Chapter 1 AP PREDICTING REACTIONS[1]
NOTE: we will cover this chapter throughout the year.

A. Points

- 15 % of Free Response or 7.5% of final grade
- Each rxn worth 5 pts for a total of 15 pts possible.
- Three points for the unbalanced net ionic equation (total nine points). **1 pt reactants, 2 pts products**
- One point for the balanced equation (I am assuming that if the net ionic is wrong, but the equation is balanced correctly, i.e. "internally consistence" then the students will get this point).
- One point for answering a question about the reaction. Possibilities include:
  - indications of chemical reaction
  - oxidizing agent
  - oxidation number
  - color of solid product
  - general chemistry interest
  - simple stoichiometry
- Points are deducted for including incorrect species, omitting species or including spectator ions.
- You don't need to include states, but I like to include (*s*) and (*g*) so I won't be tempted to dissociate incorrectly!

B. Instructions

For each of the following three reactions, in part (i) write a BALANCED equation and in part (ii) answer the question about the reaction. In part (i), coefficients should be in the **lowest whole numbers**. In all cases, a reaction occurs. Assume that solutions are aqueous unless otherwise indicated. Represent substances in solution as ions if the substances are extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction.

Example: (i) A strip of magnesium is added to a solution of silver nitrate.
(ii) Which substance is the oxidizing agent?

(i) $Mg + 2\ Ag^{+} \rightarrow Mg^{2+} + 2\ Ag$

(ii) The Ag^{+} is the oxidizing agent.

a. All reactions occur so you really don't need to worry about activity series (see single replacement)
b. The solvent for all solutions is **WATER** unless indicated otherwise.
   i. If the solute is **soluble** ionic, **strong** acid, or a **strong** base then dissociate the substance
   ii. If the solute is molecular, **weak** acid, or **weak base**, simply write the formula
c. Goal is BALANCED net ionic equation.
d. Do not need to show states, but I strongly recommend including states so you won't accidentally dissociate gases and solids.

REMEMBER YOUR OLD BIDDY (AKA – DR. LEGGETT!)

B: BALANCE
I: IONIZE (dissociate)
D: DELETE SPECTATOR IONS
D: DECOMPOSE
Y: YES – YOU DID IT!!

[1] References: Hand-outs by Todd Abronowitz , Renee McCormick, and Lisa McGaw

SOLUBILITY RULES/STRONG ACIDS/STRONG BASES AND GENERAL GUIDELINES!

SOLUBLE COMPOUNDS	EXCEPTIONS
All Group 1 salts	None
All ammonium (NH_4^+) salts	None
All NO_3^-, ClO_3^-, ClO_4^-, and $C_2H_3O_2^-$ salts	None
All Cl^-, Br^-, I^- salts	Ag^+, Hg_2^{2+} (mercury (I)), Pb^{2+}
All F^- salts	Mg^{2+} Ca^{2+}, Sr^{2+}, $Ba^2$and Pb^{2+}
All salts of SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, $Pb^{2+,}$ Ag^+, Hg_2^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
All salts of OH^-	Group I, NH_4^+, Ba^{2+}, Sr^{2+}, Ca^{2+}
All salts of S^{2-}, SO_3^{2-}, CO_3^{2-}, PO_4^{3-}, CrO_4^{2-} and any other polyatomic not named!	Group I and NH_4^+
Oxides*	Group I and Ca^{2+}, Ba^{2+}, Sr^{2+}

* these oxides are actually "soluble" because they are basic anhydrides and react with water to form a base: MO + $H_2O \rightarrow M(OH)$. More about this later!

Strong Bases: dissociate 100% in water. All hydroxides of group I and II* except beryllium and magnesium.

* Completely dissociated in solutions of 0.01 M or less. These are insoluble bases which ionize 100%. The other five in the list can easily make solutions of 1.0 M and are 100% dissociated at that concentration.

