<u>REACTIONS (Part 2)</u>

You will need to know/ memorize the following:
Solubility rules
The strong acids and bases
Oxidation states
The common oxidizing and reducing agents (and what the products are).

You should also review: naming and formulas, polyatomic ions, the diatomic elements, organic formulas and naming, balancing equations, and writing ionic and net ionic equations.

Chemical reactions are the primary means by which transformations to matter occur. Chemical equations for reactions show the rearrangements of atoms that occur during a chemical reaction. (E.U.3.A)

EVIDENCE FOR CHEMICAL CHANGE:
- Production of light or heat.
- Formation of a gas.
- Formation of a precipitate.
- Color change.

TYPES OF REACTION:

A. Combination/Synthesis Reactions
- Atoms or molecules combine to form new compounds.
- $A + B \rightarrow AB$

1. metal and nonmetal form a salt (binary ionic compound)
   - e.g. $2\ Na + F_2 \rightarrow 2\ NaF$
2. two nonmetals form a binary covalent compound
   - e.g. $S + F_2 \rightarrow SF_2$
3. metal or nonmetal combines with oxygen to form binary ionic or covalent compound
   - e.g. $2\ Mg + O_2 \rightarrow 2\ MgO$
4. nonmetallic element combines with a binary covalent compound
   - e.g. oxygen and nonmetallic oxide (sulfur dioxide + oxygen → sulfur trioxide)
   - e.g. nonmetallic halide and additional halogen (chlorine trifluoride and fluorine → chlorine tetrafluoride)
5. Combination of two compounds
   a. metallic oxide and nonmetallic oxide form an ionic compound (salt) with a polyatomic ion
      - e.g. $Na_2O + CO_2 \rightarrow Na_2CO_3$
   b. metal oxides react with water to form bases (hydroxides)
      - e.g. $CaO + H_2O \rightarrow Ca(OH)_2$
   c. nonmetal oxides react with water to form acids
      - e.g. $SO_2 + H_2O \rightarrow H_2SO_3$
   d. hydrates result when anhydrous compounds react with water to form hydrates
      - e.g. $CuSO_4 + 5\ H_2O \rightarrow CuSO_4 \cdot 5H_2O$
   e. Lewis acid (e- pr acceptor) and Lewis base (e- pr. donor)
      - e.g. boron trifluoride (e- deficient) and ammonia (lone pair on N) : $BF_3 + NH_3 \rightarrow BF_3NH_3$

B. Decomposition/ Analysis Reactions
- The reverse of synthesis.
- Molecules are decomposed, often by heating.

1. decomposition (not necessarily heated):
   a. hydrogen peroxide decomposes into water and oxygen.
   b. ammonium hydroxide decomposes into ammonia and water.
2. thermal decomposition (are heated)
   a. hydrogencarbonates (bicarbonates) yield carbonates, water and carbon dioxide (relatively low temps)
      - e.g. $NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
   b. metallic carbonates decompose into metallic oxides and carbon dioxide (high temp.)
      - e.g. $CaCO_3 \rightarrow CaO + CO_2$
   c. ammonium carbonate decomposes into ammonia, water and carbon dioxide
   d. sulfites yield oxides and sulfur dioxide
      - e.g. $FeSO_3 \rightarrow FeO + SO_2$
   e. oxides, chlorates and perchlorates yield oxygen
      - e.g. $2HgO \rightarrow Hg + O_2$
      - e.g. $2KClO_3 \rightarrow 2KCl + 3O_2$
      - e.g. $NaClO_4 \rightarrow NaCl + 2O_2$
   f. hydroxides, hydrates and some oxyacids release water
      - e.g. $Ca(OH)_2 \rightarrow CaO + H_2O$
      - e.g. $CuSO_4\text{-}5H_2O \rightarrow CuSO_4 + 5\,H_2O$
      - e.g. $H_2SO_3 \rightarrow H_2O + SO_2$
      - e.g. $H_2CO_3 \rightarrow H_2O + CO_2$
3. Electrolysis - use of electricity to decompose compounds
   Binary compounds decompose into two elements
   - e.g. $2\,NaCl \rightarrow 2\,Na + Cl_2$

C. Combustion Reactions

- $C_xH_y + O_2 \rightarrow CO_2 + H_2O$
  Combustion of hydrocarbons produces water and carbon dioxide
  - e.g. $2\,C_6H_{14} + 19\,O_2 \rightarrow 14\,H_2O + 12\,CO_2$

D. Single Replacement/ Displacement Reactions

- $A + BC \rightarrow AC + B$

A more active element replaces a less active element in a compound.

