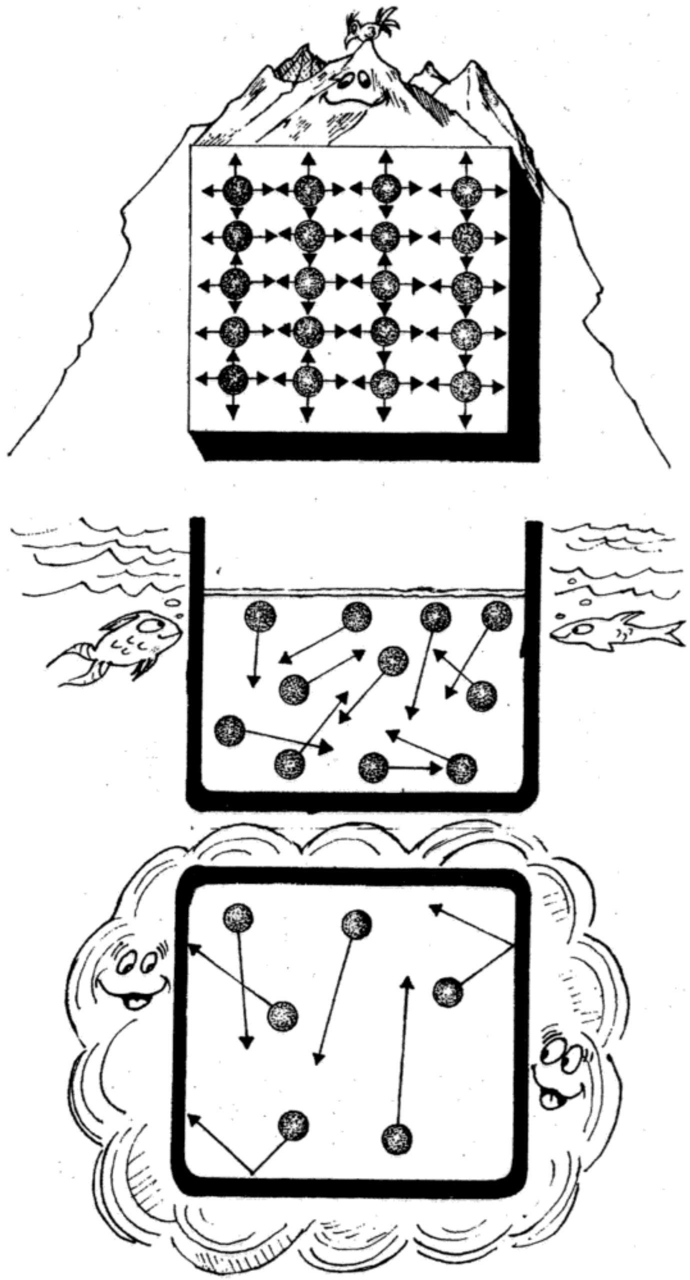
Internal and Thermal Energy

Matter in Motion



Student Booklet

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National Foundation for ENERGY EDUCATION

Composition of Matter

1. All matter is composed of extremely small atoms which are in constant motion.
2. Almost all of the atom's mass is concentrated in the center - the nucleus - of the atom.
3. The nucleus of an atom comprises an extremely small portion of the atom's volume.
4. Matter is basically empty space with large distances between neighboring nuclei.
5. The great distance between adjoining atoms results in extremely weak gravitational forces.
6. The space between matter in a neutron star is billions of times smaller than it is between matter on earth, therefore producing a much denser material.
7. The greater the charge on a particle and/or the shorter the distance between charged particles increases the force of electrostatic attraction or repulsion.
8. Atoms/molecules are held together by electrical forces of attraction between protons and electrons of adjoining atoms/molecules.

Structure of Matter

9. In a crystalline solid, the atoms are arranged in a repeating pattern.
10. The sharing of electrons between neighboring atoms brings the atoms closer together resulting in stronger electrostatic attraction and the formation of chemical covalent bond.
11. The forces of attraction between atoms in molecules are called intra-molecular forces.
12. The forces between atoms of neighboring molecules are called inter-molecular forces.
13. Intra-molecular forces are stronger than inter-molecular forces because the bond length for the intra-molecular bonds are shorter.
14. The transfer of electrons from one atom to another produces charged atoms (ions), and a strong ionic chemical bond.
15. With the strong force of attraction between oppositely charged ions, ionic solids have high melting and boiling points.
16. The more electrons that are transferred the greater the charge and the stronger the ionic bond.
17. In metals, the attractive forces are between the positively appearing charged metal atoms and a sea of free electrons.
18. The greater the density of electrons in the *sea*, the greater the force of the bond and strength of the metal.
19. Metals' attraction to a sea of electrons rather than neighboring atoms allows them to be made into foils as the atoms rearrange themselves rather than separate under pressure.
20. The distance between atoms in neighboring molecules is larger when compared to shorter distances between atoms within the actual molecule.
21. Molecules in molecular solids are held together by weaker intermolecular forces and are not chemically bonded together.
22. Molecular solids have low melting/boiling points because the intermolecular forces that keep them together are weak.

Matter in Motion

23. Atoms in solids can vibrate in three dimensions (x,y,z planes)
24. Atoms are pushed and pulled by electrostatic forces of attraction and repulsion.
25. The pushes and pulls experienced by atoms/molecules produce waves of motion in solids.
26. The levels of vibrational energy and direction of atoms/molecules are many, but not infinite, only certain values are allowed.
27. The type of atom and its crystalline structure determine the allowed vibrational energy levels of its atoms.
28. Increasing (decreasing) the temperature of a substance will decrease (increase) the force of attraction between atoms/molecules because of the greater (smaller) distances between them.
29. The kinetic energy of an object is determined by its mass and velocity.

30. Increasing (decreasing) the motion of an atom/molecule increases (decreases) its kinetic energy.
31. At the same level of vibration, the more massive atom/molecule contains more kinetic energy because of its larger mass.
32. The kinetic energy of an object's individual atoms/molecules varies greatly throughout the substance.
33. The average kinetic energy of all of a substance's atoms/molecules is known as its temperature.
34. Most of the atoms/molecules of a substance have kinetic energies near the average kinetic energy value.
35. Increasing (decreasing) the average kinetic energy of a substance's atoms/molecules will increase (decrease) its temperature.
36. The Fahrenheit scale assigns the number 32 to the freezing point of water and 212 to its boiling point.
37. The Celsius scale assigns the number 0 to the freezing point of water and 100 to its boiling point.
38. A Celsius degree measures almost twice (1.8) the temperature change than a degree of Fahrenheit change in temperature.
39. At -273 °C atomic motion is at a minimum (ground state); this is called absolute zero or 0 Kelvin.
40. The measurement of temperature change for Celsius is exactly the same as a Kelvin.
41. Water freezes at 273 K and boils at 373 K.
42. The internal energy of an object is the sum total of the kinetic and potential energies of all its atoms/molecules.
43. The internal energy of a substance is a result of both its temperature and mass.
44. Increasing (decreasing) the temperature and mass of a substance will increase (decrease) its internal energy.
45. The forces of attraction between atoms/molecules in liquids are weak; they are held together by only intermolecular forces.

Liquids and Gases

46. In liquids, atoms/molecules are free to move and spin.
47. Molecules in liquids are free to move and move with greater motion as the temperature increases.
48. The attraction between gas molecules is practically zero as the distance between gas molecules is approximately ten times more than in its liquid phase.
49. The volume of liquid increases between 250-1000 times when it is changed into a gas.

Chapter One: Atomic Structure, Internal Energy, and Temperature

One of the first associations children learn is the one between the phrase HOT! DON'T TOUCH! and the sensations of cold, warm, hot, and TOO hot to handle.

Our bodies experience thermal energy flows every second of the day and occasionally we're conscious of these flows. Perhaps it is when we sit on a hot car seat while wearing shorts. Ouch! Or when we jump into a swimming pool of very cold water. Brrrr! Or maybe it is when you take a bite *of fresh from the oven* pizza and it bums the roof of your mouth. Pizza Palate!

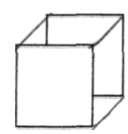
The flow of thermal energy is what makes refrigerators, automobiles, and cooling systems operate. It's also how our bodies maintain a core temperature of 37 degrees Celsius.

Size of Atoms

To understand internal energy and the flow of thermal energy we must first understand the structure of matter. Matter is made up of super small atoms and molecules that are in constant motion. How small are atoms and molecules? The number of molecules in one human breath is equal to the number of breathfuls in the entire earth's atmosphere.

In one cubic centimeter of air there are 3.0×10^{19} or, 30,000,000,000,000,000,000 air molecules.

In one cubic centimeter of aluminum there are 5.4×10^{23} or, 5,400,000,000,000,000,000,000,000 aluminum atoms.



Comparing the size of an *atom to a golf ball is like comparing a golf ball to the size of the earth.*

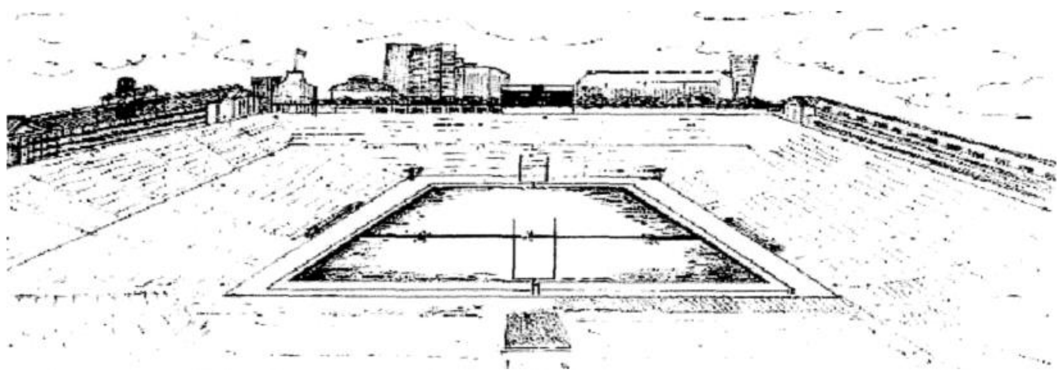
The Emptiness of Matter

Most an atom's mass is concentrated in its center, the nucleus. The nucleus of a carbon atom (with its six protons and six neutrons) contains 99.95% of the atom's mass. While the nucleus contains most of the atom's mass it only takes up one trillionth (1/1,000,000,000,000) of the atom's volume. The diagram of the carbon atom below is not drawn to scale. If it were drawn to scale the outermost electrons would be about 3000 meters away from the nucleus.

Carbon Nucleus **Outer Electron**

*If you placed a breadcrumb in the center of a large football stadium, the crumb would **represent the atom's** nucleus (having almost all the mass). The stadium's top row of seats would represent the atom's outer most electron.*

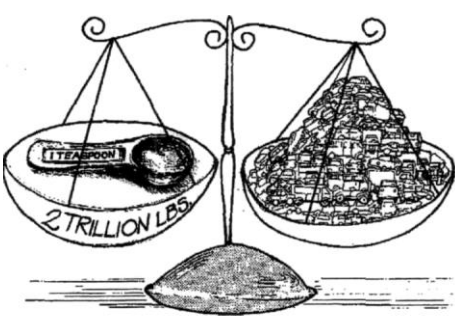
Atoms are basically made up of empty space, and when you consider the distances between the nuclei of neighboring atoms, the empty space becomes even greater. If your fist represents the nucleus of an atom, to scale, the nucleus of the next atom would be a fist almost a mile away. Matter makes up only one-trillionth (1/1,000,000,000,000) of a substance's volume, the rest is empty space.



If you could remove all the empty space inside and between all the atoms in the earth, you could fit them inside a ball with a diameter of 370 meters That's a little bigger than the size of a very large domed stadium.

This vast amount of empty space inside and between atoms of matter we see on earth is very different from the matter that makes up **neutron stars**. Very little empty space can be found between the particles in a neutron star. A teaspoon of matter from a neutron star would equal the weight of six hundred million (600,000,000) automobiles. The density of matter from a neutron star is thought be two hundred trillion (200,000,000,000,000) times more than that of water and its gravitational forces a trillion times stronger (1,000,000,000,000).

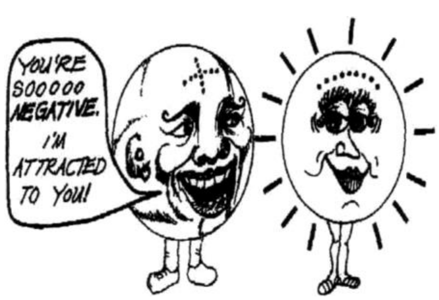
With this vast amount of empty space between the massive nuclei of atoms on earth, the resulting gravitational attraction between nuclei is extremely small. If it isn't gravitational forces that keep solids and liquids together, then what does?



Structure of Solids

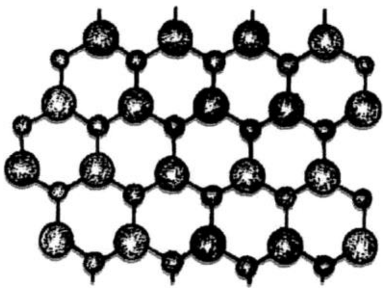
Electrostatic forces, not gravitational forces, are what keeps matter together. The positively-charged protons in an atom's nucleus are attracted to the negatively charged electrons of neighboring atoms, and the first atom's electrons are attracted to neighboring atom's protons. In solids this attraction keeps atoms in fixed positions, vibrating in three dimensions (x, y and z planes).

Many solids' atomic structure is made up of atoms that are attracted to each other to form a repeating pattern. These solids are called **crystalline solids**. Solids that do not have an organized repeating pattern are called **amorphous solids** (rubber and plastics). Crystalline solids will be examined in this treatment of internal and thermal energy.

