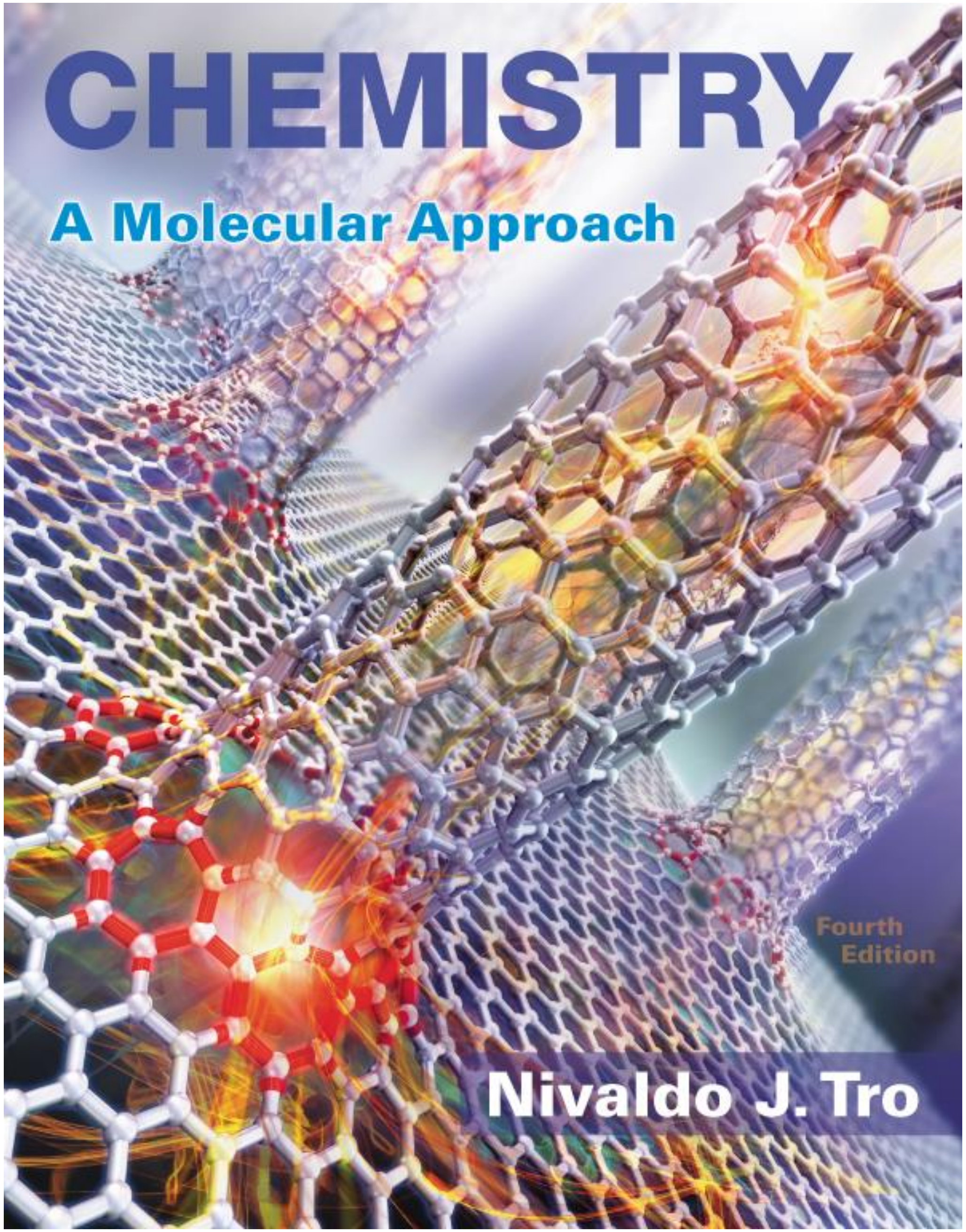


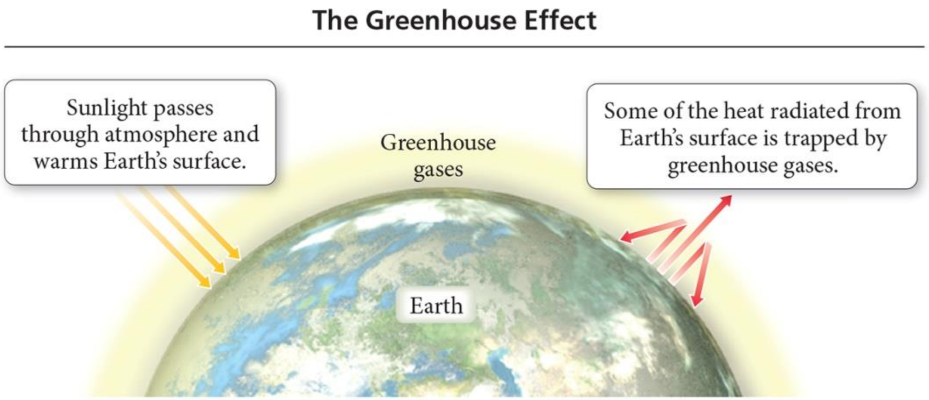
Lecture Presentation

Chapter 4

Chemical Quantities and Aqueous Reactions

The Greenhouse Effect

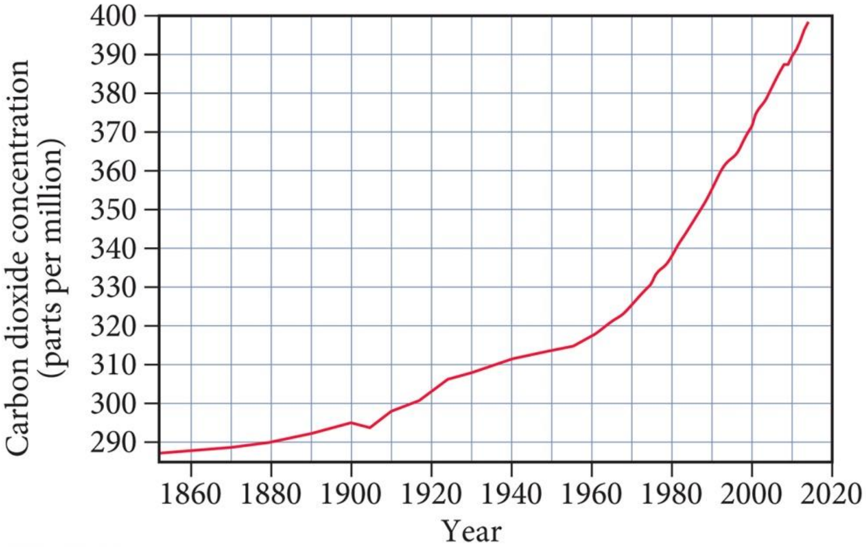
- The **greenhouse gases** in the atmosphere
  - allow sunlight to enter the atmosphere.
  - warm Earth's surface.
  - prevent some of the heat generated by the sunlight from escaping.



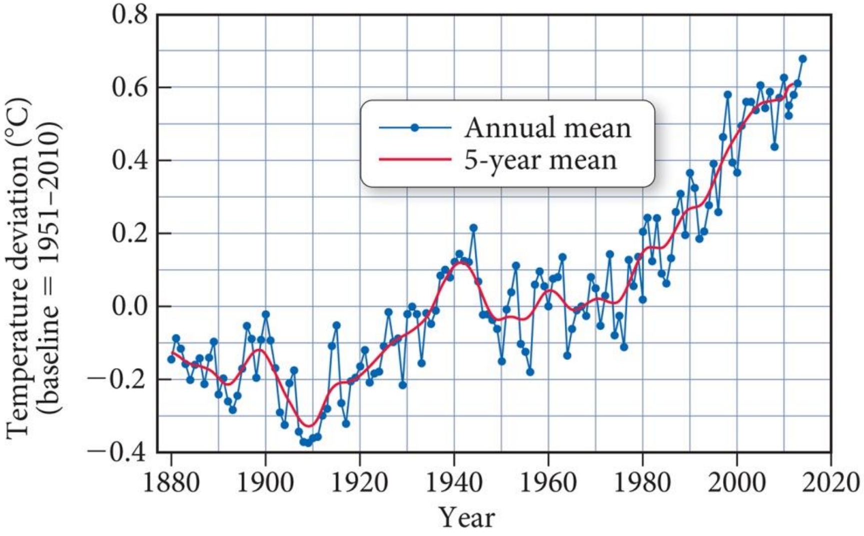
- The balance between incoming and outgoing energy from the sun determines Earth's average temperature.

Global Warming

- Scientists have measured an average 0.7 °C rise in atmospheric temperature since 1860.
- During the same period, atmospheric CO_2 levels have risen 38%.
- Are the two trends causal?



Atmospheric Carbon Dioxide



Global Temperature

Global Warming

- One source of CO_2 is the combustion reactions of fossil fuels we use to get energy.
- Another source of CO_2 is volcanic action.
  - How can we judge whether global warming is natural or due to our use of fossil fuels?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

Reaction Stoichiometry: How Much Carbon Dioxide?

- The balanced chemical equations for fossil-fuel combustion reactions provide the exact relationships between the amount of fossil fuel burned and the amount of carbon dioxide emitted.

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

  - 16 CO_2 molecules are produced for every 2 molecules of octane burned.

Quantities in Chemical Reactions

- The amount of every substance used and made in a chemical reaction is related to the amounts of all the other substances in the reaction.
  - Law of conservation of mass
  - Balancing equations by balancing atoms
- The study of the numerical relationship between chemical quantities in a chemical reaction is called **stoichiometry**.

Reaction Stoichiometry

- **The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.**

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

  - 2 molecules of C_8H_{18} react with 25 molecules of O_2 to form 16 molecules of CO_2 and 18 molecules of H_2O.
  - 2 moles of C_8H_{18} react with 25 moles of O_2 to form 16 moles of CO_2 and 18 moles of H_2O.

2 mol C_8H_{18} : 25 mol O_2 : 16 mol CO_2 : 18 mol H_2O

Making Pizza

- The number of pizzas you can make depends on the amount of the ingredients you use.

1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

This relationship can be expressed mathematically.

1 crust : 5 oz sauce : 2 cups cheese : 1 pizza

- We can compare the amount of pizza that can be made from 10 cups of cheese:

2 cups cheese : 1 pizza, then,

$$10\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$$

Making Molecules: Mole-to-Mole Conversions

- We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe.

The ratio of the coefficients acts as a conversion factor between the amount in moles of the reactants and products.

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

The ratio acts as a conversion factor between the amount in moles of the reactant, C_8H_{18}, and the amount in moles of the product, CO_2.

Mole-to-Mole Conversions

- Suppose we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

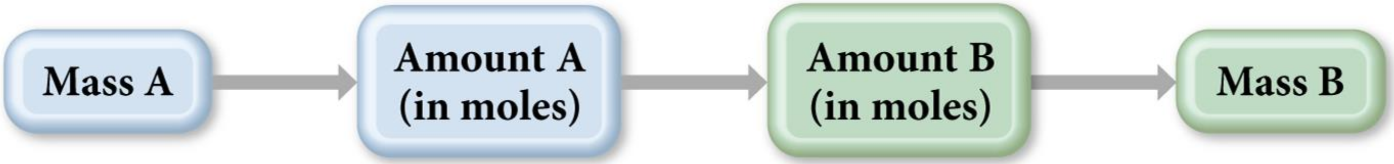
stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

$$22.0\ \cancel{\text{mol } C_8H_{18}} \times \frac{16\ \text{mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 176\ \text{mol } CO_2$$

- The combustion of 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

Making Molecules: Mass-to-Mass Conversions

- The world burned the equivalent of 3.7×10^{15} g of gasoline (octane) in 2013. We can estimate the mass of CO_2 produced based on the flow chart below.



- We use molar mass as a conversion factor between the mass given and amount in moles.
- We use coefficients as the conversion factor between the reactant, C_8H_{18}, and the amount in moles of the product, CO_2, and then molar mass as the conversion factor to get the mass of CO_2 produced.

Mass-to-Mass Conversions

- If we burn 3.7×10^{15} g C_8H_{18}, how many grams of CO_2 form?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

molar mass: C_8H_{18} 114.22 g/mol, CO_2 44.01 g/mol
stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

$$22.0\ \cancel{\text{mol}\ C_8H_{18}} \times \frac{16\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_8H_{18}}} = 176\ \text{mol}\ CO_2$$

- The combustion 3.7×10^{15} g C_8H_{18} adds 1.1×10^{16} g CO_2 to the atmosphere.

Limiting Reactant, Theoretical Yield, Percent Yield

- Recall our pizza recipe:

1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

- If we have 4 crusts, 10 cups of cheese, and 15 oz tomato sauce, how many pizzas can we make?

We have enough crusts to make

$$4\ \cancel{\text{crusts}} \times \frac{1\ \text{pizza}}{1\ \cancel{\text{crust}}} = 4\ \text{pizzas}$$

We have enough cheese to make

$$10\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$$

We have enough tomato sauce to make

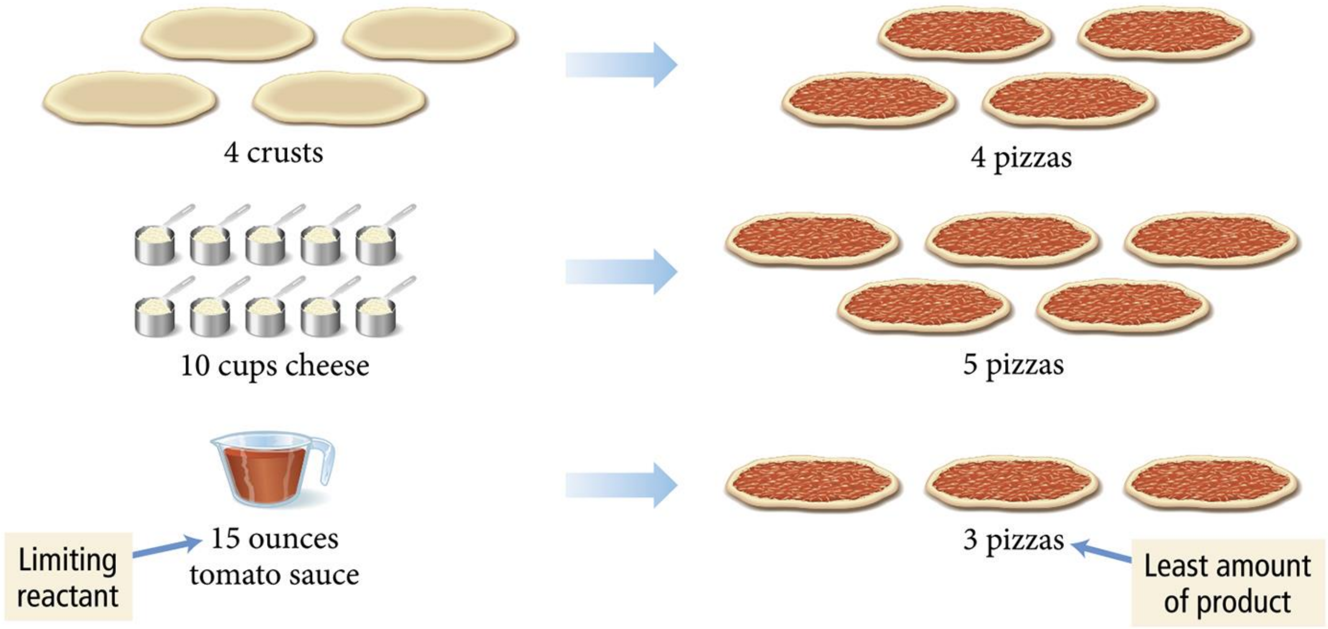
$$15\ \cancel{\text{ounces tomato sauce}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = 3\ \text{pizzas}$$

Limiting reactant

Smallest number of pizzas

Limiting Reactant

- We have enough crusts for four pizzas, enough cheese for five pizzas, but enough tomato sauce for only three pizzas.
  - We can make only three pizzas. The tomato sauce *limits* how many pizzas we can make.



