CHEMICAL REACTIONS

Chapter Seven

What is a Chemical Reaction?

- A chemical reaction involves the conversion of one or more substances into one or more different substances.
- The substance(s) which we begin with are called the reactant(s)
- The substance(s) which we end with are called the product(s)

Evidence of a Chemical Reaction

- There are many visual and/or sensory clues which can be used to detect whether or not a chemical reaction may have occurred
- The most obvious indication of a reaction is the formation of a solid, called a <u>precipitate</u>, when two chemical solutions are combined
  - Such a reaction is called a <u>precipitation reaction</u>
- Other reactions produce a gas

Solid formation

Gas formation

Evidence of a Chemical Reaction

- Many chemical reactions give off energy as heat, while others absorb heat
- Some reactions give off energy by emitting light
  - We will not consider any of these reaction in this class
- Many chemical reactions are accompanied by a change in the color of the solution
- Unfortunately, virtually none of these pieces, taken alone, can confirm that a chemical reaction has taken place
  - Physical changes may also cause some of these effects to occur

Writing Chemical Reactions

- Chemical reactions are written with the reactants to the left, the products to the right, and an arrow between them to indicate the change.
  - Occasionally symbols or values may be written over or under the arrow to indicate the reaction conditions.
- An Example:

$$H_2 + O_2 \longrightarrow H_2O$$

Balancing Chemical Equations

- Consider the last reaction:

$$H_2 + O_2 \longrightarrow H_2O$$

- There is a problem with this equation!
- It indicates that we started with two oxygen atoms, but ended with one.
- What does this contradict?

Balancing Chemical Equations

- We must balance chemical equations, which is to say that there must be equal numbers of each type of atom on either side of a chemical reaction.
- To accomplish this, we put *coefficients* in front of the chemical formulas whose atom numbers we wish to increase.
- Note that you may never change the subscripts already in place in a chemical formula!
  - Why?

Balancing Chemical Equations

- To balance chemical equations first count the number of each type of atom you have on both sides of the reaction.
- If a polyatomic ion occurs on both sides of the chemical equation you should balance it as a unit.
- If an element appears in only one compound in each side of an equation, balance it first.
  - If this is true of more than one element, balance any metals before nonmetals.
- Identify any lone elements (as opposed to compounds) in the formulas; you will balance these last.
- From here, each equation requires its own logic; by trial and error, you should be able to balance the equation.
- The only other real "tip" I can give you on this subject is that **practice makes perfect!**

Examples

Balance each of the following chemical reactions:

$Mg + O_2 \longrightarrow MgO$

$Al + Br_2 \longrightarrow AlBr_3$

$Na + MgCl_2 \longrightarrow Mg + NaCl$

$Na_3PO_4 + BaCl_2 \longrightarrow Ba_3(PO_4)_2 + NaCl$

$C_6H_{12}O_6 + O_2 \longrightarrow CO_2 + H_2O$

Examples

- Write the balanced equation that corresponds to each statement
  - When methane reacts with oxygen gas, carbon dioxide and water vapor are produced.
  - Calcium metal reacts with ferric oxide, yielding calcium oxide and iron metal.
  - Treatment of carbon monoxide gas with oxygen gas produces carbon dioxide.

Symbols Used in Chemical Equations

- It is a common practice to include the states of substances in the chemical equation

State	Abbreviation
solid	(*s*)
liquid	(*l*) or (ℓ)
gas	(*g*)
aqueous	(*aq*)

Example: $H_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow 2H_2O(l) + BaSO_4(s)$

Symbols Used in Chemical Equations

- Other symbols or information may be written above or below the arrow

Symbol/Text	Meaning
Δ	reaction is heated
$h\nu$	reaction requires light
a time span (ex. 2 h, 3 d)	how long reaction is carried out
a temperature (ex. -25°C)	temperature reaction is carried out at
a chemical formula (ex. MnO_2)	a chemical which acts as a catalyst*

*A catalyst is a substance which speeds up the rate of a chemical reaction. It is not consumed by the reaction, and therefore is ignored when balancing the equation.

Double Displacement Reactions

- The general format for a double displacement reaction involves two ionic compounds trading their anions:

$$AX + BY \longrightarrow AY + BX$$

- AX and BY are both aqueous solutions, meaning both ionic compounds are dissolved in water.
- The charges of the ions do not change in this type of reaction
- Reactions do not always occur under these conditions. We must look for certain characteristics of the products:

Double Displacement Reactions

Does a Double Displacement Reaction Occur?

