

A Classification of AP Chemistry Reactions

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General Principles of Equation-Writing

In writing the reactants:

- Write all soluble ionic compounds in aqueous solution as their constituent ions. Never write NaCl, Ca(NO₃)₂, etc.
- Write all strong acids in water solution as H⁺ or H₃O⁺ plus the anion. Never write HCl or HNO₃.
- Consider writing water as HOH instead of H₂O, especially in acid/base reactions. Sometimes water is a reactant even if it's not explicitly stated (all "solutions" are water solutions).
- If the reaction is in "an acidified solution," H⁺ is almost sure to be a reactant, and the OH⁻ should not appear in the equation.
- Don't omit any ions unless you're absolutely sure that they'll be spectator ions. For example, although the Cl⁻ in hydrochloric acid is usually a spectator, it is sometimes a reactant.

After writing the products:

- H₂ and O₂ are almost never products of chemical reactions, except in certain special cases that will be discussed below.
- Inspect all of the ions in the products to see if there are any precipitates.
- Finally, cross out any spectator ions.

Precipitation Reactions: Memorize your SOLUBILITY RULES!!! (see below)

Many of the reactions are simple reactions that involve formation of an insoluble precipitate. These reactions generally involve mixing two solutions, but may involve hydrogen sulfide gas. Watch for the possibility of two insoluble products.

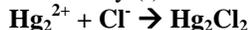
Classification based on insoluble anions; solubility rules

The soluble cation rule is: All common salts of Group 1 metals (alkali metals) and the ammonium ion are soluble. Other solubility rules are best remembered based on the solubility of certain anion. First of all, all common nitrates and acetates are soluble. Other anions are insoluble in certain cases:

Halides: Cl⁻, Br⁻, I⁻

The halides are generally soluble except with silver, mercury (I), and lead. Note that mercury (I) is written as Hg₂²⁺.

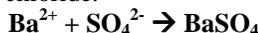
- Dilute hydrochloric acid is added to a dilute solution of mercury (I) nitrate.



Sulfates: SO₄²⁻

Sulfates are generally soluble except with calcium, barium, strontium, silver, mercury (I), and lead.

- Dilute sulfuric acid is added to a solution of barium chloride.



Hydroxides: OH⁻

Hydroxides are generally insoluble except with the soluble cations (above) and limited solubility with calcium, barium, and strontium.

- A solution of copper (II) sulfate is added to a solution of barium hydroxide.



Sulfides: S²⁻

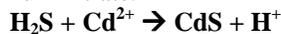
Phosphates: PO₄³⁻

Chromates: CrO₄²⁻

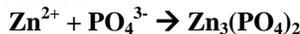
Carbonates: CO₃²⁻

Other anions are generally insoluble except with the soluble cations (above).

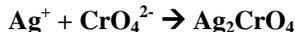
- Hydrogen sulfide gas is added to a solution of cadmium nitrate.



- Solutions of zinc sulfate and sodium phosphate are mixed.



- Solutions of silver nitrate and sodium chromate are mixed.



- **Note:** Although we are finished with simple precipitation, any of the following reactions may produce insoluble salts that precipitate. Be sure to check solubility rules for the products of every reaction.

Acid/Base Reactions

Acid/base reactions always exhibit no change in oxidation number for any of the ionic or molecular species involved in the reaction.

Special Cases

Special cases are classified based on the type of atom that exhibits special behavior: oxygen, nitrogen, hydrogen, phosphorus, aluminum, and zinc.

Oxygen: Oxides

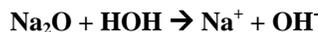
Oxides can be either acidic or basic, depending on what happens when they are added to water. Remember, there is likely to be no change in oxidation number when an oxide dissolves in water.

Oxides of non-metals are acidic:

- Sulfur trioxide gas is added to excess water.



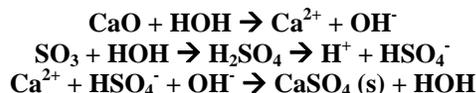
- Solid sodium oxide is added to water.



If there is no water present, oxides can be Lewis acids and bases: they combine to form a salt as if water were present. There is no change in oxidation number for such a reaction.

- Solid calcium oxide is heated in the presence of sulfur trioxide gas.

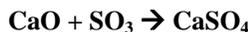
In water, we would see:



So, our overall reaction would be:

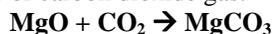


But this can happen without the water:



This is the intended answer. It is helpful, however, if students see the logic of the reaction in water solution.

