

## Common Water

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THERE is quite a lot of water in the world, either moving or lying on its surface, or soaked a little way into it, or in the air above it. There is so much of it about that we generally find it rather uninteresting, and do not consider that, so far as we are concerned, it is the most important chemical compound (or perhaps mixture of chemical compounds) in the universe.

Just to touch on the physiological side first of all. We *live* as human beings because our frames carry a multiplicity of muscles, the so-called "meat" of animals, those engines which are capable of carrying on work within the body, and of enabling the body to do work external to itself; even whilst you lie back with your eyes closed listening to the rain, even if you drop right off to sleep, a vast number of little engines are consuming fuel within your body and keeping you alive. No one of these engines can use anything but liquid fuel, and the liquid fuel is produced by mixing other chemicals, either solid or liquid, with water, and effecting some chemical changes in the mixture. This is then available for transfer in the muscles into other chemicals, with a loss of potential energy which you utilise; and so you live.

It is quite useless to offer the muscles in your arm a bun when you go out to bowl; it would be quite useless to expect muscles in your legs and thighs to do anything directly with a loin chop before playing football—but provided that, with the aid of water, you can convert the bun or chop into a suitable liquid fuel in sufficient time, you can make use of them in the game. Any liquid will not do, you must have water to prepare it; and the total amount of water in your body must be kept up. You can go a long time without food, particularly if you restrict your external output of energy by remaining at rest; but you can go only a very short time without water, even if you conserve your body supply as much as possible.

You cannot help losing water from your body; apart from the necessary wastage of that filtered through the kidneys, which carries off certain unwanted soluble salts to be excreted, evaporation is going on from the surfaces (including the lungs) all the time. That is, you are continually changing water in the liquid form to water in the vapour form; in breathing, for example, you take in air which is not saturated with moisture, and breath it out again nearly saturated, so that you change about one and a half pounds of water into vapour by evaporation in the lungs every day.

As a pint of water weighs a pound and a quarter, you pant out over a pint of water, on an average, every day.

What amount you lose from your skin by evaporation depends upon what you are doing. The engines must not run hot—they have to be controlled within very narrow limits—and you cannot sit under a shower or swim in cooling streams all the

time. When we *boil* a liquid, it passes from its liquid to vapour stage at a temperature depending on the liquid and the air pressure, and requires the constant supply of heat to effect this change from liquid to vapour; but *whenever* a liquid goes to a vapour, even if not near the boiling point, heat is taken up to effect the change. Water is even evaporating over the surface of ice. Every gramme of water evaporated at ordinary body temperatures requires about 580 calories, so that by regulating the supply of water to the exposed surface of the body more or less heat can be taken off as required. If you do much heavy labour on a hot and humid day, then your body is so concerned with presenting a maximum amount of water to evaporate that it is exuded in large quantities. The heat taken up by one gramme of the liquid in evaporating without change in temperature is called its latent heat; the latent heat of vaporization of water at body temperature is 580 calories per gramme. If for some reason we had been designed to operate on, say, chloroform—if all our seas and rivers and rains and the liquid part of us were chloroform—then we would have to evaporate ten grammes to produce the cooling effect produced by the evaporation of one gramme of water; the latent heat of vaporization of chloroform is only about  $\frac{1}{10}$ th of that of pure water, which has an outstandingly high value. If we exuded ethyl alcohol, we would require to evaporate some two and a half times as much as in the case of water.

We are adapted to deal with water, and we run on water control, but water for that reason must not be looked upon as an uninteresting material. It has such outstanding properties. Naturally, you can take your water on board in the form of milk, ginger beer, or other suspensions or solutions, or under many conditions you may *think* you are not taking on any—yet it is well distributed through most foods you eat; and of course the liquids you take as “drinks” are nearly all water.

In cold weather, when the external surroundings are very much below the 37° C. of the body, and when you are not tending to overheat the body by doing much external work, there is very little evaporation from the skin. But the loss of water from the lungs goes on, and has to be replaced; and, of course, a sufficiency must still be passed through the body to ensure that the kidneys are able to work efficiently.

This evaporation goes on from all wet surfaces, including that of the oceans; vast masses of water are daily vaporized, and go up into the atmosphere; vast masses as regularly fall to earth again as rain. This gives us one of our methods of harnessing the sun to give us our energy to do other work—some of the energy given up in vaporizing the water and raising it to higher atmospheric levels can be utilized by harnessing mountain streams and waterfalls as the condensed water falls back to its ocean levels once again. A one inch rainfall deposits a million tons of water over every fifteen square miles; when we make use of the fall of that small amount over even that small catchment, we have available over two thousand million foot pounds energy for every foot drop in level in the water to the sea.