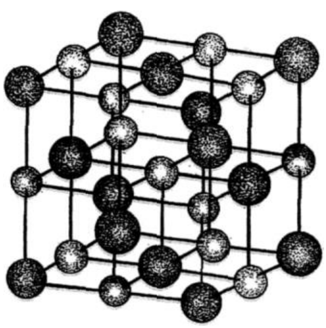


Some crystalline solids, like diamonds or quartz, have atoms that share electrons with neighboring atoms. Each of these neighboring atoms, in turn, shares electrons with still other neighbors. The sharing of electrons brings atoms closer together thus increasing the electrostatic force of attraction between atoms. The sharing of electrons produces a chemical bond called a **covalent bond.** The sharing continues countless times and in the same pattern resulting in a super structure of atoms - a covalent solid.

The patterns of the atoms and molecules in a crystalline solid repeat themselves in an orderly fashion. The type of atoms or molecules determines the shape of the pattern. These patterns are not always perfect, crystals have flaws and impurities.



Quartz Crystal – SiO_2
Covalent Bond



Sodium Chloride Crystal
NaCl - Ionic Bond

Ice Crystal - H_2O
Molecular Bond

In other crystalline solids there is a transfer of electrons from one atom to one or more neighboring atoms. The result is atoms having a positive or negative charge. Some atoms release electrons and then have more protons than electrons; these atoms have a positive charge. The atoms accepting the electrons now have an excess of electrons and have a negative charge.

The oppositely charged atoms, which are called *ions*, produce a strong force of attraction and a chemical bond called an **ionic bond**. An ionic solid is hard with high melting and boiling points. The more atoms transferred between atoms the stronger the force of attraction and the higher the melting and boiling points.

The crystalline structure for metals is similar to that of covalently bonded materials because of the sharing of electrons. The difference here is that the atoms don't always share the same exact electrons. In metals, electrons are free to move. They don't have a place to call home - they travel from atom to atom.

With this free movement of electrons, the metal atom imitates a positively charged. atom, a positive ion. However, this time the charged atom isn't attracted to a neighboring charged atom, but rather to a sea of free negatively charged electrons. Hence, when most metals are pounded they form sheets rather than a powder. In metals, the pounding force doesn't separate the atoms, it just rearranges them. This is why metals can be made into thin sheets and foils.

Molecular solids have the weakest crystalline structure. There is no sharing or transfer of electrons. With a greater distance between neighboring molecules, and a neutral electrical charge, a weaker force of attraction is present. Molecularly bonded materials are soft and have low melting and boiling points. Most gases and liquids when cooled enough will form molecular solids.

Substance	Bonding	Melting Point	Boiling Point
Diamond	Covalent	3,550°C	4,827°C
Quartz	Covalent	l,710°C	2,230°C
Sodium Chloride	Ionic	801°C	1,465°C
Calcium Fluoride	Ionic	1,423°C	2,500°C
Aluminum	Metallic	660°C	2,470°C
Copper	Metallic	1,085°C	2,927°C
Water	Molecular	0°C	100°C
Fluorine	Molecular	-220°C	-188°C

Ice is an example of a molecular bonded solid. The hydrogen atoms in the H_2O molecule are attracted to a neighboring molecule's oxygen atom, but they are not chemically bonded to the atom. There is no sharing or transfer of electrons, just a weak electrostatic force of attraction between molecules known as an intermolecular force.

The chemically bonded hydrogen atoms are almost twice as close to their oxygen atom than the neighboring oxygen atom. The smaller distance results in a stronger force of attraction for the atoms in the H_2O molecule, 25 times stronger. The force of attraction between atoms within a molecule is called an **intramolecular** force. In molecular solids the intra molecular forces (chemical bond) are 20-200 times stronger than intermolecular force that keeps the crystalline structure together.

Kinetic Energy

A moving object possesses kinetic energy. A baseball, for example, if put into motion, would contain kinetic energy. If the ball were dropped from ten centimeters above your head it would possess little kinetic energy. However, dropping the ball from a distance of 100 meters would result in a lot more kinetic energy when it strikes your head. **Velocity affects** the **kinetic energy of an object.**

Drop a sixteen-pound bowling ball instead of a baseball from the ten centimeters mark above your head and you will experience a lot more energy, as well as a big bump on the head. However, having oxygen molecules striking your head at 1000 miles per hour would have no noticeable effect on you. **Mass affects the kinetic energy of an object.**

The mass and velocity of atoms/molecules also determine their kinetic energy. In solids, atoms and molecules can vibrate, jiggle and wiggle in many directions, but they are not free to move from their positions in the crystal. Atoms in solids are pushed and pulled by their neighboring atoms.

These pushes and pulls are caused by atomic electrical forces of attraction and repulsion. Imagine atomic springs between these atoms and the pushes and pulls are like compressing or expanding the spring. The energy levels of these vibrations are many, but not infinite. Only certain vibrational energy levels are allowed, and these levels differ because of the crystalline stricture of the substance.

The pushes and pulls are not random, they are passed from atom to atom in an organized flow or rhythm. They travel in a wave motion passing energy from atom to atom in a more orderly fashion. It's like pushing one domino and having thousands more dominoes fall in a line. However, atoms bounce back and produce waves in other directions.

In a piece of aluminum, atoms at room temperature are vibrating trillions of times per second with an average velocity reaching almost 1000 miles per hour. If we had atoms of gold vibrating at the same rate as the aluminum atoms, the gold atoms would have greater kinetic energy because gold atoms are several times more massive than aluminum atoms. *More mass means more molecular kinetic energy.*

Average Molecular Kinetic Energy and Temperature

In a sample of a substance, some atoms/molecules have a great deal of kinetic energy, while others possess little. Most of the atoms/molecules will have kinetic energies somewhere in between. If we could add up the energies of all of a substance's atoms and divide this sum by the total number of atoms, we could calculate the average kinetic energy of the atoms. This average kinetic energy of a substance's atoms is called the **temperature** of the substance.

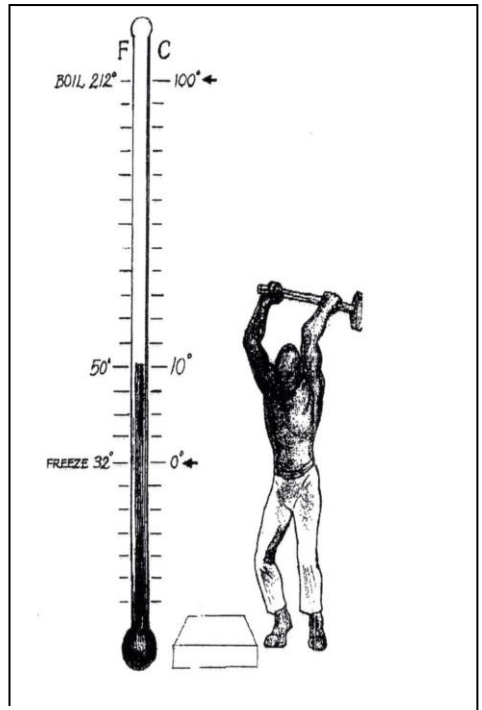
An increase in the average kinetic energy of a substance's atoms/molecules produces an increase in its temperature. A decrease in the average kinetic energy per atom/molecule produces a decrease in the temperature.

Temperature Scales

To measure the temperature of a substance, or the average kinetic energy of its atoms/molecules, we use a thermometer and two temperature scales. The most widely used scale worldwide is the Celsius scale. The freezing point of water is assigned the number 0 and the boiling point of water is assigned the number 100. Divide this difference by 100 units and you will get one Celsius degree.

Fahrenheit/Celsius Scale **Comparisons**

	⁰F	⁰C
Boiling Water	212	100
Hot Soup	150	65
Body Temperature	99	37
Cold Soda	35	2
Room Temperature	72	22
Melting Water	32	0
Ice Cube	5	-15
Scales Coincide	-40	-40
South Pole	-120	-84
Liquid Nitrogen	-320	-196
Absolute Zero	-458	-273



The United States is one of the few countries which still uses the Fahrenheit scale. Water freezes at 32 degrees and boils at 212 degrees Fahrenheit. There are 180 degrees, not 100 degrees, between the freezing and boiling points on the Fahrenheit scale. This means that one-degree Fahrenheit is a 44 percent smaller measure of temperature change than one degree Celsius. It takes 1.8 of the smaller Fahrenheit degree to equal the larger Celsius degree. A temperature increase of l8’F is equal to l0’C increase.

Let’s use an analogy of a child and adult walking 20 meters to show the differing size of a Celsius and Fahrenheit degree.

The child and adult both travel the same distance (temperature), however, because the child has a shorter stride the child must take more steps (size of degree).

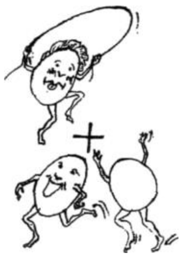
Absolute Zero

The coldest temperature that can be reached is a temperature of -273.l5’C, or absolute zero. Scientists agree you can’t go any lower. It’s not because of the equipment, it’s the laws of physics. Does this mean that the atoms are at rest at -273.15°C? No. It means that the kinetic energy of the atom is at a minimum. No more energy can be removed, no matter how hard you try.

Scientists have come within several billionths of a degree Celsius of absolute zero. Then why not to zero? To remove internal energy from an object it must flow to a cooler object. What can be colder than absolute zero? Nothing, therefore, there can be no flow of energy to a colder object.

Absolute zero is given a value of zero Kelvin (0 K). *A Kelvin is equal in size to a Celsius degree.* Starting the Kelvin scale at absolute zero (-273°C) means that the Kelvin scale has no negative temperatures. A temperature of 0 K is equivalent to - 273°C. Water freezes at 273 K or 0°C, and boils at 373 K and l00⁰C. Kelvin temperatures do not have the (°) degree symbol. Scientists decided that since the Kelvin scale is really a measure of kinetic energy it should not have a degree symbol. Earlier scientists based their scales on the freezing and boiling points of water. Then they divided this range by 100 or 180, into degrees.

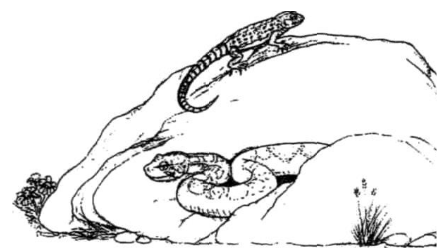
Internal Energy



By adding the kinetic and potential energies of each individual atom or molecule, you can determine the internal energy of an object.

The internal energy of a substance is the sum of all the atoms'/molecules' kinetic energies and potential energies; later on we will examine the potential energy of atoms. You can increase the internal energy of an object by causing its atoms to vibrate with greater motion. A penny has more internal energy at 100^0C than it does at 50°C. Having more mass also increases internal energy. Two pennies at the same temperature would contain twice as much internal energy as one because they would contain twice as many vibrating atoms.

The next time you look at a large rock on the ground, remember it has energy. It is not kinetic energy because the rock isn't in motion. It is not potential energy because it has no height above the ground. It is internal energy because its atoms and molecules are in motion, and they can provide a thermal energy flow to a cooler object.

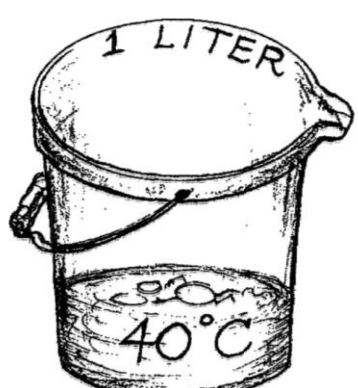


Temperature Versus Internal Energy

The internal energy of an object is the **sum total** of the kinetic energies (and potential) of all of its atoms/molecules. The temperature of the object is the average kinetic energy of its atoms/molecules. Temperature and internal energy are related but have different measures.

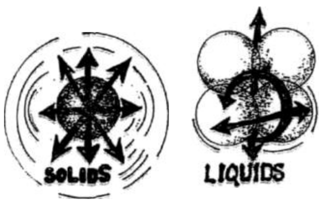
Let us examine the difference by looking at two buckets containing water at 40°C. The first bucket contains 1 liter of water; the second bucket contains 4 liters of water. The average kinetic energies of the water molecules in both buckets are the same; they are both 40^0C. However, the second bucket has four times more internal energy because it has four times the number of water molecules. If both buckets had the same amount of water, then the bucket with the hotter water would have more internal energy.

Touching a crumb-size piece of hot metal at 100°C *for ten seconds would* hardly *affect you as it contains very little internal energy. However, touching a 5-kilogram slab of the same metal for ten seconds at a lower temperature of 65°C would be very painful; it has a great deal of internal energy.*



The Atomic Structure of Liquids

Liquids also have forces of attraction which keep neighboring molecules together. They are the weaker intermolecular forces not intramolecular. When two drops of water touch, they combine to form a larger drop because of the intermolecular forces of attraction between the molecules. The positive sections of water molecules are attracted to the negative sections of neighboring molecules. Ice, on the other hand, since its molecules are in fixed positions, has strong attractive forces and can support heavy loads.



Atoms/molecules in solids are in fixed positions and can only vibrate. In liquids, atoms/molecules are free to move throughout the liquid, as well as, have the ability to spin.

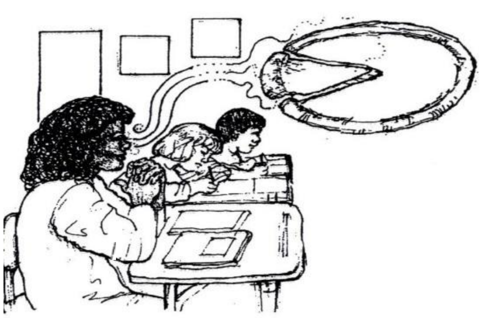
As a result of their weaker forces of attraction, atoms and molecules in liquids are free to move throughout the substance. A molecule at the bottom of a container may move to the top, middle, or side of the container. Molecules are not in fixed positions in a liquid because they possess more potential energy than they would in their solid state. This potential energy is the energy that helps them overcome the forces of attraction that would have them in fixed position as in the solid phase. Not only are the liquid's molecules free to move, they can also spin.

