Water and its Properties

Water

"Water is a chemical substance with chemical formula H_2O. (One molecule of water has two hydrogen atoms covalently bonded to a single oxygen atom.)"

Characteristics of water:

- Water is a odorless, tasteless liquid at ambient temperature and pressure.
- Liquid water has weak absorption bands at wavelengths of around 750 nm which cause it to appear to have a blue color.
- This can easily be observed in the water-filled bath or wash-basin whose lining is white. Large ice crystals, as in glaciers, also appear blue.

Water is the solvent of life

- All organisms are composed primarily of water, such as that most eukaryotic organisms are about 90 percent water while prokaryotes are about 70% water.
- No organism, not even the prokaryotes, can develop and grow without water. All reactions occur in liquid water.
- Water being polar has unique properties. These include its role as a solvent, as a chemical reactant, and as a factor to maintain a fairly constant temperature.

Bonds in water:

- Unlike other analogous hydrides of the oxygen family, water is primarily a liquid under standard conditions due to hydrogen bonding.

The molecules of water are constantly moving in relation to each other, and the hydrogen bonds are continually breaking and reforming at timescales faster than 200 femtoseconds (2×10^{-13} seconds). However, these bonds are strong enough to create many of the peculiar properties of water, some of which make it integral to life.

Water, ice, and vapour:

- Within the Earth's atmosphere and surface, the liquid phase is the most common and is the form that is generally denoted by the word "**water".**
- The solid phase of water is known as ice and commonly takes the structure of hard, amalgamated crystals, such as ice cubes, or loosely accumulated granular crystals, like snow. Aside from common hexagonal crystalline ice, other crystalline and amorphous phases of **ice** are known.

- The gaseous phase of water is known as **water vapor** (or steam. Visible steam and clouds are formed from minute droplets of water suspended in the air.

Water also form a supercritical fluid. The critical temperature is 647K and the critical pressure is 22.064 MPa. In nature this only rarely occurs in extremely hostile conditions. A likely example of naturally occurring supercritical water is the hottest parts of deep-water hydrothermal vents, in which water is heated to the critical temperature by volcanic plumes and critical pressure is caused by the weight of the ocean at the extreme depths where the vents are located. This pressure is reached at a depth of about 2200 meters: much less than the mean depth of the ocean (3800meters)

Heat capacity and heats of vaporization and fusion:

1-Heat capacity:

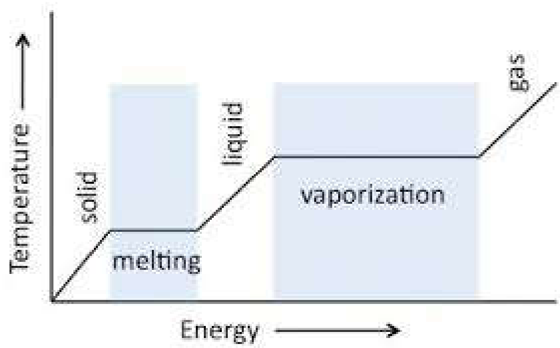
“Heat capacity or thermal capacity is a physical property of matter, defined as the amount of heat to be supplied to a given mass of a material to produce a unit change in its temperature. The SI unit of heat capacity is joule per kelvin. Heat capacity is an extensive property.”

2-Heat of vaporization:

"The enthalpy of vaporization, also known as the heat of vaporization or heat of evaporation, is the amount of energy that must be added to a liquid substance, to transform a quantity of that substance into a gas. The enthalpy of vaporization is a function of the pressure at which that transformation takes place."

3-Heat of fusion:

"The amount of **heat** required to melt one gram of solid at its melting point with no change in temperature. The molar **heat of fusion** is the amount of **heat** required to melt one mole of a solid at its melting point with no change in temperature and is usually expressed in kJ/mol."

