

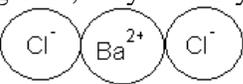
CHM 130LL: Double Replacement Reactions

One of the main purposes of chemistry is to transform one set of chemicals (the reactants) into another set of chemicals (the products) via a chemical reaction:



Many of these reactions occur in an **aqueous** environment (i.e., in a solution where ions and compounds are dissolved in water). When a reactant or product has the physical state “(aq)”, it means the **substance is dissolved in water**. When an ionic compound is in the aqueous state, it is dissociated into ions so it exists as individual ions in solution; for example, $\text{NaCl}(aq)$ exists as individual Na^+ and Cl^- ions moving around in water. $\text{BaCl}_2(aq)$ exists as three ions moving around independently in water: one Ba^{2+} and two separate Cl^- ions.

Note: Do not confuse an element’s naturally occurring state with how it appears in a compound. For example chlorine occurs as an element as a diatomic gas, $\text{Cl}_2(g)$, but when chlorine is in an ionic compound it is chloride ion, Cl^- . In barium chloride (BaCl_2) we do not have chlorine, but two chloride ions. Because the chloride ions are negative, they actually repel each other and are not bonded with

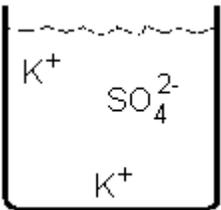
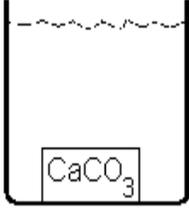
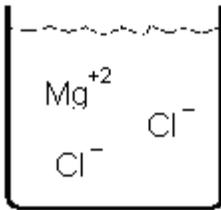
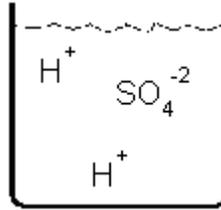
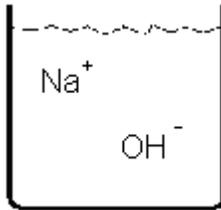
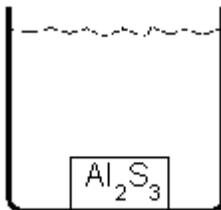
each other. Barium chloride exists as  physically but we write the formula as BaCl_2 .

Now consider putting ionic compounds into water. If the ionic compounds are **soluble**, they dissociate in water. If not they are considered **insoluble**. To determine if an ionic compound is soluble—i.e., will dissolve and dissociate—in water, we use the Solubility Rules found on your periodic table.

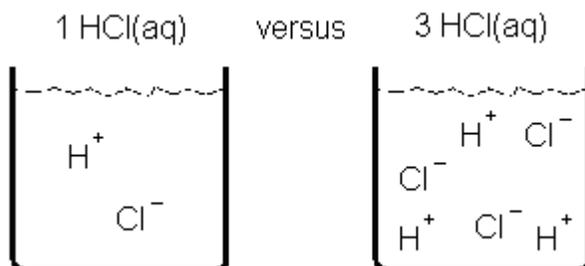
Compounds containing ions included in the left column of the Rules are generally **soluble**, with a few exceptions. Compounds containing ions included in the right column are generally **insoluble**, again with a few exceptions. The Solubility Rules allow us to determine which compounds are soluble, and thus in water are represented as aqueous: e.g., $\text{KI}(aq)$, $\text{BaCl}_2(aq)$, $\text{NaOH}(aq)$, etc. The Solubility Rules also tell us which compounds are **insoluble**—i.e., do not dissolve much in water and remain mostly as solids: e.g. $\text{BaSO}_4(s)$, $\text{AgCl}(s)$, $\text{Al}_2(\text{CO}_3)_3(s)$, etc.

