

AP Chemistry Summer Assignment 2018-2019

The AP Chemistry summer assignment is completely online! We will be using an online program called “**Mastering Chemistry**”. Once you are enrolled in the online course (via blackboard), you will be responsible for completing the **5 part summer assignment**.

NOTE: You must have a **correct email address** in *blackboard*. Information on how to login to the website will be emailed to you. Go under the “My FCPS” tab and click on “personal information” in the tools section if you need to edit your email address on blackboard.

1. Summer Assignment Part 1 – Introduction to MasteringChemistry

- This is an optional assignment. This will allow you to get familiar with the online program.

2. Summer Assignment Part 2 – Chapter 01 Practice Problems

- Note that the tutorials are not worth any points. These are optional if you need to learn the items again.
- You **SHOULD** be able to answer all problems based on your knowledge from your first year of chemistry...the tutorials will help you get familiar with everything.
- You will have 6 opportunities to complete all mandatory questions correctly before you cannot receive credit.
- You may use hints and you will not be penalized!

3. Summer Assignment Part 3 – Chapter 02 Practice Problems

- Same guidelines as part 2

4. Summer Assignment Part 4 – Chapter 03 Practice Problems

- Same guidelines as part 2

5. Summer Assignment Part 5 – Multiple Choice Test (Chapters 1 – 3)

- Once you complete parts 2, 3 and 4 you will take this test.
- The test is 50 questions multiple choice.
- You are allowed 2 attempts to get the correct answer before you lose credit.
- No hints

6. Additional Packet

- a. Please commit to memory the pages that say **MEMORIZE** this (multiple oxidation metals, polyatomic ions, strong acids/bases, and solubility rules).
- b. Look over the reactions and color guide. Since there are many reactions and we will be learning reactions in the first unit, it is in your best interest to have a solid grasp from day one!

7. Textbook

- a. This year in AP Chemistry we will be using “*Chemistry: The Central Science*” *Twelfth Edition, AP Edition*. By Brown, LeMay, Bursten, Murphy, Woodward.
- b. You will have online access to this textbook through the MasteringChemistry course.
- c. You can pick up a **hard copy** of the text from **Ms. Saccomando in Room 301** or **Ms. Malik** starting on June 15th (please note that a hard copy is not necessary to have to complete the summer assignment).

You have until August 27th at 11:59pm to complete this assignment! If you have any questions, please email Ms. Saccomando at ajsaccomando@fcps.edu or Ms. Malik at mamalik@fcps.edu .

Ionic Compounds (MEMORIZE THIS PAGE!)

Common Metal Ions with More than One Ionic Charge (MOM)

| Symbol | Stock Name | Classical Name |
|------------------|--------------------|----------------|
| Cu^+ | Copper(I) ion | Cuprous ion |
| Cu^{2+} | Copper(II) ion | Cupric ion |
| Fe^{2+} | Iron(II) ion | Ferrous ion |
| Fe^{3+} | Iron (III) ion | Ferric ion |
| Hg^{2+} | Mercury(II) ion | Mercuric ion |
| Pb^{2+} | Lead(II) ion | Plumbous ion |
| Pb^{4+} | Lead(IV) ion | Plumbic ion |
| Sn^{2+} | Tin(II) ion | Stannous ion |
| Sn^{4+} | Tin(IV) ion | Stannic ion |
| Cr^{2+} | Chromium(II) ion | Chromous ion |
| Cr^{3+} | Chromium(III) ion | Chromic ion |
| Mn^{2+} | Manganese(II) ion | Manganous ion |
| Mn^{3+} | Manganese(III) ion | Manganic ion |
| Co^{2+} | Cobalt(II) ion | Cobaltous ion |
| Co^{3+} | Cobalt(III) ion | Cobaltic ion |

***Note Silver ion and Zinc ion and Cadmium ion are all transition metal ions, but they only have 1 charge (1 oxidation state) so they do NOT need a roman numeral. (Ag^+ and Zn^{2+} and Cd^{2+})

