Kirkuk University

Science College

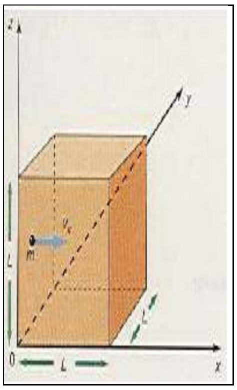
Physics Department

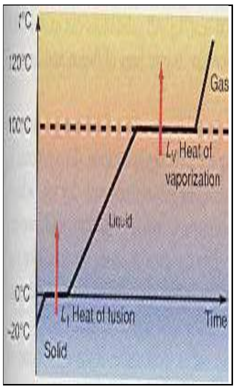
Lectures of

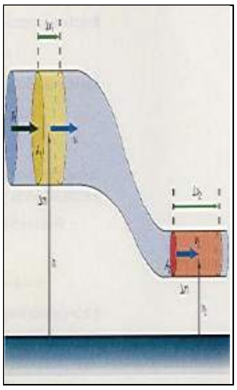
Properties of Matter

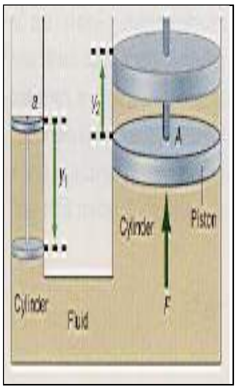
Lecture =4=

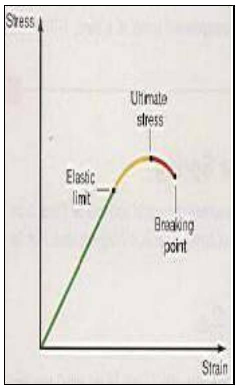
Temperature and Heat











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4.1 Temperature

The simplest and most intuitive definition of temperature is that ***temperature*** *is a measure of the hotness or coldness of a body.*

That is, if a body is hot it has a high temperature, if it is cold it has a low temperature.

This is not a very good definition, as we will see in a moment, but it is one that most people have a "feel" for, because we all know what hot and cold is. Or do we?

We place three beakers on the table, as shown in figure **(4.1)**.

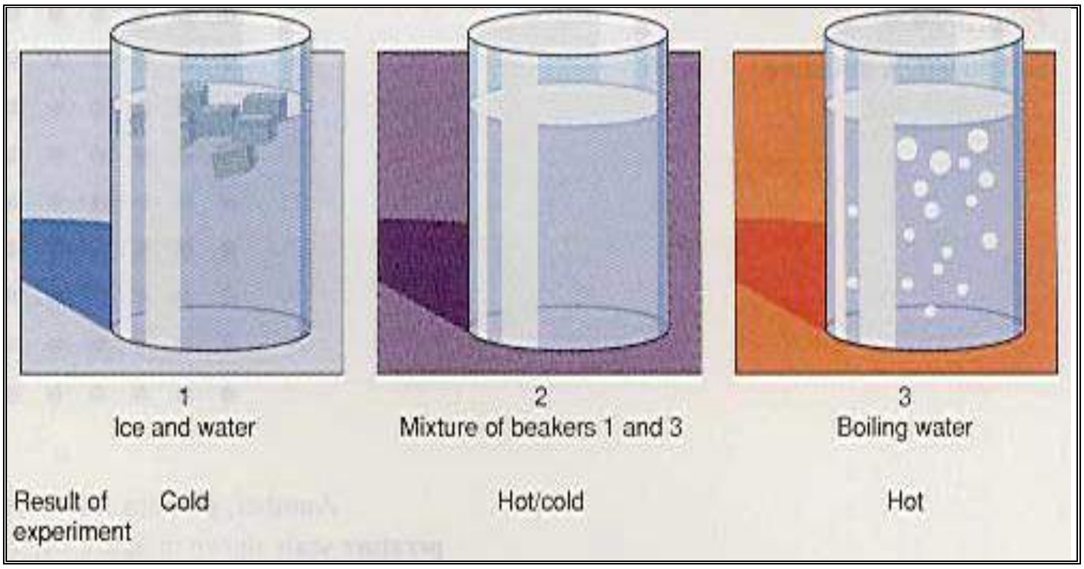

Figure (4.1) **: A "thought experiment" on temperature.**

Several ice cubes are placed into the first beaker of water, whereas boiling water is poured into the third beaker.

We place equal amounts of the ice water from beaker one and the boiling water from beaker three into the second beaker to form a mixture.

