GAS FORMATION / OTHER MOLECULES

- There are a few other molecules that can be made with exchange-type chemistry.
- Most of these molecules are unstable and can break apart to form gases.

- Formation of a weak acid:

- The formation of ANY weak acid in an exchange-type reaction can be a driving force.
- Some weak acids are unstable and can break apart into gas molecules.

$$H_2CO_3(aq) \longrightarrow H_2O(l) + \boxed{CO_2(g)}$$

Gas bubbles can leave solution!

... but how would you form carbonic acid in an exchange-type reaction?

acid + carbonate CO_3^{2-}

OR

acid + bicarbonate HCO_3^-

$$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2\boxed{H_2CO_3(aq)}$$

H^+ SO_4^{2-} Na^+ HCO_3^-

$$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + \boxed{2H_2CO_3(aq)}$$

$$H^+ \quad SO_4^{2-} \quad Na^+ \quad HCO_3^-$$

... but when we mix sulfuric acid and sodium bicarbonate, we observe BUBBLES. We need to write an equation that agrees with our observations. We know that carbonic acid decomposes, so we go ahead and put that into our equation.

$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$$

$$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$$

Other molecules of interest:

H_2SO_3: sulfurous acid - React an ACID with a SULFITE

$$H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$$

H_2S: hydrogen sulfide (gas) - React an ACID with a SULFIDE

$$H_2SO_4(aq) + Na_2S(aq) \rightarrow Na_2SO_4(aq) + H_2S(g)$$

A few more exchange examples:

*In EXCHANGE reactions, transition metals do not change their charge!

$CaCl_2(aq) + 2AgNO_3^{*}(aq) \longrightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$ ✓

$Ca^{2+}\ Cl^-\ \ \ Ag^+\ NO_3^-$

PRECIPITATION

$H_3PO_4(aq) + 3NaOH(aq) \longrightarrow 3H_2O(\ell) + Na_3PO_4(aq)$ ✓

$H^+\ PO_4^{3-}\ \ \ Na^+\ OH^-$

ACID/BASE or NEUTRALIZATION reaction, driven by formation of WATER.

$KCl(aq) + NaNO_3(aq) \longrightarrow$ ~~$KNO_3(aq) + NaCl(aq)$~~

$K^+\ Cl^-\ \ \ Na^+\ NO_3^-$

NO REACTION

Both products are soluble ionic compounds. That means that they are present in solution as free ions (just like they were before being mixed) - and there's no driving force for a reaction.

$H_2SO_4(aq) + Na_2CO_3(aq) \longrightarrow Na_2SO_4(aq) + \underline{\underline{H_2CO_3}}(aq)$

$H^+\ SO_4^{2-}\ \ \ Na^+\ CO_3^{2-}$

$\longrightarrow Na_2SO_4(aq) + H_2O(\ell) + CO_2(g)$

Driving force is the formation and release of carbon dioxide gas.

- Exchange reactions involve ions pairing up, but the ions themseves are not formed in exchange reactions. Exchanges start with pre-existing ions.

... but the ions have to be produced somehow - through a chemistry that involves the transfer of electrons.

- oxidation / reduction chemistry ("redox" chemistry) involves transfer of electrons and can make ions.

$$2Al(s) + 3Br_2(l) \longrightarrow 2AlBr_3(s)$$

Elemental, metallic aluminum. Uncharged!

Al^0

Aluminum cation

Al^{3+}

These are called "half-reactions"

$$Al \longrightarrow Al^{3+} + \underbrace{3e^-}_{\text{electron}}$$

oxidation: loss of electrons

$$Br_2 + 2e^- \longrightarrow 2Br^-$$

reduction: gain of electrons

- oxidation and reduction always occur together. In other words, we can't just make free electrons using oxidation without giving them somewhere to go.

- Many of the types of reactions that we learned about in previous courses are redox reactions!