Common Polyatomic Ions

| Name | Formula | Name | Formula |
|--------------------|------------------------|----------------------|------------------------------------|
| Ammonium | NH_4^+ | Iodate | IO_3^- |
| Nitrite | NO_2^- | Periodate | IO_4^- |
| Nitrate | NO_3^- | Acetate | $\text{C}_2\text{H}_3\text{O}_2^-$ |
| Hydrogen Sulfate | HSO_4^- | Dihydrogen Phosphate | H_2PO_4^- |
| Hydroxide | OH^- | Carbonate | CO_3^{2-} |
| Cyanide | CN^- | Sulfite | SO_3^{2-} |
| Permanganate | MnO_4^- | Sulfate | SO_4^{2-} |
| Hydrogen Carbonate | HCO_3^- | Thiosulfate | $\text{S}_2\text{O}_3^{2-}$ |
| Hypochlorite | ClO^- | Peroxide | O_2^{2-} |
| Chlorite | ClO_2^- | Chromate | CrO_4^{2-} |
| Chlorate | ClO_3^- | Dichromate | $\text{Cr}_2\text{O}_7^{2-}$ |
| Perchlorate | ClO_4^- | Hydrogen Phosphate | HPO_4^{2-} |
| Bromate | BrO_3^- | Phosphite | PO_3^{3-} |
| Hydronium | H_3O^+ | Phosphate | PO_4^{3-} |

The old way to say “hydrogen carbonate” was “bicarbonate”...if you see the prefix “bi” in front of a polyatomic ion name, then you are adding an H^+ to the formula...note that when you add an H^+ to the formula the net charge of the ion goes down one...Carbonate = CO_3^{2-} ...Bicarbonate = HCO_3^{1-} Bisulfate = HSO_4^-

MEMORIZE this page as well

TABLE OF STRONG ACIDS

TABLE OF STRONG BASES

Completely Ionized in Water to Give One (or more) Protons **Complete ionize to give one (or more) Hydroxides**

| | |
|-------------------------|---|
| HI | $\text{H}^+(\text{aq}) + \text{I}^-(\text{aq})$ |
| HBr | $\text{H}^+(\text{aq}) + \text{Br}^-(\text{aq})$ |
| HClO_4 | $\text{H}^+(\text{aq}) + \text{ClO}_4^-(\text{aq})$ |
| HCl | $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ |
| HClO_3 | $\text{H}^+(\text{aq}) + \text{ClO}_3^-(\text{aq})$ |
| H_2SO_4 | $\text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq})$ (HSO_4^- is a weak acid) |
| HNO_3 | $\text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$ |

| | |
|--------------------------|---|
| NaOH | $\text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| KOH | $\text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| LiOH | $\text{Li}^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| RbOH | $\text{Rb}^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| CsOH | $\text{Cs}^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| $\text{Ca}(\text{OH})_2$ | $\text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$ |
| $\text{Ba}(\text{OH})_2$ | $\text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$ |
| $\text{Sr}(\text{OH})_2$ | $\text{Sr}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$ |

SOLUBILITY RULES:

- 1. Alkali metal salts AND NH_4^+ salts ARE *soluble***
 - 2. NO_3^- , ClO_3^- , ClO_4^- , and CH_3COO^- are *soluble***
 3. Cl^- , Br^- , I^- are *soluble*, *except for Ag^+ , Hg_2^{+2} , Pb^{+2}
 3. F^- are *soluble*, *except for Group II metals
 4. SO_4^{-2} are *soluble*, *except for Ca^{+2} , Sr^{+2} , Ba^{+2} , Ag^+ , Pb^{+2} , Hg_2^{+2}
 5. CO_3^{-2} , PO_4^{-3} , $\text{C}_2\text{O}_4^{-2}$, CrO_4^{-2} , S^{-2} , OH^- , and O^{-2} are **INSOLUBLE**
- **unless it is group I (rule 1 takes priority!)****

It can be assumed that ionic compounds that dissolve in water are strong electrolytes and are therefore soluble.