I now take my left hand and plunge it into beaker one, and conclude that it is *cold.*

After drying off my left hand, I place it into the middle mixture.

After coming from the ice water, the mixture in the second beaker feels hot by comparison.

So I conclude that the mixture is *hot.*

I now take my right hand and plunge it into the boiling water of beaker three.

(This is of course the reason why this is only a “thought experiment.”)

I conclude that the water in beaker three is certainly *hot.*

Drying off my hand again I then place it into beaker two.

After the boiling water, the mixture feels cold by comparison, so I conclude that the mixture is *cold.*

After this relatively scientific experiment, my conclusion is contradictory.

That is, I found the middle mixture to be either hot or cold depending on the sequence of the measurement.

Thus, the hotness or coldness of a body is not a good concept to use to define the temperature of a body.

Although we may have an intuitive feel for hotness or coldness, we cannot use our intuition for any precise scientific work.

4.2 The Thermometer

In order to make a measurement of the temperature of a body, a new technique, other than estimating hotness or coldness, must be found.

Let us look for some characteristic of matter that changes as it is heated.

The simplest such characteristic is that most materials expand when they are heated.

Using this characteristic of matter we take a glass tube and fill it with a liquid, as shown in figure **(4.2)**.

When the liquid is heated it expands and rises up the tube.

The height of the liquid in the tube can be used to measure the hotness or coldness of a body.

The device will become a ***thermometer***.

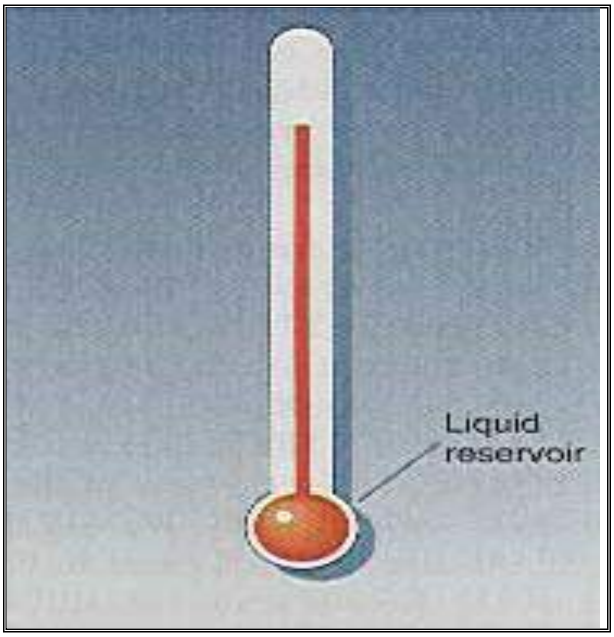

Figure (4.2) **: A thermometer.**

In order to quantify the process, we need to place numerical values on the glass tube, thus assigning a number that can be associated with the hotness or coldness of a body.

This is the process of *calibrating* the thermometer.

First, we place the thermometer into the mixture of ice and water of beaker **1** in figure **(4.1).**

The liquid lowers to a certain height in the glass tube.

We scratch a mark on the glass at that height, and arbitrarily call it **0** degrees.

Since it is the point where ice is melting in the water, we call $\mathbf{0^0}$ the melting point of ice. (Or similarly, the freezing point of water.)

Then we place the glass tube into beaker three, which contains the boiling water.(We assume that heat is continuously applied to beaker three to keep the water boiling.)

The liquid in the glass tube is thus heated and expands to a new height.

We mark this new height on the glass tube and arbitrarily call it $\mathbf{100^0}$.

Since the water is boiling at this point, we call it the boiling point of water.

Because the liquid in the tube expands linearly, to a first approximation, the distance between $\mathbf{0^0}$ and $\mathbf{100^0}$ can be divided into **100** equal parts.

Any one of these divisions can be further divided into fractions of a degree.

Thus, we obtain a complete scale of temperatures ranging from **0** to **100** degrees.

Then we place this thermometer into the mixture of beaker two.

The liquid in the glass rises to some number, and that number, whatever it may be, is the temperature of the mixture.

That number is a numerical measure of the hotness or coldness of the body.

We call this device a **thermometer,** and in particular *this scale of temperature that has* $\boldsymbol{0^0}$ *for the melting point of ice and* $\boldsymbol{100^0}$ *for the boiling point of water is called the* ***Celsius temperature scale*** and is shown in figure **(4.3(a)).**

This scale is named after the Swedish astronomer, Anders Celsius, who proposed it in **1742**.