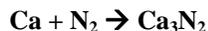
- Powdered magnesium oxide is added to a container of carbon dioxide gas.



Nitrogen

Nitrogen gas will react with some active metals to form nitride (N^{3-}) compounds.

- Calcium metal is heated strongly in nitrogen gas.



The nitride ion is very basic and pulls hydrogen from water molecules to form ammonium and hydroxide.

- Solid lithium nitride is placed in water.



Hydrogen

Hydrogen can form H^- ions with active metals. Such compounds, called ionic hydrides, are very basic and pull hydrogen off of water, forming hydrogen gas and hydroxide.

- Water is added to a sample of pure sodium hydroxide.



Phosphorus

Phosphorus halides form solutions of two acids when added to water. The phosphorus forms the weak oxyacid with the same oxidation number (+5 for phosphoric, +3 for phosphorous) while the halide forms the hydrogen halide acid.

- Solid phosphorus pentachloride is added to excess water.



Aluminum

In water solution, aluminum ions form complexes with six ligands. When base is added to the hydrated aluminum ion, $\text{Al}(\text{H}_2\text{O})_6^{3+}$, some H^+ is removed to leave $\text{Al}(\text{OH})_3(\text{H}_2\text{O})_3$, commonly written $\text{Al}(\text{OH})_3$ – a precipitate. Addition of more base results in $\text{Al}(\text{OH})_4(\text{H}_2\text{O})_2^-$, sometimes written $\text{AlO}_2^- + 4\text{H}_2\text{O}$. This means that aluminum hydroxide is amphoteric and will react with acid and base.

- Excess concentrated sodium hydroxide is added to solid aluminum hydroxide.



Zinc

Zinc ions, Zn^{2+} , coordinate with only four ligands but are amphoteric in the same way as aluminum. Zinc forms $\text{Zn}(\text{H}_2\text{O})_4^{2+}$ - $\text{Zn}(\text{OH})_2(\text{H}_2\text{O})_2(\text{s}) - \text{Zn}(\text{OH})_4^{2-}$.

- Excess concentrated potassium hydroxide is added to a precipitate of zinc hydroxide.

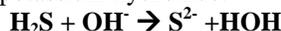


General Acid/Base Reaction

Weak Acid + Strong Base (OH⁻)

In this category a weak acid is mixed with a hydroxide. The conjugate base is produced along with water. Poly-protic acids can lose more than one H if there is excess hydroxide:

- Hydrogen sulfide is bubbled through a solution of potassium hydroxide.



Strong Acid (H⁺) + Weak Base

In this category hydronium ions are added to a weak base. The conjugate acid is produced. More than one H⁺ can be added in poly-protic situations.

- Dilute hydrochloric acid is added to a solution of potassium carbonate.



Weak Acid + Weak Base

In this category a proton is transferred from the weak acid to the weak base. If both reactants are weak acids/bases, a proton must be transferred because "in all cases a reaction occurs." A household example of this is baking soda (sodium bicarbonate) plus vinegar (acetic acid):

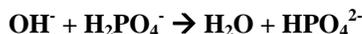
- Acetic acid solution is added to a solution of sodium hydrogen carbonate.



"Titration reactions"

These reactions fall into the above categories but require some thought as to Stoichiometry. A simple table to keep track of the moles of acid and base might be helpful.

- A solution of sodium hydroxide is added to a solution of sodium dihydrogen phosphate until the same number of moles of each compound have been added.



Hydrolysis

In hydrolysis reactions a salt of a weak acid or weak base is added to water. Water acts as a weak acid or base itself, and a small amount of H⁺ or OH⁻ is produced.

- Solid sodium cyanide is added to water.



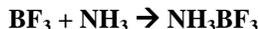
- Solid aluminum nitrate is dissolved in water.



Lewis Acid/Base

One classic Lewis acid/base reaction is:

- The gases boron trifluoride and ammonia are mixed.



Combustion: Oxygen

Oxygen gas supports combustion. A standard laboratory preparation is:

- Solid potassium chlorate is heated in the presence of manganese dioxide as a catalyst.



