

Cations \downarrow \rightarrow all anions
 CO_3^{2-} Cl^- OH^- NO_3^- PO_4^{3-} SO_4^{2-}
 Na_2CO_3 NaCl NaOH NaNO_3 Na_3PO_4 Na_2SO_4

Al^{3+} AlCl_3	PPT		PPT		PPT	
NH_4^+ NH_4Cl						
Ca^{2+} CaCl_2	PPT		PPT		PPT	PPT
Cu^{2+} CuSO_4	PPT		PPT		PPT	
Fe^{3+} FeCl_3	PPT		PPT		PPT	
Pb^{2+} $\text{Pb}(\text{NO}_3)_2$	PPT	PPT	PPT		PPT	PPT
K^+ KI						
Ag^+ AgNO_3	PPT	PPT	PPT		PPT	

PPT = Precipitate = Solid Formation

No notation = No reaction observed

Observing 48 of the possible 91 pairings \Rightarrow we are only looking at a subset.

Experiment uses a total of 14 chemicals

Grid

14 →

14 ↓

$$14 \times 14 = \frac{196 \text{ squares} - 14}{2}$$

91 unique pairings

	A	B	C	D	...
A	X	BA	CA		
B	AB	X			
C	AC		X		
D				X	
⋮					

X						
	X					
		X				
			X			
				X		
					X	
						X

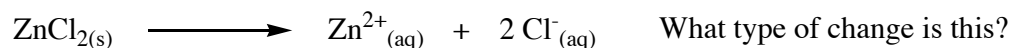
Predicting the Products of Precipitation Reactions: Solubility Rules

Goals

- Observe and record precipitation reactions.
- Derive general solubility rules from the experimental data.
- Describe precipitation reactions by writing net ionic equations.
- Understand the relationship between solubility and precipitation reactions.

Introduction

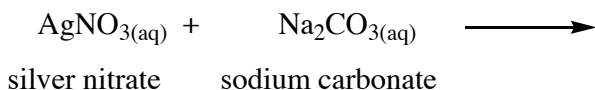
Ionic compounds in solution consist of free ions surrounded by water. In the case of zinc chloride being dissolved in water to form an aqueous solution, the dissolution of the ionic compound can be depicted as follows:



These free aqueous ions can react with other ions. Mixing of dissolved ionic compounds can lead to precipitation reactions, a double displacement reaction. The **precipitate** is a **solid product**, a new ionic compound that is different from the reactants in both composition and solubility. Solubility is defined as the amount of substance (solute) that dissolves in a given amount of solvent. Solubility is a physical property that can be useful in predicting whether the mixing of aqueous ionic compounds will lead to a precipitation reaction. The mixing of a variety of combinations leads to the formulation of general rules of solubility. Some examples of these rules include "All sodium salts are soluble in water" or "The mixing of two ionic compounds that contain a common ion will not lead to a precipitate".

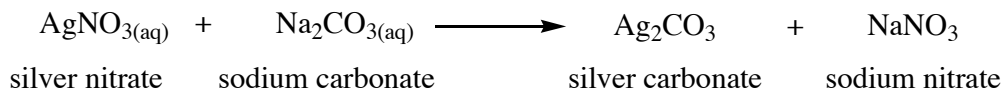
Let's look at an example to see how these solubility rules can help us. As part of the lab, aqueous solutions of silver nitrate and sodium carbonate are mixed. A foggy white precipitate is formed. To write the chemical equation:

1. *Identify the reactants and write their correct formulas:*

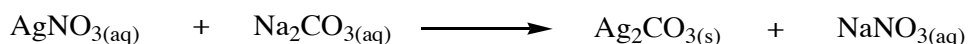


2. *"Swap" the cations of the reactants to form the products, two new ionic compounds: the products will be sodium nitrate and silver carbonate.*

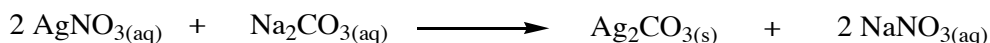
3. Write the **correct** formulas for the products after the arrow. Use the names of the products as an aid to get the formulas correct. Each product must have a net charge of zero. In other words the sum of the positive charges must equal the sum of the negative charges within each product.



