Physics 23 Chapter 12

Temperature and Heat

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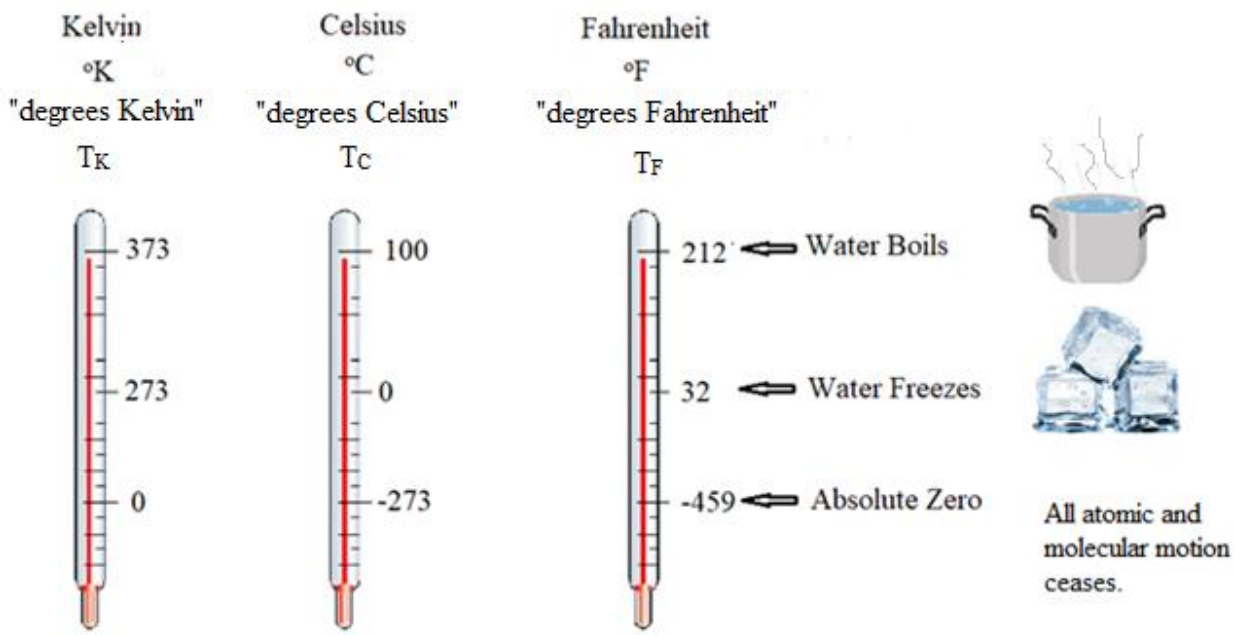


Conversion Formulas
$T_C = T_K - 273$
$T_F = (9/5)\, T_C + 32$
$T_K = T_C + 273$

<u>Example</u>:

Given that water boils at 100 °C, show that the Fahrenheit temperature at which water boils is 212 °F.

$$T_F = (9/5)\,(100) + 32$$
$$= 180 + 32$$
$$= 212\ ^{o}\text{F}.$$

Example A:	Example B:
A Fahrenheit thermometer and a Celsius thermometer are placed together in a block of ice. The Fahrenheit reading is 20 degrees higher than the Celsius reading. What are the two temperatures? $T_F = T_C + 20$ $(9/5)\ T_C + 32 = T_C + 20$ $T_C = -15\ ^oC$ $T_F = (9/5)(-15) + 32$ $= 5\ ^oF$	A Kelvin thermometer and a Celsius thermometer are placed in a liquid. The Kelvin thermometer reading is four times the Celsius thermometer reading. What are the two temperatures? $T_K = 4\ T_C$ $= 4\ (T_K - 273)$ $T_K = 364\ ^oK$ $T_C = 91\ ^oC$

Example C:	Example D:
At what Fahrenheit temperature are Celsius and Fahrenheit temperatures the same? $T_C = (9/5)\ T_C + 32$ $T_C = -40\ ^oC$ $T_F = (9/5)(-40) + 32$ $= -40\ ^oF$	What is absolute zero on the Fahrenheit Scale? Absolute zero on the Celsius scale: $T_C = -273\ ^oC$ $T_F = (9/5)\ (-273) + 32$ $= -459\ ^oF$

Remove Subscript from T_C

Henceforth in this chapter, only the Celsius temperature scale will be used, so no ambiguity will be possible if the subscript "C" is omitted. Symbol for Celsius Temperature: T

Small and Large Calories

Non-standard units of energy:

Small Calorie 1.0 calorie (cal) = 4.19 J The small calorie is sometimes called the "physicist's calorie." **Large Calorie** 1.0 Calorie (Cal) = 1000 cal The large Calorie is sometimes called the "dieter's calorie."	 1.0 Cal 1000 cal  160 Cal 160,000 cal

<u>Example</u>:

A 140-lb person eats one chocolate-frosting cupcake, which has a caloric content of 160 Cal.