Combustion of compounds; classification

There are some simple generalizations that can be made about combustion, depending on what atoms or ions are present in the substance to be burned:

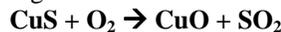
If the compound contains:

- a metal
- carbon
- hydrogen
- sulfur

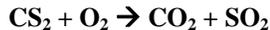
A product will generally be:

- the metal oxide
- carbon dioxide, CO₂
- water, H₂O
- sulfur dioxide, SO₂

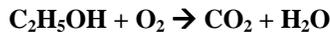
- Solid copper (II) sulfide is heated strongly in oxygen gas.



- Carbon disulfide vapor is burned in excess oxygen.



- Ethanol is completely burned in air.



Redox Reactions

Special cases

Active Metals

Active metal, like the group 1 and 2 metals, reduce water to form hydrogen gas and hydroxide:

- Sodium metal is added to water.



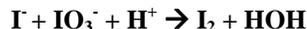
Halogens

Students should recognize the following halogen disproportionation reactions:

- Chlorine gas is bubbled into a cold solution of sodium hydroxide.



- Solutions of potassium iodide and potassium iodate are mixed in acid solution.



Hydrogen

Hydrogen gas, H_2 , is an effective reducing agent for some metal oxides.

- Hydrogen gas is passed over hot copper (II) oxide.



Electron Transfer Reactions

The first general type of redox reactions are simple electron-transfer equations. These do not involve oxygen or oxyanions. They usually have two reactants and two products, as electrons are transferred from one reactant to the other.

Non-oxidizing acids

Non-oxidizing acids are strong acids with anions that can't be reduced. There include HCl, HBr, and HI. They will react with more active metals to produce hydrogen gas and the metal ion. In these cases only the H^+ reacts. The E° tables given with the AP exam can be checked to see if the metal will react with H^+ .

- Calcium metal is added to a dilute solution of hydrochloric acid.



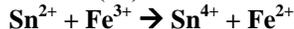
E° tables

In general, the E° tables should always be checked to determine if the two species will react. Some common examples are:

- A strip of magnesium is added to a solution of silver nitrate.



- A solution of tin (II) chloride is added to a solution of iron (III) sulfate.



- Chlorine gas is bubbled into a solution of sodium bromide.



Redox with Oxygen Compounds

The second type of redox that involve oxygen-containing compounds such as nitrates, sulfates, permanganates, dichromates, etc. First of all, since these are redox reactions, one thing must be oxidized and another must be reduced. Jotting down oxidation numbers can be helpful. Second, almost all of these reactions take place in acid solution. This means that H^+ is almost sure to be a reactant and H_2O is almost sure to be a product. These types of reactions should be visually inspected to make sure all the atoms (especially oxygens) are accounted for on both sides, and that the charge could be balanced. There cannot be only negative ions on the left and only positive ions on the right. Also, check the products for precipitates.

There are four types of redox with oxygen compounds, classified by the oxidizing agent: oxidizing acids (HNO_3 and H_2SO_4), manganese compounds (MnO_4^- and MnO_2), chromium compounds ($\text{Cr}_2\text{O}_7^{2-}$), and hydrogen peroxide (H_2O_2).

Oxidizing acids

Oxidizing acids are strong acids with anions that can be reduced. These include HNO_3 and concentrated H_2SO_4 . They will react with less active metals. In this case both the H^+ and the anion will react.

Dilute nitric acid is a good oxidizing agent and produces NO gas (colorless)

- Solid silver is added to a dilute nitric acid solution.



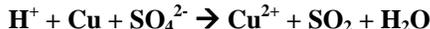
In concentrated nitric acid, brown NO_2 is produced instead.

- Zinc strips are added to a concentrated solution of nitric acid.



Concentrated sulfuric acid is also a good oxidizing agent and produces SO_2 gas:

- A strip of copper metal is added to a concentrated solution of sulfuric acid.

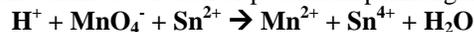


Manganese

Permanganate is MnO_4^- . It is a very good oxidizing agent. Mn has an oxidation number of +7, so it is a good candidate for reduction.

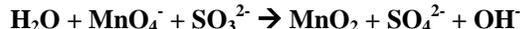
In acid solution, MnO_4^- is reduced to Mn^{2+} ion:

- A solution of tin (II) chloride is added to an acidified solution of potassium permanganate.



In basic solution, MnO_4^- is reduced to MnO_2 (s).

- A solution of potassium permanganate is mixed with an alkaline solution of sodium sulfite.



Manganese dioxide is MnO_2 . With an oxidation number of +4, it is still a good oxidizing agent.