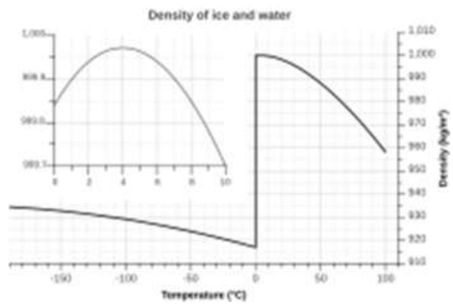


- Water stabilizes temperature in cells and ecosystems Water has a very high specific heat capacity of 4.1814 J/ (g·K) at 25 °C – **the second highest among all the hetero-atomic species (after ammonia)**, as well as a high heat of vaporization (40.65 kJ/mole or 2257 kJ/kg at the normal boiling point).
- Both of which are a result of the extensive hydrogen bonding between its molecules. These two unusual properties allow water to moderate Earth's climate by buffering large fluctuations in temperature. Most of the additional energy stored in the climate system since 1970 has accumulated in the oceans

Heat of vaporization of water from melting to critical temperature

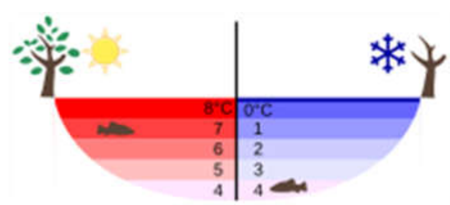
- The specific enthalpy of fusion (more commonly known as latent heat) of water is 333.55 kJ/kg at 0 °C: the same amount of energy is required to melt ice as to warm ice from −160 °C up to its melting point or to heat the same amount of water by about 80 °C. Of common substances, only that of ammonia is higher. This property confers resistance to melting on the ice of glaciers and drift ice. Before and since the advent of mechanical refrigeration, ice was and still is in common use for retarding food spoilage.

The specific heat capacity of ice at −10 °C is 2.03 J/ (g·K) and the heat capacity of steam at 100 °C is 2.08 J/ (g·K).



Density of ice and water as a function of temperature

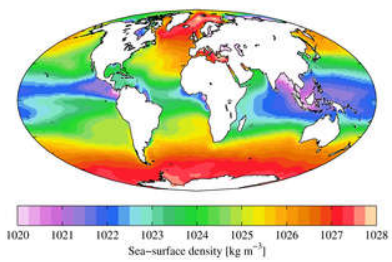
- The density of water is about **1 gram per cubic centimeter**. This relationship was originally used to define the gram. The density varies with temperature, but not linearly: as the temperature increases, the density rises to a peak at 3.98 °C (39.16 °F) and then decreases; this is unusual. Regular, hexagonal ice is also less dense than liquid water upon freezing, the density of water decreases by about 9%.
- These effects are due to the reduction of thermal motion with cooling, which allows water molecules to form more hydrogen bonds that prevent the molecules from coming close to each other. While below 4 °C the breakage of hydrogen bonds due to heating allows water molecules to pack closer despite the increase in the thermal motion above 4 °C water expands as the temperature increases. **Water near the boiling point is about 4% less dense than water at 4 °C (39 °F).**



Temperature distribution in a lake in summer and winter

The unusual density curve and lower density of ice than of water is vital to life if water were most dense at the freezing point, then in winter the very cold water at the surface of lakes and other water bodies would sink, the lake could freeze from the bottom up, and all life in them would be killed. Furthermore, given that water is a good thermal insulator (due to its heat capacity), some frozen lakes might not completely thaw in summer. The layer of ice that floats on top insulates the water below. Water at about 4 °C (39 °F) also sinks to the bottom, thus keeping the temperature of the water at the bottom constant.

Density of saltwater and ice



WOA surface density

Density:

The density, of a substance is its mass per unit volume. The symbol most often used for density is ρ, although the Latin letter D can also be used. Mathematically, density is defined as mass divided by volume: where ρ is the density, m is the mass, and V is the volume

- The density of salt water depends on the dissolved salt content as well as the temperature. **Ice still floats in the oceans**; otherwise they would freeze from the bottom up.
- However, the salt content of oceans lowers the freezing point by about 1.9 °C and lowers the temperature of the density maximum of water to the former freezing point at 0 °C. This is why, in ocean water, the downward convection of colder water is not blocked by an expansion of water as it becomes colder near the freezing point.
- The oceans' cold water near the freezing point continues to sink. So, creatures that live at the bottom of cold oceans like the Arctic Ocean generally live in water 4 °C colder than at the bottom of frozen-over fresh water lakes and rivers.

