**PDHonline Course C104 (2 PDH)**

---

# The Unique and Unusual Properties of Water

***Instructor: Charles D. Riley, Jr., PE***

**2020**

**PDH Online | PDH Center**

5272 Meadow Estates Drive
Fairfax, VA 22030-6658
Phone: 703-988-0088
www.PDHonline.com

An Approved Continuing Education Provider

# The Unique and Unusual Properties of Water

*Charles Riley, Jr. PE*
*Vice President, Research and Development*
*Hydro Service & Supplies, Inc.*
*PO Box 12197*
*Research Triangle Park, NC 27709*
*criley@hydroservice.com*

## Introduction

To the average person, water is a common and ordinary substance which is often taken for granted, that is until a drought threatens crops and drinking water supplies or a severe flood destroys life and property. Most people do not understand that without water and its unique and unusual properties, life as we know it on earth would not exist.

One of the earliest Greek philosophers, Thales of Miletus (640-546 B.C.), observed the universal nature of water. Thales believed that water was the basic element from which everything began – "the seeds of everything have a moist nature". The abundance of water was apparent, but it was also observed that water was the only substance naturally present on earth simultaneously in three distinct states or forms– solid, liquid, and gas. On a cold winter day, snow and ice cover a field, while water flows in a nearby stream and gaseous clouds float overhead.

## Forms of Matter

All substances exist in three distinct forms (solid, liquid, and gas), and the form of a substance at a given time is a function of the temperature and pressure. A solid is defined as matter with rigidity and definite shape, having a crystalline internal structure. By this definition, a substance like glass would be considered a highly viscous liquid because it lacks crystalline structure. Solids tend to resist external forces. Solids can be converted to liquids by heating. The freezing point temperature of pure water is $0^o$ C, at one atmosphere pressure. At temperatures below the freezing point, water exists as a solid - ice.

A liquid, in contrast to a solid, lacks rigidity and has no definite shape. It has a definite volume and conforms to the shape of the container in which it is stored. External forces will cause a liquid to flow. Water is a liquid between the freezing point temperature and the boiling point temperature,$100^o$ C, at one atmosphere pressure. Liquids can be converted to the gaseous phase by heating beyond the boiling point temperature.

A gas has neither rigidity nor definite volume. A gas conforms to the shape and volume of the vessel containing it. A gas greatly expands and contracts with changes in temperature and pressure and has the ability to readily diffuse into other gases.

**Boiling and Freezing Points**

Water has unusually high boiling and freezing point temperatures compared to other compounds with similar molecular structure. All other compounds with similar molecular structure are gases at ordinary temperatures. Water, with its lower molecular weight than similar compounds, would be expected to have lower boiling and freezing point temperatures. However, due to the polar nature of the water molecule and hydrogen bonding (discussed later in this article), the boiling point of water is a remarkable $100^oC$ and the freezing point is a remarkable $0^oC$. The boiling and freezing point comparisons of compounds with similar molecular structure to that of water are shown as follows:

| Compound | MW | Boiling Point | Freezing Point |
|---|---|---|---|
| Hydrogen Telluride ($H_2Te$) | 129 | $-2^oC$ | $-49^oC$ |
| Hydrogen Selenide ($H_2Se$) | 80 | $-42^oC$ | $-64^oC$ |
| Hydrogen Sulfide ($H_2S$) | 34 | $-60^oC$ | $-84^oC$ |
| Water ($H_2O$) | 18 | $100^oC$ | $0^oC$ |

**Solid Phase**

Generally, substances contract becoming denser with a decrease in temperature, and water is no exception. The density of pure water at $25^oC$ is 0.997 gm/ml. The density of water increases as the temperature decreases until maximum density (1.000 gm/ml) at about $4^oC$. In the metric system of measurement, the *kilogram* is defined as the mass of one liter of pure water at its greatest density. Between $4^oC$ and the freezing point at $0^oC$, an amazing thing happens that occurs with very few substances, water gradually expands becoming less dense. The density of ice at $0^oC$ is about 0.917 gm/ml The water molecules form tetrahedral (a four sided solid, each face an equilateral triangle) shaped ice crystals. Since the density of ice is less than that of liquid water, ice floats on water. About one-eleventh of the liquid volume is added at freezing.

