What Makes Water So Special!

First some definitions:

Atmospheric Pressure
The force produced by the gas molecules in the atmosphere. Atmospheric pressure is a function of the height and density of the atmosphere in conjunction with a planet's gravitational field. At sea level, Earth's atmosphere pushes with a force of 1013 millibars. On Mars, surface pressures are typically in the range of 6.8 millibars.

Boiling Point
Boiling is when the vapor pressure of the liquid equals the atmospheric pressure. At this point, the liquid can turn to a vapor. The bubbles ones sees are bubbles of water vapor. They arise from the bottom of a pot because this is usually where the heat is concentrated and where the particles have the most kinetic energy.

Equilibrium
When the rate of molecules leaving a particular phase equals the number returning. For example, water is in equilibrium with ice when it freezes at the same rate that the ice melts. It is in equilibrium with water vapor when it evaporates at the same rate that the vapor condenses.

Kinetic Energy
The internal energy of an atom or molecule often thought of as the vibrational energy of a particle. Higher kinetic energies translate into higher temperatures.

Phase Change Plateau
When melting or boiling, any heat added is absorbed by the particles changing state. The particles use the energy from the heat source to gain the extra kinetic energy required to change state and to maintain themselves in the new state. As a result, the temperature during these transitions never changes. On a graph, these transitions graph as plateaus. When condensing or freezing, particles give off heat, and there is a similar plateau.

Triple Point
At the triple point, all three phases are in equilibrium with one another - vapor sublimates to ice and condenses to liquid at the same rate that the liquid evaporates to vapor and freezes to ice at the same rate that the ice melts to liquid and sublimates to vapor.

Vapor Pressure
The inclination of a molecule to change phase and establish an equilibrium. Vapor pressure changes with temperature - the higher the temperature, the higher the vapor pressure.

And now some background:

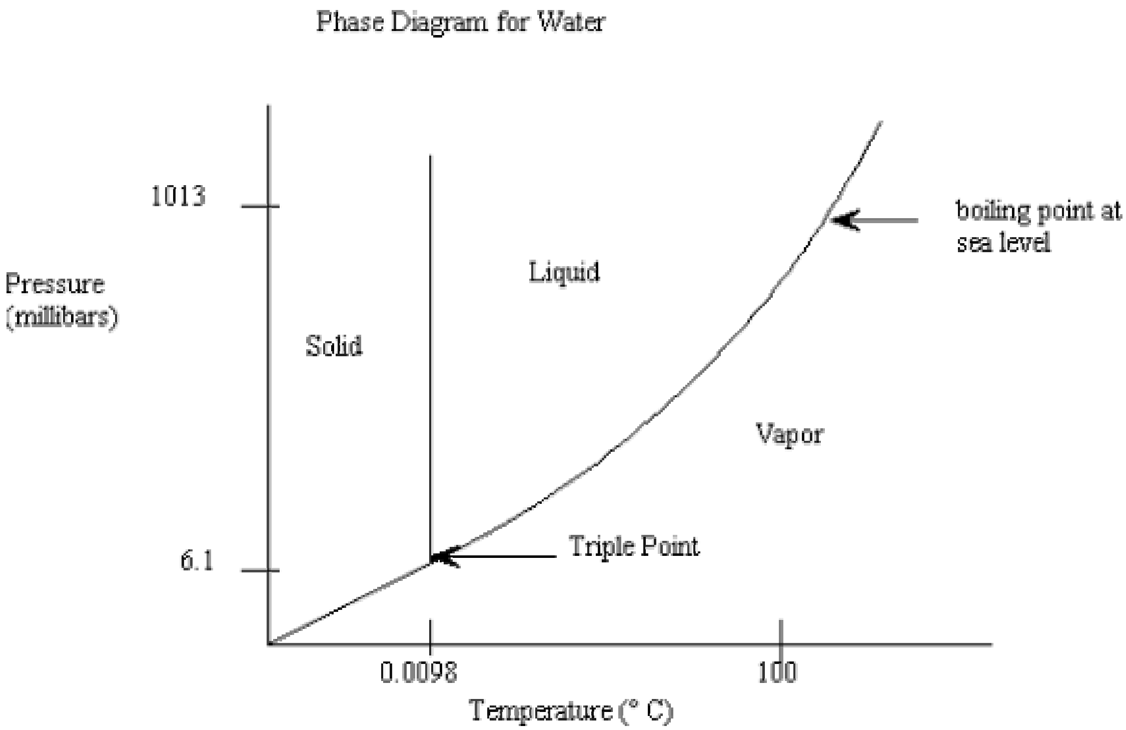
Property	Attribute due to property	Examples
Bi-polar molecule	Superior solvent of ionic compounds	• Diffuses across cell membranes to deliver nutrients and to remove wastes • Dissolves many surface materials
Forms hydrogen bonds	Chemically active molecule	• Exhibits surface tension (cohesion • Exhibits capillary action (adhesion • Present in many classes of compounds
High-density liquid	Exerts force and distributes pressure	• Provides organisms mobility and buoyancy • Erodes and transports surfaced materials
High specific heat capacity	Stores large amounts of heat	• Moderates climates on Earth • Moderates daily temperature swings • Moves equatorial heat toward the poles
Expands upon freezing	Ice is less dense than water	• Ice floats on water and protects organisms below the ice from colder temperatures • Water expands upon freezing and cracks rocks and minerals through physical weathering
Relatively low vapor given its molecular weight	Changes phase within a moderately narrow temperature range	• Exists in ail three states on Earth • Enables a water cycle that moves water through the environment • Enables cooling through evaporation
Molecule resonates at a number of frequencies	Absorbs wavelengths such as ultraviolet and infrared	• liquid water shields aquatic organisms • Water vapor shields land organisms • Acts as a greenhouse gas
Contains hydrogen and oxygen	Electrolysis can separate these elements	• Possible fuel source fox Earth-returning missions