Adding energy to a liquid also weakens the forces of attraction between its molecules. Notice the forces of attraction for vegetable oil. When you pour cold oil into a cold pan, the oil moves very slowly and looks thick. As the oil heats up, it moves more easily and looks thinner. The same is true for pancake syrup. Taken right from the refrigerator it is thick and slow. Heat the syrup, and it flows with ease. The increased kinetic energy of the molecules decreases the forces of attraction between them.

The Atomic Structure of Gases

The energy possessed by gas molecules is extremely great. Gas molecules, like liquids, are free to move. Gases will entirely fill a closed container while liquids will only do so if you fill the container to the top. Gases move with much greater velocity and spin much faster than the molecules in a liquid. The distance between molecules in a gas is also great. The molecules in gases are about ten times farther apart from each other than they are in liquids, and take up 250-1,000 times more volume. Gas molecules at room temperature are not very *sticky* and are far apart.

With the distances between gas molecules so great , there is practically no electrostatic force of attraction between them. Very low temperatures are needed to liquify gases like oxygen (-182° C) and nitrogen (-196° C).



Your sense of smell can detect this speed and free movement of gas molecules. If a class was having a pizza party down the hall, you could smell the pizza in your room. It may take several minutes for the aroma to reach your nose, even though the gas molecules are traveling at speeds above 1,000 miles per hour. The reason for the delay is that the gas molecules are striking air molecules at about six million times per second, and therefore don't travel in a straight line from the pizza box to your nose. Traveling ten meters would take months to travel and cover millions of miles. The fact that you can quickly detect the smell so quickly is due to the help given to the molecules by drafts or convection currents which move whole chunks of air rapidly from place to place.

Expansion and Contraction of Matter

1. An increase (decrease) in the temperature of a substance will decrease (increase) the forces of attraction between its atoms/molecules, usually resulting in an increase (decrease) in the distance between the atoms/molecules.

2. Substances expand and contract at different rates because of the strength of the forces of attraction between their atoms/molecules.
3. The rate at which the volume of a substance increases or decreases per degree Celsius is known as its rate of cubic expansion.
4. Liquids, with their weaker forces of attraction, expand/contract at greater rates than solids.
5. Gases have practically no force of attraction between them and expand most when heated and contract most when cooled.
6. To calculate the increase (decrease) of a substance's volume because of an increase (decrease) in temperature, multiply the original volume of the substance, by the coefficient of cubic expansion of the substance, then by the temperature change. (rise = + drop = -)
7. Water is an unusual substance; at 4^0C it expands if its temperature rises or falls; water is densest at 4 °C.
8. Thanks to water's unique properties, bodies of water freeze from top to bottom, not bottom to top.

Heat Transfer

9. The transfer of internal energy between objects is called heat or thermal energy.
10. Thermal energy travels from the substance with the higher temperature to the substance with the lower temperature.
11. At the atomic level some energy does flow from cold to hot - the net flow is from hot to cold.
12. When the net flow of thermal energy ceases between objects they are said to be in thermal equilibrium.
13. A linear relationship produces a straight line on a coordinate graph, with data points from the investigation located on this line.
14. A nonlinear relationship produces a curve with a changing slope for each data point.
15. The slope of a Temperature/Time line tells us the rate of temperature change and the direction of that change.
16. The rate of cooling (warming) of a substance depends on how much hotter (cooler) it is than its surroundings; the greater the difference, the greater the temperature decrease (increase) in any given time period.
17. The amount of cooling (warming) of a substance over time is a nonlinear relationship.
18. The greatest rate of change for cooling (warming) occurs during the first-time interval, decreases during subsequent intervals
19. Thermal energy can be transferred through direct contact by the process of conduction.
20. Conduction transfers kinetic energy through the collisions of atoms/molecules to neighboring atoms/molecules.
21. In metals, free electrons also cause the transfer of thermal energy as they collide with atoms/molecules and other free electrons.
22. Insulators are poor conductors of thermal energy because of the poor transfer of kinetic energy between neighboring atoms/molecules.
23. Air and other gases are poor conductors of thermal energy because they have large distances between neighboring molecules.
24. Wool, foam polystyrene, and feathers contain many air spaces and are good insulators.
25. In liquids and gases, thermal energy is transferred by convection, the movement of its atoms/molecules throughout the sample.
26. Increasing the kinetic energy (temperature) of atoms/molecules in liquids and gases will increase distances between the atoms/molecules and lower densities.
27. The more energetic atoms/molecules in a liquid or gas will rise to the top of the sample; the less energetic atoms/molecules will fall to the bottom of the sample.
28. Matter is not always needed to transfer energy, as radiant energy can travel through a vacuum.
29. All objects emit and absorb radiant energy.
30. Good (poor) emitters of radiant energy are also good (poor) absorbers or radiant energy.

Chapter Two: Thermal Expansion, Equilibrium, and Transfer

We learned earlier that increasing the motion of atoms/molecules decreases the force of attraction between them. This decrease in attractive forces usually causes a greater distance between neighboring atoms/molecules. Lowering the temperature increases the attractive forces and decreases the distances between atoms/molecules. This is the reason that objects usually expand when heated and contract when cooled.

We also learned that the forces of attraction are greatest for solids, followed by liquids and then least for gases. When thermal energy is added, the lack of attractive forces in gases will result in a great deal of expansion. The rate of expansion for liquids is less, and the expansion rate for solids is least.

Rate of Cubic Expansion of Some Substances at 20 ^{0}C

Substance	per ^{0}C
Air and Most Gases	$3{,}400 \times 10^{-6}$
Ethyl Alcohol	$1{,}120 \times 10^{-6}$
Gasoline	950×10^{-6}
Glycerin	500×10^{-6}
Water	210×10^{-6}
Mercury	180×10^{-6}
Ice	150×10^{-6}
Lead	87×10^{-6}
Aluminum	75×10^{-6}
Brass	56×10^{-6}
Copper	51×10^{-6}
Nickel	39×10^{-6}
Iron or Steel	35×10^{-6}
Concrete or Brick	36×10^{-6}
Glass	26×10^{-6}
Pyrex Glass	10×10^{-6}
Diamond	4×10^{-6}

The chart to the left shows the rate at which different substances expand and contract. Notice how air expands 17 times more than water, and almost 95 times more than steel when undergoing equal temperature changes.

A substance's expansion rate must be taken into account when building a bridge, filling a tooth, or checking your tire pressure on a very hot day. A bridge may undergo expansion or contraction of several inches from the hot days of summer to the ice-cold days of winter. You may have noticed the metal teeth of an expansion joint placed somewhere along the road surface of a bridge; this joint was installed to accommodate the effects of changes in weather.

If a dentist selected a substance which had a large expansion rate to fill a tooth cavity the patient would be in trouble. Hot soup and ice cream would cause tooth damage and loose fillings.

Calculating Thermal Expansion and Contraction

To calculate the cubic expansion or contraction of an object, multiply the initial volume of the object by the temperature increase/decrease (+/-) of the object, then multiply this product by the expansion rate for that substance.

Expansion = (Initial Volume) • (Temp. Change +/-) » (Expansion Rate)

A 1000 ml (1 liter) container of air undergoes a temperature change increase of 100^{0}C in an expandable container. How much did the container increase in volume? What is the resulting size of the container?

Initial Volume = 1000 ml
Temperature Change = +100^{0}C
Expansion Rate: 3,400 x 10^{-6}/C

Expansion = $(1.0 \times 10^3$ ml$)$ $(1.0 \times 10^2$ °C$)$ $(3.4 \times 10^{-3}$/°C$)$
Expansion = 3.4×10^2 ml or 340 ml

The Uniqueness of Water a 4^0C

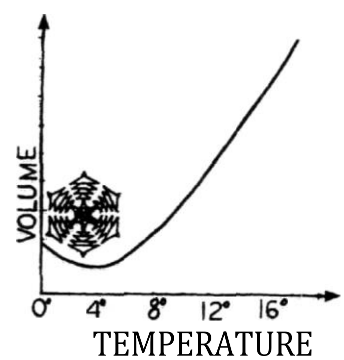
While almost all substances expand when heated and contract when cooled, water at 4^0C will expand if heated or cooled. Water is a unique substance. Like other substances, water at 4^0C expands when heated because of the increase in the motion of its molecules and the decrease in the forces of attraction.

Water molecules are free to move.

When the temperature of water starts to drop below 4°C its molecules slow down and the forces of attraction between water molecules become stronger. As these attractive forces get stronger, the oxygen atom from one water molecule forms an intermolecular bond with the hydrogen atom of a neighboring molecule to form ice crystals. The molecules are now in fixed positions. The ice formation takes up more space than the free moving molecules at higher temperatures.
This causes water to expand.

Ice crystal molecules are in fixed positions that consume more space than an equal number ofwater molecules.

When the water has been changed from water at 4°C into ice at 0'C, its volume will increase by 9 percent. However, just like other solids, ice will decrease in volume as its temperature drops below 0'C.



Water is densest at 4'C

It is fortunate that water has this unique property. If it didn't, bodies of water would freeze from the bottoms up. In moderate to deep bodies of water, the sun's energy could never melt the ice at the bottom. This would kill the organisms at the bottom. Lucky for us this doesn't happen.

Thermal Energy Vs. Internal Energy

The internal energy of a substance is the sum total of the kinetic and potential energies of all of its atoms. Placing two objects of different temperatures together will produce a transfer of **internal energy** from the hotter object to the cooler object. This transfer of internal energy is called **thermal energy.**

The motion of the atoms/molecules in the hotter object will decrease as they transfer their kinetic energy to the atoms/molecules in the cooler object. This transfer will produce an increase in the kinetic energy in the cooler object.

Temperature is the average kinetic energy of the substance's atoms7molecules -- some atoms are very energetic -- some aren't. At the atomic level, some energy does flow from the cooler object to the hotter object. The cooler substance's *more* energetic atoms/molecules can transfer energy to the hotter substance's *less* energetic atoms/molecules. However, the net flow is from warmer to cooler. If two objects at the same temperature are placed together so that they touch, there will be energy transfer at the atomic level; there's no net transfer of internal energy, and no thermal energy produced.

Thermal energy always travels from the object at a higher temperature to the object at a lower temperature.

Thermal Equilibrium

Initially, the more energetic water mole

To explain the meaning of *thermal equilibrium*, let's start with an example. We place a thermometer at room temperature of 22^0C into a beaker containing one liter of water at 75^0C. The water molecules have a higher average kinetic energy (temperature) than the thermometer's glass bulb and alcohol indicator.

The water molecules begin to transfer some of their energy to the thermometer. Eventually the kinetic energy of the molecules in the thermometer increases as its molecules vibrate with increased motion. Its temperature rises.

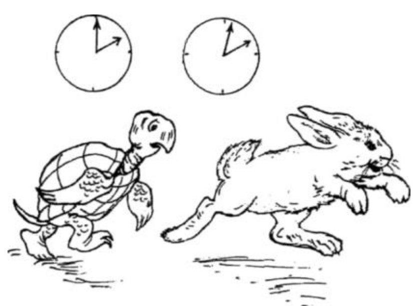
The average kinetic energy of the water molecules decreases slightly. They only decrease slightly because the mass of the water and its internal energy are so much greater than the thermometer. Eventually the temperature of the thermometer and the water will be the same. When there is no longer any net flow of energy between the two substances they are said to be in *thermal equilibrium*.

This means that an aluminum block placed in a beaker of boiling water (100^OC) will therefore eventually reach thermal equilibrium with the water which equals a temperature of 100^0C. Its also means that when the aluminum block is removed from the water, it will eventually reach equilibrium with the air and will be at room temperature.

When a warmer substance transfers its internal energy to a cooler substance to reach equilibrium, the amount of energy lost by the warmer substance is gained by the cooler substance. Place a 100-gram piece of copper at 100^0C in 500 grams of water at 20°C. When the water and the copper reach thermal equilibrium, the energy lost by the copper will be the same amount of energy gained by the water.

Energy Lost = Energy Gained

Newton's Law of Cooling



The rate at which a substance loses or gains thermal energy depends on the difference between its temperature and the temperature of its environment. The greater the difference, the greater the rate of change in temperature.

A metal block at $l00^0$C will undergo a greater drop in temperature during the first minute when it is placed in room temperature air. For example, it may cool by 30 degrees in the first minute, 8 degrees in the second minute, 5 degrees in the third minute and so on, until it reaches thermal equilibrium with the room temperature air.

The same is true for heating. A metal block at - 50'C placed at room temperature will heat faster than a block at 0^0C.

Newton's Law also affects the heating and cooling of homes. The warmer the interior of the house compared to the outside temperature, the greater the rate of heat transfer from the house to the outside air. The opposite is also true for air conditioning when heat enters the house. Keeping your home a little cooler in the winter and warmer in the summer saves energy and heating/cooling dollars.