Drawing Ionic Solids in Water

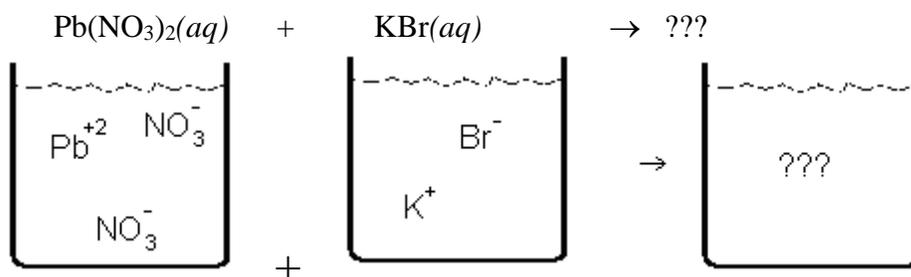
We can represent ionic solids in water as ions if they are soluble and as solids if they are insoluble. Refer to the solubility rules when looking at the examples below. Notice the solvent water molecules are not shown. **In this lab we will draw insoluble compounds completely as solids even though a few ions do form.**

Potassium sulfate formula: K_2SO_4 	Calcium carbonate formula: $CaCO_3$ 	Magnesium chloride formula: $MgCl_2$ 
Sulfuric acid formula: H_2SO_4 	Sodium hydroxide formula: $NaOH$ 	Aluminum sulfide formula: Al_2S_3 

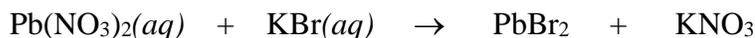
When we balance chemical reactions we use coefficients so that the “Law of conservation of matter” is not violated: Matter, or atoms, cannot be created nor destroyed. Look at the following example to see how we draw one $HCl(aq)$ versus three $HCl(aq)$. The coefficient just tells us how many of the molecules are present, not that they are bonded together somehow.



In a **precipitation reaction** (one type of double-replacement reaction), two soluble ionic compounds in aqueous solution are mixed and result in an insoluble solid compound called a **precipitate**. For example, consider the reaction between aqueous lead(II) nitrate with aqueous potassium bromide, as shown below:

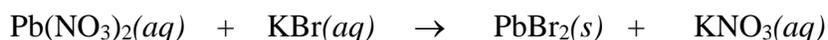


When we pour the two solutions together we have four ions in solution: $\text{Pb}^{2+}(\text{aq})$, $\text{NO}_3^-(\text{aq})$, $\text{K}^+(\text{aq})$ and $\text{Br}^-(\text{aq})$. (Remember (aq) means dissolved in water) Consider the possible combinations of ions hitting in the solution. Lead(II) and nitrate ions can hit each other but will not bond because otherwise they would not have dissociated in the first place. The same goes for potassium and bromide ions. Lead(II) and potassium ions can hit each other, but positives repel. Similarly, the negatively charged nitrate and bromide ions will not bond because negatives repel. So the only possible combinations that could form products are lead(II) ion, Pb^{2+} , with bromide ion, Br^- , to form lead(II) bromide, PbBr_2 , and potassium ion, K^+ with nitrate ion, NO_3^- , to form potassium nitrate, KNO_3 , as shown below:

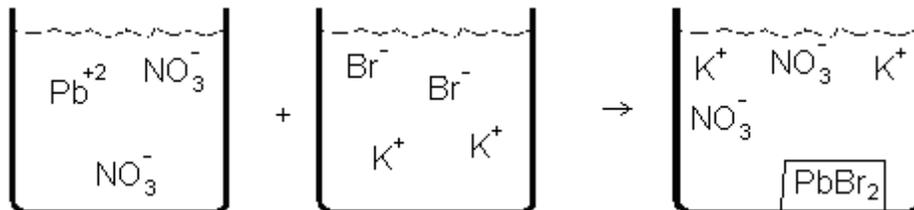
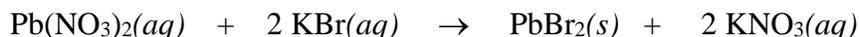


Note that the chemical formulas for the products formed are based on their charges, not how they appear on the reactant side of the chemical equation.

