KCS @ home

Summer Edition

Chemistry

Activity 4

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Sinkholes
Chemistry Goes Deep

By Brian Rohrig

On February 28, 2013, Jeff Bush went to bed in his home near Tampa, Fla. Shortly afterward, the earth swallowed him up, as his home sank into a giant sinkhole. He was never seen again.

On July 3, 2013, Pamela Knox, an elementary school principal, was driving her car down a city street in Toledo, Ohio, when the ground opened up, and her car dropped out of sight as she was suddenly engulfed in a giant sinkhole. Fortunately, she escaped without serious injury.

On August 12, 2013, a large portion of a three-story building at the Summer Bay Resort near Disney World in Florida collapsed into a giant 18 meter (59 feet) diameter sinkhole. Fortunately, no one was injured.

Central Florida is the sinkhole capital of the world, and has been dubbed "Sinkhole Alley." Between 2006 and 2010, insurance companies in this region received an average of 17 claims per day from sinkhole damage! About 20% of land in the United States is susceptible to sinkholes. Pennsylvania, Kentucky, Tennessee, Missouri, Alabama, and Texas are especially prone to sinkholes.

This aerial photo shows workers picking up remnants from a home in Seffner, Fla. A huge sinkhole opened under the bedroom in this home and swallowed the owner, Jeff Bush.

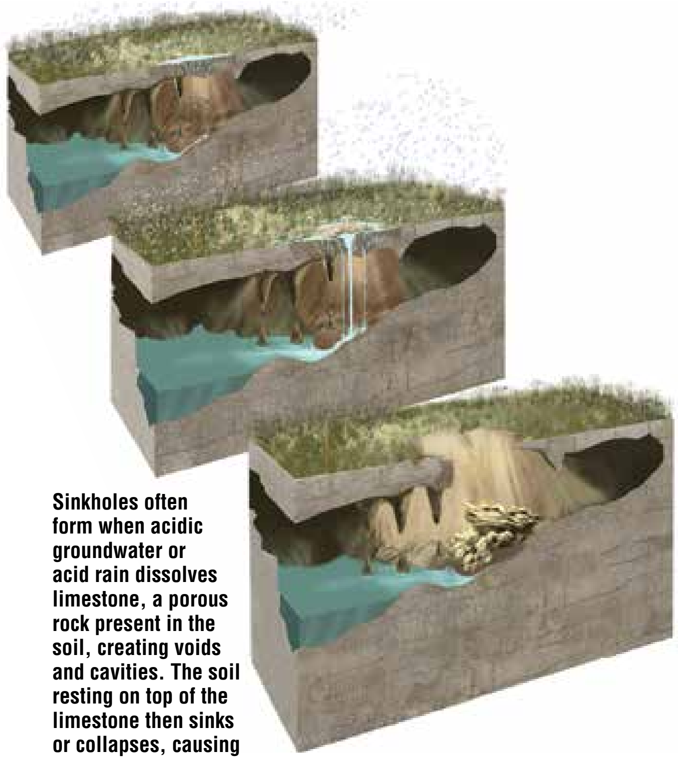
Sinkholes are a worldwide phenomenon. One of the largest natural sinkholes in the world is in Egypt. It measures 80 kilometers (km) (50 miles) long by 120 km (75 miles) wide—larger than the state of Delaware! It is 133 meters (436 feet) deep, and parts are filled with quicksand. In Chongqing, China, you can find perhaps the world's deepest sinkhole at a depth of more than seven football fields! Thousands of new sinkholes open up every day. One sinkhole in Texas, known as the Devil's Sinkhole, plunges 122 meters (400 feet) before opening into a cavern and is home to thousands of bats.

THE BALTIMORE SUN; AP IMAGES

What is a sinkhole?

A sinkhole is a depression in the ground that opens up suddenly. There are various reasons about why sinkholes form. Human activity is a major cause, and it is likely the reason the number of sinkholes is increasing in populated areas. Any time underlying soil is removed, such as through mining or construction, the soil on the surface can collapse. The sinkhole in Toledo was caused by the collapse of the roof of the underlying sewer, perhaps precipitated by a water main break.

Sinkholes can also form due to pressure from above the soil. One morning in June 2004, residents of Wildwood, Mo., woke up to find that a 23-acre lake in their development had simply disappeared. A giant sinkhole opened up under the lake and drained it dry. Ponds and lakes—usually man-made—that disappear suddenly are most likely draining into giant sinkholes. The soil underneath is simply not strong enough to support the immense weight of the water above it. Since water has a density of 1 kilogram per liter (kg/L), a cubic meter of water has a mass of 1,000 kg (2,200 pounds). The builders of the lake had no idea that there was an enormous cavity underneath the lake which it eventually emptied into, similar to pulling the plug in a bathtub.



