

Tutorial 13

Solutions

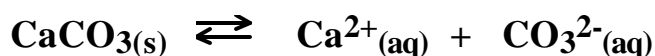
The Common Ion Effect and

Altering Solubility

Answer to Question 1 on page 4 of Tutorial 13.

1. Predict which compounds would decrease the solubility of $\text{CaCO}_3(\text{s})$ if added to a saturated solution. For each compound that does, state why it does.

Refer to the following equilibrium when reading the “Reason for Effect” below:



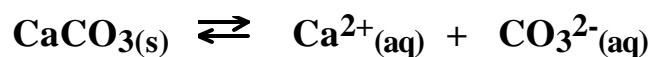
Added compound	Ions		Effect on Solubility of $\text{CaCO}_3(\text{s})$	Reason for effect
$\text{Ca}(\text{NO}_3)_2$	Ca^{2+}	NO_3^-	Decrease	The Ca^{2+} is a common ion. Increasing the $[\text{Ca}^{2+}]$ causes the above equilibrium to shift left and decrease the solubility of $\text{CaCO}_3(\text{s})$
KNO_3	K^+	NO_3^-	No Effect	Neither K^+ nor NO_3^- are common ions to the $\text{CaCO}_3(\text{s})$. Neither will have an effect on the above equilibrium.
K_2CO_3	K^+	CO_3^{2-}	Decrease	The CO_3^{2-} is a common ion. Increasing the $[\text{CO}_3^{2-}]$ causes the above equilibrium to shift left and decrease the solubility of $\text{CaCO}_3(\text{s})$
CaCO_3	Ca^{2+}	CO_3^{2-}	No Effect	Adding more of the <u>same</u> compound will have no effect on how much will dissolve. The increased surface area may have an effect on the <u>rate</u> of reaching equilibrium, but the solubility is fixed by the K_{sp} and the temperature.

Answer to Question 2 on page 7 of Tutorial 13.

2. Some buildings and statues are made of marble, which is mainly *calcium carbonate* (CaCO_3). Using the concepts in this tutorial, explain how **acid rain** can damage these structures.

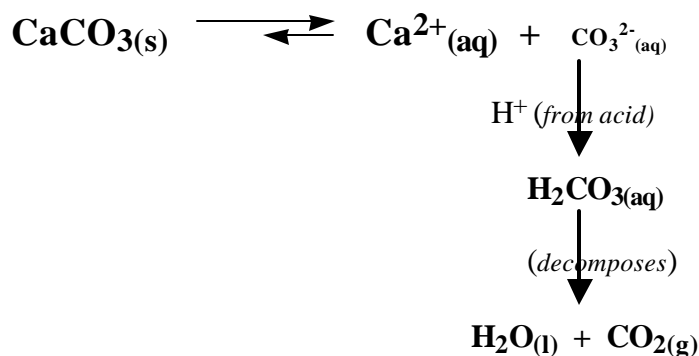
Answer:

When rain hits the marble (CaCO_3) surface, a very small amount of the CaCO_3 dissolves, forming Ca^{2+} ions and CO_3^{2-} ions and an equilibrium is established.



Acid rain, like any acid solution releases hydrogen (H^+) ions.

Hydrogen (H^+) ions react with carbonate (CO_3^{2-}) ions from this equilibrium, to form a solution of carbonic acid (H_2CO_3). The carbonic acid, being unstable in water solution, decomposes into water (H_2O) and carbon dioxide (CO_2):



This process decreases the $[\text{CO}_3^{2-}]$ in the original equilibrium. This, in turn causes the equilibrium to shift to the **right** and dissolve the solid calcium carbonate. As this happens, the statues or buildings are disfigured.

Answer to Question 3 on page 9 of Tutorial 13.

3. Suggest two different compounds which could be added to a saturated solution of calcium hydroxide ($\text{Ca(OH)}_{2(s)}$) in order to increase its solubility. Show with equilibrium equations how each one works.

First, write the equilibrium equation for the dissociation of Ca(OH)_2 .