(This is what we will be learning in September...)

Synthesis reactions $A + B \rightarrow AB$

1. A metal combines with a nonmetal to form a binary salt (IONIC COMPOUND)

Ex. A piece of magnesium ribbon is placed in a container with oxygen gas.



Net Ionic: $2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$...*Note that there is no water so the ionic compound will not be aqueous, so it will not be dissociated!!!!

2. Metallic Oxides (Ionic compounds with oxygen as the anion) added to water will form bases (Ionic compounds with Hydroxide as the anion)

Ex. Solid barium oxide is added to water.

Equation: $\text{BaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ba}(\text{OH})_2(aq)$ *Note that the compound was added to water so the product is aqueous

Net Ionic: $\text{BaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ba}^{2+} + 2\text{OH}^-$ *Note that barium hydroxide is a strong base so it will be written dissociated in the net ionic equation!

3. Nonmetallic oxides (Covalent Compounds with oxygen) added to water will form acids. The nonmetal attached to oxygen in the nonmetallic oxide will keep its oxidation number.

Ex. Carbon dioxide is bubbled through water.

Equation: $\text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{CO}_3(aq)$ *Note that the oxidation number for carbon is still +4, and that Carbonic acid is aqueous because it is in water.

Net Ionic: $\text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{CO}_3(aq)$ *Note that carbonic acid is NOT dissociated because it is a WEAK acid

4. Metallic oxides combined with nonmetallic oxides will form Ionic compounds with polyatomic ions

Ex. Crystals of sodium oxide are added to a container with sulfur dioxide gas.

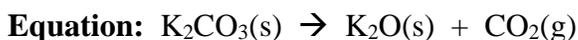
Equation: $\text{Na}_2\text{O}(s) + \text{SO}_2(g) \rightarrow \text{Na}_2\text{SO}_3(s)$ *Note that the sulfur maintains its oxidation number...this is why the product is not Na_2SO_4 . Also note that the product is a solid because there is no water...so it CANNOT BE AQUEOUS!

Net Ionic: $\text{Na}_2\text{O}(s) + \text{SO}_2(g) \rightarrow \text{Na}_2\text{SO}_3(s)$ *Note that nothing dissociates when there is no aqueous species

Decomposition Reactions $AB \rightarrow A + B$ (compounds generally start out as a solid)

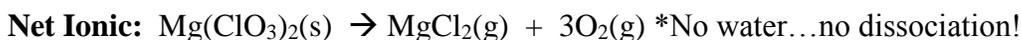
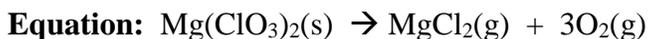
1. Metallic Carbonates (Ionic compounds with carbonate) decompose into metallic oxides and carbon dioxide.

Ex. Potassium carbonate is heated



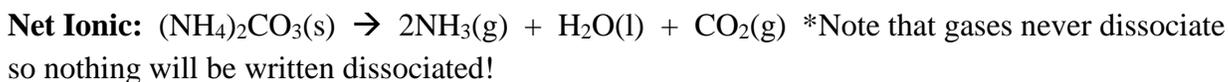
2. Metallic Chlorates (Ionic compounds with chlorate) decompose into metallic chlorides and oxygen

Ex. A sample of magnesium chlorate is heated.



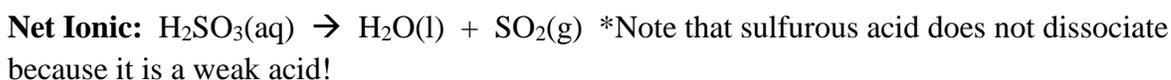
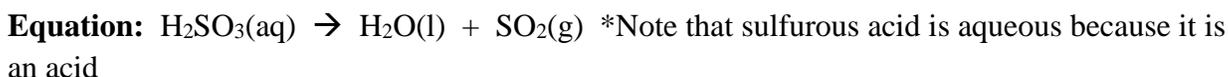
3. Ammonium Carbonate decomposes into ammonia, water and carbon dioxide.