Sinkholes often form when acidic groundwater or acid rain dissolves limestone, a porous rock present in the soil, creating voids and cavities. The soil resting on top of the limestone then sinks or collapses, causing a sinkhole.

How sinkholes form

Naturally occurring sinkholes are most commonly found in a type of terrain known as karst topography, which consists of bedrock (rock beneath the soil) filled with nooks and crannies. The underlying bedrock in karst landscapes is usually made of limestone.

A great portion of the state of Florida is, in essence, sitting atop one continuous slab of limestone, making it vulnerable to sinkholes. Limestone is composed of calcium carbonate ($CaCO_3$), which primarily comes from the remnants of corals and other types of marine organisms, whose shells are made of calcium carbonate. Limestone builds up slowly after these animals die and their shells are deposited and accumulate over time. Other substances composed of calcium carbonate include marble, chalk, Tums antacid tablets, and eggshells.

To understand how limestone bedrock contributes to sinkholes, consider what happens when you place an egg in a glass of vinegar, which contains 5% acetic acid (CH_3COOH). You will notice that little bubbles of carbon dioxide gas form almost immediately and, within a day or two, the eggshell will have completely disappeared, leaving you with the egg's translucent membrane to protect the egg. The eggshell, which is composed of calcium carbonate, does not normally dissolve in water, but in the presence of acetic acid, calcium carbonate and acetic acid react with each other, causing the eggshell to dissolve according to the following chemical reaction:

$$2\,CH_3COOH\,(aq) + CaCO_3\,(s)$$
$$\text{Acetic acid} + \text{Calcium carbonate}$$
$$\rightarrow [Ca^{2+}\,(aq) + 2CH_3COO^-\,(aq)] + H_2O\,(l)$$
$$\rightarrow \text{Calcium acetate} + \text{Water}$$
$$+\ CO_2\,(g)$$
$$+\ \text{Carbon dioxide}$$

Any substance made of calcium carbonate will react with an acid. Limestone, being made of calcium carbonate, will react with an acid and will be slowly worn away. But are there acids underground?

To answer this question, consider what happens to rainfall (which eventually becomes groundwater) as it passes through the atmosphere. While falling through the air, the rain comes into contact with carbon dioxide. Although carbon dioxide comprises

A sinkhole opened in Duluth, Minn., on June 20, 2011. Steady torrential downpours caused residents to evacuate their homes and this car to fall into the sinkhole.

only about 0.04% of the atmosphere, that is enough to make rainfall acidic, lowering its pH to about 5.6. So, by the time rainfall reaches the ground, it has turned into acid. The reaction is as follows:

$$H_2O\,(l) + CO_2\,(g) \rightarrow H_2CO_3\,(aq)$$
$$\text{Water} + \text{Carbon dioxide} \rightarrow \text{Carbonic acid}$$

Carbonic acid then dissociates to give a hydrogen ion (H^+) and a bicarbonate ion (HCO_3^-):

$$H_2CO_3\,(aq) \rightarrow H^+\,(aq) + HCO_3^-\,(aq)$$


A large sinkhole appeared suddenly in a street of Shenzhen, China. A security guard fell into it and died.

The ability of carbonic acid to dissociate by producing hydrogen ions is what makes this molecule an acid.

Over time, acidic rainwater seeps into the ground and comes into contact with limestone bedrock. Water makes its way into cracks or pockets in the rock, reacting with the limestone and eventually making holes and fissures in the rock. Sinkholes occur when acidic rainwater has eaten away so much of the underlying limestone bedrock beneath the soil that the ground collapses.

The more it rains, the greater the amount of carbonic acid leaching into the soil below. Humid areas have the most rainfall. High humidity in the air leads to cloud formation, which eventually produces rainfall. So it is no surprise that Florida leads the United States in the number of sinkholes because it has both limestone bedrock and high humidity.

The acidity of rainwater is not the only reason water in the ground is acidic. Decaying organic materials and root respiration also produce carbon dioxide, which dissolves in soil water to form carbonic acid.

