A chemical equation indicates the reactants and products of a reaction. By convention, the reactants are to the left of the arrow and the products to the right of the arrow.

Reactants → Products

The energy term may be included in the equation. The type of energy associated with physical/chemical changes is heat energy. Heat energy has the unit of joule (J). If heat is absorbed by the reaction, the heat term is on the reactant side of the equation. Such a reaction is termed endothermic. If heat is released by the reaction, the heat term is placed on the product side. Such reactions are termed exothermic.

Chemists typically include the phases of the reactants and products designated by a symbol in parenthesis after each species. The symbol designations are:

(s) = solid
(l) = liquid
(g) = gas
(aq) = aqueous (which means dissolved in water)

A key idea in writing any chemical equation is that it must obey the **law of mass conservation**; the sum total of masses of reactants must equal the sum total of masses of products:

$\sum$ MASS REACTANTS = $\sum$ MASS PRODUCTS

In writing a chemical equation we must first write the correct formulas of all reactants and products. Recall that we do this by adjusting subscripts. Once we have written all the correct formulas **we balance the equation by adjusting coefficients**. Coefficients are the numbers in front of the formulas of substances in chemical equations. They tell us the relative number of molecules or formula units taking part in a chemical reaction.

Let's consider an example. Write a balanced equation for the synthesis of ammonia gas from its elements at room temperature and pressure. In the process 91.8 kJ of heat energy is released:

skeleton equation: $N_{2(g)} + H_{2(g)} \rightarrow NH_{3(g)} + 91.8 kJ$

now balance by adjusting coefficients:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} + 91.8 kJ$

How does our balanced equation verify the law of mass conservation? Coefficients not only tell us the relative number of molecules or formula units but **coefficients also represent moles**.

Balancing Chemical Equations

Write balanced chemical equations for the following reactions and include the phases.

1. Magnesium metal reacts with tetraatomic phosphorous to produce magnesium phosphide.

$$\underline{6} Mg_{(s)} + \underline{\;} P_{4(s)} \rightarrow \underline{2} Mg_3P_{2(s)}$$

2. Silver nitrate and aluminum sulfate are soluble in water. When the solutions of these two compounds are mixed, a white precipitate (solid) of silver sulfate forms and aluminum nitrate is found in solution.

$$6AgNO_{3(AQ)} + Al_2(SO_4)_{3(AQ)} \rightarrow 3Ag_2SO_{4(s)} + 2Al(NO_3)_{3(AQ)}$$

3. Aluminum products that are used in the home, such as window frames and door frames, soon become coated with a film of aluminum oxide. This film can be removed by cleaners that contain a dilute solution of hydrochloric acid. The products are aluminum chloride and water.

$$Al_2O_{3(s)} + 6HCl_{(AQ)} \rightarrow 2AlCl_{3(AQ)} + 3H_2O$$

4. Solid calcium phosphate may be obtained by reacting aqueous solutions of phosphoric acid and calcium hydroxide. The other product is liquid water.

$$2H_3PO_{4(AQ)} + 3Ca(OH)_{2(AQ)} \rightarrow Ca_3(PO_4)_{2(s)} + 3H_2O_{(l)}$$

5. If iron pyrite, FeS_2, is not removed from coal, oxygen from the air will combine with both the iron and sulfur as the coal burns. Write a balanced chemical equation illustrating this combustion. The products are ferric oxide and sulfur dioxide.

$$2FeS_{2(s)} + 5.5O_2 \rightarrow Fe_2O_{3(s)} + 4SO_{2(g)}$$

$$4 \qquad 11 \qquad 2 \qquad 8$$

6. Ethanol liquid, having the chemical formula CH_3CH_2OH, combines with oxygen gas in a combustion reaction to yield carbon dioxide gas and water vapor.

