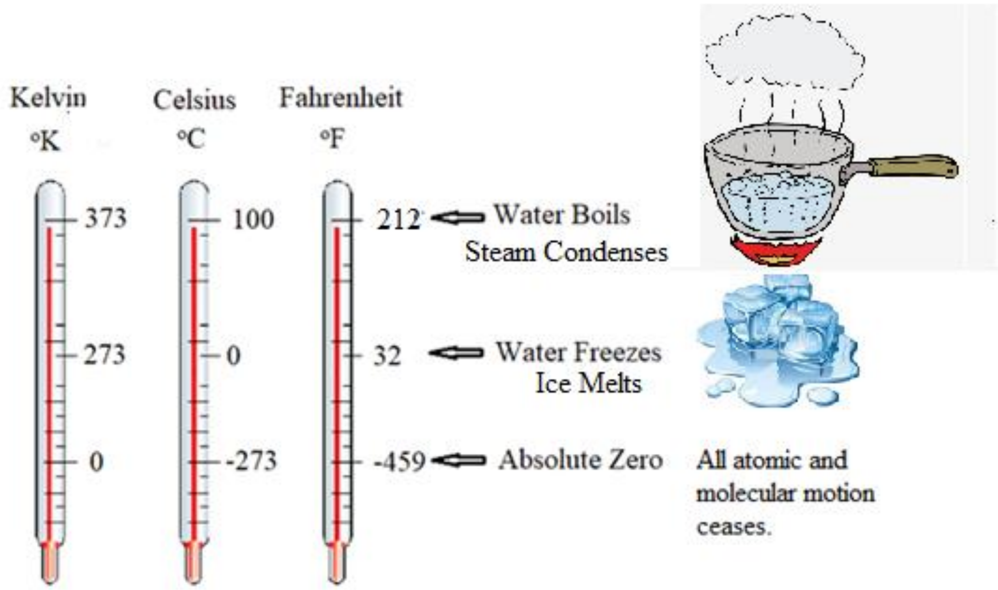
Physics 17 Part I

Temperature and Heat



T_K T_C T_F

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

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? $T_F = T_C + 20$ $(9/5)\ T_C + 32 = T_C + 20$ $T_C = \mathbf{-15\ ^oC}$ $T_F = (9/5)(-15) + 32$ $= \mathbf{5\ ^oF}$	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:

At what temperature are the Celsius and Fahrenheit temperatures the same?

$T_F = (9/5)\ T_C + 32$
Require T_F above be equal to T_C:
$T_C = (9/5)\ T_C + 32$
$T_C =$ **$-\ 40\ {}^oC$**

$T_F = (9/5)(-40) + 32$
$=$ **$-40\ {}^oF$**

Small and Large Calories

We introduce new, non-standard, units of energy:

Small Calorie 1.0 calorie (cal) = 4.19 J The small calorie is also called the “chemist’s calorie.” Physicists call it the “physicist’s calorie.” **Large Calorie** 1.0 Calorie (Cal) = 1000 cal The large Calorie is called the “dieter’s calorie.”	 1.0 Cal 1000 cal 4190 J Diet Coke	 140 Cal 140,000 cal 5.87×10^5 J Chocolate Frosting Cupcake

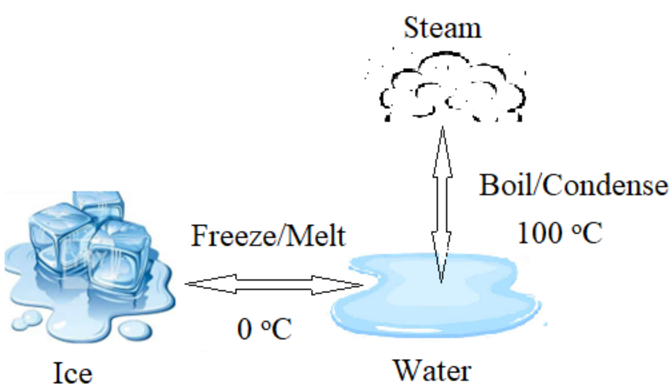
Phases and Phase Changes

Three of the more common “phases” of matter are solids, liquids, and gases. When a substance changes from one phase to another a “phase change” has occurred.