Theoretical Yield

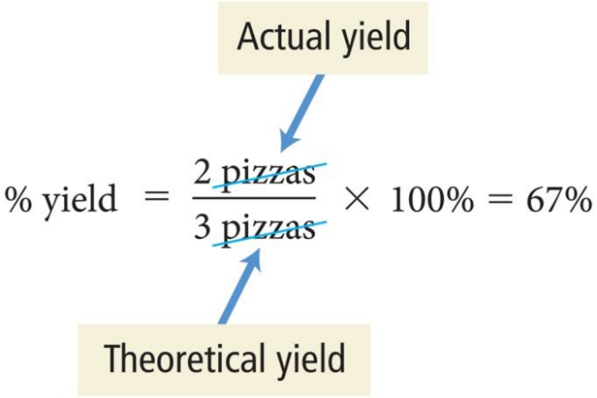
- Tomato sauce is the **limiting reactant**, the reactant that makes *the least amount of product.*
  - The limiting reactant is also known as the *limiting reagent.*
- The maximum number of pizzas we can make depends on this ingredient. In chemical reactions, we call this the **theoretical yield**.
  - This is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.

  - The ingredient that makes the least amount of pizza determines how many pizzas you can make (**theoretical yield**).

Percent Yield

Assume that while making pizzas, we burn a pizza, drop one on the floor, or other uncontrollable events happen so that we make only two pizzas. The actual amount of product made in a chemical reaction is called the **actual yield.**

We can determine the efficiency of making pizzas by calculating the percentage of the maximum number of pizzas we actually make. In chemical reactions, we call this the **percent yield**.



Actual yield

$$\%\ \text{yield} = \frac{2\ \cancel{\text{pizzas}}}{3\ \cancel{\text{pizzas}}} \times 100\% = 67\%$$

Theoretical yield

In a Chemical Reaction

- For reactions with multiple reactants, it is likely that one of the reactants will be completely used before the others.
- When this reactant is used up, the reaction stops and no more product is made.
- The reactant that limits the amount of product is called the **limiting reactant**.
  - It is sometimes called the limiting reagent.
  - The limiting reactant gets completely consumed.
- Reactants not completely consumed are called **excess reactants**.
- The amount of product that can be made from the limiting reactant is called the **theoretical yield**.

Summarizing Limiting Reactant and Yield

- The **limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.

- The **reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.

Summarizing Limiting Reactant and Yield

- The **theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- The **actual yield** is the amount of product actually produced by a chemical reaction.
- The **percent yield** is calculated as follows:

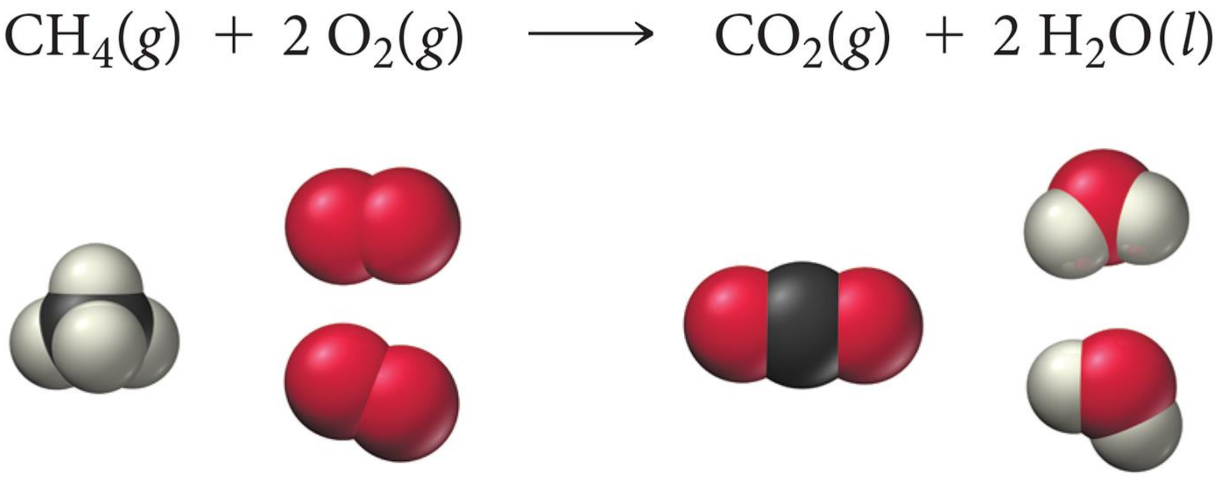
$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield

- Recall our balanced equation for the combustion of methane:

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

  - Our balanced equation for the combustion of methane implies that every one molecule of CH_4 reacts with two molecules of O_2.



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

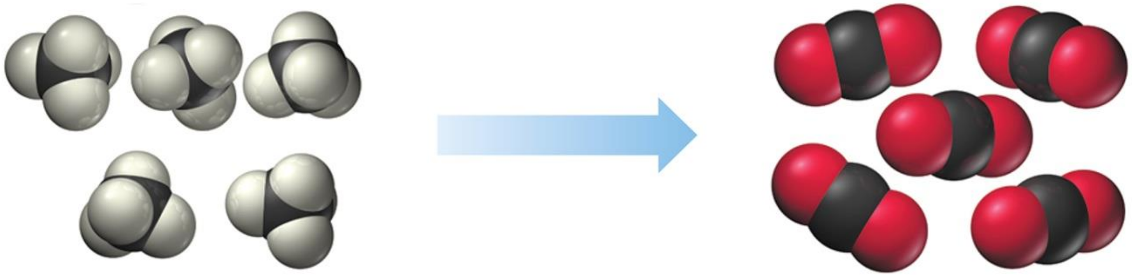
Combustion of Methane

- If we have five molecules of CH_4 and eight molecules of O_2, which is the limiting reactant?

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

  - – First we calculate the number of CO_2 molecules that can be made from five CH_4 molecules.

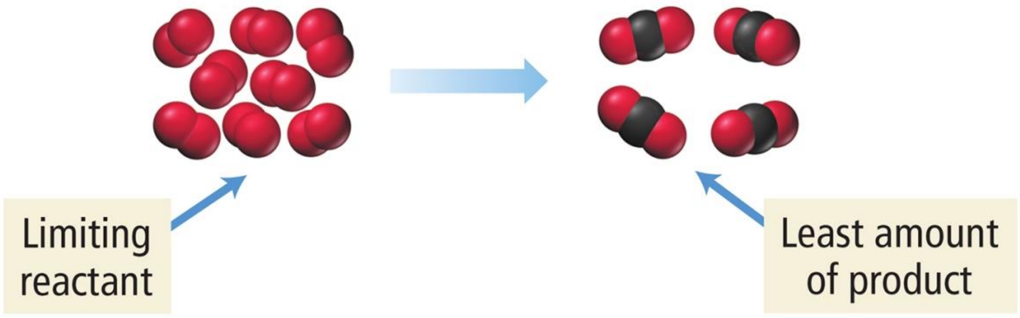
$$5\ \cancel{CH_4} \times \frac{1\ CO_2}{1\ \cancel{CH_4}} = 5\ CO_2$$



Combustion of Methane

- Then we calculate the number of CO_2 molecules that can be made from eight O_2 molecules.

$$8\,\cancel{O_2} \times \frac{1\,CO_2}{2\,\cancel{O_2}} = 4\,CO_2$$



  - We have enough CH_4 to make five CO_2 molecules and four CO_2 molecules.
  - Therefore, O_2 is the limiting reactant, and four CO_2 molecules is the theoretical yield.
  - CH_4 is in excess.

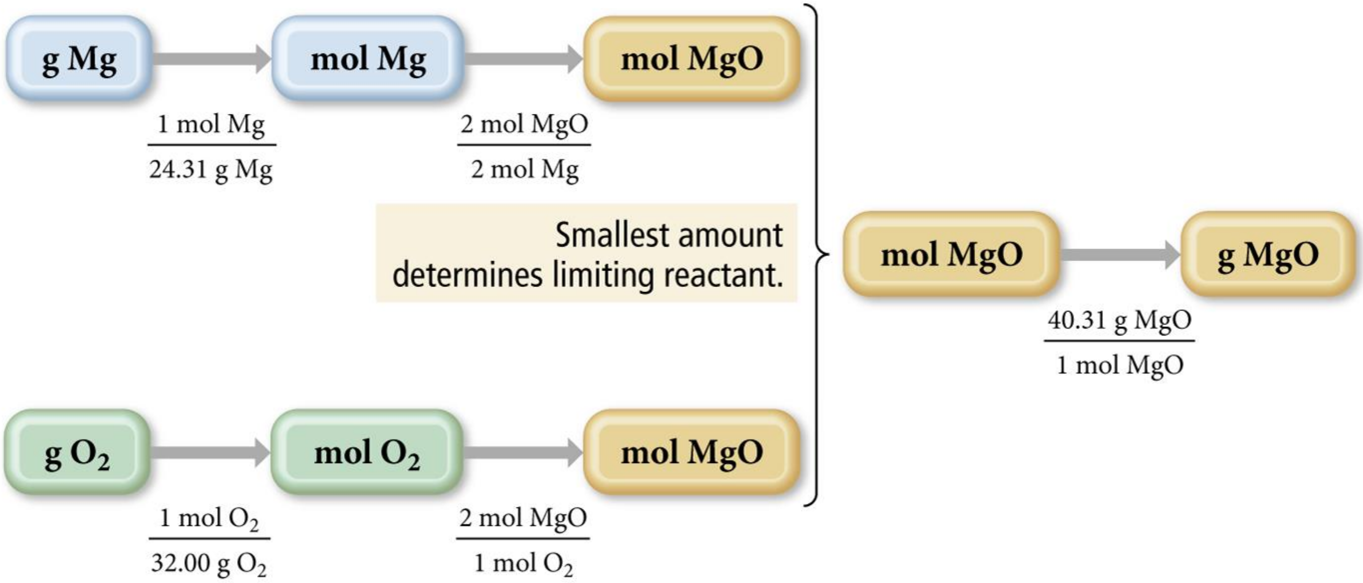
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

- When working in the lab, we normally measure reactant quantities in grams.
- To find the limiting reactant and theoretical yield, we must first convert grams to moles.

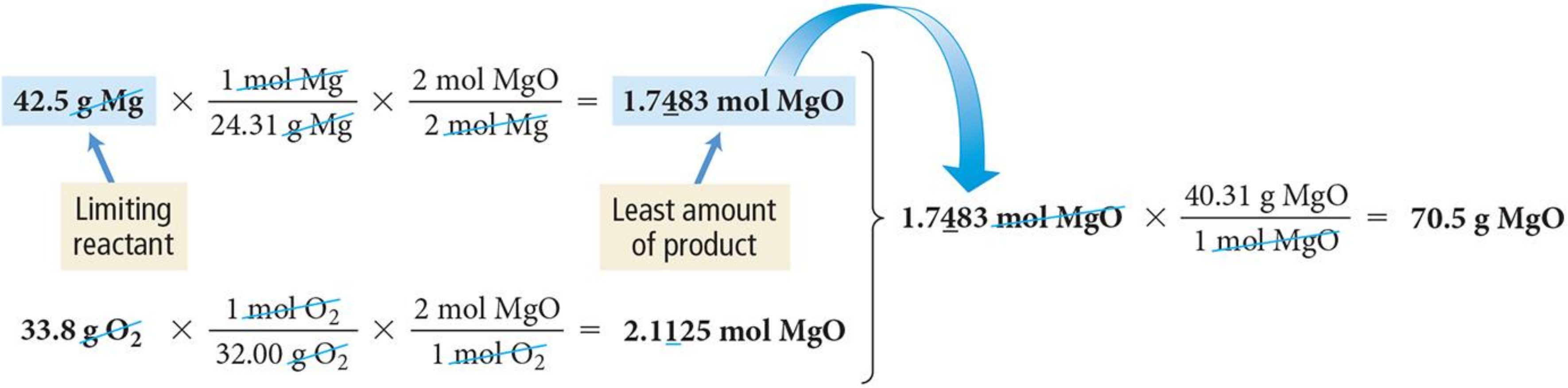
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

- A reactant mixture contains 42.5 g Mg and 33.8 g O_2. What is the limiting reactant and theoretical yield?

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



Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses



$$42.5\ \cancel{\text{g Mg}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = 1.7\underline{4}83\ \text{mol MgO}$$

Limiting reactant

Least amount of product

$$33.8\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol O}_2}} = 2.1\underline{1}25\ \text{mol MgO}$$

$$1.7\underline{4}83\ \cancel{\text{mol MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = 70.5\ \text{g MgO}$$

Solution Concentration and Solution Stoichiometry

- When table salt is mixed with water, it seems to disappear or become a liquid; the mixture is homogeneous.
  - The salt is still there, as you can tell from the taste or simply boiling away the water.
- Homogeneous mixtures are called **solutions**.
- The majority component is the **solvent**.
- The minority component is the **solute**.
- A solution in which water is the solvent is an **aqueous solution**.

Solution Concentration

- Because solutions are mixtures, the composition can vary from one sample to another.
  - Pure substances have constant composition.
  - Saltwater samples from different seas or lakes have different amounts of salt.

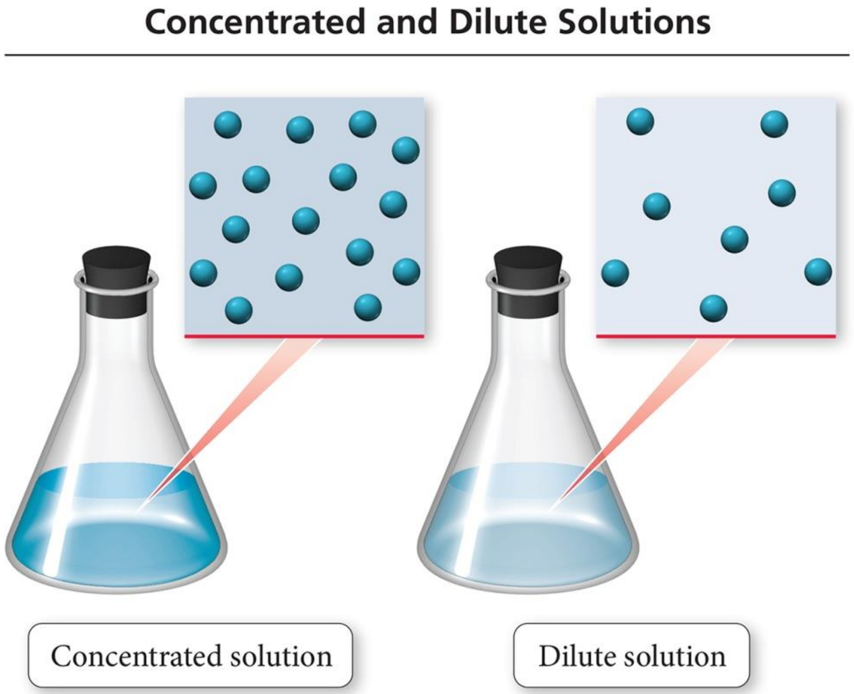
- So, to describe solutions accurately, we quantify the amount of solute relative to solvent, or **concentration of solution**.

Solution Concentration

- Solutions are often described quantitatively, as dilute or concentrated.

- **Dilute solutions** have a small amount of solute compared to solvents.

- **Concentrated solutions** have a large amount of solute compared to solvents.

