

Treating the Public Water Supply: What Is In Your Water, and How Is It Made Safe to Drink?

Inorganic Reactions Experiment

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Key Concepts:

- Aqueous Solutions (Substances Dissolved in Water)
 - Solubility
 - The Solvation Process for Ionic Crystals (yellow box) (Note: This section contains an animation.)
 - Suspensions
 - Physical Separation Techniques for Water Treatment
 - Precipitation Reactions
 - Ion-Exchange Reactions
-

Human Use of Freshwater

Humans' Need for Clean Freshwater

Water is perhaps the most important nutrient in our diets. In fact, a human adult needs to drink approximately 2 liters (8 glasses) of water every day to replenish the water that is lost from the body through the skin, respiratory tract, and urine. But some water sources cannot safely be used to meet our requirement for drinking water. In fact, 99.7% of the Earth's

water supply is not usable by humans. This unusable water includes saltwater, ice, and water vapor in the atmosphere. Only freshwater, which is contained in rivers, lakes, and underground sources, can be used for human consumption. Furthermore, many freshwater sources are not suitable for humans to drink. Many serious diseases, such as cholera, are caused by drinking water that contains parasitic microorganisms. Water containing large amounts of industrial waste or agricultural chemicals (*e.g.*, pesticides) can also be toxic and unfit for drinking. Hence, humans have a great need for a reliable source of clean freshwater for drinking.

In addition to the water needed for drinking, humans use much larger amounts of clean freshwater in other applications. These other uses of freshwater include household use (*e.g.*, cooking and cleaning), industry, agriculture (*e.g.*, irrigation), and recreation. Hence, the quality of the freshwater supply is important for virtually every aspect of our lives.

In response to this need for reliable supplies of clean and usable freshwater, governments at all levels have formed organizations and passed legislation to monitor, treat, and protect our water supplies. The Clean Water Act, passed in 1972 and revised in 1978, established the policy of contracting water-treatment facilities and implemented enforcement procedures to ensure that our water supplies would be kept clean for human use. Later, the Safe Drinking Water Act of 1974 (amended in 1986 and 1996) established minimum federal public-health standards for the water supply. The policies and standards provided by these acts, together with a large body of legislation developed at the state and local levels, have helped to ensure that residents of almost every part of the United States have access to a safe, usable supply of water for drinking and other uses.

Sources of Freshwater

Where does the water that we use come from? As stated above, saltwater is not suitable for human use. (The high salt concentration makes it unsuitable for most human applications, and removal of the salt is costly.) Water found in the forms of ice and vapor are not readily accessible for human use. This leaves only about 0.3% of the water on the Earth as liquid freshwater. The vast majority (approximately 98.5%) of the Earth's freshwater is contained in underground supplies known as **groundwater**. The remaining freshwater sources, including rivers and lakes, are collectively known as **surface water**. The amount of water that we use from groundwater and surface water varies from region to region. Generally, groundwater is cleaner and requires less treatment, but surface water is easier to obtain.

Water is Not Really Pure H₂O

Why is it necessary to treat freshwater before we drink it or use it for other applications? The answer is that our water supplies are not pure, *i.e.*, these supplies contain other species that may make the water unsuitable for human use. H₂O is certainly the largest and most

important component of any water source, but this molecule is hardly the only chemical present in the water supply. It may be surprising to learn that water treatment does not eliminate all of the impurities from water. In fact, the treatment process itself contributes additional impurities to the water. However, the impurities added during the treatment process generally help to make water more suitable for human use.

Species (Other than H₂O) Contained in Water

Chemical analysis of virtually any freshwater sample reveals that "water"-- even water that has been rigorously cleaned and treated-- is really a **solution** containing many **dissolved species**. A solution is a **homogenous** system (a system that is uniform throughout) containing more than one substance. A solution in which H₂O is the solvent is known as an **aqueous solution**. In addition to H₂O (which is the solvent), water samples may include:

- ions (*e.g.*, Na⁺, Ca²⁺, F⁻, and HSO₄⁻)
- dissolved gases (*e.g.*, O₂ and CO₂)
- other natural dissolved molecules (*e.g.*, organic by-products of decaying leaves)
- dissolved molecules from human activity (*e.g.*, industrial and agricultural wastes)

How do the dissolved species listed above get into the water supply? Many of these species originate as solids that are dissolved in water to yield the aqueous species in solution. In order to understand how a solid dissolves in the water supply, we must consider the microscopic events that occur when a solid dissolves. The process by which these materials dissolve and enter the water supply is discussed in the yellow box, below.

The Solvation Process

Consider an ionic solid, such as NaCl. Recall from the Periodic-Properties Experiment that an ionic solid is an extended three-dimensional lattice (a regular geometric arrangement) of cations and anions. NaCl (Figure 2) has the shape of a cubic crystal consisting of repeating units of Na⁺ and Cl⁻ ions. One of the repeating units (the "unit cell") is shown in Figure 1.

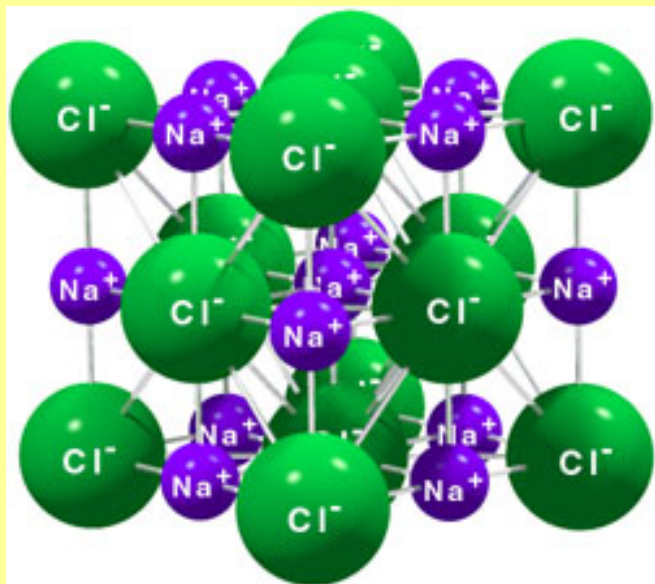
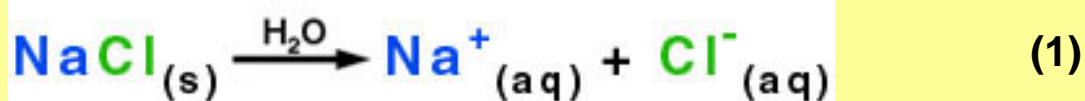


Figure 1

This is the unit cell for a sodium chloride (NaCl) crystal lattice, in which Na^+ and Cl^- ions are arranged in a regular cubic pattern. This pattern can be repeated indefinitely to make a NaCl crystal.

Note: The crystal structure was drawn using PowderCell for Windows, and the image was rendered using POV-Ray (see [References](#)).

In order for NaCl to be soluble, the Na^+ and Cl^- ions must break free from the crystal-lattice structure of the solid. When the ions are in solution, they are surrounded by water molecules, and the ions are said to be **solvated**, or dissolved in an **aqueous** solution, denoted (aq). Hence, the process of dissolving a NaCl crystal can be described by the following chemical equation (Equation 1):



What are the microscopic events that occur when the crystal dissolves in water? The ions become solvated because of (1) the favorable electrostatic interactions with the solvent molecules (as you will learn about in Chem 111) and (2) entropy (randomness,

as you will learn about in Chem 112). If the sum of these effects are stronger in the solution than in the crystal, the salt will dissolve. In this tutorial, we will describe only the effect of the electrostatic interactions on solvation.

