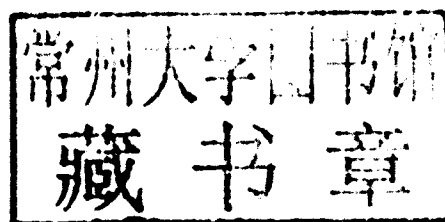# RIVER ECOSYSTEM ECOLOGY

EDITOR

GENE E. LIKENS

RIVER ECOSYSTEM ECOLOGY: A GLOBAL PERSPECTIVE

A DERIVATIVE OF ENCYCLOPEDIA OF INLAND WATERS

EDITOR

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EDITOR

Professor Gene E. Likens is an ecologist best known for his discovery, with colleagues, of acid rain in North America, for co-founding the internationally renowned Hubbard Brook Ecosystem Study, and for founding the Institute of Ecosystem Studies, a leading international ecological research and education center. Professor Likens is an educator and advisor at state, national, and international levels. He has been an advisor to two governors in New York State and one in New Hampshire, as well as one U.S. President. He holds faculty positions at Yale, Cornell, Rutgers Universities, State University of New York at Albany, and the University of Connecticut, and has been awarded nine Honorary doctoral Degrees. In addition to being elected a member of the prestigious National Academy of Sciences and the American Philosophical Society, Dr. Likens has been elected to membership in the American Academy of Arts and Sciences, the Royal Swedish Academy of Sciences, Royal Danish Academy of Sciences and Letters, Austrian Academy of Sciences, and an Honorary Member of the British Ecological Society. In June 2002, Professor Likens was awarded the 2001 National Medal of Science, presented at The White House by President G. W. Bush; and in 2003 he was awarded the Blue Planet Prize (with F. H. Bormann) from the Asahi Glass Foundation. Among other awards, in 1993 Professor Likens, with F. H. Bormann, was awarded the Tyler Prize, The World Prize for Environmental Achievement, and in 1994, he was the sole recipient of the Australia Prize for Science and Technology. In 2004, Professor Likens was honored to be in Melbourne, Australia with a Miegunyah Fellowship. He was awarded the first G. E. Hutchinson Medal for excellence in research from The American Society of Limnology and Oceanography in 1982, and the Naumann-Thienemann Medal from the Societas Internationalis Limnologiae, and the Ecological Society of America's Eminent Ecologist Award in 1995. Professor Likens recently stepped down as President of the International Association of Theoretical and Applied Limnology, and is also a past president of the American Institute of Biological Sciences, the Ecological Society of America, and the American Society of Limnology and Oceanography. He is the author, co-author or editor of 20 books and more than 500 scientific papers.

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INTRODUCTION TO RIVER ECOSYSTEM ECOLOGY: A GLOBAL PERSPECTIVE

Rivers, streams, brooks, runs, forks, kills, creeks, are among the many names for lotic (running or fluvial) ecosystems within the landscapes of the Earth. These systems facilitate the gravitational transport of water, dissolved substances, and large and small particulate materials downstream through a diversity of types of drainage networks from relatively simple channels to highly complicated "braided" channels, both above and below ground (e.g. Allan and Castillo, 2007). The tight connection in terms of structure and function between the river and its drainage basin (catchment=European usage or watershed=American usage) has been the subject of detailed study for many decades (e.g. Hynes, 1975; Likens, 1984; Allan and Castillo, 2007). The drainage area bordering the stream is called the riparian zone and is of critical importance to the function, as well as the protection and management of a river (e.g. Naiman et al., 2005).

Nevertheless, rivers and streams are far more than channels transporting water, chemicals and sediments downstream. They function as ecosystems (e.g. Fisher and Likens, 1972, 1973) with all of the varied and complicated activities and interactions that occur among their abiotic and biotic components, which are characteristic of all ecosystems (e.g. Allan and Castillo, 2007). Thus, they are not functioning just as "Teflon pipes" in the landscape that many have assumed in the past.

Rivers and streams comprise about 0.006% of the total fresh water on the Earth (Likens, 2009b), but like lakes, reservoirs and wetlands are valued by humans far out of proportion to their small size, as these systems supply diverse drinking, irrigation, waste removal, food, recreation, tourism, transportation and aesthetic services. Rivers with the largest volume of fresh water in the world, like the Amazon, Congo, Yangtze and Orinoco, are located in the tropics or semi-tropics. In fact, some 25% of the freshwater flow to the oceans of the world comes from two rivers, the Congo and the Amazon Rivers, both at approximately the same latitude (Likens, 2009b).