It is very significant that ice expands and floats on water. The consequences of this action can be seen in broken water lines in the winter and potholes in the roads. In fact, the freezing and thawing action of water is largely responsible for the fracturing of rock and the formation of soils. Also, consider the consequences if lakes and streams froze from the bottom to the top – aquatic life would not even exist, and climate and weather patterns would be altered drastically.

**Heat Capacity**

Another remarkable property of water is its extremely high capacity to absorb heat without a significant increase in temperature. For example, the summer sun at the beach will increase the temperature of the sand to the point that it is too hot to walk on; however, the water temperature is cool to the touch. Both the sand and the water absorb the same amount of heat energy, but the temperature of the sand is higher than the water temperature. An empty iron pot hanging over a fire will glow red hot quickly, but if the pot is filled with water, the temperature increases gradually. The high heat capacity of water makes it a good coolant to use in condensers and automobile radiators to keep engines from over-heating.

The heat capacity of water is the standard for comparing the heat capacity of other substances. *Specific heat* is defined as the ratio of the heat capacity of a substance to the heat capacity of water. The *specific heat* of water is 1.0 BTU (British Thermal Unit) per pound per degree Fahrenheit (1.0 calories / gram / degree C). For comparison, the specific heat of sand is about 0.2, and the specific heat of iron is about 0.1. Therefore, water can absorb about five time the amount of heat of sand and about 10 times the amount of heat of iron for an equivalent increase in temperature.

The moderate climate in coastal areas is the result of the absorbing of huge amounts of solar heat energy by water during the day and the slow release of heat energy during the night. Inland areas away from the coast typically experience much wider temperature extremes. The vast oceans on earth (about 75 percent of the surface area) are responsible for tempering the climate on earth permitting life to exist.

*Specific heat* is a useful concept and finds application in heat transfer problems. When two bodies not in thermal equilibrium are placed next to each other, heat transfer will take place from one body to the other until thermal equilibrium is established. The amount of heat transfer, Q, is

$$Q = c\, m\, (T_2 - T_1)$$

where $c$ is the *specific heat*, $m$ is the mass of the body, and $T_2$ and $T_1$ are respectively, the final and initial temperatures of the body.

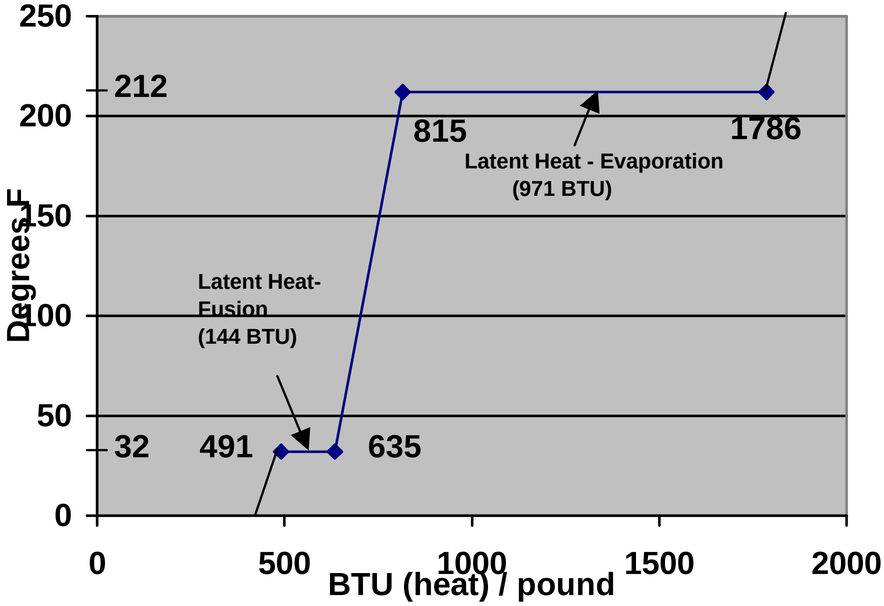
Example: Calculate the amount of heat required to increase the temperature of a flow of water at 10 GPM from $40^oF$ to $120^oF$.