H₂O is a polar molecule; i.e., the molecule has negatively and positively charged regions. These charged regions are attracted to ions with the opposite charge. Hence, the positively charged regions of water molecules are attracted to Cl⁻ ions, and the negatively charged regions of water molecules are attracted to Na⁺ ions. When several water molecules surround an ion in the crystal, the sum of the attractive forces between the water molecules and the ion may become strong enough to overcome the attractive forces between the cations and anions in the crystal. (Recall from the "[Bands, Bonds, and Doping: How Do LED's Work?](#)" tutorial that many weak interactions can result in strong bonding.) The water molecules form a **shell of solvation** around the ion, and the water-surrounded ion can break away from the crystal, as shown in Figure 2.

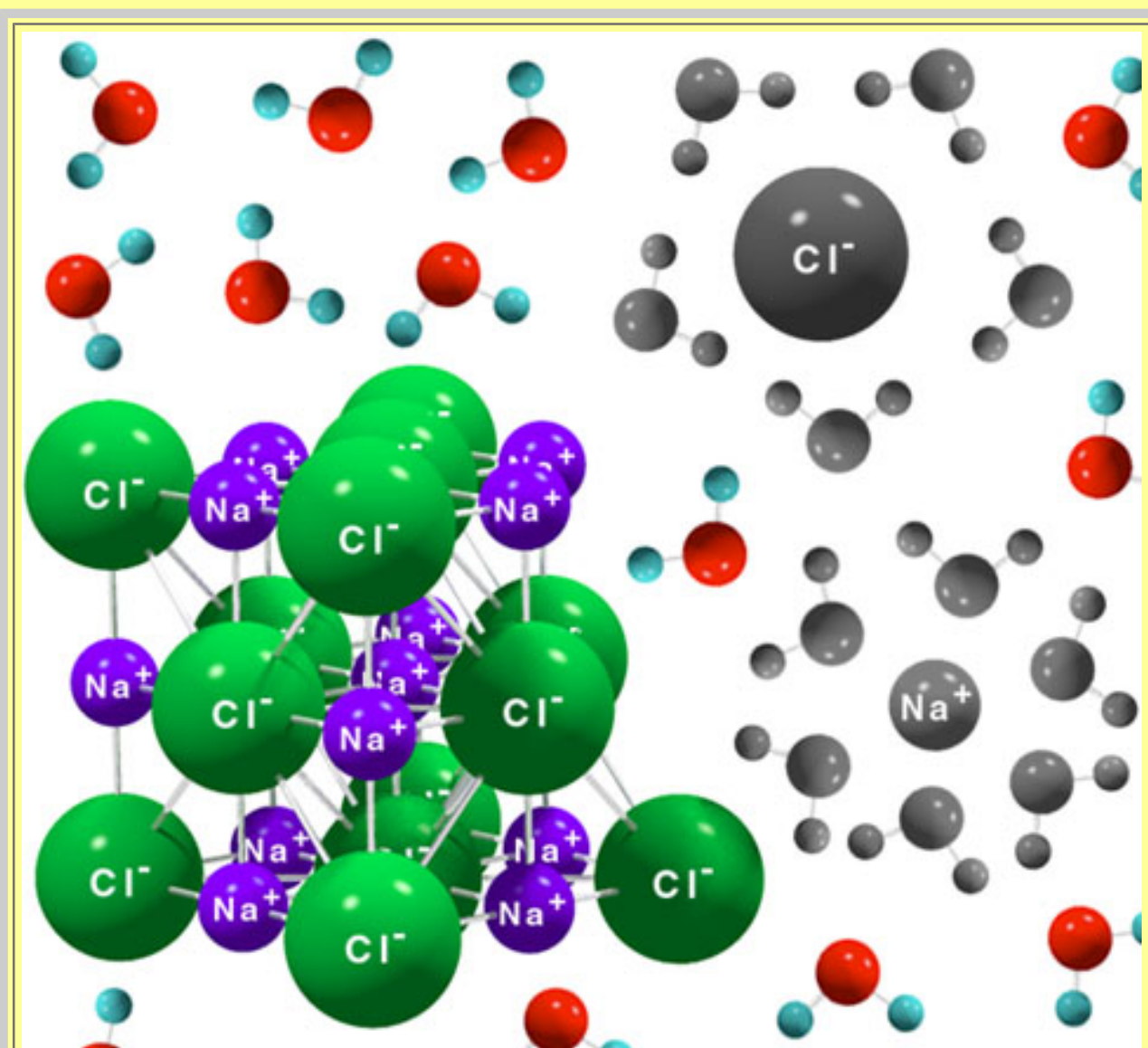


Figure 2

This illustration shows a NaCl crystal being dissolved by H₂O. Two ions have become solvated by water molecules and have broken away from the crystal. The water molecules form a "shell of solvation" around the ions. The ions that have broken away from the crystal and the water molecules forming the shells of solvation are shown in black and white in this illustration.

Note: Chloride ions are shown in green, sodium ions are shown in dark blue, oxygen atoms are shown in red, and hydrogen atoms are shown in light blue.

Eventually, all of the ions in the crystal may become solvated and break away from the crystal, forming a solution of solvated (aqueous) ions. The ions in the crystal become "free agents" in the solvent-- unattached to other ions and free to migrate about randomly in solution. In the solution, H₂O molecules are constantly moving and exchanging with one another. Three steps in the formation of an ionic solution from a crystal are shown in Figure 3. The movie link in Figure 3 allows you to watch this process.

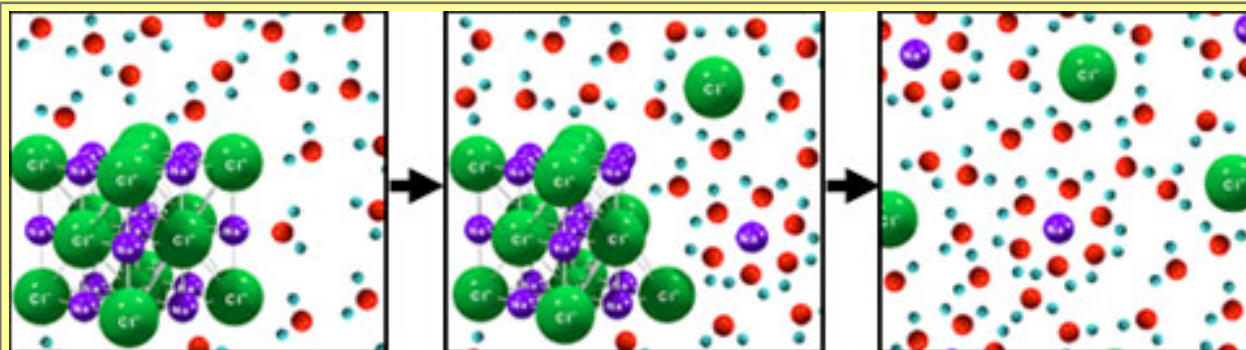


Figure 3

This scheme illustrates three steps in the formation of an ionic solution from a crystal. (1) The ionic crystal is immersed in a solvent (*e.g.*, water). (2) Ions become solvated and break away from the crystal. (3) The crystal is completely broken apart as all of the ions become solvated, and a solution of solvated (aqueous) ions results.

