Chapter 5: Thermal Energy, the Microscopic Picture

Goals of Period 5

Section 5.1: To define temperature, thermal energy, evaporative cooling, diffusion, thermal expansion, and thermodynamic systems
Section 5.2: To discuss temperature and phase changes
Section 5.3: To examine pressure, temperature, volume and the ideal gas law
Section 5.4: To define the first law of thermodynamics
Section 5.5: To investigate order, disorder, and entropy

In the previous period we discussed thermal energy and the concept of heat as thermal energy in transit. We defined ***thermal energy*** or ***internal energy*** as the kinetic energy involved in random molecular motion together with the potential energy stored in molecular attractions. We also saw that the temperature of an object is an indication of how hot or how cold an object is, but is not a measure of the amount of thermal energy the object has. In this period, we look in more detail at thermal phenomena in terms of the atomic theory of matter.

All matter is made of atoms or groups of atoms called molecules. So, it is not surprising that the study of thermal energy on an atomic or molecular basis is necessary in order to gain a fundamental understanding of the subject. In fact, we find that many phenomena such as diffusion, evaporation, thermal expansion, and changes of phase in matter can all be described on a molecular or atomic level.

5.1 Thermal Energy and Molecular Motion

Temperature and Thermal Energy

For brevity in the following discussion, we use the term “molecules” to refer either to molecules or to single atoms. Molecules are always in random motion, even when no motion is discernible to the naked eye. The fact that there is random molecular motion can be inferred from the motion of small dust particles suspended in air. The molecules or atoms making up the air are constantly in random motion, and as the dust particles in the air are struck, they also move about in a random fashion. This random motion of the dust particles is known as Brownian motion. The observation of Brownian motion was one of the first accepted experimental proofs for the molecular or atomic theory of matter.

For gases, the thermal motion of molecules is random not only in direction, but also in speed. Some molecules are fast, some are slow. However, it is possible to speak of an average molecular speed applying to the entire group of molecules making up an object. It is this average molecular speed that determines the temperature of the object. The higher the average speed of the molecules in an object, the higher the temperature. A graph of the molecular speed for the molecules of a gas at a temperature of 295 kelvin is shown in Figure 5.1. Note that the average speed is not equal to the most probable speed of the molecules.

Figure 5.1 Graph of Molecular Speeds of a Gas

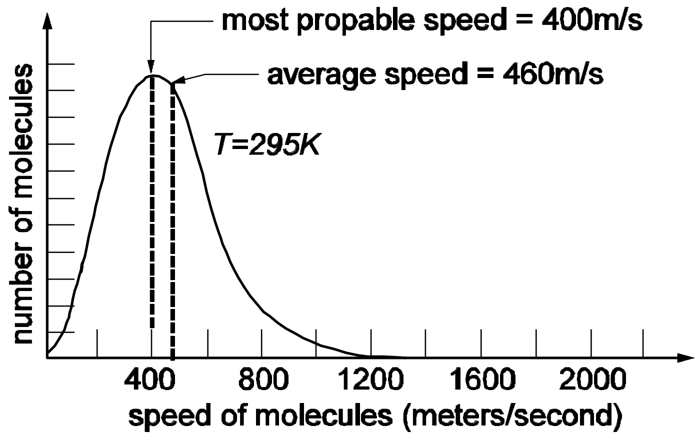
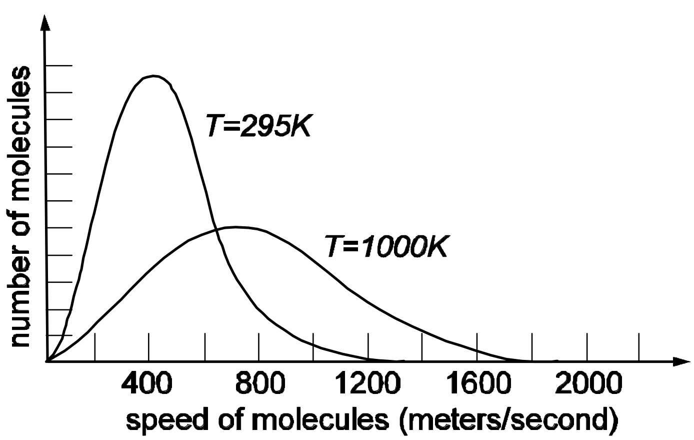


Figure 5.2 illustrates the change in the speed of the molecules of a gas when the gas temperature changes. The graph shows the molecular speed of oxygen gas at a temperature of 295 K and 1000 K.

