Unit 12 Important Types of AP Reactions:

Hints:

1) Keep it simple!
2) Reactions do not always need to be balanced.
3) Reactions do need charges and subscripts (especially precipitation reactions).
4) If H & OH around, then make H_2O.
5) **Write reactions as net ionic reactions, cancel out spectator ions!**

Keywords:

1) **Excess, Concentrated**: This is not a precipitate reaction!
2) **Excess, Concentrated or excess concentrated or excess saturated:** Usually a redox or complex ion reaction.
3) **One drop of a reactant**: complex ion or redox
4) **Solution:** Write compound as ions.
5) **Two solutions:** precipitate reaction
6) **Wire/ metal bar/ solid/ strip/ powdered metals put into solution:** redox
7) **Acidified:** usually redox reaction with H^+ as a reactant.
8) **Equimolar or Equal volumes:** usually acid/base reaction.
9) If there are no keywords and there is an ion that can cancel: precipitate reaction.
10) If all of the ions are soluble, then not a precipitate reaction.

For Precipitation Reactions: (See Unit 3)

Solubility Rules ** soluble = aqueous
marginally/slightly soluble = solid

Solubility Table

1. Nitrate (NO_3^{-1}) salts are soluble.
2. Ammonium (NH_4^{+1}) and Group I Alkali metal (Li^{+1}, Na^{+1}, K^{+1}, Cs^{+1}, Rb^{+1}) salts are soluble.
3. Chloride, Bromide, Iodide (Cl^{-1}, Br^{-1}, I^{-1}) salts are soluble, except for salts with Ag^{+1}, Pb^{2+} and Hg_2^{2+}.
4. Sulfate (SO_4^{2-}) salts are soluble, except Ba^{2+}, Ca^{2+}, Pb^{2+}, and Hg_2^{2+}.
5. Soluble Hydroxides are LiOH, NaOH and KOH. Marginally soluble: $Ba(OH)_2$, $Sr(OH)_2$ and $Ca(OH)_2$. The rest are slightly soluble.
6. Carbonates (CO_3^{2-}), Sulfides (S^{2-}), Chromates (CrO_4^{2-}) and Phosphates (PO_4^{3-}) are slightly soluble. (All with ammonium or Group 1 Alkali metals are soluble.)

For Complex Ions (see Unit 7)

Usually involve excess concentrated ammonia (NH_3 as ligand), excess concentrated HCl (Cl^{-1} as ligand), excess concentrated metal ion hydroxide (OH^{-1} as ligand), excess concentrated metal ion fluoride (F^{-1} as ligand), a drop of SCN^{-1}, etc. See ligands in Unit 7. The previous ligands will be mixed with Transition Metals to get the complex ion. Ligand number is usually 2X charge of metal.

Combustion Reactions:

Organic compound (C, H) or (C, H, O) + $O_2 \rightarrow CO_2 + H_2O$

Important Reactions:

**** You may see the reverse of these reactions too.****

Acid/Base Reactions:

For General Acid/Base Reactions (See Unit 10)

**Write weak acids not as $HC_2H_3O_2$, but as CH_3CO_2H

Weak acids $HA + OH^- \rightarrow A^- + H_2O$ (forms salt of weak acid)

Remember all variations from unit 10 (see notes): $A^- + OH^-$ etc.

Special Acid/Base Reactions:

****Metal Oxides** are **basic** (produce hydroxides)******: $M_xO_y + H_2O_{(l)} \rightarrow MOH \rightarrow M^+ + OH^-$

$$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)} \rightarrow Mg^{2+}{}_{(aq)} + OH^-{}_{(aq)}$$

If weak base, don't ionize it. Use Bronsted-Lowry (donating/accepting protons (H^+).

Metals (Grp #1 & 2) and Water: react with water to make $H_{2(g)}$ and base.
(metal oxidation)

$$Mg_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)} + H_{2(g)} \rightarrow Mg^{2+}{}_{(aq)} + OH^-{}_{(aq)} + H_{2(g)}$$

****Non-metal Oxides** are **acidic** (N,S,C Oxides)******:

Non-metal oxides usually form by burning compounds, containing N, S, or C.
(i.e. combustion produces NO_2, N_2O_5, SO_2, CO_2 etc.)

Memorize these reactions!

$2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)}$

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)} \rightarrow H^+{}_{(aq)} + HSO_4^-{}_{(aq)}$
HSO_4^- weak has K_a

$SO_{3(g)} + OH^-{}_{(aq)} \rightarrow SO_4^{2-}{}_{(aq)} + H_2O_{(l)}$
sulfate is salt of acid

$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$

** Remember strong acid if # O - # H ≥ 2
If strong break to ions, H_2SO_4 will only loose 1 H^+

$CO_{2(g)} + H_2O_{(l)} \leftrightarrow H_2CO_{3(aq)}$ Reverse: $H^+ + HCO_3^- \rightarrow H_2O + CO_2$

$2\ NO_{2(g)} + H_2O_{(l)} \leftrightarrow HNO_{3(aq)} + HNO_{2(aq)} \rightarrow H^+{}_{(aq)} + NO_3^-{}_{(aq)} + HNO_{2(aq)}$

