MEDICAL UNIVERSITY – PLEVEN
FACULTY OF MEDICINE

DIVISION OF PHYSICS AND BIOPHYSICS

LECTURE 5

INTERNAL ENERGY, HEAT AND TEMPERATURE

Internal energy and temperature. First law of thermodynamics. Thermal expansion. Methods for temperature measurement. Specific heat. Heat of combustion: the dietary calorie. The mechanical equivalent of heat

Prof. M. Alexandrova, DSc

INTERNAL ENERGY

At ordinary temperatures the molecules of solids, liquids and gases are in ceaseless, rapid motion.

Def. The energy associated with random, disordered motion of molecules is referred to as *internal energy*.

Def. The condition under which the molecular motion ceases is referred to as the *absolute zero temperature*.

If we look at a solid or a liquid object at room temperature, they may appear to be completely at rest but a microscopic view would reveal the chaotic molecular activity which contributes to the internal energy.

If T of an object is raised, the level of this chaotic activity is increased.

If a person touches a hot object, he experiences collisions with the energetic molecules of the object, which may transfer enough internal energy to his finger to cause pain and physiological damage.

In solids, the motion is restricted to back-and-forth motion about an equilibrium position, analogous to the motion of a mass on a spring.

In liquids, the molecules have unrestricted positions and move around each other under the constraint of strong mutual attractive forces. In both solids and liquids there is a considerable amount of potential energy as well as kinetic energy associated with the state.

In the gaseous state the internal energy is essentially all in the form of kinetic energy of motion, since the intermolecular attractive forces can be neglected.

It is important to keep the words "random" and "disordered" associated with the concept of internal energy as the energy of molecular motion.

Internal energy can not be associated with any type of molecular motion.

E.g. The molecules of a speeding bullet are in motion, but that motion is not classified as internal energy. The motion of the bullet is an example *of ordered* motion, since all the molecules are moving in the same direction.

Whether an object as a whole is in motion or at rest, the microscopic or molecular view of the object will reveal the random agitation of the molecules.

THE DISTINCTION BETWEEN INTERNAL ENERGY AND TEMPERATURE

The terms "**hot**" and "**cold**" are **useful relative terms**, but they are imprecise and deceptive when judged by our senses.

A block of metal
A block of wood
→ Immersed in boiling water and removed
→ The metal will seem considerably "hotter" than the wood

If the two blocks are then immersed in ice water and allowed to come to equilibrium, the metal block will now seem considerably "colder“.

Yet in both experiments T of the blocks were the same.

The difference as perceived by our sense of touch is a result of the differing thermal conductivities of wood and metal.

What is temperature?

Def. The temperature is a measure of the average kinetic energy of the disordered molecular motion, but since that is not the entire internal energy, the temperature is not directly proportional to the internal energy.

Two objects that have the same temperature do *not* necessarily have the same amount of internal energy per gram. A gram of water and a gram of a metal such as copper which are at the same temperature have drastically different internal energy.

Some materials can "hold" more internal energy than others at a given temperature.

Two objects which are at the same temperature are said to be in *thermal equilibrium*.

!!! Regardless of how much internal energy they possess, there will be no transfer of internal energy between them.

Heating is the transfer of internal energy from a high-temperature object to a lower-temperature object. At the boundary between the objects the process may be viewed as the transfer of energy by collisions of the more energetic molecules of the hotter object with the molecules of the colder object.

" More energetic" here refers only to that part of the internal energy that the temperature measures, the motion part.

!!! The lower-temperature object could have the larger total internal energy.

When two objects reach the same temperature, the molecular collisions at their boundaries will transfer no net energy in either direction.

Temperature is usually measured by allowing a substance to come to thermal equilibrium with a thermometer which is calibrated in degrees.

TEMPERATURE SCALES

In order to define a reproducible temperature scale, there must be at least two standard reference points which can be used to calibrate a thermometer.

The two most convenient reference points are the freezing point and boiling point of water.
The freezing temperature of pure water is essentially independent of atmospheric pressure.

The boiling point depends upon the atmospheric pressure, so the pressure must be specified.
The universal standard for temperature scales is the boiling point at 760 mm Hg, the standard atmospheric pressure.

