

Dougherty Valley HS • AP Chemistry  
Predicting Reactions  
CLASSIFYING REACTIONS

[Keep for Reference]

Look for these five TYPES of reactions:

- Double Replacement
- Redox
- Organic
- Complex Ion
- Lewis Acid-Base

### Double Replacement Reactions

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

a gas	memorize the double replacement products that form <b>gases</b>
a precipitate	memorize the <b>solubility rules</b>
a weak electrolyte	memorize the <b>strong acids</b> so you recognize a weak acid; memorize H <sub>2</sub> O and NH <sub>4</sub> OH as weak electrolytes

*Watch for:*

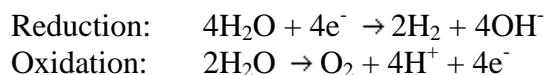
- Impt. stoichiometry key words like "equimolar", for the formation of acid salts like HPO<sub>4</sub><sup>2-</sup>
- Complex ion formation when double replacement does not seem to form new compounds. Example: Zn(OH)<sub>2</sub> + excess NaOH → Zn(OH)<sub>4</sub><sup>2-</sup>  
Key words like "excess" or "concentrated"

### RedOx (Oxidation-Reduction) Reactions

Memorize the common **strong oxidizers**, generally ions with lots of oxygens, MnO<sub>4</sub><sup>-</sup>, Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>, IO<sub>3</sub><sup>-</sup>, etc. (in the "STRONG OXIDIZERS" section on "Stuff I Should Know for the AP TEST"), **memorize** what they turn into, and look for something to oxidize.

Memorize the common **strong reducers**, **memorize** what they turn into, and look for something to reduce.

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

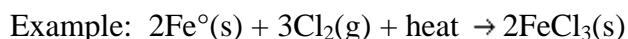


*Watch for:*

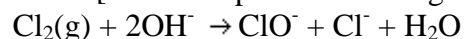
- Key words "acidified solution" or an acid included in the reactants. The H<sup>+</sup> ions form H<sub>2</sub>O with the oxygens in the oxidizer.
- Neutral elements, Cu<sup>0</sup>, O<sub>2</sub>, H<sub>2</sub>, etc. must be redox reactions and often change into common ions.

*Tips & Tricks:*

1. When reactions occur between a metal with multiple valences like Fe<sup>0</sup> (Fe<sup>2+</sup> & Fe<sup>3+</sup>), Cu<sup>0</sup>, or Sn<sup>0</sup> and gases like O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub> the ions usually form oxidize to the "-ic" ion.



2. When you identify an oxidation product, make certain you also have a reduction product. Ex: "Free halogens + dilute OH<sup>-</sup> → hypohalite ions," the halide ions (such as Cl<sup>-</sup>) as a product are required for full credit. [industrial process: making bleach]



3. When the **oxides** of an alkali metal (Family 1), Ca, Ba, or Sr dissolve in water, hydroxides form, but no gases will be released. [This is **not** redox.]  
Example:  $\text{K}_2\text{O}(\text{s}) + \text{H}_2\text{O} \rightarrow 2\text{K}^+ + 2\text{OH}^-$

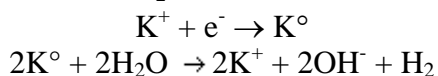
4. When the **hydrides** of an alkali metal (Family 1), Ca, Ba, or Sr dissolve in water, hydroxides will form and H<sub>2</sub> gas is released. [This **is** redox.]  
Example:  $\text{LiH}(\text{s}) + \text{H}_2\text{O} \rightarrow \text{Li}^+ + \text{OH}^- + \text{H}_2(\text{g})$

5. Look for "battery" (activity series) reactions. You can use the reduction potential chart on the AP exam; the metal with greatest potential will reduce.  
Example:  $\text{Mg}^0(\text{s}) + 2\text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag}^0(\text{s})$

