Types of Rxns

1. Synthesis or Combination Rxns: Two reactants that combine to form a single product.

Example: $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3$

2. Decomposition Rxn: One reactant that breaks down and forms two or more products.

Example: $MgCO_3 \rightarrow MgO + CO_2$

3. Single Replacement rxn: Usually a rxn where a neutral metal reacts with an ionic compound changing places with the metal cation of the ionic compound.

Examples: ① $2FeCl_3 + 3Zn(s) \rightarrow 2Fe(s) + 3ZnCl_2(aq)$

② $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$

4. Double Replacement rxn: A rxn between two reactants where two elements switch places.

① $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$

② $HCl + NaOH \rightarrow H_2O + NaCl$

5. Combustion Rxn: Any reactant that is reacted with oxygen.

① $O_2 + 2H_2 \rightarrow 2H_2O$

② $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

↑ special case of a hydrocarbon

which consists only of C, H always produces CO_2 and H_2O.

Precipitation Reactions

The rxn between two ionic solutions where a double replacement occurs forming a new ionic solution and an ionic solid. Preciptitation = formation of a solid.

Example

ag means dissolved in water

$$\underbrace{AgNO_3(aq) + NaCl(aq)}_{\text{2 ionic solutions}} \longrightarrow \underbrace{AgCl(s)}_{\text{ionic solid}} + \underbrace{NaNO_3(aq)}_{\text{new ionic solution}}$$

How does this rxn happen? First, it is important to understand that if an ionic compound dissolves in water then it has to be fully ionized - in the form of separated ions. whereas for ionic compounds that don't dissolve in water, those compounds are not ionized - they are ionic solids and the ions are not separated. Secondly, that means that everything in the above example equation is in the form of separated ions except for AgCl (s). we can show this by writing what's called the complete ionic equation:

$$Ag^+ + NO_3^- + Na^+ + Cl^- \longrightarrow AgCl(s) + Na^+ + NO_3^-$$

Since the two reactants ($AgNO_3$ and $NaCl$) are fully ionized the cations can exchange with the anions - so Ag^+ bonds to Cl^- and Na^+ bonds to NO_3^- - A double replacement.

The new products, AgCl and $NaNO_3$, differ in how they interact with water: The $NaNO_3$ dissolves and therefore is ionized. But the AgCl does <u>not</u> dissolve and remains unionized in the form of a solid. The Na^+ and NO_3^- remain in solution and unreacted. So the only ions involved in the chemical rxn were Ag^+ and Cl^-. We show this by writing a <u>net ionic equation</u>:

$$Ag^+ + Cl^- \rightarrow AgCl\,(s)$$

we call the Na^+ and NO_3^- <u>spectator ions</u> because they are not involved in the rxn.

For any precipitation rxn:

The <u>molecular equation</u>: balanced chemical equation:

$$AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(s) + NaNO_3\,(aq)$$

The <u>complete ionic equation</u>: spectator ions are Na^+, NO_3^-

$$Ag^+ + \cancel{NO_3^-} + \cancel{Na^+} + Cl^- \rightarrow AgCl\,(s) + \cancel{Na^+} + \cancel{NO_3^-}$$

The <u>net ionic equation</u>: only the ions that form the solid are in this equation.

$$Ag^+ + Cl^- \rightarrow AgCl\,(s)$$

Make sure you can write the complete ionic and net ionic equations for the following precipitation rxns:

1. $AgNO_3 + NaCl \longrightarrow$

2. $Ba(NO_3)_2 + Na_2SO_4 \longrightarrow$

3. $Pb(NO_3)_2 + 2KI \longrightarrow$

soluble = dissolves

How do you tell if an ionic compound dissolves in water (is soluble in water)?

Rules of solubility:

1) All nitrate salts are soluble - dissolve in H_2O.

$AgNO_3$, $Pb(NO_3)_2$, $Fe(NO_3)_3$ = soluble

2) All alkali metal and ammonium salts are soluble.

$NaCl$, NH_4OH, KBr = soluble

Types of Rxns

1. Precipitation rxns: we did these rxns. See previous notes, videos, and lab notes.

2. Neutralization Rxns: A neutralization rxn occurs between an acid and a base to produce water and a salt. The possible acids are: HCl, HBr, HI, $HClO_4$, H_2SO_4, HNO_3. The possible bases are $NaOH$, KOH, $Ba(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$ or any (metal)OH).