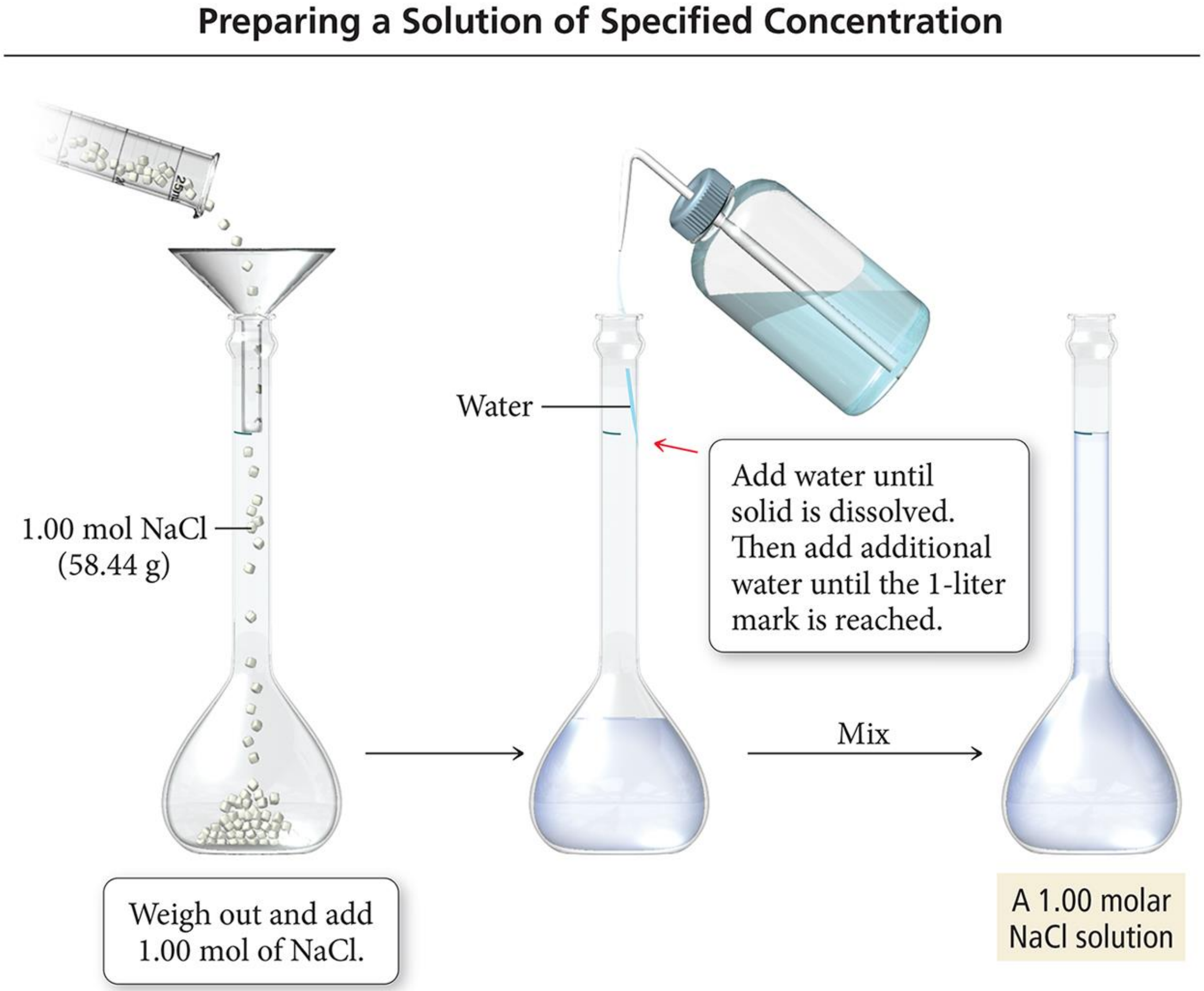


Solution Concentration: Molarity

- A common way to express solution concentration is **molarity (*M*)**.
  - Molarity is the amount of solute (in moles) divided by the volume of solution (in liters).

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

Preparing 1 L of a 1.00 M NaCl Solution



Using Molarity in Calculations

- We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution.
  - For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.

$$\frac{0.500\ \text{mol NaCl}}{\text{L solution}} \quad \textit{converts} \quad \boxed{\textbf{L solution}} \longrightarrow \boxed{\textbf{mol NaCl}}$$

$$\frac{\text{L solution}}{0.500\ \text{mol NaCl}} \quad \textit{converts} \quad \boxed{\textbf{mol NaCl}} \longrightarrow \boxed{\textbf{L solution}}$$

Solution Dilution

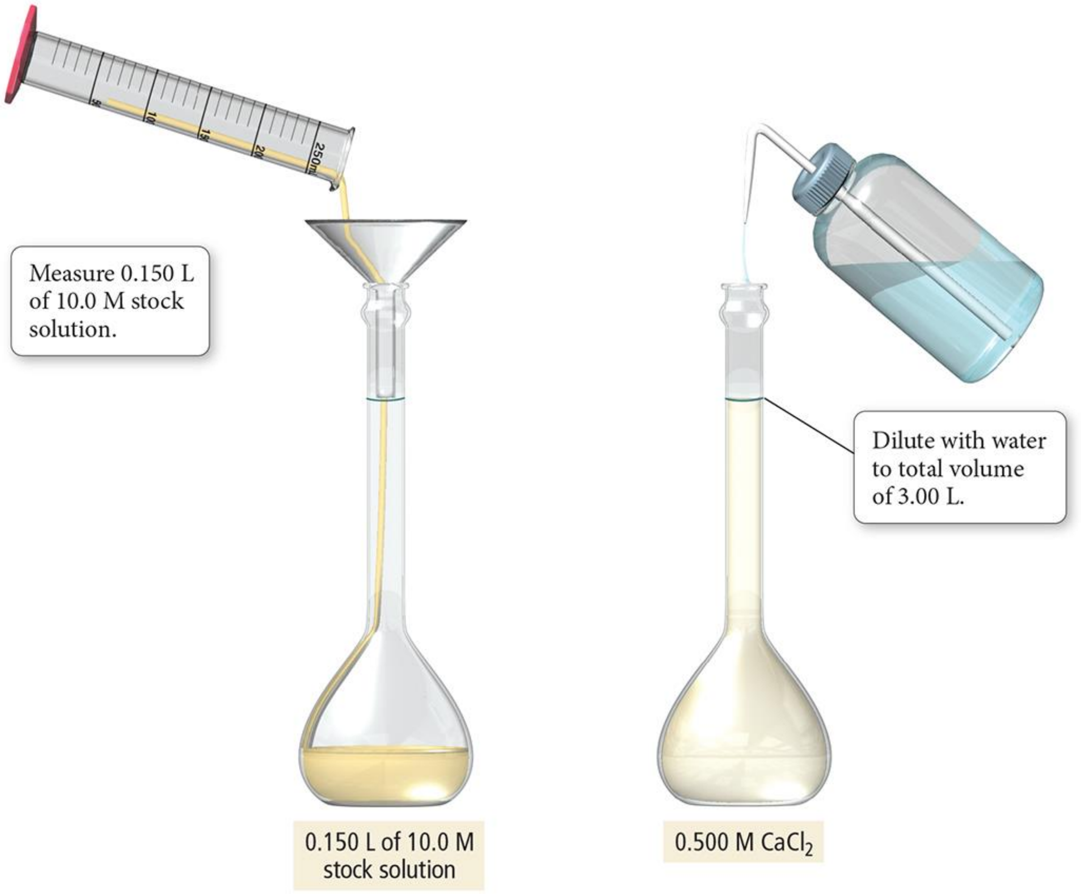
- Often, solutions are stored as concentrated **stock solutions**.
- To make solutions of lower concentrations from these stock solutions, more solvent
  is added.
  - **The amount of solute doesn't change, just the volume of solution:**

  moles solute in solution 1 = moles solute in solution 2
- The concentrations and volumes of the stock and new solutions are inversely proportional:

$$M_1 \cdot V_1 = M_2 \cdot V_2$$

Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution

Diluting a Solution



Measure 0.150 L of 10.0 M stock solution.

Dilute with water to total volume of 3.00 L.

0.150 L of 10.0 M stock solution

0.500 M $CaCl_2$

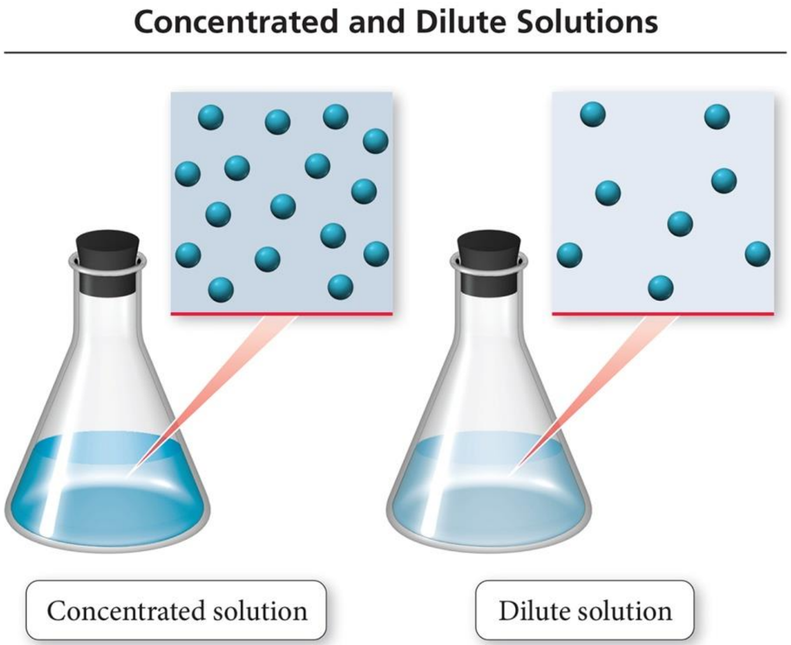
$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\ \cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

Solution Stoichiometry

- Because molarity relates the moles of solute to the liters of solution, it can be used to convert between amount of reactants and/or products in a chemical reaction.
  - The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product.

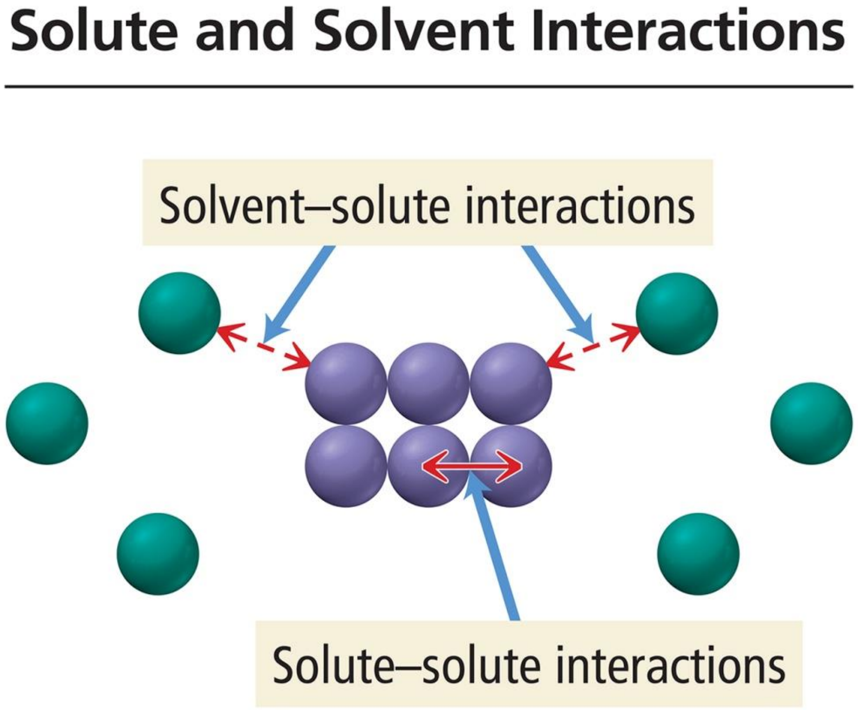


Types of Aqueous Solutions and Solubility

- Consider two familiar aqueous solutions: saltwater and sugar water.
  - Saltwater is a homogeneous mixture of NaCl and H_2O.
  - Sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O.
- As you stir either of these two substances into the water, it seems to disappear.
  - How do solids such as salt and sugar dissolve in water?

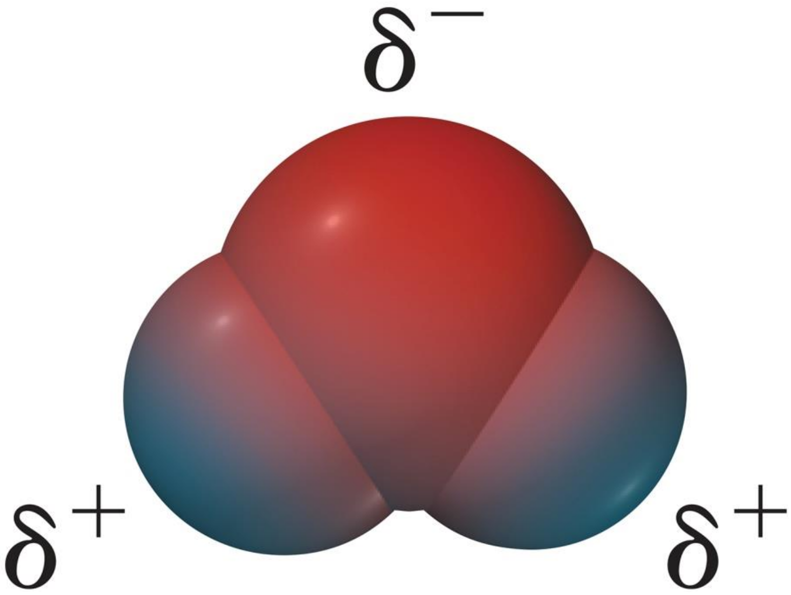
What Happens When a Solute Dissolves?

- There are attractive forces between the solute particles holding them together.
- There are also attractive forces between the solvent molecules.
- When we mix the solute with the solvent, there are attractive forces between the solute particles and the solvent molecules.
- If the attractions between solute and solvent are strong enough, the solute will dissolve.



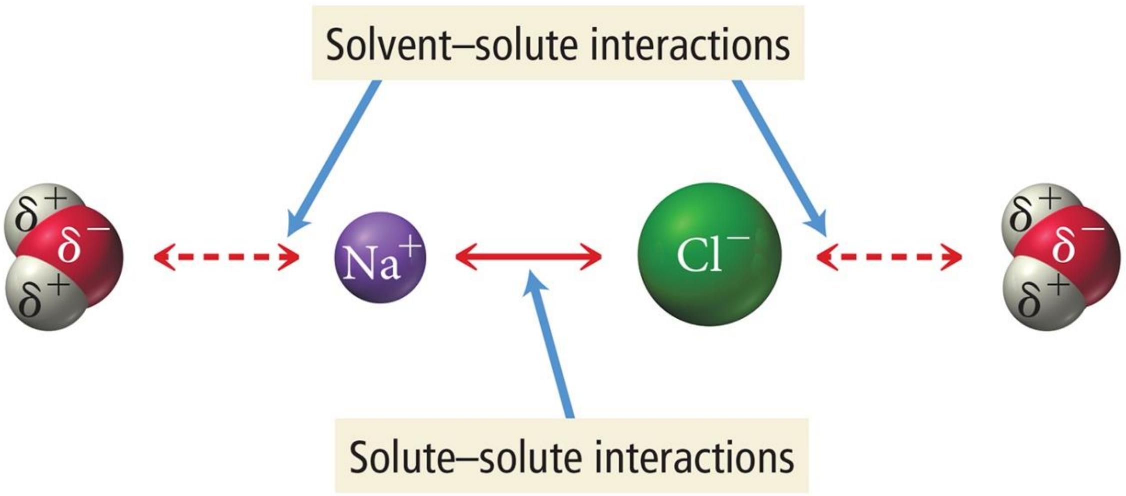
Charge Distribution in a Water Molecule

- There is an uneven distribution of electrons within the water molecule.
  - This causes the oxygen side of the molecule to have a partial negative charge (δ^-) and the hydrogen side to have a partial positive charge (δ^+).