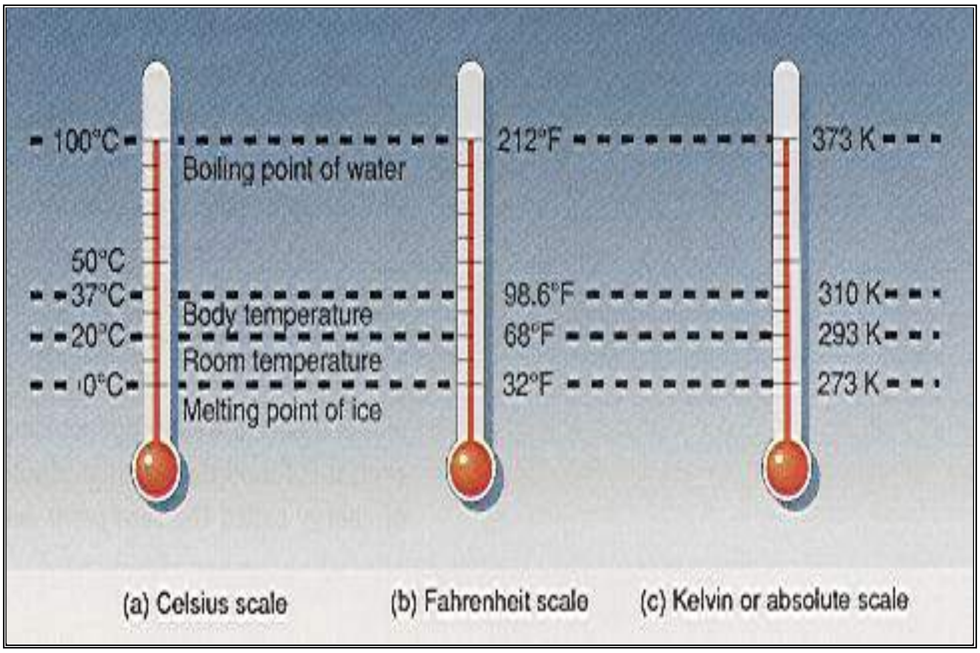

Figure (4.3) **: The temperature scales.**

Another, perhaps more familiar, temperature scale is the ***Fahrenheit temperature scale*** shown in figure **(3.3(b)).**

The melting point of ice on this scale is **32 ^{0}F** and the boiling point of water is **212 ^{0}F**.

This scale is named after the Gabriel Fahrenheit, the German physicist, who proposed it in **1714**.

In addition to the Celsius and Fahrenheit scales there are other temperature scales, the most important of which is the Kelvin or absolute scale, as shown in figure **(3.3(c)).**

The melting point of ice on this scale is **273 K** and the boiling point of water is **373 K**.

The ***Kelvin temperature scale*** does not use the degree symbol for a temperature.

To use the terminology correctly, we should say that, "zero degrees Celsius corresponds to a temperature of **273** Kelvin."

The Kelvin scale is extremely important in dealing with the behavior of gases.

In fact, it was in the study of gases that Lord Kelvin first proposed the absolute scale in **1848**.

For the present, however, the implications of the Kelvin scale can still be appreciated by looking at the molecular structure of a solid.

The simplest picture of a solid, if it could be magnified trillions of times, is a large array of atoms or molecules in what is called a lattice structure, as shown in figure **(4.4)**.

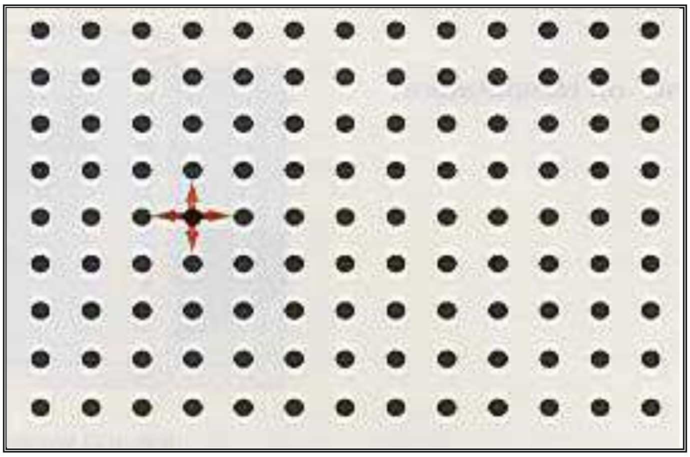

Figure (4.4) **: Simple lattice structure.**

Each dot in the figure represents an atom or molecule, depending on the nature of the substance.