The Celsius or centigrade scale - developed for the measurement of ordinary temperatures.

The freezing point of water is chosen to be 0°C and the boiling point at standard pressure is 100°C. The convenient division into **100 parts** or "**cents**" is the reason for the commonly used designation "centigrade scale."

The kelvin or absolute scale **K** maintains the same degree size as the Celsius scale. K is one of the seven base units in the SI system.

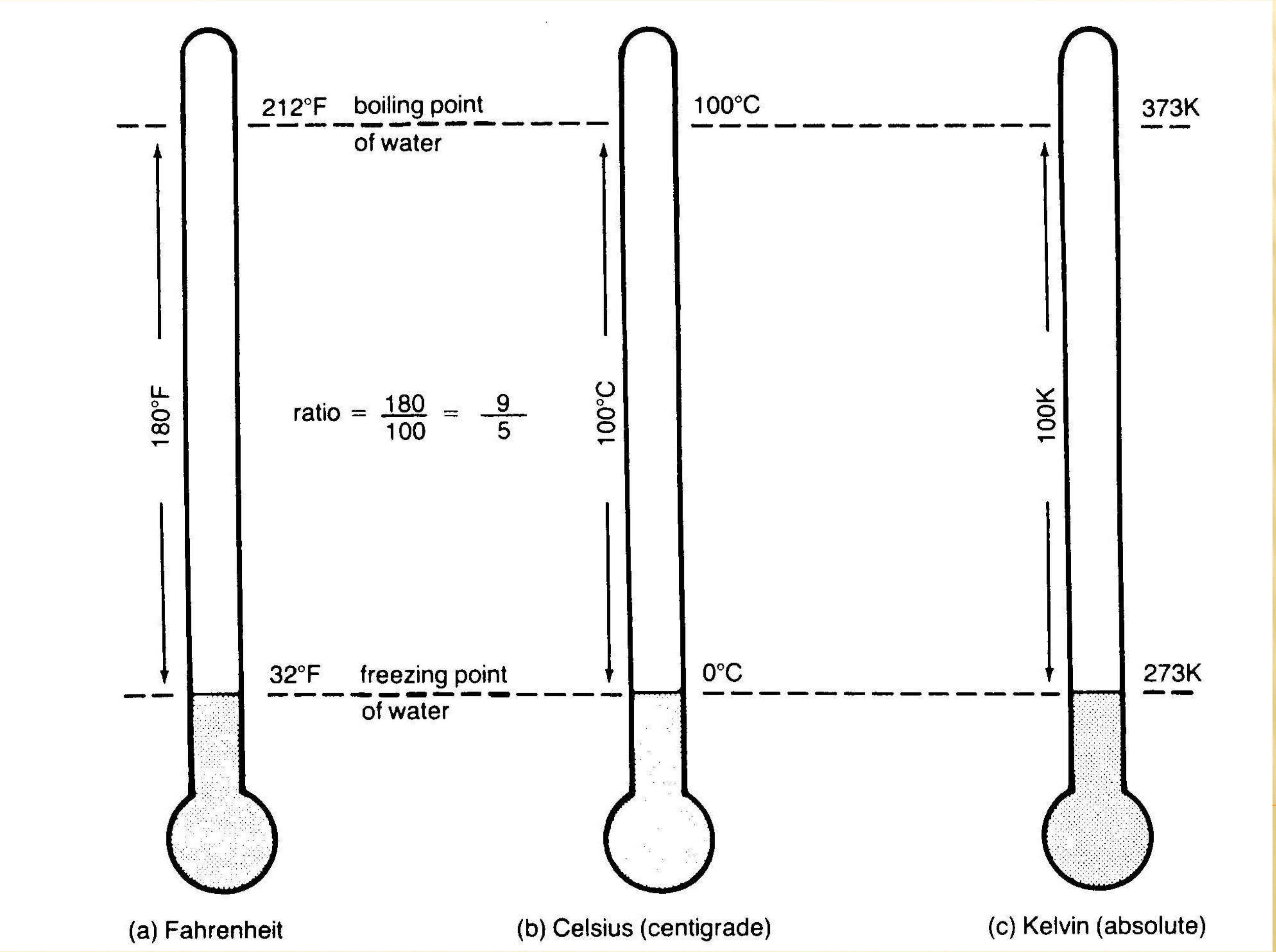
It is standard practice to refer to temperatures or temperature changes in "kelvins," not "degrees kelvin."