This volume consists of 5 sections: 1. Introduction to River Ecosystems; 2. Physical and Chemical Processes Influencing Rivers; 3. Ecology of Flowing Waters; 4. Human Impacts on Streams and Rivers; and, 5. Rivers of the World.

The articles in this volume are reproduced from the Encyclopedia of Inland Waters (Likens, 2009a). I would like to acknowledge and thank the authors of the articles in this volume for their excellent and up-to-date coverage of these important riverine topics.

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CONTENTS

HUMAN IMPACTS ON STREAMS AND RIVERS

RIVERS OF THE WORLD

PHYSICAL AND CHEMICAL PROCESSES INFLUENCING RIVERS

Contents

Physical Properties of Water

K M Stewart, State University of New York, Buffalo, NY, USA

Introduction

Water is an indispensable and remarkable substance that makes all forms of life possible. Speculation about possible past or present life on other planets within our solar system, or on any extraterrestrial body somewhere within the universe, is conditioned on the evidence for or against the existence of past or present water or ice. Humans can and did survive and evolve without petroleum products (gas and oil) but cannot survive and evolve without water. Water is the most important natural resource.

By far the greatest volume ($\sim$76%) of water on Earth is in the oceans. A smaller fraction ($\sim$21%) is found within sediments and sedimentary rocks. A still smaller fraction ($\sim$1% of the overall volume) is freshwater, and of that 1%, about 73% is in the form of ice (mostly contained within the Greenland and Antarctic ice caps), and only about 23% of that 1% is liquid freshwater. If we consider further that about one-fifth of the world's liquid freshwater is contained within the five St. Lawrence Great Lakes in North America, and another approximately one-fifth is contained within the deepest freshwater lake on Earth, Lake Baikal, in Russia, we are left with an unevenly distributed resource. It is obvious that if the expanding human populations around the world do not conserve and manage this precious resource very carefully, they put themselves at great peril.

Liquid water can be formed through some hydrogen bonding and electrostatic attraction of two slightly positively charged atoms of the gaseous hydrogen (H) and one slightly negatively charged atom of the gaseous oxygen (O) to form one molecule of water (H_2O). **Figure 1** provides two views of that polar molecule. **Figure 1(a)** and **1(b)** show the somewhat lopsided or asymmetrical arrangement of two smaller hydrogen atoms, separated by an angle of $\sim$105°, and a larger oxygen atom. **Figure 1(a)** is a simple 'ball and spoke' representation whereas **Figure 1(b)** shows the shared electron orbits, positive ($+$) and negative ($-$) poles, and the number (eight each) of protons and neutrons in the nucleus of the oxygen atom.

The relative elemental simplicity of water is somewhat deceptive because of the great influence that some of the unusual properties of water have on the physics, chemistry, and biology of the world generally, and on the distribution of life specifically. The following discussion will describe briefly some of these unusual properties and provide examples of how these properties may help us understand the world of inland waters.

Density

Density may be simply defined as the amount of weight or mass contained in a specific volume. If the volumes of all substances could be standardized to one size, e.g., one cubic centimeter (cm^3), then a measure of the weight or mass in that fixed volume gives the density. **Table 1** lists a few comparative densities (rounded to two decimals) of two liquids (water and mercury) and some selected solids.

Density differences in inland waters may be caused by variations in the concentrations of dissolved salts,
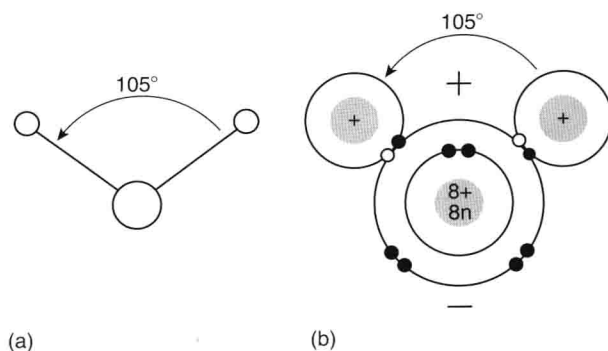
Figure 1 Two schematic representations (a) and (b) of a water molecule. (Modified from various sources.)

