

REACTIONS (Part 2)

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. $2 \text{Na} + \text{F}_2 \rightarrow 2 \text{NaF}$
 2. two nonmetals form a binary covalent compound
 - e.g. $\text{S} + \text{F}_2 \rightarrow \text{SF}_2$
 3. metal or nonmetal combines with oxygen to form binary ionic or covalent compound
 - e.g. $2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}$
 4. nonmetallic element combines with a binary covalent compound
 - e.g. oxygen and nonmetallic oxide (sulfur dioxide + oxygen \rightarrow sulfur trioxide)
 - e.g. nonmetallic halide and additional halogen (chlorine trifluoride and fluorine \rightarrow chlorine tetrafluoride)
 5. Combination of two compounds
 - a. metallic oxide and nonmetallic oxide form an ionic compound (salt) with a polyatomic ion
 - e.g. $\text{Na}_2\text{O} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3$
 - b. metal oxides react with water to form bases (hydroxides)
 - e.g. $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
 - c. nonmetal oxides react with water to form acids
 - e.g. $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
 - d. hydrates result when anhydrous compounds react with water to form hydrates
 - e.g. $\text{CuSO}_4 + 5 \text{H}_2\text{O} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
 - e. Lewis acid (e- pr acceptor) and Lewis base (e- pr. donor)
 - e.g. boron trifluoride (e- deficient) and ammonia (lone pair on N) : $\text{BF}_3 + \text{NH}_3 \rightarrow \text{BF}_3\text{NH}_3$

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. $\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$
 - b. metallic carbonates decompose into metallic oxides and carbon dioxide (high temp.)
 - e.g. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
 - c. ammonium carbonate decomposes into ammonia, water and carbon dioxide
 - d. sulfites yield oxides and sulfur dioxide
 - e.g. $\text{FeSO}_3 \rightarrow \text{FeO} + \text{SO}_2$
 - e. oxides, chlorates and perchlorates yield oxygen
 - e.g. $2\text{HgO} \rightarrow \text{Hg} + \text{O}_2$
 - e.g. $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$
 - e.g. $\text{NaClO}_4 \rightarrow \text{NaCl} + 2\text{O}_2$
 - f. hydroxides, hydrates and some oxyacids release water
 - e.g. $\text{Ca}(\text{OH})_2 \rightarrow \text{CaO} + \text{H}_2\text{O}$
 - e.g. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \rightarrow \text{CuSO}_4 + 5\text{H}_2\text{O}$
 - e.g. $\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O} + \text{SO}_2$
 - e.g. $\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2$
3. Electrolysis - use of electricity to decompose compounds
Binary compounds decompose into two elements
 - e.g. $2\text{NaCl} \rightarrow 2\text{Na} + \text{Cl}_2$

C. Combustion Reactions

- $\text{C}_x\text{H}_y + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
Combustion of hydrocarbons produces water and carbon dioxide
- e.g. $2\text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 14\text{H}_2\text{O} + 12\text{CO}_2$

D. Single Replacement/ Displacement Reactions

- $\text{A} + \text{BC} \rightarrow \text{AC} + \text{B}$

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 [H₂] Cu Ag Pt** Least active

Activity series for non-metals

Most active **F₂ Cl₂ Br₂ I₂** Least active

Active metal replaces hydrogen in water or an acid

a. metal + acid \rightarrow salt and hydrogen gas

- e.g. $\text{Mg}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$ (Net: $\text{Mg} + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2$)

b. Group 1A or 2A metal and water \rightarrow base and hydrogen gas

- e.g. $2\text{Li}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{LiOH}(\text{aq}) + \text{H}_2(\text{g})$ (Net: $\text{Li} + 2\text{H}_2\text{O} \rightarrow \text{Li}^+ + \text{OH}^- + \text{H}_2$)

1. a more active metal replaces a less reactive metal ion from solution(see activity series)

- e.g. $\text{ZnCl}_2(\text{aq}) + \text{Mg}(\text{s}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{Zn}(\text{s})$ (Net: $\text{Zn}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Zn}$)

2. a halogen replaces another halogen

- e.g. $\text{MgBr}_2(\text{aq}) + \text{F}_2(\text{g}) \rightarrow \text{MgF}_2(\text{aq}) + \text{Br}_2(\text{l})$ (Net: $2\text{Br}^- + \text{F}_2 \rightarrow 2\text{F}^- + \text{Br}_2$)

4. Hydrides of alkali metals react with water to form hydroxides :

- e.g. $\text{LiH}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{LiOH} + \text{H}_2(\text{g})$ (Net: $\text{LiH} + \text{H}_2\text{O} \rightarrow \text{Li}^+ + \text{OH}^- + \text{H}_2$)

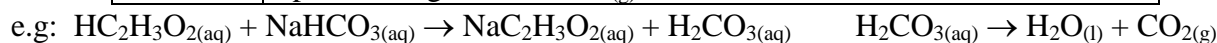
E. Double Replacement/ Metathesis Reactions

- $AB + CD \rightarrow AD + CB$
- 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 |



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. $2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$ (net: $H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$)
 - b. sulfites and acids form a salt, water and SO_2
 - e.g. $2HCl + Na_2SO_3 \rightarrow 2NaCl + H_2O + SO_2$ (net: $H^+ + SO_3^{2-} \rightarrow H_2O + SO_2$)
 - c. metallic sulfides and acids form H_2S and a salt
 - e.g. $2HCl(aq) + Na_2S(aq) \rightarrow 2NaCl(aq) + H_2S(g)$ (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).

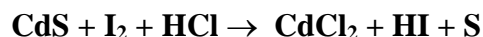
(OIL RIG or LEO says GER)

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, Cl₂, O₂, etc.) **are** redox reactions.

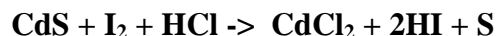
Balancing Oxidation-reduction reactions (Oxidation States Method):

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:



1. **Oxid #'s** $\text{Cd}^{2+} \text{S}^{2-} + \text{I}_2^0 + \text{H}^+ \text{Cl}^- \rightarrow \text{Cd}^{2+} \text{Cl}_2^- + \text{H}^+ \text{I} + \text{S}^0$
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.



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.



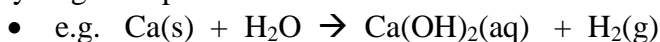
Example: Balance the following redox reaction



Note: This is an example of a **disproportionation reaction** - the same reactant (Cl₂) undergoes both oxidation and reduction.

1. Simple redox reactions :

a. Hydrogen displacement



b. Metal displacement



c. Halogen displacement



d. Combustion reactions.

- e.g. $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

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 : $\text{Sn}^{2+} + \text{Fe}^{3+} \rightarrow \text{Sn}^{4+} + \text{Fe}^{2+}$)

3. Free halogens in dilute basic solutions form hypohalite ions

- e.g. $\text{Cl}_2(\text{g}) + \text{KOH}(\text{aq}) \rightarrow \text{KClO}(\text{aq}) + \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ (Net : $\text{Cl}_2 + \text{OH}^- \rightarrow \text{ClO}^- + \text{Cl}^- + \text{H}_2\text{O}$)

4. Redox reactions involving oxyanions such as $\text{Cr}_2\text{O}_7^{2-}$

- e.g. $14\text{H}^+(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 6\text{Fe}^{3+}$

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