How much work would this person have to do to burn off this cupcake's Calories?

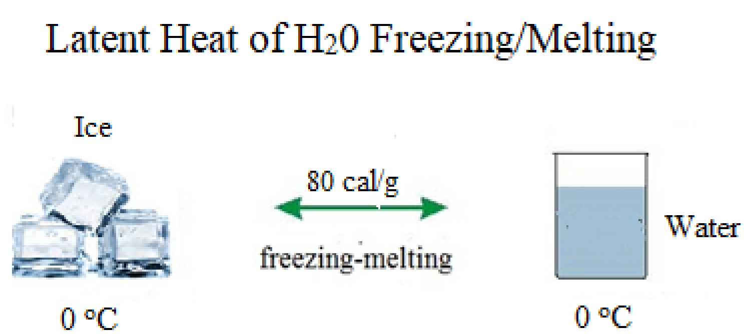
Fact: A 140-pound person burns about four Calories per minute walking at normal pace. So, in 40 minutes she would burn 160 calories.

Eating cupcakes commits a person to spend a substantial amount of her future to canceling its effect.

Phases and Phase Changes

The three common "phases" of matter are solid, liquid, and gas. When a substance changes from one phase to another, a "phase change" has occurred.

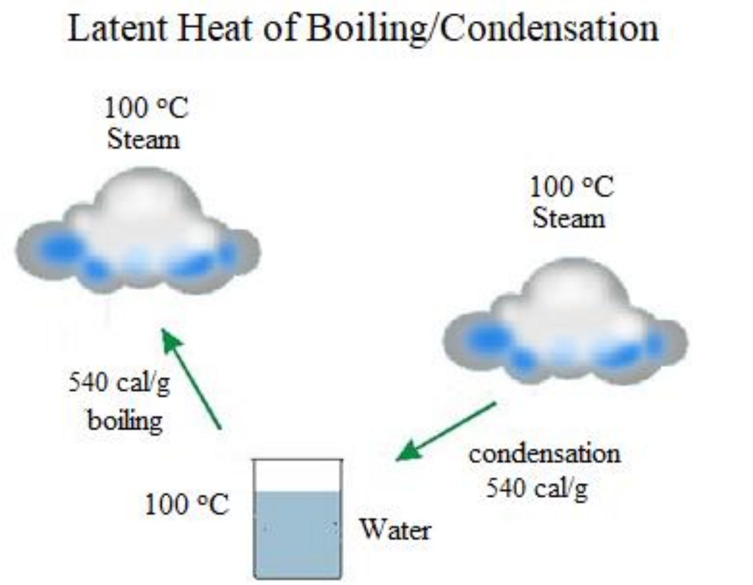
When a phase change is occurring, the temperature of the mixture does not change, even as heat is added or removed: $\Delta T = 0$.



While melting is occurring, the temperature of the ice-water mixture doesn't change: it remains at 0 ^{o}C.

Similarly, when freezing is occurring, the temperature of the ice-water mixture remains at 0 ^{o}C.

Only after all of the ice is melted, or the water frozen, and then heat is added or removed, will the temperature then change.



While steam at 100 ^oC is condensing, the temperature of the steam-water mixture doesn't change; it remains at 100 ^oC.

Similarly, when water at 100 ^oC is being converted to steam, the temperature of the steam-water mixture remains at 100 ^oC.

Only after the water is completely converted to steam, or the steam completely condensed, will the temperature change when heat is added or removed.

Solid-Liquid H_2O Phase Changes

<table>
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<td>
<u>Melting Ice</u>

Each gram of ice that is to be melted requires 80 calories of heat be added. This number--80 cal/gram--is called the "latent heat of <u>melting</u>."

