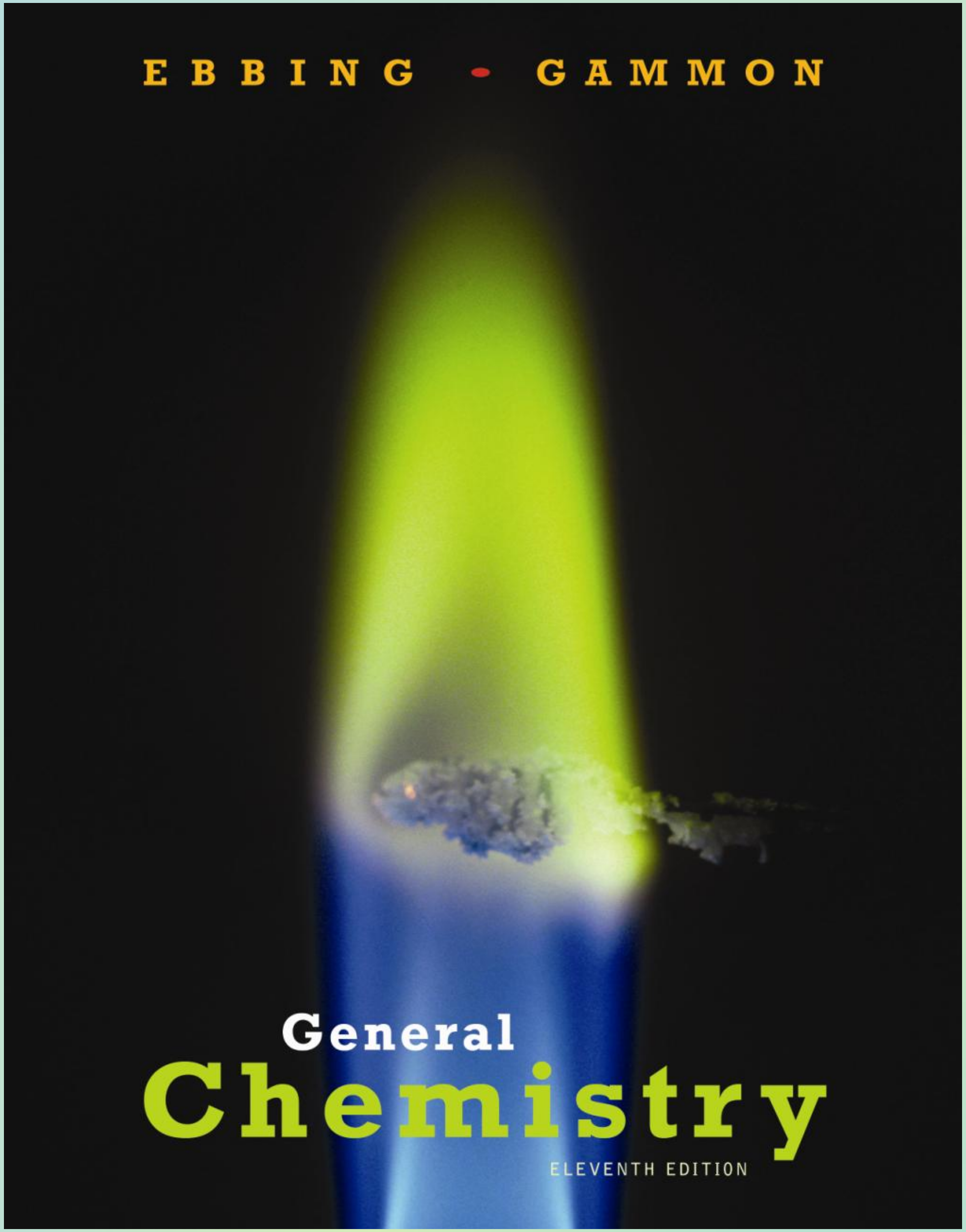


Chapter 4 Chemical Reactions

Contents and Concepts

Ions in Aqueous Solution

Explore how molecular and ionic substances behave when they dissolve in water to form solutions.