Table 1 Some comparative densities of water and other substances or elements

Substance	Densities (g cm^{-3})
Wood	
Seasoned balsa	0.11–0.14
Seasoned maple	0.62–0.75
Seasoned ebony	1.11–1.33
Water	1.00
Calcium	1.55
Aluminum	2.70
Iron	7.87
Lead	11.34
Mercury	13.55
Uranium	18.95
Platinum	21.45

Information from multiple sources.

Table 2 Comparative densities of average ocean water (salinity ~35%), freshwater ice, and pure distilled water at different temperatures

Temperatures ($^\circ$C)	Densities (g cm^{-3})
20.0	1.02760, ocean water (salinity 35%)
0.0	0.9168, freshwater ice
0.0	0.99987, pure water (from here on)
2.0	0.99997
3.98–4.00	1.00000
6.0	0.99997
8.0	0.99988
10.0	0.99973
12.0	0.99952
14.0	0.99927
16.0	0.99897
18.0	0.99862
20.0	0.99823
22.0	0.99780
24.0	0.99733
26.0	0.99681
28.0	0.99626
30.0	0.99568
32.0	0.99505

Values from Hutchinson (1957), Pinet (1992), and Weast and Astle (1979).

by changes in the water temperature, and in pressure. For the vast majority of inland lakes, only vertical differences in salt concentrations and temperatures are of significant influence to mixing processes. Fixed or uniform additions of salts to the water tend to cause *linear increases* in the density of water. In contrast, fixed or uniform changes in the temperature (both below and above 4 °C) of water cause *nonlinear changes* in the density of water (see **Table 2**). The density of pure water is maximum at a temperature of 4 °C (3.98 °C to be precise). It is at this temperature that the interatomic and intermolecular motions and intermolecular distances of water molecules are least. One consequence of this reduction is that more molecules of H_2O can fit into a fixed space at 4 °C than at any other temperature. This compaction allows the most mass per unit volume and thus the greatest density. It is especially noteworthy that the temperature at which water has the maximum density is above its freezing point.

Because the differences in densities, within a few degrees above and below 4 °C, are very slight, it takes

relatively little wind energy to induce substantial vertical mixing when water temperatures are within those ranges. An example period, for those lakes that become covered with ice in the winter, would be shortly before an ice cover develops and shortly after the ice cover departs. However, it takes much more energy to cause extensive mixing when the density differences are high, such as is common between the usually warm upper waters and colder lower waters of Temperate Zone lakes during summer. The greater the top-to-bottom differences in temperature, the greater the top-to-bottom differences in density and, consequently, greater are the energies required for wind-induced mixing.

There is an old, but still valid, cliché in the northern hemisphere that '... it is cold up north and warm down south.' Water temperatures in more northerly Temperate Zone lakes tend to average cooler than those of more southerly tropical lakes. Interestingly, although the upper-water summer temperatures in tropical lakes are somewhat higher than those of Temperate Zone lakes, the lower-water temperatures in tropical lakes are substantially higher than those ordinarily found in the lower waters of Temperate Zone lakes. It might there fore seem that there would be an easy top-to-bottom mix of the water in tropical lakes. Indeed some shallow tropical lakes, with only slight top-to-bottom temperature differences, may have this. However, because of the nonlinear increases in water density with temperature, tropical lakes can be surprisingly stable and resistant to much vertical mixing.

Table 2 provides a listing of some comparative densities. Let us consider two hypothetical lakes with just a $2\,^\circ C$ spread between their lower and upper waters. For example, if a Temperate Zone lake in the spring, not long after the ice departed, had lower and upper waters of 4.0 and $6.0\,^\circ C$, respectively, the density difference would be $1.00000 - 0.99997 = 0.00003\,g\,cm^{-3}$. In contrast, a warmer tropical lake whose lower and upper water temperatures may be 26.0 and $28.0\,^\circ Cs$ would have density differences that are much greater $(0.99681 - 0.99626 = 0.00055\,g\,cm^{-3})$. Thus, the top-to-bottom ratio or density difference of these two lakes with a temperature difference of just $2\,^\circ C$ would be 55/3 or ~ 18 times as great in the tropical lake as in the Temperate Zone lake. The example above is only hypothetical but it shows the nonlinear influence of density changes with temperature, a property of water that influences, to varying degrees, the stratification and mixing of lakes around the world.