In acid solution, MnO_2 is reduced to Mn^{2+} ion:

- A concentrated solution of hydrochloric acid is added to powdered manganese dioxide and gently heated.



Chromium

There are two chromium oxyanions with chromium in the +6 oxidation state:

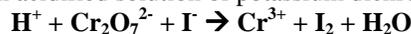
Chromate: CrO_4^{2-} (yellow)

Dichromate: $\text{Cr}_2\text{O}_7^{2-}$ (orange)

We saw chromate earlier when discussing precipitates. It is not likely to appear in redox reactions on the AP exam.

Dichromate is found in redox reactions. It is a very good oxidizing agent, and is always used in acidic solution, where it forms Cr^{3+} :

- A solution of potassium iodide is added to an acidified solution of potassium dichromate.



Hydrogen Peroxide

Hydrogen peroxide, H_2O_2 , has oxygen in the -1 oxidation state. H_2O_2 is an unstable compound, and it decomposes via the following reaction, in which O^{-1} is disproportionated into O^{-2} and O^0 :

- A solution of hydrogen peroxide is heated.



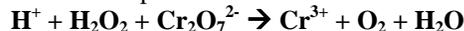
Hydrogen peroxide is a good oxidizing agent. In acid solution, it forms water:

- Hydrogen peroxide is added to an acidified solution of sodium bromide.



Hydrogen peroxide can itself be oxidized in the presence of a strong oxidizing agent (such as permanganate or dichromate). It forms oxygen gas:

- Hydrogen peroxide is added to an acidified solution of potassium dichromate.



Redox Reactions (Change in Oxidation State)

IMPORTANT OXIDIZERS:

OXIDIZER:

MnO_4^- in acid

MnO_2 in acid

MnO_4^- in neutral or basic solution

$\text{Cr}_2\text{O}_7^{2-}$ in acid

HNO_3 concentrated

HNO_3 dilute

H_2SO_4 hot conc.

Metal $-ic$ ions

Frees halogens

Na_2O_2

HClO_4

Formed in the reaction

Mn^{2+}

Mn^{2+}

MnO_2

Cr^{3+}

NO^2

NO

SO_2

metal-ous ions

halide ions

NaOH

Cl^-

IMPORTANT REDUCERS

REDUCERS:

Halide ions

Free metals

Sulfite ions ($\text{SO}_2(\text{aq})$)

Nitrite ions

Free halogens, conc. basic sol'n

Free halogens conc. basic sol'n.

Metal-ous ions

Formed in the reaction:

Free halogen

Metal ions

Sulfate ions

Nitrate ions

Hypohalite ions

Halite ions

Metal-ic ions

Redox: Reactions between an oxidizer and a reducer.

Redox reactions are often recognized by:

...,familiarization with important reducers and oxidizers.

....the clue that there is 'added acid' or the solution is "acidified".

....the use of the supplied reduction potential reference.

Examples:

1. Manganese dioxide is added to concentrated hydrochloric acid and heated.
$$\text{MnO}_2 + \text{H}^+ + \text{Cl}^- \Rightarrow \text{Mn}^{2+} + \text{Cl}_2 + \text{H}_2\text{O}$$
2. A solution of iron (II) nitrate is added to an acidified solution of potassium permanganate.
$$\text{Fe}^{2+} + \text{H}^+ + \text{MnO}_4^- \Rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}_2\text{O}$$
3. Magnesium metal is added to dilute nitric acid. One of the products contains nitrogen with an oxidation number of -3.
$$\text{Mg} + \text{H}^+ + \text{NO}_3^- \Rightarrow \text{Mg}^{2+} + \text{NH}_3 + \text{H}_2\text{O}$$

Redox: Combination reactions.

An oxidizer will react with a reducer of the same element to produce the element at an intermediate oxidation state.

Examples:

1. Solutions of potassium iodide, potassium iodate, and dilute sulfuric acid are mixed.
$$\text{I}^- + \text{IO}_3^- + \text{H}^+ \Rightarrow \text{I}_2 + \text{H}_2\text{O}$$
2. A piece of iron is added to a solution of iron (III) sulfate.
$$\text{Fe} + \text{Fe}^{3+} \Rightarrow \text{Fe}^{2+}$$

Redox: Replacement reactions

A more reactive element (often in the free state) can displace a less reactive element with similar properties from a compound.