Next, we refer to the Solubility Rules to determine if PbBr_2 and KNO_3 are soluble or insoluble. Based on Rule #3, PbBr_2 is insoluble, so we indicate it as $\text{PbBr}_2(\text{s})$. Based on Rule #1, any compound with K^+ is soluble, so we represent potassium nitrate as $\text{KNO}_3(\text{aq})$. Thus, the complete *unbalanced* equation is:



Because a precipitate formed— $\text{PbBr}_2(\text{s})$ —then a reaction occurred. We balance the equation by making sure the number of atoms and/or polyatomic ions are equal on both sides of the equation. The balanced equation is:



Note that only the lead(II) ions and bromide ions react to form the lead(II) bromide precipitate. The potassium and nitrate ions exist as individual ions before and after the reaction; because they themselves did not react, they are called “*spectator ions*,” as if they simply watched the other ions reacting.

If you carry out this reaction in a laboratory, you will observe that the lead(II) nitrate and potassium bromide solutions are both clear and colorless. When you mix them, the resulting solution is cloudy, indicating a solid is present and small particles of the solid are suspended in the solution. When enough solid has formed, it will begin to settle at the bottom of the beaker. Thus, two clear solutions becoming cloudy when combined is often interpreted as experimental evidence of a precipitate forming.

In general, **acids** are compounds that **produce hydrogen ions (H^+)**, which are **protons**, when dissolved in water. The chemical formulas for acids are most often given with the H’s at the beginning, so acids are usually easy to recognize. A few common acids are hydrochloric acid, $\text{HCl}(\text{aq})$, nitric acid, $\text{HNO}_3(\text{aq})$, and sulfuric acid, $\text{H}_2\text{SO}_4(\text{aq})$.

Note that the physical state aqueous, (aq) , must be included to distinguish the acid from other forms of a substance. For example, the formula “HCl” can be used for hydrogen chloride gas, $HCl(g)$, so to indicate hydrochloric acid, one must specify $HCl(aq)$.

Bases are compounds that *produce hydroxide ions* (OH^-) when dissolved in water. A few common bases are sodium hydroxide, NaOH, potassium hydroxide, KOH, calcium hydroxide, $Ca(OH)_2$, and barium hydroxide, $Ba(OH)_2$.

Acid-Base Neutralization Reactions

In an acid-base neutralization reaction, (another type of double-replacement reaction), an acid can react with a metal hydroxide base to produce water and a salt:

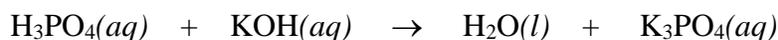


The *hydrogen ion* (H^+) from the acid reacts with the *hydroxide ion* (OH^-) from the base to form the water. The **salt** is formed by combining the *cation from the base* and the *anion from the acid*. Because water is formed in these types of acid base reactions, acids will always react with metal hydroxide bases, regardless of whether the salt produced is soluble or insoluble.

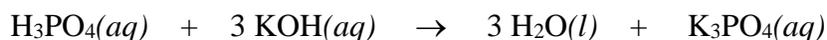
Let us consider the reaction between phosphoric acid and potassium hydroxide.



The three hydrogen ions (H^+) from the phosphoric acid, $H_3PO_4(aq)$, combine with the hydroxide ions (OH^-) from the $KOH(aq)$ to produce water. The remaining ions, the potassium ion, K^+ , from the $KOH(aq)$ and the phosphate ion, PO_4^{3-} , from the $H_3PO_4(aq)$ combine to form the salt, K_3PO_4 . Based on the Solubility Rules (#1), K_3PO_4 is soluble, so we represent it as $K_3PO_4(aq)$. Thus, the complete *unbalanced* chemical equation is shown as:



We finish the chemical equation by balancing the number of atoms and/or polyatomic ions on both sides of the equation. Thus, the balanced chemical equation is:



The progress of acid-base neutralization reactions is difficult to monitor in the lab because acid solutions and base solutions are often clear and colorless, as are the resulting water and soluble salt solutions. However, these reactions give off heat. Reactions that release heat are **exothermic** and feel warm to the touch. Thus, acid-base neutralization reactions can be identified if any heat energy change is observed.