$$CH_3CH_2OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$$

7. Aqueous solutions of sodium hydroxide and sodium hydrogen sulfate react to yield aqueous sodium sulfate and water.

$$NaOH_{(Aq)} + NaHSO_{4(Aq)} \rightarrow Na_2SO_{4(Aq)} + H_2O_{(l)}$$

8. Solid ammonium carbonate is decomposed by heating to produce ammonia gas, carbon dioxide gas and water vapor.

$$(NH_4)_2CO_{3(s)} \rightarrow 2NH_{3(g)} + CO_{2(g)} + H_2O_{(g)}$$

Reactions in Aqueous Solutions

Observing the many different kinds of chemical reactions that can occur have led chemists to recognize that there are several types of reactions that tend to occur when reactants are added. That is, these reactants tend to form a product (or are likely to take place) Chemists view such reactions as ones which involve

- formation of a solid called precipitate reactions
- formation of water these involve acid-base reactions
- transfer of electrons called oxidation-reduction reactions (redox)

Precipitate Reactions

Let us first consider precipitate reactions. These reactions generally involve aqueous solutions of two different salts that are mixed together. Salts are ionic compounds consisting of a metal-nonmetal combination in which the metal occurs as a positively charged species (called a cation) and the nonmetal occurs as a negatively charged species (called an anion). Remember these ions when we learned about writing the chemical formulas for ionic compounds? In order to understand what occurs, we need to first understand what happens when an ionic compound dissolves in water. Consider the common salt, sodium chloride. This salt occurs as a crystalline white solid. This solid appears white because it reflects all colors of visible light. When placed in water, the salt disappears! Chemists have a model that explains this disappearance. This model pictures the ions making up the solid break apart in water (or dissociate in water) into separate ions that independently "swim" around in the water. A water molecule is said to be polar – having a partial positive (hydrogen) end and a partial negative (oxygen) end. The cation is surrounded by the negative oxygen end of water molecules (making a plus-minus attraction) and the anions are surrounded by the positive hydrogen ends of water making plus-minus attractions. We say the ions become "***hydrated***" by water. In this aqueous phase, visible light passes through the ions and thus they appear colorless. The reaction involving the dissociation of sodium chloride in water can be written as:

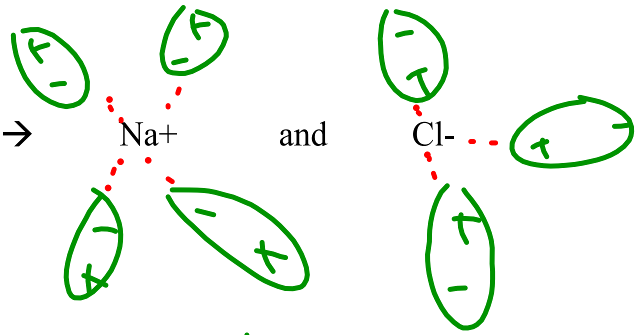
$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

white clear

- + = H_2O

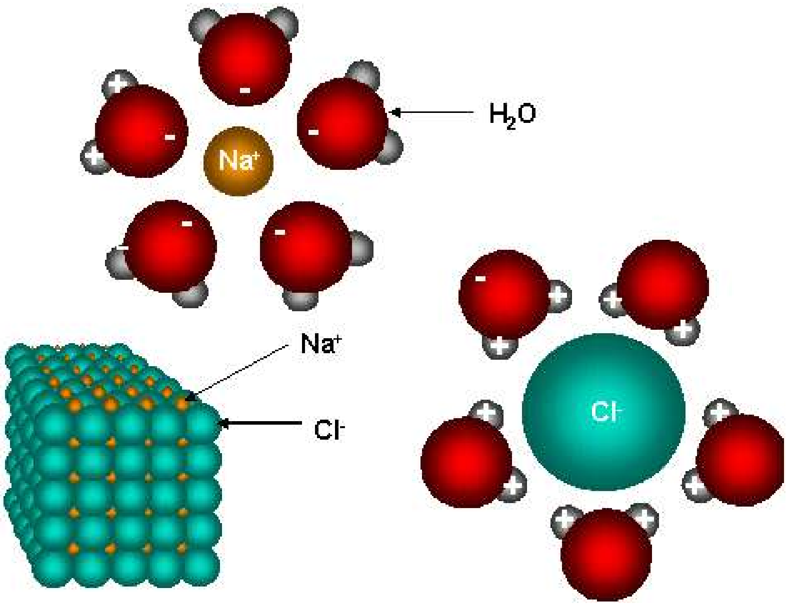
particle diagram:

Na+Cl-Na+Cl-Na+Cl-
Cl-Na+Cl-Na+Cl-Na+ →
Na+Cl-Na+Cl-Na+Cl-
Cl-Na+Cl-Na+Cl-Na+



$Na^+_{(aq)}$ $Cl^-_{(aq)}$

hydrated ions


H2O
Na+
Na+
Cl-
Cl-

Let's practice writing some equations showing the **dissociation of salts into positive and negative aqueous ions.** ~~Cl_2^-~~

Calcium chloride: $CaCl_{2(s)} \xrightarrow{H_2O} Ca^{+2}_{(AQ)} + 2Cl^-_{(AQ)}$

Sodium acetate: $NaC_2H_3O_{2(s)} \xrightarrow{H_2O} Na^+_{(AQ)} + C_2H_3O_2^-{}_{(AQ)}$

Aluminum nitrate: $Al(NO_3)_{3(AQ)} \xrightarrow{H_2O} Al^{3+}_{(AQ)} + 3NO_3^-{}_{(AQ)}$

Potassium phosphate: $\overset{+1}{K_3}\overset{-3}{PO_4}{}_{(s)} \xrightarrow{H_2O} 3K^+_{(AQ)} + PO_4^{3-}{}_{(AQ)}$

Sodium hydroxide: $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(AQ)} + OH^-_{(AQ)}$

How is our model of the existence of ions in solutions supported by experimental evidence? Chemists know that a solution will conduct electricity if the solution contains freely moving ions. Pure water does not contain any dissolved ions and does not conduct an electric current. When a salt is added to the water, the solution will now conduct an electric current.

Substances that, when dissolved in water, break apart to yield positive and negative ions are termed **electrolytes**. What is a famous example of an electrolyte developed in Florida?

Salts, acids and bases are all examples of electrolytes.

Substances that dissolve in water and break apart completely (that is, 100% of the units break apart to form ions) are called **strong electrolytes**. Sodium chloride is a strong electrolyte:

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

	$NaCl(s)$	$\rightarrow$	$Na^{+}(aq)$	$Cl^{-}(aq)$
start	100			
		100%		
after dissolving	0		100	100

Know let's look at reactions that produce a solid (called a precipitate). These reactions are also known as **double replacement** reactions. In these reactions we mix two aqueous solutions of two different salts and exchange or switch the positive and negative ions. This is like switching dance partners. This produces two new combinations of ions. The general equation is:

$$\overset{+1\ -1}{XY}(aq) + \overset{+1\ -1}{ZW}(aq) \rightarrow \overset{+1\ -1}{XW}(s) + \overset{+1\ -1}{YZ}(aq)$$

Notice that one of the products formed is a solid. This will occur if the new combination of ions is more stable with the ions combined in the solid state rather than as separated ions in the water. The solid formed is called the precipitate.

How do we know which salt will be more stable as a solid in water or more stable as separated ions in water? Chemists have developed a set of solubility rules or guidelines that will allow us to make such a prediction. If the combination of ions is such that they will be more stable as separated ions in water, the salt is termed **soluble**. If, on the other hand, it is more stable with the ions together in the solid form, the salt is termed insoluble. For the purposes of this course you will not need to memorize any solubility rules. You will be required to understand how to use the solubility tables (given on the next page) to predict the formation of a precipitate.