Examples:

1. Zinc metal reacts with tin (II) sulfate.
$$\text{Zn} + \text{Sn}^{2+} \Rightarrow \text{Zn}^{2+} + \text{Sn}$$
2. Free chlorine reacts with sodium bromide.
$$\text{Cl}_2 + \text{Br}^- \Rightarrow \text{Cl}^- + \text{Br}_2$$
3. Solid barium peroxide is added to cold sulfuric acid.
$$\text{BaO}_2 + \text{H}^+ + \text{SO}_4^{2-} \Rightarrow \text{BaSO}_4 + \text{H}_2\text{O}_2$$

Redox: Decomposition reactions

Examples:

1. A solution of hydrogen peroxide is catalytically decomposed.
$$\text{H}_2\text{O}_2 \Rightarrow \text{H}_2\text{O} + \text{O}_2$$
2. Chlorates decompose in the presence of heat.
$$\text{KClO}_3 \Rightarrow \text{KCl} + \text{O}_2$$
3. Electrolysis decomposes compounds into their elements.
$$\text{H}_2\text{O} \Rightarrow \text{H}_2 + \text{O}_2$$

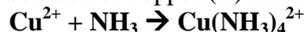
Complexation Reactions

Common Ligands

Ammonia, NH₃

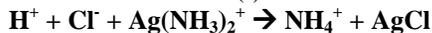
Each type of metal ion has a characteristic coordination number: Ag⁺: 2 (monodentate) ligands; Zn²⁺ and Cu²⁺: 4 ligands; Al³⁺ and Fe³⁺: 6 ligands. Complexes can be formed:

- A concentrated solution of ammonia is added to a solution of copper (II) chloride.



Also, ammonia complexes can be broken by adding acid to form the ammonia ion:

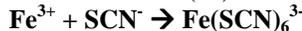
- Dilute hydrochloric acid is added to a solution of diamminesilver (I) nitrate.



Thiocyanate, SCN⁻

Thiocyanate forms a dark red complex with iron, Fe³⁺:

- A solution of ammonium thiocyanate is added to a solution of iron (III) chloride.



Thiosulfate, S₂O₃²⁻

Thiosulfate forms complexes with any metal ions:

- A solution of sodium thiosulfate is added to a precipitate of silver iodide.



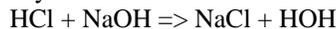
Reactions Involving No Changes in Oxidation States

Displacement reactions.

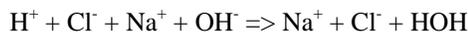
The products can be predicted by exchanging the positive ions of the two reactants.

Example:

Water solutions of hydrochloric acid and sodium hydroxide are mixed.



HCl is a strong acid, NaOH is a strong base, and NaCl is a soluble salt; these should be written as ions.



Substances that do not change ('spectator ions') are not represented in a net ionic equation.



Combination Reactions (acid and base anhydrides)

Examples:

1. Metal oxide + water => a base
 $\text{CaO}(\text{s}) + \text{HOH} \Rightarrow \text{Ca}(\text{OH})_2(\text{soluble hydroxide}) \Rightarrow \text{Ca}^{2+} + 2\text{OH}^-$
2. Nonmetal oxide + water => an acid
 $\text{SO}_2 + \text{HOH} \Rightarrow \text{H}_2\text{SO}_3$ (not a strong acid)
3. Metal oxide + nonmetal oxide = salt
 $\text{CaO} + \text{SO}_2 \Rightarrow \text{CaSO}_3$

Decomposition reactions (the reverse of combination reactions)

Examples:

1. Base => metal oxide + water
 $\text{Ca}(\text{OH})_2 \Rightarrow \text{CaO} + \text{HOH}$
 $\text{Mg}(\text{OH})_2 \Rightarrow \text{MgO} + \text{H}_2\text{O}$
2. Acid containing oxygen => nonmetal oxide + water
 $\text{H}_2\text{CO}_3 \Rightarrow \text{CO}_2 + \text{HOH}$
3. Salt containing oxygen => metal oxide + nonmetal oxide
 $\text{CaCO}_3 \Rightarrow \text{CaO}(\text{s}) + \text{CO}_2$

Reactions of coordination compounds and ions

The ligands (Lewis acids) are bonded to a central atom that is usually a transition metal ion. The most frequently occurring ligands are ammonia, NH_3 , and hydroxide ion OH^- .