Summary: Three steps for writing balanced equations:

1. Write the product formulas based on charges
2. Identify the physical states of the products
3. Balance the equation

In this experiment, you will mix solutions containing various soluble compounds and note observations to determine if a chemical reaction has occurred. Two examples of visible evidence of a chemical reaction are as follows:

- precipitate formation (the solution becomes cloudy)
- acid-base reaction (heat is given off)

NOTE: All waste for this experiment should go into the labeled waste containers in the hood.
Do not dispose of any solutions or solids down the drain.

Procedure

Part A: Mixing Known Solutions

1. Rinse your well-plate thoroughly with plenty of tap water, then rinse it briefly with deionized water and dry completely with cotton swabs to prevent any errors from contamination.
2. For each of the following combinations, place 5 drops of the first solution into a well. Now add 5 drops of the second solution. Shake the well plate gently for 10 seconds to mix the solutions and observe. Record your observations. If a reaction occurs, write the balanced reaction on your report sheet. If nothing happens, write NR (no reaction) on your report sheet.
 - a. $\text{Na}_2\text{CO}_3(\text{aq})$ and $\text{CaCl}_2(\text{aq})$ shake this one gently for 45 seconds then observe
 - b. $\text{Na}_2\text{SO}_4(\text{aq})$ and $\text{CaCl}_2(\text{aq})$
 - c. $\text{KI}(\text{aq})$ and $\text{Pb}(\text{NO}_3)_2(\text{aq})$
 - d. $\text{Na}_2\text{SO}_4(\text{aq})$ and $\text{KI}(\text{aq})$
 - e. $\text{CaCl}_2(\text{aq})$ and $\text{Pb}(\text{NO}_3)_2(\text{aq})$
 - f. $\text{Na}_2\text{SO}_4(\text{aq})$ and $\text{Pb}(\text{NO}_3)_2(\text{aq})$
 - g. $\text{CaCl}_2(\text{aq})$ and $\text{KI}(\text{aq})$
 - h. $\text{Na}_2\text{CO}_3(\text{aq})$ and $\text{Pb}(\text{NO}_3)_2(\text{aq})$
3. Dump your well plate into the waste container in the hood. Then rinse with tap water several times and replace in the tub.

Part B: Molecular Level Drawings

Refer to the examples in this lab handout and draw the molecular level drawings for the selected reactions from part A.

Part C: Questions

Answer the follow up questions

Wash your hands and your bench before you leave.

CHM 130LL:

Double Replacement Reactions

Lab Report

Name: _____

Partners: _____

Section Number: _____

Part A: Double Replacement Reactions

For the mixtures in Part A record your observation below, then write the balanced reaction that occurred including states. If you observe nothing write NR for your observation. In these reactions lead is lead(II).