Strong Acids [2] - dissociate 100% in water

Type	Formula		
Hydrogen halides (aq)	HCl	HBr	HI
Oxyacids of halogens	$HClO_3$		
	$HClO_4$		HIO_4
Sulfuric **(1^{st} H^+ only!!)**	$H_2SO_4 \rightarrow H^+ + HSO_4^-$		
Nitric Acid	HNO_3		

Weird products to memorize: If any of these show up as products, you must write them as follows:

REACTANTS →	PRODUCT	MUST DECOMPOSE (TAKE AWAY WATER!)
→	**$H_2CO_3 \rightarrow$**	**$H_2O + CO_2$**
→	**$H_2SO_3 \rightarrow$**	**$H_2O + SO_2$**
→	**$NH_4OH \rightarrow$**	**$NH_3 + H_2O$**

If the reactant says "aqueous ammonia": write $NH_3 + H_2O$, think NH_4OH

[2] http://www.science.uwaterloo.ca/~cchieh/cact/c123/stacids.html

DOUBLE REPLACEMENT

A. How to Recognize: Two ionic AND/OR ACIDS AND/OR BASES compounds react
  a. one or both will be soluble
  b. one or both of the products will be a solid, a gas, or water
  c. these are not redox reactions so oxidation numbers will remain unchanged.

B. Strategy
  a. Write the complete molecular equation, showing states so you won't carelessly ionize solids or gases.
  b. Write the solid or molecular product(s) that formed to the left of the arrow.
  c. Bring down the species needed to form the product and place on the right hand side of the arrow.
    i. If the reactant species is soluble, only bring down the ion that participates
    ii. If the reactant species was not soluble, write the entire formula then balance the product side.
  d. DON'T FORGET!

$$H_2CO_3 \rightarrow H_2O + CO_2$$
$$H_2SO_3 \rightarrow H_2O + SO_2$$
$$NH_4OH \rightarrow NH_3 + H_2O$$

You may not realize this is solid, but you should know for sure that $NaNO_3$ is soluble. Since all reactions "go", the silver chromate must be the insoluble species.

C. Examples:

1. **One Product:** Solutions of silver nitrate and sodium chromate are mixed.

Complete Formula	$2AgNO_3\ (aq) + Na_2CrO_4\ (aq) \rightarrow 2NaNO_3(aq) + Ag_2CrO_4\ (s)$	
Complete Ionic	$2Ag^+ + 2NO_3^- + 2Na^+ + CrO_4^{2-} \rightarrow 2Na^+ + 2NO_3^- + Ag_2CrO_4\ (s)$	
Net Ionic	$\mathbf{2Ag^+ + CrO_4^{2-} \rightarrow Ag_2CrO_4}$	**Spectator Ions:** NO_3^- & Na^+

2. **Two Products:** Sulfuric acid and barium acetate are mixed.

Complete Formula	$H_2SO_4\ (aq) + Ba(C_2H_3O_2)_2\ (aq) \rightarrow BaSO_4(s) + 2HC_2H_3O_2$	
Complete Ionic	$H^+ + HSO_4^-\ (aq) + Ba^{2+} + 2C_2H_3O_2^-\ (aq) \rightarrow BaSO_4(s) + 2HC_2H_3O_2\ (aq)$	
Net Ionic	$\mathbf{H^+ + HSO_4^- + Ba^{2+} + 2C_2H_3O_2^- \rightarrow BaSO_4 + 2HC_2H_3O_2}$	**SIs:** None

1st ionization is strong.

Soluble salt – ionize, but neither are spectators

Insoluble salt – don't dissociate

Weak acid – don't ionize

3. **Product decomposes:** Aqueous solutions of sodium hydroxide and ammonium chloride are mixed.

Complete Formula	$NaOH\ (aq) + NH_4Cl\ (aq) \rightarrow NaCl\ (aq) + NH_4OH\ (aq)$	
Complete Ionic	$Na^+ + OH^- + NH_4^+ + Cl^- \rightarrow Na^+ + Cl^- + NH_4OH\ (aq)$	
Net Ionic	$\mathbf{OH^- + NH_4^+ \rightarrow NH_3 + H_2O}$	**Spectator Ions:** Na^+ & Cl^-

4. **Aqueous ammonia as reactant:** A solution of ammonia is added to a solution of magnesium chloride. **Write:** NH_3 + H_2O **Think:** NH_4OH

Complete Formula	$2NH_3 + 2H_2O + MgCl_2\,(aq) \rightarrow Mg(OH)_2\,(s) + 2NH_4Cl\,(aq)$	
Complete Ionic	$2NH_3 + 2H_2O + Mg^{2+} + 2Cl^- \rightarrow Mg(OH)_2\,(s) + 2NH_4^+ + 2Cl^-$	
Net Ionic	$\mathbf{2NH_3 + 2H_2O + Mg^{2+} \rightarrow Mg(OH)_2 + 2NH_4^+}$	**Spectator Ion(s):** Cl^-