The zero of this scale is chosen to be absolute zero. Absolute zero is -273.15°C, so the zero on the Celsius scale corresponds to 273.15K, usually rounded to 273 K for practical purposes.

$$T_K = T_C + 273.15$$

The Fahrenheit scale is a much and much less convenient than the other two scales and is no longer used appreciably in scientific work . The conversion from Celsius to Fahrenheit can be reduced to the formula

$$T_F = \frac{9}{5} T_C + 32$$



212°F boiling point of water
100°C
373K
180°F
ratio = 180/100 = 9/5
100°C
100K
32°F freezing point of water
0°C
273K
(a) Fahrenheit
(b) Celsius (centigrade)
(c) Kelvin (absolute)

HEAT AND THE FIRST LAW OF THERMODYNAMICS

The first law of thermodynamics is basically the application of the conservation of energy principle to heat and internal energy.

The internal energy U may be changed by doing work on the object, as in compression of a gas, or by putting it in contact with a higher temperature object so that energy is transferred to it.

$\Delta U = Q + W$. This relationship is called the first law of thermodynamics.

Heat - "internal energy in transit" or the energy transferred by virtue of a temperature difference.

The transfer of Q is a microscopic process, whereas the work done on the system, W, is a macroscopic process.

The same change in the internal energy U could be accomplished by heat transfer *Q*, by doing work *W* *on* the system, or by a combination of heat and work, with the final results being indistinguishable.

E.g. Suppose a volume of gas at 20ºC is compressed to a volume of 200 cm^3 and its temperature is observed to rise to 40 ° C.

E.g. As an alternative, suppose we start with 200 cm^3 of gas at 20 ° C and hold its volume constant while placing it over a flame until the temperature rises to 40 ° C. Thus, the temperature change has been accomplished entirely by heat transfer *Q*.

The final state of the gas ($V = 200\ cm^3$; $T = 40\ °C$) gives no clue about the process used to get it to that state; it could have been accomplished by heat or work.

In most cases the temperature will rise when Q is transferred to a body, but this is not always true. When you are adding heat to a block of ice to melt it, its U is increasing, but its T=const until it is melted.

In this case the energy transferred by heat is used to alter the bonds which hold the ice in its solid configuration. During the melting E_k does not increase, so the temperature remains the same.

E.g. Suppose an amount of heat *Q* is added to a gas which is allowed to freely expand. The expanding gas may do an amount of work *W* on its surroundings which is exactly equal to the energy transferred to it by heat.

W<0 since it is work done *by* the system and not *on* it. Therefore $\Delta U = Q + W = 0$ - *isothermal* expansion of the gas; the energy transferred by heat enables the gas to do work but does not raise its T.

THERMAL EXPANSION

The expansion of a solid is expressed in terms of the fractional expansion $\Delta l/l$.

T increase $\longrightarrow$ the atoms vibrate back and forth over greater distances and with greater average speeds $\longrightarrow$ larger average distances between atoms $\longrightarrow$ thermal expansion.

Since the interatomic distances all expand by the same amount, then all dimensions of a solid body will increase by the same percentage.

The linear expansion (expansion in one dimension) can be summarized by the equation,

$$\frac{\Delta l}{l} = \alpha \Delta T$$

[α]= 1/°C or 1/°F.

Thermal expansion is similar to a photographic enlargement - no distortion occurs if an object is made of only one type of material.

The thermal expansion of solid objects can create destructive stresses.

When the bottom of a glass vessel is heated, it will crack because the bottom expands more than the top, creating large forces which break the brittle glass. Pyrex is less likely to crack (α is less than half that of ordinary glass). Quartz tubing is often used rather than glass in conditions of extreme temperature changes ($\frac{1}{15}$ α_{glass}).

When solid objects are constructed from more than one type of material, the different rates of thermal expansion may create stresses. A considerable amount of research went into developing the mercury-silver amalgams used by dentists to fill cavities. $\alpha_{amalgams} \approx \alpha_{tooth}$. Otherwise, painful stresses are created by unequal thermal expansion (contraction) when the tooth is exposed to high or low temperatures.