Warning signs

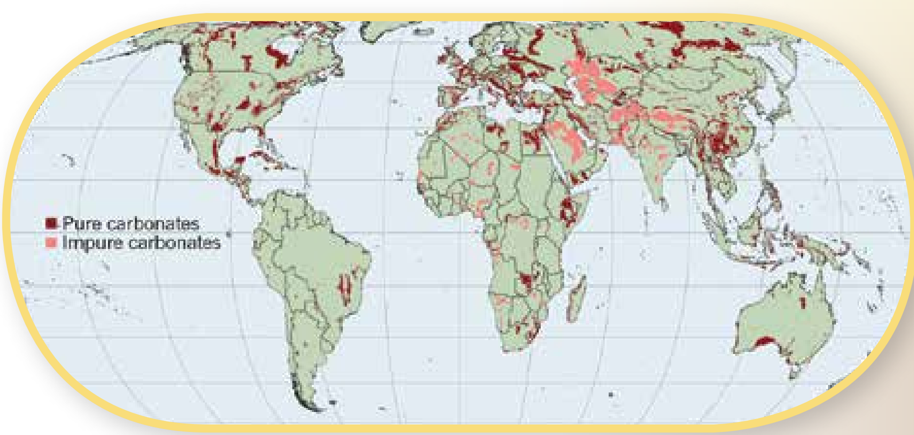
You should not be afraid of falling into a sinkhole on your way to school tomorrow or be concerned about a sinkhole swallowing up your school, because there are ways to detect a potential sinkhole. Look for dying vegetation, sudden appearance of standing water, muddy well water, cracks in the ground, and fence posts or signs that appear to be slumped over. If you live in a house, look for a crumbling foundation or doors and windows that do not close properly.

If a sinkhole is suspected, ground-penetrating radar can be used to confirm your suspicion. This method uses electromagnetic radiation in the microwave portion of the spectrum to look underground. This is the same type of radar used by stealth bombers and by the U.S. military to locate roadside bombs in Iraq and Afghanistan. If a large underground void is detected, then further investigation is warranted. It is always a good idea to have your property inspected for sinkholes by a trained professional, especially if you live in a sinkhole-prone area.

The concept of a sinkhole makes one wonder about the expression "solid ground." Just beneath our feet is an active world of physical and chemical processes. Sinkholes are a scary—and sometimes tragic—reminder that our planet is dynamic and ever-changing. Fortunately, our technology has advanced to the point where we can detect sinkholes, and, hopefully, it can be used to prevent more tragedies from occurring. *CM*

Soft Drinks and Carbonic Acid

If you read the first ingredient on the label of any type of carbonated soft drink, it will most likely read "carbonated water," which is mostly water and a small amount of carbonic acid (H_2CO_3). Soft drinks are carbonated by adding carbon dioxide gas to water, which reacts to form carbonic acid, contributing to their sour taste. Carbonic acid is a weak acid, so it is harmless when consumed as part of a soft drink.



This map shows the areas where carbonate rocks can be found throughout the world. These areas are more susceptible to erosion by running water and thus to the formation of sinkholes.

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salting Roads

THE SOLUTION FOR WINTER DRIVING

By Doris R. Kimbrough

Most of you are probably planning to be in a car at some point this winter, so be sure to wear your seatbelt, obey the speed limit, and be extra cautious in "winter driving conditions". It is actually physics that makes being in a car on an icy road hazardous; friction is an important part of keeping a car under control. However, your state and local transportation departments make use of some pretty interesting chemistry to keep roads safer for travelers in the wintertime. In addition to plowing, one of the ways that highway workers keep roads clear of ice and snow is by spreading salt on the roads. Even though salt can cause rust and corrosion on cars, bridges, and other parts of the highway, it more than makes up for this costly damage by saving lives. Let's take a look at how it works.

Freezing point depression

Pure water freezes at 0°C; adding salt to water depresses or lowers the freezing point below zero. When you remove heat from water (or any substance), the molecules slow down. The freezing process occurs when the molecules stop sliding and tumbling all over each other (liquid phase) and settle into fixed positions in a large network called a *crystal lattice,* which is the solid phase. The molecules are still moving, but in the solid phase that motion involves bonds stretching and compressing or the atoms wiggling a little bit. This is called vibrational motion.

When a solution of salt in water is cooled to a low enough temperature, the water molecules begin to stick together in an organized way to form solid crystals. The crystal framework tends not to include the salt ions because the ions would disturb this organization. So when you cool a solution enough, the ice crystals that start to form are made of pure frozen water. You can actually purify salty water like this, by freezing a portion of the solution and washing the salty water off of the ice crystals and then thawing the ice to produce pure water. Eventually, if you cooled it enough, the whole solution will freeze, but it does not have a sharply defined freezing temperature. Getting the water molecules organized into a crystal from a solution requires that you remove more energy (actually free energy) than if you are freezing pure water, so the water in a solution typically does not start freezing until it reaches a lower temperature than the normal freezing point. This is true of all solutions, not just those made with water.

The difference in temperature between where the pure solvent (water in our case) freezes and where the solution starts to freeze is called the *freezing point depression*. How low

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you go depends upon what the solvent is and how concentrated the solution is. The more concentrated the solution, the lower the freezing point. This is why if you freeze a solution completely, you have to keep lowering the temperature. As the pure water freezes away from the solution, the concentration of the solution remaining increases. How does your Department of Transportation use this chemistry to keep roadways clear of snow and ice in the wintertime?