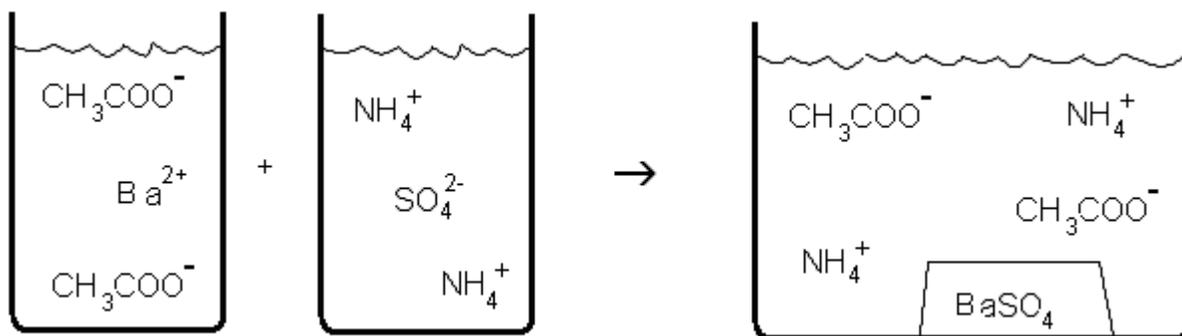
Observation	Balanced reaction if any (if not write NR)
1	___ Na ₂ CO ₃ (aq) + ___ CaCl ₂ (aq) →
2	___ Na ₂ SO ₄ (aq) + ___ CaCl ₂ (aq) →
3	___ KI(aq) + ___ Pb(NO ₃) ₂ (aq) →
4	___ Na ₂ SO ₄ (aq) + ___ KI(aq) →
5	___ CaCl ₂ (aq) + ___ Pb(NO ₃) ₂ (aq) →
6	___ Na ₂ SO ₄ (aq) + ___ Pb(NO ₃) ₂ (aq) →
7	___ CaCl ₂ (aq) + ___ KI(aq) →
8	___ Na ₂ CO ₃ (aq) + ___ Pb(NO ₃) ₂ (aq) →

Part B: Selected Molecular Level Drawings

Refer to Part A for the following selected sets of reactants and fill in the following blanks and beaker drawings. If there is no reaction because all the product ions are spectators, **still do the beaker drawings**, then put NR for the reaction type. For the reaction type choose precipitation, acid-base, or no reaction. Two examples not from this experiment are presented first.

Example: **barium acetate and ammonium sulfate**

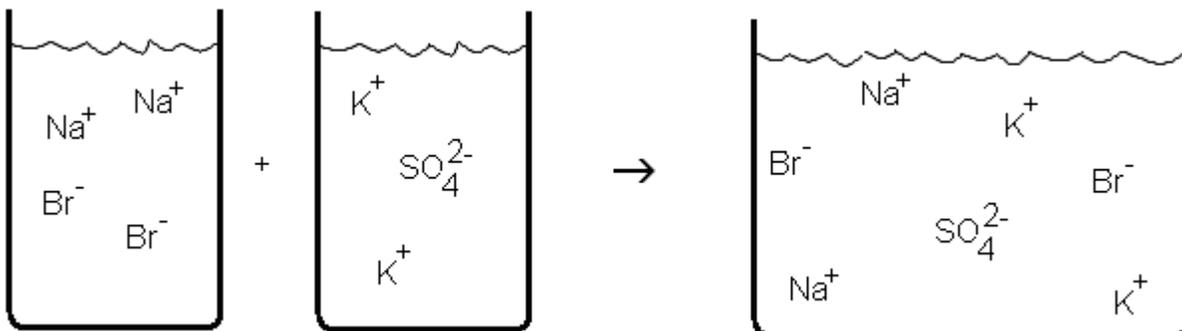
Reaction type: precipitation



Example: **sodium bromide and potassium sulfate**

Reaction type: NR

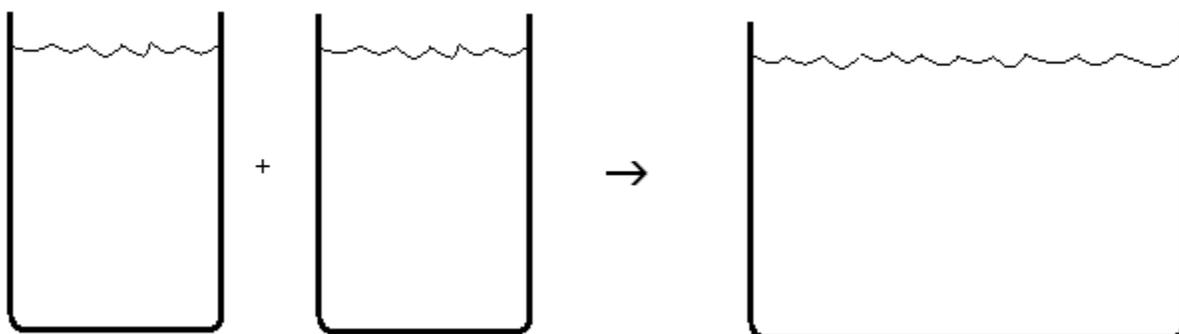
Reaction: $2 \text{NaBr(aq)} + \text{K}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + 2 \text{KBr(aq)}$ (since all ions soluble it is NR)