1. Ionic Theory of Solutions and Solubility Rules
2. Molecular and Ionic Equations

Types of Chemical Reactions

Investigate several important types of reactions that typically occur in aqueous solution: precipitation reactions, acid–base reactions, and oxidation–reduction reactions.

3. Precipitation Reactions
4. Acid–Base Reactions
5. Oxidation–Reduction Reactions
6. Balancing Simple Oxidation–Reduction Equations

Working with Solutions

Now that we have looked at how substances behave in solution, it is time to quantitatively describe these solutions using concentration.

7. Molar Concentration
8. Diluting Solutions

Quantitative Analysis

Using chemical reactions in aqueous solution, determine the amount of substance or species present in materials.

9. Gravimetric Analysis
10. Volumetric Analysis

Learning Objectives

Ions in Aqueous Solution

1. Ionic Theory of Solutions and Solubility Rules
   a. Describe how an ionic substance can form ions in aqueous solution.
   b. Explain how an electrolyte makes a solution electrically conductive.
   c. Give examples of substances that are electrolytes.
   d. Define *nonelectrolyte* and provide an example of a molecular substance that is a nonelectrolyte.

e. Compare the properties of solutions that contain *strong electrolytes* and *weak electrolytes*.
f. Learn the *solubility rules* for ionic compounds.
g. Use the solubility rules.

2. Molecular and Ionic Equations
    a. Write the *molecular equation* of a chemical reaction.
    b. From the molecular equations for both strong electrolytes and weak electrolytes, determine the *complete ionic equation*.
    c. From the complete ionic equation, write the *net ionic equation*.
    d. Write net ionic equations.

Types of Chemical Reactions

3. Precipitation Reactions
    a. Recognize *precipitation (exchange)* reactions.
    b. Write molecular, complete ionic, and net ionic equations for precipitation reactions.
    c. Decide whether a precipitation reaction will occur.
    d. Determine the product of a precipitation reaction.

4. Acid–Base Reactions
   a. Understand how an acid–base indicator is used to determine whether a solution is acidic or basic.
   b. Define *Arrhenius acid* and *Arrhenius base.*
   c. Write the chemical equation of an Arrhenius base in aqueous solution.
   d. Define *Brønsted–Lowry acid* and *Brønsted–Lowry base.*
   e. Write the chemical equation of a Brønsted–Lowry base in aqueous solution
   f. Write the chemical equation of an acid in aqueous solution using the *hydronium ion*.
   g. Learn the common *strong acids* and *strong bases*.

h. Distinguish between a strong acid and a *weak acid* and the solutions they form.
i. Distinguish between a strong base and a *weak base* and the solutions they form.
j. Classify acids and bases as strong or weak.
k. Recognize *neutralization reactions*.
l. Write an equation for a neutralization reaction.
m. Write the reaction equations for a *polyprotic acid* in aqueous solution.
n. Recognize acid–base reactions that lead to gas formation.
o. Write an equation for a reaction with gas.

6. Balancing Simple Oxidation–Reduction Equations
   a. Balance simple oxidation–reduction reactions by the half−reaction method.

Working with Solutions

7. Molar Concentration
   a. Define *molarity* or *molar concentration* of a solution.
   b. Calculate molarity from mass and volume.
   c. Use molarity as a conversion factor.
8. Diluting Solutions
   a. Describe what happens to the concentration of a solution when it is diluted.
   b. Perform the calculations associated with dilution.
   c. Describe the process for diluting a solution.

Quantitative Analysis

9. Gravimetric Analysis
    a. Determine the amount of a species by *gravimetric analysis*.
10. Volumetric Analysis
    a. Calculate the volume of reactant solution needed to perform a reaction.
    b. Understand how to perform a *titration*.
    c. Calculate the quantity of substance in a titrated solution.

Ionic theory of solutions

Arrhenious proposed the ionic theory of solutions to account for the electrical conductivity of water.