Each molecule is in equilibrium with all the molecules around it.

The molecule above exerts a force upward on the molecule, whereas the molecule below exerts a force downward.

Similarly, there are balanced forces from right and left and in and out.

The molecule is therefore in equilibrium.

In fact every molecule of the solid is in equilibrium.

When heat is applied to a solid body, the added energy causes a molecule to vibrate around its equilibrium position.

As any one molecule vibrates, it interacts with its nearest neighbors causing them to vibrate, which in turn causes its nearest neighbors to vibrate, and so on.

Hence, the heat energy applied to the solid shows up as vibrational energy of the molecules of the solid.

The higher the temperature of the solid, the larger is the vibrational motion of its molecules.

The lower the temperature, the smaller is the vibrational motion of its molecules.

Thus, the temperature of a body is really a measure of the mean or average kinetic energy of the vibrating molecules of the body.

It is therefore conceivable that if you could lower and lower the temperature of the body, the motion of the molecules would become less and less until at some very low temperature, the vibrational motion of the molecules would cease altogether.

They would be frozen in one position.

This point is called *absolute zero,* and is **0** on the Kelvin temperature scale.

From work in quantum mechanics, however, it is found that even at absolute zero, the molecules contain a certain amount of energy called the *zero point energy*.

4.3 Temperature Conversions

The Celsius temperature scale is the recognized temperature scale in most scientific work and in most countries of the world.

The Fahrenheit scale will eventually become obsolete along with the entire British engineering system of units.

For the present, however, it is still necessary to convert from one temperature scale to another.

That is, if a temperature is given in degrees Fahrenheit, how can it be expressed in degrees Celsius, and vice versa? It is easy to see how this conversion can be made.

The principle of the thermometer is based on the linear expansion of the liquid in the tube.

For two identical glass tubes containing the same liquid, the expansion of the liquid is the same in both tubes.

Therefore, the height of the liquid columns is the same for each thermometer, as shown in figure **(4.5)**.

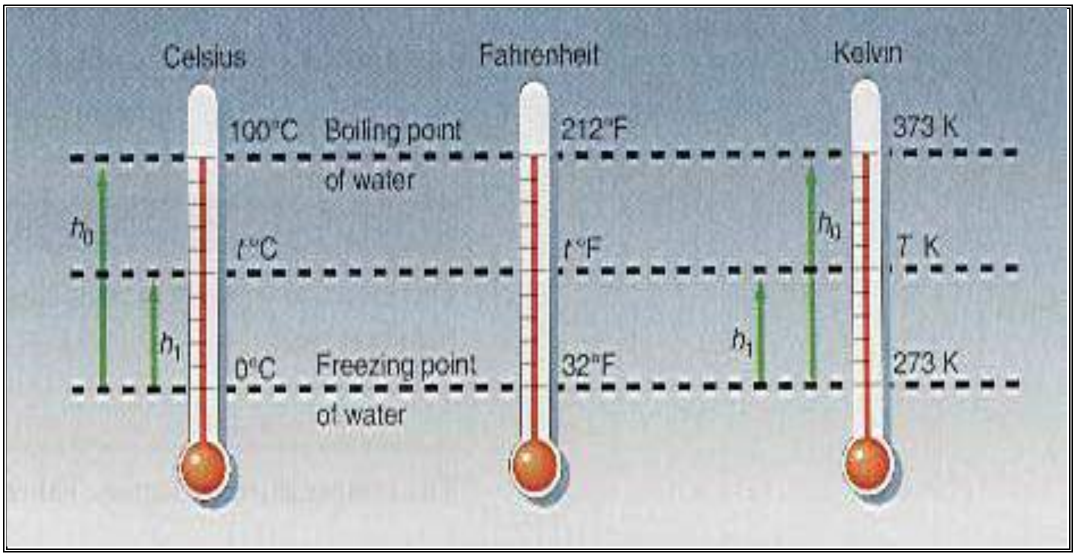

Figure (4.5) **: Converting one temperature scale to another.**

The ratio of these heights in each thermometer is also equal.