Heat Capacity/Specific Heat

Heat is a form of energy and, as such, we can measure changes in the temperature of a given volume of a substance and determine its heat capacity. Water is the common standard used and its heat capacity (arbitrarily defined as the heat needed to increase the temperature of 1 g of water by $1\,^\circ C$) is comparatively large. When the mass is also considered then the number of calories needed to raise 1 g of a substance by $1\,^\circ C$ is termed its *specific heat*. For water, the value is $1\,cal\,g^{-1}$. That quantity may not seem like much but, compared to other materials, the heat capacity or specific heat of water $(1.00\,cal\,g^{-1})$ and ammonia $(1.23\,cal\,g^{-1})$ are much greater than that of most other substances (**Table 3**). Consequently, these two liquids are commonly used to exchange heat in refrigerators and air conditioners.

Along with its ever changing and mesmerizing aesthetic qualities, inland waters are of immense importance in the storage and release of heat. In terms of freshwater lakes, the influence of their heat capacity can be seen most easily around very large lakes located in Temperate Zone latitudes and more inner continental areas. It is in these areas that even larger swings in seasonal air temperature would ordinarily occur in the absence of those lakes. Parts of the immediate surrounding areas of Lake Baikal in Russia (this is actually the world's deepest freshwater lake as well as one with the greatest volume of water) and the five St. Lawrence Great Lakes of North America are prime examples of the 'thermal buffering' these large lakes provide to their surroundings because of their large heat capacity.

For humans, this may mean some 'beneficial economic consequences' as portions of a lake's heat capacity are slowly released or 'shed' to down-wind regions as the fall and winter seasons progress. The immense thermal capacity of Lake Baikal is such that the lake and its immediate environments are roughly $10\,^\circ C$ warmer in December and January, and about $7\,^\circ C$ cooler in June and July, than in the cities of Irkutsk (about 50 km to the west of the southern half of Lake Baikal) and Ulan-Ude (about 70 km to the east of the lake). Several coastal and near-coastal regions of the St. Lawrence Great Lakes also provide impressive beneficial evidence of the influence of the Great Lake's heat capacity. There may be reduced costs associated with home and business heating in some coastal regions. An extended or milder autumnal period permits greater production in near-shore plantations of fruit trees and vineyards. Economic benefits may also accrue in some coastal regions of higher terrain during winter, when enhanced snows permit additional winter skiing, snowmobiling, and other winter sports.

However, some influences of a lake's heat capacity have 'detrimental economic consequences'. There are costs involved with snow removal, increased vehicular accidents (because of slippery roads), the corrosion of cars (attributable to road salts), and the potential long-term ecological changes associated with lake and stream salinization. There are also greater heating costs in spring as cooler water bodies extend their cooling influence inland. In late fall and winter, before an ice-cover develops, heavy snows may result when water vapor, being formed by evaporative processes off a relatively warm lake, is buoyed into much colder Arctic air (northerly Temperate Zone) crossing the lake. The rising water vapors may freeze, coalesce to ice crystals, and be carried down wind to shore areas where they fall out as snow. Perhaps the most dramatic of all the detrimental consequences is seen following the sometimes paralyzing effect of occasional, but intense, 'lake-effect'

Table 3 The specific heat $(cal\,g^{-1})$ of selected substances compared to that of ice, pure water, and ammonia

Aluminum	0.215
Copper	0.092
Gold	0.030
Lead	0.030
Silver	0.056
Zinc	0.092
Ethyl alcohol	0.60
Ice (at $0\,^\circ C$)	0.51
Water	1.000
Ammonia	1.23

Information from multiple sources.

snow storms of mesoscale proportions. The lake-effect snow storms tend to have their greatest impact at the downwind end of the St. Lawrence Great Lakes after very cold Arctic air ($\geq 13\,°C$ colder than the temperature of the lake) has moved across a long axis of the lakes and deposited its snows. These deposits or drops of snow may be in a broader synoptic pattern, but sometimes they are in very narrow bands of thick snow that may bring auto traffic, schools, and businesses to a stop. In the St. Lawrence Great Lakes region of North America, three of the better known areas where unusually heavy deposits of lake-effect snows may occur are (1) portions of the Upper Peninsula of Michigan on the southeastern shore of Lake Superior, (2) the southeasterly and easterly shores of Lake Ontario, especially the Tug Hill Plateau area of New York State, and (3) the easterly end of Lake Erie, around Buffalo, NY. Indeed, the St. Lawrence Great Lakes have been considered 'weather factories' capable of causing twists of climate found in few other parts of the world.