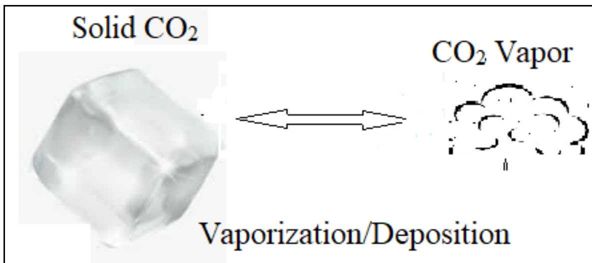
Important Fact:

When a phase change is occurring, the temperature of the substance does not change.

H_2O Phase Changes



Carbon Dioxide Phase Changes



CO_2 doesn’t exist in liquid form.

Solid-Liquid H_2O Phase Changes

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<u>Melting Ice</u>


A melting ice-water system remains at 0 ^oC until all of the ice is melted.

Each gram of ice at 0 ^oC that is to be melted requires 80 calories of heat be added. This number--80 cal/gram--is called the “latent heat of melting.”

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.

Temperature doesn’t change in a phase change.
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<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 freezing”

L = 80 cal/g

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

As the water at 0 ^oC is freezing, the temperature of the 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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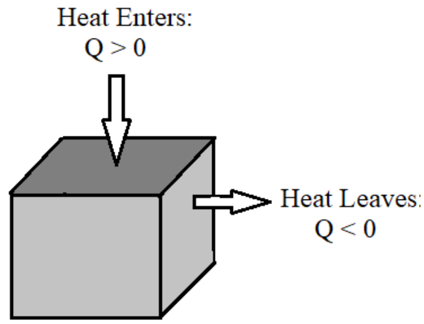


Q symbolizes the quantity of heat entering or leaving a substance.

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?	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?
Answer: Q = -mL = -200 g (80 cal/g) = -16,000 cal = -16 kilocalories (kcal)	Q = mL = (300 g)(80 cal/g) = 24,000 cal
Negative Q's indicate that heat is *removed*.	Positive Q's represent heat *added*.

Note in the two examples above:

$$Q = \pm\ (80 \text{ cal/g})\ m$$

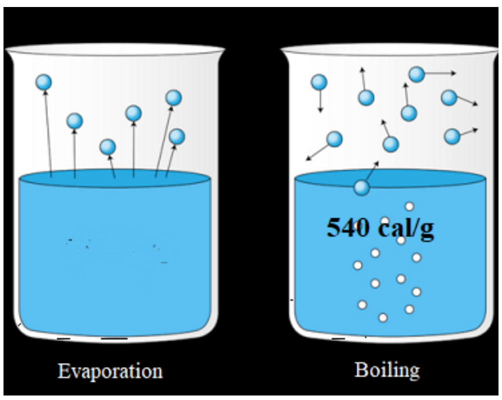
The negative sign is used when heat is removed, while the positive sign is used when heat is added.

Vaporization: Evaporation vs Boiling

There are two types of vaporization of water. The first occurs at the surface of a pool of water or on skin coated with perspiration, when water molecules with sufficient kinetic energy escape into the surrounding air and become water vapor.

This type of vaporization is called "evaporation."

The second type of vaporization is called "boiling." At the boiling point of water, 100 ^{o}C, water vapor is formed below the surface, creating bubbles, which rise to the surface, and the water vapor inside them escapes into the air. The escaping water vapor is called "steam."



To cause one gram of water to boil, 540 calories of heat energy must be absorbed. Less energy than this is necessary to cause evaporation. Evaporation problems will not be discussed in this course.

540 cal/g is called the "latent heat of vaporization," also called the "latent heat of boiling."

L = 540 cal/g

<u>Example A</u>:

How many calories of heat must be added to 200 grams of water at 100 ^{o}C to convert it to 200 grams of steam at 100 ^{o}C?