Therefore,

$$\left(\frac{h_1}{h_0}\right)_{\text{Celsius}} = \left(\frac{h_1}{h_0}\right)_{\text{Fahrenheit}}$$

These ratios, found from figure **(4.5)**, are:

$$\frac{t\ ^0\text{C} - 0^0}{100^0 - 0^0} = \frac{t\ ^0\text{F} - 32^0}{212^0 - 32^0}$$

$$\frac{t\ ^0\text{C}}{100^0} = \frac{t\ ^0\text{F} - 32^0}{180^0}$$

Solving for the temperature in degrees Celsius

$$t\ ^0\text{C} = \frac{100^0}{180^0}(t\ ^0\text{F} - 32^0)$$

Simplifying,

$$\boxed{t^\circ C = \frac{5}{9}(t^\circ F - 32^\circ) \ldots (4-1)}$$

Equation **(4.1)** allows us to convert a temperature in degrees Fahrenheit to degrees Celsius.

To convert a temperature in degrees Celsius to one in Fahrenheit, we solve equation **(4.1)** for ***t*⁰F** to obtain

$$\boxed{t^\circ F = \frac{9}{5} t^\circ C + 32^\circ \ldots (4-2)}$$

*We can also find a conversion of absolute temperature to Celsius temperatures from figure **(4.5),*** as:

$$\left(\frac{h_1}{h_0}\right)_{\text{Celsius}} = \left(\frac{h_1}{h_0}\right)_{\text{Kelvin}}$$

$$\frac{t\,^0\text{C} - 0^0}{100^0 - 0^0} = \frac{T\,\text{K} - 273}{373 - 273}$$

$$\frac{t\,^0\text{C}}{100} = \frac{T\,\text{K} - 273}{100}$$

Therefore, the conversion of Kelvin temperature to Celsius temperatures is given by:

$$\boxed{t^\circ C = TK - 273 \ldots (4-3)}$$

And the reverse conversion by:

$$\boxed{TK = t^\circ C + 273 \ldots (4-4)}$$

*For very precise work, **0 ^{0}C** is actually equal to **273.16 K**.*

*In such cases, equation **(4.3)** and **(4.4)** should be modified accordingly.*

4.4 Heat

A solid body is composed of trillions upon trillions of atoms or molecules arranged in a lattice structure, as shown in figure **(4.4)**.

Each of these molecules possess an electrical potential energy and a vibrational kinetic energy.

The sum of the potential energy and kinetic energy of all these molecules is called the ***internal energy of the body*****.**

When that internal energy is transferred between two bodies as a result of the difference in temperatures between the two bodies it is called heat.

Heat *is thus the amount of internal energy flowing from a body at a higher temperature to a body at a lower temperature.*

Hence, a body does not contain heat, it contains internal energy.

When the body cools, its internal energy is decreased; when it is heated, its internal energy is increased.

Whenever two bodies at different temperatures are brought into contact, thermal energy always flows from the hotter body to the cooler body until they are both at the same temperature.

When this occurs we say the two bodies are in ***thermal equilibrium*****.**

The traditional unit of heat was the ***kilocalorie,*** *which was defined as the quantity of heat required to raise the temperature of* ***1 kg*** *of water* ***1 ^{0}C****, from* ***14.5 ^{0}C*** *to* ***15.5 ^{0}C****.*

1 kilocalorie = 1000 calories = 4186 J

In terms of the SI unit of energy, the joule, it takes ***4186 J*** *of energy to raise the temperature of* ***1 kg*** *of water* ***1 ^{0}C,*** *from* ***14.5 ^{0}C*** *to* ***15.5 ^{0}C****.*

4.5 Specific Heat

When the temperature of several substances is raised the same amount, each substance does not absorb the same amount of thermal energy.

To handle the problem of different bodies absorbing different quantities of thermal energy when subjected to the same temperature change, *the* ***specific heat c*** *of a body is defined as the amount of thermal energy* ***Q*** *required to raise the temperature of a unit mass of the material* ***1 ^{0}C****.*

In terms of the SI unit joules, the specific heat **c** *of a body is defined as the number of joules* ***Q*** *required to raise the temperature of* ***1 kg*** *of the material* ***1 ^{0}C****.*

Thus,

$$\boxed{c = \frac{Q}{m\Delta t} \quad ...(4-5)}$$

We observe from this definition that the specific heat of water in SI units is **4186 J/kg ^{0}C**, since **4186 J** raises the temperature of **1 kg** of water **1 ^{0}C**.

All other materials have a different value for the specific heat.

Some specific heats are shown in table **(4.1)**.

Table (4.1) : Specific Heats of Various Materials

Table : (4 – 1)

Specific Heats of Various Materials	
Material	$\frac{J}{kg\,^{o}C}$
Air	1009
Aluminum	900.0
Brass	393.5
Copper	385.1
Glass	837.2
Gold	129.8
Iron	452.1
Lead	129.8
Platinum	134.0
Silver	238.6
Steel	452.1
Tin	226.0
Tungsten	134.0
Zinc	389.3
Water	4186
Ice	2093
Steam	2013

Note that water has the largest specific heat.