Overall, the charges in an ionic solution balance, so the solution is electrically neutral.

Nonionic (molecular) solids (*e.g.*, sugar) can also dissolve in water by a similar mechanism. When molecular solids dissolve, the individual molecules remain intact but become solvated by H₂O molecules (and thus break away from the solid). Gas and liquid molecules can also be solvated by H₂O molecules to form solutions.

In addition to the species that are dissolved in water, a number of particles may be found to be **suspended** in the water. A **suspension** is a system in which the dispersed particles (typically solids, but may also be liquid droplets) are larger than ordinary molecules and ions, but are still small enough to remain dispersed (suspended) in the solvent. Because the suspended particles remain in the solid state (or as droplets), the particles are not said to be dissolved, and the system is **not homogenous**. The differentiation between a suspension and a solution is based on particle size (the suspended particles typically have diameters in the range of 1-100 nm); however, the distinction between a suspension and a solution on the basis of particle size is not well-defined. Experimentally, the difference between the two types of mixtures is seen using the Tyndall effect, in which light is scattered by the suspended particles in a suspension (*i.e.*, the solution appears cloudy), but is not scattered in a homogenous solution. Examples of suspended particles in water samples may include:

- bacteria
- floating debris (*e.g.*, twigs, bits of leaves, and trash)
- sand and dirt

The types and amounts of these species in a given water sample depend on the source of the sample, and what sort of treatment the water has undergone. Some of these dissolved species, such as O₂ gas, are unavoidable and do not detract from our ability to use the water. Other species, such as F⁻, are desirable and are often added during the water-treatment process. Many of the species dissolved in water, however, are harmful for humans to drink, or otherwise render the water unsuitable for human use (*e.g.*, Ca²⁺ is responsible for "hard water", which can leave insoluble scum on appliances and industrial boilers). These products must be removed via water-treatment facilities.

Questions on Species (Other than H₂O) Contained in Water

- A small amount (1.0 g) of copper sulfate (CuSO₄), a sky-blue, ionic crystalline solid, is added to 1.0 L of water in a volumetric flask, and vigorously shaken until all of the solid is dissolved.

- a. Briefly, describe the major microscopic events that occur in this solvation process.
 - b. What species (other than H_2O molecules) are present in the resulting solution?
- Tell whether the following mixtures are solutions or suspensions:
 - a. a clear-looking mixture formed by mixing two ionic solutions
 - b. a cloudy-looking mixture formed by mixing two different ionic solutions
 - c. muddy water
 - d. hydrated Mg^{2+} and Cl^- ions in H_2O
-

Treatment of the Public Water Supply

What, then, must be removed from public water supplies, and what other chemicals are added to the water? How do public water facilities treat our water to make it safe for us to drink and appropriate for other human uses? There are six major steps in the treatment of our water: screening, sedimentation, precipitation, filtration, adsorption, and disinfection. Some of these steps, such as precipitation, involve chemical reactions among the aqueous species dissolved in the water; others, such as screening, involve only separation of particles on the basis of physical characteristics like size. Many of these steps depend on one another. For instance, precipitation generates solids in the water from particles that had been dissolved; these solids must then be removed through sedimentation or filtration. We shall discuss each of the six steps in water treatment below, and then present a schematic showing how the steps work together to produce clean, usable freshwater.

Screening

Surface water (water from lakes and rivers) often has large debris, such as sticks, leaves, fish, and trash, floating in it. These objects can clog the water-treatment system and must be removed before the water enters the treatment plant. Treatment facilities that use surface water have large screens (Figure 4) covering the site of water intake. The debris is too large to pass through the holes in the screens. Thus, as the water enters the plant, the large debris is removed. The screens must be cleaned periodically to remove any objects that have become stuck, so that they do not clog the screen and impede water flow into the plant. (Another problem for water-treatment plants in the Great Lakes and Mississippi Valley is zebra mussels, which can "congregate" at the water intake and clog the screen, eventually sealing it off. Potassium permanganate, KMnO_4 , can be used to kill these mollusks.)

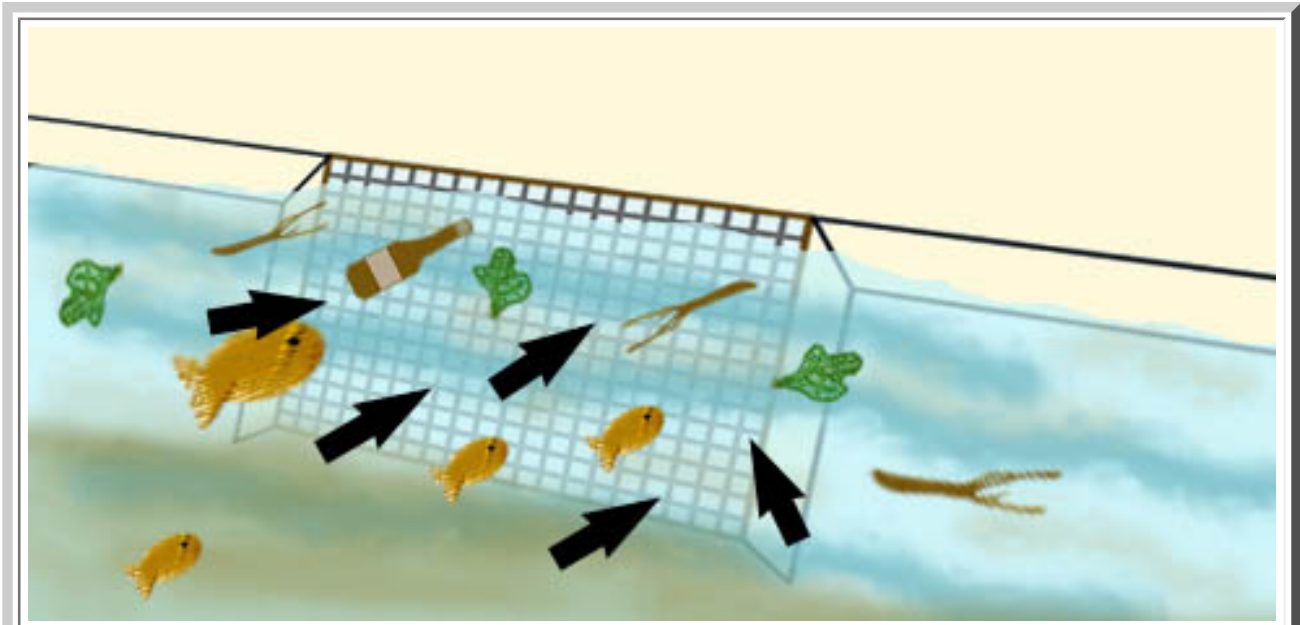


Figure 4

This drawing shows some of the large objects in surface water that are removed as the water passes through a screen into the water-treatment facility. The large black arrows show the direction of water flow through the screen.