“As the surface of salt water begins to freeze (at −1.9 °C for normal salinity seawater, 3.5%) the ice that forms are essentially salt-free, with about the same density as freshwater ice. This ice floats on the surface, and the salt that is "frozen out" adds to the salinity and density of the sea water just below it, in a process known as brine rejection.”

- This denser salt water sinks by convection and the replacing seawater is subject to the same process. This produces essentially freshwater ice at −1.9 °C on the surface. The increased density of the sea water beneath the forming ice causes it to sink towards the bottom.
- On a large scale, the process of brine rejection and sinking cold salty water results in ocean currents forming to transport such water away from the Poles, leading to a global system of currents called the thermo hyaline circulation.

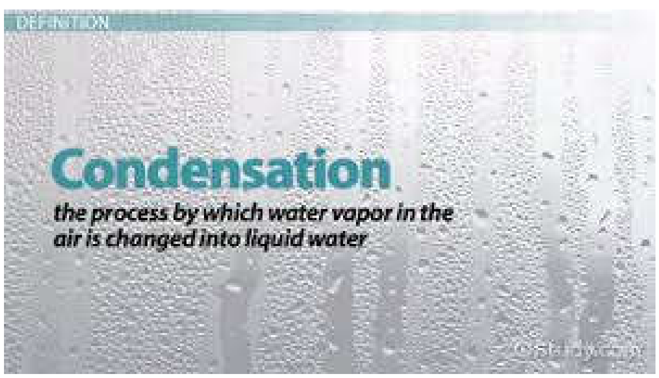
Miscibility and condensation:

1-Miscibility:

“Miscible is a fancy word for "mixable." Oil and water are not very **miscible** substances, whereas seltzer and orange juice are **miscible** and delicious! **Miscible** is a word used by chemists to explain why some liquids mix together well, while others do not.”

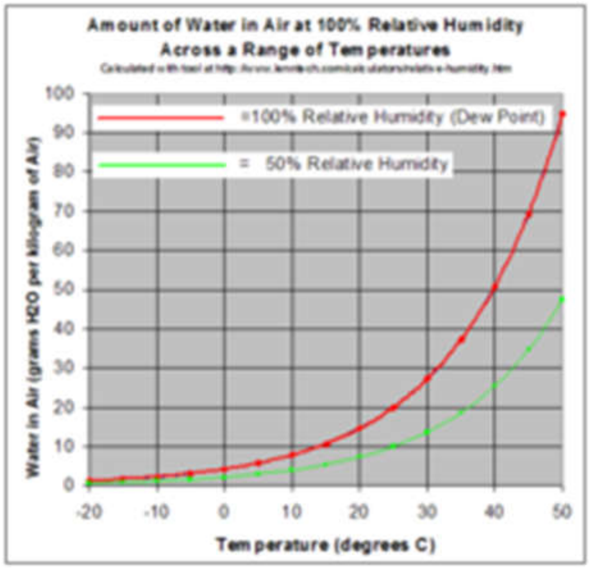
2-Condensation:

“Condensation is the process of a substance in a gaseous state transforming into a liquid state. This change is caused by a change in pressure and temperature of the substance.”



Water is miscible with many liquids, including ethanol in all proportions. Water and most oils are immiscible usually forming layers according to increasing density from the top.

- This can be predicted by comparing the polarity. Water being a relatively polar compound will tend to be miscible with liquids of high polarity such as ethanol and acetone, whereas compounds with low polarity will tend to be immiscible and poorly soluble such as with hydrocarbons.



Red line shows saturation

Vapour pressure

- As a gas, water Vapour is completely miscible with air. On the other hand, the maximum water vapor pressure that is thermodynamically stable with the liquid (or solid) at a given temperature is relatively low compared with total atmospheric pressure.
  For example:
- If the vapor's partial pressure is 2% of atmospheric pressure and the air is cooled from 25 °C, starting at about 22 °C water will start to condense, defining the dew point, and creating fog or dew. **The reverse process accounts for the fog burning off in the morning.** If the humidity is increased at room temperature, **for example**, by running a hot shower or a bath, and the temperature stays about the same, the vapor soon reaches the pressure for phase change, and then condenses out as minute water droplets, commonly referred to as steam.