4. Use solubility rules to determine which product is the precipitate: "All sodium salts are soluble"; therefore silver carbonate must be the foggy, white precipitate.

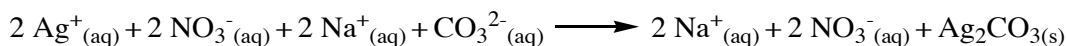


5. **LAST** but not least, **balance the equation** using whole number coefficients:

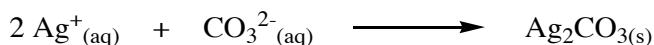


How do you convert a chemical equation into a correct net-ionic equation?

1. Rewrite the correct chemical equation **but** this time write any aqueous (not solid) ionic compound as free cations, with proper charge and aqueous symbol, and anions, with proper charge and aqueous symbol. This form is referred to as the total ionic equation. Remember that coefficients in front of a species apply to both cation and anion. Only subscripts on monatomic ions and subscripts outside of parenthesis surrounding polyatomic ions multiply against coefficients. Subscripts that are part of polyatomic ions **do not** multiply against coefficients. Look below, which subscripts changed coefficients when the aqueous compounds were written as free ions?



2. Once you are sure you have the correct total ionic equation, look at the equation again. It should be balanced and you will be able to see what ions actually changed. Any species that is exactly the same on both sides is considered to be a spectator ion. *To write the net ionic equation, eliminate the spectators and only write the species that changed:*



Why so many types of equations? Each equation is useful for different reasons. The net ionic is preferred in double displacement reactions because it focuses solely on the product.

In this lab you will mix a variety of solutions. By using sodium salts for at least one of the reactants you will be able to identify which product is the precipitate. To assist you in identifying the solid product, you will write chemical equations for all observable reactions. These chemical reactions will be translated into net ionic equations. After examining the chemical formulas of the precipitates, you will "conclude" by summarizing your results as a set of solubility rules.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.

Wear safety glasses at all times.

Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate (AgNO_3) will stain the skin.

Use microburets to dispense reagents in such a way that they do not make contact with other drops or the reaction surface.

Return any contaminated microburets to your instructor.

Materials:

Reagent Central chemicals include aqueous solutions of the compounds listed on your experimental page.

Equipment: Clean, dry transfer pipet for mixing. Labtop reaction surface

Experimental Procedure

1. Insert your experimental page under your reaction surface. Place 1 drop of each solution in the squares on your experimental page. Record what happened after mixing with the air from a clean, dry pipet. ("NVR" can indicate no visible reaction.) Please include adjectives that describe both color and texture being as specific as possible (not all white precipitates look the same).
2. After all the reactions are completed and all observations recorded, take one last look at your surface. Have any of the squares changed over time? Record any noticeable changes (there won't be many). Clean your surface by absorbing the contents onto a paper towel. Rinse the reaction surface with a damp paper towel and dry it. Dispose of paper towels in waste bin. Clean your area. Wash your hands thoroughly with soap and water.
3. Answer the questions, correctly writing both chemical and net ionic equations.
4. Draw general conclusions about the cations and anions in your experiment by formulating your own solubility rules.
5. Apply your rules to unknown combinations.