Solute and Solvent Interactions in a Sodium Chloride Solution

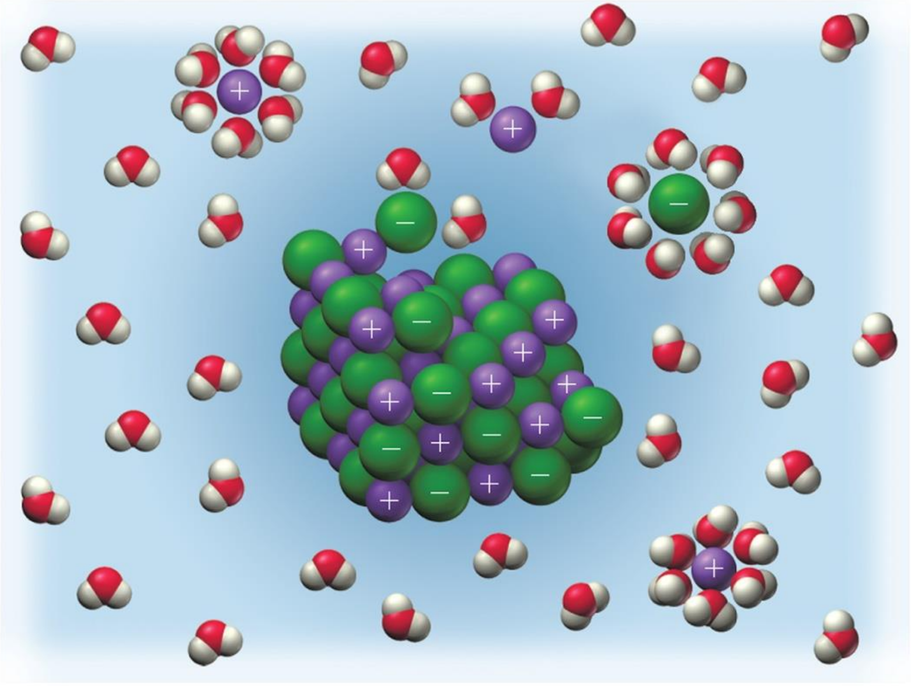
- When sodium chloride is put into water, the attraction of Na^+ and Cl^- ions to water molecules competes with the attraction among the oppositely charged ions themselves.


Interactions in a Sodium Chloride Solution

Dissolution of Ionic Compounds

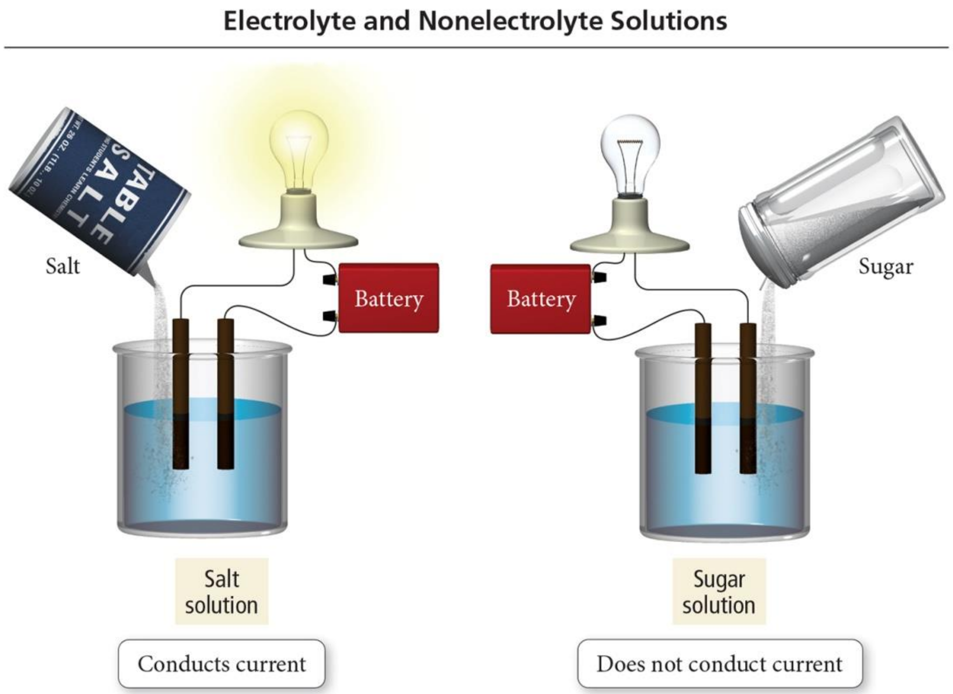
- Each ion is attracted to the surrounding water molecules and pulled off and away from the crystal.
- When it enters the solution, the ion is surrounded by water molecules, insulating it from other ions.
- The result is a solution with free-moving, charged particles able to conduct electricity.

Dissolution of an Ionic Compound



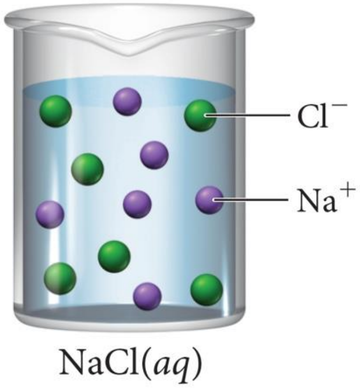
Electrolyte and Nonelectrolyte Solutions

- Materials that dissolve in water to form a solution containing ions will conduct electricity. These are called **electrolytes**.
- Materials that dissolve in water to form a solution with no ions will not conduct electricity. These are called **nonelectrolytes**.
- A solution of salt (an electrolyte) conducts electrical current. A solution of sugar (a nonelectrolyte) does not.

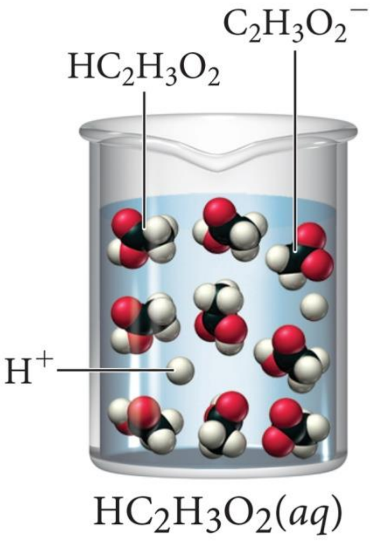


Electrolyte and Nonelectrolyte Solutions

- Ionic substances, such as sodium chloride, that completely dissociate into ions when they dissolve in water, are **strong electrolytes**.
- Except for acids, most molecular compounds, for example sugar, dissolve in water as intact molecules, or **nonelectrolytes**.
- Acids ionize to varying degrees in water. Those that completely ionize are **strong acids**. Those that don’t are **weak acids**.

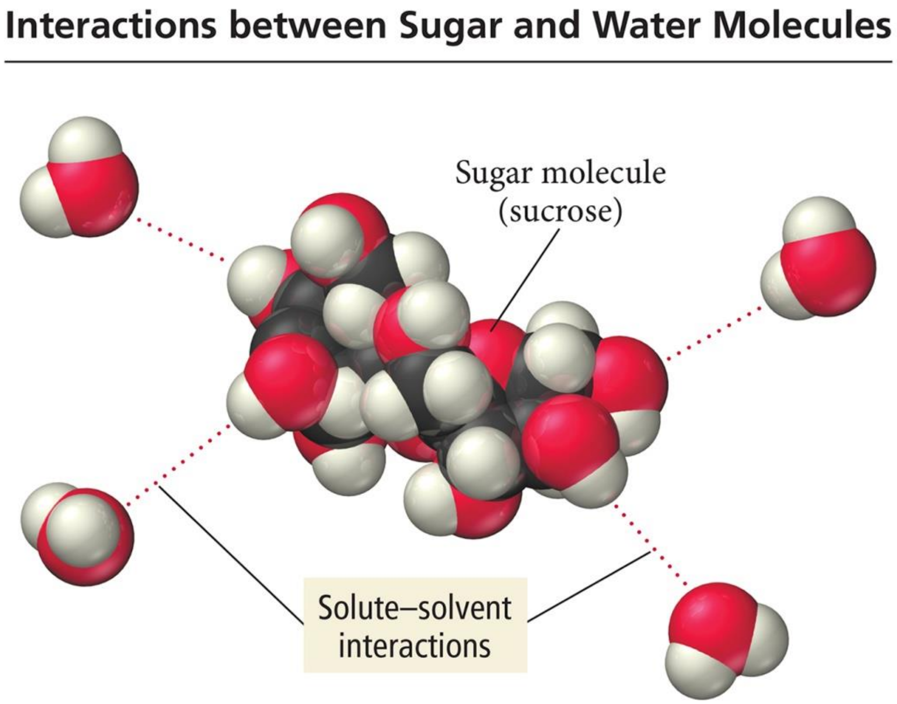


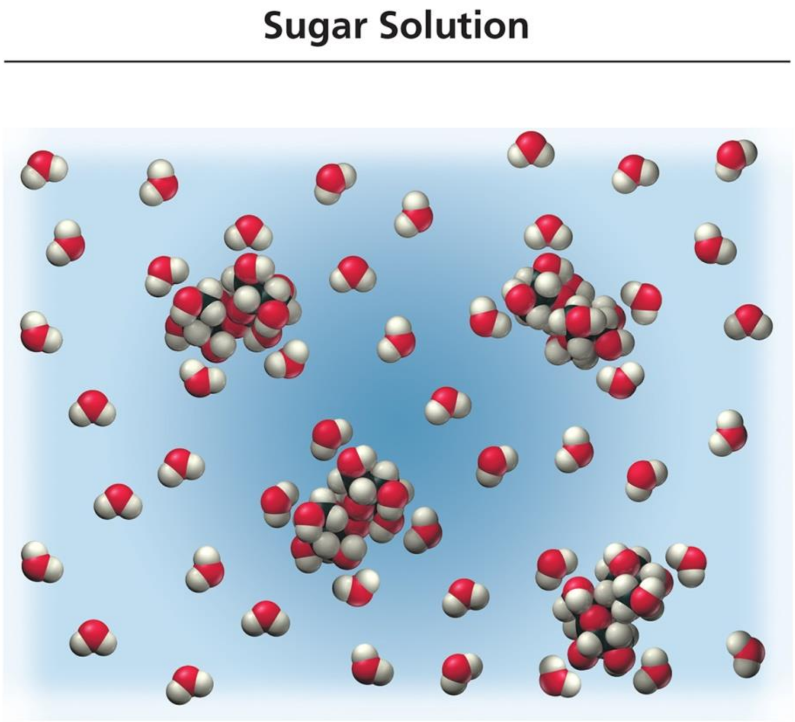
Strong electrolyte



Weak acid

Sugar Dissolution in Water


Interactions between Sugar and Water Molecules


Sugar Solution

Acids

- Acids are molecular compounds that **ionize** when they dissolve in water.
  - The molecules are pulled apart by their attraction for the water.
  - When acids ionize, they form H^+ cations and also anions.
- The percentage of molecules that ionize varies from one acid to another.
- Acids that ionize virtually 100% are called **strong acids**.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- Acids that only ionize a small percentage are called **weak acids**.

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$$

Strong and Weak Electrolytes

- **Strong electrolytes** are materials that dissolve completely as ions.
  - Ionic compounds and strong acids.
  - Solutions are good conductors of electricity.
- **Weak electrolytes** are materials that dissolve mostly as molecules but partially as ions.
  - Weak acids.
  - Solutions conduct electricity, but not well.

Dissociation and Ionization

- When ionic compounds dissolve in water, the anions and cations are separated from each other. This is called **dissociation**.

$$Na_2S(aq) \rightarrow 2\ Na^+(aq) + S^{2-}(aq)$$

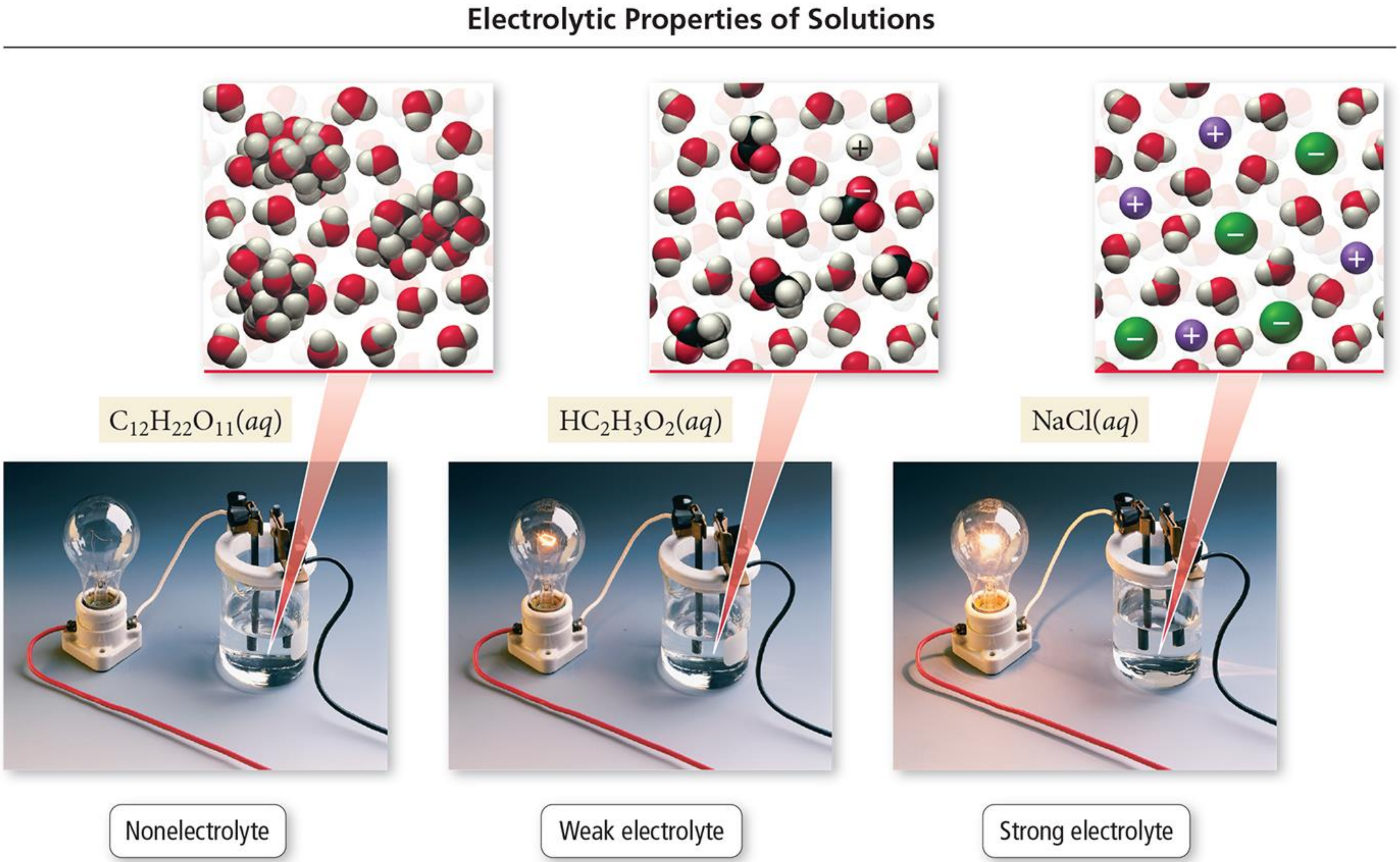
- When compounds containing polyatomic ions dissociate, the polyatomic group stays together as one ion.

$$Na_2SO_4(aq) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq)$$

- When strong acids dissolve in water, the molecule **ionizes** into H^+ and anions.

$$H_2SO_4(aq) \rightarrow 2\ H^+(aq) + SO_4^{2-}(aq)$$

Classes of Dissolved Materials

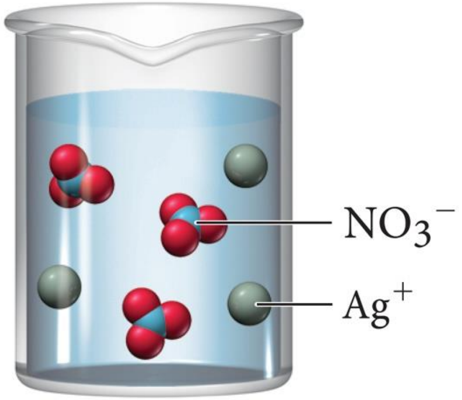


The Solubility of Ionic Compounds

- When an ionic compound dissolves in water, the resulting solution contains not the intact ionic compound itself but its component ions dissolved in water.
- However, not all ionic compounds dissolve in water. For example, AgCl remains solid and appears as a white powder at the bottom of the water.
- In general, a compound is termed **soluble** if it dissolves in water and **insoluble** if it does not.

