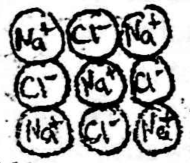
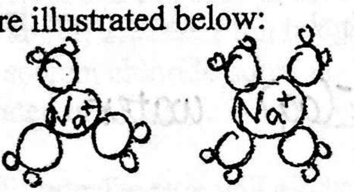
Dissolving Process, Hydration Shells, and Precipitation Name KEY

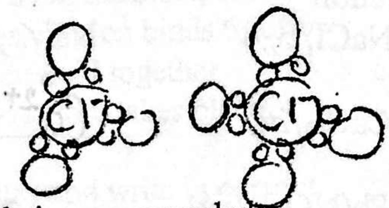
Solid sodium chloride seems to disappear when dissolved in water. However, the same amount of sodium chloride is present after dissolving as was present before. The difference is the solid is in a crystalline form, and the dissolved form has ions surrounded by water. The differences are illustrated below:



Solid NaCl (crystalline) (before dissolving)



Sodium ions surrounded by water (Na^+ from dissolved NaCl)



Chloride ions surrounded by water (Cl^- from dissolved NaCl)

Notice that the water molecules are oriented with the oxygen side facing the positively charged sodium ions, and are oriented in the opposite way when "sticking" to the negatively charged chloride ions. The opposite orientation of the water can be explained when considering the partial charges on the H_2O molecules. A water molecule has a slight negative charge on the oxygen, and slight positive charges on the hydrogen atoms (Molecules such as water with permanent partial positive and negative ends are called *polar* molecules). The water molecules surrounding and associating with an ion are collectively known as the *hydration shell*. The number of water molecules in a hydration shell depends upon the size of the ion, the charge of the ion, and the temperature.

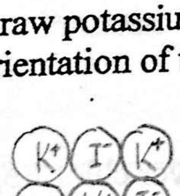
The chemical formula for water, H_2O, is typically left out of the equation representing the dissolving process. The equation for the dissolving of NaCl in water can be....

$NaCl(s) \rightarrow NaCl(aq)$ or $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$

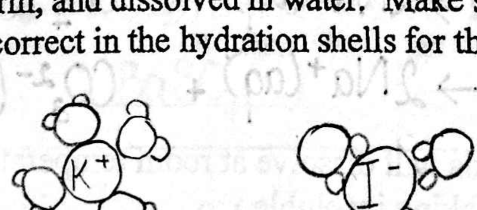
Please note that NaCl*(s)* and NaCl*(aq)* are very different! One is a pure solid in crystalline form, and the other is a mixture with individual ions surrounded by water molecules. *(s)* stands for solid; *(aq)* stands for aqueous (dissolved in water).

Practice:

Draw potassium iodide in the solid form, and dissolved in water. Make sure the orientation of the water molecules is correct in the hydration shells for the ions.



Solid KI (crystalline)



Dissolved KI, including hydration shells

Write two equations representing the dissolving of KI.

1. $KI(s) \rightarrow KI(aq)$

2. $KI(s) \rightarrow K^+(aq) + I^-(aq)$

The (aq) symbol means dissolved in water. Thus, NaCl(aq) means sodium chloride dissolved in water, or "salt water." NaCl(aq) consists of mostly water and also hydrated sodium and chloride ions. Complete the following table:

solution	components of solution		
1. $NaCl(aq)$	$Na^+(aq)$	$Cl^-(aq)$	water (mostly)
2. $CaCl_2(aq)$	$Ca^{2+}(aq)$	$Cl^-(aq)$	water
3. $Pb(NO_3)_2(aq)$	$Pb^{2+}(aq)$	$NO_3^-(aq)$	water (mostly)
4. $AgNO_3(aq)$	$Ag^+(aq)$	$NO_3^-(aq)$	Water
5. $Ba(OH)_2(aq)$	$Ba^{2+}(aq)$	$OH^-(aq)$	water
6. $Na_2SO_4(aq)$	$Na^+(aq)$	$SO_4^{2-}(aq)$	water
7. $Na_2CO_3(aq)$	$Na^+(aq)$	$CO_3^{2-}(aq)$	water

Write balanced equations for the dissolving process required to make each of the solutions represented above.

1. $NaCl(\ ) \rightarrow Na^+(aq) + Cl^-(aq)$
2. $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$
3. $Pb(NO_3)_2(s) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq)$
4. $AgNO_3(s) \rightarrow Ag^+(aq) + NO_3^-(aq)$
5. $Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$
6. $Na_2SO_4(s) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$
7. $Na_2CO_3(s) \rightarrow 2Na^+(aq) + CO_3^{2-}(aq)$

Not all ionic compounds will dissolve at room temperature. For example, silver chloride, AgCl, is, generally speaking insoluble.

AgCl can be formed if solution 4 above, containing hydrated silver ions, hydrated nitrate ions, and water, is mixed with solution 1 above, containing hydrated sodium ions, hydrated chloride ions and water. The silver ions stick to the chloride ions, forming insoluble AgCl. This insoluble substance causes a cloudy appearance in the reaction mixture, and is known as a precipitate.