L = 80 cal/g

Note: As the ice at 0 ^oC is melting, un-melted ice and melted ice water will co-exist and be at the same temperature: 0 ^oC; the melted ice water will remain at 0 ^oC until all of the ice is melted. *Temperature doesn't change in a phase change.*
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<td>
<u>Freezing Water</u>

Water will not begin to freeze until its temperature has been lowered to its freezing point, which is 0 ^oC. After that temperature is reached, additional removal of heat will cause the water to freeze. For each gram of water at 0 ^oC to be frozen, 80 cal of heat must be removed.

This number, 80 cal/g, is called the "latent heat of <u>freezing</u>"

L = 80 cal/g

Note that the latent heat of freezing is the same as the latent heat of melting, so we will use the same symbol for both latent heats.

Note: as the water at 0 ^oC is freezing, the temperature of the freezing ice and water mixture *does not change*: it remains at a temperature of 0 ^oC.

Temperature doesn't change in a phase change.
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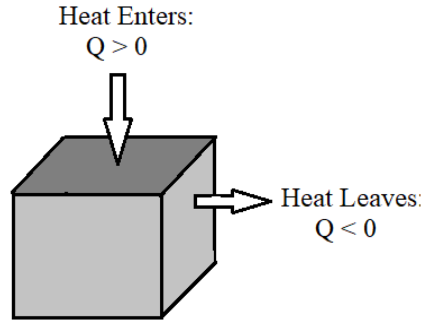
Q = Quantity of Heat



Q symbolizes the quantity of heat entering or leaving a substance.

Latent Heat of Freezing and Melting Equations

In heat problems involving a melting or freezing of H_2O the relevant equations are the ones shown below:

m = mass in grams
Q = heat energy in cal

Freezing: $Q = -(80\ \text{cal/g})\ m$ (heat is removed)

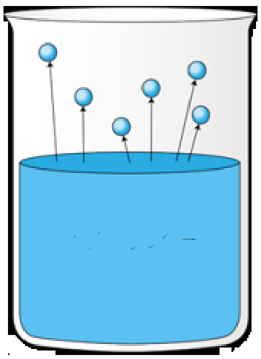
Melting: $Q = (80\ \text{cal/g})\ m$ (heat is added)

Example A:	Example B:
What quantity Q of heat must be removed from 200 grams of water at 0 ^{o}C to convert it to 200 grams of ice at 0 ^{o}C? Answer: $Q = -\ (80\ \text{cal/g})\ 200\ \text{g}$ $= -16{,}000$ cal $= -16$ kilocalories (kcal) A negative Q indicates heat was *removed*.	What quantity Q of heat must be added to 300 grams of ice at 0 ^{o}C to convert it to 300 grams of water at 0 ^{o}C? $Q = (80\ \text{cal/g})(300\ \text{g})$ $= 300\ \text{g})(80\ \text{cal/g})$ $= 24{,}000$ cal $= 24$ kcal A positive Q indicates heat was *added*.

Vaporization Phase Changes of H_2O

There are two types of vaporization phase changes of H_2O: Evaporation, and Boiling

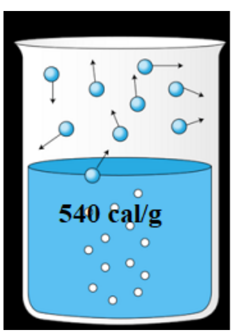
1. Evaporation



The first type of vaporization is called "evaporation, " which occurs *at the surface* of water, or on skin coated with perspiration, when water escape into the surrounding air and become water vapor. The rate of evaporation depends on the temperature difference between the surface and the air, the relative humidity, wind speed, and atmospheric pressure.

We will not attempt to perform calculations involving this type of phase change of water. We instead will concern ourselves only with the second type of vaporization of water: Boiling.

2. Boiling



The second type of vaporization is called "boiling." At the boiling point of water, 100 ^{o}C, water molecules below the surface are transformed into water vapor bubbles, which rise to the surface. The water vapor inside the bubbles escape into the air and become "steam" at 100 ^{o}C .

Each gram of 100 ^{o}C water that is to be boiled away must absorb 540 cal of heat.

This number is called the "latent heat of boiling."

$$L = 540 \text{ cal/g}$$

<u>Example</u>: How many calories of heat must be added to 200 grams of water at 100 ^oC to convert it to 200 grams of steam at 100 ^oC? Q = = (540 cal/g))(200 g) = 108,000 cal = 108 kcal	

Condensation

The process that is the reverse of vaporization is "condensation." One type of condensation that occurs is the reverse of evaporation. In this type of condensation, water moisture in the air changes to the liquid phase on the surface of automobiles parked outdoors on a cold night, or on leaves and blades of grass.