NOTES:
- ✓ you needed to supply the **water**.
- ✓ The chloride ion was a **spectator** ion.
- ✓ Need to keep ammonium since it doesn't cancel with the ammonia

SINGLE REPLACEMENT

A. How to recognize:

a. metal plus an ionic compound
b. Active metal plus water
c. Active metal plus acid
d. halogen plus an ionic halide
e. these are a subset of redox (typically simpler)

B. Don't forget:

a. Diatomics: **I Br**ing **Cl**ay **F**or **O**ur **N**ew **H**ouse
b. Metals that oxidize to their higher oxidation state: **As Sn**oopy **Fe**ll **H**u**g**e **Cu**ps **Cr**acked. All the rest oxidize to their lower oxidation state.
c. You don't need to worry about an activity series because all reactions "GO".

C. Types:

1. Metal plus an ionic compound.

Example: magnesium shavings are added to a solution of iron (II) chloride:

COMPLETE FORMULA	$Mg(s) + FeCl_2(aq) \rightarrow Fe(s) + MgCl_2(aq)$
COMPLETE IONIC	$Mg(s) + Fe^{2+}(aq) + 2Cl^-(aq) \rightarrow Fe(s) + Mg^{2+}(aq) + 2Cl^-(aq)$
NET IONIC	$Mg(s) + Fe^{2+}(aq) \rightarrow Fe(s) + Mg^{2+}(aq)$

Fe becomes solid:
LOSE THE CHARGE!

2. Active metal plus water

Example: A piece of sodium add carefully dropped into water:

COMPLETE FORMULA	$2Na(s) + 2HOH \rightarrow 2NaOH + H_2$	
COMPLETE IONIC	$2Na(s) + 2HOH \rightarrow 2Na^+ + 2OH^- + H_2$	
NET IONIC	$\mathbf{2Na(s) + 2HOH \rightarrow 2Na^+ + 2OH^- + H_2}$	**SI: None**

Writing water as HOH helps see how sodium ends up and an OH^- (DO NOT DISSOCIATE THE WATER!)

NaOH is formed, but it is a strong base so ionize the puppy!

H^+ is reduced to H^o but remember hydrogen is diatomic!

3. Active metal plus acid

Example: A piece of mossy zinc is added to hydrochloric acid:

COMPLETE FORMULA	$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	
COMPLETE IONIC	$Zn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2$	
NET IONIC	$\mathbf{Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)}$	**SI:** Cl^-

4. Halogen plus and ionic halide

Example: Chlorine gas is bubbled through a solution of chromium(II)iodide:

COMPLETE FORMULA	$Cl_2(g) + CrI_2(aq) \rightarrow CrCl_2(aq) + I_2(l)$	
COMPLETE IONIC	$Cl_2(g) + Cr^{2+}(aq) + 2I^-(aq) \rightarrow Cr^{2+}(aq) + 2Cl^-(aq) + I_2(l)$	
NET IONIC	$\mathbf{Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-}$	**SI:** Cr^{2+}

5. Hydrogen is blown over a hot oxide

Example: Hydrogen gas is blown over hot copper (II) oxide: $\mathbf{H_2 + CuO \rightarrow H_2O + Cu}$

COMBINATION/SYNTHESIS/ADDITION

A. Elememt + element OR Cmpd + element OR cmpd + comp (memorize rules for the latter)
B. Elements + element/compounds

a. Metal + nonmetal – follow oxidation number rules

Example: A stream of hot nitrogen flows over a piece of sodium: **Na (*s*) + N_2 (*g*) → Na_3N**

b. Two non-metals – often form more than one product – MEMORIZE!

Example: P_4 + $6Cl_2$ → $4PCl_3$ (in the presence of a **limited** amount of chlorine gas)
Example: P_4 + $10Cl_2$ → $4PCl_5$ (in the presence of a **excess** amount of chlorine gas)

NOTE: Phosphorus exists as the red form (P) or the white form (P_4). Sulfur's most stable form is rings with eight sulfur atoms (S_8)

c. Element plus a compound – usually add up all the of elements

Example: PCl_3 + Cl_2 → PCl_5

C. Compound + compound – simple rules. **NOTE: if there is not a source of water – do not ionize even if a soluble compound forms!**

1. Sulfur dioxide plus a metal oxide → metal sulfite

Example: calcium oxide is exposed to a stream of sulfur dioxide: **CaO + SO_2 → $CaSO_3$**

2. Carbon dioxide plus a metal oxide → metal carbonate

Example: calcium oxide is exposed to a stream of carbon dioxide: **CaO + CO_2 → $CaCO_3$**