METHODS FOR TEMPERATURE MEASUREMENT

In order to measure temperature conveniently, some easily observable physical property must be found which changes with temperature in a precise and reproducible manner.

Most common thermometers either produce an electrical signal or operate on the basis of thermal expansion.

Liquid Expansion Thermometers. Since the expansion is proportional to the temperature, the height of the column in the capillary can be calibrated directly in ^{0}C or F. This is the basis for the mercury-filled clinical thermometers and the alcohol-filled ordinary household thermometers.

Bimetallic Strip Thermometers

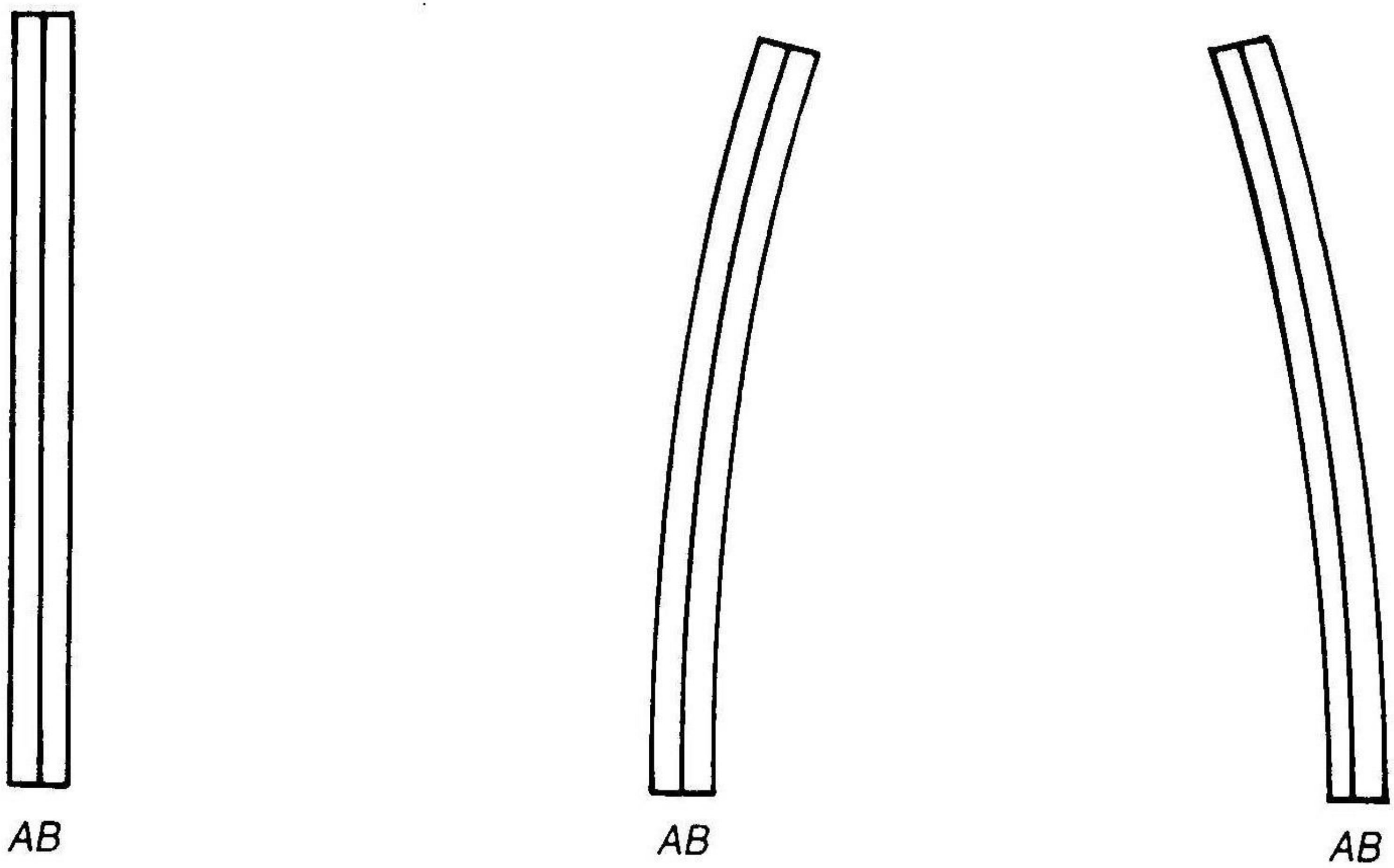
If two metals A and B which have different α are bonded together in the form of a thin strip, it can be used to measure temperature.

