Chapter 8: Chemical Reactions - Balancing Chemical Equations and Reaction Types

Topics:

- Physical vs. chemical changes

-Evidence of chemical reactions

-Writing and balancing chemical equations

-Reaction Types:
- -Combination (synthesis)
- -Decomposition
- -Single Replacement
  - -Activity Series
- -Double Displacement (precipitation; metathesis)
  - -Solubility Rules
- -Neutralization (acid-base)

Physical vs. Chemical Changes

Physical Change: The composition of the substance remains unchanged.

Examples: Melting ice; crushing a can, cutting a piece of wood; breaking a rock.

Chemical Change: Composition of a substance is changed. Atoms are rearranged and a chemical reaction takes place.

Examples: Burning a match; taking an antacid; cooking a steak; making wine.





Energy changes are common in chemistry and account for the color of fireworks.

CHEMICAL	FIREWORKS COLOR
Na compounds	orange-yellow
Ba compounds	yellow-green
Ca compounds	red-orange
Sr compounds	bright red
Li compounds	scarlet-red
Cu compounds	blue-green
Al or Mg metals	white sparks

The electrons of metals in ionic compounds are energized when the firework explodes and give off a characteristic color when the electron loses energy again.

Evidence of Chemical Reactions (Driving forces for chemical reactions)

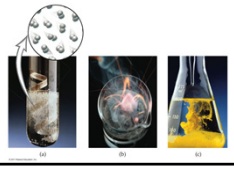
1. A gas is produced.

2. An insoluble compound is produced from the reaction between two soluble compounds.

3. A permanent color change is observed.

4. An energy change is observed (temperature change, fire, etc.)

5. Water is produced.



Chemical Reaction Equations (Represents a rearrangement of atoms)

Reaction arrow

Reactants
The substances taking part in the reaction

Products
The substances that are produced in the reaction

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

Subscript: Gives the *quantity* ratio of one element to another in a compound

State of matter: solid = (s); liquid = (l); gas = (g); aqueous = (aq)

Coefficient: Gives the relative *quantity* of a substance that takes part in a reaction (moles or molecules).

The quantity and type of each element must be "conserved" on both sides of the reaction equation.

Symbology of Reactions

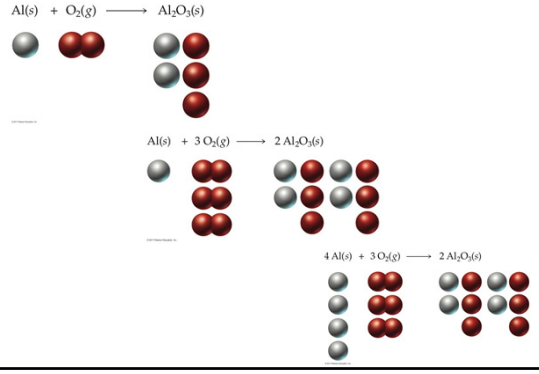
TABLE 8.1 CHEMICAL EQUATION SYMBOLS

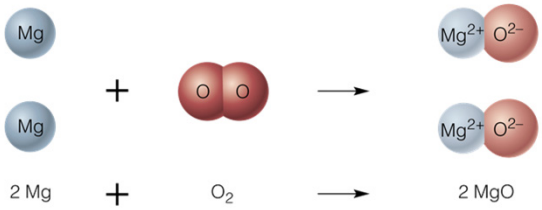
SYMBOL	INTERPRETATION OF CHEMICAL EQUATION SYMBOL
$\rightarrow$	produces, yields, gives (separates reactants from products)
+	reacts with, added to, plus (separates two or more reactants or products)
$\xrightarrow{\Delta}$	heat is a catalyst for the reaction
$\xrightarrow{Fe}$	iron is a catalyst for the reaction
NR	no reaction
(*s*)	solid substance or precipitate
(*l*)	liquid substance
(*g*)	gaseous substance
(*aq*)	aqueous solution

A catalyst speeds up a reaction without being consumed or permanently altered. It can provide an alternate reaction path.

It's easier to travel through the valley than over the mountain.

Progression of the balancing of aluminum reacting with oxygen to produce aluminum oxide.





Note that these reactions are all based on quantities, not masses.

(Don't forget about diatomic molecules, like in oxygen above.)

Writing and Balancing Chemical Equation:

1. Mass must be conserved in the chemical reaction therefore the number and type of atoms on each side of the reaction equation must be balanced.
2. Chemical equations MUST be balanced by coefficients which go in the front of a compound's formula.
3. The chemical formula for a compound must not be altered:

E.g.

Don't turn H_2O into H_2O_2 or H_23O (no coefficients in the middle of a compound)

4. Reaction equations are balanced by "inspection". It is a good idea to take inventory on the number and type of elements on each side of the reaction equation.