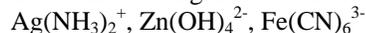
Ammonia



'Excess' hydroxide



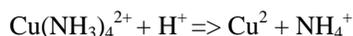
The number of ligands attached to a central metal ion is sometimes twice the oxidation number of the central metal.



The breakup of complex ions is frequently achieved by adding an acid. The products are the metal ion and the species formed when hydrogen ions from the acid react with the ligand (a Lewis base).

Example:

Tetraammine copper II ions are reacted with nitric acid.

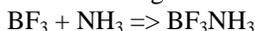


Nonaqueous definitions of acids and bases

- Bronsted reactions involve the transfer of a proton.
- Lewis reactions involve the formation of a coordinate covalent bond.

Example of a Lewis reaction:

The gases boron trifluoride and ammonia are mixed.



Hydrolysis Reactions

Hydrolysis occurs when water reacts with a compound. The water is written as HOH . Combining the ' H^+ ' from the water with negative ion from the other reactant usually gives the formula for one of the products

Example:

- Sodium acetate is dissolved in water.
$$\text{NaC}_2\text{H}_3\text{O}_2 + \text{HOH} \Rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH}$$
- Sodium salts are soluble in water.
 $\text{HC}_2\text{H}_3\text{O}_2$ is not a strong acid and is not ionized.
$$\text{Na}^+ + \text{C}_2\text{H}_3\text{O}_2^- + \text{HOH} \Rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{Na}^+ + \text{OH}^-$$
- Omit the 'spectator ions' to get the net ionic equation.
$$\text{C}_2\text{H}_3\text{O}_2^- + \text{HOH} \Rightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$$

Classifying Reactions to help with Predicting Reactions

There are five different types of reactions possible in the reactions section in the AP Test:

1. [Metathesis Reactions](#); [Redox \(Oxidation - Reduction\) Reactions](#); [Organic Reactions](#); [Complex Ion Formation](#); [Lewis Acid-Base Reactions](#)

Metathesis Reactions

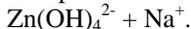
When you see two binary ionic compounds (including acids), the compounds switch partners to form two new compounds. The driving force and product is either a gas, a precipitate, or a weak electrolyte.

a gas	memorize the metathesis products that form gases
a precipitate	memorize the solubility rules
a weak electrolytes	memorize the strong acids so you recognize a weak acid ; memorize H ₂ O and NH ₄ OH as weak electrolytes

Watchout for:

Important stoichiometry...key words "equimolar", etc. for the formation of acid salts like HPO₄²⁻

Complex ion formation through metathesis does not seem to form new compounds. For example: Zn(OH)₂ + *excess* (or *concentrated*) NaOH



RedOx (Oxidation - Reduction) Reactions

Memorize the common strong oxidizers, generally ions with lots of oxygen, MnO₄⁻, Cr₂O₇²⁻, IO₃⁻, etc., **memorize** what they turn into, and look for something to oxidize.

Memorize the common strong reducers (on the handout mentioned above), **memorize** what they turn into, and look for something to reduce.

Memorize the equations for the oxidation and the reduction reactions of water during the electrolysis of water.

□



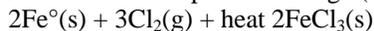
Watchout for:

Keywords "acidified solution" or an acid included in the reactants. The H⁺ ions form H₂O with the oxygens.

Anytime you see a neutral element, Cu⁰, O₂, H₂, etc. it must be redox.

Michael's Tips & Tricks:

1. When reactions occur between a metal like Fe⁰, Cu⁰, Sn⁰, etc. (multiple ions possible metals: Fe²⁺ and Fe³⁺) reacts with reactive gases like O₂, F₂, Cl₂ with the addition of heat, Fe⁰ will ionize itself to maximum positive charge (-ic metal ions). Example:



2. When applying the rule of "Free Halogens + Dilute OH⁻ Hypohalites ions," the addition of halide ions (such as Cl⁻) to the equation are required to obtain full credit on the reaction equations. Example:



3. When the **oxides** of an alkali metal (Family 1), Ca, Ba, or Sr dissolve in water, hydroxides will form, but no gases will be released.

