CHEMISTRY NOTES - Chapter 8

Chemical Reactions

Goals: To gain an understanding of:

1. Writing and balancing chemical equations.
2. Types of chemical reactions.

Notes

A chemical reaction is a reaction in which a chemical change takes place, that is one or more substances are changed into one or more new substances. Energy is always involved in chemical changes - either being absorbed or released.

A chemical equation is a shorthand notation showing a chemical reaction using formulas. In its most simplest form it would look like:

\[ \text{reactants} \rightarrow \text{products} \]

Most chemical equations will look more like:

\[ \text{H}_2\text{O} + \text{CO}_2 \rightarrow \text{H}_2\text{CO}_3 \]

This equation would read "One molecule of water plus one molecule of carbon dioxide yields one molecule of carbonic acid."

Reactants are the substances changed in a chemical reaction and are shown to the left of the "yields" sign (arrow). Water and carbon dioxide are the reactants in the above chemical equation.

Products are the substances produced in a chemical reaction and are shown to the right of the yields sign. Carbonic acid is the product of the above reaction.

Skeleton equations do not show the relative amounts of reactants and products (are "unbalanced"). Balanced equations do show the relative amounts of the reactants and products.

A balanced equation has the same kinds and numbers of atoms in the reactants as in the products in the lowest whole number ratio.

Balancing equations is necessary to show:

- the law of conservation of mass - matter is neither created nor destroyed in ordinary chemical reactions
- that chemical reactions do not change elements from one to another - chemical reactions merely rearrange the atoms of the reactants into new combinations in the products

To balance equations always be sure that you start off with a correctly written skeleton equation - one that shows the correct formulas for the reactants and products. To balance the equation make sure you have the same kinds and numbers of atoms in the reactants as in the products.

Note: Only balance by changing the coefficients - numbers in front of reactants or products. These numbers indicate how many of the reactant or products particles there are. **DO NOT CHANGE THE FORMULAS OF THE CORRECTLY WRITTEN SKELETON EQUATION.**
Example 1: Solid aluminum reacts with oxygen gas to produce aluminum oxide.

Step 1 - Write the skeleton equation - be sure all reactants and products are written with the correct formula:

\[ \text{Al}(s) + \text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \]

Step 2 - Balance the equation by adding the appropriate coefficients (remember - do not change correct formulas):

There are 2 oxygen atoms in the reactants and 3 oxygen atoms in the products. The lowest common multiple is 6, so I will put the coefficient "3" in front of the \( \text{O}_2 \) in the reactants and the coefficient 2 in front of the \( \text{Al}_2\text{O}_3(s) \) of the products to balance the oxygen atoms:

\[ \text{Al}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Al}_2\text{O}_3(s) \]

There are now 4 aluminum atoms in the products so I will write a "4" in front of the aluminum in the reactants to balance the aluminum atoms:

\[ 4 \text{Al}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Al}_2\text{O}_3(s) \]

Step 3 - Check your work to make sure all elements are balanced and are in the lowest whole number ratio. It may be helpful to make a chart as follows:

<table>
<thead>
<tr>
<th>number of atoms in reactants</th>
<th>element</th>
<th>number of atoms in products</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>Al</td>
<td>4</td>
</tr>
<tr>
<td>6</td>
<td>O</td>
<td>6</td>
</tr>
</tbody>
</table>

Example 2: Aqueous aluminum sulfate reacts with sodium metal to produce aqueous sodium sulfate and aluminum metal.

Step 1. \( \text{Al}_2(\text{SO}_4)_{3(aq)} + \text{Na}_(s) \rightarrow \text{Na}_2\text{SO}_4_{(aq)} + \text{Al}_(s) \)

Step 2. The polyatomic sulfate ion remains unchanged so it can be treated as a single unit. There are 3 sulfate ions in the reactants and 1 sulfate ion in the products. Balancing these we get:

\[ \text{Al}_2(\text{SO}_4)_{3(aq)} + \text{Na}_(s) \rightarrow 3 \text{Na}_2\text{SO}_4_{(aq)} + \text{Al}_(s) \]

Balancing the aluminum atoms/ions we get:

\[ \text{Al}_2(\text{SO}_4)_{3(aq)} + \text{Na}_(s) \rightarrow 3 \text{Na}_2\text{SO}_4_{(aq)} + 2 \text{Al}_(s) \]

And balancing the sodium atoms/ions we get:

\[ \text{Al}_2(\text{SO}_4)_{3(aq)} + 6 \text{Na}_(s) \rightarrow 3 \text{Na}_2\text{SO}_4_{(aq)} + 2 \text{Al}_(s) \]

Step 3. Checking the answer we see that it is in the lowest whole number ratio and we do have equal kinds and numbers of atoms/ions on both sides:

<table>
<thead>
<tr>
<th>Number of atoms/ions in reactants</th>
<th>species</th>
<th>Number of atoms/ions in products</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>\text{SO}_4^{2-}</td>
<td>3</td>
</tr>
<tr>
<td>2</td>
<td>Al</td>
<td>2</td>
</tr>
<tr>
<td>6</td>
<td>Na</td>
<td>6</td>
</tr>
</tbody>
</table>
The solid state of reactant or product is designated with the subscript (s) - ex. Na\(_{(s)}\).

The liquid state of reactant or product is designated with the subscript (l) - ex. H\(_2\)O\(_{(l)}\).