Having defined the specific heat by equation **(4.5)**, we can rearrange that equation into the form:

$$\boxed{Q = mc\Delta t \ldots (4-6)}$$

Equation **(4.6)** represents the amount of thermal energy $\boldsymbol{Q}$ that will be absorbed or liberated in any process.

4.6 Change of Phase

Matter exists in three states called the **phases of matter.**

They are the solid phase, the liquid phase, and the gaseous phase.

Let us see how one phase of matter is changed into another.

Let us examine the behavior of matter when it is heated over a relatively large range of temperatures.

In particular, let us start with a piece of ice at **−20.0 ^{0}C** and heat it to a temperature of **120 ^{0}C**.

We apply heat to a piece of ice, and observe the temperature as a function of time and plot it, as in figure **(4.6)**.

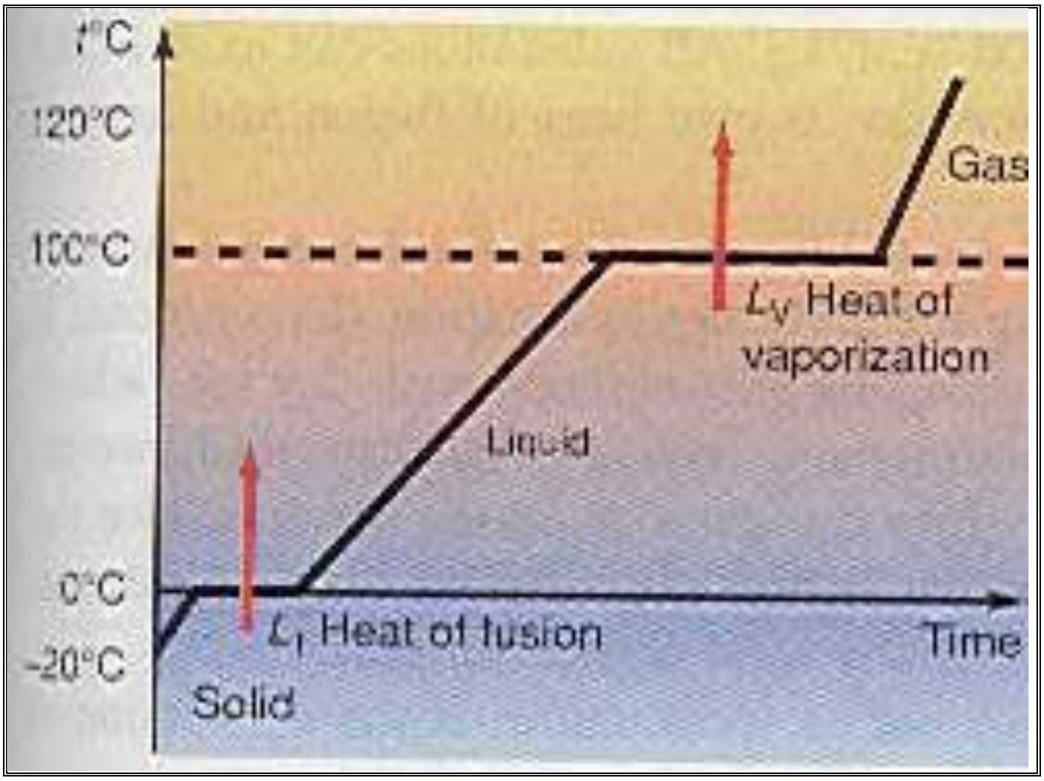

Figure (4.6) **: Changes of phase.**

*The **latent heat of fusion** L_f is the amount of heat necessary to convert **1 kg** of the solid to **1 kg** of the liquid.*

For water, it is found experimentally that it takes **334,000 J** of thermal energy to melt **1 kg** of ice.

Hence we take the latent heat of fusion of water to be :

$$L_f = 3.34 \times 10^5 \text{ J/kg}$$

*The heat that is necessary to convert **1 kg** of the liquid to **1 kg** of the gas is called the **latent heat of vaporization** L_v.*

For water, it is found experimentally that it takes **2,260,000 J** of thermal energy to boil **1 kg** of liquid water.

Hence we take the latent heat of vaporization of water to be:

$$L_v = 2.26 \times 10^6 \text{ J/kg}$$

All substances can exist in the three states of matter, and each substance has its own heat of fusion and heat of vaporization.