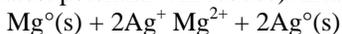
Example:



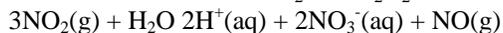
4. When the **hydrides** of an alkali metal (Family 1), Ca, Ba, or Sr dissolve in water, hydroxides will form and H₂ gas is released. Example:



5. Look for "battery" reactions (activity series of metals). When you are not sure which one will undergo changes, look at the reduction potential chart given in the AP Test (the metal with greatest potential will reduce). Examples:



6. Be aware of Disproportionation Redox Reactions. These are the reactions when a portion substance is oxidized while the rest is reduced. The same chemical substance undergoes both oxidation and reduction. NO₂ and H₂O₂ are classic chemicals that have this ability. Example:



7. (Trick #1) When you see electrolysis such as KI in water, and you are stuck on whether H₂ gas will form from which side. The following is the solution that always works. You know potassium is always going to the negative electrode. Let say K⁰ forms at the negative electrode (cathode) and immediately undergoes reaction with water:



Since all these happens on the negative electrode. So, H₂ gas forms from the negative electrode and that's exactly what happens when water reduces at the cathode.

8. (Trick #2) When CuSO₄(aq) is electrolyzed, you know that Cu⁰ metal is going to form because copper's potential is higher than water. So, positive side will attract SO₄²⁻ ions. Nevertheless, SO₄²⁻ can't further oxidize (full of oxygen and no more unshared pair of electrons possible for further oxidation). As the result, you should use the other side of the hydrolysis in Trick #1: instead of OH⁻, put H⁺; instead of H₂, put O₂. This makes sense because when you electrolyze H₂O, you get H₂, and O₂; and OH⁻ and H⁺ from each electrode will neutralize and become water again.



9. (Trick #3) When you recognize great oxidizers like Cr₂O₇²⁻, MnO₄⁻, and MnO₂, with acidified solutions in the test, but you may not remember what they turn into. So, remember the definition of an oxidizer: the ability to give off its oxygen components. So, acidified Cr₂O₇²⁻, turn into Cr³⁺ And MnO₂, and MnO₄⁻ will turn to Mn²⁺. As you can see, the oxygen components are "ripped off" and form water with H⁺ ions.

Organic Reactions

These are especially useful for the **older** tests. The "acorn book" states that organic compounds may show up as examples, but organic reactions are pretty restricted.

Oxidation - complete combustion of hydrocarbons results in CO_2 and H_2O . Milder oxidation may change a primary alcohol into an aldehyde, a secondary alcohol into a ketone, etc.

Substitution - example: Br_2 + a hydrocarbon results in HBr + a Br substitute for one of the hydrogens on a carbon.

Esterification - organic acid + alcohol results in an ester and H_2O

Addition - a double bond is broken and atoms are attached to two carbons. If H_2O is added, an H and OH add to the carbons. If Br_2 is added, a Br atom is added to each of the carbons.

Complex Ion Formation

These are usually formed from a transition metal surrounded by ligands (polar molecules or negative ions). As a "rule of thumb" you place twice the number of ligands around an ion as the charge on the ion... example: the dark blue $\text{Cu}(\text{NH}_3)_4^{2+}$ (ammonia is used as a test for Cu^{2+} ions), and $\text{Ag}(\text{NH}_3)_2^+$.

Memorize the common ligands.

Ligands	Names used in the ion
H_2O	aqua
NH_3	ammine
CO	carbonyl
NO	nitrosyl
OH^-	hydroxy
Cl^-	chloro
Br^-	bromo
CN^-	cyano
O^{2-}	oxo
CO_3^{2-}	carbonato
$\text{S}_2\text{O}_3^{2-}$	thiosulfato
SCN^-	thiocyanato (bonded through sulphur) isothiocyanato (bonded through nitrogen)
$\text{C}_2\text{O}_4^{2-}$	oxalato
NO_2^-	nitrito/nitro*

*used as "nitro" when NO_2^- is bonded through nitrogen: ONO^-

Watchout for:

Aluminum also forms complex ions as do some post transition metals. Ex: $\text{Al}(\text{H}_2\text{O})_6^{3+}$

The names are very impressive, but easy...the ions above are the tetraamminecopper(II) ion, the diamminesilver(I) ion, and hexaquoaluminum(III) ion. $\text{Zn}(\text{OH})_4^{2-}$ is the tetrahydroxyzinc(II) ion, the charge is the sum of the parts $(2+)+4(-1)=-2$.

Acid-base reactions may change NH_3 into NH_4^+ (or vice versa) which will alter its ability to act as a ligand.

Visually, a precipitate may go back into solution as a complex ion is formed. For example, Cu^{2+} + a little NH_4OH will form the light blue precipitate, $\text{Cu}(\text{OH})_2$. With excess ammonia, the complex, $\text{Cu}(\text{NH}_3)_4^{2+}$, forms.