Figure 5.2 Comparison of the Speeds of Gas Molecules at Different Temperatures



You might ask what happens if all of the molecules have the smallest amount of energy that they can have. In this case, the temperature corresponds to the lowest possible temperature, zero Kelvin, and is called absolute zero. Zero Kelvin occurs at very close to – 273 ^{o}C and to – 460 ^{o}F. It is not true, however, that at this temperature all motion of the molecules of an object will cease. If we remove from an object all of the thermal energy that can be removed, the molecules of the object will still have some residual energy, known as the zero point energy.

The molecular picture explains how mechanical energy can be transformed into thermal energy by friction. Take, for example, a block of wood sliding across a table top. As frictional forces do work to slow down the block, some of the mechanical energy of the block is converted into random motion of the individual molecules in the block and the tabletop, that is, into thermal energy.

Evaporative Cooling

We observe that a liquid in an open container will eventually disappear. This process is known as evaporation, and gives us more evidence for the motion of the molecules in the liquid. Consider the molecules just below the surface of a liquid. The fast ones are more likely to overcome the intermolecular forces and escape, while the slower molecules stay behind in the liquid. Without the fast molecules, the average molecular speed in the liquid is lower and that implies a lower temperature. The greater the rate of evaporation, the stronger this cooling effect will be. The evaporation rate of a liquid depends on what the liquid is, on the surface characteristics of the liquid, on the temperature of the liquid, and ultimately on the forces that exist between the atoms or molecules making up the liquid. Machine oil, for instance, evaporates so slowly that almost no cooling is noticeable. The cooling effect is more pronounced in liquids such as freon (the working fluid in some air conditioners and refrigerators) that evaporate quickly, because less time is available for the liquid to gain back thermal energy from its environment. You will see several examples of evaporative cooling in class.

Diffusion

If a drop of ink is placed in one corner of a water tank, the ink will gradually spread out and distribute itself all through the water. This spreading out will take place even if the water is not stirred and the temperature of the water is held uniform to prevent convection currents. This process is called ***diffusion***. Diffusion can be understood on the basis of the random motion of molecules in a substance. This is why the drop of ink diffuses in the water. The concentration of the ink molecules is highest when the drop is first placed into the water. As time goes on, the ink molecules spread throughout the container. If the temperature of the water is higher, the diffusion takes place more rapidly. Diffusion takes place more rapidly in gases than in liquids and more rapidly in liquids than in solids.

Thermal Expansion

Molecules in hot matter move faster than molecules in cold matter, and usually the average intermolecular distances are larger. Thus, almost all matter expands when heated. In the classroom we shall see several examples of thermal expansion and contraction and observe that different materials exhibit different characteristics in this regard. There are many industrial applications of this effect, for example, in obtaining a very tight fit between machine parts.

One of the very few exceptions to the general rule that objects expand when heated is water near its freezing point. Between 0 ^{o}C and 4 ^{o}C decreases in volume. The result is that the water expands by about 10 percent when it freezes. The same is true of seawater. This fact makes ice less dense than water, so ice rises to the top and floats. While this effect may seem a nuisance to homeowners, forcing them to drain outdoor water pipes for the winter, it has tremendous ecological consequences because it prevents lakes and oceans from freezing solid. This is very good for fish and all other life forms.

Thermal Equilibrium and Thermodynamic Systems

Thermal equilibrium is an important concept when discussing thermal energy. Let's consider this situation: you take a can of orange juice from your freezer and another from your table. You put both of them in a well-insulated chest and wait. What would you guess would happen after some length of time has expired? Logic tells you that your best guess would be that eventually they are at the same temperature. Why? Well, you already know that the two objects, i.e. the two cans of OJ, will exchange heat. If the cans are at different temperatures, the thermal energy in one OJ can will increase and the thermal energy in the other can will decrease. Your guess is correct. It does not matter if one can was big and the other was small. They will still wind up at the same temperature.