- Is a solid compound formed?
  - If yes, we have a precipitation reaction
- Is a gas formed?
  - If yes, we have a gas evolution reaction
- Is water or a weak acid formed?
  - If yes, we have an acid-base reaction, also called a neutralization reaction
- If none of these occur then there is no reaction

Examples of Double Displacement Reactions

$AgNO_{3(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

$3Ba(OH)_{2(aq)} + 2FeCl_{3(aq)} \longrightarrow 3BaCl_{2(aq)} + 2Fe(OH)_{3(s)}$

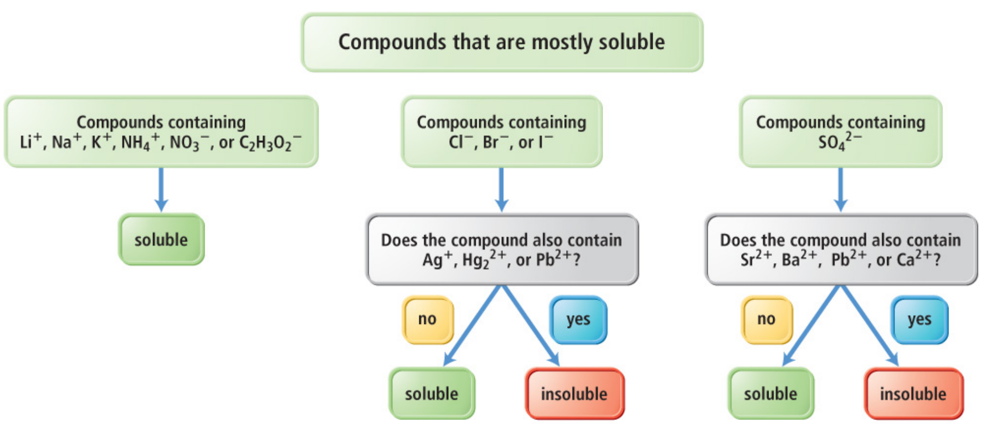
$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(l)} + NaCl_{(aq)}$

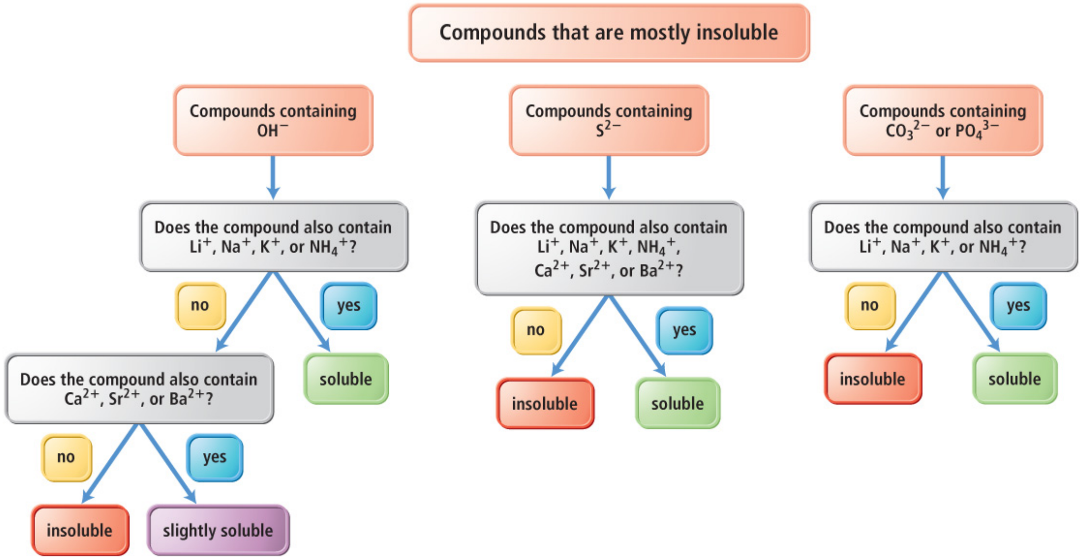
Precipitation Reactions and Solubility of Compounds

- <u>Solubility</u> is best described as the degree to which a compound will dissolve in a solvent (usually water).
  - A compound that is <u>soluble</u> will dissolve to a significant extent.
  - Compounds that are <u>insoluble</u> will not dissolve, remaining solid in solution.
- A reaction which produces an insoluble product can be described as a <u>precipitation</u> reaction; the product "falls out" of the solution like rain precipitates.