Description:
Magnesium metal reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

(Reaction Clip Link)

Reaction Equation (unbalanced):
$Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Reaction Equation (balanced):
$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Example:
Balance the following reactions:

$Re_{(s)} + Br_{2(l)} \rightarrow ReBr_{3(s)}$

$Ca(AlO_2)_2 + HCl \rightarrow AlCl_3 + CaCl_2 + H_2O$

Answers:

$2Re_{(s)} + 3Br_{2(l)} \rightarrow 2ReBr_{3(s)}$

Inventory:

Re – 2	Re – 2
Br – 6	Br – 6

$Ca(AlO_2)_2 + 8HCl \rightarrow 2AlCl_3 + CaCl_2 + 4H_2O$

Inventory:

Ca – 1	Ca - 1
Al – 2	Al - 2
O – 4	O - 4
H – 8	H - 8
Cl – 8	Cl – 8

Note: AlO_2^- is the polyatomic ion aluminate

Sometimes you will have to write the equation first before you balance:

Examples:

Write and balance the following reaction equations

1. Copper combines with sulfur to form Copper(I)sulfide.

(Note: Elemental sulfur exists as S_8)

2. Silver nitrate reacts with sulfuric acid to produce silver sulfate and nitric acid.

Answers:

1. $16Cu + S_8 \rightarrow 8Cu_2S$

2. $2AgNO_3 + H_2SO_4 \rightarrow Ag_2SO_4 + 2HNO_3$

Sample Balancing

1. Calcium hydroxide reacts with phosphoric acid to produce calcium phosphate and water

2. Hydrogen reacts with nitrogen monoxide to produce water and nitrogen.

3. Potassium phosphate and aluminum nitrate react to produce potassium nitrate and aluminum phosphate.

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Answers:

$3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6H_2O$

$2H_2 + 2NO \rightarrow 2H_2O + N_2$

$K_3PO_4 + Al(NO_3)_3 \rightarrow 3KNO_3 + AlPO_4$

Reaction Types

Combination (synthesis) Reactions:

$$A + Z \rightarrow AZ$$

Simpler substances combining into something more complex.

Example: $Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$

Decomposition Reactions:

$$AZ \xrightarrow{\Delta} A + Z$$

Complex compound being broken into simpler substances. Heat (Δ) is commonly involved.

Example: $2C_6H_2(NO_2)_3CH_{3(s)} \rightarrow 12CO_{(g)} + 5H_{2(g)} + 3N_{2(g)} + 2C_{(s)}$

Specific Decomposition Reactions:

Metal bicarbonates thermally break down into metal carbonates, water and carbon dioxide

Example:
$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + H_2O_{(g)} + CO_{2(g)}$

Metal carbonates thermally break down into metallic oxides and carbon dioxide

Example:
$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

Oxygen containing compounds can thermally break down into another substance with the release of oxygen gas

Example:
$2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

Reaction Types

Single Replacement Reaction:

$$A + BZ \rightarrow AZ + B$$

Simpler substances combining into something more complex.

Example: $Mg_{(s)} + 2AgNO_{3(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2Ag_{(s)}$

Reactions Predicted by Activity Series (Lab manual page 23)

Activity Series of Select Elements:

(Least Active) Au, Hg, Ag, Cu, H_2, Pb, Sn, Ni, Cd, Fe, Zn, Cr, Al, Mg, Na, Ca, K **(Most Active)**

Note that in the example above, magnesium can replace silver, but silver could not replace magnesium in a compound because it is less active.

The more active element always ends up as the ion.

1. If the more active element is already the ion, no reaction occurs.
2. If the more active element is not the ion, it becomes one, forcing the other element back to its elemental zero charge state.

Problem: Using the activity series below, predict what will happen for the following:
a. $Zn_{(s)} + Mg(NO_3)_{2(aq)} \rightarrow$
b. $Al_{(s)} + CuSO_{4(aq)} \rightarrow$

Activity Series of Select Elements:

(Least Active) Au, Hg, Ag, Cu, H_2, Pb, Sn, Ni, Cd, Fe, Zn, Cr, Al, Mg, Na, Ca, K **(Most Active)**

Reaction Types

Double Displacement Reactions:

$$AX + BZ \rightarrow AZ + BX$$

The reactants AX and BZ must be soluble salts in solution [aqueous (aq)] and at least one of the products must be insoluble (s).

Reactions Predicted by Solubility Rules (Lab manual page 23)

Solubility Characteristics of Ionic Compounds:

Soluble Compounds	
Almost all salts of Na^+, K^+, NH_4^+ Salts of: Nitrate, NO_3^- chlorate, ClO_3^- perchlorate, ClO_4^- acetate, $CH_3CO_2^-$	**Exceptions**
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Insoluble Compounds	Exceptions
Most salts of carbonate, CO_3^{2-} phosphate, PO_4^{3-} oxalate, $C_2O_4^{2-}$ chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations $Ba(OH)_2$ is soluble
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

Example: $AgNO_{3(aq)} + CaCl_{2(aq)} \rightarrow 2AgCl_{(s)} + Ca(NO_3)_{2(aq)}$

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Problem: Using the chart above, predict what will happen when combining solution of:
a. potassium nitrate and sodium fluoride.
b. strontium chloride and lithium sulfate.

Reaction Types

Neutralization (acid-base) reaction:

$$HX + BOH \rightarrow BX + HOH$$

Where HX is an acid (e.g. HCl or HNO_3); BOH is a base (e.g. NaOH or KOH); BX is a salt and HOH is H_2O

Remember: Acids have leading hydrogens that produce H^+ (hydrogen ions) and bases contain or produce OH^- (hydroxide ions) in solution.

Acid + Base $\rightarrow$ Salt and Water

Example:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Problem:
Predict the products in the reaction between nitric acid and potassium hydroxide.
Name the salt that is formed.

Combustion Reaction

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$

Hydrocarbons react with oxygen to produce carbon dioxide and water

Things that "burn" are going through a combustion reaction.

Examples:

Natural gas Gasoline Wood Respiration