Ex. A sample of ammonium carbonate is heated



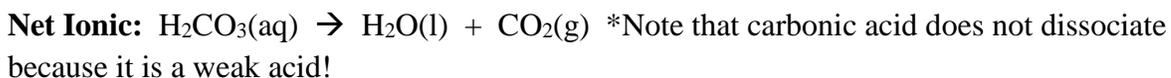
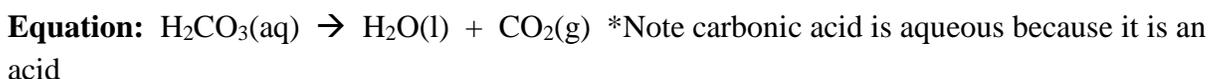
4. Sulfurous acid decomposes into sulfur dioxide and water

Ex. A sample of sulfurous acid is heated.



5. Carbonic acid decomposes into carbon dioxide and water.

Ex. A solution of carbonic acid is heated



6. A binary compound (Ionic compounds made of monatomic ions) may break down to produce two elements...usually starts with a molten (liquid) ionic compound being electrolyzed

Ex. Molten aluminum chloride is electrolyzed.

Equation: $2\text{AlCl}_3(\text{l}) \rightarrow 2\text{Al}(\text{s}) + 3\text{Cl}_2(\text{g})$ *Note this will be the only time you have a liquid ionic compound

Net Ionic: $2\text{AlCl}_3(\text{l}) \rightarrow 2\text{Al}(\text{s}) + 3\text{Cl}_2(\text{g})$

7. Hydrogen peroxide decomposes into water and oxygen.

Ex. A sample of hydrogen peroxide is heated

Equation: $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ *Note hydrogen peroxide will always be in solution so it will be aqueous

Net Ionic: $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ *Hydrogen peroxide does NOT dissociate

8. Ammonium Hydroxide decomposes into ammonia and water

Ex. Ammonium hydroxide is heated.

Equation: $\text{NH}_4\text{OH}(\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})$ *Note that ammonia hydroxide is always in an aqueous solution

Net Ionic: $\text{NH}_4\text{OH}(\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})$ *Note that ammonia hydroxide does not dissociate because it acts as weak conjugate acid!

Single Replacement Reactions ($\text{A} + \text{BC} \rightarrow \text{B} + \text{AC}$)

1. Active metals will replace less active metals from their compounds in an aqueous solution.

Ex. Calcium deposits are added to a solution of Iron (III) chloride

Equation: $3\text{Ca}(\text{s}) + 2\text{FeCl}_3(\text{aq}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{MgCl}_2(\text{aq})$

Net Ionic: $3\text{Ca}(\text{s}) + 2\text{Fe}^{3+} \rightarrow 2\text{Fe}(\text{s}) + 3\text{Mg}^{2+}$ *Note that the metal that gets replaced is now in its normal elemental form with oxidation state of zero

2. Active metals will replace the hydrogen in water...always producing a base and hydrogen gas

Ex. Lithium metal is added to water.

Equation: $2\text{Li}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{LiOH}(\text{aq}) + \text{H}_2(\text{g})$ *Note that LiOH is aqueous because the lithium was added to water so the products are in water!

Net Ionic: $2\text{Li}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{Li}^+ + 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$ *LiOH is dissociated because it is a strong base!

3. Active metals will replace hydrogen in acids to produce an aqueous ionic compound and hydrogen gas

Ex. Sodium is placed in a solution of nitric acid.

Equation: $2\text{Na(s)} + 2\text{HNO}_3\text{(aq)} \rightarrow 2\text{NaNO}_3\text{(aq)} + \text{H}_2\text{(g)}$ *Nitric acid is aqueous because it is an acid.