Sedimentation

Other suspended (insoluble) particles, such as sand and dirt, are small enough to pass easily through the screens. These particles must be removed from the water by another process known as **sedimentation** (Figure 5). When water is allowed to sit, heavy suspended particles (*e.g.*, sand) will settle to the bottom over time because they are denser than water. The water, now free of the suspended impurities, can be collected from the top without disturbing the layer of sediment at the bottom (which is eventually discarded).

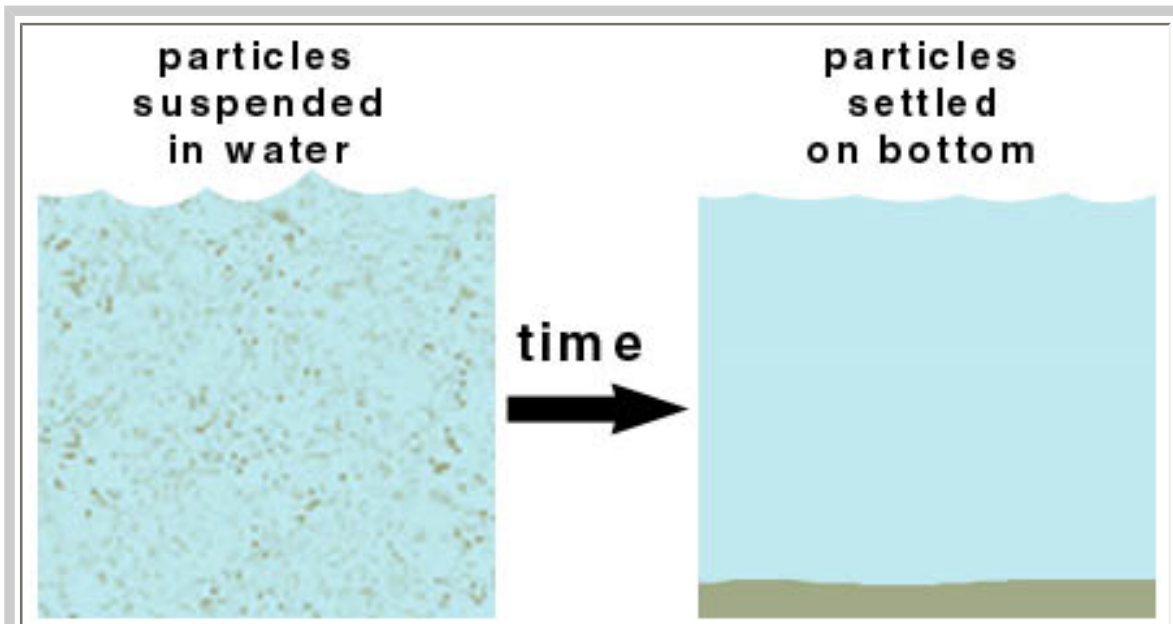


Figure 5

Particles that are insoluble in water may be suspended in the water, particularly if the water is turbulent (stirred up). If the particles are heavy enough, they will settle to the bottom when the water is allowed to sit still over time.

Sometimes the insoluble particles are too small to settle out quickly enough to use sedimentation alone. Two processes, known as **flocculation** and **coagulation**, are used to create larger particles that will settle quickly to the bottom. In flocculation, small particles with non-rigid surfaces are made to agglomerate by mixing the water (and thus bringing the particles into contact with one another so that the surfaces can become stuck together). When the agglomeration of the particles gets large enough, the aggregate can settle in still water by sedimentation. Other suspended particles do not agglomerate well by flocculation. To remove these particles from the water, coagulation must be used. Coagulation is the process of gathering particles into a cluster or clot, often achieved by the addition of special chemicals known as **coagulants**. The most common coagulant used in water-treatment facilities is **aluminum sulfate** (alum, $\text{Al}_2(\text{SO}_4)_3$). Other Al and Fe salts, including poly-aluminum chloride, ferric chloride, and ferric sulfate, may be used as well. These salts react with ions naturally found in the water to produce a solid precipitate (Equation 2). As this precipitate forms, other particles are caught in the solid, forming a mass that will settle to the bottom via sedimentation (Figure 6).



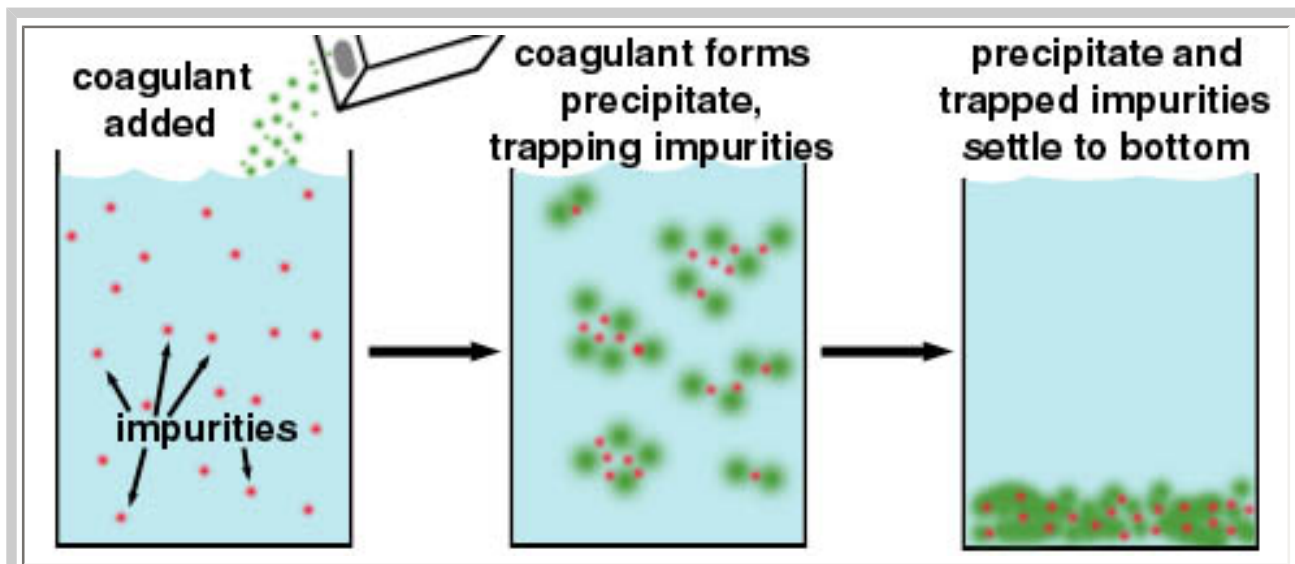


Figure 6

When coagulants such as $\text{Al}_2(\text{SO}_4)_3$ are added to the water supply, they form solid precipitates (green), as shown in Equation 2, above. These precipitates catch other impurities (red) in the water, forming a solid mass containing the precipitate formed by coagulation and the trapped impurities. This mass will settle to the bottom by sedimentation, and the water (with the trapped impurities now removed) can be drained off of the top.

Precipitation

The steps in the water-treatment process described above are used to remove insoluble particles from the water supply. But recall from above (["Species \(Other Than \$\text{H}_2\text{O}\$ \) Contained in Water"](#)) that water also contains many molecules and ions in solution. Many of the ions in solution can be removed by **precipitation**: reacting the ions (to be removed) with other ions to produce insoluble solids that can be removed by sedimentation (see above) or filtration (see below).

How does the precipitation step of the water-treatment process differ from the coagulation reaction described above?