1. **calcium chloride and potassium iodide**

Reaction type: _____

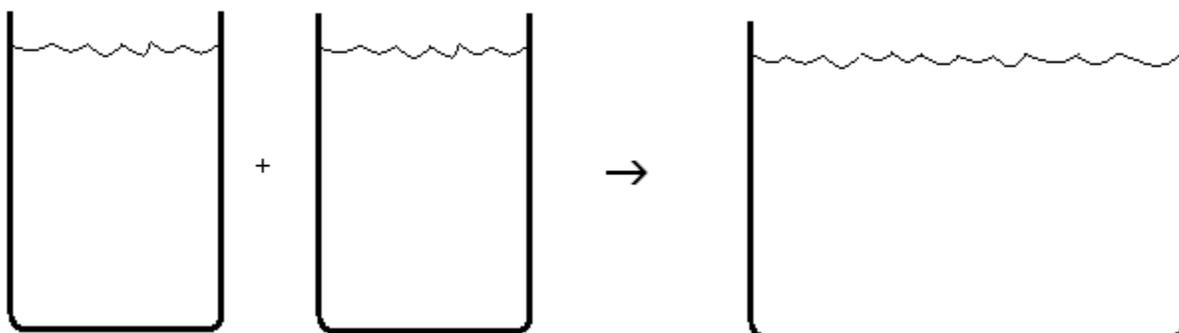
Reaction: _____



2. **nitric acid and sodium hydroxide**

Reaction type: _____

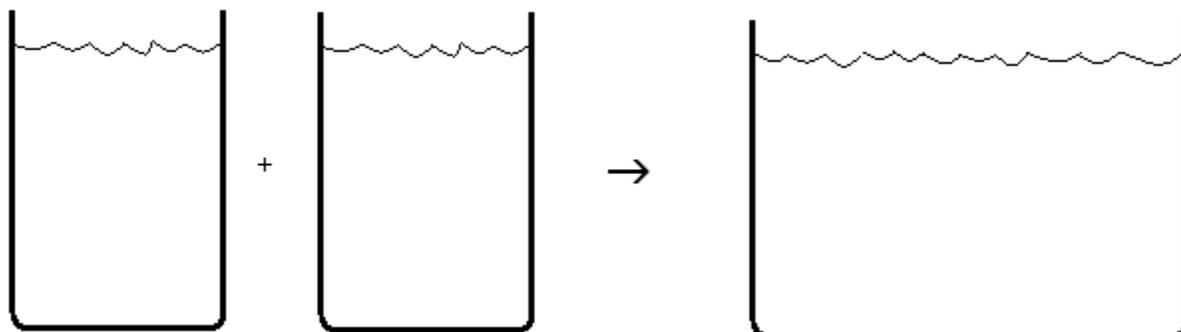
Reaction: _____



3. sodium carbonate and calcium chloride

Reaction type: _____

Reaction: _____



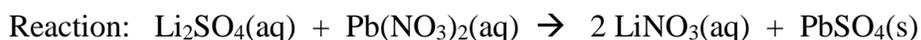
Part C: Follow Up Questions

Use the Solubility Rules (on p. 1) to predict products for each of the following sets of reactants. Write the chemical formulas first for the products, making sure you have the correct formulas, including physical states, then balance the reaction. Afterwards, indicate if a reaction occurred. If a solid, liquid or gas was produced, circle “Yes”; otherwise, circle “No”. The first two have been completed as examples.