Methods of Heat Transfer

Thermal Conductivity

Substance	ca1/(sec-meter-^{0}C)
Silver	100
Copper	92
Aluminum	50
Iron	16
Ice	0.50
Glass	0.20
Concrete or Brick	0.20
Water	0.140
Wood	0.02-0.04
Fiber Glass Insulation	0.012
Polyurethane Foam	0.010
Wool	0.010
Goose Down	0.006
Air and Most Gases	0.0055

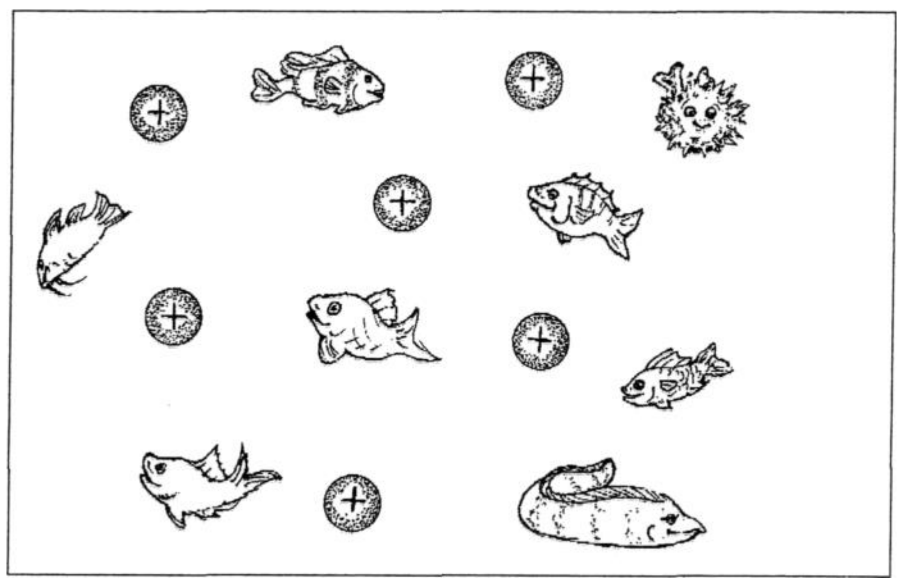
Conduction
When the more energetic atoms/molecules of one substance are put in direct contact with the less energetic atoms/molecules of a second substance, there is a flow of thermal energy. This transfer of thermal energy by direct contact is called **conduction.**

Remember the net flow of thermal energy is from warmer to cooler substances. There is no such thing as cold energy. A cold substance's atoms/molecules have less kinetic energy than those in your hand. When you touch the cold object there is a loss of energy from your hand and a signal is sent to your brain with the message that the substance is cold. The opposite is also true. Grasp a hot object and there will be a flow of thermal energy into your skin's molecules. The message: it's hot.

Free Electrons
While the vibration of neighboring molecules is the way thermal energy is transferred by conduction, in metals the transfer is mostly the result of the movement of free and excited electrons. Unlike other solids, metals such as gold, copper, silver and aluminum contain many electrons free to move among atoms without having to be the property of just one atom.

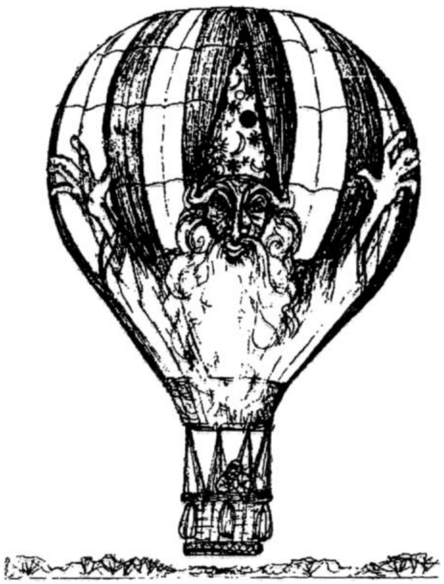
When these free electrons are excited by the addition of thermal energy, they transfer their kinetic energy to other electrons and neighboring atoms. These free electrons are also the reason that metals are good conductors of electricity.

Metals contain a sea of free electrons which are not bound to any one particular atom.



A metal spoon has so many more free electrons, it heats up faster in a hot liquid than a glass or plastic spoon. Substances that transfer thermal energy well are called **conductors.**

Some substances' atoms/molecules transfer their kinetic energy to neighboring atoms/molecules very slowly; these materials are called **insulators.** The mass, size and structures of the atoms/molecules determine the ability to transfer energy to neighboring atoms/molecules. You would drop a metal cup which was just filled with boiling water. However, you wouldn't feel any transfer of energy or pain while holding a wooden cup of the same thickness filled with boiling water.



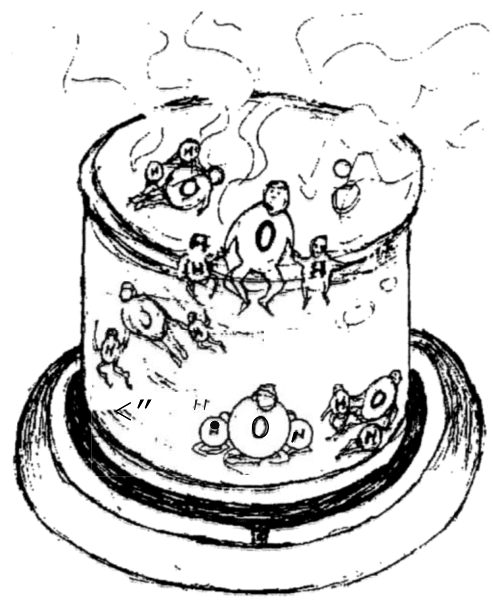
A hot air balloon rises because its gas molecules have more energy and greater distances between them than the surrounding air.

Convection

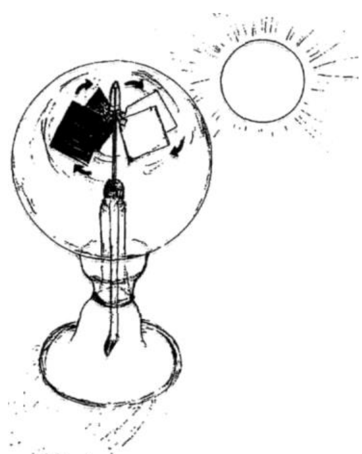
The atoms/molecules in liquids and gases are not held in fixed positions and are free to move. This ability to move is the way liquids and gases transfer their energy. The method of thermal energy transfer by the movement of its atoms/molecules is called **convection.**

When an atom/molecule receives a push or pull from a neighboring atom/molecule its kinetic energy increases and the distances between the atoms/molecules increases. This increase in distance makes these atoms/molecules less dense and they rise. Just imagine holding a wooden block in the bottom of a tank of water. Because the wood is less dense than the water, it rises. The same is true for the countless atoms/molecules in a section of a liquid or a gas.

The opposite is also true. Reduce the kinetic energy of atoms/molecules and you will increase the forces of attraction and decrease the distances between the atoms/molecules. The atoms/molecules become more dense and sink.



Our homes are heated by convection as the warm air from the registers in the floor rises and circulates throughout the room. Convection is also the way thermal energy is transferred throughout a pot of water on a stove; the more energetic water molecules rise and are replaced by less energetic molecules.



Radiant energy heats the black side of the radiometer's vanes more than the white side. Air molecules rebound off of the black and white sides of the vanes. The molecules rebound with more energy off the black side. For every action there is an ***opposite*** *and equal reaction.* ***The*** *vanes spin in the opposite direction to the black side of the* vanes.

Radiation

The third method of heat transfer, **radiation,** does not require the vibration or movement of matter to produce the transfer of thermal energy. The sun transfers its thermal energy to the earth 93 million miles away through space which contains virtually no matter.

Radiant energy is transferred by waves. All objects continually absorb and emit radiant energy. We cannot see the radiation emitted by everyday objects because the waves they emit are infrared waves. Unlike the visible light waves coming from the sun or a lamp, *infrared waves are invisible*. By using *night goggles,* we can see objects radiating thermal energy.

If an object is hot enough, some of the radiant energy it emits can be seen. A piece of iron placed in a blacksmith's fire at 500 ^{0}C will emit red light. As the object's temperature increases more wave lengths of light can be seen. At 1200 ^{0}C an object will emit white light — a mixture of all seven colors of the spectrum.

All objects absorb and emit radiant energy. Good absorbers are also good emitters of radiation. A black object will absorb radiation better than a white or shiny object, black objects will also emit thermal radiation better. The type of atom and its structure determine the frequencies of light that can be absorbed/emitted by a substance. Green objects are green because their atoms cannot absorb green light.

Specific Heat and Heat Capacity

1. An atom's/molecule's mass and structure determine how much molecular kinetic energy must be added or removed from an object to produce a given temperature change.
2. The amount of heat energy required to raise one gram of a substance 1^0C is called the **specific heat of the substance.**
3. Heat is measured most commonly in calories; one calorie can raise one gram of water one degree Celsius.
4. Heat is also measured in Kilocalories (raising one kilogram of water one degree Celsius), or joules (1 calorie = 4.184 joules), and the British Thermal Unit, Btu, (raising one pound of water one degree Fahrenheit).
5. When two substances at different temperatures are allowed to reach thermal equilibrium in a closed system, the energy gained by the cooler substance will be energy lost by the hotter substance
6. An object's mass and temperature change will determine how much thermal energy it loses or gains.
7. To calculate the heat (Q) needed to raise – lower - the temperature (ΔT) of a substance a given amount, multiply the mass (m) of the substance, times the specific heat (c) of the substance, times the desired temperature change (+ / -). ($Q = mc\Delta T$)
8. Substances with high heat capacities make great thermal storage sources.

Changes in Phase

9. At the boiling point atoms/molecules can exist in either the liquid or gaseous phases.
10. At the melting point atoms/molecules can exist in either the liquid or solid phases.
11. Kinetic energy must be absorbed to change a liquid into a gas at its boiling point.
12. Kinetic energy must be absorbed to change a solid into a liquid at its melting point.
13. Evaporation occurs when the most energetic atoms/molecules at the surface receive sufficient energy to escape the forces of attraction of neighboring atoms/molecules (kinetic energy is now potential energy) and become atoms/molecules in the gaseous phase.
14. Evaporation is a cooling process resulting from the most energetic atoms/molecules escaping the liquid, thereby lowering the average kinetic energy (temperature) of the remaining atoms/molecules.
15. Condensation occurs when atoms/molecules in the gaseous state lose so much of their potential (hidden) energy that they are changed into a liquid phase.
16. Condensation is a warming process resulting from potential energy transferred from gaseous atoms/molecules into kinetic energy on the surface of an object they strike which changes them to liquid.
17. The amount of thermal energy needed to change one gram of a substance from its liquid phase into its gaseous phase at its boiling point is called its **heat of vaporization**.
18. The amount of thermal energy needed to change one grain of a substance at its melting point from its solid phase into its liquid phase is called its heat of fusion.
19. To calculate the amount of heat (Q) required to produce a total phase change of a substance at its melting/boiling point, multiply the mass (m) of the substance by its heat of fusion/vaporization (H_F/H_V)
    $Q = m\ H_F/Hv$

Refrigeration Cycle

20. Cooling technologies have changed the way people live, work, and spend their leisure time.
21. The refrigeration cycle is made up of four components. In the expansion device the pressure is reduced; in the evaporator the liquid refrigerant changes into gas; in the compressor the gas is compressed and heated; and in the condenser the gas is changed back into a liquid refrigerant.
22. During the refrigeration cycle, the refrigerant is changed from a liquid into a gas, and then back into a liquid, as a result of thermal energy being added or removed.

Chapter Three: Specific Heat and Phase Changes

Specific Heat

The amount of thermal energy flow needed to produce a 1 ^{0}C temperature change in one gram of a substance is called its specific heat. Metals have low specific heats; a little bit of thermal energy flow can produce a 1 ^{0}C temperature change. Water has a very high specific heat and requires a great deal of thermal energy flow to produce a 1 ^{0}C temperature change.

The amount of thermal energy flow needed to ***increase or decrease*** *the temperature of a substance varies according to the atomic/molecular mass and structure of the substance. Gold, for example, a metal with many free electrons,* ***requires*** *33* ***times*** *less energy than water to increase its temperature by 1 ^{0}C.*

Specific Heat of Some Substances at 22 ^{0}C

Substance	(J/g-^{0}C)
Hydrogen	14.285
Helium	5.185
Ammonia	4.704
Water	4.182
Human Body	0.830
Ethyl Alcohol	2.450
Steam	2.007
Ice(-15'C)	2.000
Wood	1.672
Air	1.003
Aluminum	0.899
Sand	0.836
Pyrex Glass	0.786
Iron or Steel	0.460
Copper	0.385
Silver	0.234
Mercury	0.138
Gold	0.125

We use substances with low specific heat (such as metals) when we want a quick increase in temperature. Metal pots and pans heat quickly. While metals heat quickly they also cool quickly because they have a low heat capacity.

At other times we want substances with high specific heats like water or concrete for heat storage. They require a large amount of heat to increase their temperatures; however, they retain heat a lot longer. The atoms/molecules mass, size, and structure are factors determining the specific heat of a substance.

One common unit used to measure a quantity of thermal energy is the *Joule.* A total of 4.182 Joules (J) can raise one gram of water 1 ^{0}C. To raise the temperature of 10 grams of water 1 ^{0}C will require 41.82 J. The same amount of thermal energy (41.82 J) will raise 20 grams of water 0.5 °C; or l gram of water 10 ^{0}C; or 0.5 grams of water 20 ^{0}C.

The table to the left shows the amount of thermal energy in Joules needed to produce a 1°C temperature change of one gram of the substance.

(continued on page 19)

Calculating Thermal Energy Transfers

To calculate the amount of energy flow into or out of an object in calories, use the following equation: $Q = mc\Delta T$

Q = thermal energy flow required in calories -- positive for energy *in*, and negative for energy *out*
m = mass of the object in grams
c = specific heat – in J/g-^{0}C
ΔT = temperature change in degrees Celsius -- (+T) for temperature increases and (-T) for temperature decreases

Problem 1

How many Joules of thermal energy must flow into a 100 gram sample of water to increase the temperature from 30^{0}C to 100^{0}C?

Mass = m = 100 g Specific Heat = cw = 4.182J/g-^{0}C Temperature Change = ΔT = +70

Q = mcΔT

Q = (100g) x (4.182 J- /gm^{0}C) x (70^{0}C) = 29,274 J g^{0}C ÷ g^{0}C = 29,274 J

Problem 2

How many Joules of thermal energy must flow out of a 100-gram piece of aluminum to decrease the temperature from 50^{0}C to 30^{0}C?

Mass = m = 100 g Specific Heat = c_{Al} = 0. 899J/g-^{O}C Temperature Change = ΔT = -20^{0}C

Q - mcΔT

Q = (100g) x (0.899 J/g-^{0}C) x(-20^{0}C) = -1,798 J g^{o}C ÷ g-^{o}C = - 1,798 J

Problem 3

What would be the final temperature of a 200 gram piece of silver at 10 ^{0}C if 2,340 Joules of thermal energy flowed into the silver sample?