Solubility Rules

- It is important to know whether or not some common chemicals are soluble or not in order to predict how a reaction will proceed.
- You should know the following solubility rules (in water) *by memory*:
  - All nitrates and acetates are soluble.
  - All salts of Group I cations (Li^+, Na^+, etc.) and ammonium are soluble.
  - All chlorides, bromides, and iodides are soluble, except those of Ag^+, Pb^{2+}, and Hg_2^{2+}.
  - All hydroxides are insoluble except those of Group I, NH_4^+, Ba^{2+}, Sr^{2+}, and Ca^{2+}.
    - Calcium, barium, and strontium hydroxides are only *slightly* soluble, but we will not worry about this distinction for now





Examples

Indicate the products produced in each precipitation reaction.

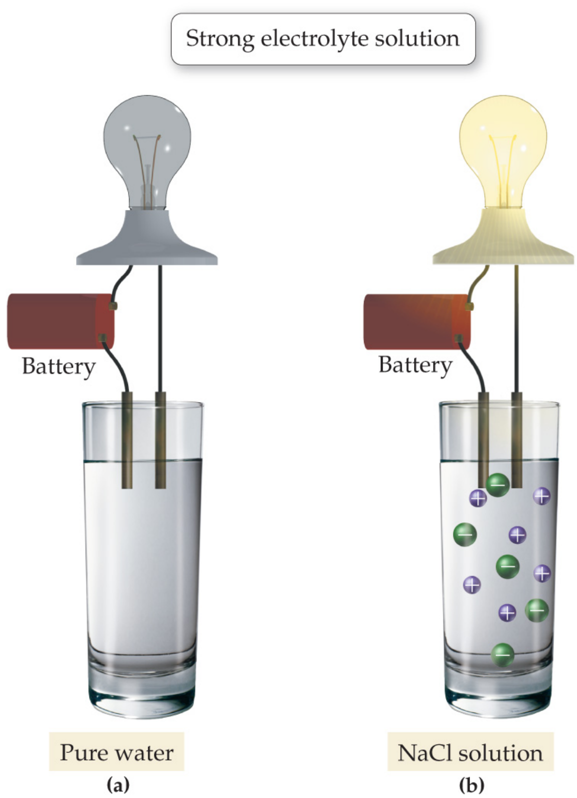
$NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow$

$Mg(NO_3)_{2(aq)} + KOH_{(aq)} \rightarrow$

$KNO_{3(aq)} + NaCl_{(aq)} \rightarrow$

Electrolytes

- When an ionic compound dissolves in water, it dissociates (separates) into ions.
- A solution containing ions is capable of conducting an electric current. We call these electrolytic solutions.
- A solution formed from a soluble ionic compound conducts a current very strongly; the compound is called a strong electrolyte.
- Some compounds stay mostly intact in solution; relatively little of it dissociates into ions. Such compounds are called weak electrolytes.
- Many compounds do not dissociate at all, or to such a small extent that they cannot conduct an electric current in solution; these compounds are called nonelectrolytes.



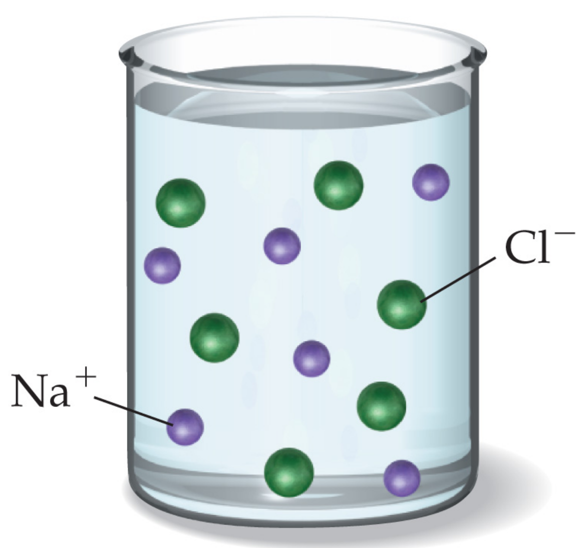
Electrolytes

Examples of strong electrolytes:

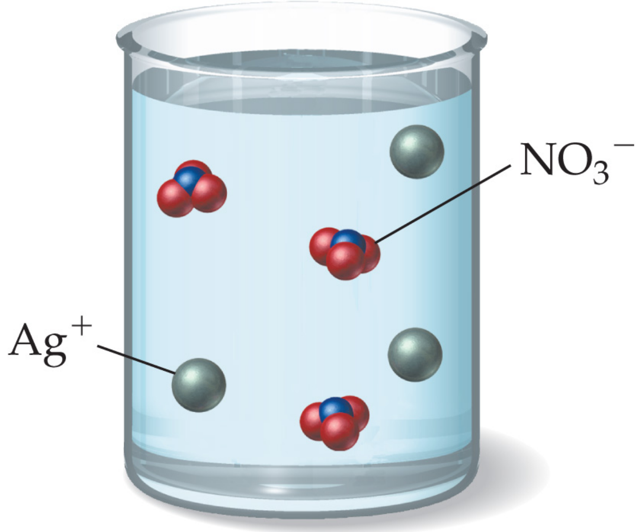
$NaCl$ KNO_3 NH_4Cl

Examples of non-electrolytes:

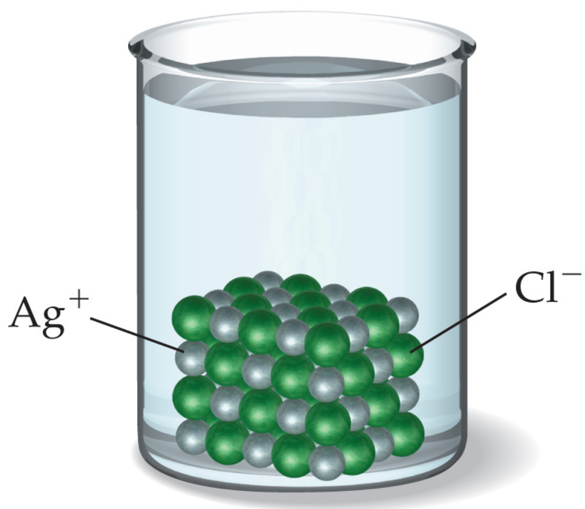
$CaCO_3$ $AgCl$ CH_3OH



A sodium chloride solution contains independent **Na^+** and **Cl^-** ions.

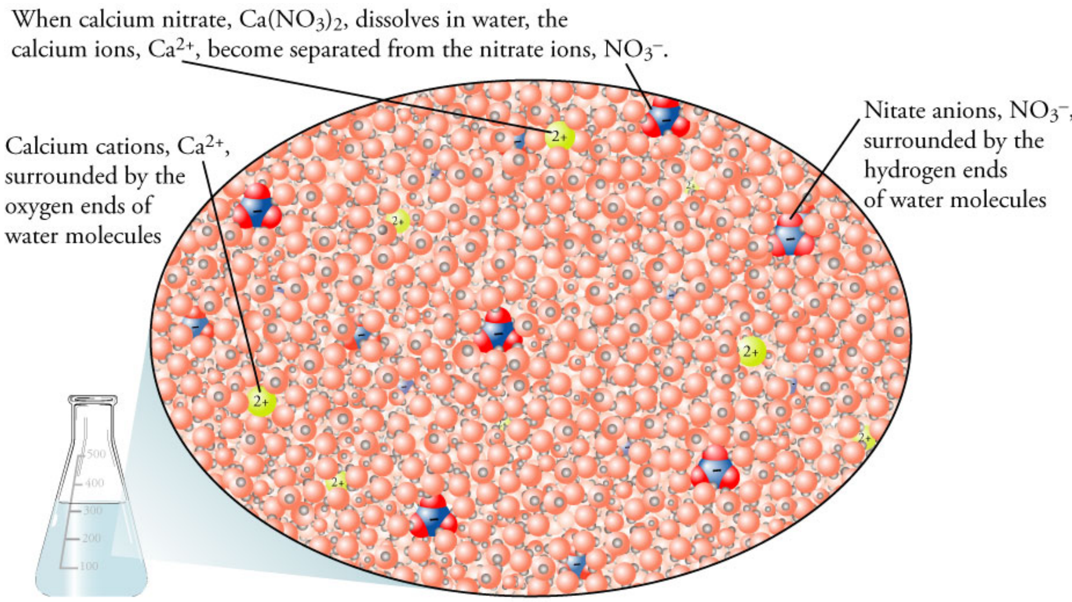


A silver nitrate solution contains independent $\mathbf{Ag^+}$ and $\mathbf{NO_3^-}$ ions.