ANSWER: In the precipitation step, the particles to be removed are part of the chemical reaction forming the precipitate. In the coagulation step, the particles to be removed are not part of the reaction forming the precipitate; they are simply trapped in the precipitate that is formed from added chemicals.

A typical precipitation reaction used to remove ions in water treatment follows the reaction shown in Equation 3, below. This is the same reaction type that you performed in the Experiment when the reaction between ions from two aqueous solutions produced a solid precipitate.



Two major classes of ions are typically removed via precipitation:

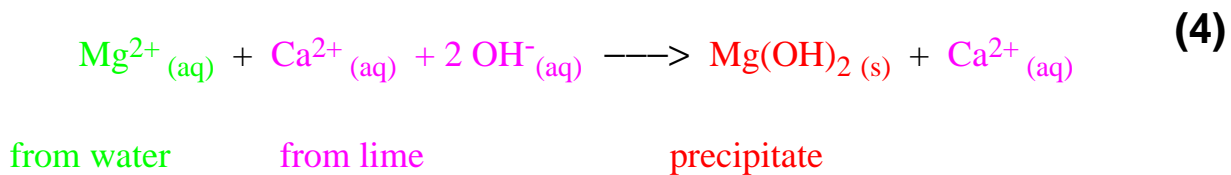
- Calcium (Ca^{2+}) ions and magnesium (Mg^{2+}) ions that have been leached from minerals in the ground cause the condition known as "**water hardness**". These ions do not pose any health threat, but they can engage in reactions that leave insoluble mineral deposits, such as scum rings on bathtubs and cooking vessels, or scale on industrial boilers, which decreases the boilers' efficiency. These deposits can make hard water unsuitable for many uses.
- Iron (Fe^{2+}) ions and manganese (Mn^{2+}) ions can stain plumbing fixtures and laundered clothes. These ions may also promote the growth of certain bacteria, which give foul tastes and odors to the water.

Treating Water Hardness

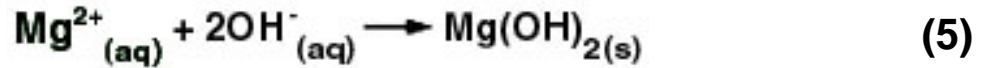
The process of removing Ca^{2+} and Mg^{2+} from the water is known as **water softening**. Two minerals, **lime** ($\text{Ca}(\text{OH})_2$) and **soda ash** (Na_2CO_3), are typically used to soften public water supplies. (Incidentally, one important source of lime is near St. Genevieve, Missouri.)

When lime is added to water, it dissolves to give three aqueous (solvated) ions: one Ca^{2+} ion and two OH^- ions for each unit of $\text{Ca}(\text{OH})_2$. Likewise, soda ash dissolves to give two Na^+ ions and one CO_3^{2-} ion for each unit of Na_2CO_3 that dissolves.

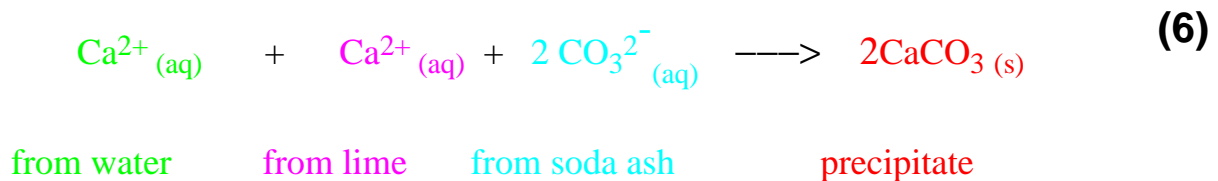
A number of reactions occur to generate the insoluble precipitates $\text{CaCO}_{3(s)}$ and $\text{Mg}(\text{OH})_{2(s)}$ from the Ca^{2+} and Mg^{2+} ions. The most important reaction for the removal of Mg^{2+} is shown in Equation 4.



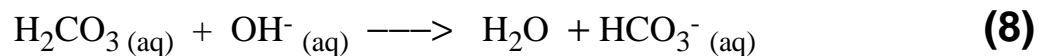
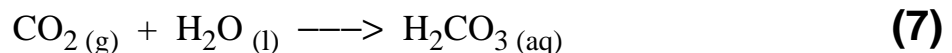
Notice that Ca^{2+} appears on both sides of Equation 4. The calcium ion from lime does not actually participate in the reaction to generate insoluble $\text{Mg}(\text{OH})_2$. Hence, this ion is called a **spectator ion** and can be omitted from the equation. We can write the reaction more correctly with the **net ionic equation**, given by Equation 5.



The important reaction for the removal of Ca^{2+} ions is given in Equation 6.



The solids generated by the water-softening precipitation reaction are then removed by sedimentation or filtration. If an excess of lime was used to precipitate magnesium ions in the water (Equation 4), some unused hydroxide (OH^{-}) ions will remain in the water after the calcium is precipitated, resulting in a high (or basic) pH. If necessary, the pH can be lowered by bubbling carbon dioxide gas through the water. The net ionic equations for this recarbonation are given in Equations 7 and 8.



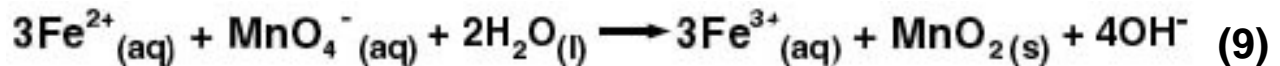
Bicarbonate (HCO_3^{-}) remaining in the water is nontoxic and does not negatively affect the flavor of the water.

Removing Iron and Manganese

Two types of precipitation reactions may be used to remove Fe^{2+} and Mn^{2+} from water.

1. The most important of these reactions is **oxidation**. Using molecular oxygen (O_2) or another oxidant such as potassium permanganate (KMnO_4), $\text{Fe}(\text{II})$ is readily

oxidized to Fe(III) in solution (Equation 9). If the solution is alkaline (high pH, basic), the Fe(III) forms Fe(OH)₃. As the concentration of Fe(OH)₃ increases, the oxygens start to coordinate between multiple iron ions, and a lattice begins to form. (Recall the definition of a lattice from the discussion above on "[The Solvation Process](#).") At some point in this lattice formation, the Fe(OH)₃ starts to look like Fe₂O₃ (rust) and precipitates. Hence, by adding an oxidant to the water and raising the water's pH at the water-treatment plant, an insoluble precipitate is formed. The insoluble rust can then be removed by sedimentation or filtration.



2. The water-softening agents described in the "Treating Water Hardness" section above can also help to make insoluble precipitates from Fe²⁺ and Mn²⁺.

Questions on Precipitation

- Explain what happens to the manganese that is added to the water for iron removal.
- Equation 4 shows how Mg²⁺ is removed using lime. However, this process introduces the new water-hardness problem of Ca²⁺ in the water. Propose a way to eliminate this problem.

Filtration

Often, the particles generated by the precipitation reactions described above are too small to settle efficiently by sedimentation. One strategy that is frequently employed to remove these solids is **gravity filtration** (Figure 7). In this process, water containing solid impurities (*e.g.*, precipitates from water softening) is passed through a porous medium, typically layers of sand and gravel. The force of gravity is used to push the water through the medium. The small water molecules pass through the holes between sand and gravel pieces. However, the solids (from precipitation) get stuck in the holes, and are thus retained in the porous medium. The water that passes through the bottom of the filter no longer contains those solid impurities.