General Background

Students can raise the boiling temperature of water by increasing the pressure above the water. To vaporize in the pressurized container, the water molecules need to increase their kinetic energy. This increase in kinetic energy translates into a higher boiling temperature.

What is the result of reducing the pressure? With a reduced pressure, the water molecules need less kinetic energy to vaporize than they do in an open container. This decrease in kinetic energy translates in a lower boiling point.

Whether water exists as a solid, liquid, or gas depends on its temperature and the pressure of the surrounding environment. Change the temperature or pressure, and water may undergo a phase change. We are familiar with how water responds to changes in temperature – at sea level, it typically freezes at 0 degrees Celsius and boils at 100 degrees Celsius. For an understanding of the effects of pressure we need to look at the following phase change diagram.



Remember that water is in equilibrium with ice when it freezes at the same rate that the ice melts and it is in equilibrium with vapor when it evaporates at the same rate that the vapor condenses.

Every substance has its own unique phase diagram. At a boundary line between two phases, two phases are in equilibrium with one another – that is, the rate of molecules leaving a phase equals the number returning. We refer to the inclination of a molecule to change phase and establish an equilibrium as its vapor pressure. Vapor pressure increases with increasing temperatures. At higher temperatures, a particle's kinetic energy is higher, and, with more energy available at higher temperatures, it is easier for particles to change phase. Even solids such as ice have a vapor pressure and can sublimate directly to the vapor phase.

If two phases are not in equilibrium, molecules will change from one phase to the other until an equilibrium is established. The water evaporating out of a lake in the desert is trying to establish an equilibrium with the dry desert air. In the case of a puddle, the water disappears completely before an equilibrium is established. Each temperature-pressure combination has its own equilibrium point. If you connect these many equilibrium points, you will have drawn the boundary lines on the phase diagram.

All three boundary lines meet at a point called the triple point. At this temperature and pressure, all three phases are in equilibrium with one another. In other words, at the triple point, vapor sublimates to ice and condenses to liquid, liquid evaporates to vapor and freezes to ice, and ice melts to liquid and sublimates to vapor all the same rate. A minuscule change in either temperature or pressure will move the phase changes away from the triple point. Away from the boundary lines, water exists in a single phase over a particular range of temperatures and pressures.