6. Be aware of **disproportionation** reactions. These are the reactions where the same chemical substance undergoes both oxidation and reduction. NO<sub>2</sub> and H<sub>2</sub>O<sub>2</sub> are classic examples:  
 $3\text{NO}_2(\text{g}) + \text{H}_2\text{O} \rightarrow 2\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{NO}(\text{g})$

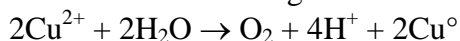
7. (Trick #1) During electrolysis of salts such as KI(aq), remember that K<sup>0</sup> will not form in water; the water is reduced. If you don't remember the reaction, **pretend** that K<sup>0</sup> is formed at the negative

electrode (where  $K^+$  is attracted) and have it react with water to form  $H_2$  and  $OH^-$ .



The combination of these two reactions is exactly what happens when water is reduced at the cathode.

8. (Trick #2) When  $CuSO_4(aq)$  is electrolyzed, you know that  $Cu^\circ$  metal is going to form because copper's potential is higher than water. So, positive electrode will attract  $SO_4^{2-}$  ions but  $SO_4^{2-}$  can not further oxidize (full of oxygen and no more unshared pair of electrons possible for further oxidation). As a result, you should use the other side of the hydrolysis in Trick #1: instead of  $OH^-$ , put  $H^+$ ; instead of  $H_2$ , put  $O_2$ . This makes sense because when you electrolyze  $H_2O$ , you get  $H_2$ , and  $O_2$ ; and  $OH^-$  and  $H^+$  from each electrode will neutralize and become water again.



9. (Trick #3) You may recognize acidified solutions of strong oxidizers like  $Cr_2O_7^{2-}$ ,  $MnO_4^-$ , and  $MnO_2$ , but may not remember what they turn into.

Remember the definition of an oxidizer: the ability to give off its oxygen components. Acidified  $Cr_2O_7^{2-}$ , turns into  $Cr^{3+}$  and  $MnO_2$ , and  $MnO_4^-$  will turn to  $Mn^{2+}$ . The oxygen components are lost and form water with  $H^+$  ions.

## Organic Reactions

These are especially useful for the older tests. The AP syllabus states that organic compounds may show up as examples, but organic reactions are pretty restricted, usually just *combustion*.

- **Oxidation** - complete combustion of hydrocarbons results in  $CO_2$  and  $H_2O$ . Milder oxidation may change a primary alcohol into an aldehydes and a secondary alcohol into a ketone.
- **Substitution** - example:  $Br_2$  + a hydrocarbon results in  $HBr$  & a Br substitutes for one of the hydrogens on a carbon.
- **Esterfication** - organic acid + alcohol results in an ester and  $H_2O$
- **Addition** - a double bond is broken and atoms are attached to the two carbons. If  $H_2O$  is added, an H and OH add to the carbons (alcohol). 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* atoms 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...

Examples: dark blue  $Cu(NH_3)_4^{2+}$  and  $Ag(NH_3)_2^+$ . (Note: Ammonia is used as a test for  $Cu^{2+}$  ions)

Memorize the common ligands.

Ligands	Names used in the ion
$H_2O$	aqua
$NH_3$	ammine
$OH^-$	hydroxy
$Cl^-$	chloro
$Br^-$	bromo
$CN^-$	cyano
$SCN^-$	thiocyanato (bonded through sulphur) isothiocyanato (bonded through nitrogen)

### Watch for:

Aluminum also forms complex ions as do some post transitions metals. Ex:  $AlF_6^{3+}$

The names are impressive, but easy..the ions above are the tetraamminecopper(II) ion, the diamminesilver(I) ion, and hexafluoroaluminum(III) ion.  $Zn(OH)_4^{2-}$  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) and 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 ammonia ( $NH_4OH$ ) will form the light blue precipitate,  $Cu(OH)_2$ . With excess ammonia, the complex,  $Cu(NH_3)_4^{2+}$ , forms.