When solutions containing dissolved ionic compounds (hydrated ions) are mixed, a precipitate will form if a positive ions from one solution can bind tightly to the negative ions from the other solution. In a simple case where each solution contains one species of hydrated cation and one of hydrated anion, there are two possibilities for precipitate formation (the positive ions from one can bind to the negative ions from the other, and vice versa). In the example above, the silver ion in the silver nitrate solution binds tightly to the chloride ion from the sodium chloride solution. If no pairs can hold together tightly, all ions remain hydrated in solution. No changes, thus no reaction, takes place.

Fill in the following table. Eventually, you will circle the precipitates and write in correct states of matter as shown in the example, but do not be concerned with this yet. Write both potential precipitates with the solid state of matter *(s)*.

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1. Solution A: $AgNO_3$ *(aq)* — hydrated ions in A: Ag^+ *(aq)* NO_3^- *(aq)* — Solution B: Na_2CO_3 *(aq)* — hydrated ions in B: Na^+ *(aq)* CO_3^{2-} *(aq)*

Potential Precipitates when Sol. A ($AgNO_3$*(aq)*) is mixed with Sol. B (NaCl*(aq)*).

Ag_2CO_3 *(s)* $NaNO_3$*(s)*

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2. Solution A: Na_2CO_3*(aq)* — hydrated ions in A: $Na^+(aq)$ $CO_3^{2-}(aq)$ — Solution B: $CaCl_2$*(aq)* — hydrated ions in B: $Ca^{2+}(aq)$ $Cl^-(aq)$

Potential Precipitates when Sol. A (Na_2CO_3*(aq)*) is mixed with Sol. B ($CaCl_2$*(aq)*).

$NaCl(aq)$ $CaCO_3(s)$

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3. Solution A: Na_2SO_4*(aq)* — hydrated ions in A: $Na^+(aq)$ $SO_4^{2-}(aq)$ — Solution B: $Ba(OH)_2$*(aq)* — hydrated ions in B: $Ba^{2+}(aq)$ $OH^-(aq)$

Potential Precipitates when Sol. A (Na_2SO_4*(aq)*) is mixed with Sol. B ($Ba(OH)_2$*(aq)*).

$NaOH(aq)$ $BaSO_4(s)$

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4. Solution A: $MgCl_2$*(aq)* — hydrated ions in A: $Mg^{2+}(aq)$ $Cl^-(aq)$ — Solution B: $Al(NO_3)_3$*(aq)* — hydrated ions in B: $Al^{3+}(aq)$ $NO_3^-(aq)$

Potential Precipitates when Sol. A ($MgCl_2$*(aq)*) is mixed with Sol. B ($Al(NO_3)_3$ *(aq)*).

$AlCl_3(aq)$ $Mg(NO_3)_2(aq)$

5. Solution A hydrated ions in A Solution B hydrated ions in B

$MgSO_4(aq)$ $Mg^{2+}(aq)$ $SO_4^{2-}(aq)$ $NaOH(aq)$ $Na^+(aq)$ $OH^-(aq)$

Potential Precipitates when Sol. A ($MgSO_4(aq)$) is mixed with Sol. B ($NaOH(aq)$).

$Mg(OH)_2(s)$ $Na_2SO_4(aq)$

The identity of substances that are insoluble, and that thus form precipitates if the ions from which they are composed are supplied when two solutions are mixed, has been determined from experiments. The results of these experiments are summarized in what are called "solubility rules." The solubility rules can be used to determine whether or not an ionic compound is soluble or insoluble.

A complete discussion of solubility rules will not be included here. However, a few rules will be shared to allow you to determine the identity of precipitates and soluble substances from those in examples 1-5 above.

1. All nitrates are soluble (in other words, nitrate will not form a precipitate with anything). Return to 1-5 above, and cross out the *(s)* state of matter (representing solid and thus a precipitate) for any "potential precipitates" containing nitrate. Switch the *(s)* to *(aq)*, meaning the substance will remain dissolved.
2. All chlorides except those of silver, mercury I, and lead (remember as small, medium, and large!) are soluble. Adjust "potential precipitates" as appropriate.
3. All sodium compounds are soluble. Adjust "potential precipitates" as appropriate.
4. All hydroxides except those of the alkali metals and calcium, strontium, and barium are insoluble. Adjust "potential precipitates" as appropriate.

For examples 1-5 above, any of the substances you have not marked as soluble (*(aq)*) are insoluble, and thus will be precipitates if the solutions shown are mixed. CIRCLE ALL PRECIPITATES, LEAVING THE *(s)* STATE OF MATTER INTACT. Hint—in one of the combinations, no reaction takes place, thus no precipitate is formed.

A common equation used to show a precipitation reaction includes just the formulas for the hydrated ions involved in the reaction and the formula for the precipitate formed. This type of an equation is called a net ionic equation. Write balanced net ionic equations for the mixing of solutions as represented in 1-5 on the previous page.

1. $2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$
2. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$
3. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
4. No reaction
5. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$