If $l_A = l_B$ at T=0°C, the strip will be straight at that T.

If $\alpha_A > \alpha_B$, then at T >0°C $l_A > l_B$ - bending toward B to keep them bonded together . The amount of bending is proportional to T.

Most oven thermometers and thermostat controls are bimetallic strips, because they are quite rugged and can tolerate wide ranges of temperatures.

The bimetallic strip can also be used to measure low T, since when it is cooled below 0°C strip A contracts faster and becomes shorter than B.



(a) $T = 0°C$ (b) T greater than 0°C (c) T less than 0°C

Constant Volume Gas Thermometer - based upon the ideal gas law.

When V=const, the pressure of an enclosed gas is directly proportional to the Kelvin T:

$P = BT$, where B is a numerical constant.

If a pressure gauge is attached to the closed volume, then the pressure gauge face can simply be marked off in Kelvin rather than mm of Hg or some other pressure unit.

This kind of thermometer is easy to calibrate, but it is large and cumbersome compared to the tiny liquid-in-glass thermometers.

INTERNAL ENERGY AND SPECIFIC HEAT

Two substances which are at the same T will in general possess different amounts of U/g.

Def. The specific heat of a substance c is defined as the amount of heat in calories required to raise the temperature of 1g of the substance by 1°C.

The calorie as a unit of heat or internal energy can be defined as the heat or internal energy required to raise the temperature of 1 g of water by one degree Celsius.

The dietary Calorie is 1000 cal, or 1 kcal.

HEAT OF COMBUSTION: THE DIETARY CALORIE

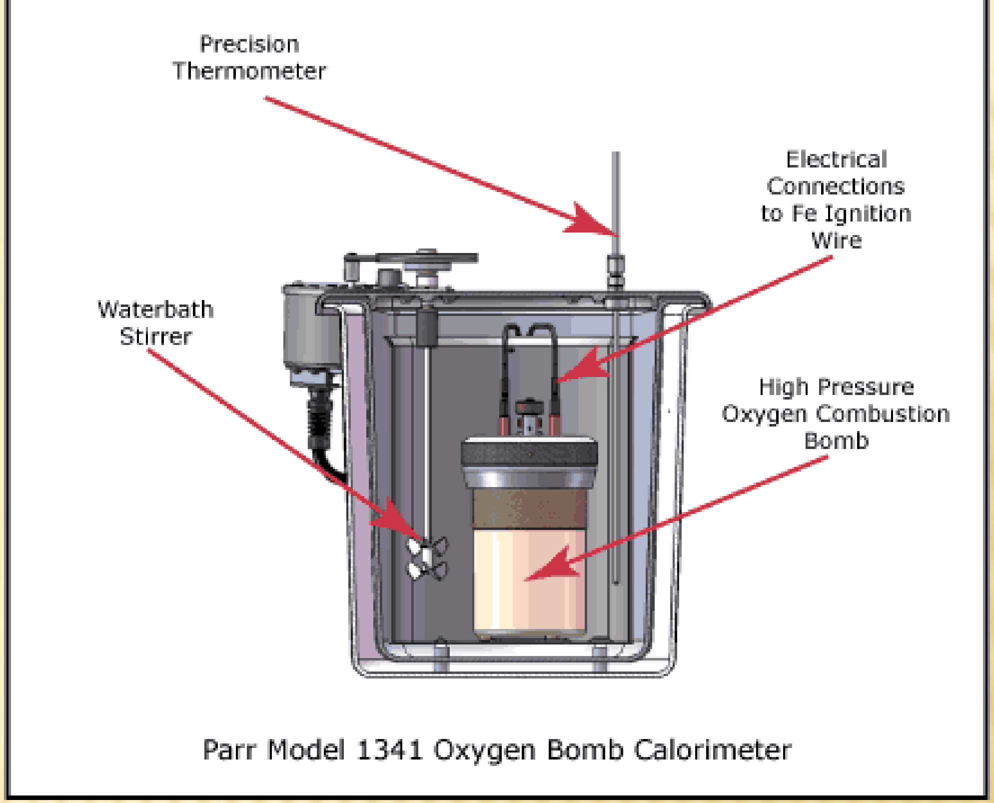
The energy value of foods taken into the body is expressed in terms of the Calorie.