Salting roads

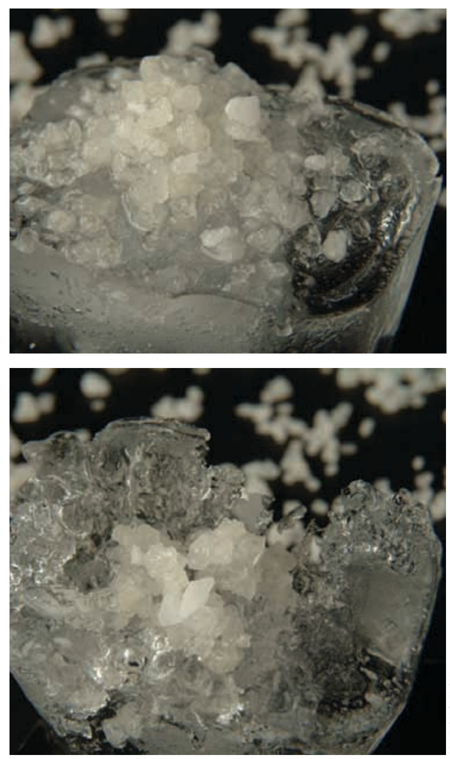
Highway workers use salt in two ways: 1) to melt ice that is already on the roadway and 2) to prevent ice from forming on the roadway. The second one is a little easier to understand, so we will start there. Let's say a snowstorm is forecast for your town. Municipal workers get out and spread salt on the roads. As the storm hits, snow starts to fall, but the road surface is warmer than the air, so the first flakes melt. As they melt, the salt dissolves in the liquid water. Now you have a solution of salty water, which has a lower freezing point than pure snow, so that even though the additional snow might cool the road enough to "stick" to the road surface, the temperature will not get cold enough to freeze the solution the way it could freeze pure water. In the end, the real snow removal is done by the plows, but salt plays a crucial role in preventing snow and ice from bonding to the pavement.

"Hold it!" you say. Suppose the temperature does get cold enough to freeze the solution. Or suppose that enough snow falls so that the salt water solution gets too dilute to work, what happens then? In both of those cases, snow could build up on the road. This is why your highway workers are out there round-the-clock, plowing and spreading more salt as long as the storm is in the area. Communities with really cold temperatures, like parts of Canada and Alaska, where it can get to –20°C (below zero on the Fahrenheit scale), don't even bother with salting the roads, because it doesn't help. They typically plow off as much as they can and use gravel or sand to add traction.

How does it work if the road already has snow or ice on it? You may have seen how spreading salt on an icy sidewalk will cause the snow or ice to melt. The very beginning of the melting process begins where solid ice contacts solid salt. At the surface of ice crystals, water molecules do not have the same stable arrangement they have on the interior; they are more mobile and more reactive. So the surface ice reacts with the surface of the salt crystals, allowing a small amount of salty

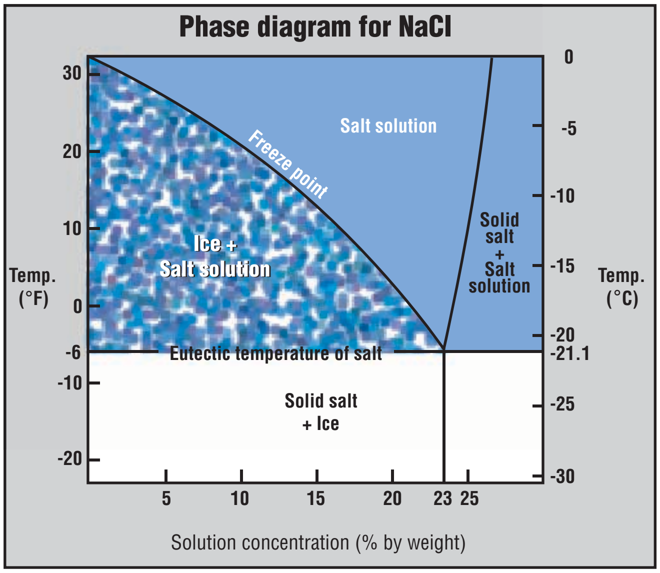
MIKE CIESIELSKI

Sodium chloride and calcium chloride.



MIKE CIESIELSKI

Water freezes at lower temperatures when it is saltier. But solubility of salt in water decreases with decreasing temperature. So at some point, a solution is too cold to hold the salt that keeps it from freezing. The *eutectic* temperature is the lowest temperature at which a mixture of two or more substances can stay liquefied. For a NaCl and water mixture, the eutectic temperature is –21.1 °C. For road ice, –10 °C is the practical limit for salt.