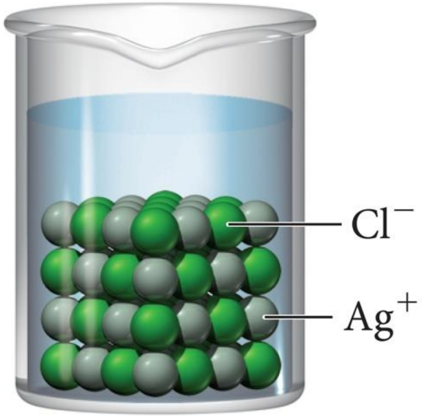
Solubility of Salts

- If we mix solid $AgNO_3$ with water, it dissolves and forms a strong electrolyte solution.
- Silver chloride, on the other hand, is almost completely insoluble.
  - If we mix solid AgCl with water, virtually all of it remains as a solid within the liquid water.



$AgNO_3(aq)$

Soluble



$AgCl(s)$

Insoluble

When Will a Salt Dissolve?

- Whether a particular compound is soluble or insoluble depends on several factors.
- Predicting whether a compound will dissolve in water is not easy.
- The best way to do it is to conduct experiments to test whether a compound will dissolve in water, and then develop some rules based on those experimental results.
  - – We call this method the **empirical method**.

Solubility Rules

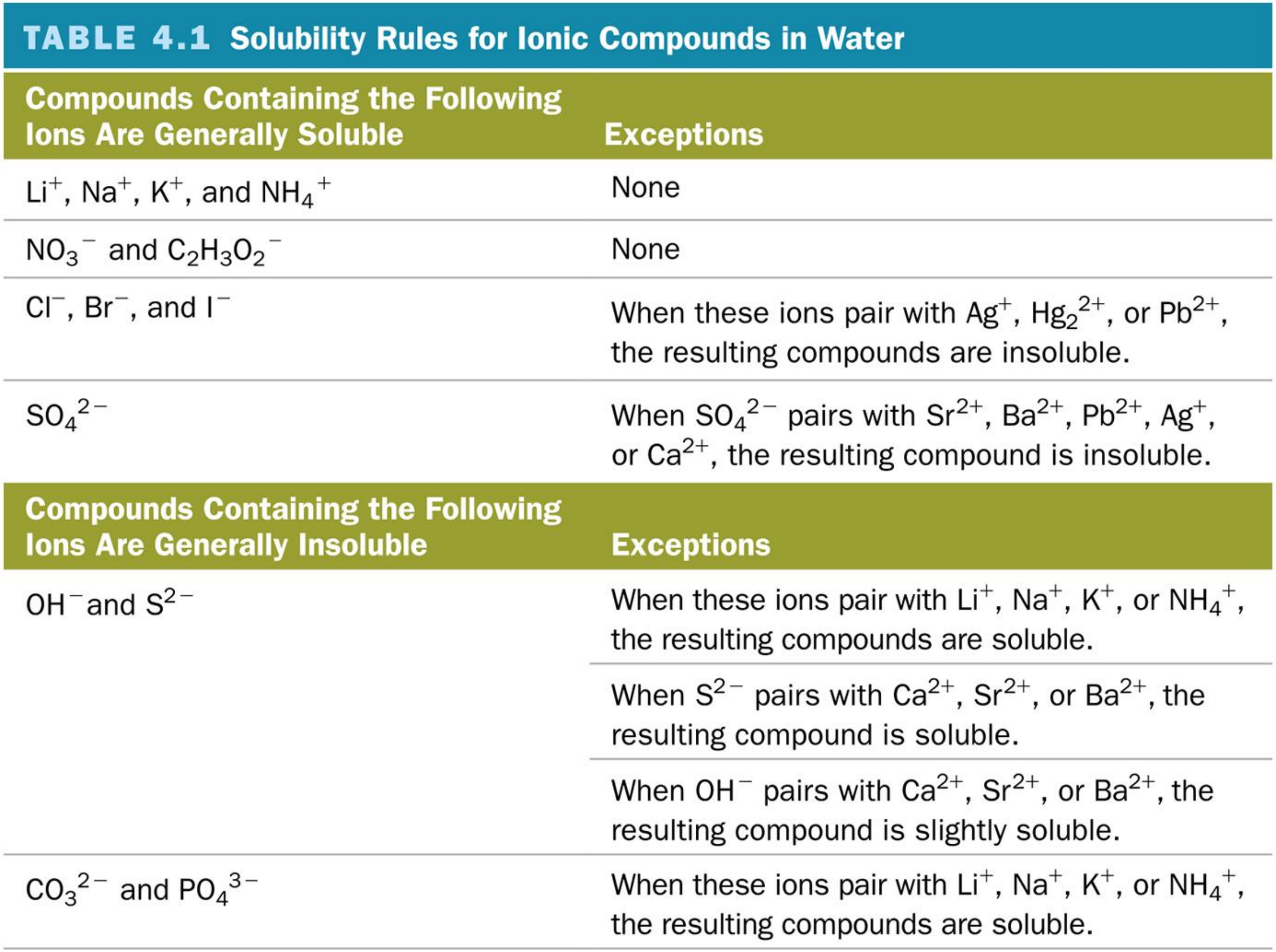


TABLE 4.1 Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li^+, Na^+, K^+, and NH_4^+	None
NO_3^- and $C_2H_3O_2^-$	None
Cl^-, Br^-, and I^-	When these ions pair with Ag^+, Hg_2^{2+}, or Pb^{2+}, the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, Ag^+, or Ca^{2+}, the resulting compound is insoluble.
Compounds Containing the Following Ions Are Generally Insoluble	**Exceptions**
OH^- and S^{2-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.
	When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble.
	When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is slightly soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.

Precipitation Reactions

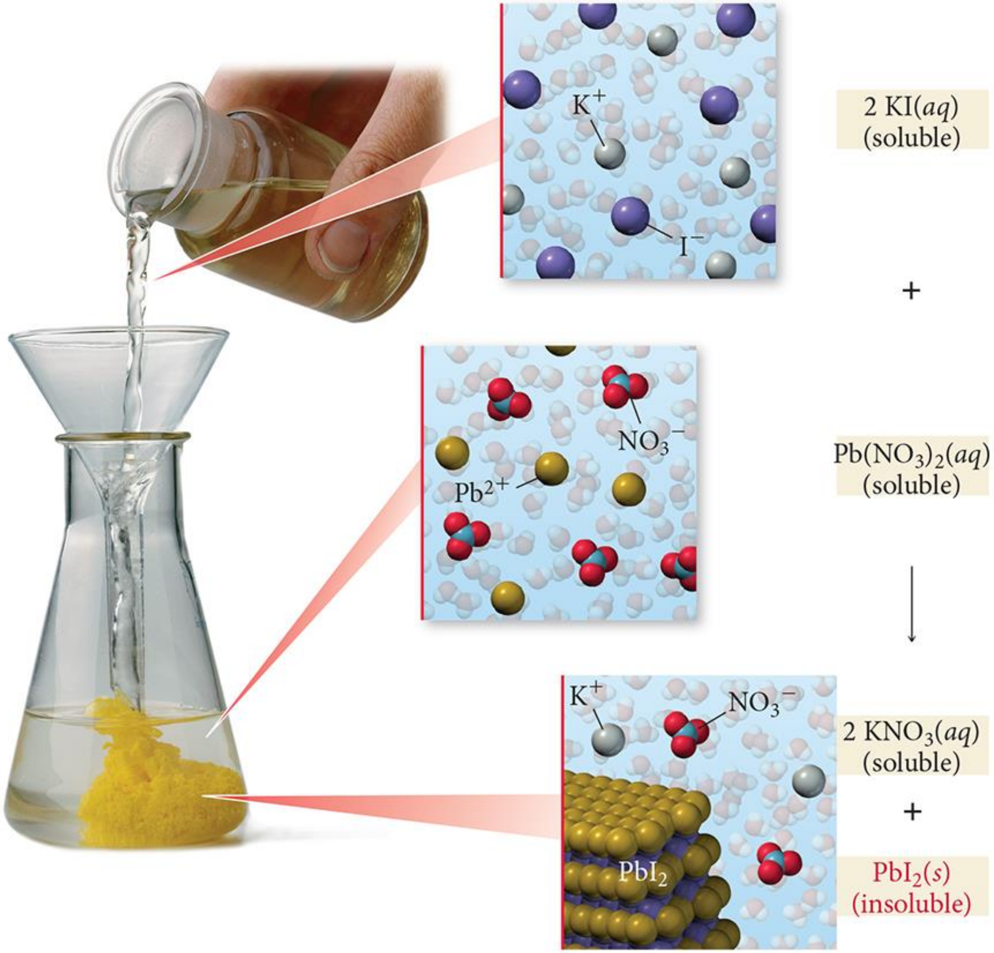
- **Precipitation reactions** are reactions in which a solid forms when we mix two solutions.
  - Reactions between aqueous solutions of ionic compounds produce an ionic compound that is insoluble in water.
  - The insoluble product is called a **precipitate**.

Precipitation of Lead(II) Iodide

Precipitation Reaction

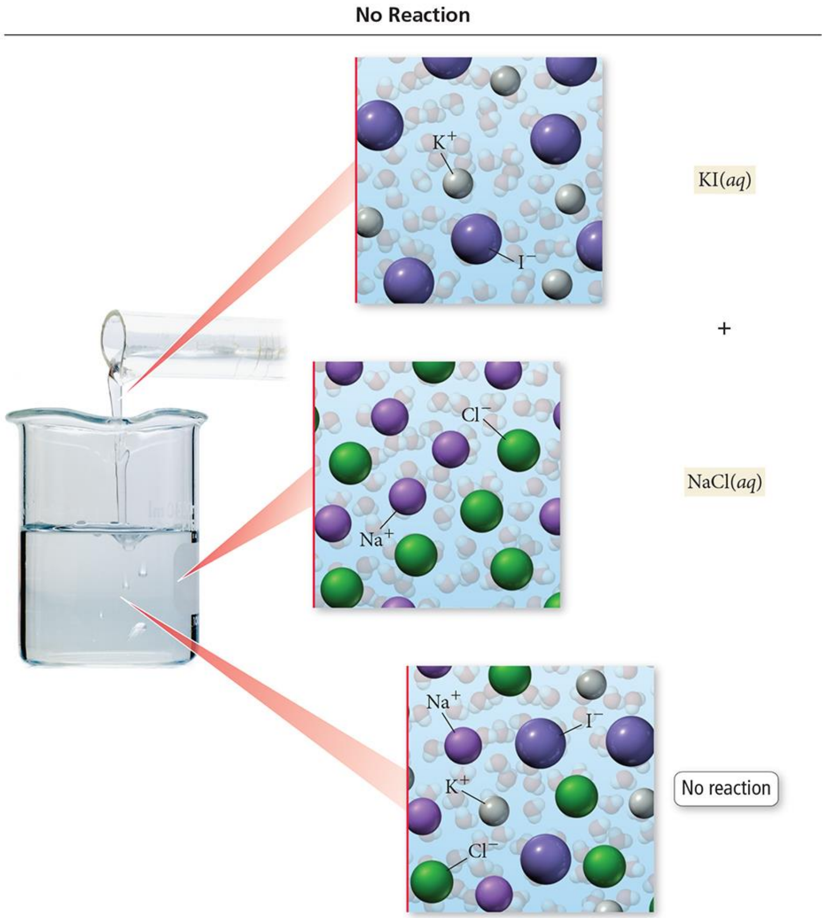
$$\underset{\text{(soluble)}}{2\ \mathrm{KI}(aq)} + \underset{\text{(soluble)}}{\mathrm{Pb(NO_3)_2}(aq)} \longrightarrow \underset{\text{(insoluble)}}{\mathrm{PbI_2}(s)} + \underset{\text{(soluble)}}{2\ \mathrm{KNO_3}(aq)}$$

When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.



No Precipitation Means No Reaction

- Precipitation reactions do not always occur when two aqueous solutions are mixed.
  - – Nothing happens when combining solutions of KI and NaCl.



KI(*aq*) + NaCl(*aq*) → No Reaction

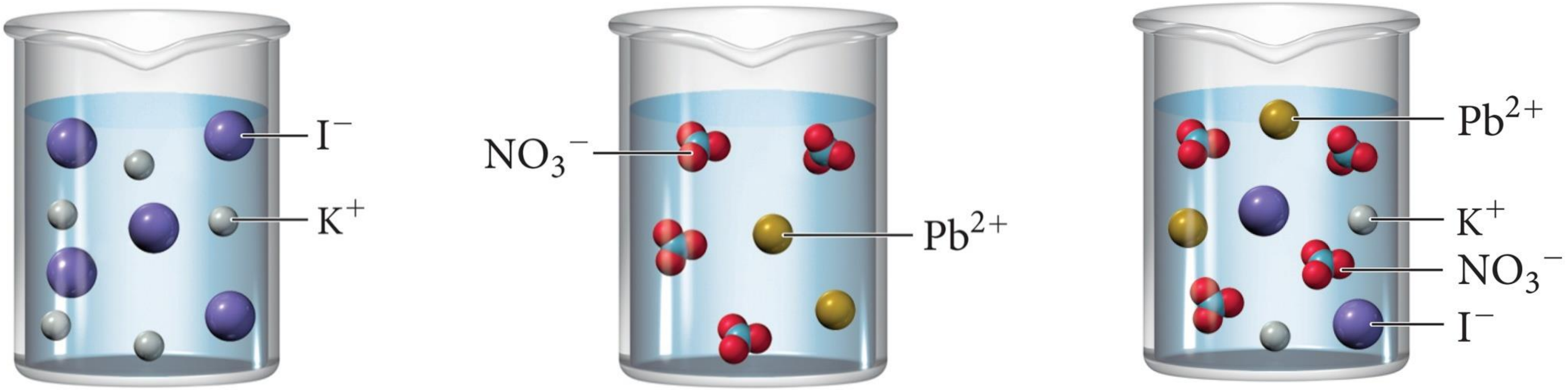
Predicting Precipitation Reactions

1. Determine what ions each aqueous reactant has.
2. Determine formulas of possible products.
   - Exchange ions.
     - (+) ion from one reactant with (–) ion from other
   - Balance charges of combined ions to get the formula of each product.
3. Determine solubility of each product in water.
   - Use the solubility rules.
   - If the product is insoluble or slightly soluble, it will precipitate.
4. If neither product will precipitate, write **no reaction** after the arrow.

Predicting Precipitation Reactions

5. If any of the possible products are insoluble, write their formulas as the products of the reaction using **(*s*)** after the formula to indicate **solid**. Write any soluble products with **(*aq*)** after the formula to indicate ***aqueous***.
6. Balance the equation.
   - Remember to change only coefficients, not subscripts.