Reaction Guide: Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

Sol'ns	Na_2CO_3	NaCl	NaOH	NaNO_3	Na_3PO_4	Na_2SO_4
AlCl_3	✗	✗	✗	✗	✗	✗
NH_4Cl	✗	✗	✗	✗	✗	✗
CaCl_2	✗	✗	✗	✗	✗	✗
CuSO_4	✗	✗	✗	✗	✗	✗
FeCl_3	✗	✗	✗	✗	✗	✗
$\text{Pb}(\text{NO}_3)_2$	✗	✗	✗	✗	✗	✗
KI	✗	✗	✗	✗	✗	✗
AgNO_3	✗	✗	✗	✗	✗	✗

Data Organization Suggestions

There are a number of important facets to this experiment. Learning how to predict the products of precipitation reactions, learning to write proper balanced chemical equations and net ionic equations, and observing solubility trends. To make all of these goals easier use the reaction guide on the previous page as a template for a data table. Use an entire page of your notebook for the table. Use the ruler provided in your chemistry kit to make clean lines. When a precipitation reaction is observed, indicate the color and use one or two adjectives to describe the product (milky, cloudy, marbled, granular...). Where no visible reaction is observed you can simply write "NVR" for "no visible reaction" (remember to define abbreviations somewhere on the page).

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

1. Look across the rows, what cations are always soluble (never form precipitates)?
2. Look down columns, what anions are always soluble (never form precipitates)?
3. Look down columns, what anions are usually soluble (seldom form precipitates)? Which cation(s) will form precipitates with these anions (your looking for exceptions to the general trend)?
4. Look down columns, what anions are usually insoluble (usually form precipitates)? Which cation(s) will not form precipitates with these anions (your looking for exceptions to the general trend)?
5. Given that you have answered questions two through five correctly, would you expect to see a precipitate form in the reaction of aqueous magnesium chloride with aqueous potassium phosphate? Write out the complete chemical equation for this reaction. Defend your answer by stating your reasoning.
6. Would you expect to see a precipitate form in the reaction of aqueous sodium chloride with aqueous ammonium hydroxide? Write out the complete chemical equation for this reaction. Defend your answer by stating your reasoning.
7. For all reactions in which there was an observable change (precipitate formation) write out a complete correct chemical equation to describe the reaction, followed by the net ionic equation. Refer back to the introduction if you need assistance. There should be 22 reactions in all (go back and look at the $\text{CaCl}_{2(\text{aq})}$ with $\text{Na}_2\text{SO}_{4(\text{aq})}$). Allow enough space between equations in your notebook so that your work is easily readable.

Data Analysis

- ① Look across the rows, what cations never form precipitates (PPT) \Rightarrow Empty rows

potassium K^+ \leftarrow
ammonium NH_4^+
sodium Na^+ \leftarrow Group IA is important

Group IA Cations never form precipitates

\Rightarrow Group IA & NH_4^+ are highly soluble cations & never form precipitates

- ② Look down the columns, which anions never form precipitates? \Rightarrow empty columns.

Nitrate NO_3^-

Nitrate is an always soluble anion, it never forms precipitates

③ Look down a column, which anions are usually soluble, form a few precipitates?

{ Chloride Cl^- are usually soluble
Sulfate SO_4^{2-}

They each have 2 exceptions

⇒ Chloride Cl^- will form a ppt with Pb^{2+} & Ag^+

⇒ Sulfate SO_4^{2-} will form a ppt with Ca^{2+} & Pb^{2+}

④ Looking down a column, what anions usually form precipitates? And what are their exceptions?

Carbonate CO_3^{2-}
Phosphate PO_4^{3-}
Hydroxide OH^- } usually form ppt
⇒ usually insoluble

Important exceptions are NH_4^+ & K^+
that are soluble with these 3 anions.

Solubility Table

	<u>Rule</u>	<u>Ions</u>	<u>Exceptions</u>
①	Always Soluble Cations	Group 1A (Li^+ , Na^+ , K^+) & NH_4^+	None
②	Always Soluble Anions	NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$	None
③	Usually Soluble Anions	Cl^- , Br^- , I^-	Pb^{2+} , Ag^+ , Hg^{2+}
	Usually Soluble Anions	F^- , SO_4^{2-}	Pb^{2+} , Ca^{2+} , Ba^{2+}
④	Usually Insoluble Anions	OH^- , CO_3^{2-} , PO_4^{3-}	Group 1A & NH_4^+