Q = 2,340 J Mass = m = 200 g Specific Heat = c_{Al} = 0.234 J/g-C^O

$Q = mc_{Ag}\Delta T$

2,340 J = (200gm) x (0.234 J/g -^{0}C) x (ΔT^{0}C) = (46.8 J-g/g ^{0}C) x (ΔT ^{0}C}

2,340 J ÷ 46.8 J-g/g ^{0}C = (ΔT^{0}C) = 50 ^{0}C Final Temperature = 10^{0}C + 50 ^{0}C = 60^{0}C

Problem 4

An unknown metal with a mass of 100 grams at 200^{0}C is placed in 100 grams of water at 30^{0}C. When thermal equilibrium is reached, the temperature is 35^{0}C for both substances. What is the specific heat and name (check chart on page 17) of the unknown metal?

Energy Lost by Metal (Q) = Energy Gained by Water (Q) $Q = mc\Delta T$

(100g) x (c_m) x (165^{0}C) = (100g) x (4.182 J/g-^{0}C) x (5^{0}C)
(16,500g-^{0}C) x (c_m) = 2,091 J
(c_m) = 2,091 J ÷ 16,500 g-^{0}C = 0.126 J/g^{0}C) = gold

Notice how the three forms of H_2O have different specific heats. Water has almost twice the specific heat of steam or ice. The structure of water varies in all three phases, so the amount of energy needed to increase the kinetic energy of H_20 molecules differs. Also, notice that the seven lowest specific heats on the chart on page 17 are all metals. Metals, with so many free electrons, require very little energy to increase the average kinetic energy (temperature) of their atoms.

Her molecules are losing their attraction for each other.

Melting Point

In solids, the atoms/molecules of a substance are kept in fixed positions because of the strong forces of attraction between them. As the temperature of the substance increases the atoms/molecules vibrate with greater motion and kinetic energy.

With this increase in temperature comes a decrease in the force of attraction between the atoms/molecules. At a certain temperature, the attractive forces decrease enough to free the atoms/molecules from their fixed positions. The substance changes from a solid to a liquid. This temperature is called *the melting point of the substance.* Molecular solids like water and carbon dioxide have low melting points because they are held together by weaker intermolecular forces. The stronger ionic and covalently bonded materials have higher melting points.

The melting point of a substance could also be called its freezing point. If you remove the kinetic energy from the atoms/molecules of a liquid you slow their motion and reduce their kinetic energy. The distances between the atoms/molecules decrease and the forces of attraction become stronger. The previously free atoms/molecules join together to organize themselves into fixed positions resulting in a solid.

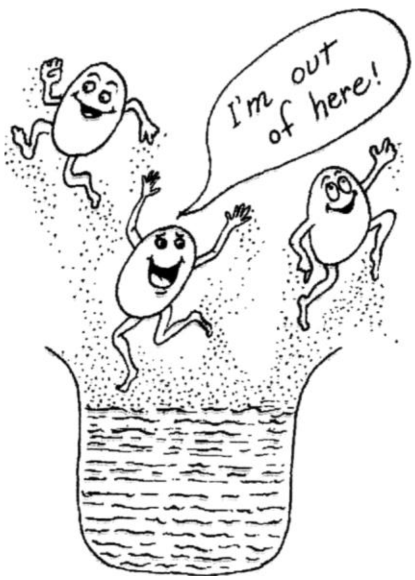
Boiling Point

Liquids have weaker forces of attraction between their atoms/molecules than solids, but the forces still exist. When the kinetic energy of the atoms/molecules of the liquid increases to a certain temperature, they free themselves from their neighbors and become a gas. Once again, the opposite is true; lower the temperature of the gas atoms/molecules and the forces of attraction become stronger and the individual atoms/molecules join together to form a liquid.

Some Heats of Transformation

Substance	Melting Point (^{0}C)	Heat of Fusion (J/g)	Boiling Point (^{0}C)	Heat of Vaporization (J/g)
Aluminum	660	396	2,467	10,530
Copper	1,083	206	2,566	4,730
Ethyl Alcohol	- 114	198	78	846
Gold	1,063	63	2,808	1,577
Iron	1,808	247	3,023	1,520
Lead	327	23	1,750	208
Mercury	39	11	357	295
Oxygen	- 218	14	- 183	213
Silver	961	105	2,193	2,356
Water	0	335	100	2,272

Evaporation



We learned earlier that the temperature of a substance is the average kinetic energy of its countless atoms/molecules. Some of the atoms/molecules have higher (and some have lower) kinetic energies than the average; most are somewhere around the average.

When an atom/molecule possesses *a lot more energy than the average* is at the surface of the liquid, it can free itself from the attractive forces of its neighbors. It becomes an individual atom/molecule in the gaseous phase. This process is called **evaporation.**

Since the most energetic molecules escape, or evaporate, their departure will reduce the average kinetic energy (or temperature) of the liquid.

If your best batter is out of the lineup, the team batting average will drop. When calculating the average height of classmates, what happens to the average if several of the tallest students are not in class that day? The average becomes lower. If the average is lower when the most energetic atoms/molecules escaped, the temperature has dropped. Remember, temperature is the measure of the average kinetic energy of a substance's atoms/molecules.

Evaporation is a cooling process. The loss of the substance's most energetic atoms/molecules results in a decrease in temperature. We've all experienced this cooling process. Remember when you started shivering after getting out of a heated swimming pool or a hot shower? Perspiring cools the body in the same way as perspiration evaporates -- by taking internal energy from the body.

Condensation

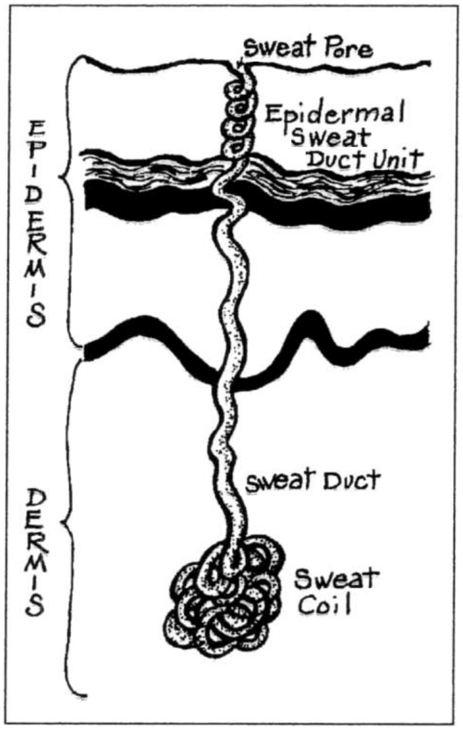
When gas atoms/molecules strike the surface of a liquid or a solid they may lose enough kinetic energy from the collision to change back into a liquid. The loss in kinetic energy results in a decrease in motion of the atoms/molecules. This decrease results in an increase in the attractive forces between it and neighboring atoms/molecules. This increase in the force of attraction transforms individual atoms/molecules into the liquid phase.

We've all witnessed the formation of water drops on the outside of a glass containing a cold beverage; we know that the water isn't coming through the glass. *This is called condensation*. As countless water molecules in the air strike the cold glass and transfer some of their kinetic energy to the glass, this changes them into the liquid phase.

Since the atoms/molecules in the glass are gaining energy from the energetic gaseous water molecules, the temperature of glass increases. *Condensation is a warming process.* When the baseball team gets several new players with great batting averages the team average will increase. The average height of the class will increase with the addition of several tall students.

When gaseous water molecules lose a substantial amount of kinetic energy they transform directly into the solid phase. We've all seen how frost appears on something we've just taken from the freezer.

The condensation of water is the reason that on a very humid day, a day with a large number of water molecules in the air, you feel so uncomfortable. Countless more water molecules are condensed on your skin, raising the average kinetic energy of the atoms/molecules in your skin.



*540 calories of thermal **energy** are required from your body to change the water into a gaseous **phase** when a gram of water evaporates from the skin.*

Heat of Vaporization

Energy is required to overcome the attractive forces of atoms/molecules in a liquid and to change the atoms/molecules to gas. When gas atoms/molecules transfer some of their energy they change into liquids. The questions are how much energy, and is it the same amount for all substances? The answers: a lot of energy, and no.

Since the forces of attraction between atoms/molecules differ from substance to substance, the energy needed to escape from or return to the substance also differs.

The amount of energy required to overcome all the forces of attraction between neighboring atoms/molecules is great. The amount of thermal energy needed to change one gram of a liquid at its boiling point totally into a gas is known as the heat of vaporization of the liquid.

The table on the bottom of page 19 lists the heat of vaporization of several common substances. Notice that 2,272 Joules of thermal energy are needed to change one gram of water at a temperature of 100 ^{0}C entirely into steam at a temperature of 100 ^{0}C. The addition of energy does not change the temperature of the water, only its phase (liquid-gas). The energy is used to break the forces of attraction.

To calculate the number of Joules needed to change a liquid into a gas at its boiling point, multiply the number of grams of the sample by the heat of vaporization for that sample.

Energy Required (Released) = Mass x Heat of Vaporization

The amount of energy released by a gram of gas atoms/molecules when it totally changes into a liquid can be calculated the same way. These energetic gas atoms/molecules must transfer some of their energy if they are to change into a liquid phase. When a gram of water condenses on your skin, 2,272 Joules of thermal energy are added to your body.

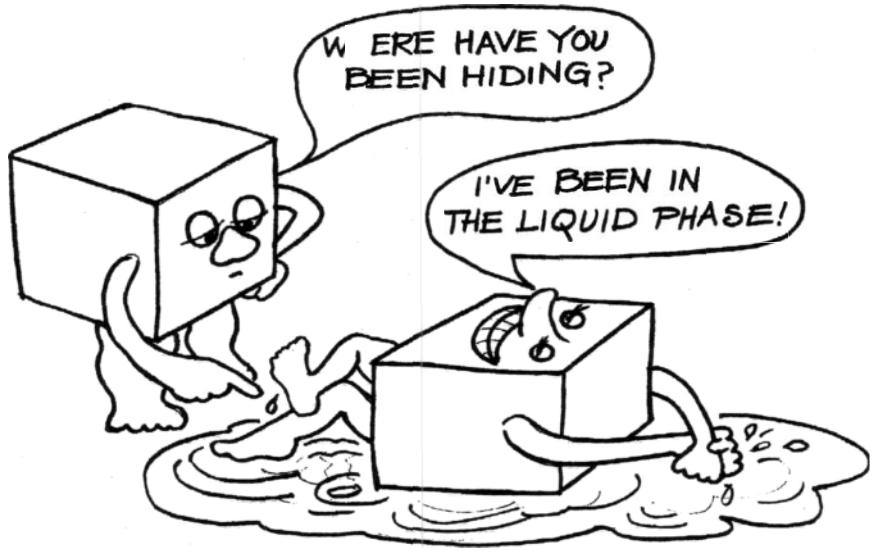
Heat of Fusion

In solids the atoms/molecules are held in fixed positions; they can only vibrate. Energy is required to weaken these forces of attraction and allows the atoms/molecules to move freely. While the forces of attraction are greatly reduced in the liquid phase, they still exist.

The amount of energy required to change one gram of a solid totally into a liquid at its melting point is called its *heat of fusion*. The chart on page 19 lists the heat of fusion for several common substances.

Notice that it takes substantially more energy to change from a liquid into a gas than to change from a solid into a liquid. Why? The energy required to overcome almost all the forces of attraction as liquids turn into gases is much greater than the energy needed to free atoms/molecules from their fixed positions in solids.

The energy required for a phase change is hidden heat energy. It appears when a substance changes from a gas to a liquid, or from a liquid into a solid. The energy hides when changing from a liquid into a gas, or from a solid into a liquid.



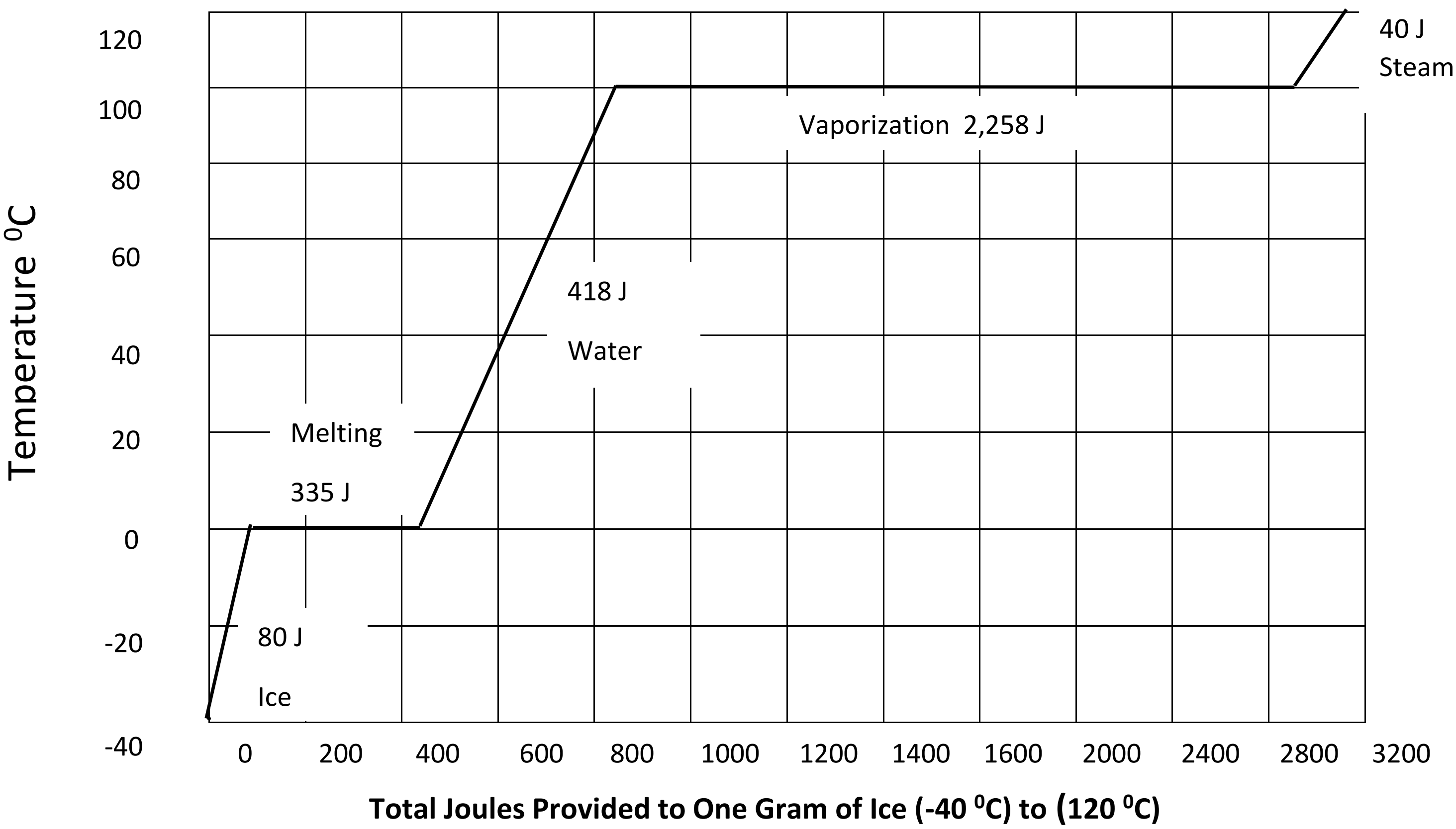
It takes 335 Joules of thermal energy to change one gram of ice at 0 °C into one gram of water at $0\ ^0$C. The opposite is also true. One gram of water at $0\ ^0$C will require the removal of 335 Joules of thermal energy to produce ice at $0\ ^0$C. Ten grams of water will require the removal 3,350 Joules of thermal energy to produce ice at $0\ ^0$C.