Table F
Solubility Guidelines

Ions That Form *Soluble* Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
perchlorate (ClO_4^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}

Ions That Form *Insoluble* Compounds	Exceptions
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, or Sr^{2+}

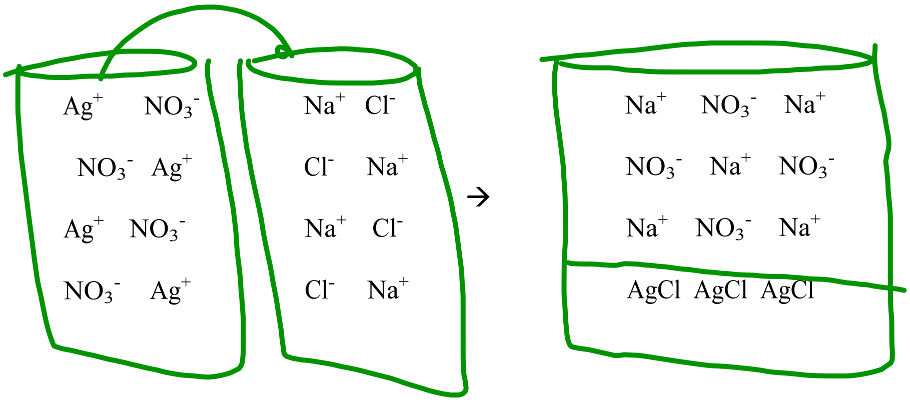
Let's try some precipitate reactions!

Aqueous solutions of silver nitrate and sodium chloride are mixed:

$$\overset{+1}{Ag}\overset{-1}{NO_3}{}_{(aq)} + \overset{+}{Na}\overset{-}{Cl}{}_{(aq)} \rightarrow \overset{+}{Ag}\overset{-}{Cl}{}_{(s)} + \overset{+}{Na}\overset{-}{NO_3}{}_{(aq)}$$

Diagram:

Key: Ag^{+}(aq) NO_3^-(aq) Na^+(aq) Cl^-(aq) AgCl(s)



Writing Net Ionic Equations

- The first step is to write a balanced molecular equation (BME). This is what we have been doing so far. For example

  Write a balanced molecular equation for the reaction of aqueous solutions of calcium chloride and sodium carbonate

$$\overset{+2}{Ca}\overset{-1}{Cl}_{2(aq)} + \overset{+}{Na}_2\overset{-2}{CO_3}_{(aq)} \rightarrow \overset{+2}{Ca}\overset{-2}{CO_3}_{(s)} + 2\overset{+}{Na}\overset{-}{Cl}_{(aq)}$$

- The next step involves writing a complete ionic equation (CIE). To do this, we must show how the ions actually occur in solution. That is, wherever we have an (aq) phase, we must separate the ions and show **how they occur as *separate* aqueous ions**. For example, $CaCl_2$(aq) means one separate calcium +2 ion and two separate -1 chloride ions. For those in the aqueous phase ask "what is the positive ion?" and then "what is the negative ion?" The formula of chloride ion is Cl^-, not Cl_2^-!

  $\cancel{Cl_2}$

$$Cl^- \text{ --- } Ca^{2+} \text{ --- } Cl^- \rightarrow Ca^{2+} + 2\,Cl^-$$

  Substances in the solid, liquid or gas phase occur as such and are "carried down" in the CIE.

$$Ca^{+2}_{(aq)} + \cancel{2Cl^-_{(aq)}} + \cancel{2Na^+_{(aq)}} + CO_{3\,(aq)}^{2-} \rightarrow \overset{+2}{Ca}\overset{-2}{CO}_{3(s)} + \cancel{2Na^+_{(aq)}} + \cancel{2Cl^-_{(aq)}}$$

- To write the net ionic equation (NIE) we must **cancel spectator ions** – those ions that do not change in either oxidation number or phase (i.e. they stay the same). What are the spectator ions in the above CIE‘

$$Ca^{+2}_{(aq)} + CO_{3\,(aq)}^{2-} \rightarrow CaCO_{3(s)}$$

Practice Problems- write BME, CIE and NIE for the following reactions:

1. Aqueous solutions of sodium hydroxide and copper(II)chloride are mixed.

$$NaOH_{(AQ)} + CuCl_{2(AQ)} \rightarrow 2NaCl_{(AQ)} + Cu(OH)_{2(s)}$$

$$Na^{+}_{(AQ)} + OH^{-}_{(AQ)} + Cu^{2+}_{(AQ)} + 2Cl^{-}_{(AQ)} \rightarrow 2Na^{+}_{(AQ)} + 2Cl^{-}_{(AQ)} + Cu(OH)_{2(s)}$$

$$Cu^{2+}_{(AQ)} + 2OH^{-}_{(AQ)} \rightarrow Cu(OH)_{2(s)}$$

2. The reaction between aqueous lead(II)nitrate and aqueous potassium sulfate.

$$Pb(NO_3)_{2(AQ)} + K_2SO_{4(AQ)} \rightarrow PbSO_{4(s)} + 2KNO_{3(AQ)}$$

$$Pb^{2+}_{(AQ)} + 2KNO^{-}_{3(AQ)} + 2K^{+}_{(AQ)} + SO^{2-}_{4(AQ)} \rightarrow PbSO_{4(s)} + 2K^{+}_{(AQ)} + 2NO^{-}_{3(AQ)}$$

$$Pb^{2+}_{(AQ)} + SO^{2-}_{4(AQ)} \rightarrow PbSO_{4(s)}$$

3. Aqueous solutions of aluminum sulfate and calcium iodide are mixed.

$$Al_2(SO_4)_{3(AQ)} + 3CaI_{2(AQ)} \rightarrow 3CaSO_{4(s)} + 2AlI_{3(AQ)}$$

$$2Al^{3+}_{(AQ)} + 3SO_4^{2-}{}_{(AQ)} + 3Ca^{2+}_{(AQ)} + 6I^-_{(AQ)} \rightarrow 3CaSO_{4(s)} + 2Al^{3+}_{(AQ)} + 6I^-_{(AQ)}$$

$$3Ca^{2+}_{(AQ)} + 3SO_4^{2-}{}_{(AQ)} \rightarrow CaSO_{4(s)}$$

4. The reaction between aqueous sodium phosphate and aqueous magnesium nitrate.

$$\overset{+1}{2Na_3}\overset{-3}{PO_{4(AQ)}} + \overset{+2}{3Mg}\overset{-1}{(NO_3)_{2(AQ)}} \rightarrow \overset{+2}{Mg_3}\overset{-3}{(PO_4)_{2(s)}} + \overset{+}{6Na}\overset{-}{NO_{3(AQ)}}$$

$$\cancel{6Na^+_{(AQ)}} + 2PO_4^{3-}{}_{(AQ)} + 3Mg^{+2}_{(AQ)} + \cancel{6NO_3^-{}_{(AQ)}} \rightarrow Mg_3(PO_4)_{2(s)} + \cancel{6Na^+_{(AQ)}} + \cancel{6NO_3^-{}_{(AQ)}}$$

$$2PO_4^{-3}{}_{(AQ)} + 3Mg^{+2}_{(AQ)} \rightarrow Mg_3(PO_4)_{2(s)}$$

5. Aqueous solutions of potassium sulfide and aluminum acetate are mixed.

$$3K_2S_{(AQ)} + 2Al(C_2H_3O_2)_{3(AQ)} \rightarrow Al_2S_{3(s)} + 6KC_2H_3O_{2(AQ)}$$

$$6K^+_{(AQ)} + 3S^{2-}_{(AQ)} + 2Al^{3+}_{(AQ)} + 6C_2H_3O^-_{2(AQ)} \rightarrow Al_2S_{3(s)} + 6K^+_{(AQ)} + 6C_2H_3O^-_{2(AQ)}$$

$$2Al^{3+}_{(AQ)} + 3S^{2-}_{(AQ)} \rightarrow Al_2S_{3(s)}$$

6. The reaction between aqueous sodium phosphate and aqueous magnesium nitrate.

Reactions That Form Water

Here we will limit our discussion to acid-base reactions. Recall we studied the nomenclature of acids previously. But just what is an acid?