This is appropriate because the process of using the foods is basically an oxidation process assumed to be similar to ordinary burning in terms of the chemical energy released.

An approximation of the nutritional energy available from a food may be obtained by measuring the heat of combustion of the food.

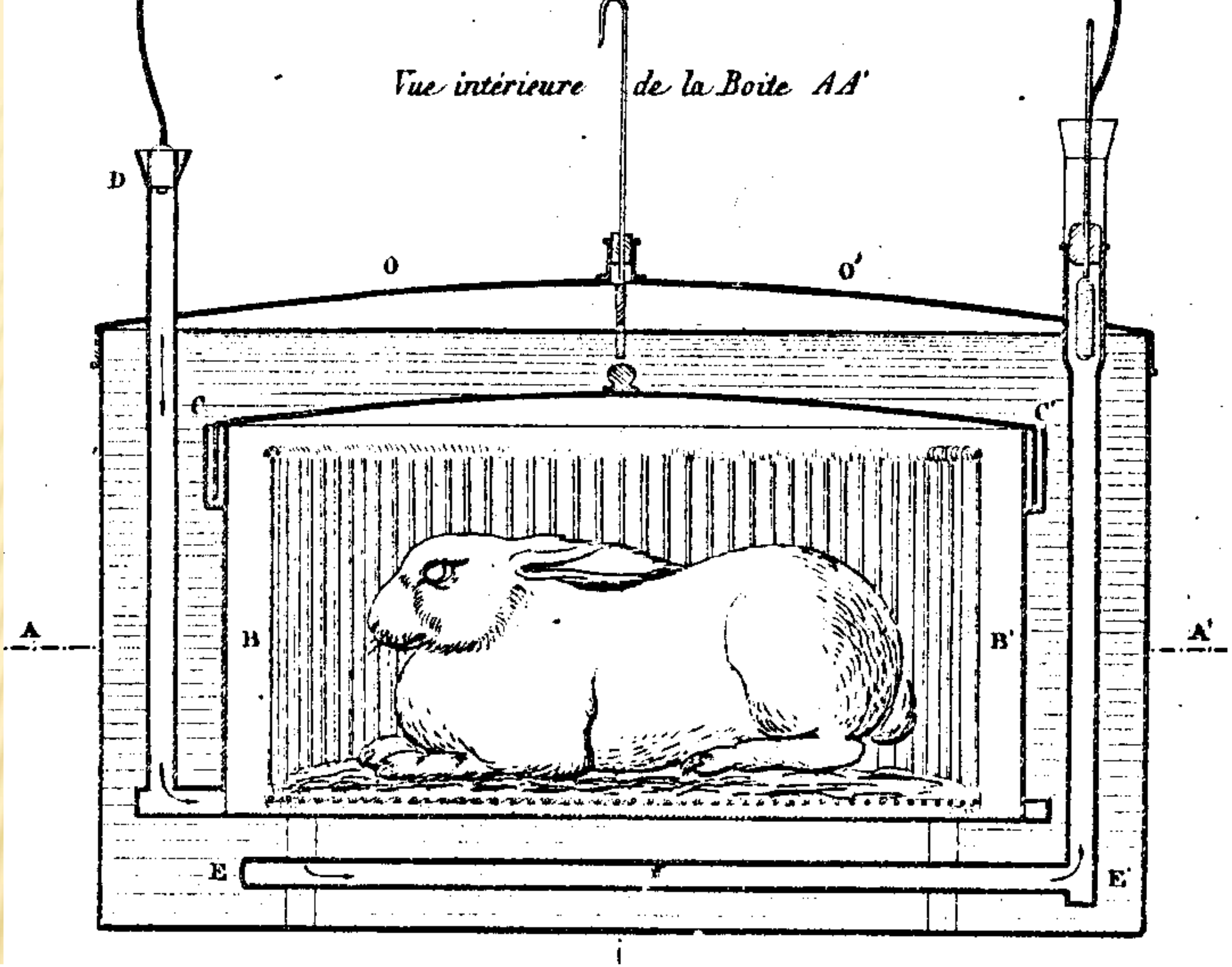
This is normally done by burning the food in a pure oxygen atmosphere inside a calorimeter.

