Chapter 5 – Chemical Reactions

5.3 Chemical Reactions – Evidence of a Chemical Change

Chemical reactions involve rearrangement and exchange of atoms to produce new molecules

Remember: matter can neither be created nor destroyed. The atoms are simply rearranged to produce a new substance.

For example:

HCl(aq) + NaOH(aq) → HOH(l) + NaCl(aq) (+ heat)

Chemical reactions require a driving force.
The most common driving forces are:

- formation of a solid (precipitation reaction or double displacement)
- formation of water (acid / base reaction)
- transfer of electrons (oxidation / reduction reaction)
- formation of a gas

Chemical reactions will be classified into broad categories based on the driving force or product involved later in this chapter.

Evidence of a chemical change:

- Visual clues (permanent)
  - color change
  - precipitate formation
  - gas bubbles
  - flames
  - heat release
  - cooling (heat absorbed)
  - light

- Other clues
  - new odor
  - permanent new state (solid, liquid, or gas)

Examples:
Nearly all chemical reactions involve some sort of transfer of energy, sometimes in the form of heat.

Heat is produced: Methane + Oxygen:

Figure 6.3d: Methane gas reacts with oxygen to produce a flame in a bunsen burner.

$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ + heat + light

Heat is produced:

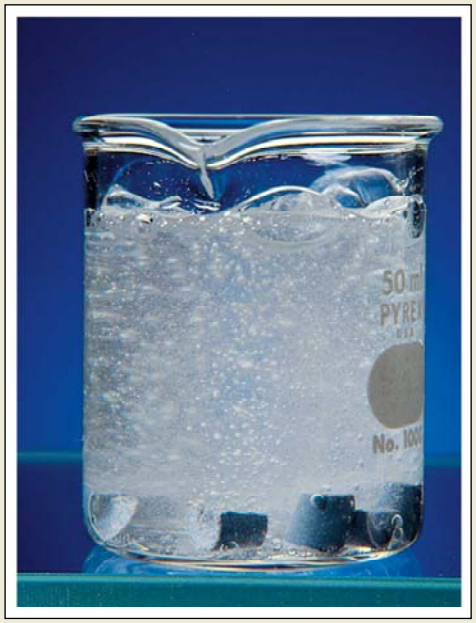
A hot pack used to warm hands and feet in winter. When the package is opened, oxygen from the air penetrates a bag containing solid chemicals. The resulting reaction produces heat for several hours.

Heat is absorbed:

An injured girl wearing a cold pack to help prevent swelling. The pack is activated by breaking an ampule; this initiates a chemical reaction that absorbs heat rapidly, lowering the temperature of the area to which the pack is applied.

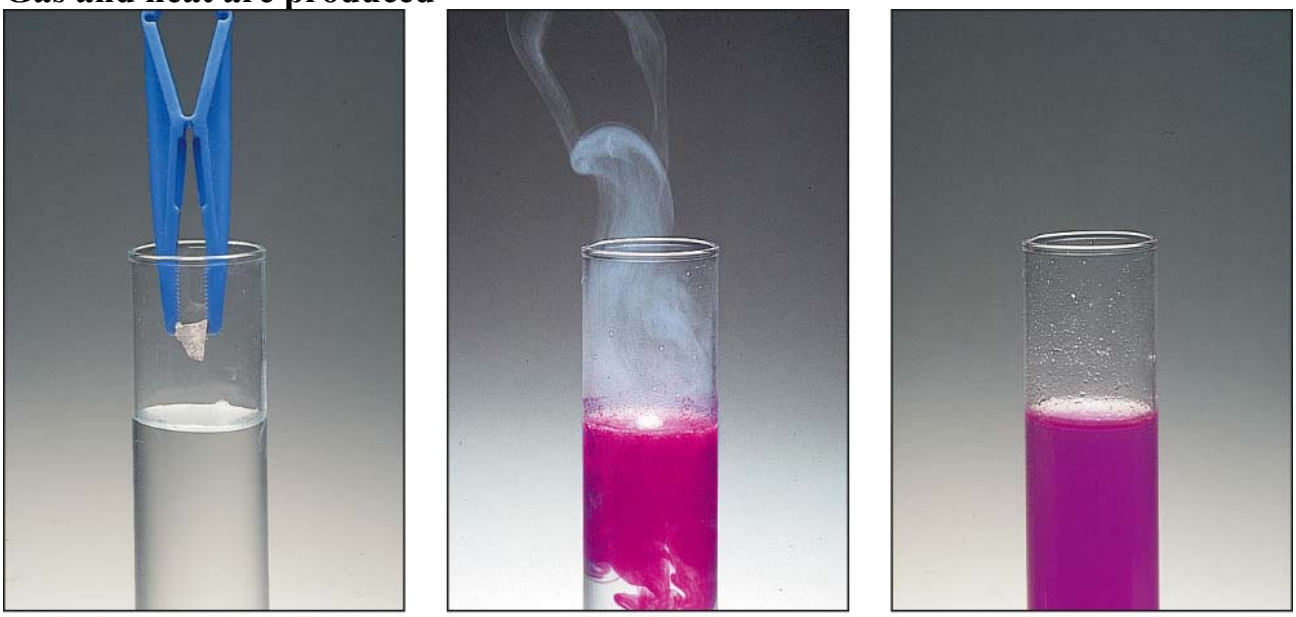
Gas is produced:

Figure 6.3c: Bubbles of hydrogen gas form when calcium metal reacts with water.



$$Ca(s) + H_2O(l) \rightarrow H_2(g) + Ca(OH)_2(aq)$$

Gas and heat are produced



(a) (b) (c)

Sodium metal reacts with water containing a colorless dye, phenolphthalein. The reaction generates heat and sodium hydroxide. The dye reacts with hydroxide ion (OH-) to produce a pink colored compound.

- $Na(s) + H_2O(l) \rightarrow H_2(g) + NaOH(aq)$
- $OH^- +$ phenolphthalein $\rightarrow$ pink color

Gas is produced:

Figure 6.1: Bubbles of hydrogen and oxygen gas form when an electric current is used to decompose water.

$H_2O(l) + e- \rightarrow H_2(g) + O_2(g)$

A solid (precipitate) forms:

Figure 6.3b: A solid forms when a solution of sodium dichromate is added to a solution of lead nitrate.

$$Na_2Cr_2O_7(aq) + Pb(NO_3)_2(aq) \rightarrow PbCr_2O_7(s)\downarrow + NaNO_3(aq)$$

5.4 Chemical Equations

- Shorthand way of describing a reaction

- Provide information about the reaction
  - **Qualitative** Information:
    - Formulas of reactants and products
    - **States** of reactants and products (solid, liquid, gas)
  - **Quantitative** Information:
    - Relative numbers of reactant and product molecules that are required
    - Can be used to determine weights of reactants used and of products that can be made (see Chapter 7: Chemical Formula Relationships)

Writing a chemical equation

First: some symbols used in writing equations.

TABLE 5.5 Some Symbols Used in Writing Equations

Symbol	Meaning
+	Separates two or more formulas
$\longrightarrow$	Reacts to form products
Δ	The reactants are heated
(*s*)	Solid
(*l*)	Liquid
(*g*)	Gas
(*aq*)	Aqueous

Begin with a ***"Word"*** equation:

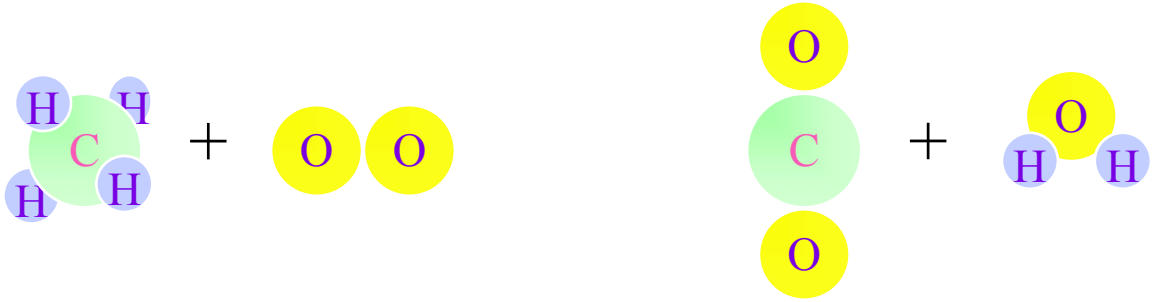
Example: *Methane gas burns to produce carbon dioxide gas and liquid water:*

(Note: whenever something **burns** it combines with O_2(g), so O_2 is a reagent)

- Identify the reagents (what you start with) and the products (what you end up with)
- Determine the chemical formula for each
- Reagents and reactants are separated by an arrow
- Reagents are written to the left of the arrow, and products to the right of the arrow