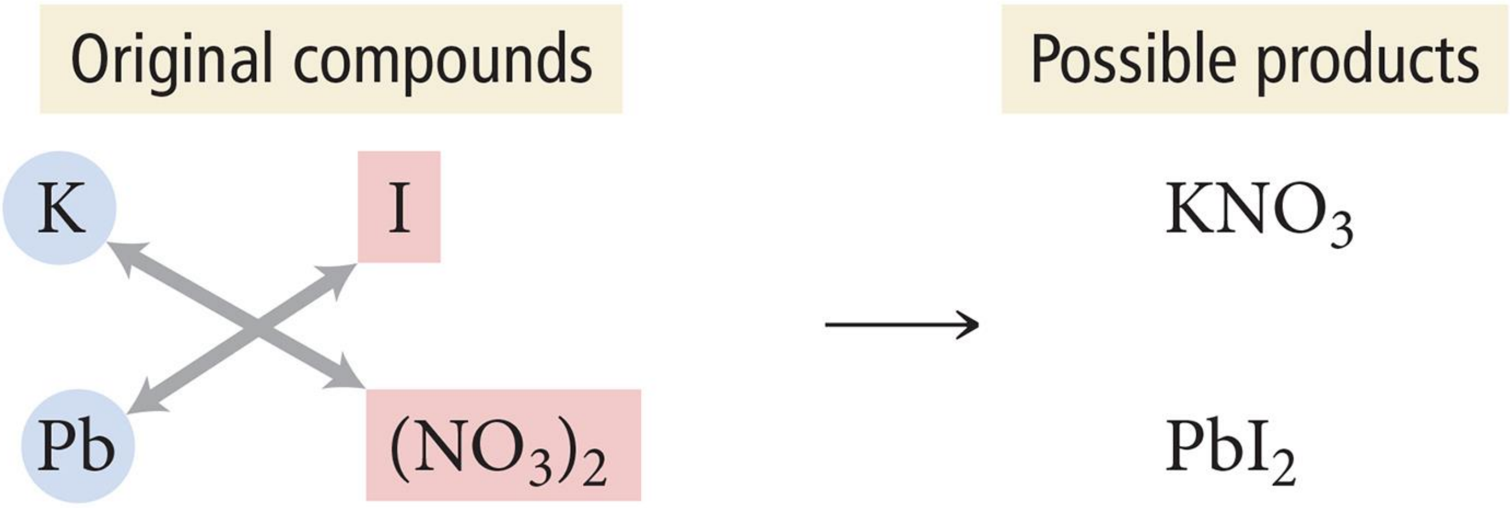
Predicting Precipitation Reactions



KI(*aq*)

$Pb(NO_3)_2(aq)$

KI(*aq*) and $Pb(NO_3)_2(aq)$



Original compounds

K

I

Pb

$(NO_3)_2$

$\longrightarrow$

Possible products

KNO_3

PbI_2

Representing Aqueous Reactions

- An equation showing the complete neutral formulas for each compound in the aqueous reaction as if they existed as molecules is called a **molecular equation**.

  $2\ KOH(aq) + Mg(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + Mg(OH)_2(s)$

- In actual solutions of soluble ionic compounds, dissolved substances are present as ions. Equations that describe the nature of the dissolved species in solution are called **complete ionic equations**.

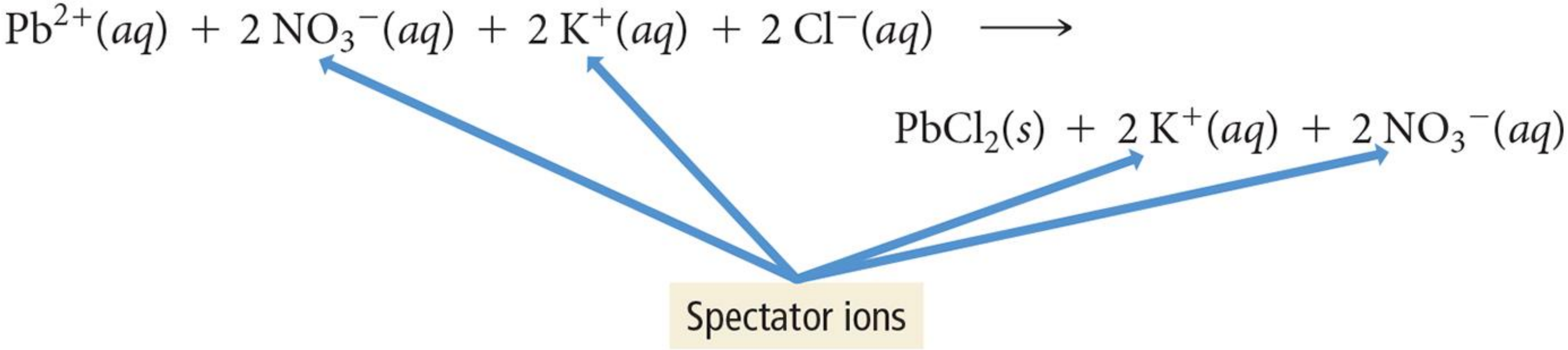
Rules for Writing a Complete Ionic Equation

- Aqueous strong electrolytes (soluble salts, strong acids, strong bases) are written as ions.
- Insoluble substances, weak electrolytes, and nonelectrolytes are written in molecule form.
  - Solids, liquids, and gases are not dissolved, hence molecule form

$$2\ K^+_{(aq)} + 2\ OH^-_{(aq)} + Mg^{2+}_{(aq)} + 2\ NO_3{}^-_{(aq)} \rightarrow$$
$$2\ K^+_{(aq)} + 2\ NO_3{}^-_{(aq)} + Mg(OH)_{2(s)}$$

Ionic Equation

- Notice that in the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation.
- These ions are called **spectator ions** because they do not participate in the reaction (soluble salts, strong acids, and strong bases).



$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Net Ionic Equation

- An ionic equation in which the spectator ions are removed is called a **net ionic equation**.

$$\cancel{2\,\mathrm{K}}^{+}{}_{(aq)} + 2\,\mathrm{OH}^{-}{}_{(aq)} + \mathrm{Mg}^{2+}{}_{(aq)} \cancel{+ 2\,\mathrm{NO}_3}^{-}{}_{(aq)} \rightarrow$$
$$2\,\mathrm{K}^{+}{}_{(aq)} + 2\,\cancel{\mathrm{NO}_3}^{-}{}_{(aq)} \cancel{+ \mathrm{Mg(OH)}}_{2(s)}$$

- The net ionic equation is

$$2\,\mathrm{OH}^{-}{}_{(aq)} + \mathrm{Mg}^{2+}{}_{(aq)} \rightarrow \mathrm{Mg(OH)}_{2(s)}$$

Examples

- Write the ionic and net ionic equation for each of the following:

1. $K_2SO_4(aq) + 2\ AgNO_3(aq) \rightarrow 2\ KNO_3(aq) + Ag_2SO_4(s)$

$$2\ K^+(aq) + SO_4^{2-}(aq) + 2\ Ag^+(aq) + 2\ NO_3^-(aq) \rightarrow$$
$$2\ K^+(aq) + 2\ NO_3^-(aq) + Ag_2SO_4(s)$$
$$2\ Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$$

2. $Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + CO_2(g) + H_2O(l)$

$$2\ Na^+(aq) + CO_3^{2-}(aq) + 2\ H^+(aq) + 2\ Cl^-(aq) \rightarrow$$
$$2\ Na^+(aq) + 2\ Cl^-(aq) + CO_2(g) + H_2O(l)$$
$$CO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

Acid–Base and Gas-Evolution Reactions

- Two other important classes of reactions that occur in aqueous solution are
  1. acid–base reactions and
  2. gas-evolution reactions.
- Acid–base reaction:
  - An **acid–base reaction** is also called a **neutralization reaction**.
  - An acid reacts with a base, and the two neutralize each other, producing water (or in some cases a weak electrolyte).

Acid–Base and Gas-Evolution Reactions

- In a **gas-evolution reaction**, a gas is produced, resulting in bubbling.
- In both acid–base and gas-evolution reactions, as in precipitation reactions, the reactions occur when the anion from one reactant combines with the cation of the other.
- Many gas-evolution reactions are also acid–base reactions.

Acid–Base Reactions

Arrhenius Definitions:

- Acid: Substance that produces H^+ in aqueous solution. $HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$
- In solution, H^+ bonds with water to produce the **hydronium ion**, H_3O+.
- **Polyprotic acids** contain more than one ionizable proton and release them sequentially.
- The first ionizable proton is strong while subsequent ionizable protons are weak.
- Base: Substance that produces OH^- ions in aqueous solution.

$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$

Acid–Base Reactions

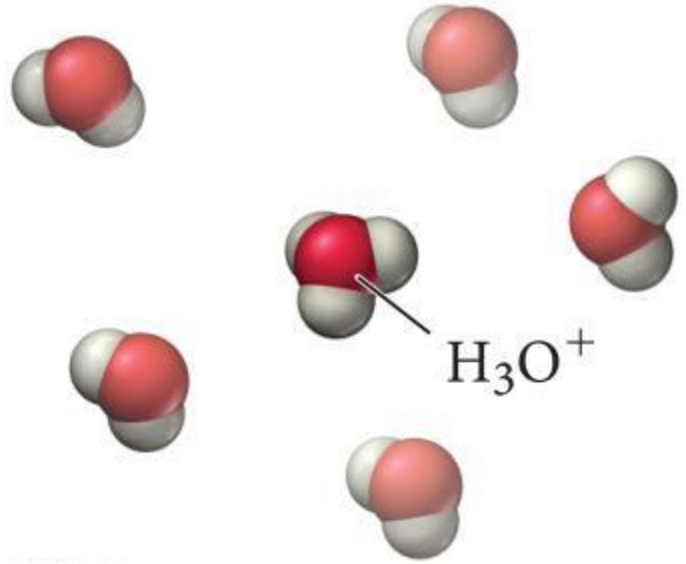
- These reactions are called **neutralization reactions** because the acid and base neutralize each other's properties.

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$$

- The net ionic equation for an acid–base reaction is

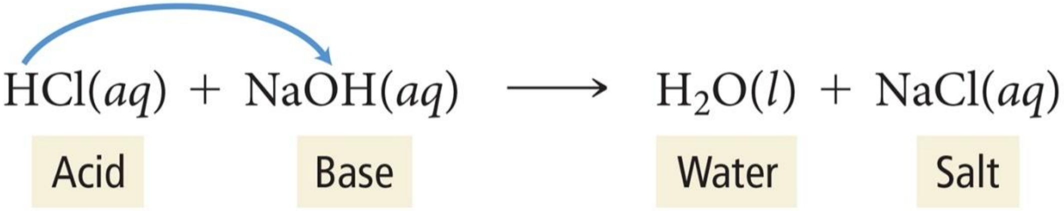
$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

  - As long as the salt that forms is soluble in water.



Acids and Bases in Solution

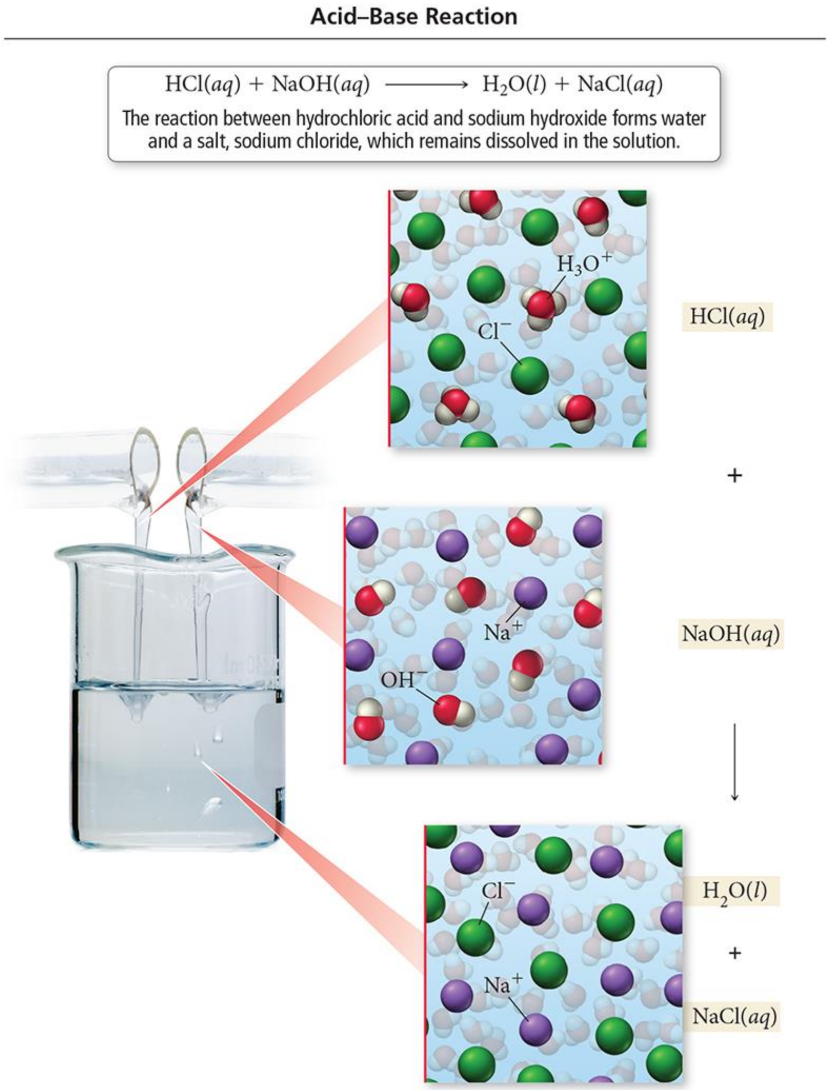
- Acids ionize in water to form H^+ ions.
  - More precisely, the H^+ from the acid molecule is donated to a water molecule to form **hydronium ion, H_3O^+**.
    - Most chemists use H^+ and H_3O^+ interchangeably.
- Bases dissociate in water to form OH^- ions.
  - Bases, such as NH_3, that do not contain OH^- ions, produce OH^- by pulling H^+ off water molecules.
- In the reaction of an acid with a base, the H^+ from the acid combines with the OH^- from the base to make water.
- The cation from the base combines with the anion from the acid to make the salt.



$$\text{HCl}(aq) + \text{NaOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$$

Acid–Base Reaction

- HCl(*aq*) + NaOH(*aq*) → NaCl(*aq*) + H_2O(*l*)



Acid–Base Reaction

HCl(*aq*) + NaOH(*aq*) ⟶ H_2O(*l*) + NaCl(*aq*)

The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.