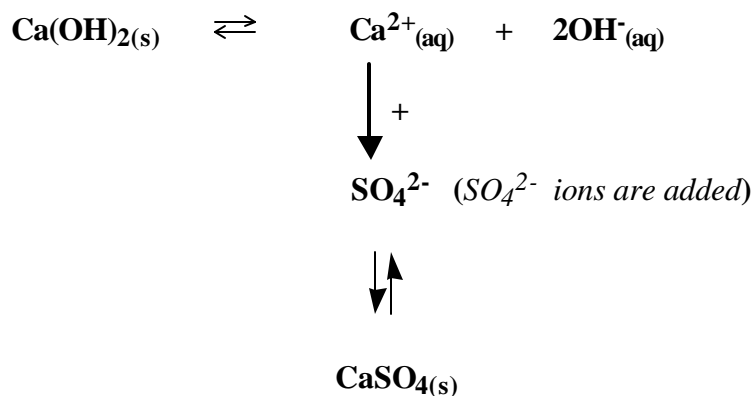
So first we could add an ion (other than OH^-) that forms a precipitate with Ca^{2+} . Looking at the solubility table, such ions could include sulphate (SO_4^{2-}), phosphate (PO_4^{3-}), carbonate (CO_3^{2-}), or sulphite (SO_3^{2-}).

So compounds that could be added would include any of the sodium or potassium salts of any of these ions: eg. Na_2SO_4 , Na_3PO_4 , Na_2CO_3 , Na_2SO_3 etc.

IMPORTANT NOTE: When you are writing the formula for a compound containing an ion you want to use, be really, really, really sure that you write the correct formula for it! Take the time necessary to look up the charges. This is a Grade 9 level skill and it looks really bad when Chemistry 12 students don't write the correct formula for a compound. I have sadly had to take marks off on tests when a student writes something like "NaCO₃" for a compound rather than the correct Na₂CO₃.

To explain why one of these compounds does increase the solubility of Ca(OH)_2 use the following example (assuming that Na_2SO_4 (SO_4^{2-}) was chosen).

(The answer would also be correct if any of the other correct answers are substituted, eg. Na_3PO_4 etc.)



The added sulphate (SO_4^{2-}) ions precipitate with the calcium (Ca^{2+}) ions to form solid CaSO_4 . This decreases the $[\text{Ca}^{2+}]$ in solution, causing the $\text{Ca}(\text{OH})_2$ equilibrium to shift to the right, thus increasing the amount of $\text{Ca}(\text{OH})_2$ that dissolves.

Looking at the equilibrium equation again:



If we add something that forms a precipitate with hydroxide (OH^-) ions (other than Ca^{2+}), then we can decrease $[\text{OH}^-]$, shifting this equilibrium to the right and dissolving the solid $\text{Ca}(\text{OH})_2$.

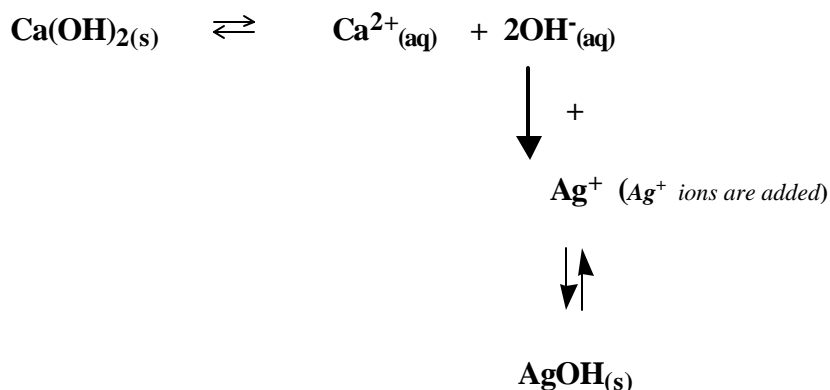
Looking at the solubility table, the following positive ions will form a precipitate with OH^- ions:

(anything BUT alkali ions, H^+ , NH_4^+ , or Sr^{2+}).

Some common positive ions that are not in this group are Be^{2+} , Mg^{2+} , Ag^+ , Pb^{2+} , Fe^{2+} etc. etc.

It is best to use nitrate compounds of these, so some compounds that are correct could include $\text{Be}(\text{NO}_3)_2$, $\text{Mg}(\text{NO}_3)_2$, AgNO_3 , $\text{Pb}(\text{NO}_3)_2$, $\text{Fe}(\text{NO}_3)_2$ etc.

If you chose AgNO_3 (Ag^+), the explanation could be like the following: (*Remember that any of the other answers, and many other possible ones could be substituted here. If you have a answer and you don't know whether it is correct, ask the teacher or another student!*)



(Don't worry about the '2' on the OH^- here. We're just looking at WHAT reacts, not how much.)