The same is true if you have another object, it could be another can of OJ or some other object, that is at the SAME temperature. Then you very quickly put it in the chest and shut it and all three objects are in equilibrium with each other. Two objects, each in thermal equilibrium with a third object, are all in equilibrium. This rule is called the Zeroth law of thermodynamics.

In learning about thermodynamics we use the word ***system*** when we use the transfer of energy from one object to another by heat (thermal energy in transport) or work. The objects are things that we can perceive by our senses, such as the size, mass and temperature of the object. Such objects are considered macroscopic (a piece of matter), not microscopic (atoms and molecules). The system could be any physical system such as a heat engine that will be discussed in the next period, or a biological system such as your friend. Thermodynamics is the study of these systems.

A ***thermodynamic system*** is any collection of objects, and the rest of the universe is the system's environment. The thermodynamic system impacts the environment around it by heat and/or work. This causes an energy exchange with the environment, and the system's thermal energy, the internal energy of all objects in the system, may change. The ***internal energy*** is the total kinetic energy and potential energy of all objects that are in the system. The system may also have kinetic and potential energies due to outside forces such as gravity. We will discuss this when we look at changing the phase of an object, such as changing water to ice.

5.2 Temperature and Phase Changes

States of Matter

In a solid body, the molecules are tightly bound to each other by intermolecular forces (which are electromagnetic in nature) and can only vibrate about their equilibrium positions. In a liquid, the molecules are still close together and strongly attracted to each other, but they can freely slide past one another and their average vibrations are faster than in the solid state. In a gas, the molecules move so fast and are so far apart that they fly around relatively freely, experiencing intermolecular forces only when they are colliding with each other.

Temperature and Phase Changes

As thermal energy is added to, or taken away from, a substance, two things can happen. First, as you have already seen, the temperature can change. However, the amount that the temperature will change is not the same for all objects, even when the amount of thermal energy added is the same. You already know, for example, that a cup of water will get much hotter than a pan of water, if the same amount of thermal energy is added to each of them. This is because temperature change depends on the amount of matter being heated. Temperature change also depends on the type of material being heated. It takes more energy to raise the temperature of water than to raise the temperature of an equal amount of metal, for example.

The amount of energy that must be added to an object to raise its temperature by one degree is known as the ***heat capacity*** (H_{cap}) of that object. The thermal energy required to change the temperature of an object by several degrees is given in Equation 5.1. If thermal energy is measured in calories and temperature in degrees Celsius, then heat capacity is given in calories/degree C or in joules/degree C.

Change in thermal energy = (Heat capacity) x (Change in Temperature)

(Equation 5.1)

$$Q = H_{cap} \times \Delta T$$

where

Q = heat added or subtracted (calories or joules)

H_{cap} = heat capacity (calories/ ^{o}C or joules/ ^{o}C)

ΔT = change in temperature = $T_{final} - T_{initial}$ (Celsius degrees)

(Example 5.1)

What is the heat capacity of one kilogram of iron if 9,000 joules of thermal energy are required to increase the temperature of the iron by 20 ^{o}C?

Solve Equation 5.1 for the heat capacity, H_{cap}

$$H_{cap} = \frac{Q}{\Delta T} = \frac{9{,}000 \text{ J}}{20 \ ^{o}\text{C}} = 450 \text{ J}/^{o}\text{C}$$

Concept Check 5.1

What is the heat capacity of one kilogram of copper if 7,800 joules of thermal energy are required to increase the temperature of the copper by 20 ^{o}C?

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If we divide the heat capacity of an object by its mass, we obtain a quantity known as the ***specific heat*** (s_{heat}) of the object. The specific heat does not depend on the size or shape of an object, but only on the material from which it is made. Water has a large specific heat of 1 calorie per gram per degree Celsius or 4,186 joules per kilogram degree Celsius. Ice floats because the volume of water increases when it freezes. This is connected to the change in the specific heat of water near 0 ^{o}C. Table 5.1 lists the specific heat of some common materials.