Some Common Acids and Bases

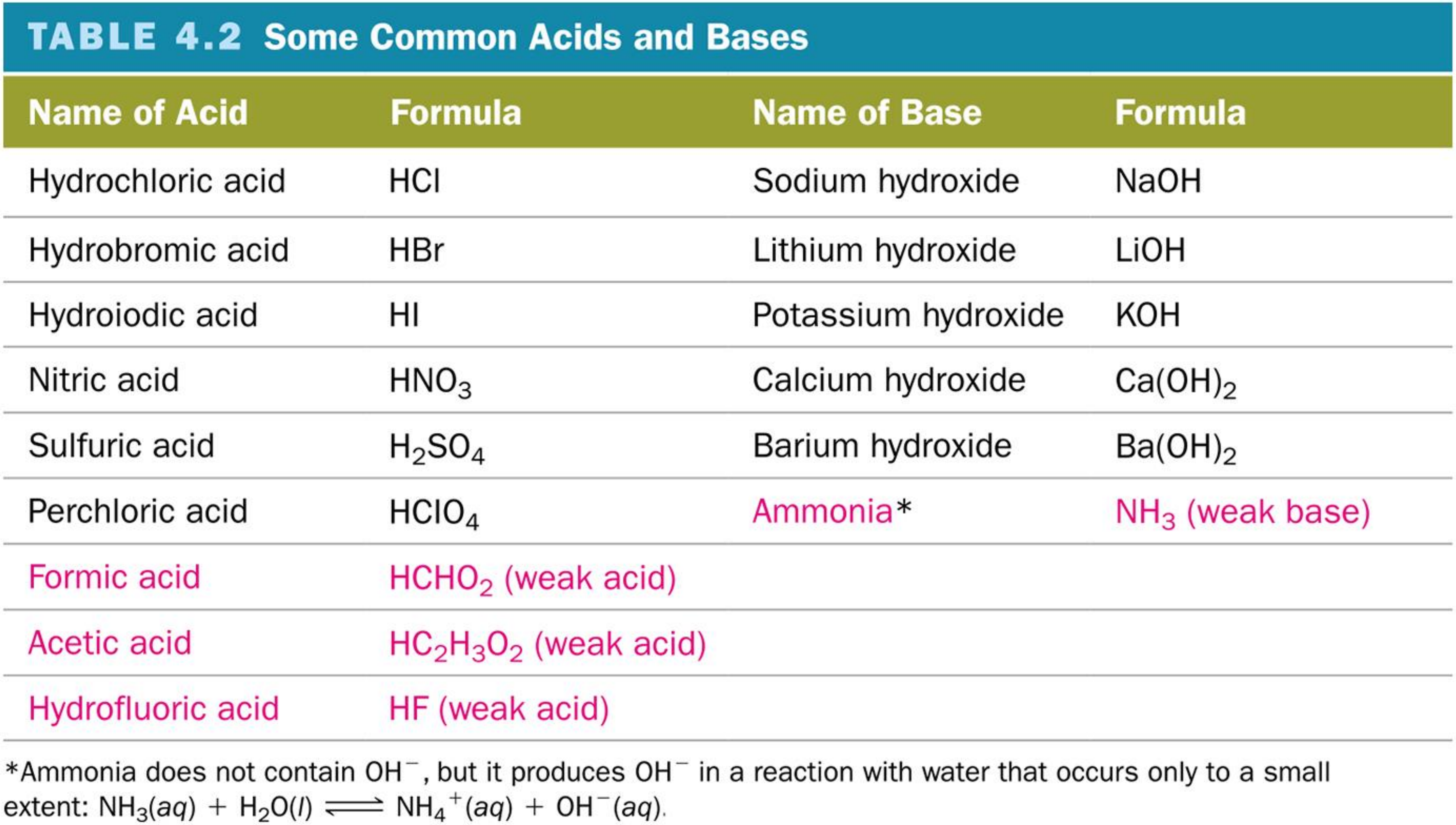


TABLE 4.2 Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Formic acid	$HCHO_2$ (weak acid)		
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

Predict the Product of the Reactions

1. $HCl(aq) + Ba(OH)_2(aq) \rightarrow$

$(H^+ + Cl^-) + (Ba^{2+} + OH^-) \rightarrow (H^+ + OH^-) + (Ba^{2+} + Cl^-)$

$HCl(aq) + Ba(OH)_2(aq) \rightarrow H_2O(l) + BaCl_2$

$2\ HCl(aq) + Ba(OH)_2(aq) \rightarrow 2\ H_2O(l) + BaCl_2(aq)$

2. $H_2SO_4(aq) + LiOH(aq) \rightarrow$

$(H^+ + SO_4^{2-}) + (Li^+ + OH^-) \rightarrow (H^+ + OH^-) + (Li^+ + SO_4^{2-})$

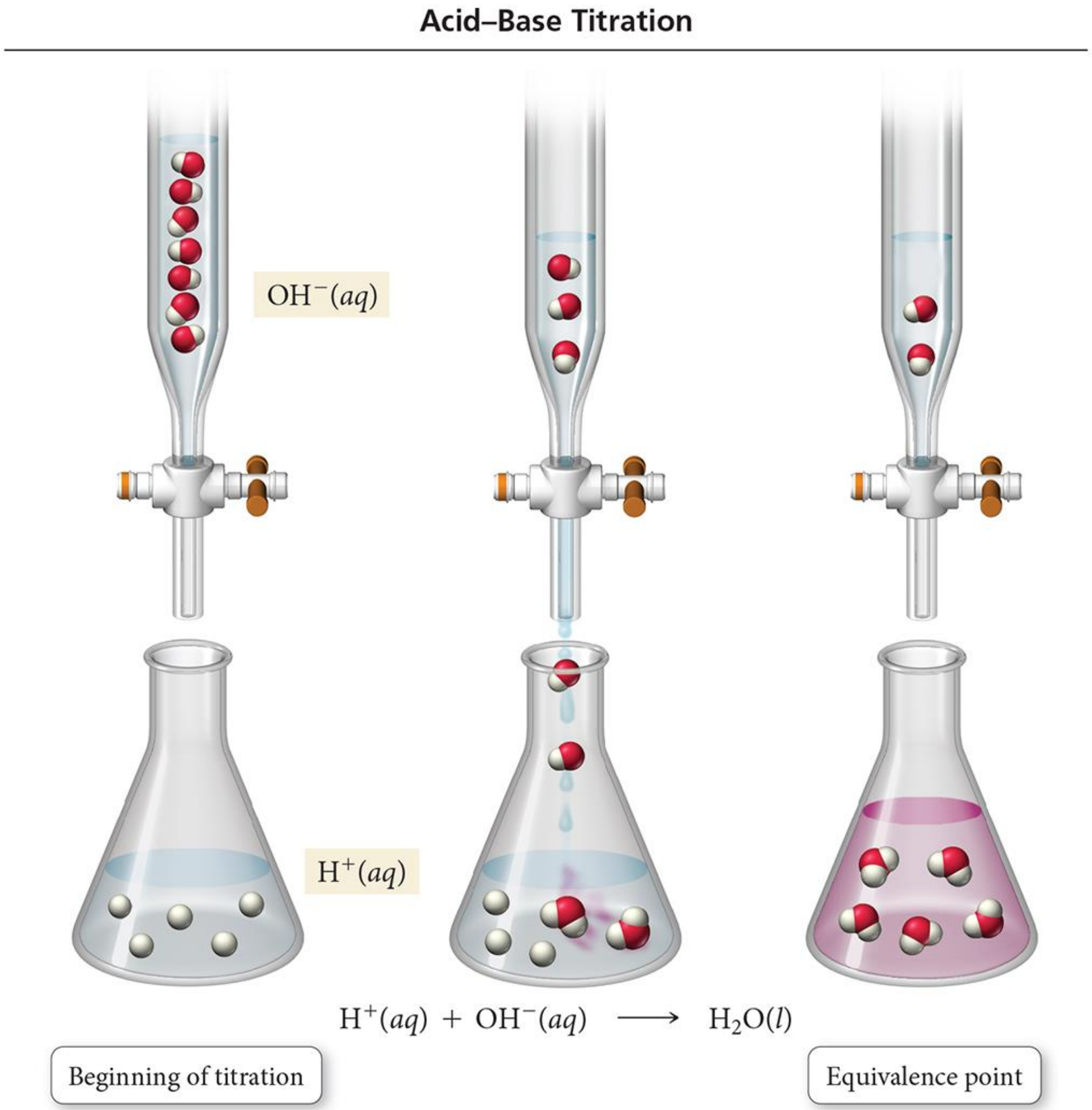
$H_2SO_4(aq) + LiOH(aq) \rightarrow H_2O(l) + Li_2SO_4$

$H_2SO_4(aq) + 2\ LiOH(aq) \rightarrow 2\ H_2O(l) + Li_2SO_4(aq)$

Acid–Base Titrations

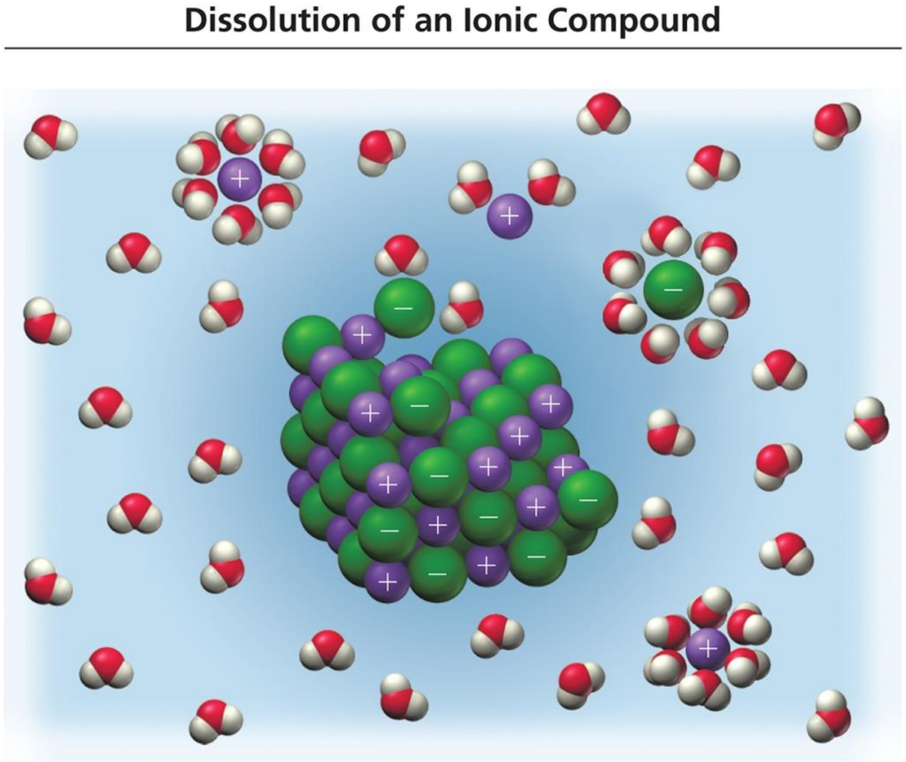
- A **titration** is a laboratory procedure where a substance in a solution of known concentration (**titration**) is reacted with another substance in a solution of unknown concentration (**analyte**).
- The **equivalence point** is the point in the titration when the H^+ and OH^- from reactants are in their stoichiometric ratio and are completely reacted.
- An **indicator** is a dye whose color depends on the acidity or basicity of solution.

Acid–Base Titration



Titration

In this titration, NaOH is added to a dilute HCl solution. When the NaOH and HCl reach stoichiometric proportions (the equivalence point), the phenolphthalein indicator changes color to pink.



Gas-Evolving Reactions

- Some reactions form a gas directly from the ion exchange.

$$K_2S(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2S(g)$$

- Other reactions form a gas by the subsequent decomposition of one of the ion exchange products into a gas and water.

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2CO_3(aq)$$

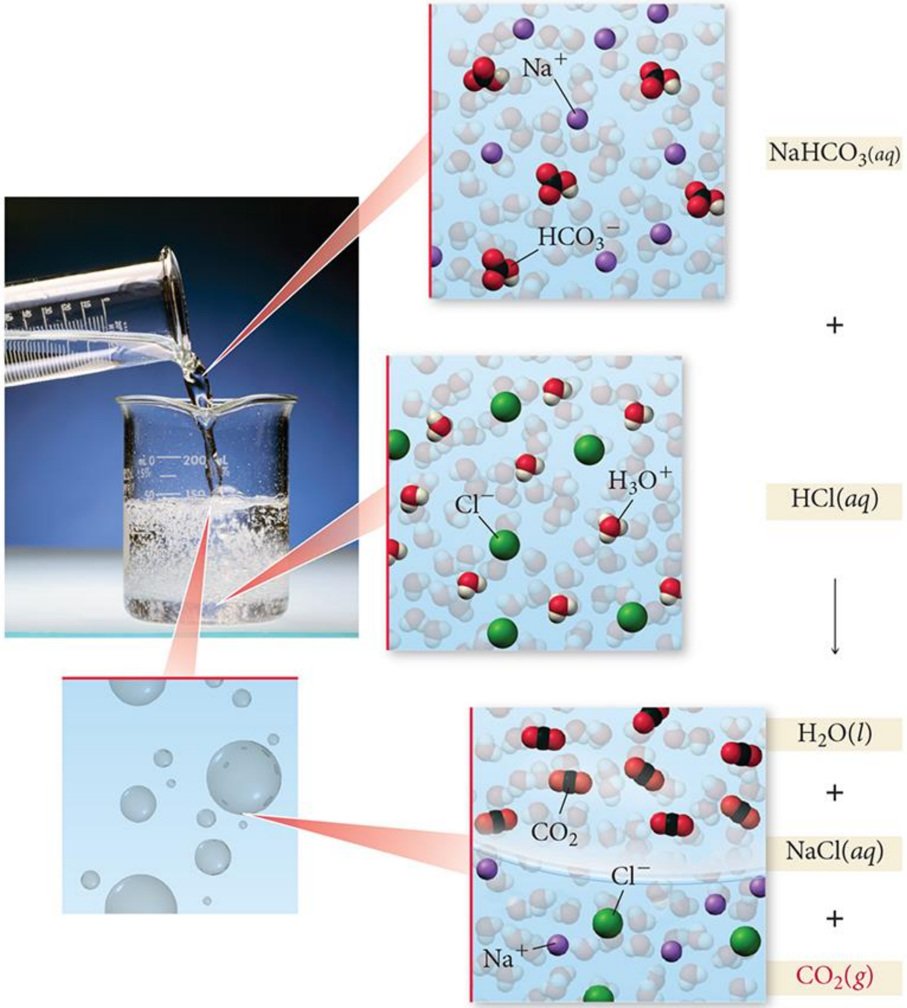
$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Gas-Evolution Reaction

Gas-Evolution Reaction

$NaHCO_3(aq) + HCl(aq) \longrightarrow H_2O(l) + NaCl(aq) + CO_2(g)$

When aqueous sodium bicarbonate is mixed with aqueous hydrochloric acid, gaseous CO_2 bubbles are the result of the reaction.



Types of Compounds That Undergo Gas-Evolution Reactions

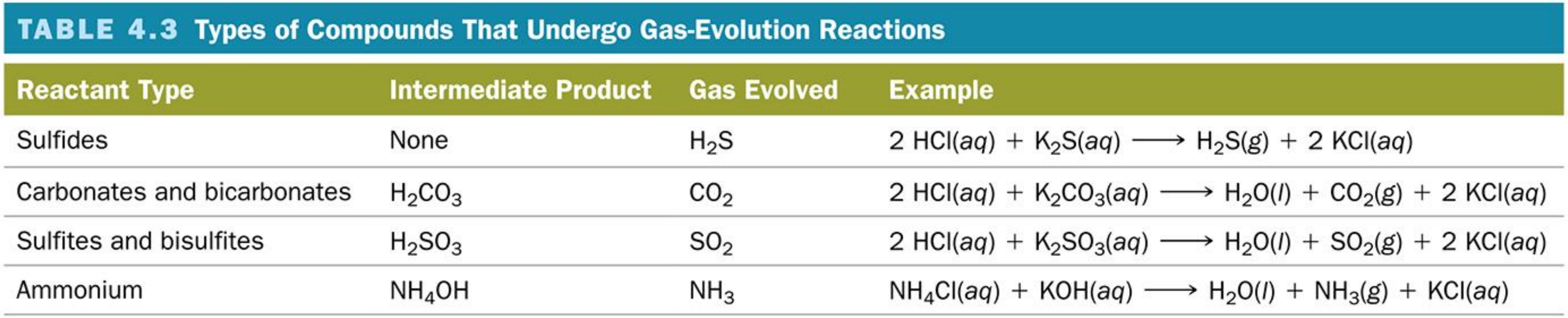


TABLE 4.3 Types of Compounds That Undergo Gas-Evolution Reactions

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	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

Oxidation–Reduction Reactions

- The reactions in which electrons are transferred from one reactant to the other are called **oxidation-reduction reactions**, or **redox reactions**.
- Many redox reactions involve the reaction of a substance with oxygen.