The equation for this chemical reaction is:

$$\mathbf{CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(\mathit{l})}$$



<u>Writing a chemical equation: key points</u>

1. Identify the elements or molecules involved in the reaction
   - identify diatomic elements
   - identify polyatomic ions

2. Write the equation as a word equation first, and then as a formula

3. Symbols are used after the chemical formula to indicate the state
   - (g) = gas; (*l*) = liquid; (s) = solid
   - (aq) = aqueous, dissolved in water

Example:

- Sodium reacts with water to produce hydrogen gas and sodium hydroxide solution, and heat.
- $Na + H_2O \rightarrow H_2 + NaOH$
- $Na(s) + H_2O(l) \rightarrow H_2(g) + NaOH(aq) + heat$

Problem:

Zinc metal added to aqueous hydrogen chloride forms gaseous hydrogen and zinc chloride.

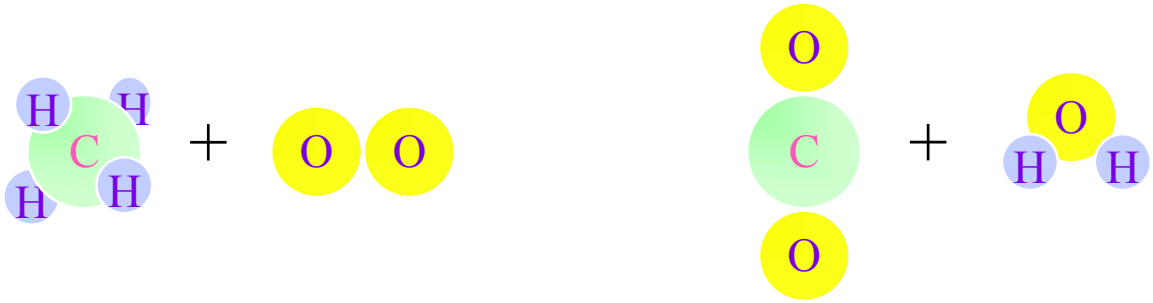
Problem:

An aqueous solution of hydrogen fluoride reacts with silicon dioxide to form gaseous silicon tetrafluoride plus liquid water.

Balancing Chemical Equations

Recall the following equation:

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$



Do you notice that something is wrong with the above equation, and also some of the equations written earlier?

What's wrong with this equation as written?

It's only a ***qualitative*** expression of the chemical reaction.

This equation as written violates the law of conservation of matter. It must be ***balanced*** so that the correct numbers of atoms are on each side of the equation.

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O(l)$$

Now, this equation is both a qualitative and quantitative expression of this chemical reaction.

Balancing a chemical equation: key points

The proper equation should be balanced

- obey Law of Conservation of Mass
- all elements found on reactants side are also on product side
- equal **numbers** of each element are on the reactant side as well as on product side

Specific steps:

Balancing equations is a trial and error process. One approach:

- Start with the compound with the greatest number of elements
- Wait to balance uncombined elements (e.g. Na, O_2, etc)
- Add coefficients in front of other ***compounds*** to balance the elements in the starting compound
- Add coefficients in front of uncombined elements
- Clear any fractional coefficients by multiplying all coefficients by the lowest common denominator
- Remove any "1" coefficients
- Double check the number of each element on both sides of the equation

Don't ……..

- Change a correct formula in order to make an element number balance
- Insert a coefficient within the correct formula to balance an element

Writing and balancing a chemical equation

Example: *Magnesium metal burns to form a white powder, magnesium oxide.*

Note: Burning always means a reaction with oxygen (O_2)

- Write out the chemical equation in words
- Identify the state of each chemical

magnesium(s) + oxygen(g) $\rightarrow$ magnesium oxide(s)

Write the equation as a formula with symbols of the elements

- identify diatomic elements
- identify polyatomic ions
- determine formulas

$$\mathbf{Mg(s) + O_2(g) \rightarrow MgO(s)}$$

Employ coefficients to balance the number of elements on both sides of the reaction arrow.

$$\mathbf{Mg(s) + O_2(g) \rightarrow MgO(s)}$$
$$\mathbf{1 \leftarrow Mg \rightarrow 1}$$
$$\mathbf{2 \leftarrow O \rightarrow 1}$$

- Multiply MgO by 2
- Then go back and multiply Mg by 2
- Verify that the number of each element are the same on both sides of the equation

Example: *Ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide and gaseous water*

- Write the equation in words
- Identify the state of each chemical

ammonia(g) + oxygen(g) $\rightarrow$ nitrogen monoxide(g) + water(g)

Write the equation as a formula

- identify diatomic elements
- identify polyatomic ions
- determine formulas

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$$

Employ coefficients to balance the number of elements on both sides of the reaction arrow.