Table 5.1 Specific Heat of Common Materials

Material	Specific Heat (J/kg ^{O}C)	Specific Heat (cal/g ^{O}C)
Water (liquid, 15 ^{O}C)	4,186	1.00
Water (ice, – 5 ^{O}C)	2,100	0.50
Water (steam, 110 ^{O}C)	2,010	0.48
Wood	1,700	0.40
Aluminum	900	0.22
Glass	840	0.20
Iron	450	0.11
Copper	390	0.093
Silver	230	0.056

The heat required to change the temperature of an object can be expressed as

Change in thermal energy = (Specific heat) x (Mass) x (Change in Temperature)

(Equation 5.2)

$$Q = s_{heat} \times M \times \Delta T$$

where

Q = heat added or subtracted (calories or joules)

s_{heat} = specific heat (calories/gram ^{o}C or joules/kilogram ^{o}C)

M = mass (grams or kilograms)

ΔT = change in temperature = $T_{final} - T_{initial}$ (Celsius degrees)

If thermal energy is given in calories, mass in grams and temperature in degrees Celsius, then specific heat is given in calories/(gram degree Celsius).

(Example 5.2)

How much thermal energy is needed to raise 2 kilograms of iron by 10 ^{o}C? The specific heat of iron is 450 J/kg ^{o}C.

$$Q = s_{heat} \times M \times \Delta T = (450 \text{ J/ kg } ^{o}\text{C}) \times 2 \text{ kg} \times 10 \ ^{o}\text{C} = 9{,}000 \text{ J}$$

What is the heat capacity of this object?

Solve Equation 5.1 for the heat capacity, H_{cap}

$$H_{cap} = \frac{Q}{\Delta T} = \frac{9{,}000 \text{ J}}{10 \ ^{o}\text{C}} = 900 \text{ J/}^{o}\text{C}$$

The second thing that can happen when thermal energy is transferred to or from a system is that the state of the system can change. Changes from the solid to the liquid or from the liquid to the gaseous state, and vice versa, are called ***phase changes***. They always involve a transfer of heat, even though the temperature of the substance undergoing the phase change stays constant. As discussed earlier, the heat flow from one object to another can change either the average kinetic energy of the random motion of the molecules, which changes the temperature of the object, or can change the average potential energy of the molecules, which causes the phase of the object to change. Consider what happens when, for instance, a pot of water is heated on a stove. At first, the temperature rises. Upon reaching 100 ^{o}C (212 ^{o}F) the temperature stops increasing, even though the flame keeps supplying heat at the same rate as before. We know that the thermal energy supplied goes into breaking the bonds between the molecules, while the kinetic energy of the molecules remains unchanged. Gradually, more and more molecules gain sufficient energy to overcome the intermolecular forces binding the molecules one to the other. A similar phenomenon occurs when ice melts.

We call the heat required to produce a phase change the ***latent heat*** $\boldsymbol{(L_{heat})}$. Two examples of latent heat are the heat of freezing and the heat of vaporization. The heat of freezing is the amount of thermal energy *given off* as a liquid freezes, and the heat of vaporization is the amount of thermal energy that must be *added* to change a liquid to a gas.

Heat added or subtracted for a phase change = Latent heat x Mass

$$Q = L_{heat}\ M \qquad \textbf{(Equation 5.3)}$$

where

Q = heat (calories or joules)
L_{heat} = latent heat (calories/gram or joules/kilogram)
M = mass (grams or kilograms)

If liquid water at 100 ^{o}C is changed into steam, the heat added (the latent heat of vaporization) is 540 calories for every gram of water. If steam at 100 ^{o}C is changed into water at 100 ^{o}C, 540 calories for every gram of steam must be subtracted. If ice at 0 ^{o}C is changed into liquid water at 0 ^{o}C, the heat added (the latent heat of melting) is 80 calories for every gram of ice. If liquid water at 0 ^{o}C is change into ice at 0 ^{o}C, 80 calories for every gram of liquid water must be subtracted.

Latent heats can be very large. For example, the latent heat of vaporization of water is 540 cal/g and the latent heat of freezing of water is 80 cal/g. Therefore, changing a given quantity of water to steam requires 5.4 times as much heat as warming it from 0 ^{o}C (+32 ^{o}F) to 100 ^{o}C (212 ^{o}F), and melting ice requires as much heat as warming water from 20 ^{o}C (68 ^{o}F) to 100 ^{o}C.