Net Ionic: $2\text{Na(s)} + 2\text{H}^+ \rightarrow 2\text{Na}^+ + \text{H}_2\text{(g)}$ *Nitric acid dissociates because it is a strong acid, Sodium nitrate dissociates because it is soluble. Nitrate is not included because it is a spectator ion

4. Active nonmetals (mostly halogens) will replace less active nonmetals (mostly halogens) from their compounds in aqueous solution.

Ex. Fluorine gas is bubbled through a solution of rubidium iodide..

Equation: $\text{F}_2\text{(g)} + 2\text{RbI(aq)} \rightarrow 2\text{RbF(aq)} + \text{I}_2\text{(s)}$ *Iodine is a solid at standard conditions

Net Ionic: $\text{F}_2\text{(g)} + 2\text{I}^- \rightarrow 2\text{F}^- + \text{I}_2\text{(s)}$ *Rubidium Iodide and rubidium fluoride are both dissociated because they are soluble.

Double Replacement Reactions (Formation of Precipitate)

1. A precipitate is an insoluble substance (solid) formed by the reaction of two aqueous substances (most of the time aqueous ionic compounds). It is the result of ions bonding together so strongly that the solvent (water) cannot pull them apart.

Ex. Potassium bromide is added to a solution of silver nitrate

Equation: $\text{KBr(aq)} + \text{AgNO}_3\text{(aq)} \rightarrow \text{AgBr(s)} + \text{KNO}_3\text{(aq)}$ *Silver Bromide is insoluble

Net Ionic: $\text{Br}^- + \text{Ag}^+ \rightarrow \text{AgBr(s)}$ *Potassium and Nitrate are spectators

Things to recognize about this reaction:

1. The number of moles of potassium ions in solution and nitrate ions in solution will remain the same
2. The concentrations of the potassium ions and nitrate ions will be diluted (get smaller) because the volumes of the solutions are being added together.
3. The moles of silver ions and bromide ions in solution will decrease because they are precipitating out of solution. This is good to remember when dealing with limiting reactant/excess reactant stoichiometry and Le Chatelier's Principle.

Double Replacement Reactions (Formation of a gas)

Gases may form directly in a double replacement reaction or from the decomposition of one of the products. The gases will bubble off or evolve from solution.

1. Any sulfide salt (Ionic compound with sulfide, S^{2-}). Solid or aqueous when added to an acid will form $H_2S(g)$ and an ionic compound with the acid's anion.

Ex. Solid potassium sulfide is added to hydrochloric acid.

Equation: $K_2S(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2S(g)$ *Hydrochloric acid is aqueous because it is an acid.

Net Ionic: $K_2S(s) + 2H^+ \rightarrow 2K^+ + H_2S(g)$ *Hydrochloric acid dissociates because it is a strong acid. Potassium chloride dissociates because it is soluble in solution. Chloride is a spectator ion

2. Any carbonate salt (Ionic compound with carbonate) solid or aqueous added to any acid will form carbonic acid which will break down to $H_2O(l) + CO_2(g)$.

Ex. Potassium carbonate is added to a solution of acetic acid.

Equation: $K_2CO_3(aq) + 2HC_2H_3O_2(aq) \rightarrow H_2CO_3(aq) + 2KC_2H_3O_2(aq)$

Net Ionic: $CO_3^{2-} + HC_2H_3O_2(aq) \rightarrow H_2O(l) + CO_2(g) + C_2H_3O_2^-$ *Acetic acid does not dissociate because it is a WEAK acid. Potassium is a spectator ion.

3. Any sulfite salt (Ionic compound with sulfite (SO_3^{2-})) solid or aqueous added to any acid will form sulfurous acid which will break down into $H_2O(l) + SO_2(g)$.

Ex. Hydrochloric acid is poured onto solid sodium sulfite.