3. Sulfur trioxide plus a metal oxide → metal sulfate

Example: sulfur trioxide is blown over hot aluminum oxide: **Al_2O_3 + SO_3 → $Al_2(SO_4)_3$**

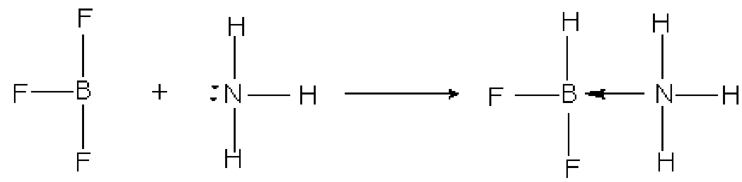
4. Metal oxide plus water → metal hydroxide (NOTE – this is a basic anhydride)

5. Non-metal oxide plus water → acid (NOTE – this is an acidic anyhydride)

Weird exception: Carbon dioxide is bubbled through water: **CO_2 + H_2O → H_2CO_3**
Earlier you were told that H_2CO_3 automatically decomposes to form CO_2 + H_2O, but in this case, we would just get our reactants back ie NO RXN! However, in this type of reaction, the pressure of CO_2 is sufficiently high to keep the reaction shifted to the product side (remember the French Guy??!!)

***Random addition*:** The gases boron trifluoride and ammonia are mixed. **BF_3 + NH_3 → F_3BNH_3.**
Ammonia is Lewis base and BF_3 is a Lewis acid so make sure the B and N are adjacent. Hydrogens can

be replaced by halogens and "P" may be used instead of "B". Since the nitrogen (or phosphorus) contributes both electrons to the bond, the bond is called a **coordinate covalent bond.**



DECOMPOSITION

A. **How to recognize**: A single compound is heated (Δ), electrolyzed, decomposed – often in the presence of a catalyst such as MnO_2. Heat, electricity, and catalysts are shown above the arrow in the reaction.
B. **Compound → 2 elements**
C. **Compound → element + compound**
D. **Compound → 2 simpler compounds** – these rules are typically the opposite of the synthesis rules.

- Metal carbonates decompose to form metal oxides plus CO_2
- Metal sulfites decompose to form metal oxides plus SO_2
- Metallic chlorates break down to yield metallic chlorides and oxygen.
- Ammonium Carbonate decomposes to form ammonia, water, and CO_2
- Hydrogen peroxide decomposes to form H_2O and O_2
- Carbonic acid SPONTANEOUSLY decomposes to form H_2O and CO_2
- Sulfurous acid SPONTANEOUSLY decomposes to form H_2O and SO_2
- Sodium hydrogen carbonate decomposes to form sodium carbonate, CO_2, and H_2O
- Hydrated salts → salt + water when GENTLY heated.

COMBUSTION

A. How to recognize:

a. adding oxygen to a compound
b. Look for words such as “burned”, “undergoes combustion”. NOTE: You need to supply the oxygen because it will not be explicitly given!

NOTE: you are not in Pre-AP any more. The products of combustion are only CO_2 and water for **hydrocarbons!**

B. Strategy: oxygen will combine with each element in the other reactant to form the most common compound.

element	Most common oxygen containing cmpd
C	CO in limited oxygen
C	CO_2 in excess oxygen (Assume if lim or xs not indicated)
S_8	SO_2 in limited oxygen (Assume if lim or xs not indicated)
S_8	SO_3 in excess oxygen
N	NO in limited oxygen
N	NO_2 in excess oxygen(Assume if lim or xs not indicated)
P_4	P_2O_5 or P_4O_{10}
H	H_2O
Metal	Metallic oxide

C. Samples:

Example: Propane is combusted in excess oxygen: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
Example: Ethane is burned in the presence of limited oxygen: $2C_2H_6 + 5O_2 \rightarrow 4\,CO + 6H_2O$
Example: Magnesium sulfide is burned in air: $2MgS + 3O_2 \rightarrow 2MgO + 2SO_2$
Example: Carbon disulfide is burned in excess oxygen: $CS_2 + 4O_2 \rightarrow CO_2 + 2SO_3$
Example: Ammonia is heated with oxygen gas in the presence of platinum:

$$4NH_3 + 7\,O_2 \xrightarrow{Pt} 4NO_2 + 6H_2O$$

COMPLEX IONS/COORDINATION COMPOUNDS

A. How to recognize: Transition metal or aluminum with ammonia, hydroxide, cyanide, or fluoride (not common)

B. Definitions:

a. Complex ion:
  i. Metal ion that acts as a **Lewis Acid** (accepts electron pairs)
  ii. Ligand: group with non-bonded pair of electrons that can ace as a **Lewis base** (donates electron pairs)

b. Coordination complex: Complex ion with a counter ion to balance charge to form a neutral complex

c. Coordination number: The number of bonds a metal ion is able to form with ligands. Depends on size, oxidation state and electron configuration.

d. Coordinate Covalent Bond: A covalent bond in which one of the bond participants supplies both of the electrons.

