Qualitative Solubility Guidelines for Ionic Solids

Ion contained in the compound	**Solubility**	**Exceptions**
Group 1A [Li^+, Na^+, K^+, etc.]	soluble	
ammonium [NH_4^+]	soluble	
acetate [$CH_3CO_2^-$]	soluble	
nitrate [NO_3^-]	soluble	
chloride [Cl^-], bromide [Br^-], iodide [I^-]	soluble	Ag^+, Pb^{2+}, Hg_2^{2+}
sulfate [SO_4^{-2}]	soluble	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
carbonate [CO_3^{2-}]	insoluble[1]	(see above)[2]
phosphate [PO_4^{3-}]	insoluble	(see above)
hydroxide [OH^-]	insoluble	Ca^{2+}, Sr^{2+}, Ba^{2+}, (see above)

NOTES:

1 All ionic compounds (even the least soluble ones, dissolve to some slight extent in water. So ... the “insoluble” classification really means ionic compounds that have a very limited solubility.

2 By “(see above)” I mean the following (as an example): If **all** ionic solid compounds containing Li^+ are soluble, then lithium carbonate [Li_2CO_3] must be soluble - even though most carbonates are insoluble.

Ionic Equation and Net Ionic Equation

Consider the reaction we used in the Sea Water Experiment when you mixed an aqueous solution of silver nitrate [$AgNO_3$] with the sea water - which contained dissolved sodium chloride [NaCl]. One can represent this reaction in several ways. First a "**molecular equation**" involving the formula unit for the chemicals used (and made):

$$AgNO_3 \quad + \quad NaCl \quad \text{---------->} \quad AgCl \quad + \quad NaNO_3$$

This is a correctly balanced double replacement reaction, BUT it does not represent the actual process taking place. Both the $AgNO_3$ and the NaCl were *dissolved in water* (that is, aqueous solutions). When an ionic solid "dissolves" in water, the ions actually become separated and surrounded by water molecules. So **solid** sodium chloride is $NaCl_{(s)}$, where the "s" represents "solid." [liquid and gas are represented with (l) and (g)] But upon mixing solid sodium chloride with water the solid dissolves and becomes discrete Na^+ ions and Cl^- ions and we represent these dissolved ions as $Na^+_{(aq)}$ and $Cl^-_{(aq)}$, where the "aq" indicates "aqueous" - the ion dissolved in water. All dissolved ionic compounds are actually present in the form of individual *separate* ions. So the reactants really are:

$$Ag^+_{(aq)} \; + \; NO_3^-{}_{(aq)} \quad \text{and} \quad Na^+_{(aq)} \; + \; Cl^-_{(aq)} \quad \text{---------->}$$

Let us now look at the products of this double replacement reaction: the AgCl and $NaNO_3$.

The solubility guidelines show that:
1) although most chlorides would be soluble, *one exception is silver chloride* [AgCl]. This means that the AgCl will not dissolve, and therefore will be a solid! [You may remember that in your experiment when you mixed the two solutions a white solid precipitate formed.]
2) all compounds containing the Na^+ ion are soluble - so this means that the sodium nitrate [$NaNO_3$] is actually present in the form of individual separate ions.

So now our original molecular equation can be written [more correctly] as an "**ionic equation**" involving the ions and solid that are actually involved:

$$Ag^+_{(aq)} + NO_3^-{}_{(aq)} \; + \; Na^+_{(aq)} + Cl^-_{(aq)} \quad \text{---------->} \quad AgCl_{(s)} \; + \; Na^+_{(aq)} + NO_3^-{}_{(aq)}$$

But notice that the $Na^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ ions appear unaltered on both sides of the equation. These ions are called *spectator ions* because they do not take part in the reaction (they simply "watch" as the other ions form a solid). We can then exclude such ions - and if we rewrite the ionic equation without the spectator ions, we get an equation that is called the "**net ionic equation**" which involves only those ions that actually participate in the reaction:

$$Ag^+_{(aq)} \; + \; Cl^-_{(aq)} \quad \text{---------->} \quad AgCl_{(s)}$$

Leaving out the spectator ions does not imply that they are totally unimportant. Indeed the Ag^+ ion cannot exist alone in solution - a negative ion (anion) must be present as well! But net ionic equations are widely used because they illustrate the similarity between large numbers of possible reactions. The net ionic equation shows that more than one set of reactants can give the same reaction. That is, mixing *any* aqueous solution containing the Ag^+ ion with *any* other aqueous solution containing the Cl^- ion will always produce the solid AgCl product!