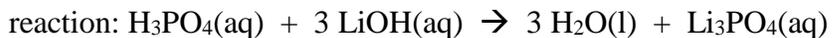
		Reaction?
1.	$2 \text{NaCl(aq)} + 1 \text{K}_2\text{SO}_4\text{(aq)} \rightarrow 1 \text{Na}_2\text{SO}_4\text{(aq)} + 2 \text{KCl(aq)}$	Yes <input type="radio"/> No <input checked="" type="radio"/>
2.	$1 \text{KBr(aq)} + 1 \text{AgNO}_3\text{(aq)} \rightarrow 1 \text{KNO}_3\text{(aq)} + 1 \text{AgBr(s)}$	<input checked="" type="radio"/> Yes <input type="radio"/> No
3.	$\text{NaCl(aq)} + \text{KOH(aq)} \rightarrow$	Yes <input type="radio"/> No <input type="radio"/>
4.	$\text{Ba(NO}_3)_2\text{(aq)} + \text{K}_3\text{PO}_4\text{(aq)} \rightarrow$	Yes <input type="radio"/> No <input type="radio"/>
5.	$\text{BaCl}_2\text{(aq)} + \text{AgNO}_3\text{(aq)} \rightarrow$	Yes <input type="radio"/> No <input type="radio"/>
6.	$\text{H}_2\text{SO}_4\text{(aq)} + \text{NaOH(aq)} \rightarrow$	Yes <input type="radio"/> No <input type="radio"/>

Stoichiometry Calculations – show all your work, each step, as learned in lecture

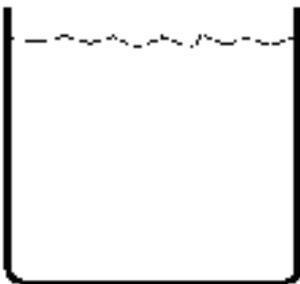
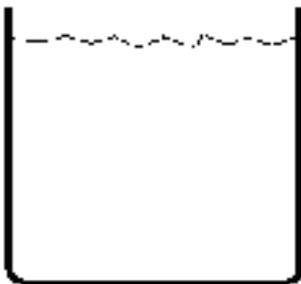
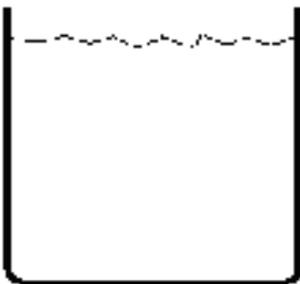
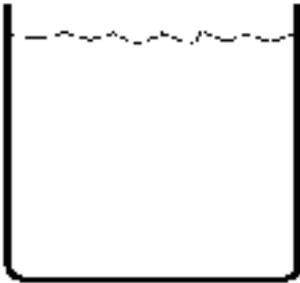
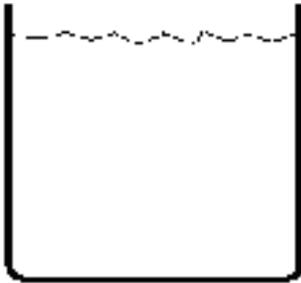
- Calculate how many grams of $\text{PbSO}_4\text{(s)}$ would be produced if 5.50 grams of LiNO_3 were produced.



2. Calculate how many grams of water one could produce if 3.35 grams of H_3PO_4 react completely.



For each compound listed below write the chemical formula and draw a representation of the compound in water similar to the drawings on page two. Do not show the solvent water molecules.

<p>1. calcium bromide</p> <p>formula: _____</p> 	<p>2. lithium acetate</p> <p>formula: _____</p> 	<p>3. lithium phosphate</p> <p>formula _____</p> 
<p>4. ammonium sulfide</p> <p>formula _____</p> 	<p>5. silver iodide</p> <p>formula _____</p> 	<p>6. barium hydroxide</p> <p>formula _____</p> 