A few more exchange examples: *

* Transition metals DO NOT change their charge in exchange reactions!

$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\underline{AgCl(s)}$

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

PRECIPITATION of AgCl drives this reaction!

$H_3PO_4(aq) + 3NaOH(aq) \rightarrow 3\underline{H_2O(l)} + Na_3PO_4(aq)$

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

Formation of WATER is the driving force. Detect this reaction by looking for HEAT RELEASE

$KCl(aq) + NaNO_3(aq) \rightarrow \cancel{KNO_3(aq) + NaCl(aq)}$ NO REACTION

K^+ Cl^- Na^+ NO_3^-

Both "products" are water-soluble ionic compounds; present in water as FREE ions (same as they were before being mixed). There's no DRIVING FORCE for a reaction here.

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

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

$H_2CO_3 \rightarrow H_2O + \boxed{CO_2}$

$H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + Na_2SO_4(aq)$

Driving force is the formation of CARBONIC ACID and its decomposition into WATER MOLECULES and CARBON DIOXIDE GAS (which escapes as bubbles)

OXIDATION / REDUCTION CHEMISTRY

- 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.

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

Elemental, metallic aluminum. Uncharged!

Al^0

Aluminum cation

Al^{3+}

These are called "half-reactions"

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

oxidation: loss of electrons

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

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)

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

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

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

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

- COMBUSTION

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

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

$$O_2(g) + 4e^- \rightarrow 2O^{2-}$$ 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:

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

$$\overset{0}{Zn}(s) + \overset{+1}{H_2}SO_4(aq) \longrightarrow \overset{+2}{Zn}SO_4(aq) + \overset{0}{H_2}(g)$$

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.

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

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

END OF CHAPTER 4 MATERIAL