Gravity filters at water-treatment plants have a pipe feeding into the under drain, the bottom

layer where the clean water is collected. By adding water to the filter through this pipe, clean water can be forced upward through the filter to remove the solids that have collected in the filter. This process is used to clean the filter.

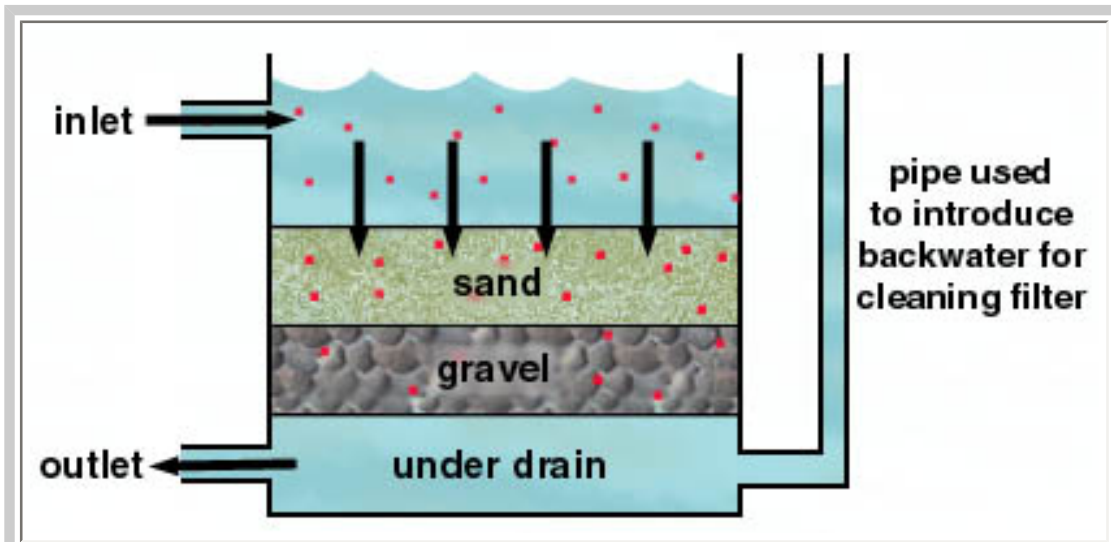


Figure 7

Water containing solid impurities (red) enters the filter through an inlet at the top and is forced by gravity through layers of sand and gravel. The solids get trapped between the sand and gravel pieces. The water that emerges into the under drain at the bottom of the filter is cleaned of these solids and exits the filter through an outlet at the bottom.

Adsorption

Dissolved organic compounds in water (*e.g.*, atrazine, an herbicide, and industrial waste products) can pose a significant health threat, and may affect the taste and odor of drinking water. To remove them, the process of **adsorption** is used. Adsorption is a process in which one substance is attached to the surface of another substance (Figure 8). **Powdered activated carbon (PAC)**, a finely ground charcoal, is used for this process. When PAC is added to the water, the organic compounds attach to the surface of the powder granules. The granules of PAC have irregularly shaped surfaces, which gives PAC a very large surface area to attract organic compounds. It is estimated that 1 pound of PAC has a surface area of 100 acres! The carbon can then be removed by filtration, taking the unwanted organic compounds with it.

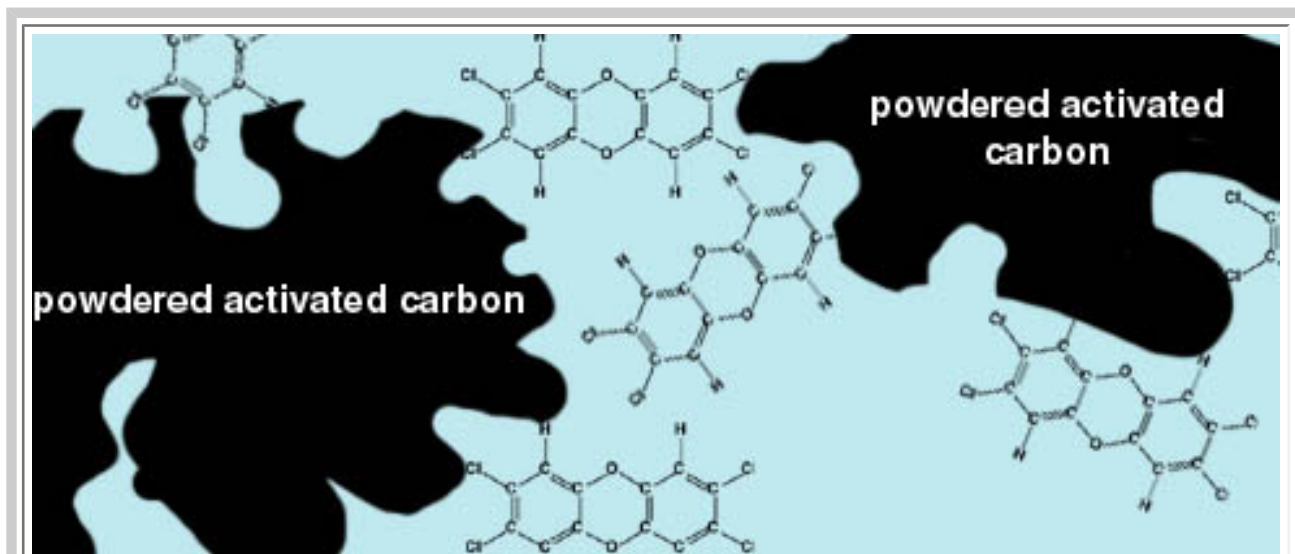


Figure 8

This figure shows the irregular surface area of powdered activated carbon (PAC). Organic molecules, such as the dioxin (a toxic byproduct of chemical bleaching and a precursor to powerful pesticides) shown here, adhere to the surface of the PAC granules. When the PAC is removed by filtration, the adsorbed organic molecules are removed with the carbon.

Disinfection

In many water supplies, the most serious health threats are posed not by chemicals, but by infectious organisms (bacteria) in the water. Chlorine (Cl_2) is a major disinfectant that is cheap and kills most of the serious disease-causing bacteria in the water. However, chlorine disinfection results in a wide variety of by-products. One class of chlorination by-products, known as trihalomethanes (THM's), are suspected carcinogens. Because of concern about these by-products in the water supply, chlorine is now kept to minimum levels, and other methods of disinfection are being used more frequently. Chloramines form more stable disinfectants and pose less risk of harmful by-products, but cost more to use. Other methods focus on removing the organisms through coagulation, sedimentation, and improved filtration.

Addition of Other Chemicals to the Water Supply

Certainly a principal objective of the water-treatment process is to remove substances from water that are harmful, or that otherwise make the water unsuitable for human use. However, another important component of the process is the addition of chemicals that make the water better for human use. For example, fluoride (F^-) is routinely added to public water supplies to protect the teeth of those who drink the water. Cities that add appropriate

amounts of fluoride to their drinking-water supplies have successfully reduced the incidence of cavities among the children who inhabit those cities.