A saturated gas

- Or one with **100% relative humidity** is when the vapor pressure of water in the air is at equilibrium with vapor pressure due to (liquid) water; water (or ice, if cool enough) will fail to lose mass through evaporation when exposed to saturated air. Because the amount of water vapor in air is small, relative humidity, the ratio of the partial pressure due to the water vapor to the saturated partial vapor pressure, is much more useful**. Vapor pressure above 100% relative humidity is called super-saturated and can occur if air is rapidly cooled, for example, by rising suddenly in an updraft.**

Compressibility:

“The compressibility of water is a function of pressure and temperature.”

- At 0 °C, at the limit of zero pressure, the compressibility is 5.1×10^{-10} Pa^{-1}. At the zero-pressure limit, the compressibility reaches a minimum of 4.4×10^{-10} Pa^{-1} around 45 °C before increasing again with increasing temperature. As the pressure is increased, the compressibility decreases, being 3.9×10^{-10} Pa^{-1} at 0 °C and 100 mega Pascal’s.

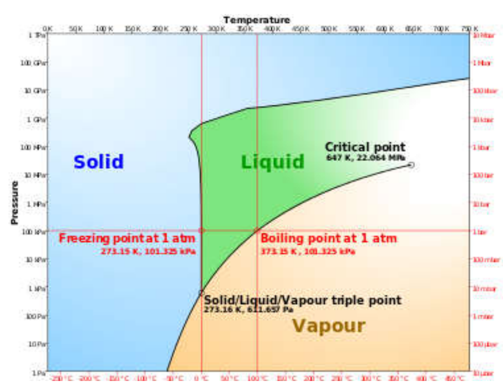
The bulk modulus of water is about 2.2 GPa. The low compressibility of non-gases, and of water in particular, leads to their often being assumed as incompressible. The low compressibility of water means that even in the deep oceans at 4 km depth, where pressures are 40 MPa, there is only a 1.8% decrease in volume.

Triple point:

“The temperature and pressure at which ordinary solid, liquid, and gaseous water coexist in equilibrium is a triple point of water.”

- Since 1954, this point had been used to define the base unit of temperature, the kelvin but, starting in 2019, the kelvin is now defined using the Boltzmann constant, rather than the triple point of water.

Due to the existence of many polymorphs (forms) of ice, water has other triple points, which have either three polymorphs of ice or two polymorphs of ice and liquid in equilibrium. Gustav Heinrich Johann Apollon Tammann in Gottingen produced data on several other triple points in the early 20th century. Kamb and others documented further triple points in the 1960s.



The Solid/ Liquid/Vapor triple point of liquid water, ice and water vapor in the lower left portion of a water phase diagram.

Melting point:

" The melting point of a substance is the temperature at which it changes state from solid to liquid. At the melting point the solid and liquid phase exist in equilibrium. The melting point of a substance depends on pressure and is usually specified at a standard pressure such as 1 atmosphere or 100."

- The melting point of ice is 0 °C (32 °F; 273 K) at standard pressure; however, pure liquid water can be super cooled well below that temperature without freezing if the liquid is not mechanically disturbed. It can remain in a fluid state down to its homogeneous nucleation point of about 231 K (−42 °C; −44 °F).
- The melting point of ordinary hexagonal ice falls slightly under moderately high pressures, by 0.0073 °C (0.0131 °F)/atm or about 0.5 °C (0.90 °F)/70 atm as the stabilization energy of hydrogen bonding is exceeded by intermolecular repulsion, but as ice transforms into its polymorphs (see crystalline states of ice) above 209.9 MPa (2,072 atm), the melting point increases markedly with pressure, i.e., reaching 355 K (82 °C) at 2.216 GPa (21,870 atm) (triple point of ice).