All are balanced except H. *Balance H by swapping the H subscripts and making them coefficients.*

$$2\ NH_3(g) + O_2\ (g) \rightarrow NO(g) + 3\ H_2O(g)$$

Multiply NO by 2 to balance N

$$2\ NH_3(g) + O_2\ (g) \rightarrow 2NO(g) + 3\ H_2O(g)$$

2 x 3		**← H →**		**2 x 3**
2		**← N → 2**		
	2	**← O → 2**	**+**	**3**

Now, O remains unbalanced.

- A trick of the trade, when you are forced to attack an element that is in 3 or more compounds – find where it is uncombined. You can find a factor to make an uncombined element any amount you want, even if that factor is – *temporarily* – a fraction!
- We want to make the O on the left equal 5, therefore we will multiply it by 2.5

$$2\ NH_3(g) + 2.5\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$$

2 ← N →2

6 ← H → 6

2.5 x 2 ← O → 2 + 3

Finally, multiply all the coefficients by a number to eliminate fractional molecules.

$$4\ NH_3(g) + 5\ O_2\ (g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$

$$4 \leftarrow N \rightarrow 4$$
$$12 \leftarrow H \rightarrow 12$$
$$10 \leftarrow O \rightarrow 10$$

For practice, return to the "writing equations" examples above and balance the equations.

For a tutorial on balancing equations, go to:

http://www.preparatorychemistry.com/Bishop_Balancing_Equations_frames.htm

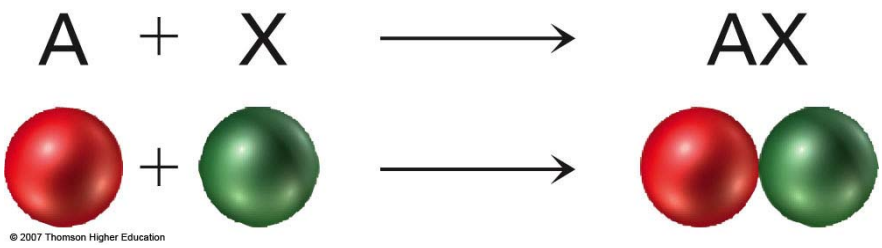
5.5 Writing Chemical Equations: *Classification of Reactions*

Reactions can be classified into several broad categories.

Note that categories frequently overlap, i.e. more than one classification can apply.

Combination (Synthesis) Reactions

- Occurs when the number of products is fewer than the number of reactants.
- The reactants are elements
- In symbolic form:



Example: the formation of water:

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(l)}$$

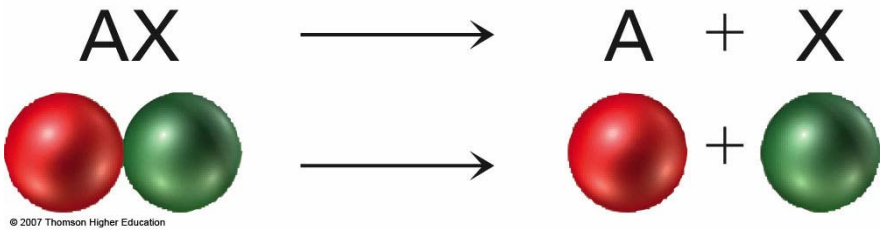
Note that in addition to classification as a combination or synthesis reaction, this reaction can also be classified [as discussed below] as:

- Oxidation – reduction (because they involve oxygen or a transfer of electrons)
- Combustion (because a flame is produced)

Problem:
Nitrogen gas plus oxygen gas react to form nitrogen oxide.
Write the balanced equation.