C. Common Coordination numbers

2	2 & 4	4	4 & 6	6
Ag^+	Cu^+	Zn^{2+}	Al^{3+}	Fe^{3+}
	Au^+	Cd^{2+}	Co^{2+}	Fe^{2+}
			Ni^{2+}	Cr^{3+}
			Cu^{2+}	Co^{3+}

NOTE: AP will accept any reasonable oxidation state. For most of these it means that you can use a coordination number that is simply twice the charge! Fe^{2+} is the only exception.

B. Common Ligands:

Ligand	Name	Charge balance
CN^-	Cyano	must balance + and – charge
OH^-	Hydroxo	must balance + and – charge
F^-	Fluoro	must balance + and – charge
Cl^-	Chloro	must balance + and – charge
Br^-	Bromo	must balance + and – charge
I^-	Iodo	must balance + and – charge
SCN^-	thiocyano	must balance + and – charge
NH_3	Ammine	charge is the same as the + ion
H_2O	Aqua	charge is the same as the + ion

C. Sample reactions:

Metal	Ligand	Complex ion	Name	Counter ion
Al***, Al^{3+}, or $Al(OH)_3$	OH^-	$[Al(OH)_4]^-$ or $[Al(OH)_6]^{-3}$	tetrahydroxoaluminum ion	+ ion such as K^+
$Zn(OH)_2$	OH^-	$[Zn(OH)_4]^{2-}$	tetrahydroxozinc ion	+ ion such as K^+
Ag^+	NH_3	$[Ag(NH_3)_2]^+$	diamminesilver ion	(-) ion such as Cl^- or SO_4^{2-}
Zn^{2+}	NH_3	$[Zn(NH_3)_4]^{2+}$	tetramminezinc ion	(-) ion such as Cl^- or SO_4^{2-}
Cu^{2+}	NH_3	$[Cu(NH_3)_4]^{2+}$	tetramminecopper (II) ion*	(-) ion such as Cl^- or SO_4^{2-}
Fe^{3+}	***SCN^-(weird)***	***$[FeSCN]^{2+}$***	***thiocyanoiron (III) ion*****	(-) ion such as Cl^- or SO_4^{2-}

*note that roman numerals must be shown for metals with more than one common oxidation number!
** sometimes called thiocyanoferrate (III)
*** when the Al metal reacts, H_2 gas is also formed (Al is oxidized, H^+ is reduced)

D. Dissociation of complex ions: Adding Acid – can typically approach like a double replacement

Ex. Dilute hydrobromic acid is added to a solution of diamminesilver nitrate.

COMPLETE FORMULA	$HBr + Ag(NH_3)_2NO_3 \rightarrow AgBr + NH_4NO_3$	
COMPLETE IONIC	$H^+ + Br^- + [Ag(NH_3)_2]^+ + NO_3^- \rightarrow AgBr + NH_4^+ + NO_3^-$	
NET IONIC	$4H^+ + Br^- + [Ag(NH_3)_2]^+ \rightarrow AgBr\ (s) + 2NH_4^+$	**SI:** NO_3^-

ANHYDRIDES

A. How to recognize: Anhydride means "without water" so we are taking a substance that is without water and re-hydrating it so to speak

B. Rules:

1. Metal oxides are "**basic anhydrides"**

    a. Form the base, keeping the oxidation number of the metal constant from reactant to product.
    b. If a strong base is formed, show the base in its ionized form.

Example: Water is added to iron (III) oxide: $\mathbf{Fe_2O_3 + 3H_2O \rightarrow 2Fe(OH)_3}$

Example: Water is added to potassium oxide: $\mathbf{K_2O + H_2O \rightarrow 2K^+ + 2OH^-}$

2. Non-metal oxides are "**acidic anhydrides"**

    a. Form the oxyacid, **keeping the oxidation number** of the non-metal **constant** from reactant to product.
    b. If a strong acid is formed, who the acid in its ionized form.