This phase diagram illustrates the impact of NaCl concentration and temperature on the phase of an aqueous NaCl solution.

solution to form. This first step is relatively slow, but then the growing solution continues to dissolve more salt and melt more ice. Passing vehicles warm the slush through friction, which speeds the dissolution of the ice and may crush the salt and ice together, which will increase the surface areas of the particles in contact with each other. Some communities use "prewetted" salt (usually rock salt with a $CaCl_2$ solution sprayed on it) to speed the process.

Road salt history, fun facts, and current technology

Putting salt on the roads to lessen the buildup of snow and ice began in the 1930s, and by the 1960s, it was used by most communities where snow and ice are a problem. Concerns about the effect of the use of sodium chloride (common table salt) on the environment have prompted some state and local road crews to explore the use of more environmentally friendly salts, such as magnesium chloride and calcium chloride. These two salts have the advantage of being more effective at lowering the freezing point and there is some interesting chemistry behind this benefit.

As you probably know, sodium chloride has the formula NaCl. When it dissolves in water, it dissociates into its ions: Na^+ and Cl^-, producing two ions in solution for every NaCl formula unit. Magnesium chloride ($MgCl_2$) and calcium chloride ($CaCl_2$) dissociate to three ions each because the metal has a 2^+ charge and there are two chlorides per metal ion:

$$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$$
$$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$$

It is the number of dissolved particles that determine the extent of the lowering of the freezing point of a solution. So although NaCl produces two ions, $MgCl_2$ and $CaCl_2$ each produce three, making them more effective. Other variations include mixtures of the magnesium and calcium chlorides, as well as magnesium and calcium acetates, $Mg(C_2H_3O_2)_2$ and $Ca(C_2H_3O_2)_2$.

	Practical Melt Temp.	Eutectic Temp.
$CaCl_2$	−32°C	−56°C
$MgCl_2$	−15°C	−33°C

The technology of salting roads has become fairly sophisticated. Often, these salts are dissolved in water or some other solvent so that they can be sprayed onto the road surface. Having the deicing substance in a solution (i.e., fluid) form makes it possible to pump through hoses, which allows for a more targeted application. In addition, various anticorrosive substances are added to protect highways and cars from the damage the salts can cause over time. Some 15 million tons of deicing salt is used each year in the United States and about 4–5 million tons in Canada.

You may have seen signs that warn about bridges freezing before road surfaces. This is because bridges are more exposed and not insulated by the ground from underneath like the rest of the highway. Some high-tech highway bridges have been constructed with deicing sprayers built right into the pavement, complete with sensors that detect when conditions are right (e.g., cold temperatures, high wind speeds, high humidity) for ice to form. The sensors detect the possible formation of ice, and the deicing sprayers go to work to keep the roadway from freezing.

Highway engineers have been working with other interesting variations. One deicing material that is currently on the market mixes magnesium chloride with sugar cane or sugar beet molasses. The sticky molasses keeps the magnesium chloride from getting blown or washed off the road surface. There are also substances that are added directly to the top layer of concrete or asphalt when the road is built or repaved that help prevent ice from forming. Highway workers can then get away with using less salt than before, which is cheaper, easier on the environment, and helpful in preventing corrosion. Scientists and engineers continue to develop new ways to keep winter highways safe while minimizing expense and environmental harm. Just another way that chemistry is keeping you out of harm's way.

MIKE CIESIELSKI

Doris Kimbrough teaches chemistry at the University of Colorado-Denver. Her last article, "Einstein's Miraculous Year", appeared in the December 2005 issue.

KCS Chemistry I Problem Set- Concentrations Name

Molarity (M) = $\dfrac{\text{moles of solute}}{\text{Liters of solution}}$	Dilution Equation: $M_1V_1 = M_2V_2$

For all problems, assume that the volume of the solution is the same as the volume of the solvent.

1. Determine the molarity of a solution prepared by dissolving 0.700 moles of sodium chloride (NaCl) in 2.00L of water.

2. Determine the molarity of a solution prepared by dissolving 20.0g of sodium hydroxide (NaOH) in 1.00L of water.

3. What is the molarity of a salt solution made by dissolving 280.0g of NaCl in 250.0 mL of water?

4. How many moles of calcium chloride ($CaCl_2$) are in 250.0mL of a 1.50 M solution of $CaCl_2$?

5. What mass of magnesium bromide ($MgBr_2$) would be required to prepared 720.0mL of a 0.0939M aqueous solution? (Find moles first, then convert to grams)

6. How many moles of $Cu(NO_3)_2$ are needed to prepare 2.35L of a 2.00 M solution?

7. How many grams of magnesium carbonate ($MgCO_3$) are needed to prepare 3.00L of a 0.750M solution?

8. What is the molarity of a solution of ammonium chloride prepared by diluting 50.00mL of a 3.79 M NH_4Cl solution to 2.00L? (Dilution equation)

9. What volume of water would be added to 16.5mL of a 0.0813 M solution of sodium borate in order to get a 0.0200M solution? (Dilution equation)

10. Describe how you would prepare 1.00L of a 0.100M solution of HCl (hydrochloric acid) if you had a stock solution of 6.00M HCl. How would this be different than preparing a 0.100M solution of NaCl from the solid?