Keywords such as "excess" and "concentrated" of any solution *may* indicate complex ions.  $AgNO_3$  +  $HCl$  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 because it is commonly used in a common first-year equilibrium lab.

## Lewis Acid-Base Reactions

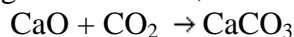
"HAVE PAIR WILL SHARE" -- Lewis Base.

The formation of a coordinate covalent bond between  $\text{NH}_3$  and  $\text{BF}_3$  to form  $\text{H}_3\text{N}:\text{BF}_3$  is a classic example of a Lewis acid-base reaction.

Oxides of nonmetals (acid anhydrides) and oxides of metals (base anhydrides) are also examples:



When you get one of each, the result is a salt:



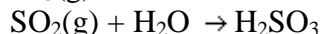
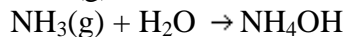
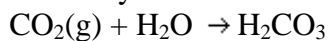
Aluminum ions (and  $\text{Fe}^{3+}$  ions) make a solution acidic because they are Lewis acids and draw electrons away from the O in  $\text{H}_2\text{O}$ , making the O-H bond more polar and increasing the leaving power of  $\text{H}^+$ .

### Watch for:

1. When you see  $\text{NO}_3^-$ ,  $\text{NO}_2^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{SO}_3^{2-}$ ,  $\text{SO}_4^{2-}$  ions in solids that are heated in a vacuum, this is *not* combustion (burning in air is combustion). The gases:  $\text{NO}_2$ ,  $\text{NO}$ ,  $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{SO}_3$  are evolved.

Example:  $\text{CaCO}_3(\text{s}) + \text{heat} \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$   
This is how they make lime (CaO) for cement and is the opposite of the reaction at the top of this

2. (Trick #4) When you see a gaseous compound bubble through any solution, change the gas into the Lewis acid or base they become:



Some reactions involve *industrial processes* such as the formation of lime (above) or making bleach (sodium hypochlorite) ( $\text{ClO}^-$  or  $\text{OCl}^-$  ion) from  $\text{Cl}_2$  and dilute  $\text{NaOH}$  or the reduction of iron using heat and hydrogen gas.

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 or  $\text{NH}_4\text{OH}$  for double replacement or acid-base reactions.
- Water may be  $\text{H}_2\text{O}$  or you might think of it as  $\text{H}^+$  and  $\text{OH}^-$  for hydrolysis or redox reactions.
- $\text{HNO}_3$  may be an acid (donating  $\text{H}^+$ ) or it may be an oxidizer (forming  $\text{NO}$  or  $\text{NO}_2 + \text{H}_2\text{O}$ ).
- In the same way, dilute  $\text{H}_2\text{SO}_4$  is a source of  $\text{H}^+$  and  $\text{SO}_4^{2-}$  ions whereas concentrated  $\text{H}_2\text{SO}_4$  is an oxidizer forming  $\text{SO}_2 + \text{H}_2\text{O}$ .

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### Some other tips and final touches:

1. This all may seem like too much, but remember, you only need to recognize five out of the eight reactions on any given exam.
2. You earn a point for just writing the reactants in chemical form.
3. Don't forget to cancel out spectators.
4. Get familiar with the chemicals: what are they going to do, function as, how will they change.
5. Make sure you are very clear with which type of acid is being described: Hydrochloric acid ( $\text{HCl}$ ), Chloric acid ( $\text{HClO}_3$ ), and Chlorous Acid ( $\text{HClO}_2$ ) all sound similar. Misunderstanding of the names will ruin your chance of getting partial credit points.
6. Read very carefully.
7. Do lots of practice problems and be careful. (<http://www.chemmybear.com/preactions.html>)
8. If it were easy, it wouldn't be AP.
9. Be confident and proud! You have made it this far and you are going to make it thru the test.

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This handout is revised from a group project by Michael Lo, Ruth Cusick, & Mr. Groves (1998) and revised in 2003 and 2005.