Note also that another process is possible whereby a solid can go directly to a gas and vice versa without ever going through the liquid state.

This process is called **sublimation.**

Just as there is a latent heat of fusion L_f and latent heat of vaporization L_v there is also a latent heat of sublimation L_s.

Its value is given by:

$$L_s = 2.83 \times 10^6 \text{ J/kg}$$

Thus, *the heat that is necessary to convert **1.00 kg** of the solid ice into **1.00 kg** of the gaseous water vapor is called the **latent heat of sublimation*** L_s**.**

Q: Compute the thermal energy that is necessary to convert **5.00 kg** of ice at **−20.0 ^{0}C** to superheated steam at **120 ^{0}C**.

The necessary thermal energy is given by **main equation**:

$$\boxed{Q = Q\mathbf{i} + Q\mathbf{f} + Q\mathbf{w} + Q\mathbf{v} + Q\mathbf{s}}$$

where:

Q_i **:** is the energy needed to heat the ice up to **0 ^{0}C**

Q_f **:** is the energy needed to melt the ice

Q_w **:** is the energy needed to heat the water to **100 ^{0}C**

Q_v **:** is the energy needed to boil the water

Q_s **:** is the energy needed to heat the steam to **120 ^{0}C**

(1) The necessary thermal energy to warm up the ice from −20.0 ^{0}C to 0 ^{0}C is found from:

$$\boxed{Q_i = m_i c_i[0^0 - (-20.0\ ^0C)]}$$

The latent heat of fusion is the amount of heat needed per kilogram to melt the ice.

The total amount of heat needed to melt all the ice is the heat of fusion times the number of kilograms of ice present.

Hence, **(2) the thermal energy needed to melt the ice** is:

$$\boxed{Q_f = m_i L_f \ldots(4.7)}$$

(3) The thermal energy needed to warm the water from 0 ^{0}C to 100 ^{0}C is:

$$\boxed{Q_w = m_w c_w(100\ ^0C - 0\ ^0C)}$$

The latent heat of vaporization is the amount of heat needed per kilogram to boil the water.

The total amount of heat needed to boil all the water is the heat of vaporization times the number of kilograms of water present.

Hence, **(4) the thermal energy needed to convert the liquid water at 100 ^{0}C to steam at 100 ^{0}C is:**

$$\boxed{Q_v = m_w L_v \quad ...(4.8)}$$

and

$$\boxed{Q_s = m_s c_s(120\ ^0\mathrm{C} - 100\ ^0\mathrm{C})}$$

is **(5) the thermal energy needed to convert the steam at 100 ^{0}C to superheated steam at 120 ^{0}C**.

Substituting all these equations into **main equation** gives:

$$Q = m_i c_i[0\ ^0\mathrm{C} - (-20\ ^0\mathrm{C})] + m_i L_f + m_w c_w(100\ ^0\mathrm{C} - 0\ ^0\mathrm{C}) + m_w L_v + m_s c_s(120\ ^0\mathrm{C} - 100\ ^0\mathrm{C}) \quad)$$

Using the values of the specific heat from table **(4.1)**, we get:

$$Q = (5.00\ \mathrm{kg})\left(\frac{2093\ \mathrm{J}}{\mathrm{kg}\ ^0\mathrm{C}}\right)(20\ ^0\mathrm{C}) + (5.00\ \mathrm{kg})\left(3.34\times10^5\frac{\mathrm{J}}{\mathrm{kg}}\right)$$
$$+ (5.00\ \mathrm{kg})\left(\frac{4186\ \mathrm{J}}{\mathrm{kg}\ ^0\mathrm{C}}\right)(100\ ^0\mathrm{C}) + (5.00\ \mathrm{kg})\left(2.26\times10^6\frac{\mathrm{J}}{\mathrm{kg}}\right)$$
$$+ (5.00\ \mathrm{kg})\left(\frac{2013\ \mathrm{J}}{\mathrm{kg}\ ^0\mathrm{C}}\right)(20.0\ ^0\mathrm{C})$$

$$= 0.209\times10^6\ \mathrm{J} + 1.67\times10^6\ \mathrm{J} + 2.09\times10^6\ \mathrm{J} + 11.3\times10^6\ \mathrm{J} + 0.201\times10^6\ \mathrm{J}$$
$$= 15.5\times10^6\ \mathrm{J}$$

Therefore, we need **15.5×10^6 J** of thermal energy to convert **5.00 kg** of ice at **−20.0 ^{0}C** to superheated steam at **120 ^{0}C**.