Decomposition Reactions

- The opposite of a combination reaction
- This occurs when the number of products is more than the number of reactants.
- The compounds are broken down to simpler compounds, or all the way down to the elements.
- In symbolic form:



Example: the decomposition of water $2\ H_2O_{(l)} \rightarrow 2\ H_{2(g)} + O_{2(g)}$

Example: the decomposition $PbO_{2(s)} \rightarrow Pb_{(s)} + O_{2(g)}$

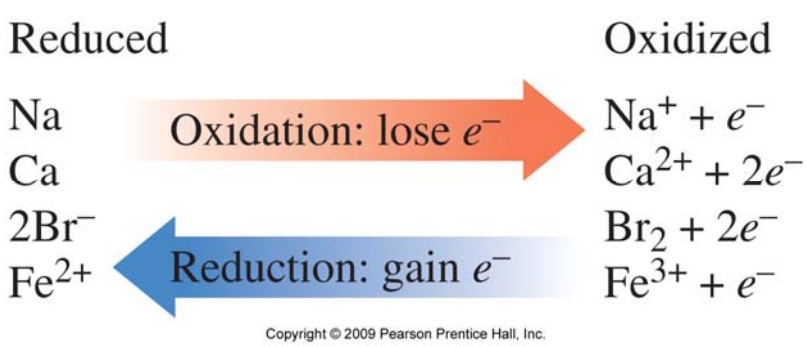
- Note that this is also an oxidation – reduction reaction: the Pb^{4+} and O^{2-} ions are changed from ions to the uncharged elements by the transfer of electrons.

Problem:
Solid calcium carbonate decomposes into calcium oxide and carbon dioxide gas. Write the balanced equation.

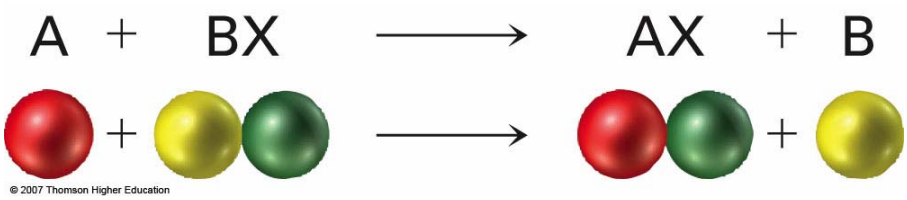
5.6 Single Replacement (or Single Displacement) Oxidation-Reduction Reactions

All reactions that involve a transfer of one or more electrons are called **oxidation-reduction reactions**. One type of oxidation-reduction reaction involves a metal and a non-metal.

- The metal loses electrons and becomes a cation
- This process is called **oxidation**
- The nonmetal gains electrons and becomes an anion
- We call this process **reduction**
- In the reaction, electrons are transferred from the metal to the nonmetal
- One element replaces another element in a compound.



In symbolic form:



We say that the substance that loses electrons in the reaction is **oxidized** and the substance that gains electrons in the reaction is **reduced.**

An acronym for remembering this is: ***OIL RIG***
OIL – Oxidation is a Loss (of electrons)
RIG – Reduction is a Gain (of electrons)

Oxidation-reduction reactions can involve two metals. It may be necessary to consult an "activity series" table to predict if a reaction may take place. A metal listed above another metal in the table will replace the metal below it

Example:

$$\mathbf{2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3}$$

(Al oxidized → Al^{3+} , Fe^{3+} reduced → Fe)

This type of reaction is sometimes called a "redox" reaction.

Table 9.2 Activity Series

Li K Ba Sr Ca Na	Will replace H_2 from liquid water, steam, or acid
Mg Al Mn Zn Cr	Will replace H_2 from steam or acid
Fe Ni Sn Pb	Will replace H_2 from acid
H_2	
Sb Cu Hg Ag Pd Pt Au	Will not replace H_2 from liquid water, steam, or acid

Example:

$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

Example:

$Na(s) + H_2O(l) \rightarrow H_2(g) + NaOH(aq)$

Example *(also a synthesis reaction)*:

$2Na + Cl_2 \rightarrow 2NaCl$ (Na oxidized → Na^+ , Cl_2 reduced → Cl^-)

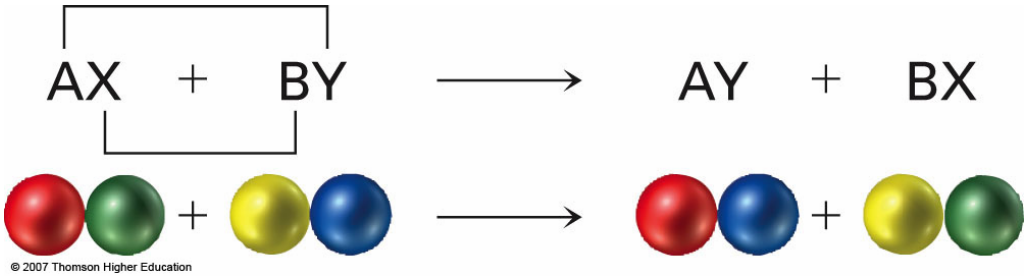
Problem:

Copper reacts with silver nitrate (assume copper(II) is formed). Write the equation.

