

Predicting Reactions

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

Watch for:

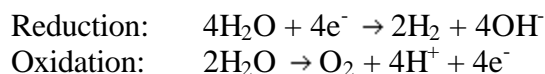
- Impt. stoichiometry key words like "equimolar", for the formation of acid salts like HPO₄²⁻
- Complex ion formation when double replacement does not seem to form new compounds. Example: Zn(OH)₂ + excess NaOH → Zn(OH)₄²⁻
Key words like "excess" or "concentrated"

RedOx (Oxidation-Reduction) Reactions

Memorize the common **strong oxidizers**, generally ions with lots of oxygens, MnO₄⁻, Cr₂O₇²⁻, IO₃⁻, 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.



CLASSIFYING REACTIONS

Watch for:

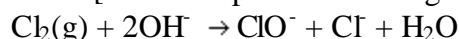
- Key words "acidified solution" or an acid included in the reactants. The H⁺ ions form H₂O with the oxygens in the oxidizer.
- Neutral elements, Cu⁰, O₂, H₂, etc. must be redox reactions and often change into common ions.

Tips & Tricks:

1. When reactions occur between a metal with multiple valences like Fe⁰ (Fe²⁺ & Fe³⁺), Cu⁰, or Sn⁰ and gases like O₂, F₂, Cl₂ 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⁻ → hypohalite ions," the halide ions (such as Cl⁻) 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: K₂O(s) + H₂O → 2K⁺ + 2OH⁻

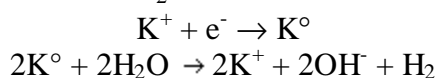
4. When the **hydrides** of an alkali metal (Family 1), Ca, Ba, or Sr dissolve in water, hydroxides will form and H₂ gas is released. [This **is** redox.]
Example: LiH(s) + H₂O → Li⁺ + OH⁻ + H₂(g)

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: Mg⁰(s) + 2Ag⁺ → Mg²⁺ + 2Ag⁰(s)

6. Be aware of **disproportionation** reactions. These are the reactions where the same chemical substance undergoes both oxidation and reduction. NO₂ and H₂O₂ are classic examples:
3NO₂(g) + H₂O → 2H⁺(aq) + 2NO₃⁻(aq) + NO(g)

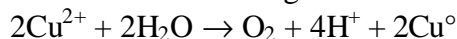
7. (Trick #1) During electrolysis of salts such as KI(aq), remember that K⁰ will not form in water; the water is reduced. If you don't remember the reaction, **pretend** that K⁰ is formed at the negative

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° 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

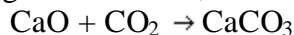
"HAVE PAIR WILL SHARE" -- Lewis Base.

The formation of a coordinate covalent bond between 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 nonmetals (acid anhydrides) and oxides of metals (base anhydrides) are also examples:



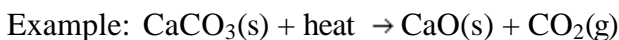
When you get one of each, the result is a salt:



Aluminum ions (and Fe^{3+} ions) make a solution acidic because they are Lewis acids and draw electrons away from the O in H_2O , making the O-H bond more polar and increasing the leaving power of H^+ .

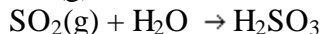
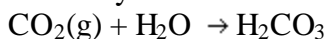
Watch for:

1. When you see NO_3^- , NO_2^- , CO_3^{2-} , SO_3^{2-} , SO_4^{2-} ions in solids that are heated in a vacuum, this is *not* combustion (burning in air is combustion). The gases: NO_2 , NO , CO_2 , SO_2 , SO_3 are evolved.



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) (ClO^- or OCl^- ion) from Cl_2 and dilute 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 NH_3 (aq) for complex ions or NH_4OH for double replacement 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 $\text{NO}_2 + \text{H}_2\text{O}$).
- In the same way, dilute H_2SO_4 is a source of H^+ and SO_4^{2-} ions whereas concentrated H_2SO_4 is an oxidizer forming $\text{SO}_2 + \text{H}_2\text{O}$.

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 (HCl), Chloric acid (HClO_3), and Chlorous Acid (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.

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