$$Q = (1)\ \text{BTU/lb–}^o\text{F}\ (10)\ \text{gal/min}\ (60)\ \text{min/hr}\ (8.34)\ \text{lb/gal}\ (120 - 40)\ ^o\text{F}$$

$$Q = 400{,}320\ \text{BTU/hr}$$

**Latent Heats of Fusion and Evaporation**

Related to heat capacity is latent heat. Latent heat is the quantity of heat energy in BTU per pound or calories per gram absorbed or released by a substance undergoing a change in phase (liquid to solid and vice versa, and liquid to gas or vice versa) without a change in temperature. The latent heats of fusion and evaporation of water are unusually high as illustrated in the following graph.




In the above graph, the amount of heat added to water is referenced from $0^{o}$ Kelvin (absolute zero). The latent heat of fusion of water is 144 BTU per pound. In the process of freezing, water releases the same amount of heat (144 BTU / pound) that is absorbed in the process of melting.

The use of ice to keep drinks cold in an insulated cooler is a practical example of the use of the latent heat of fusion of water. In the process of melting, ice absorbs the heat energy in the drinks, thus keeping the drinks cool. A tub of water in a greenhouse on a cold winter night will moderate the temperature in the greenhouse because of the heat released by the water in the process of freezing.

The condensation of water vapor releases the same amount of heat absorbed by the water in the process of evaporating. As illustrated in the above graph, the latent heat of vaporization is over five times the amount of heat required to raise the water temperature from freezing to boiling. This large amount of stored heat energy is what makes steam heating so effective for space heating homes, factories, etc. The steam in the process of condensing to water releases this stored heat energy. An afternoon thunderstorm on a hot summer day is another example of the often violent release of heat energy in the upper atmosphere from the condensation of hot moist air. Even a hurricane, one of the strongest forces in nature, is an example of the effect of the redistribution of huge amounts of heat energy absorbed in the tropic oceans.

Evaporative cooling systems work just the opposite. Water in the process of evaporating absorbs the heat energy from the air, cooling the air. A water misting system on a hot, dry summer day keeps your skin cool because the water absorbs heat energy from the body in the process of evaporating.

Example: How much boiling water is required to melt 100 pounds of ice at $32^oF$ with the equilibrium temperature of the resulting water at $68^oF$? The heat lost by the boiling water must equal the heat gained by the ice,

$$m_w\, c\, \Delta T_w = L_i\, m_i + m_i\, c\, \Delta T_i$$

where $m_w$ is the mass of the boiling water, $m_i$ is the mass of the ice, $c$ is the specific heat of water, and $L_i$ is latent heat of fusion.

$$m_w\,(1)\,(212 - 68) = (144)\,(100) + (100)\,(1)\,(68 - 32)$$

$$m_w\,(144) = 14400 + 3600$$

$$m_w = 125 \text{ lbs.}$$

Example: How much steam, at $212^oF$, must be mixed with 100 pound of water to raise the temperature of the water from $32^oF$ to $212^oF$? The heat lost by the steam must equal the heat gained by the water,

$$L_s\, m_s + m_s\, c\, \Delta T_s = m_w\, c\, \Delta T_w$$

where $m_s$ is the mass of the steam, $m_w$ is the mass of the water, $c$ is the specific heat of water, and $L_s$ is the latent heat of evaporation.

$$(971)\, m_s + m_s\,(1)\,(212 - 212) = (100)\,(1)\,(212 - 32)$$

$$(971)\, m_s + 0 = 18000$$

$$m_s = 18.5 \text{ lbs.}$$

## Universal Solvent

A solvent is a substance capable of dissolving another substance (solute) to form a homogeneous mixture (solution) at the molecular level. The highly polar nature of water makes it an excellent solvent, especially for other polar compounds – salts, alcohols, carboxyl compounds, etc. More substances dissolve in water than any other solvent. More than half of the known elements can be found in water, some in high concentrations, and others only in trace amounts. For example, the saturation concentration of sodium chloride is about 36 grams per 100 ml, but the saturation concentration of calcium carbonate is about 0.0015 grams per 100 ml. The ability of water to dissolve a substance depends on the chemical composition, chemical bond strengths of the elements, temperature, and pH.