Note: Oxidation – reduction reactions will be discussed in greater detail in Chapter 19.

Double Replacement (or Double Displacement): Precipitation Reactions

- Two sets of ions in ionic compounds switch places in a reaction.
- If one of the products precipitates, this double displacement reaction is known more specifically as a **precipitation reaction**
- If the reactants are an acid and a base, this double displacement reaction is known as an **acid-base reaction**
- The general equation in symbolic form:



More accurately, representing the compounds as they exist in solution:

- $AX(aq) + BY(aq) \rightarrow AY(s) + BX(aq)$

Either BX or AY may be insoluble and precipitate

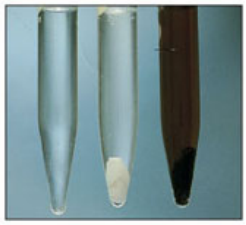
Example:

$$2\ NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\ NaNO_{3(aq)} + PbCl_{2(s)}$$

Note: The reaction will not occur unless one of the products either
(1) precipitates, or
(2) is water

We can determine in advance whether a chemical reaction will proceed by producing an insoluble product (precipitate). We do this be consulting a table of solubility guidelines

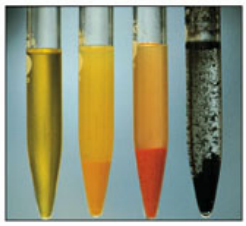
SILVER COMPOUNDS



$AgNO_3$ AgCl AgOH

(a) Nitrates are generally soluble, as are chlorides (except AgCl). Hydroxides are generally not soluble.

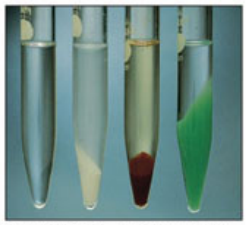
SULFIDES



$(NH_4)_2S$ CdS Sb_2S_3 PbS

(b) Sulfides are generally not soluble (exceptions include salts with NH_4^+ and Na^+).

HYDROXIDES



NaOH $Ca(OH)_2$ $Fe(OH)_3$ $Ni(OH)_2$

(c) Hydroxides are generally not soluble except when the cation is a Group 1A/1 metal.

SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^-; chlorate, ClO_3^-; perchlorate, ClO_4^-; acetate, $CH_3CO_2^-$	Acetates of Al^{3+}, Ag^+
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
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	$Ba(OH)_2$ is soluble

Example:

$$\mathbf{HCl(aq) + NaOH(aq) \rightarrow HOH(l) + NaCl(aq) \quad (+\ heat)}$$

Note: This double displacement is also classified as

- an acid – base reaction
- a synthesis reaction

Note: Double replacement precipitation reactions are discussed in greater detail in Chapter 9 under Net Ionic Equations.

Acid – Base Reactions

When the reaction is between an acid (a compound that produces hydrogen ions) and a base (any compound that forms hydroxide ions), water is formed as one of the products. This specific type of double displacement reaction is also called a Neutralization reaction.

General example: **acid + base → water + ionic compound**

Example: $H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ H_2O_{(l)} + Na_2SO_{4(aq)}$

Problem:
Write the balanced equation for the reaction of sulfuric acid and solid aluminum hydroxide.

Note: Acid – base reactions will be discussed in greater detail in Chapter 17.

Combustion

Rapid reactions that involve oxygen and produce a flame or heat are called combustion reactions. Typically, combustion occurs when a hydrocarbon reacts with oxygen to produce carbon dioxide and water, and heat, usually in the form of a flame. Hydrocarbons are a class of compounds that primarily consist of hydrogen and carbon.

Example:

General example: $C_xH_y + O_2 \rightarrow CO_2 + H_2O$

Actual example: $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

This is also an oxidation – reduction reaction (see Chapter 9.6 and Chapter 19).

You should be familiar with the following common chemical reactions:

- acid + base → water + ionic compound
- metal + oxygen → ionic compound
- metal + acid → hydrogen gas + ionic compound
- ionic compound1 + ionic compound2 → ionic compound3 + ionic compound4
- acid + carbonate → carbon dioxide + water + ionic compound
- metal1 + ionic compound1 → metal2 + ionic compound2
- hydrocarbon + oxygen → carbon dioxide + water

Exercise:

ClassificationTypesOfReactions – Ch5 web page exercise with answers