Example

$$\underset{\text{acid}}{HCl} + \underset{\text{base}}{NaOH} \rightarrow \underset{\text{water}}{H_2O} + \underset{\text{salt}}{NaCl}$$

$$\underset{\text{acid}}{2HNO_3} + \underset{\text{base}}{Ba(OH)_2} \rightarrow \underset{\text{water}}{2H_2O} + Ba(NO_3)_2$$

3. Oxidation-Reduction (Redox): A chemical rxn that involves the exchange of e^-. Usually you can tell a Redox rxn based on identifying an element that goes from being neutral to being ionized. Most single replacements rxns are Redox:

$$\underset{\text{ionized to } Fe^{+3}}{2FeCl_3} + \underset{\text{neutral}}{3Zn} \rightarrow \underset{\text{neutral}}{2Fe} + \underset{\text{ionized to } Zn^{+2}}{3ZnCl_2}$$

In the above example Fe^{+3} gained $3e^-$ to become Fe and Zn lost $2e^-$ to become Zn^{+2}.

Here is another example:

neutral, lost e^-

$$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$$

neutral Br^- Cl^-, gained e^-

gains e^- loses e^-

Combustion rxns are redox rxns. Combustion rxns are rxns of elements or compounds with oxygen:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Redox + combustion

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

For organic compounds (C_3H_8, CH_3CH_2OH, etc.) the products of combustion are always CO_2 and H_2O.

oxidation/reduction rxns

oxidation/Reduction = REDOX

Redox Rxns

oxidation: the loss of one or more e^-

reduction: the gain of one or more e^-

Redox chemical rxns require that at least one e^- is transferred from one component to another component in the chemical equation. As an example consider the following rxn between sodium metal and chlorine gas:

$$2\,Na(s) + Cl_2(g) \longrightarrow 2\,NaCl(s)$$

you can see that the neutral metal reacted with the neutral gas to form an ionic compound. Both the Na(s) and $Cl_2(g)$ are uncharged so we say that they have a charge of zero = 0. But in NaCl the Na has become Na^+ and the Cl has become Cl^-. So one electron was given up by Na and accepted by the chlorine.

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

So we say that Na metal was <u>oxidized</u> and chlorine gas, $Cl_2(g)$, was <u>reduced</u>.

Usually you can identify <u>redox</u> rxns if the reaction is combustion or at least one of the reactants is oxygen gas, $O_2(g)$.

(A) <u>Examples of rxns that are Redox</u>: one reactant is oxygen

$$2H_2 + O_2 \rightarrow 2H_2O$$

$$2CH_4 + 3O_2 \rightarrow 2CO_2 + 4H_2O$$

$$N_2 + O_2 \rightarrow 2NO$$

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

$$S + O_2 \rightarrow SO_2$$

$$S_8 + 12O_2 \rightarrow 8SO_3$$

(B) Classic combustion: <u>Know</u> Know Know!!!

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

$$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$$

$$2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O$$

$$C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$$

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

So we say that the sodium metal was <u>oxidized</u> and the chlorine gas was <u>reduced</u>.

Here is another example:

$$Mg(s) + 2AgNO_3(aq) \longrightarrow 2Ag(s) + Mg(NO_3)_2(aq)$$

In this case Mg metal is oxidized and silver ion, Ag^+, is reduced to metal ($Ag(s)$). We can use a simple set of rules to make this job easier. <u>See the next page for oxidation rules.</u>

Let's apply the rules to this rxn:

Stays the same

$$Mg(s) + 2AgNO_3(aq) \longrightarrow 2Ag(s) + Mg(NO_3)_2(aq)$$

Assign # $\rightarrow$ 0   +1 $\longrightarrow$ 0   +2

+1 because we know nitrate is NO_3^- and so Ag has to be +1.

Note that Ag goes from +1 to zero as the metal. Mg goes from zero to +2 as the ion. Since the oxidation numbers = charges <u>changed</u>, then the rxn is a <u>Redox</u>.

7th Grade Oxidation Number Rules

1. Oxidation number for an element in its natural form is zero (0). Examples: Na(s), H_2(g), O_2(g), Hg(l), etc.

2. All 7A elements are -1 when they appear in a compound except when bonded to oxygen.
3. Group 1A is +1, Group 2A is +2.
4. Oxygen is -2 when it appears in a compound.
5. H in covalent compounds is +1.
6. For neutral compounds (charge is 0) the sum of oxidation numbers must be zero.