The added silver (Ag^+) ions precipitate with the hydroxide (OH^-) ions to form solid AgOH . This decreases the $[\text{OH}^-]$ in solution, causing the $\text{Ca}(\text{OH})_2$ equilibrium to shift to the right, thus increasing the amount of $\text{Ca}(\text{OH})_2$ that dissolves.

Answers to Self-Test starting on page 11 of Tutorial 13

1. The following table shows some compounds with low solubility in the left column. In column 2, a solution (reagent) is added. In column 3, indicate whether the solubility of the compound on the left will be increased, decreased or not affected. In column 4 give a brief explanation for your answer. You don't need to include equilibrium equations in your explanations in this case.

<i>Low Solubility Compound</i>	<i>Added Reagent</i>	<i>Effect on Solubility of Compound in Column 1</i>	<i>Explanation for Effect</i>
SrSO_4	$\text{Ba}(\text{NO}_3)_2(\text{aq})$	increase	The Ba^{2+} from the barium nitrate will precipitate the SO_4^{2-} ion from the SrSO_4 equilibrium, forming BaSO_4 . Decreasing the $[\text{SO}_4^{2-}]$ will cause the SrSO_4 equilibrium to shift to the ion side, thus dissolving the solid.
Ag_2S	$\text{AgNO}_3(\text{aq})$	decrease	The Ag^+ ion from the silver nitrate is a common ion to the Ag^+ from the Ag_2S equilibrium. The increased $[\text{Ag}^+]$ causes the Ag_2S equilibrium to shift toward the solid side and decrease the solubility of Ag_2S .
SrCO_3	$\text{HNO}_3(\text{aq})$ (nitric acid)	increase	The H^+ ions from the acid react with the carbonate (CO_3^{2-}) ions from the SrCO_3 equilibrium, forming carbonic acid (H_2CO_3). The carbonic acid decomposes into carbon dioxide and water. The decreased $[\text{CO}_3^{2-}]$ in the SrCO_3 equilibrium causes a shift toward the ion side, thus increasing the solubility of SrCO_3 .
AgBr	$\text{Pb}(\text{NO}_3)_2(\text{aq})$	increase	The Pb^{2+} from the lead nitrate will precipitate the Br^- ion from the AgBr equilibrium, forming PbBr_2 . Decreasing the $[\text{Br}^-]$ will cause the AgBr equilibrium to shift to the ion side, thus dissolving the solid.

PbCl_2	$\text{KCl}_{(\text{aq})}$	decrease	The Cl^- ion from the potassium chloride is a common ion to the Cl^- from the PbCl_2 equilibrium. The increased $[\text{Cl}^-]$ causes the PbCl_2 equilibrium to shift toward the solid side and decrease the solubility of PbCl_2 .
$\text{Be}(\text{OH})_2$	$\text{NaCl}_{(\text{aq})}$	no effect	The NaCl has no common ions to the $\text{Be}(\text{OH})_2$ and neither the Na^+ nor the Cl^- forms any precipitates with either of the ions from $\text{Be}(\text{OH})_2$.
PbCO_3	$\text{HCl}_{(\text{aq})}$	increase	The H^+ ions from the acid react with the carbonate (CO_3^{2-}) ions from the PbCO_3 equilibrium, forming carbonic acid (H_2CO_3). The carbonic acid decomposes into carbon dioxide and water. The decreased $[\text{CO}_3^{2-}]$ in the PbCO_3 equilibrium causes a shift toward the ion side, thus increasing the solubility of PbCO_3 . -or- The Cl^- from the HCl will precipitate the Pb^{2+} ion from the PbCO_3 equilibrium, forming PbCl_2 . Decreasing the $[\text{Pb}^{2+}]$ will cause the PbCO_3 equilibrium to shift to the ion side, thus dissolving the solid.