Svante Arrhenius described an acid as a substance that when dissolved in water breaks apart to produce hydrogen ion (H^+) as the positive ion and a negative ion. Acids are molecular substances (as opposed to ionic). Molecular substances like acids are not made up of ions but form ions when dissolved in water. Technically we say that acids ***ionize*** in water. Salts are already made up of ions and simply ***dissociate*** in water.

Arrhenius described bases as substances that when dissolved in water dissociate to form a positive ion and a hydroxide (OH^-) ion.

The acid-base reactions we will deal with involve the reactions of strong acids and strong bases. What does it mean to be a strong acid or base? If you thought strength reflects how concentrated the acid or base is you would be incorrect. Acid strength refers to how easily an acid ionizes in water; base strength refers to how a base dissociates in water. The definition of a strong acid or base is one that ionizes/dissociates virtually 100% when placed in water. This means that if you placed 100 molecules of in acid in water, all 100 will ionize to give 100 H^+ and 100 negative ions. Likewise, if you placed 100 units of a strong base in water, all 100 will dissociate to give 100 OH^- ions and 100 positive ions:

	$HCl(g)$	→	$H^+(aq)$	+	$Cl^-(aq)$ (strong acid)
start	100				
after dissolving	0	→	100		100

	$NaOH(g)$	→	$Na^+(aq)$	+	$OH^-(aq)$
start	100				
after dissolving	0	→	100		100

Strong acids and strong bases are thus strong electrolytes.

The common strong acids you need to memorize are:

- hydrochloric
- nitric
- sulfuric
- perchloric
- hydrobromic
- hydroiodic

The common strong bases you need to memorize are:

- sodium hydroxide
- potassium hydroxide
- calcium hydroxide

Acid base reactions are also double replacement type reactions that are also known as "neutralization" reactions and follow the general equation:

acid + base → HOH + salt

Write the BME, CIE and NIE for the following acid-base reactions:

1. nitric acid and aqueous sodium hydroxide:

2. hydrochloric acid and aqueous potassium hydroxide.

3. sulfuric acid and aqueous sodium hydroxide.

4. perchloric acid and aqueous calcium hydroxide.

Oxidation-Reduction Reactions (Electron Transfer)

These reactions involve electron(s) transferred from metals to nonmetals resulting in positive metal ions and negative nonmetal ions. The bond formed by this transfer of electron(s) is called an ionic bond. The transfer of electrons is not limited to metal-nonmetal reactions. To recognize a redox reaction we need to assign charges to the atoms of the elements in a compound. We did this back when we learned how to write the formulas of ionic and molecular compounds. An element in its free state (uncombined) is assigned a charge of zero. If the element occurs as an aqueous ion, the charge assigned is the charge of the ion. We can recognize a redox reaction has occurred if the elements undergo a change in charge (oxidation number).

Some redox reactions can also be classified as combustion, synthesis, decomposition or single replacement reactions. Let's take a look at some examples.

Synthesis (Combination) Reactions

In a synthesis reaction two or more simple substances are combined to form one new or more complex substance. Here the general form is

element or compound + element or compound $\rightarrow$ compound

general equation: $X + Y \rightarrow XY$

The following are some general types of synthesis reactions.

1) A binary ionic compound is formed when a metal reacts with a nonmetal.

e.g. $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

what substance was oxidized?

What substance was reduced?