(Example 5.3)

How many calories of heat are required to convert 500 grams of water at a temperature of 25 ^{o}C into steam at 100 ^{o}C?

First, use Equation 5.2 to find the heat required to raise the temperature of the water to 100 ^{o}C. The specific heat of liquid water is 1.00 calories/gram ^{o}C.

$$\boldsymbol{Q} = \boldsymbol{s}_{heat} \times \boldsymbol{M} \times \Delta \boldsymbol{T} = (1.00 \text{ cal/g } ^{o}\text{C}) \times 500 \text{ g} \times (100 \, ^{o}\text{C} - 25 \, ^{o}\text{C}) = 37{,}500 \text{ cal}$$

Next, use Equation 5.3 to find the heat required for the phase change of 500 grams of water at 100 ^{o}C into steam at 100 ^{o}C. The latent heat of evaporation of water is 540 calories/gram.

$$\boldsymbol{Q} = \boldsymbol{L}_{heat} \times \boldsymbol{M} = 540 \text{ cal/g} \times 500 \text{ g} = 270{,}000 \text{ cal}$$

The total heat is the sum of the heat required to heat the water to 100 ^{o}C and the heat required to convert the liquid water into steam.

Total heat = 37,500 cal + 270,000 cal = 307,500 cal

Concept Check 5.2

How many calories heat are required to convert 200 grams of ice at 0 ^{o}C into liquid water at 30 ^{o}C ? The latent heat of melting of ice is 80 calories/gram. The specific heat of ice is 0.5 calories/gram ^{o}C .

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Pressure can influence phase changes. Water expands when it becomes vapor, but by applying enough pressure one can compress the vapor to water. Therefore the boiling temperature of water depends on pressure, and is 100 ^{o}C only at normal atmospheric pressure. In a pressure cooker that is providing a pressure of twice the normal atmospheric pressure, the boiling point of water is raised to about 120 ^{o}C (248 ^{o}F), which shortens cooking times. At the top of Mount Everest, where the pressure is about one third that of normal atmospheric pressure, water boils at 71 ^{o}C (160 ^{o}F).

Similarly, ice can be melted by applying enough pressure. This facilitates ice skating. Under the pressure of a person's weight on the thin blade a little bit of ice melts and the water provides a good lubricant between the blade and the ice surface. If a wire is looped around a piece of ice with the ends of the wire attached to weights, it will work its way through the ice. Under the pressure exerted on it by the wire, the ice beneath the wire melts and the wire moves down a little into the block of ice. The water resulting from the ice that was melted is pushed up above the wires and, since it is no longer under pressure, refreezes. By this process the wire can move through the ice, leaving solid ice again behind it.

5.3 Pressure, Temperature, Volume and the Ideal Gas Law

If energy is added to a substance in a solid or liquid state and if the substance remains in that state, the temperature of the substance will increase. The amount of energy required to increase the temperature of a unit of mass of a substance is obtained by measuring the specific heat of the substance. Specific heat varies substantially from material to material. In general, the specific heat of a material depends on the temperature of the material and on whether the material is in a solid, liquid, or gaseous state.

Visually, we think of the states of matter in terms of their large-scale or macroscopic behavior – solid, liquid, or gas. When studying the phases of matter at the microscopic or atomic level, the phases are quite different. We know that objects such as a chair or a TV set are held together by some attractive force between the molecules that make up the objects. If we were just looking at molecules flying around randomly, we would not see them as a block of iron or a chair. It is very good that molecules in an object attract each other because of the forces between the molecules. In solids the attractive forces are strong enough to keep the molecules close together, sometimes in very regular arrangements. In the liquid state, the attractive forces between the molecules are not as strong and the molecules move around each other. But for gasses, the high speed molecules do not stay close to each other. They move in a random manner, and when the molecules do hit each other, the attractive forces are not large enough to keep them close together. This is why a gas in a closed container will fill the container. However, if you put a liquid or solid in a container, it will stay at the bottom of the container.