Activity series for metals:
Most active **Li Ca Na Mg Al Zn Fe Pb [H_2] Cu Ag Pt** Least active

Activity series for non-metals
Most active **F_2 Cl_2 Br_2 I_2** Least active

Active metal replaces hydrogen in water or an acid
a. metal + acid → salt and hydrogen gas
   - e.g. $Mg(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ (Net : $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$)

b. Group 1A or 2A metal and water → base and hydrogen gas
   - e.g. $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$ (Net : $Li + 2H_2O \rightarrow Li^+ + OH^- + H_2$)

1. a more active metal replaces a less reactive metal ion from solution(see activity series)
   - e.g. $ZnCl_2(aq) + Mg(s) \rightarrow MgCl_2(aq) + Zn(s)$ (Net : $Zn^{2+} + Mg \rightarrow Mg^{2+} + Zn$)
2. a halogen replaces another halogen
   - e.g. $MgBr_2(aq) + F_2(g) \rightarrow MgF_2(aq) + Br_2(l)$ (Net : $2\,Br^- + F_2 \rightarrow 2\,F^- + Br_2$)

4. Hydrides of alkali metals react with water to form hydroxides :
   - e.g. $LiH(s) + H_2O(l) \rightarrow LiOH + H_2(g)$ (Net : $LiH + H_2O \rightarrow Li^+ + OH^- + H_2$)

E. Double Replacement/ Metathesis Reactions

- $AB + CD \rightarrow AD + CB$
- For a double replacement reaction to occur, ions must be removed from the solution by formation of a precipitate, a gas, or a molecular compound.

Memorize solubility rules, strong acids and bases, and the following gas formers:

Common Gas Forming Reactions:

Common Gases	
H_2S	Any sulfide (S^{2-}) + any acid form $H_2S_{(g)}$ and a salt
CO_2	Any carbonate (CO_3^{2-}) + any acid form $CO_{2(g)}$, H_2O, and a salt
SO_2	Any sulfite (SO_3^{2-}) + any acid form $SO_{2(g)}$, H_2O, and a salt
NH_3	Any ammonium salt (NH_4^+) + a soluble strong hydroxide react upon heating to form $NH_{3(g)}$, H_2O, and a salt

e.g: $HC_2H_3O_{2(aq)} + NaHCO_{3(aq)} \rightarrow NaC_2H_3O_{2(aq)} + H_2CO_{3(aq)}$ $\quad H_2CO_{3(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

1. formation of an insoluble ionic compound (precipitate).
   - e.g. $BaCl_{2(aq)} + MgSO_{4(aq)} \rightarrow BaSO_{4(s)} + MgCl_{2(aq)}$ (Net :$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_{4(s)}$)
2. neutralization reaction between an acid and a base (forms water)
   a. strong acid/strong base :
   - e.g. $HCl + NaOH \rightarrow H_2O + NaCl$ (net : $H^+ + OH^- \rightarrow H_2O$)

   b. strong acid/weak base
   - e.g. $HCl + NH_3 \rightarrow NH_4Cl$ (net : $H^+ + NH_3 \rightarrow NH_4^+$)
   - e.g. $HCl + CH_3NH_2 \rightarrow CH_3NH_3Cl$ (net : $H^+ + CH_3NH_2 \rightarrow CH_3NH_3^+$)

   c. weak acid/strong base
   - e.g. $HC_2H_3O_2 + NaOH \rightarrow H_2O + C_2H_3O_2^-$ (net : $HC_2H_3O_2 + OH^- \rightarrow H_2O + C_2H_3O_2^-$)

   d. weak acids and weak bases
   - e.g. $NH_3 + HF \rightarrow NH_4^+ + F^-$ (net)
3. reactions with acids :
   a. carbonates or bicarbonates and acids form a salt, water and CO_2
   - e.g. $2HCl + Na_2CO_3 \rightarrow 2\,NaCl + H_2O + CO_2$ (net : $H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$)

   b. sulfites and acids form a salt, water and SO_2
   - e.g. $2\,HCl + Na_2SO_3 \rightarrow 2\,NaCl + H_2O + SO_2$ (net : $H^+ + SO_3^{2-} \rightarrow H_2O + SO_2$)

   c. metallic sulfides and acids form H_2S and a salt
   - e.g. $2\,HCl(aq) + Na_2S(aq) \rightarrow 2\,NaCl(aq) + H_2S(g)$ (net : $H^+ + S^{2-} \rightarrow H_2S$)

   d. metallic hydrides and acids form H_2 and a salt
   - e.g. $HCl(aq) + LiH(aq) \rightarrow LiCl(aq) + H_2(g)$ (net : $H^+ + H^- \rightarrow H_2$)
4. strong acids with salts of weak acids (salts containing anions such as : OCl^-, ClO_2^-, F^-, NO_2^-, CN^-, $C_2H_3O_2^-$)
   a. metallic acetates and acids form acetic acid and a salt
   - e.g. $HCl(aq) + NaC_2H_3O_2(aq) \rightarrow NaCl(aq) + HC_2H_3O_2(aq)$ (net : $H^+ + C_2H_3O_2^- \rightarrow HC_2H_3O_2$)

   b. metallic nitrites and acids form nitrous acid and a salt
   - e.g. $HCl(aq) + NaNO_2(aq) \rightarrow HNO_2(aq) + NaCl(aq)$ (net : H+ + $NO_2^- \rightarrow HNO_2$)
5. reactions with bases :
   a. ammonium salts and soluble bases yield ammonia, water and a salt
   - e.g. $NH_4Cl + LiOH \rightarrow NH_3 + H_2O + LiCl$ (net : $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$)
6. Amides with water
   - e.g. sodium amide and water produce sodium hydroxide and ammonia
     - Net : $NaNH_2 + H_2O \rightarrow Na^+ + OH^- + NH_3$