Q = 200 (540)
    = 108,000 cal
    = 108 kcal

<u>Example B</u>:

Five grams of perspiration evaporates from an athlete's skin. How many calories of heat leave the person?

<u>Answer</u>:

Q = - (540 cal/g)(5 g)
    = 2700 cal

Condensation

"Condensation" is the reverse of vaporization. To condense one gram of water vapor at 100 ^{o}C, 540 cal must be removed.

L = 540 cal/g

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

<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?

$$\begin{aligned} Q &= -mL \\ &= -(50\text{ g})\,(540) \\ &= -27{,}000\text{ cal} \\ &= -27\text{ kilocalories (kcal)} \end{aligned}$$

Note in the examples above concerning boiling versus condensation,

$$Q = \pm\,(540\text{ cal/g})\ m$$

The negative sign is used when heat is removed, while the positive sign is used when heat is added.

Celsius Temperature Changes

Symbol: ΔT

Note: To simplify the expression, we suppress the subscript "C" that normally appears on Celsius temperature symbols: We will use T instead of T_C.

T_o = 30 ^oC (initial temperature)
T = 75 ^oC (final temperature)
$\Delta T = T - T_o$ (change in temperature)
= 45 C^o ("Celsius degrees")

Note: Celsius degrees (C^o) are not the same as degrees Celsius (^oC). The former are units of temperature *change*, while the latter are *temperatures*.

Heat Capacity

Recall: When a phase change is occurring as heat is added or removed, the temperature doesn't change. However, heat is added or removed from a substance whose phase is *not* changing, a temperature change will occur.

All substances have a property called "heat capacity." The other name for this property is longer: "specific heat capacity." We will primarily use "heat capacity."

The greater the heat capacity of a substance, the more heat it can absorb without its effect being as noticeable, i.e., without it experiencing a larger temperature change.

The units of heat capacity are cal/g-C^o.

Substance	**c cal/g-C^o**
Water	1.0
Ice	0.5
Steam	0.5
concrete	0.2
soil	0.2
aluminum	0.2
bone	0.1

$$Q = mc\Delta T$$

$$\Delta T = \frac{(Q/m)}{c}$$

Example A:

Prove that substances with larger heat capacities experience smaller temperature changes, i.e., they cool down less rapidly, and likewise warm up more slowly, than substances with smaller heat capacities.

Proof:

Let the heat capacity of one substance be c_1, and let the higher heat capacity of a second substance be c_2. Show that ΔT_2 is smaller than ΔT_1 when equal amounts of heat Q enter or leave equal masses m.

$\Delta T_2 = (Q/m)/c_2$
$\Delta T_1 = (Q/m)/c_1$
Divide:

$\Delta T_2 / \Delta T_1 = c_1/c_2$
$\Delta T_2 = (c_1/c_2)\, \Delta T_1$
$c_1/c_2 < 1$ so
$\Delta T_2 < \Delta T_1$

Example B:

40×10^3 calories of heat enter an 8000-gram block of concrete (0.20 cal/g-C^o); another 400 cal enters 8000 grams of water (c = 1.00 cal/g-C^o). What are the temperature increases each substance experiences?

Concrete:

$40 \times 10^3 = 8000\ (0.20)\ \Delta T$
$\Delta T = 25\ C^o$

Water:

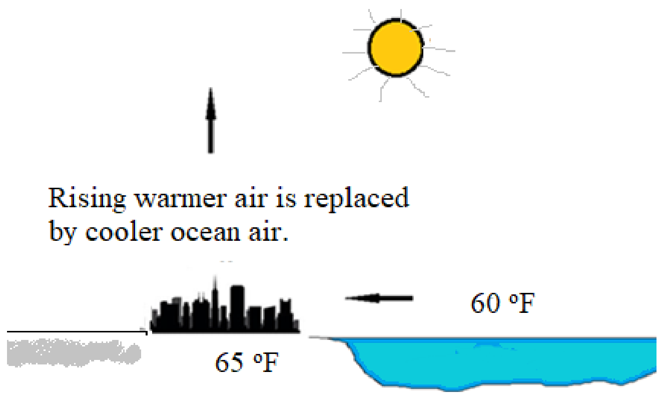
$40 \times 10^3 = 8000\ (1.00)\ \Delta T$
$\Delta T = 5\ C^o$

The temperature rise of the water is one-fifth of the temperature rise of the concrete block.