Heat of Fusion/Melting

This is just the amount of heat exchanged during a phase shift from either liquid water to solid ice, or from solid ice to liquid water. One gram of water at $0.0\,°C$ can be converted to ice at $0.0\,°C$ if 80 cal ($79.72\,cal\,g^{-1}$ to be precise) are released in the process. The same quantity, i.e., 80 cal, is required to melt that 1 g of ice back to 1 g of water. No further caloric additions or subtractions are needed to effect the phase shift.

Because of the heat needed to melt ice, researchers might intuitively expect to see a brief but substantial drop in the mean or weighted lake-water temperature when the ice cover of a lake melts in the spring season. For example, assume there is a hypothetical northerly latitude and a 20-m deep lake in late winter (March). Consider that the lake is covered with 50 cm of ice at $0.0\,°C$. Consider further that the weighted mean temperature of the 1950 cm (essentially 1950 g) water column below the ice is $3.0\,°C$. The heat content of that water column would be 5850 cal ($1950\ g \times 3\,cal\,g^{-1} = 5850\,cal$). Assuming that there are no further gains or losses of heat to the lake, the amount of heat required to melt the ice would be 3680 cal ($80\,cal\,g^{-1} \times 50\,cm$ of ice $\times\ 0.92\,g\,cm^{-1}$, allowing for density of pure ice rounded to two decimals $= 3680$ cal). If some of the caloric content of the water column could be used to melt all the ice, the total caloric content would drop to 2170 cal ($5850\,cal\ -\ 3680\,cal = 2170\,cal$). If those 2170 cal were now equally distributed within a $1\text{-}cm^2$ square and 20-m (2000 cm, essentially 2000 g) deep water

column, the mean water temperature would need to drop from 3 to $1.08\,°C$ ($2170\,cal/2000\,cal = 1.08\,°C$). A drop of about $2\,°C$ during the melting of ice would be large!

As it turns out, the hypothetical example in the above paragraph is not realistic. Some background follows. Many years ago as a graduate student, I took daily measurements of ice thickness and top-to-bottom water temperatures for two winters and right through the spring ice break up in a Midwestern U.S. lake. From conversations with others, I was told to expect, and did anticipate, a substantial drop in mean water temperature as the ice melted... especially in the last few days of ice cover when the ice thinned rapidly. However, I did not measure any big drops in lake temperature and, in retrospect, should not have anticipated them. The reasons researchers do not see large decreases in lake temperatures with ice loss reflect some interacting physics. For example, there may be somewhat differing weather patterns each spring. The ice generally melts over an extended period of time, from several days to several weeks, not suddenly. Half or more of the total ice thickness may be lost from the top of the ice by melting from warming air temperatures above the ice, not necessarily from waters that are just above freezing below the ice. Because of its *albedo* (*percent of incoming solar radiation that is reflected back into space*) dark or open water generally reflects only a small fraction of the incoming solar radiation, whereas white snow cover on a frozen lake can reflect a large fraction of incident radiation. Indeed, snow cover extending into the spring period can delay the date the ice disappears. However, with increasing amounts of solar radiation, rising air temperatures, melting snows, and darkening ice, the water below the ice may be gaining some heat from solar inputs at the same time it is losing some heat in melting an overlying ice cover. Moral of the story: Do not expect a big drop in mean water temperature as an ice cover melts on a lake.

Heat of Vaporization/Condensation

As was the case for 'Heat of Fusion/Melting,' the heat of vaporization/condensation also represents the amount of heat exchanged during a phase shift. For vaporization, it is the quantity of heat (540 cal g^{-1}) needed to convert 1 g of water to 1 g of water vapor. The same amount of heat is exchanged or released in the phase shift during the condensation of 1 g water vapor to 1 g of water.

Aquatic scientists may be naturally impressed with the large amount of heat exchanged ($80\,cal\,g^{-1}$) in the phase shift from water to ice, or from ice to water, but

the amount of heat exchanged ($540\,\mathrm{cal\,g^{-1}}$) in the phase shift from water to water vapor, or water vapor to water is 6.75 times larger ($540/80 = 6.75$). Although the importance of this large amount of heat exchange via vaporization or condensation may be underappreciated by humans, it is huge. On a small but critical scale for life, water evaporating off perspiring warm-blooded animals, including humans, helps maintain body temperatures within narrow survivable limits. On a global scale, the seemingly endless phase shifts between liquid water and water vapor in the atmosphere are key determinants in the redistribution of water and heat within the hydrological cycle around the world.