Energy Required (Released) = (Mass) x (Heat of Fusion)

The hidden energy in liquids is the reason that ice cubes cool your beverage so quickly. When an ice cube is placed in a glass of water, the temperature of the ice (approximately $-15\ ^0$C) starts to rise by a thermal energy flow from the warmer water. The water cools.

When the ice at the surface of the cube reaches 0 °C, 335 Joules of the water's thermal energy are required to change one gram of the ice into water at 0 'C. Remember, 335 Joules is a lot of energy. The removal or addition of 335 Joules of thermal energy can lower/raise the temperature of one gram of water $80\ ^0$C. Here the 335 Joules are used only to produce a phase change.

So if the size of an average ice cube is 60 grams, 20,100 Joules of thermal energy are required to cause a phase change. The opposite is also true. Your freezer must remove 20,100 Joules of thermal energy from the water to change it from water at $0\ ^0$C to ice at $0\ ^0$C.



Q = Thermal Energy Added or Removed m = mass ΔT = Temperature Change
c = Specific Heat H_f = Heat of Fusion H_v = Heat of Vaporization

First Step: We start with a one-gram sample of ice at -40^{0}C. We add thermal energy to the ice sample to increase the motion of the H_2O molecules. The molecules stay in fixed positions until the sample reaches 0^{0}C. During this time, 80 Joules of energy are required to cause the ice to increase in temperature by 40^{0}C.

$Q = mc_{ice}\,\Delta T$ $Q = (1\ g)\ (2.0\ J/g^0C)\ (40^0C) = 80\ J$

Second Step: Three hundred thirty-five Joules of thermal energy are needed to melt one gram of ice. This energy does not affect the temperature. The energy is just to change the phase from a solid into a liquid. The H_2O molecules which were once in a repeating solid crystalline pattern with strong forces of attraction between molecules are now free to move.

$Q = m\ H_f$ $Q = (1\ g)\ (335\ J/g) = 335\ J$

Third Step: To increase the one gram of water from 0^{0}C to 100^{0}C will require 418.2 Joules. It required 335 J just to melt the ice. Now 418 J will raise that one gram of water 100^{0}C.

$Q = (1\ g)\ (4.18\ J/g^0C)\ (100^0C) = 418\ J$

Fourth Step: To change one gram of water into steam will require 2,258 Joules. This energy does not affect the temperature. The energy is just to change the phase from a liquid into a gas. The H_2O molecules in the gaseous form are now at great distances from each other and have practically no forces of attraction between them.

$Q = m\ Hv$ $Q = (1\ g)\ (2{,}258\ J/g) = 2{,}258\ J$

Fifth Step: A total of 41.18 Joules are required to increase the temperature of steam from 100^{0}C to 120^{0}C.

$Q = mc_{steam}\,\Delta T$ $Q = (1\ g)\ (2.0\ J/g^0C)\ (20^0C) = 40.0\ J$

Total Joules needed = 80J + 335 J + 418 J + 2,258 J + 40 J = <u>3,131 J</u>

The History and Mystery of Cool

Keeping buildings warm (or warm enough) has always been achieved by using a sufficient number of fireplaces, stoves or furnaces in your home or castle. The rooms and people in those rooms would be heated by thermal convection and radiation produced by these sources of heat. However, providing cool air into a building environment is a more complicated task.

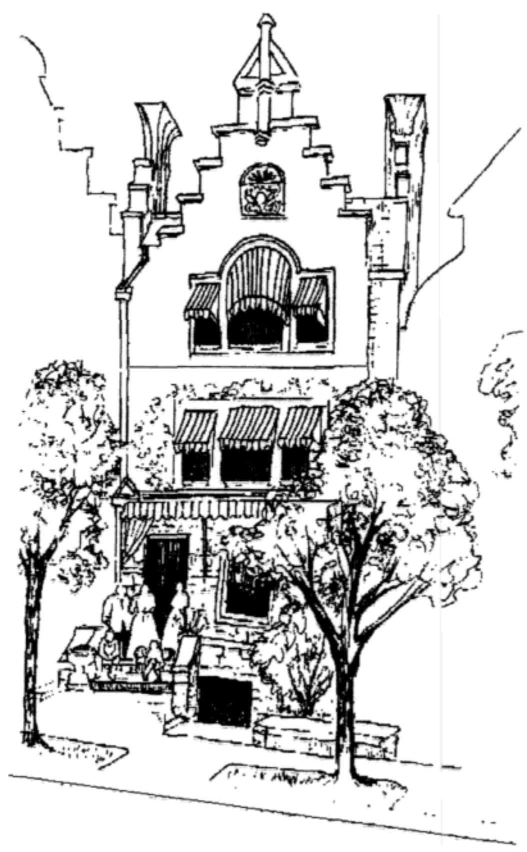
We live in a temperature-controlled environment. On the hottest day of the year we can step from one air-conditioned environment into another, spending only minutes in the heat of the day. What was it like in the *olden days*, the days before air-conditioning† It was horrible! There were weeks and months of constant heat and humidity in many regions of the nation and world without relief.

Can you image spending weeks and months in a classroom or in a factory, at temperatures of 90^{0}F (35^{0}C) and above? The inability to cool environments led to lower productivity. It made certain areas of the country less desirable to establish businesses and to build communities.

Natural Cooling

Buildings were often constructed to replace hot air with cooler air using natural thermal convection. Buildings were made with breezeways, center hallways and courtyards to allow this natural flow of air. Homes with large windows were built in warmer climates to allow cool breezes to circulate while removing hot air from the room. The breeze might also evaporate a bead or two of perspiration (trillions and trillions of water molecules), further cooling uncomfortable inhabitants.

To eliminate radiation from entering a building, awnings above the windows were used to block out the sun. Trees were also used to prevent solar radiation from striking houses during the summer.

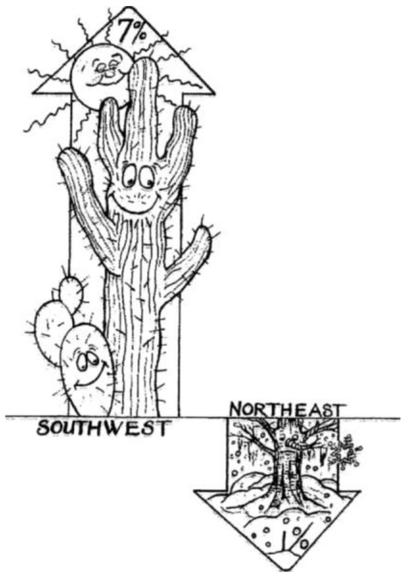


If the breeze was small, or not at all, you could fan yourself. In the late 1880's the very rich who had electricity could use electric fans. Others sat out on the porch to get relief from the heat. On really hot nights, people slept on their porches, balconies, and fire escapes.

Prior to the 1920's few people went to the movies during the summer; it was just too hot. But it was in movie theaters in the 1920's that Americans first got their taste of air conditioning. Instead of being places to avoid, movie theaters became places to escape from the heat and humidity, even if for just a few hours.

Engineered Air

Perhaps one of the earliest attempts to produce cold air using electricity was in the summer of 1881 during an attempt to save the life of President James A. Garfield. President Garfield was running a very high temperature because of an illness resulting from an assassin's bullet. To cool his room, a fan was used to push air through thin cotton screens onto which melted ice dripped. The cooled air was ducted into the president's bedroom, resulting in as much as a 11^0C (20^0F) temperature drop.



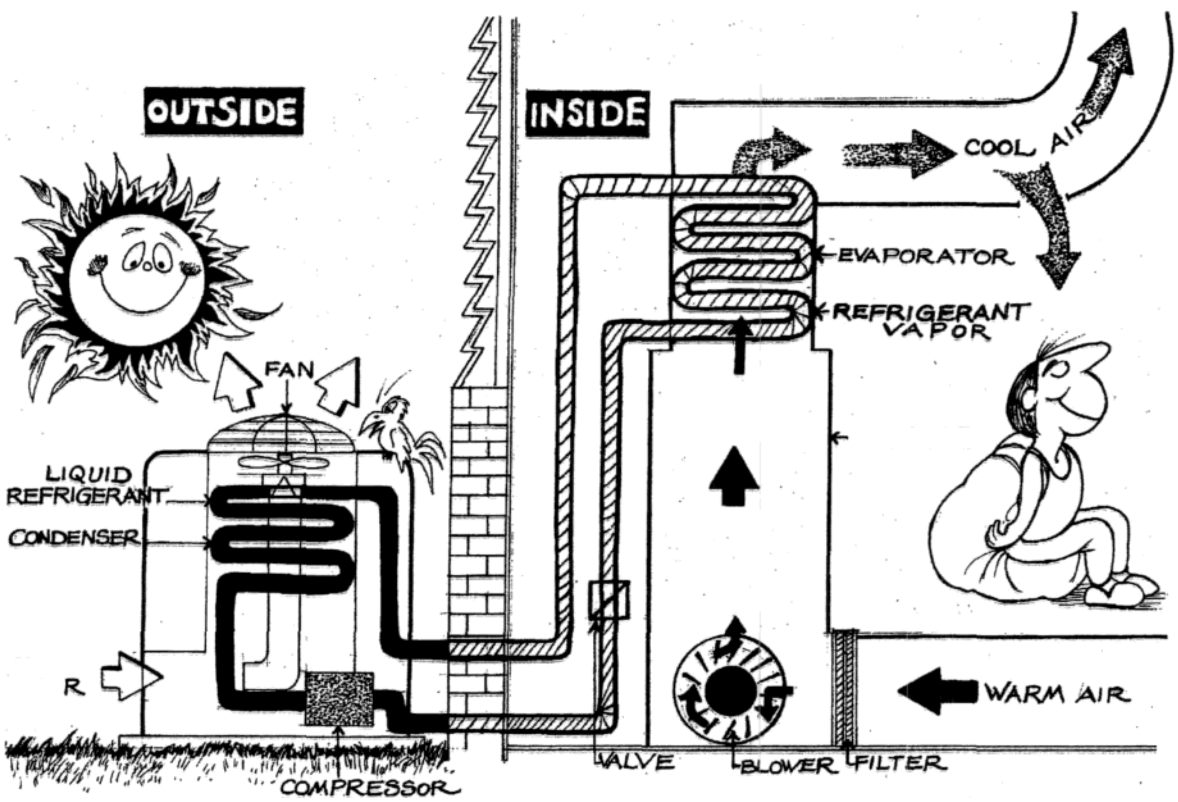
During the past 50 years the population growth of the southwestern states has been at a rate of 7%, in northeastern states the rate has actually declined by 1%.

Cooling technology has come a long way since the 1880's. The ability to control the climate inside buildings and automobiles has totally changed the way we live. Climate-controlled environments have allowed us to build indoor shopping malls, giant indoor stadiums, and hospital operating rooms with safe environments.

People are now able to live comfortably in regions previously considered too hot. The growth of population in these areas greatly outpaced the growth in colder climates. Would you move to cities like Las Vegas, Houston, or Miami if there were no such thing as air conditioning for your home, business or car?