As I said at the beginning, we must not look upon water as being uninteresting because it is abundant. The very fact that it is an “ordinary” substance should make us suspect its enormous interest and importance to us. Probably absolutely

pure "water" has never been met, even in the laboratory. One can, with great care, prepare water which is chemically pure, but which contains traces of impurities, detectable by physical methods, which are negligible for nearly all purposes. Ordinary water contains a lot of impurities—dissolved salts, which in our local drinking water are not sufficiently pronounced to be noticed in drinking, but which are readily made apparent by chemical tests—and organic impurities, which are a good indication of the relative safety of the water for drinking purposes. This million tons of water for a one inch rainfall over a fifteen square miles catchment is more frequently of interest to us in Australia as something to drink, to bathe in, or to give to plants than as a source of energy in flowing to the sea; so we have to face the fact that it may pick up lots of things, alive or dead, on the way. Most water supplies for towns pass the outflow stream through filter beds of gravels and sands to remove impurities which have not actually gone into solution; and then the waters are tested physically and chemically in their later journeys to ensure that the difference between chemically pure "water" and the liquid that issues from our taps is not such as to cause us concern. Tests for bacteria and other living matter present are also made; one of the simplest tests for the contamination of water by bacteria is a chemical test of the nitrogenous organic matter originally present; free ammonia in water, even to the extent of five parts in 100 million, gives rise to suspicion as to its contamination.

This material, water, used once to be regarded as an "element", even when the word had obtained its present scientific significance, and even after it had been decomposed by different methods. Later, Cavendish burnt pure hydrogen in air, and formed water; then it was recognized as being more complicated, and a compound. By mass, 2.015 parts of hydrogen unite with 16.000 parts of oxygen to form water. By volume, two volumes of hydrogen unite with one volume of oxygen.

Further complications arose; whilst many of us display our knowledge of chemistry by talking about taking a drink of  $H_2O$ , or going for a swim in the  $H_2O$ , we should really take cognisance of the fact that water may not even be a simple compound—it is possibly a mixture of compounds. (This discussion of possible "structure" of "water molecules" that follows is strongly opposed by many chemists, and there are other theories.)

There is certainly  $H_2O$ , a molecule consisting of an atom of oxygen and two atoms of hydrogen, present, but there is possibly much more of the molecules  $H_2O_4$  and  $H_6O_3$ . Pure steam, as a gas, is probably  $H_2O$ . Pure ice, the solid, is possibly  $H_6O_3$ . The ice molecule would thus be a relatively big packet, and one would have to supply heat energy to melt the ice—that is, to break down some of the  $H_6O_3$  into  $H_4O_2$ . Below  $4^\circ C$ . there is probably a very large percentage of the more bulky packets—the water actually begins to expand from  $0^\circ C$ ., but the effect of the breaking down of those bulky molecules is to produce an actual contraction in volume until  $4^\circ C$ . is reached, after which expansion wins, and the water expands with rise in temperature till it reaches  $100^\circ C$ ., when, under standard pressure, it turns to steam. Thus ice is less dense than water, and we have a point of maximum density at  $4^\circ C$ . One litre (approximately 1,000 c.cs.) of ice at  $0^\circ C$ . has a mass of 917.5 grms.; one litre of water at  $0^\circ C$ . has a mass of 999.87 grms.; one litre of water at  $4^\circ C$ . a mass

of 1000·027 grms. ; one litre of water at 10° C. a mass of 999·73 grms. ; and one litre of water at 100° C. has a mass of 958·5 grms.

The change in volume when the compound ice changes to the mixture of compounds, water, is considerable ; the difference in mass between a litre of ice and a litre of water, both at 0° C., is over 82 grms. ; the mass of a litre *increases* by a *little more*—by 0·16 gm.—at 4° C., and then decreases in mass till at 100° C., the boiling point at standard pressure, the litre of water is 42 grms. less mass than when at maximum density. Bulky packets take up more room for the same mass.

Different research workers have tried to determine the percentages of these different types of molecules present in water. Traube suggested that water contained some simple H<sub>2</sub>O molecules, which he called "gasogenic" molecules, because undoubtedly they are the molecules comprising the gas "steam", and "liquidogenic" molecules, which were complex molecules of the general formula H<sub>4</sub>O<sub>2</sub> and H<sub>6</sub>O<sub>3</sub> ; but probably most chemists today incline to a theory (too much to put here) due to Bernal and Fowler.

All this is of interest in endeavouring to explain the peculiar physical properties of water—its high latent heats of fusion of ice, and of vaporization ; and its unusual changes in density with temperature. We know also that the actual chemical properties of water vary with temperature, and chemists have enjoyed themselves for many years assigning different positions to the atoms in the complex molecules which we have here called H<sub>6</sub>O<sub>3</sub> and H<sub>4</sub>O<sub>2</sub>.

There are other complexities yet to come, but they may be deferred for a few minutes whilst we consider ice, the solid formed when water freezes. We have to subtract energy to the extent of 80 calories, or 336 joules, from every gramme of water at 0° C. to convert it to the gramme of ice at 0° C. We say that the latent heat of fusion of ice is 80 calories per gramme. Examining a pond of water, we find that the top layers of water are cooled, become denser, and sink in freezing weather, until the surface reaches 4° C. From this temperature further cooling decreases the density ; so that the pond stirs itself up only till the whole body of water reaches 4° C., after which the top layer stays in place and cools down to 0° C., after which a further removal of heat causes a skin of ice to form. Water is a bad conductor of heat, ice is nearly as good an insulator, so that in the absence of stirring heat is subtracted but slowly from the water below the ice surface, and it takes a very prolonged freezing of the upper surface, with an atmosphere below freezing point, to cause a pond to freeze to any considerable depth. This is of considerable importance to pond life, and consequently to all other life.