Isotopes

An isotope is one of two or more forms of the same chemical element. Different isotopes of an element have the same number of protons in the nucleus, giving them the same atomic number, but a different number of neutrons giving each elemental isotope a different atomic weight. Isotopes of the same element have different physical properties (melting points, boiling points) and the nuclei of some isotopes are unstable and radioactive. For water (H_2O), the elements hydrogen (atomic number 1) and oxygen (atomic number 16) each have three isotopes: ^1H, ^2H, and ^3H for hydrogen; ^{16}O, ^{17}O, and ^{18}O for oxygen. In nature, the ^1H and ^{16}O (usually just given as O) isotopes are by far the most common. In water, the water molecule may be given as ^1H_2O or hydrogen oxide, ^2H_2O or deuterium oxide, and ^3H_2O or tritium oxide, the radioactive one. Both of the latter two are sometimes called heavy water because of their increased mass. However, the phrase 'heavy water' gained notoriety primarily because of the association of ^2H_2O or deuterium oxide, also called the deuterated form of water, in the development of nuclear weapons. Many elements have isotopes, but the isotopes of hydrogen and oxygen are of particular interest because fractionation occurs in vapor–liquid–solid phase changes. Heavier molecular 'species' tend to be enriched in the condensation phase and lighter molecular 'species' in the vapor phase. Some isotopes can be used to great advantage as tracers in understanding water movements and exchanges within atmospheric, oceanic, lake, stream, and ground water systems.

Sublimation

Water is said to be sublimated, sublimed, or undergo sublimation when it passes directly from a solid (ice) stage to a gas (vapor stage) without becoming a liquid in between. The latent heat of sublimation, i.e., the heat required to make the form of water change from ice to a water vapor, is $679\,\mathrm{cal\,g^{-1}}$. This quantity is larger than the heat required to melt ice ($80\,\mathrm{cal\,g^{-1}}$) and vaporize water ($540\,\mathrm{cal\,g^{-1}}$) combined ($80 + 540 = 620\,\mathrm{cal\,g^{-1}}$). Because there may be multiple heat sources and sinks (e.g., the air above the ice and the water below the ice) associated with changing ice thickness on frozen Temperate Zone lakes, it is a challenge to assess the quantitative role that sublimation may play in those changes.

Some practical effects of sublimation may be visualized by observing a reduction in the volume of some dry ice (solid CO_2) or camphor. In another example, after several weeks of continuing subfreezing temperatures and deep frost, and assuming that no deicing salts were used, sublimation is most likely responsible for the slow disappearance of an ice sheet over the surface of a frozen sidewalk. Sublimation is also the main process by which wet clothes, which were hung out to dry in subfreezing temperatures, may dry. In the latter case, the water on the clothing quickly freezes to ice, but then slowly vaporizes through sublimation, and the clothes dry. In more recent years, freeze-dried vegetables, fruits, and other products (including instant coffee) provide other examples where the practical application of sublimation is utilized to both market and preserve food.

Surface Tension and Cohesiveness

Surface tension may be regarded as the resistance offered by liquid water to forces attempting to deform or break through the surface film of water. It is an interesting property and, for water, the surface tension measured in Newton's per meter ($\mathrm{N\,m^{-1}}$), is high and shows a slight *increase* as the temperature falls from 100 ($0.0589\,\mathrm{N\,m^{-1}}$) to $0\,^\circ\mathrm{C}$ ($0.0765\,\mathrm{N\,m^{-1}}$). The molecules of water are strongly attracted to each other through their *cohesiveness* (attraction of like substances). The properties of surface tension and cohesiveness work together in water in shaping the small rounded water droplets seen on a table top or a car windshield. The same properties help to form the slightly flattened to spherically-shaped raindrops as they fall through the air.

The primary force for restoring larger wind-generated surface and internal waves of lakes is gravity, but the primary force for restoring the much smaller capillary waves or ripples on a lake's surface seems to be surface tension of the water itself.

The surface tension of water is sometimes used to advantage in parlor games in which someone claims that he/she can float a more dense (than water) steel

needle on less dense water. When the needle is lowered slowly and carefully with its long axis paralleling the surface of the water, it may be possible to 'float the needle' because the high surface tension of the water may prevent the needle from sinking. Do not try this by lowering one of the sharp ends of the needle first because a point application of the needle will exceed the surface tension of the water film, and the needle will sink rapidly.

