1.11: Temperature, Heat, and Energy

Learning Objectives

- To identify the different between temperature and heat
- To recognize the different scales used to measuring temperature

Temperature

Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.

Figure 1.11.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.

Figure 1.11.2: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right)

The Fahrenheit Temperature Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as 32^oF and the boiling point as 212^oF. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Temperature Scale

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701 - 1744). The Celsius scale sets the freezing point and boiling point of water at 0^oC and 100^oC respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".

The Kelvin Temperature Scale

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824 - 1907). It is based on molecular motion, with the temperature of $0\,K$, also known as **absolute zero**, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is $273.15\,K$, while the boiling point is $373.15\,K$. Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales where temperatures are referred to as "degrees F" or "degrees C", we simply designated temperatures in the Kelvin scale as kelvin.

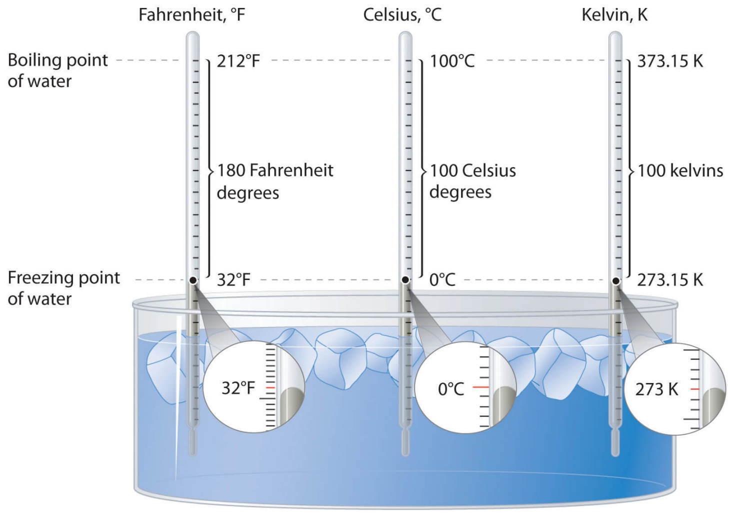

Figure 1.11.3: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in }^\circ\text{C}) + 273.15 = T\,(\text{in K}) \tag{3.10.1}$$

$$T\;(\text{in K}) - 273.15 = T\;(\text{in }^\circ\text{C}) \tag{3.10.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales are as follows:

$$^\circ C = \frac{5}{9} \times (^\circ F - 32) \tag{3.10.3}$$

$$^\circ F = \frac{9}{5} \times (^\circ C) + 32 \tag{3.10.4}$$

Example 1.11.1: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 3.10.3:

$$\begin{aligned} ^\circ C &= \frac{5}{9} \times (103.5^\circ F - 32) \\ &= 39.7\ ^\circ C \end{aligned}$$

Converting from Celsius to Kelvin requires the use of Equation 3.10.1:

$$\begin{aligned} K &= 39.7\ ^\circ C + 273.15 \\ &= 312.9\ K \end{aligned}$$

> ❓ Exercise 1.11.1
>
> Convert each temperature to °C and °F.
>
> a. the temperature of the surface of the sun (5800 K)
> b. the boiling point of gold (3080 K)
> c. the boiling point of liquid nitrogen (77.36 K)
>
> **Answer (a)**
>
> 5527 K, 9980 °F
>
> **Answer (b)**
>
> 2807 K, 5084 °F
>
> **Answer (c)**
>
> -195.79 K, -320.42 °F

Heat

While the concept of temperature may seem familiar to you, many people confuse temperature with heat. as discussed above, **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures.

When scientists speak of ***heat***, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot – the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents.

Energy

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we're awake or asleep.