Write balanced equations for the following reactions and identify the substances oxidized and reduced:

a) sodium reacts with chlorine gas:

b) magnesium metal undergoes a reaction with nitrogen gas:

Decomposition Reactions

When energy in the form of heat, electricity, light, or mechanical shock is supplied, a compound may decompose to form simpler substances. The general form for this type of reaction is

Compound → two or more substances

general equation: XY → X + Y

The following are some general types of decomposition reactions.

1. Binary covalent compounds, and some binary ionic compounds, decompose under unique conditions into the separate elements of which they are made.

   e.g. Water is decomposed by electrolysis:

   $$2H_20(l) \rightarrow 2H_2(g) + O_2(g)$$

   The group 1 and group 2 metals are obtained in their elemental form by the electrolysis of their molten (melted) salts:

   $$2NaCl(l) \rightarrow 2Na(s) + Cl_2(g)$$

2. Some compounds may be decomposed by simple heating.

   e.g. a sample of mercuric oxide is heated and decomposes to its elements:

   indicate the substances oxidized and reduced:

3. Metal chlorates decompose to metal chlorides and oxygen gas when heated.

   e.g. a sample of potassium chlorate is heated:

4. Metal carbonates decompose to metal oxides and carbon dioxide gas when heated.

   e.g. a sample of calcium carbonate is heated:

Write balanced equations for the following reactions:

a) a sample of aluminum chloride undergoes electrolysis:

b) a sample of magnesium chlorate is heated:

c) a sample of lithium carbonate is heated:

Combustion Reactions

In a combustion reaction a substances reacts with oxygen gas to produce heat energy so rapidly that a flame results. Hence these reactions are highly exothermic.

The combustion of hydrocarbons (compounds containing only carbon and hydrogen) and alcohols (hydrocarbons containing an –OH group attached to a carbon atom) results in the formation of carbon dioxide gas and water vapor:

e.g. $C_3H_8(l) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
(propane)

Write balanced equations for the following reactions:

a) the combustion of octane gas, C_8H_{18}:

b) the combustion of liquid ethanol, C_2H_5OH:

Single Replacement reactions

One element replaces or displaces another element in a compound. A single replacement has the general form

element + compound → element + compound

general equation: $\overset{+1}{X}\overset{-1}{Y} + \overset{0}{Z} \rightarrow \overset{+1}{Z}\overset{-1}{Y} + \overset{0}{X}$

A more active metal will replace a less active metal (or plus replaces plus). A more active nonmetal will replace a less active nonmetal (minus replaces minus). To determine whether or not single replacement reactions will occur, consult Table J, which lists metals and nonmetals in order of decreasing activity.

Table J
Activity Series**

Most ↓	Metals	Nonmetals	Most ↓
	Li	F_2	
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	**H_2		
	Cu		
	Ag		
	Au		
Least			Least

**Activity Series based on hydrogen standard

1) An active metal will replace the metallic ion in a compound of a less active metal.

e.g. $Fe(s) + CuCl_2(aq) \rightarrow FeCl_2(aq) + Cu(s)$

2) The group 1 metals (alkali metals) and the group 2 metals (alkaline earth metals) react with water to produce the associated metallic hydroxide and hydrogen.

e.g. $Mg(s) + 2HOH(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

3) An active nonmetal will replace a less active nonmetal.

e.g. $Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$

Write BME, CIE, and NIE for the following reactions:

a) zinc metal is added to a solution of tin(II)chloride:

b) iron fillings are sprinkled into a solution of nitric acid:

c) potassium metal is added to water:

d) fluorine gas is bubbled through a solution of lithium chloride:

Identifying phases of substances

- metallic elements are all solids except mercury (Hg)
- the nonmetal elements that occur as diatomic molecules and their phases are: H_2(g), N_2(g), O_2(g), F_2(g), Cl_2(g), Br_2(l) I_2(s)
- most of the compounds we will encounter will be metal-nonmetal combinations, or salts. Salts are all solids. If a salt i formed in a reaction involving water or some aqueous phase, we need to check the solubility table to determine if it is soluble (indicate aqueous phase) or insoluble (indicate solid phase).