When responding to a 'fire call' in fire trucks, water is the most common and practical substance used by firemen. Water is cool, it suppresses heat, it puts out fires and sometimes there is much water to spare. However, the *high surface tension of water can reduce its effectiveness in suppressing some fires.* Surfactants are compounds that reduce the surface tension of water. In their response to a 'fire call' firemen often quickly attach hoses to street fire hydrants and spray water from that source on a burning structure. Although the addition of tiny quantities of surfactants to water may help put out fires, it is not practical (or safe) to add surfactants to an entire distribution system of a city. However, the addition of tiny quantities of surfactants to the volume (roughly 1.89 m^3 or 500 gallons in the United States) of water being carried in the fire truck would make that truck water 'wetter.' Some combustibles could be penetrated more easily by this wetter water of reduced surface tension and selected fires could be put out more rapidly.

There is a specialized community of organisms, sometimes called neuston, associated with the surface film. For many observers of nature, it is always fascinating to see small insects such as pond skaters or water-striders (*Gerris* sp., within the insect Order Hemiptera), and whirligig beetles (*Gyrinus* sp. and *Dineutes* sp., within the insect Order Coleoptera), running around on the surface of ponds, sheltered lakes, and some streams. Because of padded ends to the long middle and hind feet of water striders, and the much shortened but paddle-like feet of the whirligig beetles, the high surface tension of the water is such that the insects may dimple, but not break through, the surface film.

One of the easiest ways of getting popcorn into your mouth is by touching your tongue to some popcorn in a container. Here again it is the surface tension of the water on your tongue that lets you 'hold on' to the light popcorn easily.

Viscosity

This property may be thought of as the internal friction or resistance exerted on one substance (gas,

liquid, or solid) as that substance tries to flow or move through the same or another liquid. One way of visualizing the influence that liquids or semiliquids of progressively greater viscosities might exert would be to take three glass marbles (same diameter and density) and drop one in each of three similar-sized glasses, one glass containing water, one light oil, and one honey, all at the same temperature. The marble would descend quite rapidly in water, more slowly in the light oil, and very much more slowly in the glass of honey. In this example, honey would obviously exert the most friction or resistance to movement through it and have the greatest viscosity. Viscosity is usually measured in poises (N s m^{-2}) or centipoises (= 0.01 P). Water at 20 °C has a viscosity of 0.01002 P or 1.002 cP.

The rate of passive descent through a liquid reflects the density of the liquid itself as well as the surface area and density of the substance moving through it. Viscosity changes with water temperature in that viscosities decrease as water temperatures rise and increase as water temperatures fall. Many fish are powerful enough, slippery from mucous on their skin, and shaped so they can 'slip through' water relatively easily. In contrast tiny zooplankton, with multiple projections on their body, are ordinarily challenged as they attempt to move in any direction and particularly so when moving in cool waters.

Colligative Properties

These are the four special properties of water that are significantly altered or modified when solutes are added to and dissolve in water. The alterations or modifications of a colligative property (regarded as a binding property) may be predictable in dilute solutions when the number of solute particles is known. It is the number of solute particles, not their chemical nature, that determines the extent to which a property is modified.

The four colligative properties of water are *vapor pressure* (when water is in equilibrium with its own vapor), *osmotic pressure* (the pressure controlling the diffusion of a solvent across a semipermeable membrane), *boiling point* (the temperature at which water undergoes a phase shift to a gas), and *freezing point* (the temperature at which water undergoes a phase shift to a solid). Even at standardized pressures and temperatures, the extent to which a property is modified depends on the number of solute particles added. Generally, if we add a fixed number of solute particles of a sugar or salt to a liter of pure water, there would be some consequences. The vapor pressure would be lowered but the osmotic pressure would rise. The boiling point (also termed *boiling-point*

elevation) would be elevated a bit above the usual boiling point of 100.0 °C. In the latter case, a watery mixture with solutes (e.g., a well-mixed soup being heated for a meal) would have to get hotter than the boiling point of pure water before it would boil. The freezing point would be lower than 0.0 °C. A practical application of this (also termed *freezing-point depression*) is easily seen, in parts of the northerly Temperate Zone in winter, following the application of deicing salts to melt the ice and snow on roads and sidewalks. Although not a colligative property as such, a simple increase in physical pressure also lowers the melting point of ice ($\sim$0.007 °C/atm) and helps form snowballs (when the snow is not too cold) and form a lubricating layer of water under the blade of an ice skate.

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