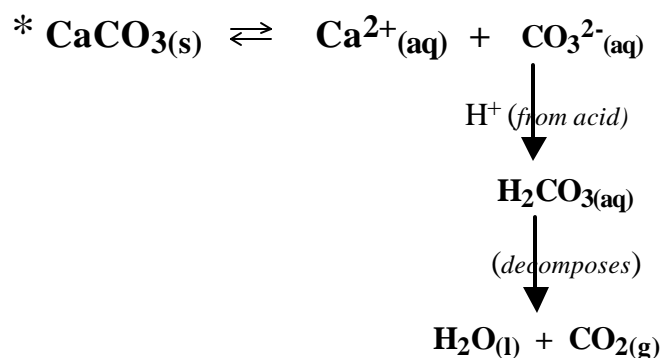
CuI	$\text{CaI}_2_{(\text{aq})}$	decrease	The I^- ion from the calcium iodide is a common ion to the I^- from the CuI equilibrium. The increased $[\text{I}^-]$ causes the CuI equilibrium to shift toward the solid side and decrease the solubility of CuI .
Ag_2CO_3	$\text{Na}_2\text{S}_{(\text{aq})}$	increase	The S^{2-} from the Na_2S will precipitate the Ag^+ ion from the Ag_2CO_3 equilibrium, forming Ag_2S . Decreasing the $[\text{Ag}^+]$ will cause the Ag_2CO_3 equilibrium to shift to the ion side, thus dissolving the solid.
$\text{Ca}_3(\text{PO}_4)_2$	$\text{K}_2\text{SO}_4_{(\text{aq})}$	increase	The SO_4^{2-} from the K_2SO_4 will precipitate the Ca^{2+} ion from the $\text{Ca}_3(\text{PO}_4)_2$ equilibrium, forming CaSO_4 . Decreasing the $[\text{Ca}^{2+}]$ will cause the $\text{Ca}_3(\text{PO}_4)_2$ equilibrium to shift to the ion side, thus dissolving the solid.

2. Given that natural rainwater is slightly acidic, explain why rain will slowly dissolve limestone ($\text{CaCO}_3(\text{s})$) over a period of time. Give a full explanation including relevant equilibrium equations.

In any water, CaCO_3 dissolves a slight amount, releasing Ca^{2+} ions and CO_3^{2-} ions into the solution and establishing an equilibrium:



Because natural rainwater is slightly acidic, it releases some H^+ ions. These H^+ ions combine with the CO_3^{2-} ions in the CaCO_3 equilibrium.



As shown above, the H^+ ions combine with CO_3^{2-} ions to form carbonic acid (H_2CO_3), which decomposes to water and carbon dioxide.

Since, in the process, the $[\text{CO}_3^{2-}]$ is decreased, the top (*) equilibrium will shift to the right, *increasing* the solubility of $\text{CaCO}_3(\text{s})$.

After a long time, a large amount of CaCO_3 (limestone) will be dissolved.

3. Silver sulphate is a white precipitate with low solubility. When a solution of ammonium sulphide ($(\text{NH}_4)_2\text{S}_{(\text{aq})}$) is added, the white precipitate slowly dissolves and a black precipitate forms on the bottom. Using equilibrium equations and clear explanations, indicate what happened here.

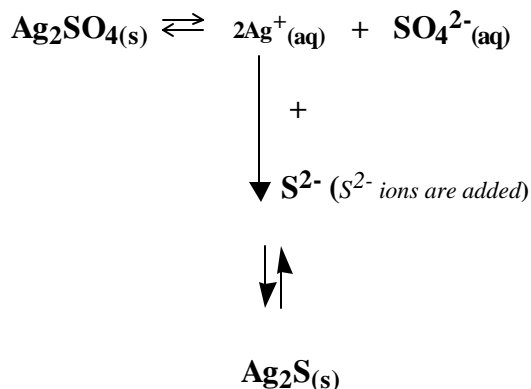
The equilibrium equation for the white precipitate (silver sulphate) is:



Ammonium sulphide ($(\text{NH}_4)_2\text{S}$) is a highly soluble compound which releases ammonium (NH_4^+) ions and sulphide (S^{2-}) ions to the solution. NH_4^+ ions are spectators in this case.

If you check the solubility table, you will see that sulphide (S^{2-}) will precipitate with silver (Ag^+) ions in the solution, forming the black precipitate silver sulphide ($\text{Ag}_2\text{S}_{(\text{s})}$).

This will cause a decrease in the $[\text{Ag}^+]$, shifting the above equilibrium to the right, causing the white $\text{Ag}_2\text{SO}_{4(\text{s})}$ to dissolve.

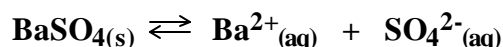


4. Name two compounds (not just ions) that can *decrease* the solubility of $\text{BaSO}_4(\text{s})$ and explain why each one of them works.

Any compound that contains either Ba^{2+} ions or SO_4^{2-} ions (other than BaSO_4) would work.

