PROPERTIES AND REACTIVITY OF INORGANIC COMPOUNDS FOR CHEM 101A

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 need to know 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 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: $3\ Fe^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$

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:

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

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 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)*

Any ionic compound that contains OH^- ion is a **base**. Most of these are insoluble in water. The following <u>strong</u> <u>bases</u> are soluble in water and 100% ionized:

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

(acid-base reactivity, continued)

When solutions of strong acids and strong bases are mixed, the net ionic equation is simply:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

If a weak acid is used, you must write out the complete formula of the acid, because most of the molecules of the acid are not ionized. "The majority rules" when writing net ionic equations. Here is the reaction of acetic acid with a strong base.

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$$

Note that sulfuric acid (H_2SO_4) ionizes completely to H^+ and HSO_4^- ions. The ion HSO_4^- is a weak acid, so H^+ and HSO_4^- will both react independently with a base. Therefore, reactions involving sulfuric acid occur in two steps whenever the base is in excess. The base first reacts with the free H^+ ions. Any remaining base reacts with the HSO_4^- ions. Here are the reactions that occur when solutions of sulfuric acid and NaOH are mixed:

step 1: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
step 2: $HSO_4^-(aq) + OH^-(aq) \rightarrow H_2O(l) + SO_4^{2-}(aq)$

Weak acids that contain more than one hydrogen also react in steps. Here are the reactions that occur when solutions of phosphoric acid and NaOH are mixed:

step 1: $H_3PO_4(aq) + OH^-(aq) \rightarrow H_2PO_4^-(aq) + H_2O(l)$
step 2: $H_2PO_4^-(aq) + OH^-(aq) \rightarrow HPO_4^{2-}(aq) + H_2O(l)$
step 3: $HPO_4^{2-}(aq) + OH^-(aq) \rightarrow PO_4^{3-}(aq) + H_2O(l)$

Writing a combined reaction such as the one below is only correct if the hydroxide ions are in excess. If the acid is in excess, this reaction is not valid and will lead to incorrect predictions about the product mixture.

$H_3PO_4(aq) + 3\ OH^-(aq) \rightarrow PO_4^{3-}(aq) + 3\ H_2O(l)$ *(only valid if the base is in excess!!!)*

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

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

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!)

3) Note on weak electrolytes

All weak acids are weak electrolytes (by definition). There are a few other weak electrolytes, but you do not need to learn them.

Weak electrolytes will be formed whenever their constituent ions are mixed. Here are two examples:

Mixing solutions of HCl and $KC_2H_3O_2$:

$H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq)$

($HC_2H_3O_2$ forms because it is a weak acid.)

Mixing solutions of HNO_3 and NaF:

$H^+(aq) + F^-(aq) \rightarrow HF(aq)$

(HF forms because it is a weak acid.)

If a weak electrolyte is a reactant, it must be written as a molecule, not as ions. Here is an example of an acid-base reaction that involves a weak acid (acetic acid):

Mixing solutions of $HC_2H_3O_2$ and KOH:

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

4) 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.