Example: Sulfur trioxide is bubbled through water.

$SO_3 + H_2O \rightarrow H_2SO_x$ the result is H_2SO_4 which needs to be written as $H^+ + HSO_4^-$

We know this is a "2" because all sulfur oxyanions have a -2 charge

We can now calculate the number of oxygens required to maintain the same oxidation state on sulfur (+6)

Example: dinitrogen trioxide is added to water: $\mathbf{N_2O_3 + H_2O \rightarrow 2HNO_2}$ (weak acid – don't ionize!)

3. Metallic hydrides plus water yield metallic hydroxides and hydrogen gas.

Example: solid sodium hydride is added to water.: $\mathbf{NaH + HOH \rightarrow Na^+ + OH^- + H_2}$

This looks a little like a double replacement but isn't a typical one since we have redox happening! Writing water as HOH is frequently helpful in working reactions.

4. Phosphorus halides and phosphorus oxyhalides react with water to produce two acids: phosphorus oxyacid and a hydrohalic acid (HCl, HBr, HI).

Example: Phosphorus trichloride is added to water: $\mathbf{PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3H^+ + 3Cl^-}$

Note that the phosphorus is +3 just like in PCl_3.

This looks a little like a double replacement

5. Group I and II nitrides react with water to form a base and ammonia

Example: Strontium nitride is added to water: $\mathbf{Sr_3N_2 + 6HOH \rightarrow 3Sr^{2+} + 6OH^- + 2NH_3}$

6. Metal Carbides react with water to form metal hydroxides and methane.

ACID BASE NEUTRALIZATION (SALT HYDROLYSIS IS THE OPPOSITE!)

A. How to Recognize:

a. Look for an acid plus a base. This is really a subset of double replacement with products typically being a salt and water.
b. If the acid is polyprotic, the product may still be an acid.
c. Look for the words "equal molar", "equal volume", "same moles", "twice the moles", "excess". You will need to consider whether the acid or base is fully or partially neutralized.
d. BE CAREFUL! Sometimes anhydrides are mixes with acid/base neutralization.

B. Strategy

a. Follow a similar strategy as double replacement.
b. If the words such as "equal molar" are used, then put the mole ratio under the species in the complete molecular.

C. Types:

1. ***Fully Neutralized***

Example: Hydrogen sulfide gas is bubbled through a solution of excess potassium hydroxide.

Complete Formula	H_2S (*g*) + 2KOH (*aq*) → K_2S (*aq*) + H_2O	
Complete Ionic	H_2S (*g*) + $2K^+$ + $2OH^-$ → $2K^+$ + S^{2-} + $2H_2O$	
Net Ionic	**$H_2S(g)$ + $2OH^-$ → S^{2-} + $2H_2O$**	**Spectator Ions: K^+**

K_2S is exception to sulfide rule and is soluble: dissociate and K^+ becomes spectator

H_2S is a **gas** and a weak acid – don't ionize!

Strong base: ionize!
K^+ is spectator – don't include!

2. ***Partially Neutralized***

Example: Equimolar volumes of phosphoric acid and sodium hydroxide solutions are mixed.

H_3PO_4 (*aq*) + LiOH(*aq*) → LiH_2PO_4 (*aq*) + H_2O

Mole ratio: **1** **1**

H_3PO_4 + OH^- → $H_2PO_4^-$ + H_2O

NOTES:
- H_3PO_4 weak
- LiOH strong, Li^+ spectator
- One OH^- can only neutralize one H^+

3. ***Anhydride*: metal oxide + acid OR nonmetal oxide + base**

Example: Sulfur dioxide is bubbled through a solution of excess strontium hydroxide.

Complete Formula	$SO_2 + H_2O + Sr(OH)_2\ (aq) \rightarrow H_2SO_3 + Sr(OH)_2\ (aq) \rightarrow SrSO_3\ (s) + 2H_2O$	
Complete Ionic	$SO_2 + H_2O + Sr^{2+} + 2OH^- \rightarrow H_2SO_3 + Sr^{2+} + 2OH^- \rightarrow SrSO_3\ (s) + 2H_2O$	
Net Ionic	$\mathbf{SO_2 + H_2O + Sr^{2+} + 2OH^-\ (aq) \rightarrow SrSO_3\ (s) + 2H_2O}$	**Spectator Ions: None**

BE CAREFUL! Not all salts from a neutralization will be soluble!

Example: Solid barium oxide is dropped into a solution of excess phosphorous acid.