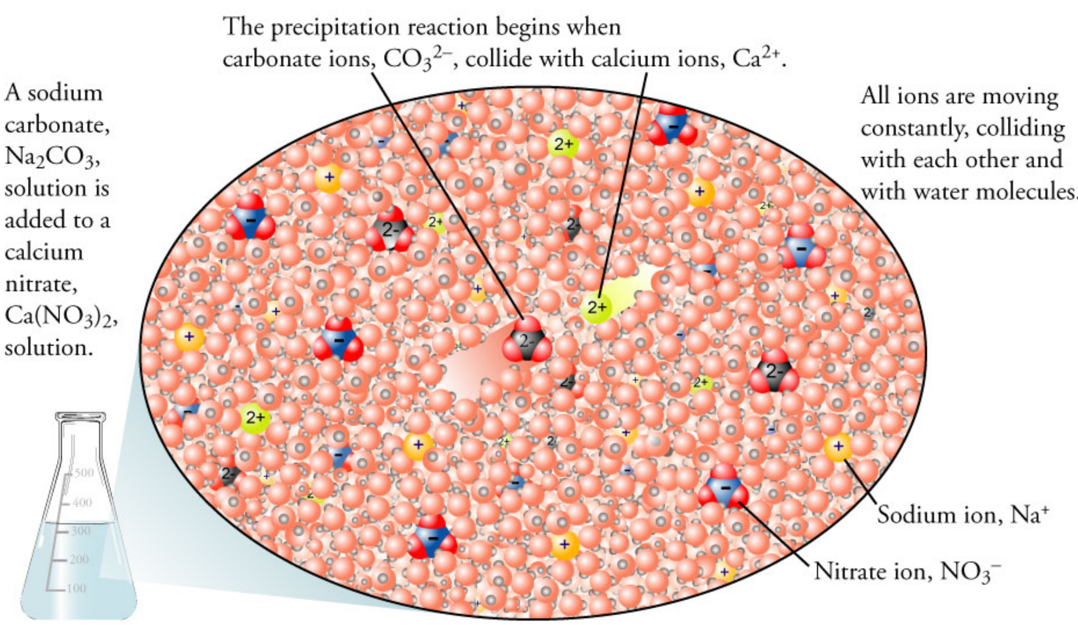


When silver chloride is added to water, it remains as solid AgCl—it does not dissolve into independent ions.

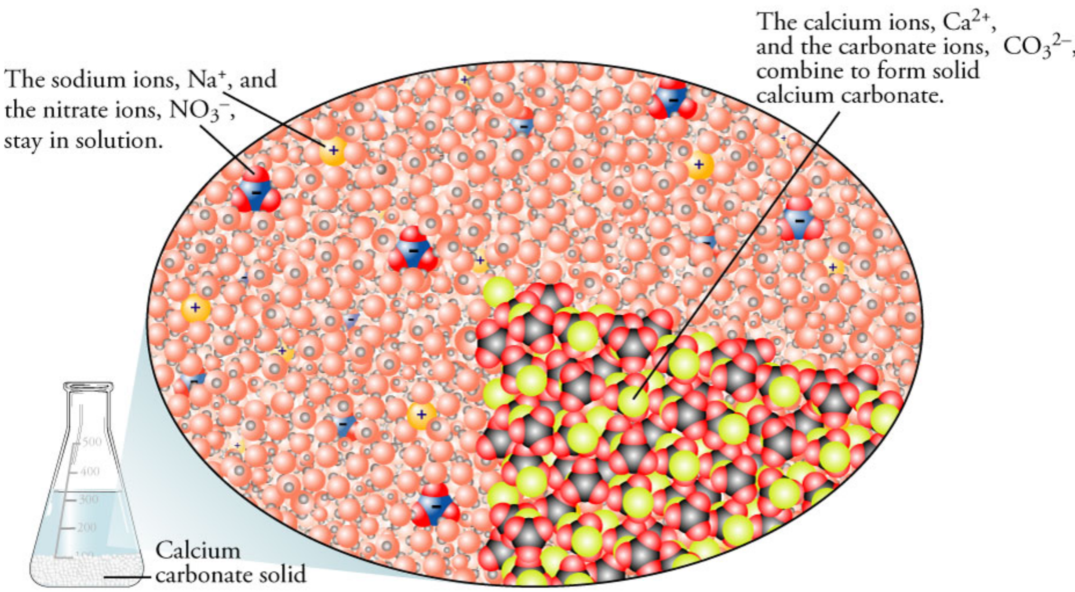
Solution of $Ca(NO_3)_2$



Solution of $Ca(NO_3)_2$ and Na_2CO_3 at the time of mixing, before the reaction



Product Mixture for the reaction of $Ca(NO_3)_2$ and Na_2CO_3



Complete Ionic Equations

- <u>Complete ionic equations</u> show all species in solution for a given reaction.
  - The term <u>species</u> refers to the specific ions and/or compounds in the solution.
- Examples of species:

  For $NaCl_{(aq)}$, $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$
  For $KNO_{3(aq)}$, $K^{+}_{(aq)}$ and $NO_3^{-}{}_{(aq)}$

Complete Ionic Equations

Example:

$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Complete Ionic Equation:

$Na^{+}_{(aq)} + Cl^{-}_{(aq)} + Ag^{+}_{(aq)} + NO_3{}^{-}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}_{(aq)} + NO_3{}^{-}_{(aq)}$

Spectator Ions

- A *spectator* of sports is someone who watches the game from the sidelines, but does not participate.
- Similarly, in chemical reactions, spectator ions "hang out" in a solution but do not actively participate in the reaction itself.
  - In other words, any ion which is both on the reactants and products side of a reaction is a spectator ion, for it has not undergone a chemical change.
- The ions' main purpose is to maintain constant charge in the solution.