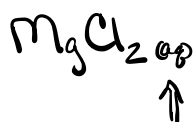
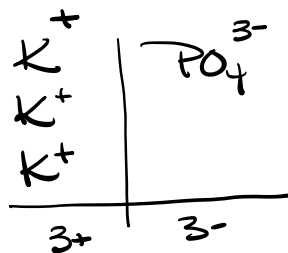
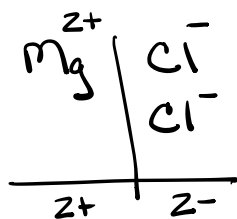
PPT

Insoluble

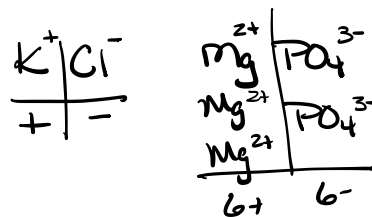
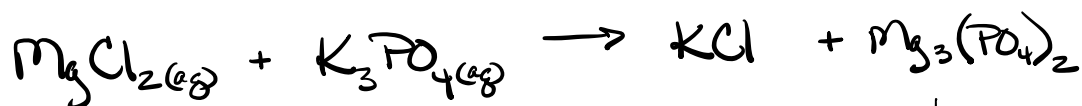
No Reaction

Soluble

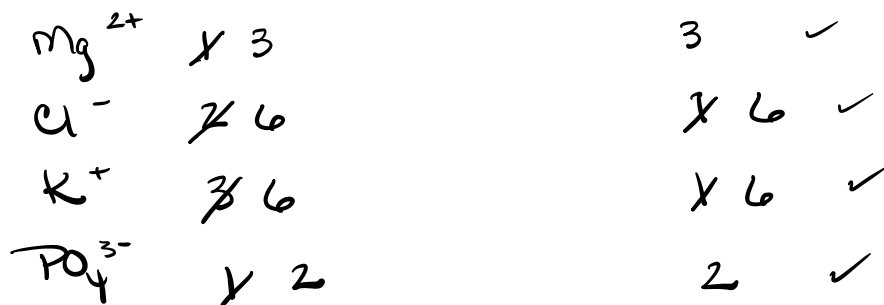
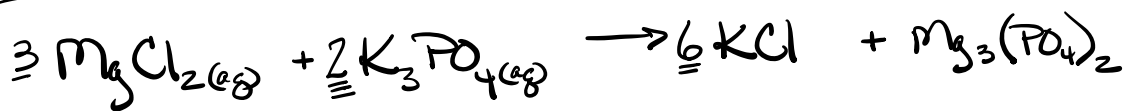
⑤ magnesium chloride + Potassium Phosphate → ppt?



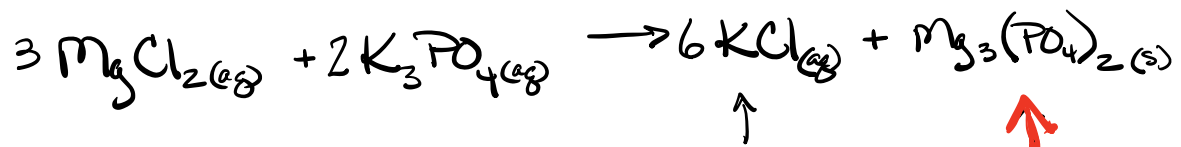
↑
Reactant will
be aqueous



Balance



Solubility



(aq) = Soluble in solution

(s) = solid = Insoluble

↑
PPT

KCl(aq)

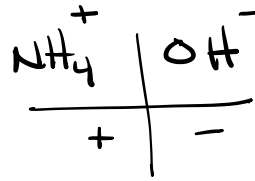
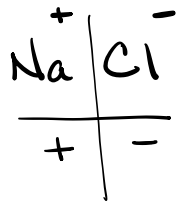
Rule K^+ is Group IA cation is always soluble.

or Cl^- is usually soluble & K^+ is not an exception to chlorides rule.