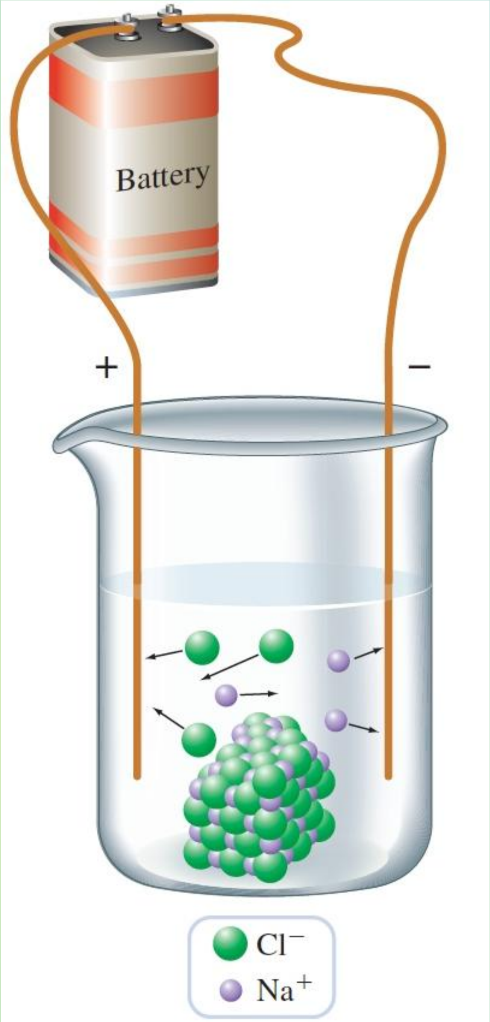
The theory proposed that certain substances produce freely moving ions when they dissolve in water, and these ions conduct electric current in an aqueous solution.

Electrolyte

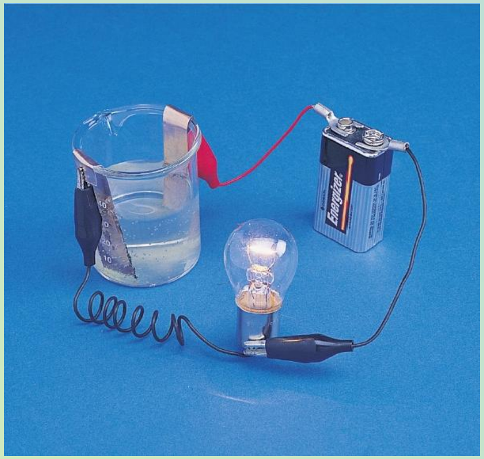
It dissolves in water to give an electrically conducting solution.

Most ionic solids dissolve in water. Ions in previously fixed states are free to move about, forming an electric current.

A nonelectrolyte dissolves in water, producing a poorly conducted solution.



Solutions that are poor conductors have less or no ions.



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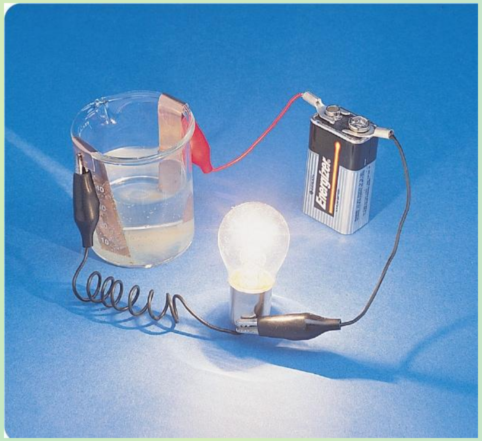
Solutions with good conductivity have plenty of ions.



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A **strong electrolyte** exists in solution almost entirely as ions. This includes almost all soluble ionic substances.

A **weak electrolyte** dissolves in water, yielding a relatively small percentage of ions. Most molecular substances are weak electrolytes.