Net Ionic Equations

- Net ionic equations only show those chemicals which participate in the reaction. Spectator ions are not included.
- To write a net ionic equation, first write down the total ionic equation.
- Then, cancel anything which appears *identically* on both sides of the reaction.

Net Ionic Equations

- Consider the complete ionic equation

$Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow AgCl_{(s)} + Na^+_{(aq)} + NO_3^-{}_{(aq)}$

- Now, factor out the spectator ions

$Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow AgCl_{(s)} + Na^+_{(aq)} + NO_3^-{}_{(aq)}$

- The net ionic equation is left over.

$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

Examples

Write chemical equations, complete ionic equations, and net ionic equations for the following:

a. a solution of barium chloride reacts with a sodium sulfate solution.

b. solutions of ferric chloride and lithium hydroxide are combined

Gas Evolution Reactions

- When produced in a reaction, some compounds will immediately decompose into other products.
  - Carbonic acid (H_2CO_3) will decompose into $CO_{2(g)}$ and $H_2O_{(l)}$.
  - Ammonium hydroxide (NH_4OH) will decompose into $NH_{3(aq)}$ and $H_2O_{(l)}$.
  - Sulfurous acid (H_2SO_3) will decompose into $SO_{2(g)}$ and $H_2O_{(l)}$.
- Notice that each produces water and a gaseous compound formed by the atoms left after water has been removed from the starting formula.
- If you produce any of these three compounds in a reaction, cancel it out and replace it with the decomposition products.
- Hydrogen sulfide, H_2S(*g*), produced by the reaction of a soluble sulfide salt (like Na_2S) with an acid, may also be a product of a gas evolution reaction

TABLE 7.4 Types of Compounds That Undergo Gas Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
sulfides	none	H_2S	$2\,HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\,KCl(aq)$
carbonates and bicarbonates	H_2CO_3	CO_2	$2\,HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,KCl(aq)$
sulfites and bisulfites	H_2SO_3	SO_2	$2\,HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,KCl(aq)$
ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

NH_3 is very soluble in water, even though enough gas escapes the solution that you can detect its smell. I prefer to indicate that it is aqueous, while your book prefers to call it a gas.

Acids and Bases

- You have heard the term *acid* applied to several compounds.
  - All of these compounds mentioned so far contain H^+, called a proton.
- There are at least three common methods for defining acids and bases.
- The method named for Swedish chemist Svante Arrhenius defines them as follows:
  - Acids are compounds which produce H^+ ions in solution
  - Bases are compounds which produce OH^- ions in solution

Neutralization Reactions

- Neutralization reactions are a subclass of double-displacement reactions.
- The general form of this reaction is

  acid + base → water + salt

  where the salt is any ionic compound.
- In most cases, this can be further simplified to the net-ionic equation

  $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Examples

$HBr_{(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KBr_{(aq)}$

$2HNO_{3(aq)} + Ba(OH)_{2(aq)} \rightarrow 2H_2O_{(l)} + Ba(NO_3)_{2(aq)}$

Write the balanced equations showing:

a) the neutralization of lithium hydroxide by perchloric acid.

b) the reaction of sulfuric acid with strontium hydroxide.

Oxidation & Reduction: An Introduction

- Many common reactions can be classified as <u>oxidation</u> or <u>reduction</u> reactions
  - We call this broad category of reactions "redox" for short
- In the most basic sense, oxidation can be described as a loss of electrons, while reduction is the gain of electrons
  - For example, the reaction between the elements sodium and chlorine is a redox reaction:
    - Sodium loses electrons, becoming Na^+
    - Chlorine gains electrons, becoming Cl^-
- An oxidation reaction is always accompanied by a reduction reaction.

