermodynamics was formulated.

Sadi Carnot is credited with having formulated the Second Law of Thermodynamics in 1824

Heat flows naturally from a hot object to a cold object; Heat never flows spontaneously from a cold object to a hot object.

Heat Engines
A heat engine is a thermodynamic system that converts heat into mechanical energy, which can then be used to do mechanical work.

Denis Papin (about 1690), inventor of the pressure cooker, first described three basic components of any heat engine:
- a high-temperature reservoir
- a low-temperature reservoir
- an engine containing gas or steam

How a heat engine works:
- The high-temperature reservoir transfers an amount of heat $Q_H$ to the engine
- In the engine, part of the heat is transformed into work $W$ (during the expansion of gas)
- The rest of the heat, $Q_L$, is exhausted to the low-temperature reservoir

The efficiency $e$ of any heat engine can be defined as a ratio of work $W$ to the heat input $Q_H$.

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

One of the important challenges in engine design was that of increasing efficiency.

This question was answered in 1824 by the French engineer Sadi Carnot.

Carnot proposed an idealized heat engine (The Carnot Engine) that would provide the maximum possible efficiency consistent with the second law of thermodynamics.
The Carnot engine operates between a high temperature reservoir at $T_H$ and a low temperature reservoir at $T_L$.

The Carnot engine consists of:
- two reversible isothermal processes $A \Rightarrow B$ and $C \Rightarrow D$,
- two reversible adiabatic processes $B \Rightarrow C$ and $D \Rightarrow A$.

Carnot Theorem

No heat engine operating between two temperatures $T_H$ and $T_L$ engine can have a greater efficiency than a Carnot engine operating between the same two temperatures.

$$e = \frac{T_H - T_L}{T_H}$$  ideal efficiency

Note: temperatures are in Kelvin.

53. A Carnot engine moves 1000 J of heat from a 500 K reservoir to a 300 K reservoir. With what efficiency will the engine produce work during this process?

A 10%
B 20%
C 30%
D 40%
E 50%

Answer: D

$$e = \frac{500 - 300}{500} = 0.40$$

W = 400 J

54. A Carnot engine moves 1000 J of heat from a 500 K reservoir to a 300 K reservoir. How much work did the engine perform?

A 400 J
B 800 J
C 1000 J
D 1600 J
E 200 J

Answer: A

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

0.4 = W/1000

W = 400 J
Entropy and Disorder

Entropy

Kinetic energy of macroscopic objects (a ball, pendulum...) is associated with organized, coordinated motions of many molecules.

In contrast, heat transfer involves changes in energy of random, disordered molecular motion.

Therefore conversion of mechanical energy into heat involves an increase of randomness or disorder. The energy is conserved, its just no longer usable to do work in the system.

Entropy provides a quantitative measure of this disorder.

Entropy and the Second Law

The total entropy of an isolated system never decreases

The mixing of the two liquids cannot be undone.

Entropy increased by mixing and it can't be reduced.

55

- Heat will flow until thermal equilibrium is reached
- No net heat flow occurs once thermal equilibrium is reached
- Entropy is reduced during the process
- The natural flow of heat will be from the warmer system to the colder system

55

- Heat will flow until thermal equilibrium is reached
- No net heat flow occurs once thermal equilibrium is reached
- Entropy is reduced during the process
- The natural flow of heat will be from the warmer system to the colder system

Answer: C
When ice freezes, its molecules become much more structured. Does this break the second law of thermodynamics?

- Yes. The second law of thermodynamics is wrong. All scientists should be fired right now.
- No, because the energy of the molecules increased
- No, because the entropy of the system as a whole (ice + ice's environment) increased
- No, because the density of water decreased

Answer: C

No, because the entropy of the system as a whole (ice + ice's environment) increased.