We will not concern ourselves with this type of condensation. Instead, we will deal only with the condensation process that is the reverse of boiling, i.e., the condensation of steam at 100 ^oC into water at 100 ^oC.

To condense one gram of steam at 100 ^oC, 540 cal must be removed.

$$L = 540 \text{ cal/g}$$

This number is called the "latent heat of condensation," and is the same number as the latent heat of boiling.

Latent Heat of Boiling and Condensation

In heat problems involving boiling of water or condensation of steam, the operational equations are the ones shown below:

m = number of grams
Q = heat energy in calories
Condensation: $Q = -(540\ \text{cal/g})\ m$ (heat is removed)
Boiling: $Q = (540\ \text{cal/g})\ m$ (heat is added)

<u>Example</u>:

How many calories of heat must be removed from 50 g of steam at 100 ^{o}C to convert it to 50 g of water at 100 ^{o}C?

$$Q = (540\ \text{cal/g})\ (50\ \text{g})$$
$$= -27{,}000\ \text{cal}$$
$$= -27\ \text{kcal}$$

Celsius Temperature Changes

Symbol: ΔT
Units: C^{o} ("Celsius degrees")

Note that the units of temperature *change* are not the same as the units of temperature; the degree superscript appears after the letter "C," so instead of saying, "degrees Celsius," we say, "*Celsius degrees.*"

$T_o = 30\ ^{o}C$
$T = 75\ ^{o}C$
$\Delta T = T - T_o$
$= 45\ C^{o}$

Similarly, if Fahrenheit or Kelvin temperatures are used, the corresponding temperature change symbols are F^{o} and K^{o}, respectively.

Specific Heat Capacity

All substances have a property called "specific heat capacity," symbolized as "c" and measured in units of cal/g-C^o. *

$$Q = mc\,\Delta T$$

$$\Delta T = Q/(mc\,\Delta T)$$

This equation shows that the higher the heat capacity c, the larger will be the denominator of the ratio above, and therefore the smaller will be the temperature change ΔT when heat Q is added or removed.

We can summarize by saying "*things with high heat capacity cool down and warm up more slowly*."

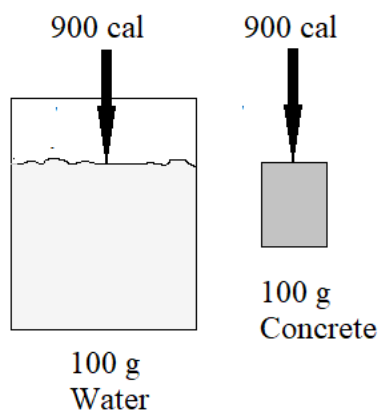
As the table shows, among all solids and liquids, water has the second-highest heat capacity. Only ammonia has a higher heat capacity.

Substance	c cal/g-C^o
Water	1.0
Ice	0.5
Steam	0.5
concrete	0.2
glass	0.2
aluminum	0.2
bone	0.1
ammonia	1.1
air	0.2

* We usually just call it, "heat capacity."

Example:

Calculate and compare the temperature changes of the water and concrete in the diagram below.



$$\Delta T = Q / (mc)$$

Water: $mc = (100\text{ g})(1.0\text{ cal/g-C}^{o})$
$= 100\text{ cal/C}^{o}$

$\Delta T = Q / (mc)$
$= 900\text{ cal} / 100\text{ cal/C}^{o}$
$= 9\text{ C}^{o}$