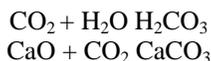
Keywords such as "excess" and "concentrated" of any solution **may** indicate complex ions. $\text{AgNO}_3 + \text{HCl}$ forms the white precipitate, AgCl . With excess, concentrated HCl , the complex ion, AgCl_2^- , forms and the solution clears.

The odd complex ion, FeSCN^{2+} , shows up once in a while simply because it is commonly used in the CHEMStudy first-year equilibrium lab. Transitional metals, such as Iron, Zinc and Chromium, can form complex ions. Aluminum can form complex ions as well.

Lewis Acid-Base Reactions

"HAVE PAIR WILL SHARE" -- Lewis Base. The formation of a coordinate covalent bond between NH_3 and BF_3 to form $\text{H}_3\text{N}:\text{BF}_3$ is a classic example of a Lewis acid-base reaction.

Oxides of metals and nonmetals are also examples:



Watchout for:

Some reactions come from industrial processes such as the formation of bleach (ClO^- ion) from Cl_2 and dilute NaOH . Note these as you come across them and add them to your "bag of tricks".

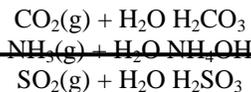
You can write (or think about) chemicals in different ways for different reactions. Ammonia may be $\text{NH}_3(\text{aq})$ for complex ions, NH_4OH for metathesis or acid-base reactions; water may be H_2O or you might think of it as H^+ and OH^- for hydrolysis or redox reactions; HNO_3 may be an acid (donating H^+) or it may be an oxidizer (forming NO or $\text{NO}_2 + \text{H}_2\text{O}$).

Tips & Tricks:

1. When you see NO_3^- , NO_2^- , CO_3^{2-} , SO_3^{2-} , SO_4^{2-} ions in solids and are heated in vacuum, $\text{NO}_2(\text{g})$, $\text{NO}(\text{g})$, $\text{CO}_2(\text{g})$, $\text{SO}_2(\text{g})$, $\text{SO}_3(\text{g})$ are evolved. Example:



2. (Trick #4) When you see a compounded gas bubble through any solution, they should be changed to "The Dissolved" state which is actually a Lewis Acid-Base Reaction:

**Solubility Rules (another version)**

The dividing line between soluble and insoluble is 0.1-molar at 25 °C.

Any substance that can form 0.1 M or more concentrated is soluble. Any substance that fails to reach 0.1 M is defined to be insoluble.

This value was picked with a purpose. VERY FEW substances have their maximum solubility near to 0.1 M. Almost every substance of any importance in chemistry is either much MORE soluble or much LESS soluble.

In the past, some teachers would have a third category: slightly soluble. For the most part, that category has been cast aside.

Solubility rules that apply to water solution:

- (1) All alkali metal (lithium, sodium, potassium, rubidium, and cesium) and ammonium compounds are soluble.
- (2) All acetate, perchlorate, chlorate, and nitrate compounds are soluble.
- (3) Silver, lead, and mercury (I) compounds are insoluble.
- (4) Chlorides, bromides, and iodides are soluble
- (5) Carbonates, hydroxides, oxides, phosphates, silicates, and sulfides are insoluble. (6) Sulfates are soluble except for calcium and barium.

These rules are to be applied in the order given. For example, PbSO_4 is insoluble because rule 3 comes before rule 6. In like manner, AgCl is insoluble because rule 3 (the smaller) takes precedence over rule 4 (the larger).

Please be aware that these rules are guidelines only. For example, there are some alkali metal compounds that are insoluble. However, these are rather exotic compounds and can be safely ignored at an introductory level.

Solubility Rules (Another Version)

1. All common compounds of Group I and ammonium ions are soluble.
2. All nitrates, acetates, and chlorates are soluble.
3. All binary compounds of the halogens (other than F) with metals are soluble, except those of Ag, Hg(I), and Pb. Pb halides are soluble in hot water.)
4. All sulfates are soluble, except those of barium, strontium, calcium, lead, silver, and mercury (I). The latter three are slightly soluble.
5. Except for rule 1, carbonates, hydroxides, oxides, silicates, and phosphates are insoluble.
6. Sulfides are insoluble except for calcium, barium, strontium, magnesium, sodium, potassium, and ammonium.