  - COMBINATIONS (often but not always redox)

  - DECOMPOSITIONS (often redox)

  - SINGLE REPLACEMENT (always redox)

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

$$Cu \rightarrow Cu^{2+} + 2e^-$$ oxidation

$$2Ag^+ + 2e^- \rightarrow 2Ag(s)$$ reduction

net ionic → $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

  - COMBUSTION

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

$$2Mg(s) \rightarrow 2Mg^{2+} + 4e^-$$ oxidation

$$O_2(g) + 4e^- \longrightarrow 2O^{2-}$$ reduction

A review of the reaction types we just mentioned:

① COMBINATION REACTIONS

- Reactions that involve two or more simple substances COMBINING to form a SINGLE product

- Often involve large energy changes. Sometimes violent!

- Form: $\boxed{A + B + \ldots \longrightarrow C}$

Example:

$$2\,Al(s) + 3\,Br_2(l) \longrightarrow 2\,AlBr_3(s)$$

② DECOMPOSITION REACTIONS

- Reactions where a SINGLE REACTANT breaks apart into several products

- Form: $\boxed{A \longrightarrow B + C + \ldots}$

Example:

$$2\,H_2O_2(l) \longrightarrow 2\,H_2O(l) + O_2(g)$$

* This reaction is NOT a combustion reaction, even though O_2 is involved!

* Combustion reactions CONSUME O_2, while this reaction PRODUCES O_2

③ COMBUSTION REACTIONS

- Reactions of substances with MOLECULAR OXYGEN (O_2) to form OXIDES.
- Combustion forms an OXIDE of EACH ELEMENT in the burned substance!

- Form: $\boxed{AB + O_2 \longrightarrow AO + BO}$

Oxide: a compound containing OXYGEN and one other element!

* Combustion of hydrocarbons makes carbon dioxide and water, if enough oxygen is present. In low-oxygen environments, carbon monoxide is made instead!

Examples:

Oxides!

*

$$C_3H_8(g) + 5O_2(g) \longrightarrow 4H_2O(g) + 3CO_2(g)$$

$$\underbrace{2Mg(s) + O_2(g) \longrightarrow 2MgO(s)}$$

This reaction can also be called a combination!
Two reactants form a single product.

④ SINGLE REPLACEMENT REACTIONS

- Reactions where one element REPLACES another element in a compound.

- Can be predicted via an ACTIVITY SERIES (p151, 9th edition)

- Form: $\boxed{A + BC \longrightarrow AC + B}$ "A" and "B" are elements., often metals.

- Easy to spot, since there is an element "by itself" on each side of the equation.

Examples:

$$\overset{0}{Cu}(s) + 2\overset{+1}{Ag}NO_3(aq) \longrightarrow \overset{+2}{Cu}(NO_3)_2(aq) + 2\overset{0}{Ag}(s)$$

$$\overset{0}{Zn}(s) + \overset{+1}{H}_2SO_4(aq) \longrightarrow \overset{+2}{Zn}SO_4(aq) + \overset{0}{H}_2(g)$$

REDOX LANGUAGE

"oxidizer"

- "Oxidation" is loss of electrons, but an OXIDIZING AGENT is something that causes ANOTHER substance to lose electrons. An oxidizing agent is itself reduced during a redox reaction.

- "Reduction" is gain of electrons, but a REDUCING AGENT is something that causes ANOTHER substace to gain electrons. Reducing agents are themselves oxidized during a redox reaction.

$$2\overset{0}{Al}(s) + 3\overset{0}{Br_2}(l) \longrightarrow 2\overset{+3}{Al}\overset{-1}{Br_3}(s)$$

Aluminum is OXIDIZED during this process. We say that metallic aluminum is a REDUCING AGENT!

Bromine is REDUCED during this process. We say that bromine is an OXIDIZING AGENT!

* Strong oxidizers (oxidizing agents) can cause spontaneous fires if placed into contact with combustibles (safety issue!).

* Reactive metals tend to be REDUCING AGENTS, while oxygen-rich ions like NITRATES tend to be OXIDIZING AGENTS. HALOGENS (Group VIIA) also tend to be OXIDIZING AGENTS