Schematic of a Water-Treatment Plant

The processes of screening, sedimentation, precipitation, filtration, adsorption, and disinfection work together to remove the unwanted substances from our water supply, making it safe to drink and appropriate for other uses. Addition of other chemicals, such as fluoride, further enhance the quality of the water for drinking. Figure 9, below, depicts a flowchart showing how these processes work together. Once the water is treated, it is sent to storage chambers and then distributed to household consumers, businesses, and industries.

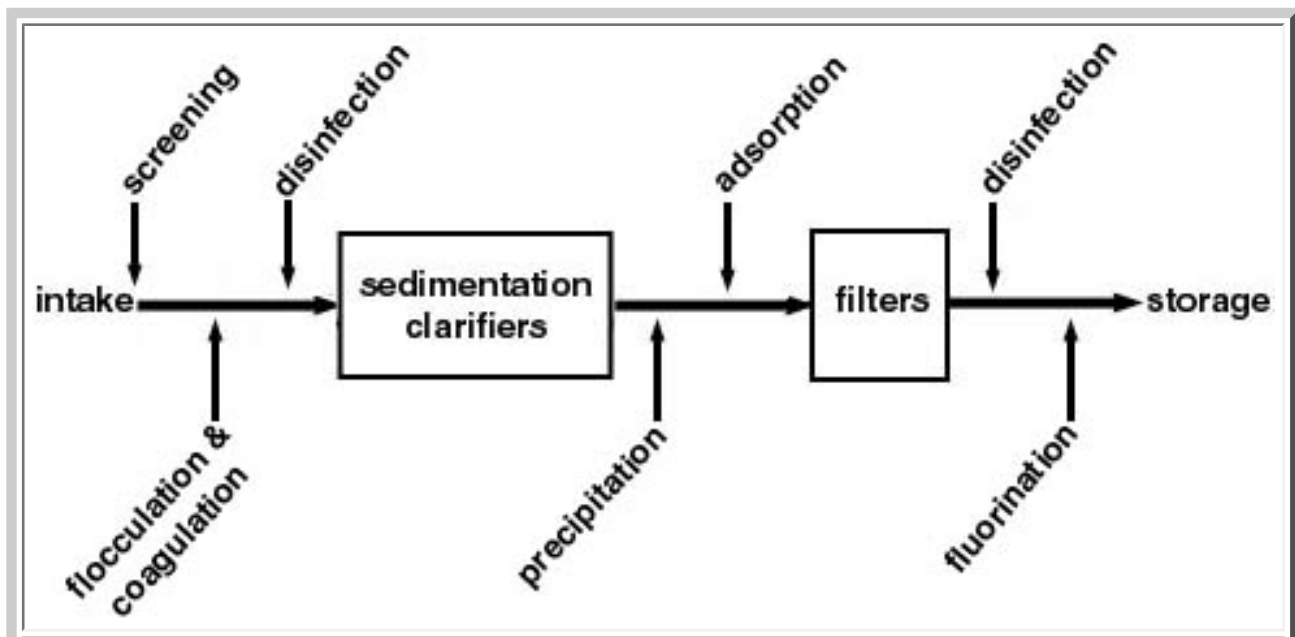


Figure 9

This flowchart shows the path that water takes from the intake of the water treatment plant (from the freshwater source) to the storage tank, from which it is pumped to homes, businesses, and industries. The specific steps and their sequence may vary somewhat from one treatment plant to another.

Questions on the Schematic of a Water-Treatment Plant

- Figure 9, "Filtration in the water-treatment process," shows that adsorption typically occurs before filtration. Briefly, explain why this order of events is important.

- How would the staff at a water-treatment plant know when the filter needs to be cleaned?
-

Point-of-Use Water Treatment

In most communities, the water that reaches our homes, businesses, and industries is clean and free from impurities that detract from the water's suitability for human use. However, water-treatment facilities in some communities do not adequately treat the water to make it safe to drink and appropriate for other uses. And even when the public water supply is considered good by most people, some users have special requirements that necessitate further treatment of the water. In these situations, special **point-of-use** treatment procedures are employed. Point-of-use water treatment includes any treatment of the water that occurs at the location where the water is to be used (*e.g.*, in the home or at an industrial site, as opposed to in a community plant from which the water will be distributed to many locations). Two of the most common types of point-of-use treatment are water softening and adsorption filtration.

Point-of-Use Water Softeners

Household water softeners (Figure 10) typically use a different process from the precipitation reaction described above (["Treating Water Hardness" section](#)), known as ion exchange. Ion-exchange devices consist of a bed of plastic (polymer) beads covalently bound to anion groups, such as -COO^- . The negative charge of these anions is balanced by Na^+ cations attached to them. When water containing Ca^{2+} and Mg^{2+} is passed through the ion exchanger, the Ca^{2+} and Mg^{2+} ions are more attracted to the anion groups than the Na^+ ions. Hence, they replace the Na^+ ions on the beads, and so the Na^+ ions (which do not form scummy residues) go into the water in place of the Ca^{2+} and Mg^{2+} .