For compounds with a charge (NO_3^-) the sum of the oxidation numbers must equal the charge

How do we tell that a rxn is redox?

Consider the following rxn: $2KClO_3 \rightarrow 2KCl + O_2$

To be redox we must determine if e^- have been exchanged. We use the rules along with simple equations to identify the charge = oxidation number of all elements in equation and then we compare numbers of the individual atoms looking for a change.

So for $2KClO_3 \rightarrow 2KCl + O_2$

<u>Step 1</u>

Find elements in natural state and assign zero.

$$2KClO_3 \rightarrow 2KCl + \overset{0}{O_2}$$

<u>Step 2</u>

Identify simple ionic compounds and assign charges.

$$2KClO_3 \rightarrow 2\overset{+1}{K}\overset{-1}{Cl} + \overset{0}{O_2}$$

<u>Step 3</u>

For complex compounds set up equation set to zero and solve unknown. For $KClO_3$ we know K and O but not Cl. Since $KClO_3$ is uncharged

$$\overset{+1}{K} + Cl + 3(\overset{-2}{O}) = 0$$

$$+1 + Cl + -6 = 0$$

$$Cl - 5 = 0$$

$$Cl = +5$$

Step 4

Now evaluate equation:

$$2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \rightarrow \overset{+1}{K}\overset{-1}{Cl} + \overset{0}{O_2}$$

we see that the charges changed for Cl and O.

So rxn is **Redox** !

Step 5

To determine oxidation numbers = charges of elements in a charged molecule follow this procedure:

For MnO_4^-: we will find Mn.

$$Mn + 4(\overset{-2}{O}) = -1$$

$$Mn - 8 = -1$$

$$Mn = +7$$

Complete sample problems on next page.

Chem 7 Acids and Bases

A

Definitions

Acid: Any substance that can donate a hydrogen ion in a proton transfer rxn.

Examples: HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$

Base: Any substance that can accept a hydrogen ion in a proton transfer rxn.

Examples: $NaOH$, KOH, $Ba(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$

Strong acids and Strong bases

Strong: means that the acid or base in question ionizes to 100% in water.

Examples:

	$HCl \rightarrow$	H^+	+ Cl^-	$NaOH \rightarrow$	Na^+	+ OH^-
Before →	1 M	0	0	1 M	0	0
ionization ↓				↓		
After	0	1M	1M	0	1M	1M

Make sure you can write ionization equations for all the strong acids and strong bases.

(c) Ionization Equations for all six strong acids and 5 strong bases:

$HCl \rightarrow H^+ + Cl^-$

$HBr \rightarrow H^+ + Br^-$

$HI \rightarrow H^+ + I^-$

$HNO_3 \rightarrow H^+ + NO_3^-$

$H_2SO_4 \rightarrow H^+ + HSO_4^-$

$HClO_4 \rightarrow H^+ + ClO_4^-$

$NaOH \rightarrow Na^+ + OH^-$

$KOH \rightarrow K^+ + OH^-$

$Ba(OH)_2 \rightarrow Ba^{+2} + 2OH^-$

$Ca(OH)_2 \rightarrow Ca^{+2} + 2OH^-$

$Sr(OH)_2 \rightarrow Sr^{+2} + 2OH^-$

(D) Neutralization Rxns and Titrations

Neutralization: The rxn between an acid and a base that produces water and a neutral salt. The pH is equal to 7 when the neutralization rxn is complete.

Examples: $HCl + NaOH \rightarrow H_2O + NaCl$

$2HNO_3 + Ba(OH)_2 \rightarrow 2H_2O + Ba(NO_3)_2$

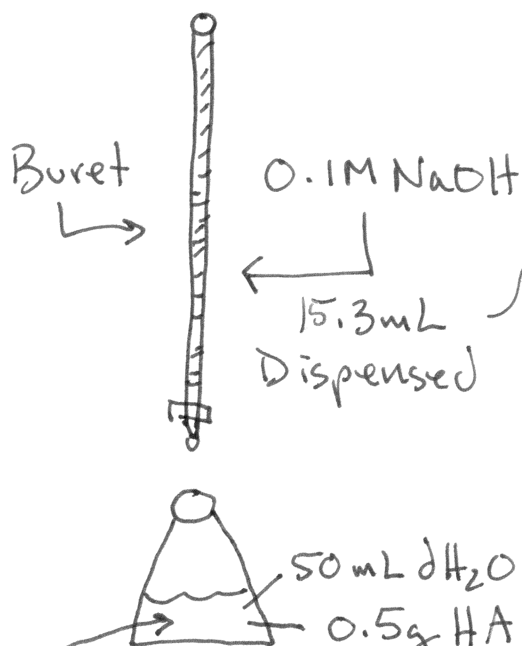
Make sure you can write a neutralization equation for any combination of strong acid and strong base.