The freezing point of water *decreases* with *increase* of pressure, but it requires big changes in pressure to cause appreciable effect—the change is only about  $\frac{7}{1000}$ th of a centigrade degree per atmosphere. Even such a slight effect is of economic importance—glaciers slide past obstacles in their paths because the increase in local pressure produces a lowering in the freezing point so that melting occurs at that place, so that the glacier slides past or over the obstacle ; on the removal of the extra

pressure caused by the obstruction the water is a liquid at a temperature at which it should be a solid for the present pressure, and it refreezes. This is termed regelation.

The solubility of materials in water varies with temperature—ordinary tap water has a lot of air dissolved in it. As you know, continued boiling will drive off most of the dissolved gases. If water be cooled very slowly, then air is given off because it is less soluble in a given volume at lower temperatures; but an ordinary block of ice is opaque because the cooling is so quick that the air coming out of solution cannot escape, but is trapped as little bubbles throughout the body of the ice.

The whole subject of what happens to materials when they are dissolved in water is of great interest and economic importance, but is too lengthy a matter to introduce here.

Ice crystallises in the hexagonal system, and the grouping of the atoms and molecules within the crystal has been investigated by many workers—notably by Sir W. H. Bragg—and is very interesting. Very big pressures—thousands of atmospheres—cause definite regrouping of the molecules to form different types of ice, which are all heavier than water. That was not thought to be of much economic value, because the pressures were not such as to occur naturally, but a recent discovery has led to the reopening of the question. Greenland's icy mountains are now found, by geophysical survey, not to be ice-covered rock, but to be vast ranges of ice. The pressure at the bottom should be so great that the freezing point would be lowered below what is thought to be the ice temperature there; so that Greenland's icy mountains should have behaved like the glaciers, and have slid off the land and the probable rim around it, into the sea. The problem is a very interesting one, and is being investigated.

Sea water has, in the most favourable circumstances, to be cooled to  $-2^{\circ}$  C. before it freezes; the ice formed is nearly free from salt. Icebergs are usually masses of ice that have broken off the ends of glaciers, or from a polar ice sheet. You can imagine a glacier in the polar regions poking its nose beneath the surface of the water, until the total upthrust—Archimedes again—is sufficient to break it off, and away it floats. Only  $\frac{1}{9}$ th of the mass of the ice is above water. Many icebergs of glacial origin carry soil and rocks, which they strew along the ocean floor as they drift away from polar regions in the ocean currents.

I mentioned that water was a thermal insulator—that is, it is a bad conductor of heat; copper, a good conductor, transmits heat nearly 1,000 times as well. *Pure* water is also a very bad conductor of electricity, but many substances on dissolving in water dissociate, or split up into electrified particles called ions, so that the solution will then permit the passage of electricity much more readily. Water spilt on surfaces, or which has soaked through materials, is fairly certain to have dissolved materials dissociated in it, so that water outside the officially pure water in bottles in the laboratories must always be regarded as a conductor of electricity. In other words, *pure* water is a laboratory curiosity. If it be subjected to pressure, water decreases in volume; but its molecules are already close to one another, so that big pressures

are required to produce small changes in volume; the change in volume of a unit volume, for one atmosphere change in pressure, is called the "compressibility per atmosphere". This changes with the pressure applied, and is not a constant; for small pressures up to 25 atmospheres, which is 370 pds. wgt./sq. inch, the value is about  $50 \times 10^{-6}$ , that is 50 millionths. The pressure increases by about one atmosphere for every 30 ft. we sink in the sea, so if pushed 30 ft. down, a cubic foot at the top would be decreased only by 50 millionths of a cubic foot. You can see that the change in density would therefore be negligible. At a pressure of 1,000 atmospheres, the compressibility of atmosphere has decreased to  $40 \times 10^{-6}$ , and at 3,000 atmospheres to  $30 \times 10^{-6}$ . These are figures at  $15^\circ \text{C}$ .; the values vary with temperature.

The deepest ocean depths are about 30,000 feet; the pressure down there would be the uncomfortable one of about a thousand atmospheres. The change in volume per unit volume would be  $40 \times 10^{-6} \times 1000 = 40 \times 10^{-3}$ , or  $\frac{1}{400}$ th. So the cubic foot

would only be squashed by  $\frac{1}{400}$ th of a cubic foot, and the density would therefore have increased approximately by one part in 400—a very small change. You can understand that the stories of ships sinking in the sea until the density of the water is so great that they sink no further are quite absurd—water is so incompressible that the change in density is very slight even in the greatest ocean deeps.

I have not yet referred to the complication of the presence in water of the oxide of our isotope hydrogen, the new heavy "water" deuterium oxide—there was an article on that in ENVIRONMENT, Vol. I, No. 2—but I hope that I have given enough to indicate that water is a very interesting and important material.

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