$N_2O_{5(s)} + H_2O_{(l)} \rightarrow HNO_3 \rightarrow H^+{}_{(aq)} + NO_3^{-1}{}_{(aq)}$

White phosphorus: $P_{4(s)} + 5\ O_{2(g)} \rightarrow P_4O_{10(s)}$ $P_4O_{10(s)}$ formerly known as $P_2O_{5(s)}$

$P_4O_{10(s)} + H_2O_{(l)} \rightarrow H_3PO_{4(aq)}$

Same but with limited Oxygen: $P_{4(s)} + 5\ O_{2(g)} \rightarrow P_4O_{6(s)}$

$P_4O_{6(s)} + H_2O_{(l)} \rightarrow H_3PO_{3(aq)}$

Hydrides:

Forming hydrides: metal + $H_2 \rightarrow$ metal$^+$ H^-

$Na_{(s)} + H_{2(g)} \rightarrow 2\ NaH_{(s)}$

****Hydrides** are **strong bases**:

$NaH_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} + H_{2(g)}$

Redox (see Unit 3):

Oxides:

Forming metal oxides (add O_2 to metal): Heat/burn

(**metal oxidation**) Metal + $O_2 \rightarrow M_xO_y$

$2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$

$4\ Al_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Al_2O_{3(s)}$

**unique: $PbS_{(s)} + O_{2(g)} \rightarrow PbO + SO_2$

$2\ HgO_{(s)} \xrightarrow{heat} 2\ Hg_{(l)} + O_{2(g)}$

Carbon Reactions:

some Metal Oxides + C: $2\ PbO_{(s)} + C_{(s)} \rightarrow 2\ Pb_{(s)} + CO_{2(g)}$

Carbide: $CaC_{2(s)} + 2\ H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + C_2H_2$

Calcium carbide

Decompositions: compound + heat → compound + compound

Carbonates: $(NH_4)_2CO_{3(s)} \xrightarrow{heat} 2\ NH_{3(g)} + CO_{2(g)} + H_2O_{(l)}$

** $CaCO_{3(s)} \xrightarrow{heat} CaO_{(s)} + CO_{2(g)}$

**$NaHCO_{3(s)} \xrightarrow{heat} Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

Strong acid: $H^+ + ZnCO_3 \rightarrow Zn^{2+} + CO_2 + H_2O$

Weak acid: $HA + CaCO_3 \rightarrow Ca^{2+} + CO_2 + A^- + H_2O$

Sulfites: $FeSO_{3(s)} \xrightarrow{heat} FeO_{(s)} + SO_{2(g)}$

Chlorates: $2\ KClO_{3(s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2(g)}$

Hydrogen Peroxide: $2\ H_2O_{2(l)} \xrightarrow{Heat} 2\ H_2O_{(l)} + O_{2(g)}$

Ammonia: $NH_{3(l)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(s)} \xrightarrow{heat} N_2O_{(g)} + 2\ H_2O_{(g)}$

$NH_{3(l)} + SO_{2(g)} + H_2O_{(l)} \rightarrow (NH_4)_2SO_{3(s)}$

$NH_3 + metal \rightarrow complex\ ion$

$NH_3 + HCl \rightarrow NH_4Cl$

$OH^- + NH_4Cl \rightarrow NH_4OH + Cl^- \rightarrow NH_3 + OH^- + Cl- + H_2O$

So $OH^- + NH_4Cl \rightarrow NH_3 + Cl^- + H_2O$

Acid Redox Reactions:

Hydrochloric Acid:

$HCl_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + H_{2(g)}$

Making weak acids with HCl (Not redox):

$HCl(g) + KCN(aq) \rightarrow HCN + Cl^-$ Cancel K^+

$HCl(aq) + NaS \rightarrow HS^- + H^+ \rightarrow H_2S + Na^+$ Cancel Na^+

Sulfuric Acid:

Dilute/warm: $Ag_{(s)} + H_2SO_{4(aq)} \rightarrow Ag^+ + SO_{2(g)} + H_2O_{(l)}$

Stong/hot: $Cu_{(s)} + H_2SO_{4(aq)} \rightarrow Cu^{2+} + SO_{2(g)} + H_2O_{(l)}$

Cold: $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow Zn^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + H_{2(g)}$

Nitric Acid:

$Cu + HNO_3 \xrightarrow{dilute} Cu^{2+}_{(aq)} + NO_{(g)} + H_2O_{(l)}$

$Cu + HNO_3 \xrightarrow{strong} Cu^{2+}_{(aq)} + NO_{2(g)} + H_2O_{(l)}$

Redox: Electrolysis of Water: $2\ e^- + H_2O_{(l)} \rightarrow H_{2(g)} + OH^-_{(aq)}$

Oxidation of MnO_4^-: $5\ e^- + 8\ H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\ H_2O$

Organic Chem: (see Unit 12)

alcohol + carboxylic acid → ester + H_2O

End of Notes