Non-polar compounds including most hydrocarbons are difficult to dissolve in water and dissolve in low or trace amounts. For example, oils tend to float on the surface of water.

## Surface Tension

With the exception of mercury, water has the highest surface tension of any other liquid. Surface tension is the attractive force exerted by the molecules below the surface on those at the liquid-air interface. This inward force tends to restrain the liquid from flowing. Polar compounds tend to have much higher surface tension than non-polar compounds. Hydrogen bonding of water molecules is attributed to the exceptionally high surface tension of water. The surface tension of water is 73 dynes per cm at 18$^{o}$C compared to ethyl alcohol at 24 dynes per cm.

Without external forces, a drop of water will pull itself into the shape of a sphere (geometric shape with the least surface area per unit volume) because of the surface tension of the water. This can be seen as a drop of water hangs from the spout of a water faucet. The high surface tension of water makes rain drops tiny bullets, that given enough time will erode rock. Objects heavier than water can actually float on water under the right conditions; insects can be seen walking on water or a razor blade will float on the surface.

Hydrogen bonding is attributed to the ability of water to adhere to or “wet” most surfaces; such substances are said to be *hydrophilic* (water-loving). Water appears to crawl up the side of glass and other containers. Other substances such as oils, fats, waxes, and many synthetics (polypropylene, etc.) will not “wet” with water; these substances are *hydrophobic* (water-fearing). Membrane filter cartridges for water filtration, made from *hydrophobic* polymers and with sub-micron pore sizes (< 1 micron), must be manufactured with wetting agents to lower the surface tension of the water to allow the water to “wet” the pores of the filter. Once the pores are filled with water, water will stay in the pores due to the surface tension; this is termed *capillary action*. Capillary action is responsible for the movement of water through soils, blood through blood vessels, and water carrying nutrients through the roots of plants.

### Medium of Life

Water is an essential ingredient for the existence of life as we know it. This explains the recent interest in discovering water in other parts of the universe. All known biochemical processes occur in aqueous environments. The composition of most living things is about 70 to 80 percent water by weight.

Water plays a significant role in the process of photosynthesis. In photosynthesis, plants utilize radiant energy from the sun to convert two inorganic substances, water and carbon dioxide, into carbohydrates.

$$6CO_2 + 6H_2O + 672 \text{ kcal} \rightarrow C_6H_{12}O_6 + 6O_2$$

Photosynthesis is the most basic and significant chemical reaction on earth. It supplies the primary nutrients, directly or indirectly, for all living organisms and is the primary source of atmospheric oxygen.

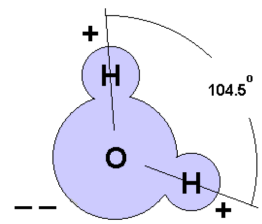
### Chemical Bonding

To understand the causes of the unusual properties of water, a basic understanding of chemical bonding and the structure of the water molecule is necessary.

The ability of *elements* (basic building blocks of all matter at the atomic level) to combine and form *compounds*, depends on the ability of their atoms to give up or acquire *electrons* (negatively charged subatomic particles that orbit the nucleus of the atom). Elements that tend to give up electrons become positively charged ions (*cations*); and elements that tend to acquire electrons become negatively charged ions (*anions*).

The power of an element to combine with other elements to form compounds is termed the *valence* of the element. Valence is a positive or negative whole number (based on the number of electrons gained or lost), and for inorganic compounds, the algebraic sum of the valence numbers of the combining elements is zero. For example, sodium readily gives up one of its electrons to become a *cation* with a valence of +1, and chloride tends to attract one electron to form an *anion* with a valence of –1. The oppositely charged ions attract forming a molecule of sodium chloride [(+1) + (-1) = 0]. The electrostatic attraction of oppositely charged ions to form a compound is termed *ionic bonding*.

Both of the elements that combine to make water, hydrogen and oxygen, exist separately in molecules containing two atoms each ($H_2$ and $O_2$). The two atoms are held together by sharing an electron pair in a chemical bond termed a *covalent bond*. Covalent bonds are much stronger than the ionic bond. The two atoms held together by the covalent bond make a molecule that is much more stable than the individual atoms. The chemical bonds in the water molecule are covalent bonds since the hydrogen atoms combine with the oxygen atom in shared electron pairs. It is the unique distribution of the electrons in the resulting chemical bond that causes the hydrogen atoms to bond with the oxygen atom at a bond angle of $104.5^{o}$.