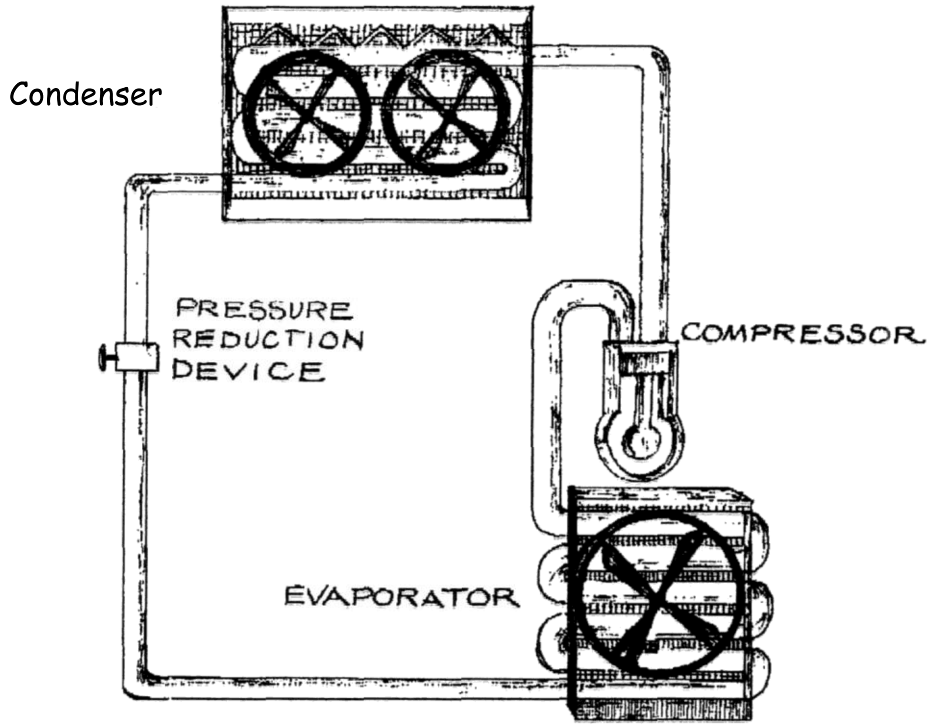
The first big demand for air conditioning, or climate-controlled environments, came from the industrial and agricultural industries. In order to increase productivity and decrease spoilage; climate-controlled environments were created.

The great growth in air conditioning started after World War II, in the late 1940's. The technology of cooling had improved to the point that home and business owners could afford to purchase air-conditioning units. In the early days of air conditioning, these units were plug-in appliances which were placed in spaces in window frames. Today, most new buildings and half the nation's homes are cooled by one or several air-conditioning units; this is called central air conditioning.



Four Steps to Being Cool

The Refrigeration Cycle



High and Low Sides

The refrigerant is under pressure between a point in the compressor, through the condenser, to the pressure reduction device. This is called the high side of the refrigeration cycle.

The refrigerant is at low pressure between the pressure reduction device, through the evaporator, to a point in the compressor. This is called the low side of the refrigeration cycle.

Step One: The Pressure Reduction Device

The key to refrigeration and air conditioning is the fluid that is sent through the system. The fluid is called a refrigerant. Good refrigerants boil at low temperatures which, at normal atmospheric pressure, is somewhere around -25^0C. At room temperature most refrigerants should be in their gaseous phase. However, by keeping the refrigerant under pressure and forcing the refrigerant's molecules close together, the attractive forces between the molecules cause the refrigerant to be in its liquid phase.

The liquid refrigerant is sent through a pressure reduction device, changing the liquid into a spray of liquid and gas. You may have experienced the same process when you allow water to escape from a garden hose with a spray nozzle.

Step Two: The Evaporator

When the tiny droplets inside the coil of tubing are exposed to room temperatures (approximately 45 ^{0}C above the refrigerant's boiling point) they evaporate. The thermal energy required to change the liquid into gas is provided by the tubing which gets very cold.

Blowing air over this cold coil reduces the temperature of the air. The air molecules transfer a portion of their energy to the colder tubing. The air warms the tubing.

Step Three: The Compressor

After the refrigerant's gas molecules have done their job of cooling, they are sent to the compressor. Here the molecules are forced together using a mechanical device powered by a motor. This compression causes the temperature of the gas to rise. The gas molecules gain kinetic energy from the motion of the compressor, just like a baseball gains kinetic energy when it is struck by a moving bat.

You probably have heard the motor of your refrigerator start and stop. The motor is providing the compressor with kinetic energy. This part of your refrigerator is actually hot, not cold.

Step Four: The Condenser

The hot refrigerant gas is sent to the condenser. In the condenser the speed of the molecules is reduced by cooling them. Often this is accomplished by blowing air over the condenser coil. Water is also used to cool the condenser coil.

The air or water absorbs thermal energy from the coil. The kinetic energy of the refrigerant's molecules is reduced. This reduction in kinetic energy allows the attractive forces between the refrigerant's molecules to change them from the gaseous phase back into their liquid phase.

ITEMM Homework & Investigations

Chapter One Questions

l. When you touch a substance you can determine if it is hot or cold. For the following seven temperature descriptions, provide an object or substance that corresponds to that temperature word or phrase. For example a Very Hot substance might be the vegetable oil in French fry cooker.

1. Very Hot 2. Hot 3. Warm

4. Room temperature

5. Cool 6. Cold 7. Very Cold

Answer questions two and three using the information found on page 3 concerning the number of atoms in a cubic centimeter of aluminum and the emptiness of matter,

2.How many atoms do you estimate are in a grain of sand?

3. You are making a scale model of an atom and you are using a grain of sand to represent its nucleus. Estimate how far from this grain you would have to place a second grain of sand to represent the neighboring atom's nucleus?

4. Why is the density of matter in a neutron star so great?

5. Why are the gravitational forces between neighboring nuclei so weak?

6. What causes the forces of attraction between neighboring atoms/molecules?

7. How do crystalline solids differ from amorphous solids?

8. Why does the sharing of electrons in covalent solids produce a strong force of attraction between atoms?

9. Why does the transfer of electrons in ionic solids produce a greater force of attraction between atoms?

10. How does the force attraction between atoms affect the melting and boiling points of solids?

11. How are metallic bonds similar and different from covalent and ionic bonds.

12. Why can metals be made into sheets and foils?

13. What forces keep molecular solids together?

14. Compare the intermolecular and intramolecular air forces between hydrogen and oxygen atoms in ice.

15. What two factors determine the kinetic energy of an object?

16. What two factors determine the kinetic energy of an atom/molecule?

17. What causes atoms to vibrate?

18. Can atoms vibrate at any energy level?

19. How is the vibrational energy transferred in a solid?

20. At room temperature, how many times a second do atoms vibrate?

22. How does the mass of an atom/molecule affect kinetic energy?

22. Why do the forces of attraction increase or decrease between atoms/molecules with temperature change?

23. What term is used to describe the average kinetic energy of atoms/molecules of a substance?

24. Using both the Fahrenheit and Celsius scales, at what temperature does water freeze and boil?

25. Estimate the temperature of the water coming out of your hot water facets.

26. A 20^0 C increase would mean how much of an increase in the Fahrenheit scale?

27. At what temperature do both the Fahrenheit and Celsius scales coincide?

28. What happens to the amount of kinetic molecular energy in a substance as you approach absolute zero?

29. Why can't a technology be developed to reach absolute zero?

30. What type of clothing would a person wear if the temperature was 275 K?

31. How is the internal energy of an object determined?

32. What effect does mass and temperature have on internal energy?

33. Does a stone at room temperature and at ground level contain any energy?

34. How can you hold a piece of lead at lOO^OC in your hand?

35. What molecular force holds liquids together?

36. Describe the motion of atoms/molecules in liquids.

37. Why are the forces of attraction between gas molecules so weak?

38. Approximately how much more space is taken up by the gaseous form of a substance compared to the liquid phase of that same substance

39. Give an example of how gas molecules are very free to move.

40. How many collisions might a gas molecule in encounter in one second?

41. If it weren't for whole chunks of air moving in a room, how long would it take for an odor to travel 10 meters?

42. Explain why the forces of attraction between atoms/molecules increase or decrease with temperature.
43. Why do substances in the liquid state expand more than when in the solid state?
44. Give two examples of how the expansion and contraction of materials are used or considered when building or producing a product.
Consult the **Cubic Expansion Table** on page 11 to answer questions 45-49.
45. Which substance increases the most when heated?
46. Which substance increases the least when heated?
47. Which substance decreases the most when cooled?
48. Which substance decreases the least when cooled?
49. Reinforced concrete contains metal rods that add strength to a structure. To prevent the concrete from cracking as a result of extreme heat or cold, which type of metal would you use and why?
50. At what temperature is water the densest?
51. The volume of a water sample at 4^0C expands. Has the water sample increased, decreased, or can you not tell from this information?
52. Why is it beneficial to have water expand when the temperature drops below 4^0C?
53.Explain the difference between the terms internal energy and thermal energy.
54. Will there be a net flow of thermal energy between two blocks of metal at 300^0C when they are placed together? Why or why not?
55. Explain what happens when an object at 1000C is placed in room temperature air at 22^0C?
56. Water with a mass of 200 grams and at a temperature of 22^0C is poured into an insulated container. When a 100 gram sample of a metal is placed in the water, below the water level, the metal loses 100 units of internal energy. How many units of thermal energy flowed out of the metal, and how many units were gained by the water?
57. Develop an investigation that proves Newton's Law of Cooling.
58. How might the inside temperature of a house determine how much heating or cooling energy must be used to maintain the set thermostat temperature?
59. Explain how thermal energy is transferred at the atomic level by conduction.
60. The tips of two metal rods, aluminum, and copper are placed in a flame. Compare the rate of thermal energy flow from the tip to the other ends of both rods.
61. Why are metals such good conductors of thermal and electrical energy?
62. Based on what you know about air and gases, why are they such poor conductors of thermal energy?
63. Why do the more energetic liquid and gas molecules rise?
64. Give an example of how radiant energy can be transferred without using matter in motion.
65. Give an example of where night goggles or a camera with an infrared lens is used.
66. Do black objects radiant energy? Why or why not?
67. An object is a good absorber of radiant energy. What kind of emitter of radiant energy is the object?
68. Do all substances require the same amount of energy to increase their temperature?
69. Referring to the chart on page 17, which type of substance has a low specific heat capacity?
70. The specific heat capacity of an object is measured by the amount of energy to raise one gram of a substance how many degrees?
71. Compare the energy needed to raise one gram of ice or steam 1^0C to the amount of thermal energy needed to raise one gram of water 1^0C .
72. Why are metals used to make pots and pans?
73. What must happen at the atomic/molecular level for a substance to freeze?
74. How are the melting point and freezing point of a substance related?
75. Of the metals listed on the chart at the bottom of page 19, which would be best to withstand the high temperatures of a furnace?
76. Which metal from the chart on page 19 could you melt in your kitchen oven?
77.What occurs at the atomic/molecular level when a liquid evaporates?
78. Why is evaporation a cooling process?
79. Describe an occasion when you experienced the cooling effect of evaporation.
80. How does a gas molecule condense and become a liquid?
81.Describe an occasion when you experienced the warming effect of condensation.
82. Give an example of how a gas can change into a solid and skip the liquid phase.
83. Why do certain substances require more thermal energy to change them from a liquid into a gas?
84. Which substance listed on the chart on page 19 requires the greatest amount of thermal energy to change one gram of it from a liquid to a gas?
85.Why does it take more energy to change a liquid into a gas than a solid into a liquid?
86. Before the use of air conditioning how would people keep themselves cool during hot weather?

Student Investigations

1. Fill a large pitcher or container with hot water from the tap. Place three large coffee mugs or foam cups on the table. Carefully pour the water into the first container making it 3/4 full, the second container 1/2full, and the third 1/3 full. Place an ice cube of approximately the same size into each container. Record how long it takes for each cube to melt. Why do you think the ice cubes melted in this order?

2. Pour equal amounts of water at different temperatures into three large coffee mugs or foam cups so each is about 2/3 full of water. One container should have very cold water, the second warm water, and the third hot water. Place an ice cube of approximately the same size into each container. Record how long it takes for each cube to melt. Why do you think the ice cubes melted in this order?

3. Pour about 50 ml of pancake syrup, or molasses or vegetable oil into three containers. Refrigerate one container, keep one at room temperature and heat the third one in a hot water bath using the hot water from the faucet. Compare the ease at which each one pours.

4. Place several ice cubes in a large coffee cup and fill the cup with water and let melting take place for about three minutes. During the three minutes boil water and pour into another cup of the same size. Pour both water samples and compare

5. Put a glass in the freezer for 15 minutes. Remove the glass from the freezer and place on a table. Observe and record what happens to the surface of the glass over a period of 5 minutes. Observe the glass two hours later and record your observations.

Thermal Calculations

1. The temperature change, from high to low, during a day is 5°C. What would the temperature change be for this day in Fahrenheit?

2. Propane in its liquid state takes up 1/270 the space than it does when in its gaseous state. How many liters of propane gas will be produced by 10 liters of liquid propane?

For questions 3 **and** 4 refer **to the chart on page 11.**

3. A container with 1,000 liters of water at 10°C is heated to 60°C. What is the volume of the water at 60^0C?

4. At 0°C, a tank contains 100 liters of gasoline. The temperature of the gasoline increases and the tank now contains 103.8 liters of gasoline. Approximately how many degrees Celsius did the temperature of the gasoline increase?

For question 5 refer to page 12, paragraph 3.

5. A sample of water at 4^0C has a volume of 20 liters. What would be the volume of this water if it were to be changed into ice at 0°C? Water expands by 9% when changed into ice at O^OC.

For questions 6-10, refer to the chart on page 17.

How many Joules of thermal energy must be added to a 600 gram sample of gold to increase its temperature 50^0C?

6. A 400 gram sample of copper absorbs 736 Joules of heat. How many degrees Celsius will the sample increase in temperature?

8. A quantity of thermal energy raises a 100 gram sample of water 10^0C. The same amount of thermal energy when added to a 100 gram sample of glass will produce what temperature change for the glass?

9. A 200 gram piece of wood at 10^OC has had 16,720 Joules of heat added to its mass. What will be the temperature of the wood now?

10. An unknown metal with a mass of 50 grams is at 20^0C. The metal is placed in to 50 grams of water at 80^0C. After two minutes the water-metal mixture reaches an equilibrium temperature of 76.8^0C. What is the metal?

11. How many Joules of thermal energy are needed to melt 50 grams of ice at O^OC?

12. How many Joules of thermal energy will be released by 10 grams of steam at 100^OC when it condenses to water at 100^0C?