Concrete: $mc = (100\text{ g})(0.20\text{ cal/g-C}^{o})$
$= 20\text{ cal/C}^{o}$

$\Delta T = 900 / 20\text{ cal/C}^{o}$
$= 45\text{ C}^{o}$

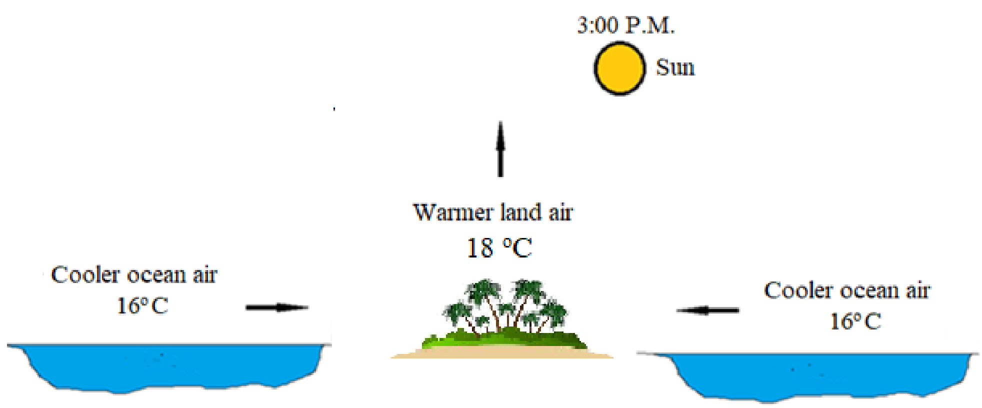
The heat capacity of concrete is only one-fifth that of water, so concrete's temperature rise is five times greater.

Ocean Breezes During the Day

Land, having much lower heat capacity than water, warms up and cools down more than water, which has a comparatively high heat capacity. These facts explain ocean and land breezes, as discussed below

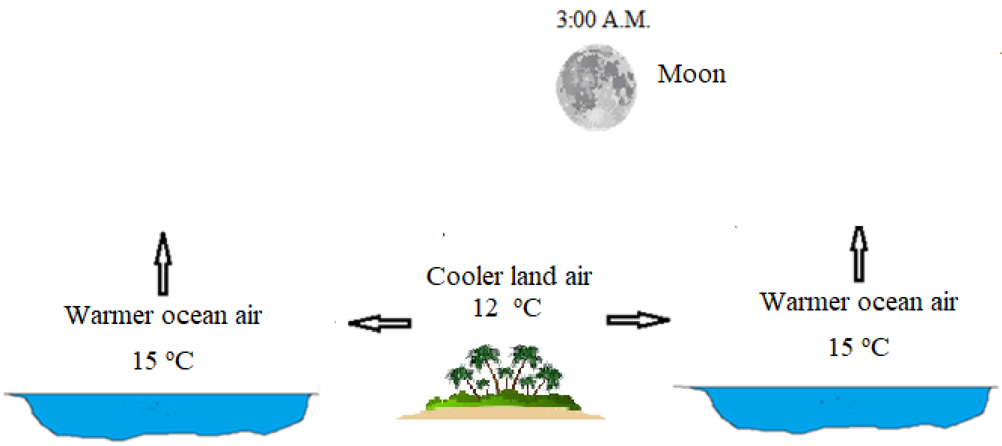
The figure below shows land surrounded by ocean water at mid-afternoon; both land and water are being warmed by sunlight. Water's temperature rise is less than land's rise. The warmer less dense air above land, immersed in surrounding cooler more dense air, rises just like the rising warmer less dense air above a barbeque pit immersed in surrounding more dense cooler air.

Nearby cooler ocean air moves toward land to take the place of the rising land air. The resulting air current, blowing from ocean to land is called an "ocean breeze."



Land Breezes at Night

During the night the ocean water and the land both lose heat and experience a temperature decrease, but the water's temperature drops less (-1 C^o) than does the land temperature (-6 C^o). The air over the water is now warmer than the air over land, so the more buoyant warmer air rises and the land air travels from land to ocean to take the place of the rising air, thereby creating a "land breeze" blowing from the land to the ocean, in a direction which is opposite to the flow of air in an ocean breeze.



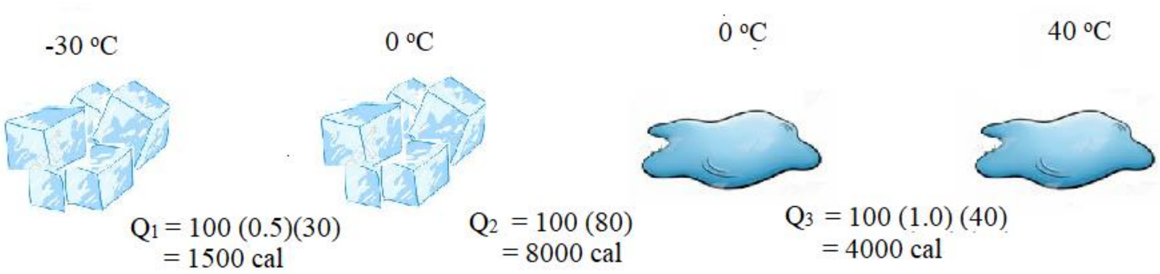
Example A:	Example B:
How much heat must be added to 100 grams of water at 30 ^{o}C to raise its temperature to 70 ^{o}C? $\Delta T = 70 - 30$ $= 40$ C^o $Q = mc\,\Delta T$ $= (100\text{ g})(1.0\text{ cal/g-C}^o)\ 40\text{ C}^o$ $= 4000$ cal The positive sign indicates that heat entered the water.	How much heat must be removed from 60 grams of ice at -30 ^{o}C to cool it down to -90 ^{o}C? $\Delta T = -90 - (-30)$ $= -60$ C^o $Q = mc\Delta T$ $= 60\ (0.5)\ (-60)$ $= -1800$ cal The negative sign indicates that heat left the ice.