Because a phase diagram shows so clearly how water changes phase in relation to pressure and temperature levels, it is the key to helping students understand the current situation with liquid water on Mars.

On the phase diagram, notice that below 6.1 millibars, liquid water cannot exist, irrespective of the temperature. Water's vapor pressure is just too high to remain a liquid below this level. When atmospheric pressure falls below 6. 1 millibars, water can only exist as ice or vapor, depending on the temperature. This fact is significant in a study of Mars because the atmospheric pressure at the Martian surface hovers just above 6.1 millibars. Any water that might form on a warm afternoon from melting ice would quickly disappear in the desiccated Martian atmosphere. If the vapor pressure of the warmed water exceeded atmospheric pressure, it would boil. If, instead, its vapor pressure stayed below atmospheric pressure, the water would evaporate. The temperature and pressure combinations on Mars make liquid water theoretically possible on an occasional basis. However, the desiccated atmosphere and the short-term nature of the appropriate temperature-pressure conditions make the existence of significant amounts of water on the Martian surface impossible.

On Earth, propane, butane, dry ice, ether, and freon are familiar materials whose vapor pressure is considerably higher than Earth's atmospheric pressure. Toy stores often see freon-filled "perpetual drinking birds" and globes containing freon that can boil when held in one's hand. On Mars, water would behave the way freon and butane behave on Earth.

Something to be alert to when discussing weather-related pressure changes is the fact that humid air weighs less than dry air. This is because water vapor weighs less than gaseous nitrogen (N_2). Because liquid water weighs more than air, students invariably say that humid air weighs more than dry air. This response reveals that they make no distinction between liquid and gaseous water, even though we teach about states of matter. So, there are actually two reasons why low pressures are associated with rainy weather - a humid air mass weighs less than a dry air mass, and low pressure system rises to a lower altitude than a high pressure system (that is, it is not as tall).

Whenever the pressure changes, there is a corresponding change in water's boiling temperature. People living at high altitudes use pressure cookers, cook foods longer, or modify their recipes to compensate for the lower boiling temperatures. In fact, Galileo used differences in boiling temperatures to calculate elevation.

Other aspects of water:

Viscosity:

Viscosity deals with the resistance to internal friction between molecules. Some liquids like water have a low viscosity where other liquids like honey or shampoo have a high viscosity. Viscosity will be affected by the temperature. At higher temperatures the viscosity decreases as the molecules take on more kinetic energy allowing them to move past each other faster.

Sample Viscosity Lab:

Objective: to determine the viscosity of Karo syrup

Theory: An object falling through a viscous medium will reach a terminal velocity (constant velocity, no acceleration) when the sum of the buoyant force and the viscous force equals the force of gravity. For a sphere of radius (r) in a fluid of density (ρ),

$$F_{buoyancy} + F_{viscosity} = F_{gravity}$$

$F_b = 4/3\,\pi\, r^3 \rho\, g$ $\quad$ $F_v = 6\,\pi\,\eta\, r\, v_t$ $\quad$ $F_g = m\, g$

where v_t is the terminal velocity and η **is the viscosity**. g is given as 9.807 m/s^2

Procedure:

1. Determine the density of the Karo syrup. Record this and the temperature of the syrup.

2. Measure diameter of ball bearings (two sizes). Record radius of each type bearing.

3. Find average mass of each type ball bearing.

4. Using rulers, stop watches, graduated cylinders determine the average terminal velocity of 10 of the small ball bearings. Velocity is defined as distance traveled divided by time taken. Then repeat for the larger ball bearings. Record all data.