Oxidation and Reduction in the formation of Binary Ionic Compounds

- When a metal reacts with a non-metal, the process can usually be simplified as follows:
  - The metal is oxidized (it loses electrons)
  - The non-metal is reduced (it gains electrons)
- Consider again the reaction between sodium and chlorine
- We can separate the main reaction into a pair of <u>half-reactions</u>, one describing the oxidation, the other the reduction

Oxidation and Reduction in the formation of Binary Ionic Compounds

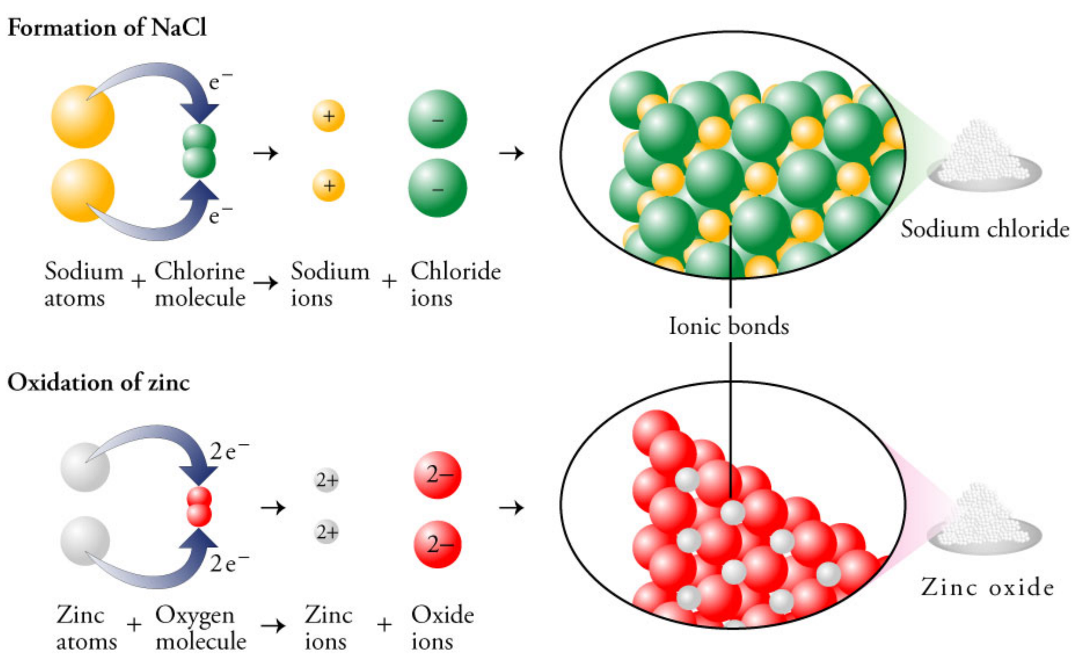
Chemical Equation: $2\ Na + Cl_2 \rightarrow 2\ NaCl$

Oxidation: $2\ Na \rightarrow 2\ Na^+ + 2\ e^-$

Reduction: $Cl_2 + 2e^- \rightarrow 2\ Cl^-$

(half-reactions)

Note that the half-reactions are balanced, and the total charge on the left side and right side of the reactions are balanced



What are the half-reactions corresponding to the reaction between zinc and oxygen?

Combustion Reactions

- In a combustion reaction, a chemical reacts with oxygen gas, forming various products.
- In this class, we will only consider the combustion of organic compounds containing C, H, and sometimes O.
- In these reactions, the compound reacts with oxygen gas, producing carbon dioxide and water vapor.

[organic compound] + $O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Examples of Combustion

- Combustion of benzene
  $2C_6H_{6(l)} + 15O_{2(g)} \rightarrow 12CO_{2(g)} + 6H_2O_{(g)}$

- Combustion of formaldehyde
  $CH_2O_{(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

- What is the balanced equation for the combustion of glycerol ($C_3H_8O_{3(l)}$)?

Single Displacement Reactions

- In a single displacementreaction, an element replaces another element which is present as an ion in a compound.
- There are two types of displacements we should be aware of:
  - displacement of one metal with another
  - displacement of one halogen with another

Single Displacement Reactions

- Format of the first type of single displacement reactions:

  $A + BX \rightarrow AX + B$

  Where A and B are metals/metal ions, and X is an anion.
- Here are some examples:

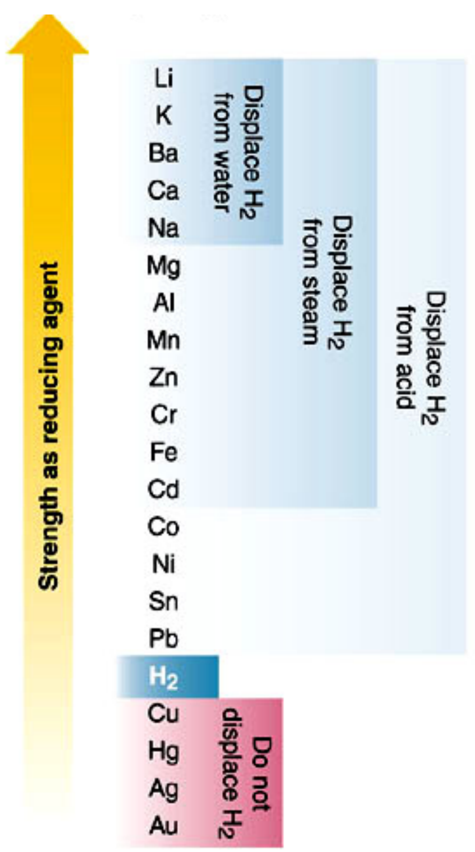
  $Mg + ZnCl_2 \rightarrow MgCl_2 + Zn$

  $Zn + 2AgNO_3 \rightarrow Zn(NO_3)_2 + 2Ag$

  $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Activities of Metals

- We can use the <u>activity</u> of a metal to describe how readily it loses electron(s) in a reaction.
  - The more active the metal, the more readily it loses its electrons.
- A more active metal will displace a less active metal in a single displacement reaction; the reverse will not occur.
  - In other words, an active metal will force the cation of a less active metal to take its electrons away from it.
- We look to the <u>activity series</u> to see the relative activities of the metals.
  - The activity series should not be memorized, but you should become familiar with trends within it.



The Activity Series of the Metals

Displacing Hydrogen

- Notice that many metals can displace the H^+ from acids, changing it into H_2.
  - Notice that the charged hydrogen ion is transformed into the neutral hydrogen molecule.
- Some very active metals can even displace H^+ from water, leaving OH^- behind.
  - In these cases, think of water as HOH.

When Does the Reaction "Go"?

- So let's consider one single displacement that works well…

  $$Zn + 2AgNO_3 \rightarrow Zn(NO_3)_2 + 2Ag$$
- Since zinc is more active than silver (higher on the activity series), zinc will displace silver.
- Consider the reverse reaction…

  $$Ag + Zn(NO_3)_2 \rightarrow \text{no reaction}$$
- Since silver is less active than zinc, it cannot displace it; therefore, no reaction can occur.

Examples

- Consider the following single displacement reaction:

$$3Mg + 2FeCl_3 \rightarrow 3MgCl_2 + 2Fe$$

  - What is being oxidized?
  - What is being reduced?
- Try this reaction:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

  - What is being oxidized?
  - What is being reduced?

Single displacement Reactions: displacement of Halogens

- Anions derived from halogens (Group VII) can be displaced by a more active halogen.
- Activity of the halogens decreases down the group:

$$F_2 > Cl_2 > Br_2 > I_2$$

most active — least active

- An example:

$$Cl_2 + 2NaI \rightarrow 2NaCl + I_2$$

- Will the reverse reaction proceed?

Examples

Predict the products of the following reactions. Write "no rxn." if none is expected to occur.

$Ba + CoBr_3 \rightarrow$

$Ni + NaCl \rightarrow$

$I_2 + KF \rightarrow$

$Li + H_2O \rightarrow$

$Ni + HNO_3 \rightarrow$

Combination Reactions

- In a combination reaction, two chemicals combine into one new chemical.
- These reactions have the general form

  $$A + B \rightarrow C$$

- It will not always be possible to predict the products of combination reactions at this level of preparation, so we will study a few general cases.

Oxide Formation

- Metals often react with oxygen to form a metal oxide.

  Ex. $4\ Na + O_2 \rightarrow 2\ Na_2O$
- Nonmetals often react with oxygen to form a nonmetal oxide.

  Ex. $C + O_2 \rightarrow CO_2$
  - It is often difficult to predict the products in this case, as CO was another possible oxide you might have considered.

Reactions of Oxides

- Metal oxides often react with water to form metal hydroxides.

  Ex. $Na_2O + H_2O \rightarrow 2NaOH$

- Nonmetal oxides often react with water to form oxyacids.

  Ex. $CO_2 + H_2O \rightarrow H_2CO_3$

  $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

  $SO_3 + H_2O \rightarrow H_2SO_4$

  $N_2O_5 + H_2O \rightarrow 2HNO_3$

Reactions of Oxides

- Metal oxides and nonmetal oxides often combine to form a salt.

$$Na_2O + CO_2 \rightarrow Na_2CO_3$$

$$CaO + SO_3 \rightarrow CaSO_4$$

Decomposition Reactions

- Decomposition reactions are simply the reverse of combination reactions.
- Their general form is

$$Z \rightarrow X + Y$$

- Like combination reactions, predicting the products of these reactions is often difficult.
- For now, simply decompose a given compound into two products which could have produced it from a method we discussed earlier.
  - This is *very* oversimplified, and not always correct, but it is the best we can do at this level.

Examples

$2HgO \rightarrow 2Hg + O_2$

$Na_2CO_3 \rightarrow Na_2O + CO_2$

$Li_2SO_4 \rightarrow Li_2O + SO_3$