Example:

100 grams of ice at -30 ^oC is to be converted to 100 grams of water at 40 ^oC. How much heat must be added to the ice to accomplish this?

We may imagine that the heat is added in three stages, when in practice the heat would be added all at once.

1. Q_1: Warm the ice up to its melting point at 0 ^oC.
2. Q_2: Melt the 0 ^oC ice.
3. Q_3: Warm the melted 0 ^oC ice water up to 40 ^oC.



$$Q = Q_1 + Q_2 + Q_3$$

$$= 13{,}500 \text{ cal}$$
$$= 13.5 \text{ kcal}$$

Mixture Problems

When two substances at different temperatures are mixed together, one of them gains heat, while the other loses heat. Whatever one gains, the other loses. One of the Q's is positive, while the other is the negative of the first one. The sum of the two Q's is zero. $Q_1 + Q_2 = 0$ When "equilibrium" is reached, both substances will be at the same temperature.	<u>Example</u>: Sixty grams of aluminum (c = 0.20 cal/g-C^o) at 75 ^{o}C is added to 300 g of water (c = 1.0 cal/g-C^o) at 30 ^{o}C. What is the equilibrium temperature? $Q_1 = 60\ (0.20)\ (T - 75)$ $Q_2 = 300\ (1.00)\ (T - 30)$ $Q_1 + Q_2 = 0$ T = 31.7 ^{o}C

Example

25 grams of ice at -20 ^{o}C is placed in 160 grams of water at 25 ^{o}C that is inside a 70-gram glass bowl, also at 25 ^{o}C.

All the ice melts.

What is the equilibrium temperature of the mixture?

Latent Heat of Melting: 80 cal/g
Ice: c = 0.5 cal/g-C^{o}
Water: c = 1.0 cal/g-C^{o}
Glass: c = 0.2 cal/g-C^{o}

1. Warm the Ice: $Q_1 = 25\ (0.5)\ (20)$
2. Melt the Ice: $Q_2 = 25\ (80)$
3. Heat the Melt Water: $Q_3 = 25\ (1.0)\ (T - 0)$
4. Cool Bowl's Water: $Q_4 = 160\ (1.0)\ (T - 25)$
5. Cool Glass Bowl: $Q_5 = 70\ (0.2)\ (T - 25)$

$$Q_1 + Q_2 + Q_3 + Q_4 + Q_5 = 0$$

$$T = 10.55\ ^{o}C$$

Example:

3.0 grams of steam at 100 ^oC is injected into 140 grams of room temperature (20 ^oC) water; all of the steam condenses. What is the equilibrium temperature?

Steam condenses: $Q_1 = -(3.0)(540)$
$= -1620$ cal

Hot condensed water cools: $Q_2 = (3.0)(1.0)(T - 100)$

Room temperature water warms: $Q_3 = (140)(1.0)(T - 20)$

$$Q_1 + Q_2 + Q_3 = 0$$

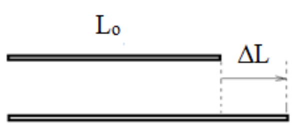
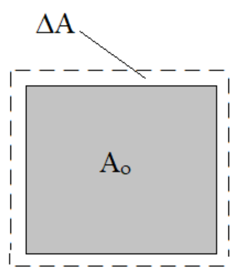
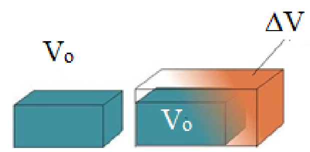
$$-1620 + (3.0)(1.0)(T - 100) + (140)(1.0)(T - 20) = 0$$

$$T = 33\ ^oC$$

Thermal Expansion and Contraction of Solids and Liquids

Solid and liquid substances have a property called the "coefficient of thermal expansion." The larger its value, the greater is the amount by which the object's length, area, or volume changes when the temperature changes.