13. How many Joules of thermal energy must be added to change 10 grams of ice at 0^0C into steam at 100^OC?

14. How many Joules of thermal energy must be added to change 10 grams of ice at -10^0C into steam at 110^0C?

Internal and Thermal Energy - Matter in Motion

Self Evaluation

NFEE

Before you begin the ITEMM unit, take this self evaluation of your knowledge of internal and thermal energy. This completed evaluation will be for your eyes only, so answer honestly. As much of this information will be new to you, don't be surprised if your level of knowledge and understanding is low.

Circle the number corresponding to the statement which best represents your **Understanding (U)** and how **Knowledgeable (K)** you are about the internal and thermal energy words or phrases listed below:

(5) Extremely High U & K (4) Very High U & K (3) Moderate U & K
(2) Little U & K (1) No U & K

Selecting many Little or No U & K responses is expected at the start of the unit.

	Extremely High		Moderate		No Knowledge
Emptiness of Matter	(5)	(4)	(3)	(2)	(1)
Atomic and Molecular Forces of Attraction	(5)	(4)	(3)	(2)	(1)
Kinetic Molecular Theory . .	(5)	(4)	(3)	(2)	(1)
Temperature	(5)	(4)	(3)	(2)	(1)
Internal Energy . . .	(5)	(4)	(3)	(2)	(1)
Thermal Energy . . .	(5)	(4)	(3)	(2)	(1)
Thermal Expansion and Contraction .	(5)	(4)	(3)	(2)	(1)
Uniqueness of Water at 4 ^{o}C .	(5)	(4)	(3)	(2)	(1)
Thermal Equilibrium . . .	(5)	(4)	(3)	(2)	(1)
Laws of Cooling and Heating .	(5)	(4)	(3)	(2)	(1)
Conduction	(5)	(4)	(3)	(2)	(1)
Free Electrons . . .	(5)	(4)	(3)	(2)	(1)
Conductors	(5)	(4)	(3)	(2)	(1)
Insulators	(5)	(4)	(3)	(2)	(1)
Convection	(5)	(4)	(3)	(2)	(1)
Radiation	(5)	(4)	(3)	(2)	(1)
Specific Heat . . .	(5)	(4)	(3)	(2)	(1)
Heat Capacity	(5)	(4)	(3)	(2)	(1)
Thermal Storage . . .	(5)	(4)	(3)	(2)	(1)
Melting and Freezing Points. . .	(5)	(4)	(3)	(2)	(1)
Boiling and Condensation Points. .	(5)	(4)	(3)	(2)	(1)
Condensation . . .	(5)	(4)	(3)	(2)	(1)
Evaporation	(5)	(4)	(3)	(2)	(1)
Heat of Fusion . . .	(5)	(4)	(3)	(2)	(1)
Heat of Vaporization . . .	(5)	(4)	(3)	(2)	(1)
Refrigeration Cycle. . . .	(5)	(4)	(3)	(2)	(1)
History Temperature Regulation . .	(5)	(4)	(3)	(2)	(1)

Upon completion of the ITEMM unit evaluate your level of knowledge and understanding by **darkening** in the number corresponding to the same statements listed above.

Aluminum Sample - An Atomic View

10	8	10	10	6	11	6	13	9	11
9	12	9	7	9	9	9	7	9	10
11	8	7	14	6	8	11	7	10	8
9	8	9	7	6	8	8	8	8	11
12	8	10	9	9	12	5	7	10	13

Energy Level	x	Number at Level	= Total Energy
14	x	________________	= __________
13	x	________________	= __________
12	x	________________	= __________
11	x	________________	= __________
10	x	________________	= __________
9	x	________________	= __________
8	x	________________	= __________
7	x	________________	= __________
6	x	________________	= __________
5	x	________________	= __________

Total ________________ __________

Number of Atoms ______________

Total Energy ______________

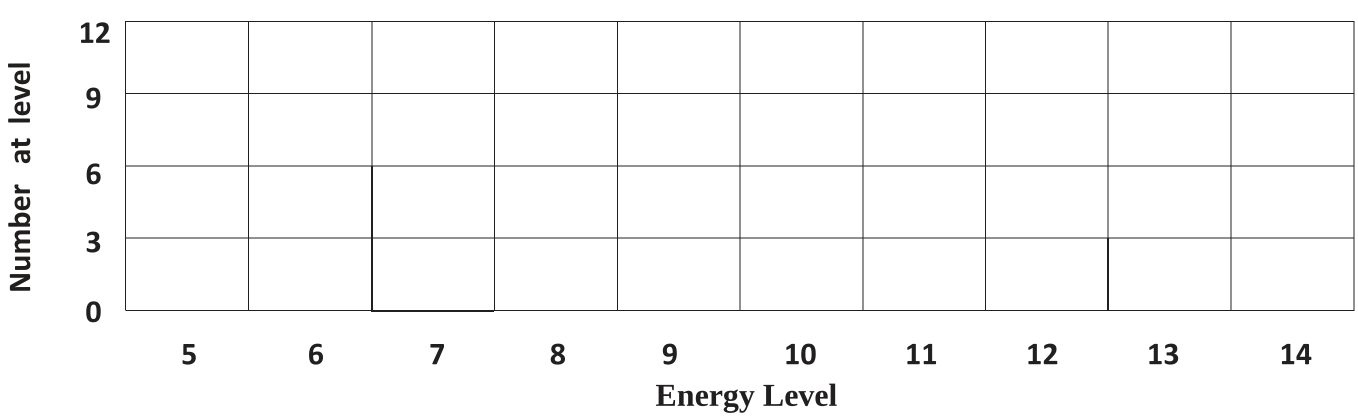
Average Energy ______________

Number at Average ______________

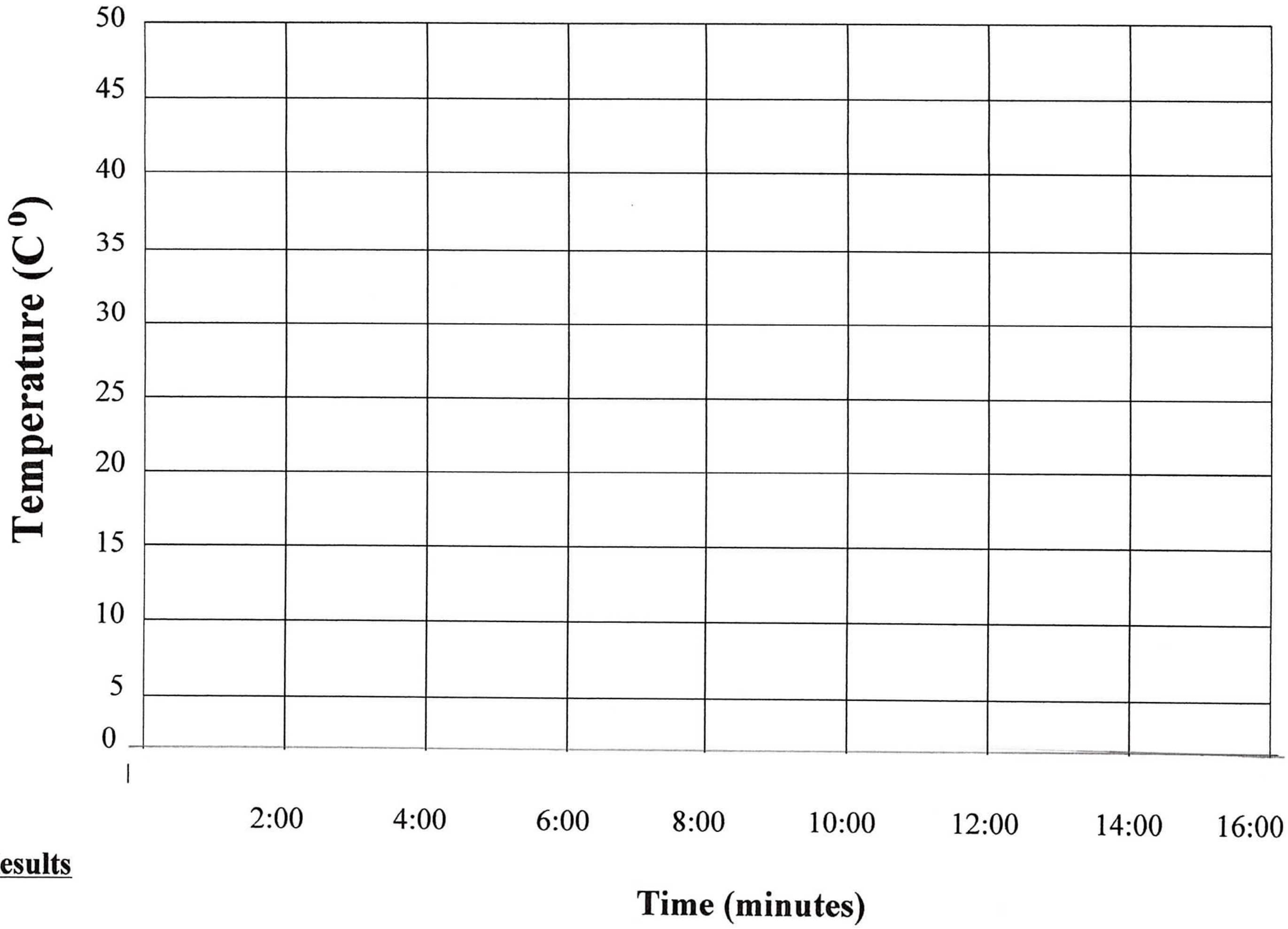
Number Below Average ____________

Number Above Average ____________

Kinetic Energy Distribution of Atoms



Thermal Equilibrium Graph -- Temperature vs. Time



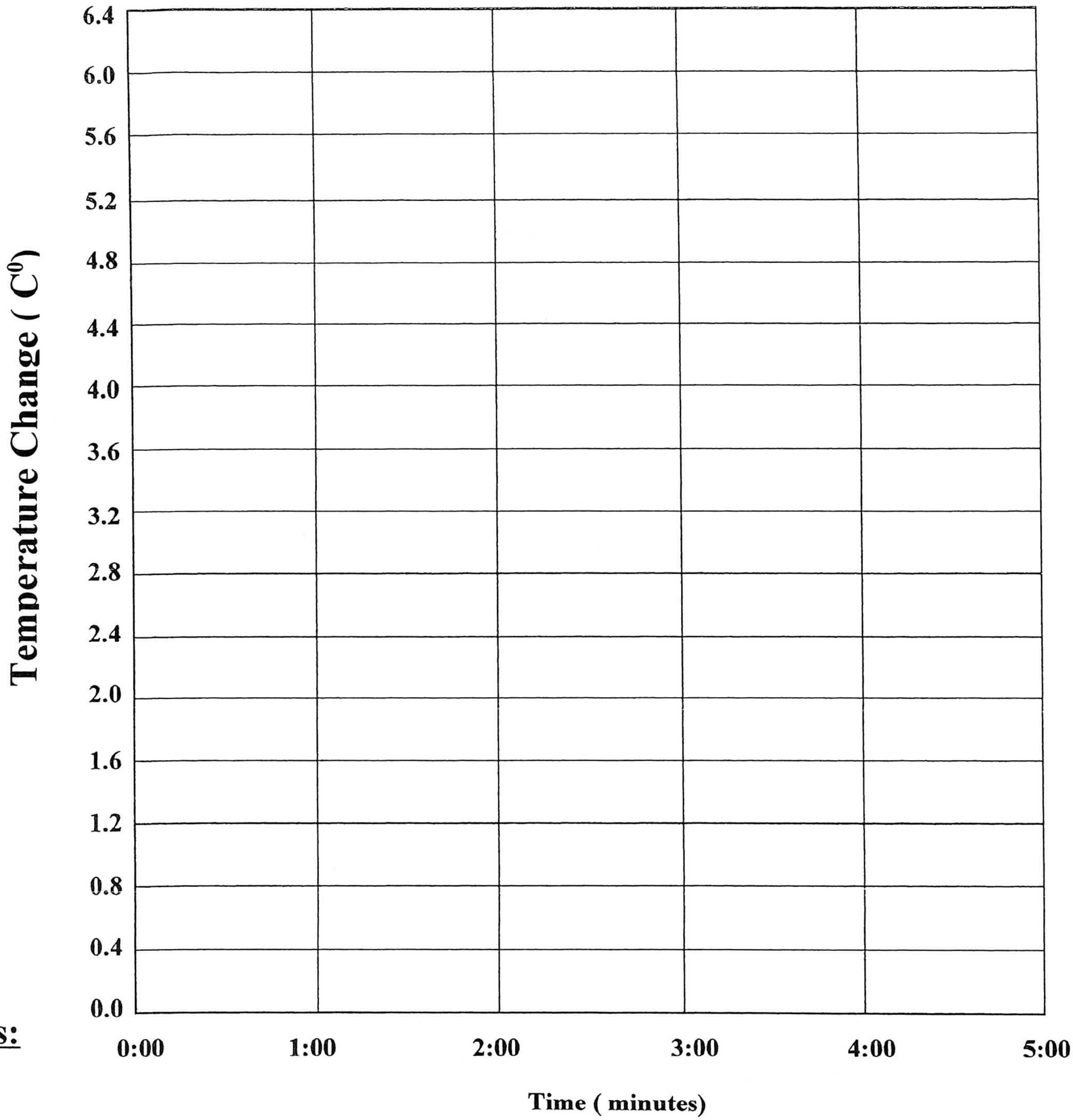
Results

Thermal Equilibrium Data Chart

Time minutes:seconds	Can Water Temperature	Can Water Temperature 2:00 Change	Container Water Temperature	Container Water Temperature 2:00 Change	Temperature Difference Can/Container
0:00		-----		-----	
2:00					
4:00					
6:00					
8:00					
10:00					
12:00					
14:00					
16:00					

Law of Cooling Investigation

NfEE



Results:

Data Chart

Time minutes:seconds	Temperature Starting 1:00 Interval	Change From Previous 1:00 Interval	Total Temperature Change from Start
0:00		---	---
1:00			
2:00			
3:00			
4:00			
5:00			

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