As was discussed in Section 5.1 and shown in Figures 5.1 and 5.3, the higher the average molecular speed in a material, the higher the temperature T. Experiments have shown that when the size of the container of gas is held fixed, the ***pressure P*** of the gas, which is the force per unit area that the gas exerts on the walls of its container, is proportional to the temperature of the gas. So, at a fixed volume,

$$P \propto T$$

We also know that if we increase the temperature of a gas but do not exert any force to hold the volume of its container fixed, the pressure will stay fixed and the volume of the gas will expand. Therefore, the volume V of the gas in a container is also proportional to the gas temperature. Thus, at fixed pressure,

$$V \propto T$$

These two experimental results can be combined into the single relationship shown in Equation 5.4.

(Equation 5.4)

$$PV \propto T$$

where

$\propto$		is a symbol indicating "proportional to"
P	=	pressure the gas exerts (newtons/meters^2)
V	=	the volume of the gas (meters^3)
T	=	temperature (Kelvin)

Concept Check 5.3

a) Use ratio reasoning to determine how the pressure exerted by a gas would change if the temperature of the gas was doubled (and all other variables were held constant).

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b) Use ratio reasoning to determine how the pressure exerted by a gas would change if the volume of the gas was doubled (and all other variables were held constant).

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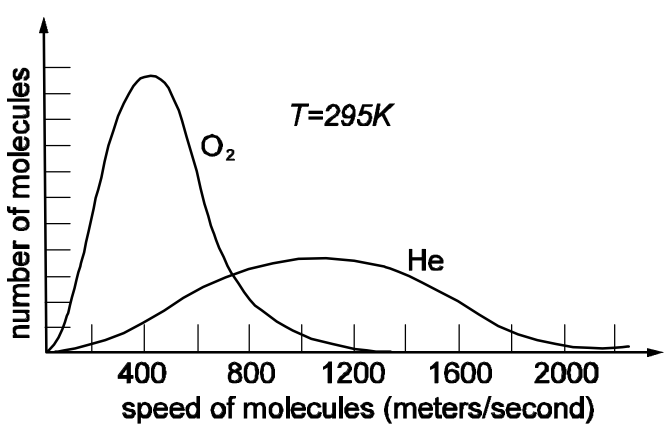
The average speed of the gas molecules depends on both the temperature of the gas and the mass of the molecules. Figure 5.3 illustrates the average speed of two different gases at the same temperature. The oxygen gas molecules are more massive and move randomly with a lower average speed than the molecules of the less massive helium molecules.

Figure 5.3 Comparison of Molecular Speeds of Two Gases



Because the shape of these graphs is known, the average kinetic energy of a molecule of the gas can be calculated and is found to be directly proportional to the absolute temperature of the gas. Recall that kinetic energy can be expressed as

$$E_{kin} = \frac{1}{2} M v_{avg}^2 \qquad \textbf{(Equation 5.5)}$$

Experimentally, the average kinetic energy of a molecule of gas has been found to be 3/2 ***kT***. Therefore,

(Equation 5.6)

$$\frac{1}{2} M v_{avg}^2 = \frac{3}{2} k T$$

where

M = mass of a gas molecule (kilograms)
v_{avg} = the average velocity of the molecules (meters/sec)
k = Boltzmann's constant (1.38×10^{-23} J/K)
T = temperature of the gas (Kelvin)

The internal energy is the energy due to the random motion of the molecules inside the container, rather than any possible motion of the container as a whole. If there are ***N*** number of molecules of a gas in a container of volume ***V***, then the total energy of the gas in the container (the total internal energy of the system) is 3/2 ***N k T***.

By experiment, the pressure of the gas times the volume of the container equals 2/3 of the total energy, 3/2 ***N k T***. Therefore, we can replace the proportional relationship in Equation 5.4 by the equation

$$P V = \frac{2}{3} \left(\frac{3}{2} N k T \right)$$

This results in the Ideal Gas Law as shown in Equation 5.7.

(Equation 5.7)

$$P V = N k T$$

where

P = pressure the gas exerts (newtons/meters^2)
V = the volume of the gas (meters^3)
N = the number of gas molecules
k = Boltzmann's constant (1.38×10^{-23} J/K degree)
T = temperature (Kelvin)

Using ratio reasoning, we see that Equation 5.7 seems very reasonable. If we keep the volume of a gas constant, increasing the number of molecules of gas in the volume or increasing the temperature of the gas would be expected to increase the pressure. Likewise, if you want to keep the pressure and temperature constant, increasing the volume of the gas would require increasing the number of molecules of the gas filling that volume.

