AQUEOUS REACTIVITY for CHEM 101A TOPIC B

1) Solubility rules and precipitation reactions

All compounds that contain the cations Na^+, K^+ or NH_4^+, or the anions NO_3^- or $C_2H_3O_2^-$ are soluble in water. It is best to memorize these. Beyond this, solubilities are normally classified using the anion in the compound. Here are the rules that you will use in Chem 101A:

Anion	Cations that produce **soluble** compounds	Cations that produce **insoluble** compounds
NO_3^-, $C_2H_3O_2^-$	All	None
Cl^-, Br^-, I^-	Most	Ag^+ Pb^{2+} Hg_2^{2+}
SO_4^{2-}	Most	Ag^+ Pb^{2+} Hg_2^{2+} Ca^{2+} Sr^{2+} Ba^{2+} (the heavier IIA elements)
OH^-	Na^+ K^+ (NH_4^+ reacts with OH^-: see the acid-base section below) Ba^{2+}	All others (see note below on Ag^+)
CO_3^{2-}, PO_4^{3-}	Na^+ K^+ NH_4^+	All others
S^{2-}	Na^+ K^+ NH_4^+ Mg^{2+} Ca^{2+} Sr^{2+} Ba^{2+} (group IIA)	All others (the reactions of sulfide with 3+ ions are not simple precipitations: you do not need to know these)

An insoluble salt will be produced whenever its constituent ions are mixed. Here are two examples:

Mixing solutions of $AgNO_3$ and K_2CO_3: $2\ Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$

Mixing solutions of $FeCl_2$ and Na_3PO_4: $3\ Fe^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$

Note that the reaction of Ag^+ with OH^- produces Ag_2O (and water), not AgOH. This is a "quirk" of the chemistry of silver ions. The net ionic equation for this reaction is:

$$2\ Ag^+(aq) + 2\ OH^-(aq) \rightarrow Ag_2O(s) + H_2O(l)$$

Note that ammonia dissolves in water to produce a small concentration of hydroxide ions, which can participate in precipitation reactions. The combination $NH_3 + H_2O$ is equivalent to $NH_4^+ + OH^-$ when you are considering reactivity for precipitation reactions. However, $NH_4^+ + OH^-$ are not the major species in solution and should not be written in net ionic equations. Instead, you must include the major species $NH_3 + H_2O$ as the reactants. Remember, "the majority rules" when writing net ionic equations. (Note that $NH_4OH(aq)$ does not exist.) Here are two examples, using Mg^{2+} and Al^{3+}:

$$Mg^{2+}(aq) + 2\ NH_3(aq) + 2\ H_2O(l) \rightarrow Mg(OH)_2(s) + 2\ NH_4^+(aq)$$

$$Al^{3+}(aq) + 3\ NH_3(aq) + 3\ H_2O(l) \rightarrow Al(OH)_3(s) + 3\ NH_4^+(aq)$$

The reactions below would be INCORRECT:

$Mg^{2+}(aq) + 2\ NH_4OH(aq) \rightarrow Mg(OH)_2(s) + 2\ NH_4^+(aq)$ (INCORRECT!)

$Al^{3+}(aq) + 3\ NH_4OH(aq) \rightarrow Al(OH)_3(s) + 3\ NH_4^+(aq)$ (INCORRECT!)

2) Acid-base reactivity

Any compound that contains hydrogen and can lose it (in the form of H^+) is an **acid**. The following acids are <u>strong</u> (100% ionized in aqueous solution):

HCl *(hydrochloric acid)*	HBr *(hydrobromic acid)*	HI *(hydroiodic acid)*
HNO_3 *(nitric acid)*	$HClO_4$ *(perchloric acid)*	H_2SO_4 *(sulfuric acid)*

You may assume that any other acid you encounter in Chem 101A is <u>weak</u>. Here are three common weak acids whose formulas you should know:

$HC_2H_3O_2$ *(acetic acid)*	H_2CO_3 *(carbonic acid)*	H_3PO_4 *(phosphoric acid)*

A compound that binds to H^+ is a **base**. Hydroxide ions (OH^-) bind to H^+ and so any ionic compound that contains OH^- ion is a **base**. Most of these compounds are insoluble in water. The following are soluble in water and 100% ionized. These are <u>strong</u> <u>bases</u>:

$NaOH$ *(sodium hydroxide)*	KOH *(potassium hydroxide)*	$Ba(OH)_2$ *(barium hydroxide)*

You may assume that any other base you encounter in Chem 101A is a <u>weak</u> base. Here is the most common weak base that you will encounter in Chem 101A whose formula you should know:

NH_3 *(ammonia)*

Ammonia (NH_3) binds to H^+ and is therefore a base. It will react with any acid "HX":

$NH_3 + H^+ \rightarrow NH_4^+$	or	$NH_3 + HX \rightarrow NH_4^+ + X^-$
(if HX is a strong acid)		(if HX is a weak acid)

Note that ammonia can also participate in precipitation reactions! See section 1.

3) Special reactions of carbonate and bicarbonate ions

Carbonate ion and bicarbonate ion react with H^+ to form H_2CO_3 (carbonic acid). However, carbonic acid can only exist at very low concentrations. Under normal circumstances, carbonic acid decomposes into CO_2 and H_2O. Therefore, carbon dioxide and water are the normal products whenever carbonate or bicarbonate react with acids. Here are some examples:

Mixing solutions of $NaHCO_3$ and HCl:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$$

Adding a solution of HCl to solid $CaCO_3$:

$$CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

Mixing solutions of $HC_2H_3O_2$ and K_2CO_3 (with the acid in excess):

$$CO_3^{2-}(aq) + 2\,HC_2H_3O_2(aq) \rightarrow H_2O(l) + CO_2(g) + 2\,C_2H_3O_2^-(aq)$$

You can see an example of this kind of reaction for yourself by mixing baking soda ($NaHCO_3$) and vinegar (a solution of acetic acid) in your kitchen. The mixture will bubble vigorously as gaseous carbon dioxide forms.