Quantitative Titration: using neutralization to quantitate unknowns.

you will be given a mixture that contains solid NaCl mixed with a solid strong acid (HA). you will use a solution of 0.1M NaOH to neutralize the acid. Since you will know the moles of NaOH reacted, you will know the moles of acid:

strong acid (204.23)

$$\underline{1}\,NaOH + \underline{1}\,HA \rightarrow H_2O + NaA$$



Determine the % mass of acid in the 0.5g sample:

$$0.0153\,\cancel{L} \times \frac{0.1\text{ mole NaOH}}{\cancel{L}} =$$

$$= 0.00153\text{ mole NaOH} = HA$$

$$= 0.00153\,\cancel{\text{mole}}\text{ HA} \times \frac{204.23\text{g}}{\cancel{\text{mole}}} =$$

$$= 0.3125\text{g HA}$$

$$\%\text{ mass} = \frac{0.3125\text{g}}{0.5\text{g}} \times 100 = 63\%$$

4-5 drops bromthymol blue (BTB) => yellow → green → blue

color indicator that changes color when rxn is over.

acidic | pH=7 H_2O | basic

Net Ionic Equations

To write a net ionic equation first write the complete ionic equation. The complete ionic equation is written by separating any aqueous strong acids or soluble compounds into their constituent ions. Do not separate pure substances or ionic solids or weak acids into ions.
(insoluble)
not

Example 1

For the neutralization rxn:

$$HCl(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCl(aq)$$

$$H^+ + \cancel{Cl^-} + \cancel{Li^+} + OH^- \longrightarrow H_2O + \cancel{Li^+} + \cancel{Cl^-}$$

$$H^+ + OH^- \rightarrow H_2O$$ Net ionic equation

Example 2

$$MgS(aq) + CuCl_2(aq) \rightarrow CuS(s) + MgCl_2(aq)$$

$$\cancel{Mg^{+2}} + S^{-2} + Cu^{+2} + \cancel{2Cl^-} \rightarrow CuS + \cancel{Mg^{+2}} + \cancel{2Cl^-}$$

$$Cu^{+2} + S^{-2} \rightarrow CuS$$ Net ionic equation

Example 3

$$HC_2H_3O_2(aq) + KOH(aq) \rightarrow H_2O(l) + KC_2H_3O_2(aq)$$

$$HC_2H_3O_2 + \cancel{K^+} + OH^- \rightarrow H_2O + \cancel{K^+} + C_2H_3O_2^-$$

$$HC_2H_3O_2 + OH^- \rightarrow H_2O + C_2H_3O_2^- \leftarrow$$ Net ionic equation

Example 4

$$2Na_3PO_4(aq) + 3NiCl_2(aq) \rightarrow Ni_3(PO_4)_2(s) + 6NaCl$$

$$(\cancel{Na^+} + 2PO_4^{-3} + 3Ni^{+2} + 6\cancel{Cl^-} \rightarrow Ni_3(PO_4)_2 + 6\cancel{Na^+} + 6\cancel{Cl^-}$$

$$3Ni^{+2} + 2PO_4^{-3} \rightarrow Ni_3(PO_4)_2$$

Example 5

$$2HBr(aq) + NiS(s) \rightarrow NiBr_2(aq) + H_2S(g)$$

$$2H^+ + 2\cancel{Br^-} + NiS \rightarrow Ni^{+2} + 2\cancel{Br^-} + H_2S$$

$$2H^+ + NiS \rightarrow Ni^{+2} + H_2S$$

Example 6

$$NH_4I_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(g)} + NaI_{(aq)}$$

$$NH_4^+ + \cancel{I}^- + \cancel{Na}^+ + OH^- \rightarrow H_2O + NH_3 + \cancel{Na}^+ + \cancel{I}^-$$

$$NH_4^+ + OH^- \rightarrow H_2O + NH_3$$