Symbol: α
Units: $(C^o)^{-1}$

1-Dimensional Solids	2-Dimensional Solids	3-Dimensional Solids
 $\Delta L = \alpha L_o \Delta T$	 $\Delta A = (2\alpha) A_o \Delta T$	 $\Delta V = (3\alpha) V_o \Delta T$

Note: The multiplier of α (1, 2, or 3) in the three equations above corresponds to the solid's dimensionality.

<u>Example</u>:

The temperature of a 20-meter aluminum bar changes from 20 ^{o}C to 200 ^{o}C. What is the new length of the bar?

The coefficient of thermal expansion of aluminum is $2.3 \times 10^{-5} (C^o)^{-1}$.

$L = L_o + \alpha L_o \Delta T$
$= 20.00 \text{ m} + (2.3 \times 10^{-5} /C^o)(20 \text{ m})(180 \text{ C}^o)$
$= 20.08 \text{ m}$

Practical Examples of Linear Expansion

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When concrete slabs in bridges, overpasses, or sidewalks expand in hot spells beyond the space in the joints, they press against each other, causing the surfaces to buckle at the joint.

To prevent buckling, expansion joints, such as those shown at the right, allow the structures to expand and contract without distortion.

Below: Highway concrete buckling occurred in record high temperatures.


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Below: Bridge expansion joint

Below: Sidewalk expansion joint

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Area Expansion

<u>Example A</u>:

A circular aluminum plate has an area
$A_o = 8.0425\ cm^2$.

What will be the new area if the plate's temperature is increased by 100 C^o ?

The coefficient of thermal expansion of aluminum is $2.3 \times 10^{-5}\ (C^o)^{-1}$.

<u>Solution:</u>

$A = A_o + 2\alpha A_o \Delta T$

$= 8.0425 + 2\ (2.3\text{x}10^{-5})\ (8.0425)\ (100)$
$= 8.0795\ cm^2$

Volume Expansion of Solids

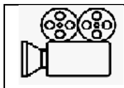
<u>Example B</u>:

The temperature of a 20-m^3 aluminum block changes from 20 ^{o}C to 200 ^{o}C.

What is the new volume of the block?

<u>Solution:</u>

$V = V_o + 3\alpha V_o \Delta T$
$= 20 + 3\ (2.30 \times 10^{-5})\ (20)\ (180)$
$= 20.25\ m^3$

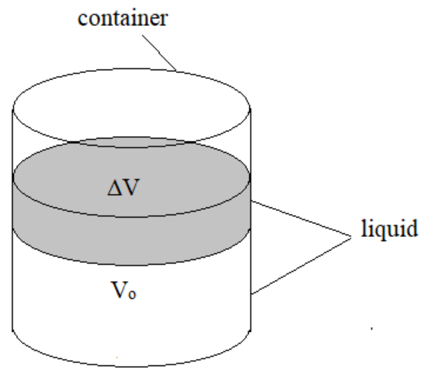
	Video Diameter of metal ball changes when its temperature is changed.

Volume Expansion of Liquids

Liquids don't expand linearly, so there is no coefficient of linear expansion, α, and therefore the volume expansion coefficient is not 3α as it is for solids. Instead, liquids have a property called "the coefficient of *volume* expansion."

Symbol: β
Units: $(C^o)^{-1}$

$\Delta V = \beta V_o \Delta T$



Example:

An aluminum can, open at the top, has volume of 2000 milliliters (mL). It is filled to the top with mercury.

Mercury: $\beta = 1.82 \times 10^{-4} (C^o)^{-1}$ (Volume expansion coefficient)
Aluminum: $\alpha = 2.3 \times 10^{-5} (C^o)^{-1}$ (Linear expansion coefficient)

What volume of mercury will spill over the top when the temperature of the can and its contents are increased by 275 C^o ?

Solution:

The mercury's expansion will be greater than the can's expansion, so the excess will overflow the top of the can.

$$
\begin{aligned}
\text{Amount Spilled} &= \Delta V_{liquid} - \Delta V_{can} \\
&= \beta V_o \Delta T - 3\alpha V_o \Delta T \\
&= 1.82 \times 10^{-4} (2000)(275) - 3(2.3 \times 10^{-5})(2000)(275) \\
&= 62 \text{ mL}
\end{aligned}
$$