Chem. Acid/Base WS#2 **Name** ______________________

1. Determine the pH of a 0.00118 M solution of HBr.

2. Determine the pH of a solution with a hydrogen (hydronium) ion concentration of $5.6x10^{-5}$ M.

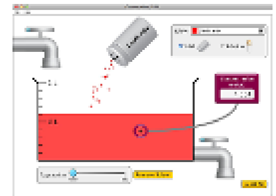
3. What is the pH of a solution with the hydrogen ion concentration of $9.6x10^{-9}$M?

4. A hydrochloric acid solution has a pH of 1.70. What is the $[H_3O^+]$ is this solution? Considering that HCl is a strong acid, what is the HCl concentration of the solution?

Author: Chris Bires, revised 3/2012 Simulations at http://phet.colorado.edu/ Name: ________________________

Concentration and Molarity PhET-Chemistry Labs

Introduction: Everyone likes candy. Have you ever wondered how that candy is produced? How do they get all that delicious sugar into those tiny packages? Could you make hard candy like those you can buy? It's easier than you think. Web searching for "rock candy" will yield a number of delicious recipes you can try at home.



Concentration

Some handy vocabulary for you to define:

Solute ____Solvent____

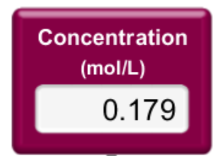
Moles ____

Molarity ____

Saturated (not fats)____

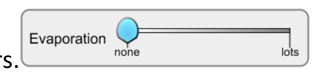
Unsaturated (not fats)____

Supersaturated____



Procedure: *PhET→Play with the Sims →Chemistry →Concentration* 

Part 1: Dissolution and Saturation

Take some time to play and familiarize yourself with the simulation. Click on everything. Move all the sliders. Notice what happens to the concentration as solid solute  is added and when evaporation occurs. 

How does the concentration change as solid solute is added? ____

How does the concentration change as additional water is added? ____

How does the concentration change as evaporation occurs? ____

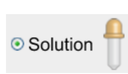
How do you know when a solution is ***saturated***? ____

When a solution *is* saturated, and additional solid solute is added, what happens? ___________________________________

Why do you think this is? ____

How does adding this additional solute change the concentration of this saturated solution? ____________________________

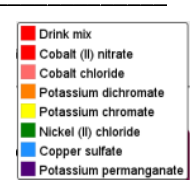
How does evaporation change the concentration of a saturated solution? ____

Part 2: Concentrated Solutions 

Adding a concentrated solution… describe a way to determine the concentration of the solution in the spigot. Write your plan here: ____

____

Using your plan…how might you get that concentrated solution to become saturated? ___________



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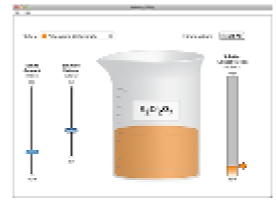
Does your plan work for all the other solutions too? __________ Why? / Why Not?_____________________

____

Author: Chris Bires, revised 3/2012   Simulations at http://phet.colorado.edu/   Name: ___________________________

$$Molarity\,(M) = \frac{amount\ of\ solute\,(mol)}{volume\ of\ solution\,(L)}$$

Part 3: Molarity *…PhET→Play with the Sims → Chemistry → Molarity* Run Now!



Molarity

Molarity is **moles per Liter**, that is, how many moles of solute (entire salt) is dissolved per Liter of solution.

First, determine the **saturation concentration** of each of the solutions, that is, how concentrated can you get each solution before the solution is saturated. If you can't determine the concentration using the simulation "Molarity", try using the simulation "Concentration" (You will use this information again in ***Part 5***, if your instructor requires it)

Cobalt (II) nitrate	Saturation concentration	Potassium chromate	Saturation concentration
Cobalt chloride		Nickel (II) chloride	
Potassium dichromate		Copper sulfate	
Gold (III) chloride		Potassium permanganate	

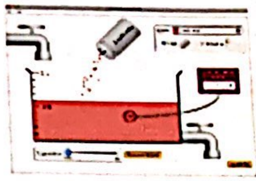
Part 4: Calculating Molarity   Using the simulation and the formula for Molarity above, complete the table below.