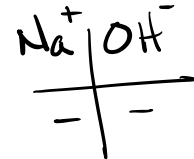
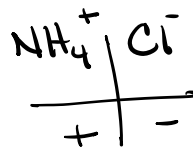
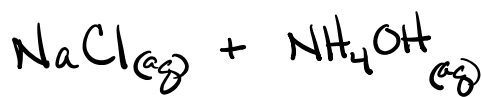
$\text{Mg}_3(\text{PO}_4)_2(\text{s})$

Rule: PO_4^{3-} is usually insoluble. Mg^{2+} is not an exception to the rule.

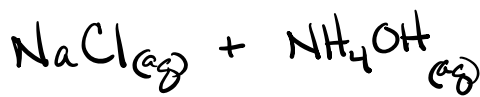
(c) Sodium Chloride + ammonium hydroxide



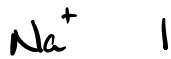
Products



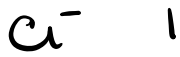
Balance



+

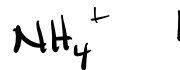


|



|

✓



|



|

Solubility of products



Both (aq) \Rightarrow no ppt
 \Rightarrow no Reaction

NaOH_(aq) = Soluble

Na⁺ Group IA \Rightarrow always Soluble

OH⁻ Usually insoluble anion \Rightarrow exception is group IA
or ammonium
 \Rightarrow Na⁺ Group IA

NH₄Cl_(aq) = Soluble

NH₄⁺ always Soluble Cation

Cl⁻ usually Soluble anion \Rightarrow NH₄⁺ is not
an exception

Solubility Table

	<u>Rule</u>	<u>Ions</u>	<u>Exceptions</u>
①	Always Soluble Cations	Group 1A (Li^+ , Na^+ , K^+) & NH_4^+	None
②	Always Soluble Anions	NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$	None
③	Usually Soluble Anions	Cl^- , Br^- , I^-	Pb^{2+} , Ag^+ , Hg^{2+}
	Usually Soluble Anions	F^- , SO_4^{2-}	Pb^{2+} , Ca^{2+} , Ba^{2+}
④	Usually Insoluble Anions	OH^- , CO_3^{2-} , PO_4^{3-}	Group 1A & NH_4^+

PPT

Insoluble

No Reaction

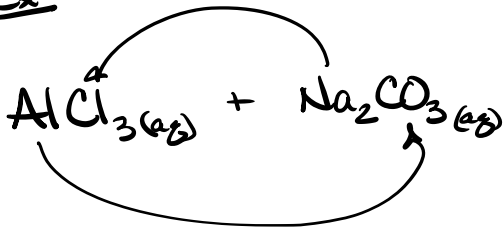
Soluble

Predict the solubility of the following
ionic compounds. aq = soluble or s = insoluble

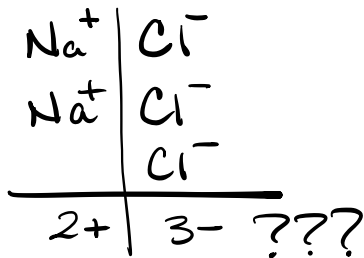
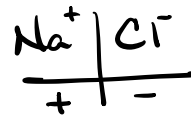
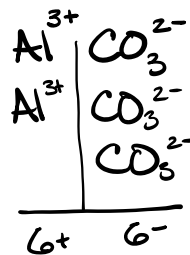
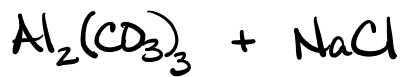
	<u>Solubility</u>	<u>Rule</u>
Na_2SO_4	Soluble (aq)	Na^+ is always soluble cation.
$FeCl_3$	Soluble (aq)	Cl^- usually soluble Fe^{3+} is not an exception
$Al_2(CO_3)_3$	Insoluble (s)	CO_3^{2-} is usually insoluble & Al^{3+} is <u>not</u> an exception
$Pb(NO_3)_2$	Soluble (aq)	NO_3^- always soluble anion
$AgBr$	Insoluble (s)	Br^- is usually soluble but Ag^+ is an exception