Energy is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by 1^oC. For example, raising the temperature of 100 g of water from 20^oC to 22^oC would require $100 \times 2 = 200 \text{ cal}$.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains 85 kcal or $85,000 \text{ cal}$ In order to make the distinction, the dietary calorie is written with a capital C.

$$1 \text{ kilocalorie} = 1 \text{ Calorie} = 1000 \text{ calories} \tag{1.11.1}$$

To say that the snack "contains" 85 Calories means that 85 kcal of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1 \text{ J} = 0.2390 \text{ cal or } 1 \text{ cal} = 4.184 \text{ J} \tag{1.11.2}$$

We can calculate the amount of heat released in kilojoules when a 400. Calorie hamburger is digested.

$$400 \text{ Cal} = 400 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 1.67 \times 10^3 \text{ kJ} \tag{1.11.3}$$

Heat Capacity and Specific Heat

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. Because of its much larger mass, the swimming pool of water has a larger "heat capacity" than the wading pool. Similarly, different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot.

We would say that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by $1^{o}C$). Water is very resistant to changes in temperature, while metals in general are not. The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1^{o}C$. The symbol for specific heat is c_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree $(J/g^{o}C)$ or calories per gram per degree $(cal/g^{o}C)$ (Table $1.11.1$). This text will use $J/g^{o}C$ for specific heat.

$$\text{specific heat} = \frac{\text{heat}}{\text{mass} \times \text{cal}/g^{o}C} \tag{1.11.4}$$

Notice that water has a very high specific heat compared to most other substances.

Table 1.11.1: Specific Heat Capacities

Substance	Specific Heat Capacity at 25^{o}C in J/g ^{o}C	Substance	Specific Heat Capacity at 25^{o}C in J/g ^{o}C
H_2 gas	14.267	steam @ 100^{o}C	2.010
He gas	5.300	vegetable oil	2.000
$H_2O(l)$	4.184	sodium	1.23
lithium	3.56	air	1.020
ethyl alcohol	2.460	magnesium	1.020
ethylene glycol	2.200	aluminum	0.900
ice @ 0oC	2.010	Concrete	0.880
steam @ 100oC	2.010	glass	0.840

Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see table above). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

Figure 1.11.4: This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.

Example 1.11.1: Measuring Heat

A flask containing 8.0×10^2 g of water is heated, and the temperature of the water increases from 21 °C to 85 °C. How much heat did the water absorb?

Solution

To answer this question, consider these factors:

- the specific heat of the substance being heated (in this case, water)
- the amount of substance being heated (in this case, 800 g)
- the magnitude of the temperature change (in this case, from 21 °C to 85 °C).

The specific heat of water is 4.184 J/g °C, so to heat 1 g of water by 1 °C requires 4.184 J. We note that since 4.184 J is required to heat 1 g of water by 1 °C, we will need 800 times as much to heat 800 g of water by 1 °C. Finally, we observe that since 4.184 J are required to heat 1 g of water by 1 °C, we will need 64 times as much to heat it by 64 °C (that is, from 21 °C to 85 °C).

This can be summarized using the equation:

$$\begin{aligned} q &= c \times m \times \Delta T \\ &= c \times m \times (T_{\text{final}} - T_{\text{initial}}) \\ &= (4.184\ \text{J}/\ \cancel{\text{g}}\ ^\circ\text{C}) \times (800\ \ \cancel{\text{g}}) \times (85 - 21)^\circ\text{C} \\ &= (4.184\ \text{J}/\ \cancel{\text{g}}\ ^\circ\ \cancel{\text{C}}) \times (800\ \ \cancel{\text{g}}) \times (64)^\circ\ \cancel{\text{C}} \\ &= 210,000\ \text{J} (= 210\ \text{kJ}) \end{aligned}$$

Because the temperature increased, the water absorbed heat and q is positive

Exercise 1.11.1

How much heat, in joules, must be added to a 5.00×10^2 g iron skillet to increase its temperature from 25 °C to 250 °C? The specific heat of iron is 0.451 J/g °C.

Answer

5.05×10^4 J

Summary

- Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).
- Heat capacity is the amount of heat required to raise the temperature of an object by 1°C).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C.

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