5. Using the formula: $F_{buoyancy} + F_{viscosity} = F_{gravity}$ solve for the viscosity of the syrup first using your terminal velocity average for the small ball bearings and then for the larger bearings.

final formula registration: $4/3\,\pi\, r^3 \rho\, g + 6\,\pi\,\eta\, r\, v_t = m\, g$ (and you must solve for η)

Surface Tension

This is a condition existing at the free surface of a liquid, resembling the properties of an elastic skin under tension. The tension is the result of intermolecular forces exerting an unbalanced inward pull on the individual surface molecules; this is reflected in the considerable curvature at those edges where the liquid is in contact with the wall of a vessel. More specifically, the tension is the force per unit length of any straight line on the liquid surface that the surface layers on the opposite sides of the line exert upon each other.

The tendency of any liquid surface is to become as small as possible as a result of this tension, as in the case of mercury, which forms an almost round ball when a small quantity is placed on a horizontal surface. The near-perfect spherical shape of a soap bubble, which is the result of the distribution of tension on the thin film of soap, is another example of this force; surface tension alone can support a needle placed horizontally on a water surface. Surface tension is important at zero gravity, as in space flight. Liquids cannot be stored in open containers because they run up the vessel walls.

One way to think of surface tension is in terms of energy. The larger the surface, the more energy there is. To minimize energy, most fluids assume the shape with the smallest surface area. This is why small drops of water are round, for instance – a sphere is the shape with the minimum surface area for a given volume. Soap bubbles also tend to form themselves into shapes with minimal surface area.

It takes work to increase the surface area of a liquid. The surface tension can be defined in terms of this work (W), as follows:

surface tension = $\gamma = W / \Delta A$, where ΔA is the change in surface area.

If you have a thin film of fluid, and try to stretch it, the film resists. The surface tension can also be defined as the force (F) per unit length (L) tending to pull the surface back:

surface tension = $\gamma = F / L$

Water is often used for cleaning, but the surface tension makes it hard for water to penetrate into small crevices or openings, such as are found in clothes. Soap is added to water to reduce the surface tension, so clothes (or whatever else) get much cleaner.

Surface Tension of Liquids

Surface tension is the force at the surface of a liquid due to adhesive forces of the liquid molecules for the walls of the container and the attractive forces of the molecules of liquid for each other. When the adhesive forces of the molecules for the walls of the container are greater than the attractive forces between the liquid molecules, then the surface of a liquid confined to a narrow diameter container will curve downward forming a concave surface called a meniscus. Most important examples are water solutions. The water adheres to the surface of the container greater than the water molecules are attracted to each other. We do not see this downward curvature when the surface area is great, but if the liquid is confined to a small diameter tube such as a graduated cylinder, pipette, burette, or volumetric flask then the surface tension is great enough to noticeably distort the surface. In such cases when we are trying to read the liquid surface level such as measuring a liquid in a graduated cylinder, then one should make the reading at eye level and the lowest curvature of the meniscus should be read.

When the adhesive forces against the walls of the container are less than the intermolecular forces, then the surface of a confined liquid will bulge upward slightly forming a convened surface. Again, such a surface should be read at eye level and the topmost part of the surface should be read. Surface tension helps to explain why the feathers of a duck can help the duck float on water.

Consequences of Surface Tension:

1) capillary action – this is related to the adhesive properties of water. You can see capillary action 'in action' by placing a straw into a glass of water. The water climbs inside the straw. What is happening is that the water molecules are attracted to the straw molecules. When one water molecule moves closer to a straw molecule the other water molecules (which are cohesively attracted to other water molecules) also move up into the straw. Capillary action is limited by gravity and the size of the straw. The thinner the straw or tube the higher the capillary action will pull the water. Plants take advantage of capillary action to pull water from the ground into themselves. From the roots, water drawn through the plant is done by transpiration.

height of fluid in capillary tube: $h = (2\gamma / \rho g r) \cos\theta$ where γ = surface tension, ρ = density, g = gravitational acceleration, r = radius of tube and θ is angle formed with meniscus.

2) cohesion - attractive force between like molecules (as in water to water attraction)

3) adhesion - attractive force between unlike molecules (as in water to container walls)