Complete Formula	$3BaO\ (s) + 2H_3PO_3\ (aq) \rightarrow 3Ba(OH)_2(aq) + 2H_3PO_3(aq) \rightarrow Ba_3(PO_3)_2 + 6\ H_2O$	
Complete Ionic	$3BaO\ (s) + H_3PO_3 \rightarrow 3Ba^{2+} + 6OH^- + 2H_3PO_3 \rightarrow Ba_3(PO_3)_2 + 6H_2O$	
Net Ionic	$\mathbf{3BaO\ (s) + H_3PO_3 \rightarrow Ba_3(PO_3)_2 + 6H_2O}$	**Spectator Ions: None**

OXIDATION/REDUCTION

A. How to Recognize:

a. Look for the terms "acidic", "acidified", "basic", "alkaline", "concentrated", "dilute"
b. NOTE: you will have to supply the H^+ or OH^- if not given explicitly as a reactant.
c. Alternatively, the H^+ or OH^- may end up on the product side. Water is used to balance extra hydrogens or oxygens.
d. There are some key oxidizing and reducing agents that only appear in redox reactions. PRACTICE will help you become more familiar with these agents!

B. Strategies:

a. Memorize the products for common agents:

Oxidizing agents (ie will be reduced)	Products formed	Δ in oxidation state
$\mathbf{Mn}O_4^{1-}$ (acidic sol'n)	Mn^{2+}	$+7 \rightarrow +2$
$\mathbf{Mn}O_4^{1-}$ (neutral or basic sol'n)	MnO_2(s)	$+7 \rightarrow +4$
$\mathbf{Mn}O_2$ (acidic sol'n)	Mn^{2+}	$+4 \rightarrow +2$
$\mathbf{Cr}_2O_7^{2-}$ (acidic sol'n)	Cr^{3+}	$+6 \rightarrow +3$
$H\mathbf{N}O_3$ (concentrated)	NO_2	$+5 \rightarrow +4$
$H\mathbf{N}O_3$ (dilute)	NO	$+5 \rightarrow +2$
$H_2\mathbf{S}O_4$ (Concentrated, hot)	SO_2	$+6 \rightarrow +4$
$Na_2\mathbf{O}_2$ (peroxide)	$Na\mathbf{O}H$	$-1 \rightarrow -2$
$H_2\mathbf{O_2}$ (peroxide)	$H_2\mathbf{O}$ (or HOH)	$-1 \rightarrow -2$
ClO_4^- (In $HClO_4$ – str acid)	Cl^-	$+7 \rightarrow -1$
Free Halogens (F_2, Cl_2, Br_2, I_2)	Halide ion	$0 \rightarrow -1$
Metal-ic ions (higher oxidation state)	Metal-ous ions (lower)	Ex. $Pb^{4+} \rightarrow Pb^{2+}$ $Cu^{2+} \rightarrow Cu^+$

Reducing Agent (ie will be oxidized)	Product formed	Δ in oxidation state
C$_2O_4^{2-}$ (oxalate ion)	CO_2	+3 → +4
HCOOH (formic acid)	CO_2	+2 → +4
H_2**O**$_2$	O_2	-1 → 0
Halide ions (F^-, Cl^-, Br^-, I^-)	Free halogen	-1 → 0
Free metals	metal ion*	0 → memorize!
Metal-ous	Metal-ic	Memorize!
SO_3^{2-} (sulfite) or **S**O_2	SO_4^{2-} (sulfate)	+4 → +6
NO_2^- (nitrite)	**N**O_3^- (nitrate)	+3 → +5
Free halogens (dilute basic) ex. Cl_2	Hypohalite (ClO^-)	0 → +1
Free Halogens (conc. Basic) ex. Cl_2	Halate ion (ClO_3^-)	0 → +5

*Metals that oxidize to their **higher** oxidation state: **As Sn**oopy **Fe**ll **H**u**g**e **Cu**ps **Cr**acked. All the rest oxidize to their lower oxidation state.

b. We will work on balancing when we get to the electrochemistry unit, but the following will help:
   i. If it is basic solution (this is simplified and won't work with metal hydroxides):
   - add twice as many OH^- as needed to the side deficient in oxygen
   - then H_2O to the other side to balance hydrogens

   ii. If it is acidic:
   - add H_2O to the side deficient in oxygen to balance oxygens
   - add H^+ to the other side to balance the hydrogens