F. Oxidation Reduction Reactions (redox):

Memorize oxidation states.
Terminology:

Oxidation - the oxidation number of one or more elements increases (it loses electrons).
Reduction - the oxidation number of one or more of the elements decreases (it gains electrons).
(OIL RIG or LEO says GER)

Redox Reactions - involve a transfer of electrons from the species that is oxidized to the species that is reduced. Single replacement and combustion reactions are always redox reactions. Combination and decomposition reactions are sometimes redox reactions. If a reaction takes place in an acidic or basic solution it is most likely a redox reaction. Reactions involving elements in their natural states (e.g. Al, Fe, Cl_2, O_2, etc.) **are** redox reactions.

Balancing Oxidation-reduction reactions (Oxidation States Method):

1. Assign oxidation numbers to all atoms in the equation
2. Identify the substance being oxidized and determine the number of electrons lost.
3. Identify the substance being reduced and determine the number of electrons gained.
4. Use coefficients to balance the atoms in the substances oxidized and reduced.
5. Use coefficients to balance the electrons gained and lost.
6. Use coefficients to balance the non redox substances.

Example:

$\mathbf{CdS + I_2 + HCl \rightarrow CdCl_2 + HI + S}$

1. **Oxid #'s** $Cd^{2+}\ S^{2-} + I_2^{0} + H^{+}\ Cl^{-} \rightarrow Cd^{2+}\ Cl^{2-} + H^{+}\ I^{-} + S^{0}$
2. **Oxidation** Sulfur changes from a 2- to a 0 oxidation state, 2 electrons are lost.
3. **Reduction** Iodine changes from 0 to -1 oxidation state, 1 electron is gained.
4. **Balance atoms** Because iodine is diatomic on the reactant side and both atoms must be reduced, there must be 2 HI molecules formed on the product side.
   $\mathbf{CdS + I_2 + HCl \rightarrow CdCl_2 + 2HI + S}$

5. **Balance e's** Since sulfur loses 2 electrons and each iodine gains 1 (for a total of two), the number of electrons lost and gained is balanced.
6. **Balance atoms** Balance the hydrogen and chlorine atoms.
   $\mathbf{CdS + I_2 + 2HCl \rightarrow CdCl_2 + 2HI + S}$

Example: Balance the following redox reaction

$\mathbf{Cl_2 + Ca(OH)_2 \rightarrow CaCl_2 + Ca(ClO_3)_2 + H_2O}$

Note: This is an example of a **disproportionation reaction** - the same reactant (Cl_2) undergoes both oxidation and reduction.

1. Simple redox reactions :
   a. Hydrogen displacement
   - e.g. $Ca(s) + H_2O \rightarrow Ca(OH)_2(aq) + H_2(g)$

   b. Metal displacement
   - e.g. $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$ (Net : $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$)

   c. Halogen displacement
   - e.g. $Cl_2(g) + MgBr_2(aq) \rightarrow MgCl_2(aq) + Br_2(l)$ (Net : $Cl_2 + Br^- \rightarrow Cl^- + Br_2$)

   d. Combustion reactions.

- e.g. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

e. Combination and decomposition reactions (see above)

2. Reactions involving transition metals with multiple oxidation states :
    - e.g. tin(II) ion with Fe(III) ion (net : $Sn^{2+} + Fe^{3+} \rightarrow Sn^{4+} + Fe^{2+}$)
3. Free halogens in dilute basic solutions form hypohalite ions

- e.g. $Cl_2(g) + KOH(aq) \rightarrow KClO(aq) + KCl(aq) + H_2O(l)$ (Net : $Cl_2 + OH^- \rightarrow ClO^- + Cl^- + H_2O$)

4. Redox reactions involving oxyanions such as $Cr_2O_7^{2-}$
    - e.g. $14H+(aq) + Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 7\,H_2O(l) + 6Fe^{3+}$

The only way to be able to predict these types of reactions is to memorize common oxidizing and reducing agents.

5. Atypical redox reactions

- hydrogen reacts with a hot metallic oxide to produce the elemental metal and water
- a metal sulfide reacts with oxygen to produce the metallic oxide and sulfur dioxide
- chlorine gas reacts with *dilute* sodium hydroxide to produce sodium hypochlorite, sodium chloride, and water
- Copper reacts with *concentrated* sulfuric acid to produce copper (II) sulfate, sulfur dioxide, and water
- Copper reacts with *dilute* nitric acid to produce copper (II) nitrate, nitrogen monoxide, and water
- Copper reacts with *concentrated* nitric acid to produce copper (II) nitrate, nitrogen dioxide, and water