Equation: $Na_2SO_3(s) + 2HCl(aq) \rightarrow H_2SO_3(aq) + 2NaCl(aq)$

Net Ionic: $Na_2SO_3(s) + 2H^+ \rightarrow H_2O(l) + SO_2(g) + 2Na^+$ *Hydrochloric acid dissociates because it is a strong acid. Sodium chloride is aqueous because it is soluble in solution based on solubility rules.

4. Any ammonium salt (Ionic compound with Ammonium as the cation) added to a strong base (soluble ionic compound which contains hydroxide) will form ammonium hydroxide which will break down to form $NH_3(g) + H_2O(l)$

Ex. Ammonium chloride is added to lithium hydroxide.

Equation: $NH_4Cl(aq) + LiOH(aq) \rightarrow NH_4OH(aq) + LiCl(aq)$

Net Ionic: $NH_4^+ + OH^- \rightarrow NH_3(g) + H_2O(l)$ *Lithium and chloride end up being spectator ions.

Double Replacement (Acid-Base Neutralization)

Acids react with bases to produce salts (ionic compounds) and water. One mole of hydrogen ions (H⁺) will react with one mole of hydroxide ions (OH⁻) to produce one mole of water. Strong acids and bases will dissociate. Weak acids and bases will stay together. You must pay attention to the amount of acid and base that react with each other so you get the correct ratio.

1. An acid is added to a base, but amounts are not specified. Go with the basic double replacement reaction.

Ex. Nitrous acid is added to potassium hydroxide.



Net Ionic: $\text{HNO}_2(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NO}_2^-$ *Nitrous acid stays together because it is a weak acid.

2. When amounts are given, you must figure out the ratio of H⁺ to OH⁻.

Ex. Equal moles of phosphoric acid and sodium hydroxide are mixed.

Reactants: $1\text{H}_3\text{PO}_4(\text{aq}) + 1\text{NaOH}$ *Note that there is one hydroxide and three hydrogen ions.

What is left over after neutralization: Na^+ and H_2PO_4^-

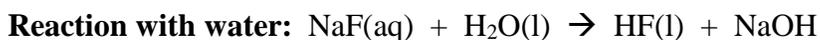


Net Ionic: $\text{H}_3\text{PO}_4(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l}) + \text{H}_2\text{PO}_4^-$ * Phosphoric acid does not dissociate because it is a weak acid!

Is the solution acidic, basic or neutral?

1. If the solution contains Fluoride ions (F⁻), carbonate ions, cyanide ions (CN⁻), sulfide ions, nitrite ions, or any other ions that when attached to H⁺ form a weak acid...then your solution will be basic.

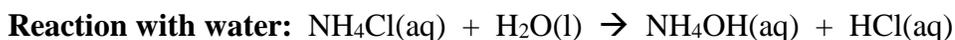
Ex. Sodium fluoride in solution.



Net Ionic: $\text{F}^- + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HF}(\text{aq}) + \text{OH}^-$ *HF does not dissociate because it is a weak acid. The result of the reaction is OH⁻ is produced therefore making the solution basic.

2. If the solution contains the ammonium ion, or any other ion that when combined with the OH⁻ from water forms a weak base...then the solution will be acidic.

Ex. Ammonium chloride in solution.



Net Ionic: $\text{NH}_4^+ + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4\text{OH}(\text{aq}) + \text{H}^+$ *Ammonium hydroxide stays together because it is a weak base. The result is that H⁺ gets produced in solution which **makes it acidic**.

Combustion of Hydrocarbons

There are two types of combustion...complete combustion...this occurs in excess oxygen. And incomplete combustion, where this is limited oxygen. You can assume complete combustion unless told other wise. These reactions are always exothermic (Release Energy)

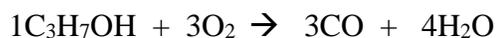
Complete Combustion of a hydrocarbon: Products are always CO₂ and H₂O....CO₂ will be a gas, H₂O you can write as liquid or gas. Please note that the ΔH for a combustion reaction is going to be more negative when you express H₂O as a liquid rather than a gas, because you are including the energy released when H₂O(g) → H₂O(l)

Example: Propanol burns in oxygen.