Moles of Compound (mol)	Liters of Solution (L)	Molarity of Solution (M)	Moles of Compound (mol)	Liters of Solution (L)	Molarity of Solution (M)
.53	.79			.78	.59
.86	.34		.88		1.8
1.0	.20		3.5	8.4	
.67	.67			6.4	8.5

Author: Chris Bires, revised 3/2012 Simulations at http://phet.colorado.edu/ KEY Name: [redacted]

Concentration and Molarity PhET-Chemistry Labs

Introduction: Everyone likes candy. Have you ever wondered how that candy is produced? How do they get all that delicious sugar into those tiny packages? Could you make hard candy like those you can buy? It's easier than you think. Web searching for "rock candy" will yield a number of delicious recipes you can try at home.



Concentration

Some handy vocabulary for you to define:

Solute the component in a solution that dissolves into the solvent Solvent the liquid or gas that dissolves the solute

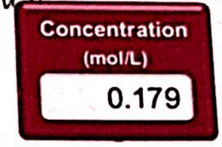
Moles SI unit for amount of substance

Molarity the number of moles of solute per liter of solution

Saturated (not fats) a substance in which the atoms are linked by single bonds

Unsaturated (not fats) a molecule/substance with double or triple bonds

Supersaturated a solution that has more of the solute than the solvent can dissolve



Procedure: *PhET→Play with the Sims → Chemistry → Concentration* Run Now!

Part 1: Dissolution and Saturation

Take some time to play and familiarize yourself with the simulation. Click on everything. Move all the sliders. Notice what happens to the concentration as solid solute  is added and when evaporation occurs. 

How does the concentration change as solid solute is added? The concentration increases

How does the concentration change as additional water is added? The concentration decreases

How does the concentration change as evaporation occurs? The concentration increases

How do you know when a solution is *saturated*? The solution is saturated when the concentration reaches its max.

When a solution *is* saturated, and additional solid solute is added, what happens? It piles at the bot

Why do you think this is? The solvent cannot dissolve anymore solute

How does adding this additional solute change the concentration of this saturated solution? It does not change i

How does evaporation change the concentration of a saturated solution? It also does not change the concentration

Part 2: Concentrated Solutions ⊙ Solution 

Adding a concentrated solution... describe a way to determine the concentration of the solution in the spigot. Write your plan here: Drain all the water out and just pour the solutio in.

Using your plan...how might you get that concentrated solution to become saturated? Pour the solid solute into the solution.

Does your plan work for all the other solutions too? Yes Why? / Why Not? None of the solutions start fully saturated, so adding more solute will cause them to become saturated.

$$Molarity\ (M) = \frac{amount\ of\ solute\ (mol)}{volume\ of\ solution\ (L)}$$

***Part 3: Molarity** ...PhET→Play with the Sims → Chemistry → Molarity* Run Now!

Molarity is moles per Liter, that is, how many moles of solute (entire salt) is dissolved per Liter of solution.

First, determine the **saturation concentration** of each of the solutions, that is, how concentrated can you get each solution before the solution is saturated. If you can't determine the concentration using the simulation "Molarity", try using the simulation "Concentration" (You will use this information again in ***Part 5***, if your instructor requires it)

Molarity

Cobalt (II) nitrate	Saturation concentration 5.000 mol/L	Potassium chromate	Saturation concentration 3.000 mol/L
Cobalt chloride	4.000 mol/L	Nickel (II) chloride	5.000 mol/L
Potassium dichromate	0.500 mol/L	Copper sulfate	1.000 mol/L
Gold (III) chloride	0.452 mol/L	Potassium permanganate	0.400 mol/L

Part 4: Calculating Molarity Using the simulation and the formula for Molarity above, complete the table below.

Moles of Compound (mol)	Liters of Solution (L)	Molarity of Solution (M)	Moles of Compound (mol)	Liters of Solution (L)	Molarity of Solution (M)
.53	.79	.67 M	.461 mol	.78	.59
.86	.34	2.5 M	.88	1.49	1.8
1.0	.20	5.0 M	3.5	8.4	.42
.67	.67	1.0 M	1.3	6.4	8.5

Part 5: (Extension Exercise) Total Ion Concentration (this will be important for Equilibrium, Kinetics, and Acid-Base)

Just as an entire solution has a concentration, so does each individual ion. For instance, since there are three ions when a Calcium Chloride $CaCl_2$ molecule dissolves into solution, a 3.0 M solution of $CaCl_2$ is 3.0 M with respect to Ca^{2+} ions and 6.0 M with respect to Cl^- ions, for an overall ion concentration (solubility) of 9.0 M (3.0 M + 6.0 M).

Using what you know about inorganic nomenclature and common ions, complete the table below

Compound	Saturated Concentration (from Part 3)	Cation Molarity	Anion Molarity	Total Ion Solubility
$Co(NO_3)_2$	5.000 mol/L			
$CoCl_2$				
$K_2Cr_2O_7$				
$AuCl_3$				
K_2CrO_4				
$NiCl_2$				
$CuSO_4$				

KCS Chemistry I Problem Set- Concentrations

Name KEY

Molarity (M) = $\frac{\text{moles of solute}}{\text{Liters of solution}}$	Dilution Equation: $M_1V_1 = M_2V_2$

For all problems, assume that the volume of the solution is the same as the volume of the solvent.