Ocean Breezes During the Day

During the daytime sunlight warms not only a city, but also the nearby ocean. The result is a cool "ocean breeze" directed from the water toward the city. The explanation for this breeze is provided below.

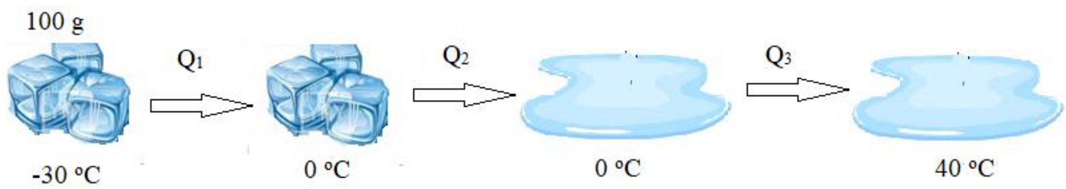
Ocean water has a much higher heat capacity ($c = 1.0$ cal/g-C^o) than the city's asphalt, concrete and metal buildings streets ($c = 0.20$ cal/g-C^o), so the ocean water warms up more slowly than does the city, so city air is hotter than ocean air. In the discussion and diagram below, it's important to note that the warmer air above the city is less dense than the surrounding cooler air and is therefore like a wood block under water, rising.



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	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

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?



$Q_1 = 100\ (0.5)\ (30)$
$= 1500$ cal

$Q_2 = 100\ (80)$
$= 8000$ cal

$Q_3 = 100\ (1.0)\ (40)$
$= 4000$ cal

$$Q = Q_1 + Q_2 + Q_3$$
$$= 13{,}500 \text{ cal}$$
$$= 13.5 \text{ kcal}$$

Mixture Problems

When two substances at different temperatures are mixed together, one of them gains the heat lost by the other.

One of the Q's is the negative of the other one, so 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.

Example:

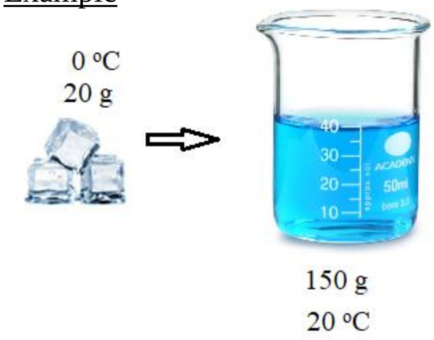
Sixty grams of aluminum at 75 ^oC is added to 300 g of water at 30 ^oC.

What is the equilibrium temperature?

$Q_1 = 60\ (0.20)\ (T - 75)$
$Q_2 = 300\ (1.00)\ (T - 30)$

$Q_1 + Q_2 = 0$
$T = 31.7\ ^oC$

<u>Example</u>



Twenty grams of ice at 0 ºC is added to 150 grams of water at 20 ºC. All of the ice melts.

What is the equilibrium temperature?

1. Heat absorbed in melting the ice:

$$Q_1 = 20\ (80)$$
$$= 1600 \text{ cal}$$

2. Heat absorbed by melted ice water:

$$Q_2 = 20\ (1.0)\ (T - 0)$$
$$= 20\ T$$

3. Heat lost by cooling beaker water:

$$Q_3 = 150\ (1.0)\ (T - 20)$$
$$= 150\ T - 3000$$

$$Q_1 + Q_2 + Q_3 = 0$$
$$1600 + 20\ T + 150\ T - 3000 = 0$$

$$T = 8.24\ ^{o}C$$