	Acidic	Basic
Oxygens	+ H_2O per deficient oxygen	+ $2OH^-$ per deficient oxygen
Hydrogens	+ H^+ per deficient hydrogen	+ H_2O per deficient hydrogen

MODEL: A solution of potassium permanganate is added to an acidified solution of iron (II) chloride.
Oxidation Inspection Method:
Step one: Assign oxidation numbers

$$\overset{+1\ +7\ -2}{KMnO_4} + \overset{+2\ -1}{FeCl_2} \rightarrow \overset{+3\ -1}{FeCl_3} + \overset{+2}{Mn^{2+}}$$

Step two: Ionize, dissociate, eliminate spectator ions!

$$\overset{+7}{MnO_4^-} + \overset{+2}{Fe^{2+}} \rightarrow \overset{+3}{Fe^{3+}} + \overset{+2}{Mn^{2+}}$$

Step three: Balance electrons

$$\overset{+7}{MnO_4^-} + \overset{+2}{Fe^{2+}} \rightarrow \overset{+3}{Fe^{3+}} + \overset{+2}{Mn^{2+}}$$

(-1e⁻)5

(+5e⁻)1

$$MnO_4^- + 5Fe^{2+} \rightarrow 5Fe + Mn^{2+}$$

Step four: Balance "O" deficiencies & "H" deficiencies

$\mathbf{8H^+} + MnO_4^- + Fe^{2+} \rightarrow Fe^{3+} + Mn^{2+} + 4\mathbf{H_2O}$ (since acidic add water to oxygen deficient and then H^+ to the hydrogen deficient side)

Half Reaction Method

Step one: Assign oxidation numbers

$$\overset{+1\ +7\ -2}{KMnO_4} + \overset{+2\ -1}{FeCl_2} \rightarrow \overset{+3\ -1}{FeCl_3} + \overset{+2}{Mn^{2+}}$$

Step two: Ionize, dissociate, eliminate spectator ions!

$$\overset{+7}{MnO_4^-} + \overset{+2}{Fe^{2+}} \rightarrow \overset{+3}{Fe^{3+}} + \overset{+2}{Mn^{2+}}$$

Step three: Write balanced half-reactions for elements that were oxidized or reduced. I like to bring down subscripts at this point.

$$Mn^{7+} + 5\,e^- \rightarrow Mn^{2+}$$
$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

Step four: Balance electrons

$$1 \text{ x } (Mn^{7+} + 5\,e^- \rightarrow Mn^{2+}) \text{ becomes } Mn^{7+} + 5\,e^- \rightarrow Mn^{2+}$$

$$5 \text{ x } (Fe^{2+} \rightarrow Fe^{3+} + e^-) \text{ becomes } 5\,Fe^{2+} \rightarrow 5\,Fe^{3+} + 5\,e^-$$

Step five: Move balance coefficients into original equation

$$MnO_4^- + 5\,Fe^{2+} \rightarrow 5\,Fe^{3+} + Mn^{2+}$$

Step six: Balance "O" deficiencies & "H" deficiencies

$\mathbf{8H^+} + MnO_4^- + Fe^{2+} \rightarrow Fe^{3+} + Mn^{2+} + 4\mathbf{H_2O}$ (since acidic add water to oxygen deficient and then H^+ to the hydrogen deficient side)

SOME ODD RULES TO MEMORIZE! You can figure them out using the rules above, but they are tricky!

1. ***Chlorine gas reacts with DILUTE sodium hydroxide to produce sodium hypochlorite, sodium chloride, and water.***: $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$
2. ***Chlorine gas reacts with CONCENTRATED sodium hydroxide to produce sodium hypochlorite, sodium chloride, and water.***: $Cl_2 + 2OH^- \rightarrow ClO_3^- + Cl^- + H_2O$
3. ***Copper reacts with CONCENTRATED sulfuric acid to produce copper (II) sulfate, sulfur dioxide, and water***: $Cu + 2H^+ + 2HSO_4^- \rightarrow Cu^{2+} + SO_4^{2-} + SO_2 + 2H_2O$
4. ***Copper reacts with DILUTE nitric acid to produce copper (II) nitrate, nitrogen MONOXIDE, and water.***: $8H^+ + 2NO_3^- + 3Cu \rightarrow 3Cu^{2+} + 2NO + 4H_2O$
5. ***Copper reacts with CONCENTRATED nitric acid to produce copper (II) nitrate and nitrogen DIOXIDE, and water.***: $4H^+ + 2NO_3^- + Cu \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$