The oxygen atom exerts a relatively strong pull on the shared electron pair causing the hydrogen atoms to become electropositive regions and the oxygen atom to become an electronegative region. Because the positive and negative regions are not evenly distributed around a center point, the water molecule is termed a *polar molecule*.

The polar nature of the water molecule causes it to become electrostatically attractive to other water molecules as well as other ions in solution and contact surfaces with electrostatic sites. The electropositive hydrogen atoms on the water molecule will be attracted to the electronegative oxygen atoms of adjacent water molecules. This “bridging” phenomenon is called *hydrogen bonding*. Hydrogen bonding is only about 10 percent of the strength of the covalent bond, but it is responsible for most of the unusual properties of water (high freezing and boiling points, high heat capacity, high heats of fusion and evaporation, solvency, and high surface tension).

Hydrogen bonding is responsible for maintaining the integrity of the water molecule during chemical reactions. While other compounds undergo chemical changes (ionization), the water itself will maintain its chemical integrity. In pure water, a relatively small number of molecules will ionize into hydrogen and hydroxyl ions. Thus, pure water is a relatively poor conductor of electrical current. The specific resistance of theoretically pure water is 18.3 megohm-cm, while most potable water supplies have a resistivity of less than 10,000 ohm-cm. Therefore, the purity of water can be readily measured with a conductivity or resistivity meter.

Hydrogen bonding is the reason for the lower density of ice relative to water. At freezing, the water molecules arrange themselves along the directional lines of the hydrogen bonds causing water to expand and become less dense. For this reason, ice floats on water. Increased pressure lowers the melting point of water. The pressure applied to ice by the blade of an ice skate melts the ice providing a layer of water for the ice skater to gracefully glide along the ice. Even at extremely cold temperatures, high pressure will weaken the crystal lattice; this is the reason that huge ice masses such as glaciers will gradually flow.

The polar nature of the water molecule causes the molecule to align in an electric or magnetic field. The electronegative oxygen atom aligns toward the positive pole, and the electropositive hydrogen atoms align toward the negative pole. Water has an exceptionally large *dipole moment* ($1.87 \times 10^{-18}$ e.s.u.) relative to most other inorganic compounds. Dipole moment is the product of the distance between the charges multiplied by the magnitude of the charge in electrostatic units (e.s.u.).

The *dielectric constant* is another property related to dipole moment. Water molecules, by aligning in an electric field, tend to neutralize the field and create resistance to the transmission of an electrostatic charge between charged bodies. The dielectric constant of a substance is defined by $\varepsilon$ in the following equation:

$$F = Q_1Q_2 / \varepsilon r^2$$

where F is the force between two charges Q, separated by the distance r in the medium. As the dielectric constant increases, the force between the charges decreases. For reference, the dielectric constant of a vacuum is 1.0; glass is 3.0; ethyl alcohol is 25; and water is 81. The high dielectric constant of water reduces the attractive force of ionized substances in water helping to account for the remarkable ability of water to dissolve a wide variety of substances.

## Conclusion

To the average person, water is an ordinary substance often taken for granted. Even though the cause of these unique and unusual properties is explainable at the atomic level, water is truly a remarkable substance. From our examination of these properties, it is evident that water is essential for life, as we know it, to exist on earth. Water is the mediator of chemical and biochemical processes. Water shapes our natural environment and even mediates our climate and weather.

## References

1. Murphy, Daniel B., and Viateur Rousseau. *Foundations of College Chemistry.* New York: The Ronald Press Company, 1969.

2. *Handbook of Chemistry and Physics*, 56th Edition. edited by Robert Weast. Cleveland: CRC Press.
3. Davis, Kenneth S., and John Arthur Day. *Water the Mirror of Science*. Garden City, New York: Doubleday & Company, Inc., 1961.
4. *The Condensed Chemical Dictionary*, 9th Edition. revised by Gessner G. Hawley. New York: Van Nostrand Reinhold Company, 1977.