The aqueous state (dissolved in water) of reactant or product is designated with the subscript (aq) - ex. NaCl\(_{(aq)}\).

The gaseous state of reactant or product is designated with the subscript (g) - ex. Cl\(_2(g)\).

A catalyst is a substance which speeds up a chemical reaction, but remains unchanged. Since it remains unchanged it is not considered a reactant or product and is written above the yields sign. An example of a catalyst is manganese dioxide (MnO\(_2\)) which speeds up the decomposition of hydrogen peroxide (H\(_2\)O\(_2\)) into water and oxygen gas.

\[
\text{MnO}_2 \\
2 \text{H}_2\text{O}_{(aq)} \rightarrow 2 \text{H}_2\text{O}_{(l)} + \text{O}_2(g)
\]

**Combination reactions or Synthesis reactions** - chemical reactions in which two or more reactants combine to form a single product.

- Combination reactions usually are exothermic - give off energy
- General equation: A + B \(\rightarrow\) AB
- Example: 4 Na\(_{(s)}\) + O\(_2(g)\) \(\rightarrow\) 2 Na\(_2\)O\(_{(s)}\)

**Decomposition reactions** - chemical reactions in which a single compound is broken down into two or more simpler products.

- Most decomposition reactions require energy in the form of heat, light or electricity (at least to start the reaction)
- General equation: AB \(\rightarrow\) A + B
- Example CaCO\(_3(s)\) + heat \(\rightarrow\) CaO\(_(s)\) + CO\(_2(g)\)

**Single replacement reactions** - chemical reactions in which atoms of a more reactive element replace the atoms of a second element in a compound.

- Which elements are more reactive are determined experimentally and can be arranged in an activity series (note Table 7-2 Activity Series of Metals on p. 155). the more reactive metals will replace the less reactive metals, but not vice versa.
- Halogen activity series (in order of decreasing activity): F\(_2\), Cl\(_2\), Br\(_2\), I\(_2\) (e.g Fluorine will replace Cl\(_2\), Br\(_2\) or I\(_2\), Chlorine will replace Br\(_2\) or I\(_2\) and so on, but not vice versa.
- The hydrogen cation of water (H\(^+\)OH\(^-\)) is replaced by metals Li through Na of Activity series of metals to form metallic hydroxides
- General equation: AB + C \(\rightarrow\) AC + B or AB + X \(\rightarrow\) XB + A
- Example 1: MgCl\(_2(s)\) + F\(_2(g)\) \(\rightarrow\) MgF\(_2(s)\) + Cl\(_2(g)\)
- Example 2: MgCl\(_2(s)\) + 2 Na\(_{(s)}\) \(\rightarrow\) 2 NaCl\(_{(s)}\) + Mg\(_{(s)}\)
- Example 3: MgCl\(_2(s)\) + I\(_2(g)\) \(\rightarrow\) No reaction (iodine is less reactive then chlorine)
Double-replacement reactions - chemical reactions which involve an exchange of positive ions between two compounds.

- For a double replacement reaction to occur one of the products usually leaves the reaction medium by forming an insoluble precipitate, forming a gas or by forming a molecular compound such as water.
- General equation: \( AB + XY \rightarrow XB + AY \)
- Example 1 (formation of precipitate):
  \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq) \)
- Example 2 (formation of a gas):
  \( \text{FeS}(s) + 2 \text{HCl}(aq) \rightarrow \text{H}_2\text{S}(g) + \text{FeCl}_2(aq) \)
- Example 3 (formation of a molecular compound):
  \( \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \)

Combustion reactions - chemical reactions in which oxygen reacts with another substance.

- Often release energy in the forms of heat and light
- Often involve hydrocarbons (compounds made up of hydrogen and carbon) and form water and carbon dioxide (may form carbon monoxide and/or elemental carbon [soot] if not enough oxygen is available)
- General equation for the combustion of a hydrocarbon:
  \( \text{C}_x\text{H}_y + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \)
- Example: Combustion of butane (in “Bic” lighters)
  \( \text{C}_4\text{H}_{10}(g) + 13 \text{O}_2(g) \rightarrow 8 \text{CO}_2(g) + 10 \text{H}_2\text{O}(l) \)

Many ionic compounds dissociate (split apart into ions) when they dissolve in aqueous (water) solution. Ionic equations are equations that show those reactants and/or products in their dissociated states. Net ionic equations show only the ions which undergo a change in a reaction. Here is an example showing a traditional (unionized equation), ionic equation and a net ionic equation. The symbol (aq) indicates an aqueous (dissolved in water) state and the symbol (s) indicates a solid (formed by being precipitated out of solution).

**Traditional equation:** \( 2 \text{AgNO}_3(aq) + \text{BaCl}_2(aq) \rightarrow 2 \text{AgCl}(s) + \text{Ba(NO}_3)_2(aq) \)

**Ionic equation:**

\[
2 \text{Ag}^+(aq) + 2 \text{NO}_3^-(aq) + 2 \text{Ba}^{2+}(aq) + 2 \text{Cl}^-(aq) \rightarrow 2 \text{AgCl}(s) + \text{Ba}^{2+}(aq) + 2 \text{NO}_3^-(aq)
\]

**Net ionic equation:**

\[
\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)
\]

Site for more info and practice on ionic and net ionic equations:
http://www.chem.vt.edu/RVGS/ACT/notes/Notes_on_Net_ionic_rxns.html