Period 5 Summary

5.1: Temperature is the average kinetic energy of the molecules in a substance. Molecules move in random directions with random velocities.

Thermal energy (or internal energy) is the kinetic energy involved in random molecular motion together with the potential energy stored in molecular attractions.

Heat is thermal energy in transit between two objects at different temperatures.

Evaporation cools a substance because the molecules with the most kinetic energy are most likely to leave (evaporate) from the surface of a substance. When the molecules with the most energy leave, the average kinetic energy of the remaining molecules goes down and the temperature is lower.

Diffusion occurs due to the random (Brownian) motion of molecules. Diffusion occurs more rapidly at higher temperatures because warm molecules have greater average kinetic energy.

5.2: Phase changes occur when matter changes from one state of matter (solid, liquid, or gas) to another state.

Heat capacity: The amount of heat needed to change an object's temperature by one Celsius degree. $Q = H_{cap}\ DT$

Specific heat**:** The amount of heat needed to change the temperature of one gram of a substance by one Celsius degree. $Q = s_{heat}\ M\ DT$

Latent heat: The amount of heat needed to change the phase of one gram of a substance. $Q = L_{heat}\ M$

5.3: The ideal gas law describes the relationship among the pressure, temperature, and volume of a gas and number of molecules in a gas: $P\,V = N\,k\,T$

Period 5 Exercises

E.1 An increase in the temperature of a solid usually

a) decreases the average molecular separation.
b) causes the molecules to melt.
c) increases the average molecular separation.
d) causes the electrons to transfer to lower energy levels.
e) NONE of the statements is correct.

E.2 Evaporation is a process

a) that increases the temperature of liquids.
b) where slow molecules increase their speed.
c) caused by cooling.
d) that results in a decrease of the temperature of liquids.
e) NONE of the statements is correct.

E.3 When water is cooled to form ice there is a decrease in

a) the kinetic energy of the molecules.
b) the latent heat of the water.
c) the intermolecular force.
d) molecular contraction.

E.4 When you transfer heat to a substance, you always increase its

a) latent heat.
b) specific heat.
c) temperature.
d) energy.

E.5 Brownian motion provided evidence for

a) electronic shells of atoms.
b) atomic weights of atoms.
c) molecular motion.
d) nuclear charges of atoms.

E.6 Container A contains air at a temperature of 100 ^{0} C and container B contains air at a temperature of 200 ^{0} C. Which of the following is true?

a) The air molecules in container A are moving faster, on average than those in container B.
b) The air molecules in container B are moving faster, on average than those in container A.
c) There is not enough information to say anything about the average molecular speeds.
d) The air molecules in both containers have the same average speed.

E.7 One can change a substance from a liquid to a solid by

a) removing thermal energy from the substance.
b) adding thermal energy to the substance.
c) adding the latent heat of vaporization to the substance.
d) adding the latent heat of fusion to the substance.

E.8 Absolute zero is

a) defined as zero degrees on the Kelvin scale of temperature.
b) the temperature at which all motion stops.
c) the temperature of liquid nitrogen.
d) defined as zero on the Celsius scale.
e) Both a) and b) are correct.

E.9 Consider two pails of water at the same temperature. Pail A contains 80 kg of water and Pail B contains 40 kg of water. Which one of the following statements is TRUE?

a) The water in pail A has a larger specific heat than the water in pail B.
b) The water in pail A has a greater thermal conductivity than the water in pail B.
c) The water in pail A has a greater heat capacity than the water in pail B.
d) The water in pail A has a smaller specific heat than the water in pail B.

Period 5 Review Questions

R.1 In class, we discussed convection, conduction, and diffusion. Which of these three processes can occur in solids, liquids and in gases?

R.2 "Wind chill" refers to the cooling effect of wind. Why does it feel colder on a windy day than it does on a calm day?

R.3 Why do car tires require less air in summer than in winter?

R.4 What does the specific heat of an object depend upon? The shape of the object? The material from which the object is made? The mass of the object? The temperature of the object?

R.5 Which contains more thermal energy – a cup full of hot coffee or a bathtub full of warm water? Why?