Examples could be barium nitrate ($\text{Ba}(\text{NO}_3)_2$) or sodium sulphate (Na_2SO_4).

Looking at the equilibrium equation for BaSO_4 dissolving:



Increasing either $[\text{Ba}^{2+}]$ or $[\text{SO}_4^{2-}]$ will cause the equilibrium to shift left and form more solid BaSO_4 . (Thus decreasing it's solubility.)

5. Name a substance (not just an ion) which could *increase* the solubility of $\text{BeCO}_3(\text{s})$. Explain why this substance works.

The solubility equilibrium for BeCO_3 is:



Adding any substance which will decrease either $[\text{Be}^{2+}]$ or $[\text{CO}_3^{2-}]$ will shift this equilibrium to the right and increase the solubility of BeCO_3 .

One way would be to find a solution which would form another precipitate with one of these ions.

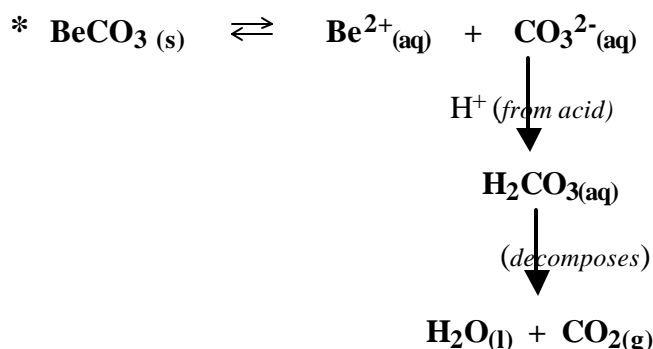
Looking at the solubility table, the following ions could form a precipitate with Be^{2+} ions: OH^- , PO_4^{3-} or SO_3^{2-} . (Notice that we can't use carbonate. I hope you can see why not!)

So 3 suitable compounds could be: NaOH , Na_3PO_4 or Na_2SO_3 .

As an example, the OH^- ions from NaOH would precipitate with the Be^{2+} ions in the BeCO_3 equilibrium, forming $\text{Be}(\text{OH})_2(\text{s})$ and decreasing the $[\text{Be}^{2+}]$. This shifts the equilibrium to the right, increasing the solubility of $\text{Be}(\text{OH})_2(\text{s})$.

There are two ways we could decrease the $[\text{CO}_3^{2-}]$:

One way would be to add an acid (eg. HCl , H_2SO_4 , HNO_3). As we have found before, adding an acid to an equilibrium with carbonate (CO_3^{2-}) ions will donate H^+ ions which will join the carbonate to form carbonic acid, which will then decompose to give carbon dioxide and water:



As shown above, the H^+ ions combine with CO_3^{2-} ions to form carbonic acid (H_2CO_3), which decomposes to water and carbon dioxide.

Since, in the process, the $[\text{CO}_3^{2-}]$ is decreased, the top (*) equilibrium will shift to the right, increasing the solubility of $\text{CaCO}_3(\text{s})$.

Another way to decrease the $[\text{CO}_3^{2-}]$ is to add a compound with an ion which would form another precipitate with it.

Checking the solubility table, some ions that precipitate with carbonate are any ions other than alkali, NH_4^+ , or H^+ ions. So there are many possibilities.

They could include Ag^+ , Pb^{2+} , Cu^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} , Mg^{2+} and many others.

Some possible compounds containing these ions could be:

AgNO_3 , $\text{Pb}(\text{NO}_3)_2$, CuNO_3 , $\text{Ca}(\text{NO}_3)_2$, $\text{Sr}(\text{NO}_3)_2$, $\text{Ba}(\text{NO}_3)_2$, $\text{Mg}(\text{NO}_3)_2$

As an example, the Ag^+ ions from AgNO_3 would precipitate with the CO_3^{2-} ions in the BeCO_3 equilibrium, forming $\text{Ag}_2\text{CO}_3(\text{s})$ and decreasing the $[\text{CO}_3^{2-}]$. This shifts the equilibrium to the right, increasing the solubility of $\text{Be}(\text{OH})_2(\text{s})$.

6. Briefly explain what is meant by the *common ion effect*.

A compound of low solubility forms two ions in a saturated solution. The addition of either of these two ions (from a compound or solution with an ion in common) will decrease the solubility of the compound with low solubility.