1. Determine the molarity of a solution prepared by dissolving 0.700 moles of sodium chloride (NaCl) in 2.00L of water.

$$\frac{.700\text{ moles}}{2.00L} = \boxed{.350\ M}$$

2. Determine the molarity of a solution prepared by dissolving 20.0g of sodium hydroxide (NaOH) in 1.00L of water.

$$\frac{20.0g\ NaOH \mid 1\ mol}{\mid 40.0g\ NaOH} = 0.500 mol \qquad \frac{0.500\ mol}{1.00L} = \boxed{.500\ M}$$

3. What is the molarity of a salt solution made by dissolving 280.0g of NaCl in 250.0 mL of water?

$$\frac{280.0g\ NaCl \mid 1\ mol\ NaCl}{\mid 58.5g\ NaCl} = 4.79\text{ moles} \qquad \frac{4.79\text{ moles}}{0.250\ L} = \boxed{19.1M}$$

4. How many moles of calcium chloride ($CaCl_2$) are in 250.0mL of a 1.50 M solution of $CaCl_2$?

$$M = \frac{mol}{L} = 1.50\ M = \frac{X\text{ moles}}{0.250L} = \boxed{0.375\text{ moles } CaCl_2}$$

5. What mass of magnesium bromide ($MgBr_2$) would be required to prepared 720.0mL of a 0.0939M aqueous solution? (Find moles first, then convert to grams)

$$M = \frac{mol}{L} \quad 0.0939M = \frac{X\ mol}{0.720L} = \frac{.0676\ mol\ MgBr_2 \mid 184.1g\ MgBr_2}{\mid 1\ mol}$$

$$\boxed{12.4g\ MgBr_2}$$

6. How many moles of $Cu(NO_3)_2$ are needed to prepare 2.35L of a 2.00 M solution?

$$2.00M = \frac{x\ mol}{2.35L} = \boxed{4.70\ mol\ Cu(NO_3)_2}$$

7. How many grams of magnesium carbonate ($MgCO_3$) are needed to prepare 3.00L of a 0.750M solution?

$$.750M = \frac{x\ mol}{3.00L} = \frac{2.25\ mol\ MgCO_3 \mid 84.3g\ MgCO_3}{\mid 1\ mol\ MgCO_3}$$

$$\boxed{190.\ g\ MgCO_3}$$

8. What is the molarity of a solution of ammonium chloride prepared by diluting 50.00mL of a 3.79 M NH_4Cl solution to 2.00L? (Dilution equation)

$$\frac{3.79M \cdot .050L}{2.00L} = \frac{M_2 \cdot \cancel{2.00L}}{\cancel{2.00L}}$$

$$M_2 = \boxed{0.0948\ M}$$

9. What volume of water would be added to 16.5mL of a 0.0813 M solution of sodium borate in order to get a 0.0200M solution? (Dilution equation)

$$.0813M \cdot 16.5mL = .0200M \cdot V_2$$

$V_2 = 67\ mL$ total $67ml - 16.5mL = \boxed{50.5mL\ \text{water}}$

10. Describe how you would prepare 1.00L of a 0.100M solution of HCl (hydrochloric acid) if you had a stock solution of 6.00M HCl. How would this be different than preparing a 0.100M solution of NaCl from the solid?

$$6.00M \cdot V_1 = 0.100M \cdot 1.00L$$

$$V_1 = .0167L \text{ or } 16.7mL$$

1. Measure out 16.7mL of the stock acid
2. Add to a 1.00L volumetric flask
3. Add Distilled water to the mark and mix.

Chem. Acid/Base WS#2 **Name** KEY

1. Determine the pH of a 0.00118 M solution of HBr.

$[H^+]$

$pH = -\log[H^+]$
$= -\log[.00118] = \boxed{2.93}$

2. Determine the pH of a solution with a hydrogen (hydronium) ion concentration of $5.6x10^{-5}$ M.

$pH = -\log[5.6\times10^{-5}] = \boxed{4.25}$

3. What is the pH of a solution with the hydrogen ion concentration of $9.6x10^{-9}$M?

$pH = -\log[9.6\times10^{-9}] = \boxed{8.02}$

4. A hydrochloric acid solution has a pH of 1.70. What is the $[H_3O^+]$ is this solution? Considering that HCl is a strong acid, what is the HCl concentration of the solution?

$pH = -\log[H^+]$
$1.70 = -\log[H^+]$
$-1.70 = \log[H^+]$
$10^{-1.70} = [H^+]$
$[H^+] = \boxed{0.0200 \text{ or } 2.0\times10^{-2} M}$