7) For 10 reactions with ppt write out the
 Chemical } all 3 equations
 Ionic }
 Net ionic }

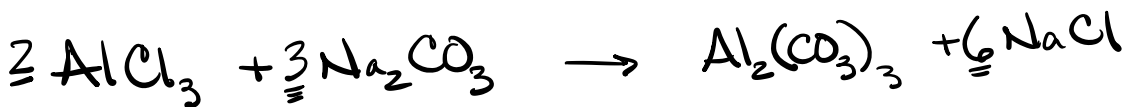
Ex



① products → properly balanced



step 2 balance the equation (coefficients)



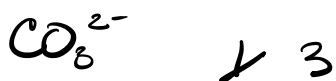
2



$\times 6$

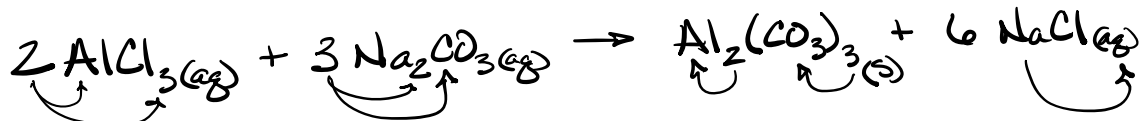


$\times 6$



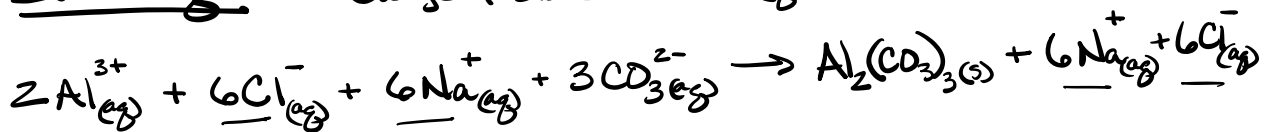
3

Step 3 Solubility of products

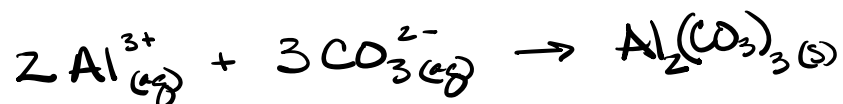


This is the full chemical Eq.

Ionic Eq → charge & state on each (aq) ion

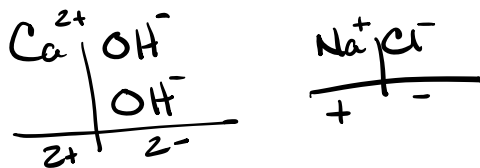
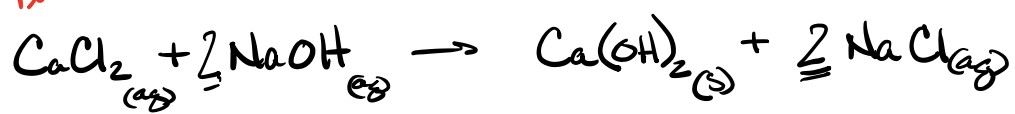


Net Ionic Eq

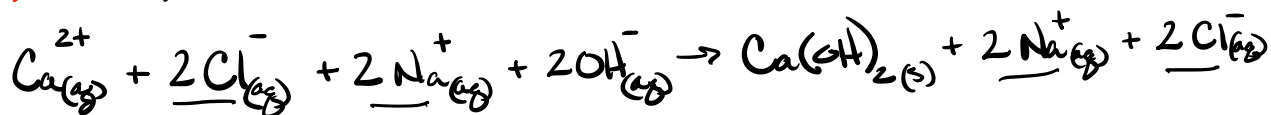


Ex

Balanced Chemical



Ionic



net ionic



Assignment \Rightarrow 10 more with PPT!