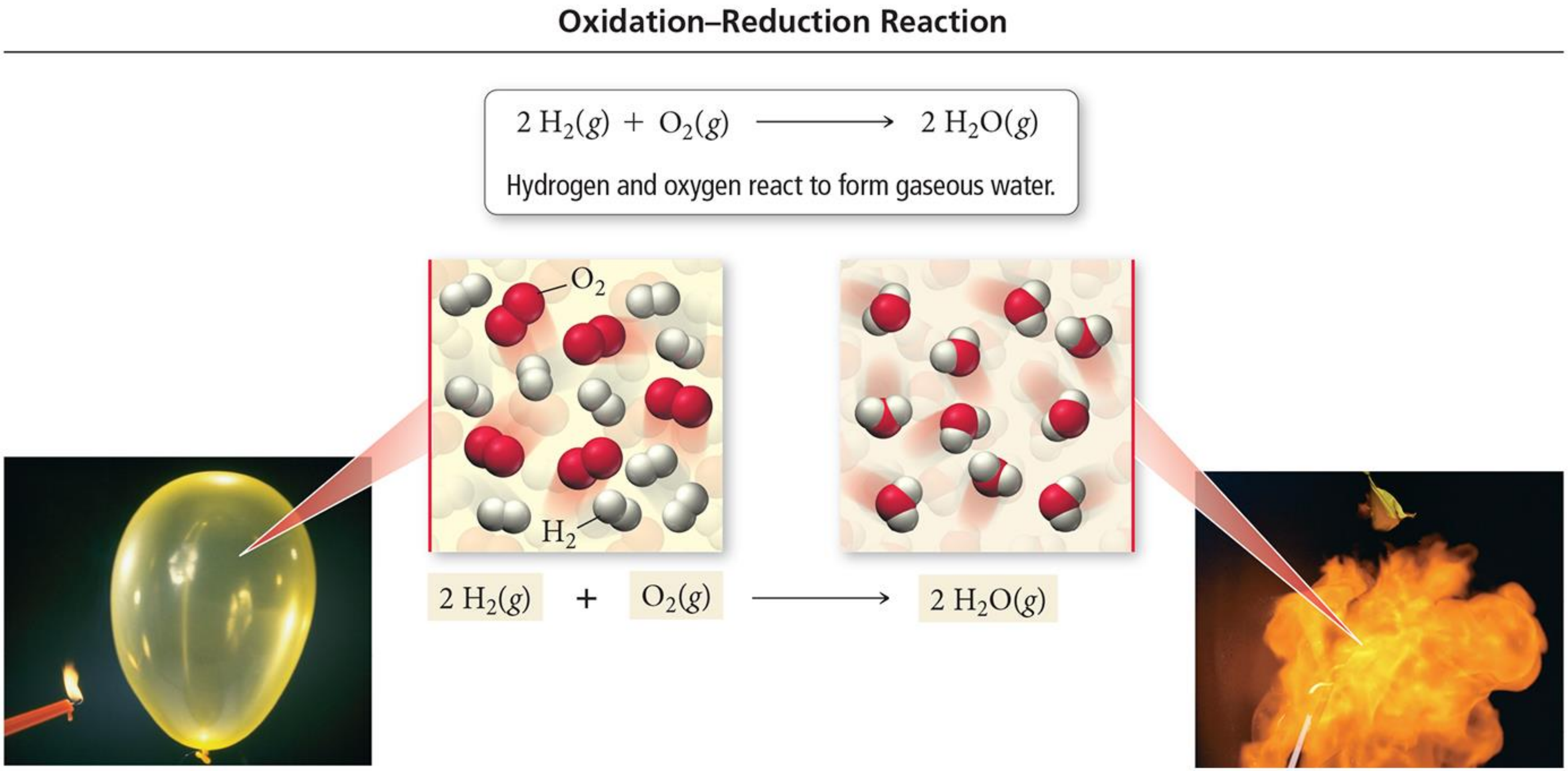
$$4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)\ \text{(rusting)}$$

$$2\ C_8H_{18}(l) + 25\ O_2(g) \longrightarrow 16\ CO_2(g) + 18\ H_2O(g)\ \text{(combustion)}$$

$$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g)$$

Combustion as Redox

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$



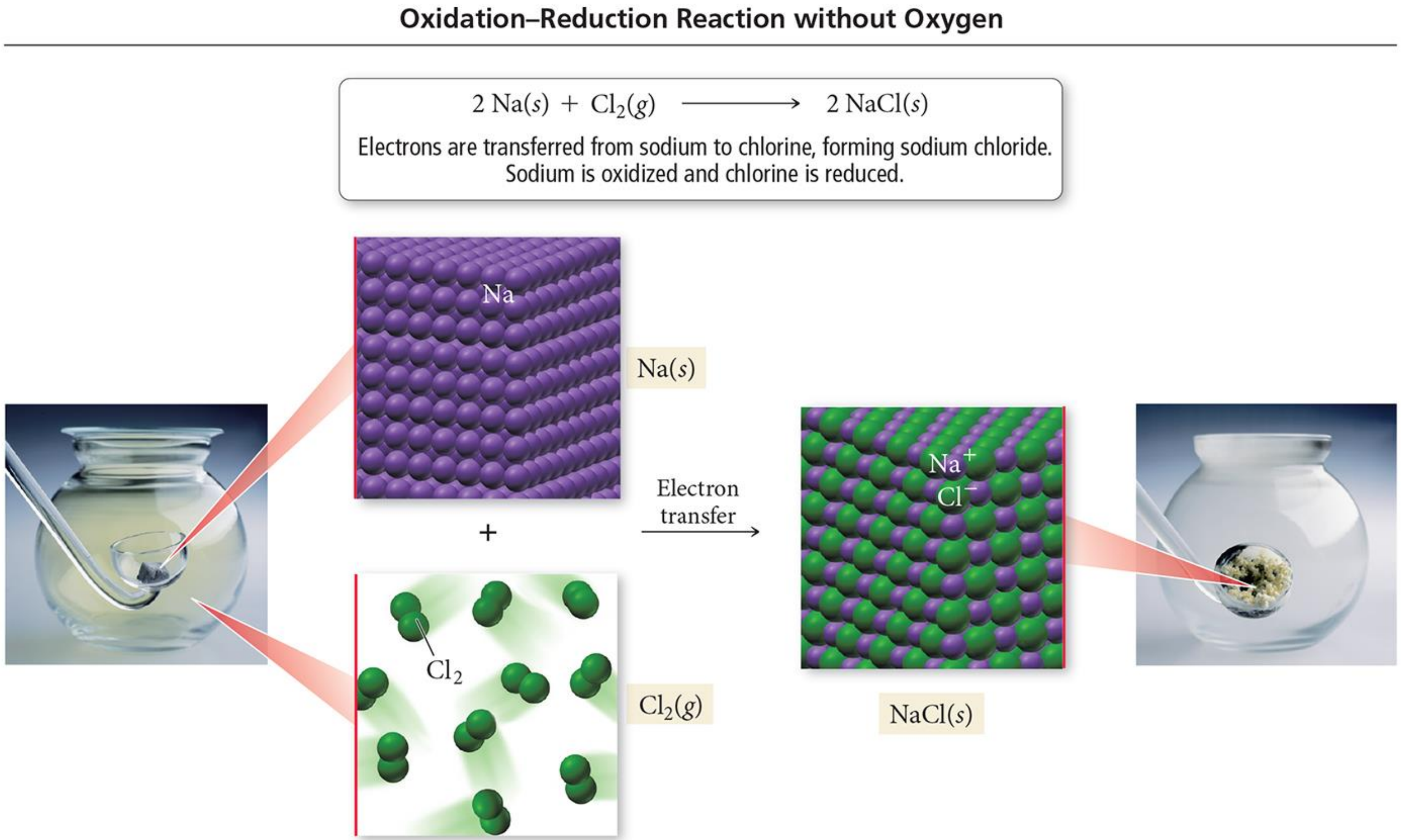
Oxidation–Reduction Reaction

$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g)$

Hydrogen and oxygen react to form gaseous water.

Redox without Combustion

$$2\ \mathrm{Na}(s) + \mathrm{Cl}_2(g) \rightarrow 2\ \mathrm{NaCl}(s)$$



Oxidation–Reduction Reaction without Oxygen

$2\ \mathrm{Na}(s) + \mathrm{Cl}_2(g) \longrightarrow 2\ \mathrm{NaCl}(s)$

Electrons are transferred from sodium to chlorine, forming sodium chloride. Sodium is oxidized and chlorine is reduced.

Reactions of Metals with Nonmetals

- Consider the following reactions:

$$4\ \mathrm{Na}(s) + \mathrm{O_2}(g) \rightarrow 2\ \mathrm{Na_2O}(s)$$

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{NaCl}(s)$$

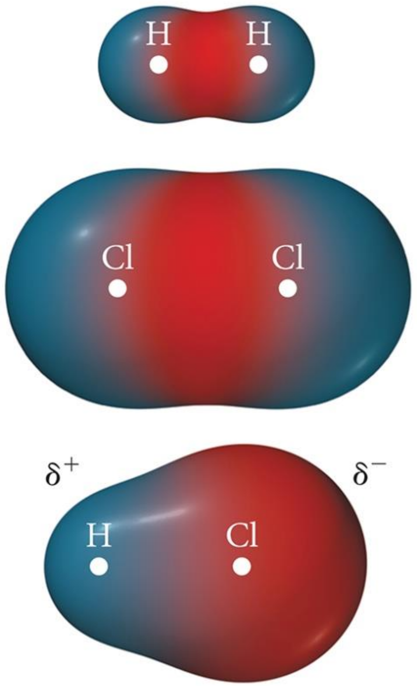
- The reactions involve a metal reacting with a nonmetal.
- In addition, both reactions involve the conversion of free elements into ions.

$$4\ \mathrm{Na}(s) + \mathrm{O_2}(g) \rightarrow 2\ \mathrm{Na^+{}_2O^{2-}}\ (s)$$

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{Na^+Cl^-}(s)$$

Redox Reaction

- Electron transfer does not need to be a *complete* transfer for the reaction to qualify as oxidation–reduction. Example: $H_2(g) + Cl_2(g)\ 2 \longrightarrow HCl(g)$
- There is uneven sharing of electrons when hydrogen bonds to chlorine, resulting in an increase of electron density (reduction) for chlorine and a decrease in electron density (oxidation) for hydrogen.



Hydrogen loses electron density (oxidation) and chlorine gains electron density (reduction).

Oxidation and Reduction

- To convert a free element into an ion, the atom must gain or lose electrons.
  - If one atom loses electrons, another atom must accept them.
- Reactions where electrons are transferred from one atom to another are redox reactions.
- Atoms that lose electrons are being oxidized, while atoms that gain electrons are being reduced.

$$2\ \mathrm{Na}(s) + \mathrm{Cl}_2(g) \rightarrow 2\ \mathrm{Na}^+\mathrm{Cl}^-(s)$$

$$\mathrm{Na} \rightarrow \mathrm{Na}^+ + 1\ e^- \quad \text{(oxidation)}$$

$$\mathrm{Cl}_2 + 2\ e^- \rightarrow 2\ \mathrm{Cl}^- \ \text{(reduction)}$$

Oxidation States

- For reactions that are not metal + nonmetal or do not involve O_2, we need a method for determining how the electrons are transferred.
- Chemists assign a number to each element in a reaction called an **oxidation state** that allows them to determine the electron flow in the reaction.
- Even though they look like them, **oxidation states are not ion charges!**
  - Oxidation states are imaginary charges assigned based on a set of rules.
  - Ion charges are real, measurable charges.

Rules for Assigning Oxidation States

The following rules are in order of priority:

1. Free elements have an oxidation state = 0.

   Na = 0 and Cl_2 = 0 in 2 Na(*s*) + Cl_2(*g*)

2. Monatomic ions have an oxidation state equal to their charge.

   In NaCl Na = +1 and Cl = −1

3. (a) The sum of the oxidation states of all the atoms in a compound is 0.

   Na = +1 and Cl = −1 in NaCl, (+1) + (−1) = 0

Rules for Assigning Oxidation States

(b) The sum of the oxidation states of all the atoms in a polyatomic ion equals the charge on the ion.

$N = +5$ and $O = -2$ in NO_3^-, $(+5) + 3(-2) = -1$

4. (a) Group I metals have an oxidation state of +1 in all of their compounds.

   $Na = +1$ in $NaCl$

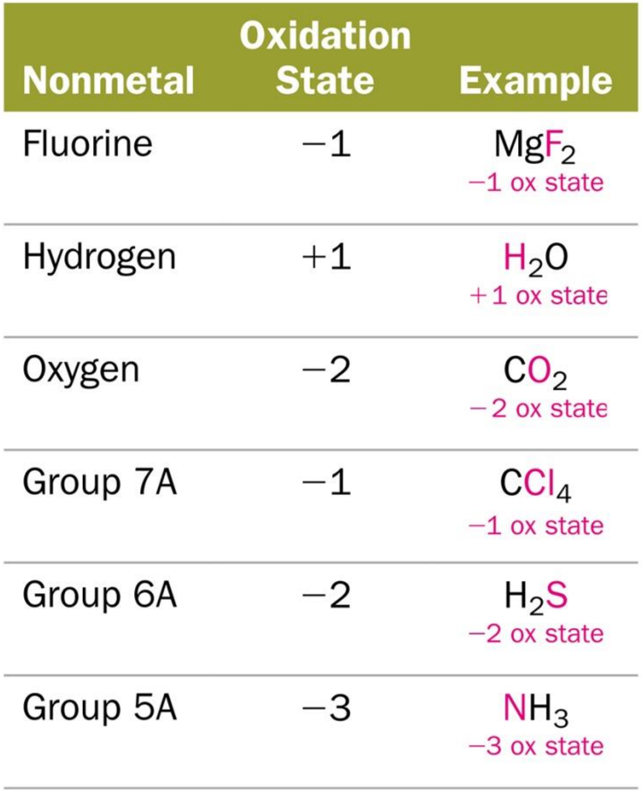
   (b) Group II metals have an oxidation state of +2 in all of their compounds.

   $Mg = +2$ in $MgCl_2$

Rules for Assigning Oxidation States

5. In their compounds, nonmetals have oxidation states according to the table below.

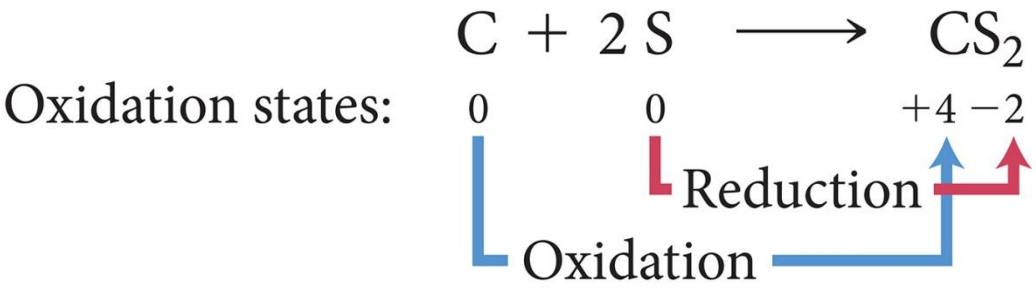
   Nonmetals higher on the table take priority.



Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 −1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 −2 ox state
Group 7A	−1	CCl_4 −1 ox state
Group 6A	−2	H_2S −2 ox state
Group 5A	−3	NH_3 −3 ox state

Identifying Redox Reactions

- Oxidation: An increase in oxidation state
- Reduction: A decrease in oxidation state



$$C + 2\,S \longrightarrow CS_2$$

  - Carbon changes from an oxidation state of 0 to an oxidation state of +4.
    - Carbon *loses electrons* and is *oxidized.*
  - Sulfur changes from an oxidation state of 0 to an oxidation state of –2.
    - Sulfur *gains electrons* and is *reduced.*

Redox Reactions

- Oxidation and reduction must occur simultaneously.
  - If an atom loses electrons another atom must take them.
- The reactant that causes reduction in another reactant is called the **reducing agent**.
  - The reducing agent contains the element that is oxidized.
- The reactant that causes oxidation in another reactant is called the **oxidizing agent**.
  - The oxidizing agent contains the element that is reduced.

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{Na^+Cl^-}(s)$$

Na is oxidized, while Cl is reduced.

Na is the reducing agent, and Cl_2 is the oxidizing agent.

Combustion Reactions

- **Combustion reactions** are characterized by the reaction of a substance with O_2 to form one or more oxygen-containing compounds, often including water.
  - – Combustion reactions also emit heat.
- For example, as you saw earlier in this chapter, natural gas (CH_4) reacts with oxygen to form carbon dioxide and water:

$$\begin{array}{lccccccccc} & \mathrm{CH_4}(g) & + & 2\,\mathrm{O_2}(g) & \rightarrow & \mathrm{CO_2}(g) & + & 2\,\mathrm{H_2O}(g) \\ \text{Oxidation state:} & -4\ \ +1 & & 0 & & +4\ -2 & & +1\ \ -2 \end{array}$$

Combustion

- Ethanol, the alcohol in alcoholic beverages, also reacts with oxygen in a combustion reaction to form carbon dioxide and water.

$$C_2H_5OH(l) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 3\,H_2O(g)$$