Note the relative size of each term's contribution to the total thermal energy.

The Language of Physics

Lecture 4: (Temperature and Heat)

Temperature

The simplest definition of temperature is that temperature is a measure of the hotness or coldness of a body.

A better definition is that temperature is a measure of the mean kinetic energy of the molecules of the body .

Thermometer

A device for measuring the temperature of a body .

Celsius temperature scale

A temperature scale that uses $\mathbf{0^0}$ for the melting point of ice and $\mathbf{100^0}$ for the boiling point of water .

Fahrenheit temperature scale

A temperature scale that uses $\mathbf{32^0}$ for the melting point of ice and $\mathbf{212^0}$ for the boiling point of water .

Kelvin temperature scale

The absolute temperature scale.

The lowest temperature attainable is absolute zero, the ^{0} K of this scale.

The temperature for the melting point of ice is **273 K** and **373 K** for the boiling point of water .

Internal energy

The sum of the potential and kinetic energy of all the molecules of a body .

Heat

The flow of thermal energy from a body at a higher temperature to a body at a lower temperature.

When a body cools, its internal energy is decreased; when it is heated, its internal energy is increased .

Thermal equilibrium

Whenever two bodies at different temperatures are touched together, thermal energy always flows from the hotter body to the cooler body until they are both at the same temperature.

When this occurs the two bodies are said to be in thermal equilibrium .

Kilocalorie

An older unit of heat. It is defined as the amount of thermal energy required to raise the temperature of **1 kg** of water **1 ^{0}C** .

Specific heat

A characteristic of a material.

It is defined as the number of joules of energy required to raise the temperature of **1 kg** of the material **1 ^{0}C** .

The specific heat of water is **4186 J/kg ^{0}C**.

Phases of matter

Matter exists in three phases, the solid phase, the liquid phase, and the gaseous phase .

Change of phase

The change in a body from one phase of matter to another.

As an example, melting is a change from the solid state of a body to the liquid state.

Boiling is a change in state from the liquid state to the gaseous state .

Latent heat of fusion

The amount of heat necessary to convert **1 kg** of the solid to **1 kg** of the liquid .

Latent heat of vaporization

The amount of heat necessary to convert **1 kg** of the liquid to **1 kg** of the gas .

Latent heat of sublimation

The amount of heat necessary to convert **1 kg** of the solid ice to **1 kg** of the gaseous water vapor .

Summary of Important Equations

Lecture 4: (Temperature and Heat)

Convert Fahrenheit temperature to Celsius	$t\ ^0\text{C} = \frac{5}{9}(t\ ^0\text{F} - 32^0)$
Convert Celsius temperature to Fahrenheit	$t\ ^0\text{F} = \frac{9}{5}t\ ^0\text{C} + 32^0$
Convert Celsius temperature to Kelvin	$T\ \text{K} = t\ ^0\text{C} + 273$
Convert Kelvin temperature to Celsius	$t^{\circ}C = TK - 273$
Thermal energy absorbed or liberated	$Q = mc\Delta t$
The thermal energy needed to melt the ice (Fusion)	$Q_f = m_i L_f$
Vaporization	$Q_v = m_w L_v$

Problems for Lecture

Lecture 4: (Temperature and Heat)

Problem 4.1

Fahrenheit to Celsius

If room temperature is **(68 ^{0}F)**, what is this temperature in Celsius degrees?

Answer : (20 ^{0}C)

Problem 4.2

Celsius to Fahrenheit

A temperature of **(−5 ^{0}C)** is equivalent to what Fahrenheit temperature**?**

Answer : (23 ^{0}F)

Problem *4.3*

Celsius to Kelvin

Normal room temperature is considered to be **(20 ^{0}C)**, find the value of this temperature on the Kelvin scale.

Answer : (293 K)

Problem 4.4

Absorption of thermal energy

A steel ball at room temperature is placed in a pan of boiling water. If the mass of the ball is **(200 g)**, how much thermal energy is absorbed by the ball**?**

Answer : (7230 J)

Problem 4.5

The final temperature

If a **(500 g)** aluminum block at an initial temperature of **(10 ^{0}C)** absorbs **(85500 J)** of energy in a thermal process, what will its new temperature be**?**

Answer : (200 ^{0}C)

Problem 4.6

Converting ice to steam

Compute the thermal energy that is necessary to convert **(5 kg)** of ice at **(–20 ^{0}C)** to superheated steam at **(120 ^{0}C)**.

Answer : (15.5 x 10^6 J)