Incomplete Combustion of a hydrocarbon: Not enough oxygen to make CO₂, so the products are CO and H₂O...You could also include all three...CO₂, CO, and H₂O...it is your call.

Example: Propanol burns in a limited supply of oxygen.



Or



AP chemistry color guide. SAVE THIS SHEET Need to know for SAT II if needed

The SAT II often asks questions in the multiple choice section which require knowing colors. These are generally “you know it or you don’t type questions”. Some common chemistry colors that you may need (Some of which you have seen or will see this year).

Remember: Violet = high energy (dangerous ultraviolet rays from the sun), high frequency, low wavelength.

Red = low energy, low frequency, high wavelength.

1. Solutions. There are zillions of colored solutions. Here are some you might want to know.

| | |
|---|---|
| permanganate, MnO_4^- : purple | bromine: orangish |
| chlorine: greenish yellow | iodine: yellow/orange/brown |
| Fe^{3+} : yellow-orange-brown | FeSCN^{2+} : blood red |
| Cu^{2+} : light blue | $\text{Cu}(\text{NH}_3)_4^{2+}$: dark blue |

More on Ions in solution: As a general rule, alkali and alkaline earth metal ions, post transition metal ions (Pb^{2+} , Al^{3+} , Sn^{2+} , Bi^{3+}) and anions are colorless (an exception is MnO_4^- above, another is CrO_4^{2-} , yellow). Most colored ions are transition metals or their complexes. Other transition metals include Mn^{2+} and Co^{2+} , both pink, or Cr^{3+} and Ni^{2+} , both green. However, some transition metal ions (notably Zn^{2+} , Ag^+ , and Hg_2^{2+}) are NOT colored.

2. Indicators. Remember that many common indicators change from acid to base in the direction of the rainbow (clear-red-orange-yellow-green-blue-purple). Also remember that the color at the pK_a of the indicator is a mix of the acid and base colors. You probably only need to know a few acid-base indicators:

| | | |
|------------------|-----------------------------------|---|
| Phenolphthalein: | clear - pink - red | $\text{pK}_a = 9$ (use for Weak Acid – Strong Base) |
| Bromthymol blue: | yellow - green - blue | $\text{pK}_a = 7$ (use for Strong Acid – Strong Base) |
| Methyl red: | red - orange - yellow | $\text{pK}_a = 5$ (use for Strong Acid – Weak Base) |
| Litmus: | red in acid, blue in base | |
| Others: | starch + iodine = dark blue-black | |

3. Solids (mostly precipitates):

Most precipitates are white including BaSO_4 , AgCl , and CaCO_3 , $\text{Ba}(\text{OH})_2$.

Yellow precipitates include AgI , PbI_2 , and some chromates.

Sulfide precipitates show a variety of colors. $\text{Fe}(\text{OH})_3$ is rust colored.

Non-metals: P = red, S = yellow, I = lustrous, purple-black crystals

4. Gases. Most gases that we use are colorless. This is because colored gases tend to be dangerous (and often have yucky odors). Here are some common examples you should know.

| | | |
|---------------------------|--|----------------|
| chlorine: yellow-green | bromine: reddish | iodine: purple |
| NO_2 : red-brown | (note: N_2O is colorless laughing gas) | |

Note: Some colorless gases are also dangerous: ex. carbon monoxide, ammonia

5. Liquids. Very few chemicals are liquid at room temperature. Other colored “liquids” are really solutions (see above):

Bromine: red-brown

Mercury: silvery (like most metals)

6. Flame test. The following metal ions give the following colors: Sodium- yellow; Potassium- violet; Copper- green; Most others-red or reddish orange (Lithium, Strontium, and Calcium); Barium- yellow/green.