The calorimeter is composed of an oxidation chamber surrounded by a measured volume of water. When the food has burned completely, the heat of combustion is given to the water. Applying the principle of conservation of energy,



the energy given off by the combustion must equal the heat energy gained by the calorimeter if it is sufficiently isolated so that no heat escapes.

From the measurement of the calorimeter temperature before and after the combustion, the energy released by the combustion can be calculated. This heat of combustion is usually tabulated in kcal/g.



The diagram showing a bunny in a calorimeter, is taken from an article published in 1841 in a French scientific journal, *Annales de chimie et de physique*. The figure is captioned "Apparatus of Mr. Dulong for [measuring] animal heat," To obtain accurate measurements of the energy available to the body, corrections would have to be made for substances such as cellulose which will burn in a calorimeter but which is not metabolized by the body.

The three basic types of foods (proteins, carbohydrates, and fats) are shown to produce about equal amounts of Q when burned outside the body and when utilized within the body.

A small correction must be made for proteins, since they are incompletely oxidized by the body, and energy-containing residues of the proteins are excreted in the urine. Although the process of converting food into available energy for the body is similar to combustion, not all of this energy is in the form of heat, since the body does mechanical work.

THE MECHANICAL EQUIVALENT OF HEAT

The British physicist James Joule demonstrated that when mechanical energy was expended on a paddle system which stirred some water, T of the water was raised, indicating that energy was being transferred to the water in the form of heat.

Careful experiments led him to the measurement of the mechanical equivalent of heat. He found that

$$1 \text{ cal} = 4.186 \text{ J}.$$

This equivalence implies that any heat transfer measurement could be expressed in terms of the standard energy unit, J. The rate of production of heat could be expressed in J/s or W. This equivalence is very important in demonstrating the conservation of energy principle.

REVIEW QUESTIONS

1. When internal energy is defined as the energy associated with the random, disordered motion of molecules, why must the words "random" and "disordered" be included?
2. Will substances with the same mass and same temperature always have the same amount of internal energy? Explain.
3. Why is the sense of touch an unreliable test for temperature?
4. When a metal object with a hole in it is heated, will the hole get bigger or smaller?

REVIEW QUESTIONS

5. Why does a bimetallic strip bend when heated? How can this effect be utilized to make a thermometer?
6. What is the advantage of Pyrex over ordinary glass for handling hot substances?
7. Why is water a good substance to use in a hot water bottle?
8. What significance does the large specific heat of water have in relation to the maintenance of a constant body temperature?

REVIEW QUESTIONS

9. How is the dietary Calorie related to the physical calorie?

10. The dietary Calorie equivalent of foods is determined by burning them in a calorimeter. Explain why this is feasible as a method for determining food value.

Problems

1. If a steel meter stick is precisely one meter long at 20°C, how much error will be made by using it for measurement at the freezing temperature, 0°C? (α=0.000013/^{0}C).

2. How much heat is required to raise the temperature of 250 g of water by 10°C? (250 g is about one cup.) (c=1 cal/g^{0}C)

3. How much heat is required to raise the temperature of 250 g of brass by 10°C? (c=0.09 cal/g^{0}C)

Problems

4. Suppose a 75 kg man consumed a normal diet of 2500 kcal, which was completely released in the form of heat. If the body were assumed to have a specific heat of $c=0.8$ cal/g°C and no heat were allowed to escape from the body, by how much would the temperature of the body rise?

Problems

5. A certain calorimeter has a heat capacity equivalent to that of 5 000 g of water. When a 50 g slice of white bread is burned in the calorimeter, the water temperature rises from 25°C to 51.6°C. Calculate the heat of combustion of white bread in kcal/g.

Answer: Q/M=2.66kcal/g.