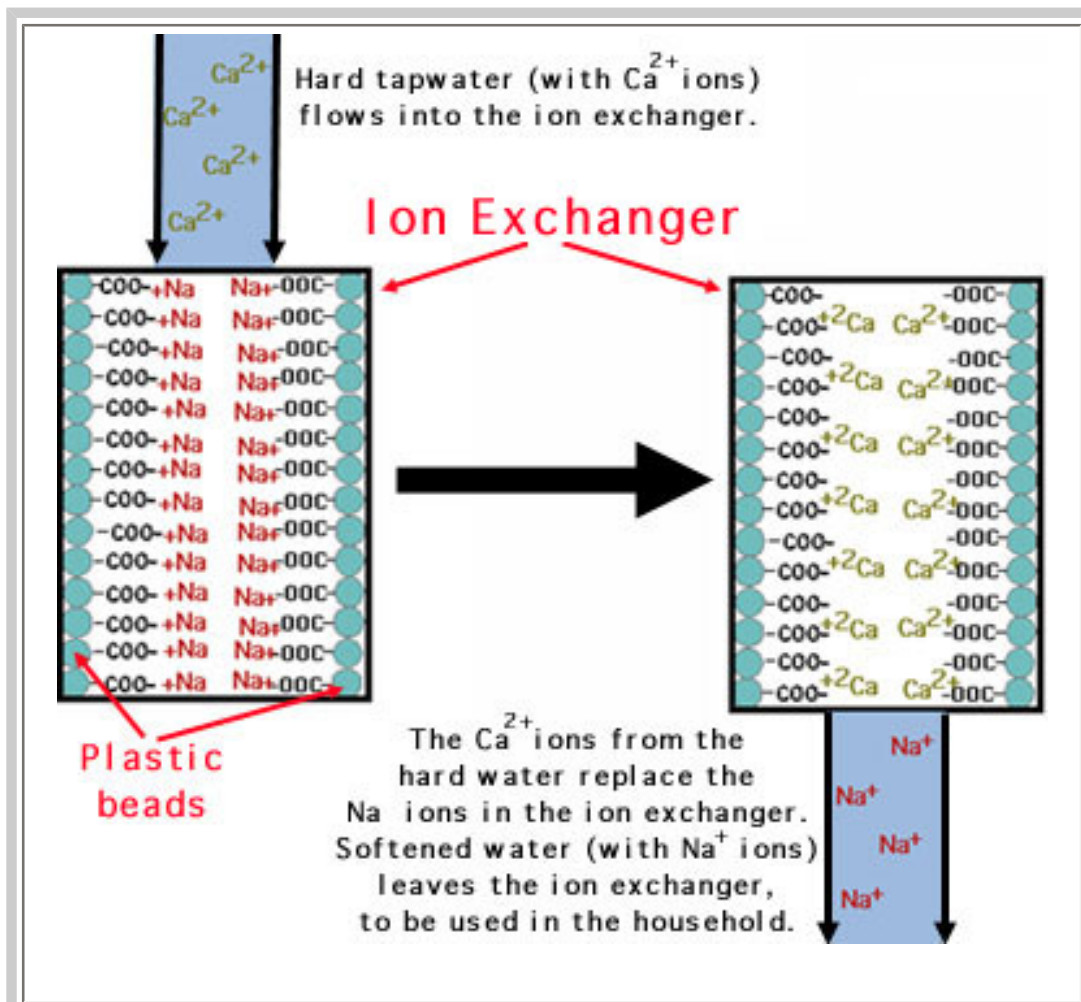


Figure 10

When hard tapwater passes through the ion exchanger, the calcium ions from the tapwater replace the sodium ions in the ion exchanger. The softened water, containing sodium ions in place of calcium ions, can be collected for household use.

Unfortunately, many people with high blood pressure or other health problems must restrict their intake of sodium. Because water softened by this type of ion exchange contains many sodium ions, people with limited sodium intakes should avoid drinking water that has been softened this way. Several new techniques for softening water without introducing sodium ions are beginning to appear on the market.

Point-of-Use Adsorption Filters

Many of the contaminants that make our drinking water unsafe or unpleasant to drink, such as lead (which may be leached into the water from lead pipes) or organic molecules producing offensive odors and tastes, can be removed by adsorption-filtration devices

installed at the tap. These devices have filters containing powdered activated carbon, which adsorbs the offending contaminants in the water. The PAC (with the unwanted contaminants attached) is strained out of the water exiting the device (for consumption) by the filter. Periodically, the filter must be replaced so that it does not become clogged and ineffective.

Questions on Point-of-Use Water Treatment

- Eventually, an ion-exchange water-softening device becomes ineffective, because all of the Na^+ ions have been replaced by Ca^{2+} and Mg^{2+} ions. Briefly, describe a way that you could "refresh" the device (i.e., flush out the Ca^{2+} and Mg^{2+} ions and replenish the Na^+ ions) using a common kitchen product.
 - Can an ion-exchange water-softening device be used to remove organic contaminants? Briefly explain your answer.
-

Completing the Cycle: What Happens to Water After We Use It?

Once water has been used, it must somehow re-enter the freshwater supply. Some of the water is evaporated (*e.g.*, if it is used to generate steam for industry, or if we drink the water and then sweat). The evaporated water eventually collects in clouds and returns to the earth via precipitation. However, most of the water that we use remains in the liquid state, and is returned to the freshwater supply directly (as runoff) or via wastewater treatment facilities.

Wastewater from municipalities (i.e., household and business use), industries, and agriculture typically follow different pathways in their return to the freshwater supply, although municipal and industrial wastewater may sometimes be combined. Wastewater from all of these sources may contain contaminants that are harmful both to humans and to the environment. Hence, water-treatment facilities (sewage facilities) must be employed to remove these harmful products before the water is returned to the freshwater supply in the environment.

Municipal wastewater may contain bacteria-harboring fecal material, as well as small amounts of grease, suspended solids, and chemicals from household use. The water passes through a wastewater-treatment facility before it is then disposed of into rivers, lakes, or estuaries, where any remaining contaminants will be diluted. Other methods of disposal, including irrigation and evaporation, are used more rarely. The wastewater-treatment facilities use many of the same techniques described above for treating water to deliver to consumers. The degree to which the water must be treated before it is returned to the

environment depends on the quality of the wastewater, as well as how the water into which the wastewater is disposed will be used. For instance, wastewater that is emptied into lakes that are used for recreational swimming and fishing must be treated particularly carefully.

Industries typically do not actually consume much water (the water is used for applications, such as cooling and processing, that do not significantly diminish the amount of water), so the amount of industrial wastewater is generally very large. Industrial wastewater may contain many contaminants, such as toxic metals, organic chemicals, and radioactive materials. Before it can be returned to the water supply, this wastewater must be carefully evaluated and treated, using many of the same techniques for water treatment described above. Then, the industrial wastewater, like municipal wastewater, is released into rivers, lakes, or estuaries.

Agricultural wastewater often contains soil sediment, as well as potentially harmful materials such as pesticides, fertilizers (which may cause an overgrowth of algae), and animal wastes (which may harbor disease-causing organisms). Unfortunately, the wastewater from agriculture often drains directly into streams, rivers, and lakes. Hence, a great deal of concern has arisen about the types of pesticides and fertilizers that are used.

Summary of Chemistry Behind Water Treatment

An understanding of chemistry is so important to the water-treatment process that water-treatment facilities hire many chemists to analyze the quality of the water and oversee its treatment. One of the most fundamental chemical principles in water treatment is **solubility**. It is imperative to understand which contaminants are soluble (forming solutions) and which are insoluble (forming suspensions), in order to determine how they can be effectively removed. Insoluble contaminants can usually be removed by physical-separation processes, including screening, sedimentation, and filtration. These physical processes may be aided by chemical processes, such as coagulation, that help to entrap the suspended particles. Soluble contaminants, on the other hand, must be removed by chemical methods that render them insoluble, so that they can then be removed by physical means, such as sedimentation and filtration. Furthermore, an understanding of solubility is essential in choosing reactants that will generate insoluble precipitates with the dissolved contaminants. Solubility guidelines, such as those listed in the introduction to this Experiment, can be instrumental in choosing the proper reagents for water treatment.

Another critical chemical concept used in water treatment is **concentration** (i.e., the amount of a substance contained per unit volume). The concentrations of contaminants in the water supply must be carefully measured (typically using chemical-analysis methods that are not discussed in this tutorial) in order to determine whether they fall within acceptable ranges, and to determine how much must be removed. An understanding of concentration is also necessary to calculate the proper amount of a treatment chemical that must be added to achieve the desired effect in treating the water.

It is interesting to note that many of the water-treatment methods described in this tutorial, especially precipitation, filtration, and adsorption, are routinely used (on a smaller scale) in the chemistry laboratory. You will perform many of these procedures in your chemistry-laboratory (and biology) courses, as well as in many research laboratories you might join.

Additional Links:

- The [St. Louis County Water Company's](#) web site gives information about the analysis and treatment procedures that are used for our local water supply.
 - This site, from the [Doulton Water Filters](#) company, contains a wealth of information about contaminants in the water and how the contaminants are removed. Check out their "[Six Chapter Training Guide to Water and its Problems](#)" in particular.
 - This tutorial developed at Washington University in St. Louis discusses the chemistry behind [acid rain](#), an important contributor to contamination in the water supply.
 - Another tutorial from Washington University in St. Louis provides more a detailed explanation of [water hardness](#), and the strategies for dealing with this problem.
 - This site contains a [virtual tour](#) of the Ames, Iowa Water Treatment Plant.
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