Table 4.1 Solubility Rules for Ionic Compounds

Rule	Applies to	Statement	Exceptions
1	Li^+, Na^+, K^+, NH_4^+	Group 1A and ammonium compounds are soluble.	—
2	$C_2H_3O_2^-$, NO_3^-	Acetates and nitrates are soluble.	—
3	Cl^-, Br^-, I^-	Most chlorides, bromides, and iodides are soluble.	$AgCl$, Hg_2Cl_2, $PbCl_2$, $AgBr$, $HgBr_2$, Hg_2Br_2, $PbBr_2$, AgI, HgI_2, Hg_2I_2, PbI_2
4	SO_4^{2-}	Most sulfates are soluble.	$CaSO_4$, $SrSO_4$, $BaSO_4$, Ag_2SO_4, Hg_2SO_4, $PbSO_4$
5	CO_3^{2-}	Most carbonates are insoluble.	Group 1A carbonates, $(NH_4)_2CO_3$
6	PO_4^{3-}	Most phosphates are insoluble.	Group 1A phosphates, $(NH_4)_3PO_4$
7	S^{2-}	Most sulfides are insoluble.	Group 1A sulfides, $(NH_4)_2S$
8	OH^-	Most hydroxides are insoluble.	Group 1A hydroxides, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

Determine whether the following compounds are soluble or insoluble in water.

a. Hg_2Cl_2

b. KI

a. Hg_2Cl_2 is an exception to Rule 3 of the solubility rules. Therefore, it is not soluble in water.
b. According to Rules 1 and 3, most iodides are soluble. Therefore, it is soluble in water.

Molecular Equations

They are chemical equations in which the reactants and products are written as if they were molecular substances, even though they may actually exist in solution as ions.

For example

$$Ca\ OH_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaOH(aq)$$

It indicates the states of the reactants and products.

Ionic equations

In a **complete ionic equation**, strong electrolytes are written as separate ions in the solution. Other reactants and products are written in molecular form.

State symbols are included.

Consider the reactants:

$$Ca^{2+}(aq) + 2OH^{-}(aq) + 2Na^{+}(aq) + CO_3^{2-}(aq) \longrightarrow$$

With the products, the ionic form is:

$$Ca^{2+}(aq) + 2OH^{-}(aq) + 2Na^{+}(aq) + CO_3^{2-}(aq) \longrightarrow$$

$$CaCO_3\ s \ + \ 2Na^{+}(aq) \ + \ 2OH^{-}(aq)$$

Net Ionic equation

It is an ionic equation from which spectator ions have been omitted.

For example:

$$\boxed{Ca^{2+}(aq)} + \cancel{2OH^{-}(aq)} + \cancel{2Na^{+}(aq)} + \boxed{CO_3^{2-}(aq)} \longrightarrow \boxed{CaCO_3(s)} + \cancel{2Na^{+}(aq)} + \cancel{2OH^{-}(aq)}$$

– The resulting equation is

$$\boxed{Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)}$$

Write a net ionic equation for the following molecular equation:

$2HClO_4(aq) + Ca(OH)_2(aq) \rightarrow Ca(ClO_4)_2(aq) + 2H_2O(l)$

Ionic equation

$\boxed{2H^+(aq)} + \cancel{2ClO_4^-(aq)} + \cancel{Ca^{2+}(aq)} + \boxed{2OH^-(aq)} \longrightarrow \cancel{Ca^{2+}(aq)} + \cancel{2ClO_4^-(aq)} + \boxed{2H_2O(l)}$

Net ionic equation

$\boxed{H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)}$

Precipitation reactions

A solid ionic substance forms from the mixture of two solutions of ionic substances.

Acid–base reactions

They are the reactions that involve the transfer of a proton between reactants.

Oxidation–reduction reactions

They are the reactions that involve the transfer of electrons between reactants.

Precipitation Reactions

They occur in aqueous solution.

A **precipitate** is an insoluble solid compound formed during a chemical reaction in solution.

A precipitation reaction, when written as a molecular reaction, has the form of **an exchange reaction**.

An exchange reaction, when written as a molecular equation, appears to involve the exchange of parts between the two reactants.

Predicting Precipitation Reactions

1. Predict the products (exchange of parts).
2. Determine the state of each product: (*s*), (*l*), (*g*), and (*aq*).
3. If all products are aqueous (*aq*), no net reaction will occur.

Determine if a precipitation reaction occurs between aqueous solutions of ammonium sulfate and sodium hydroxide. If yes, write the balanced molecular equation and the net ionic equation.

Balanced equation

$$\boxed{Al_2(SO_4)_3(aq)} + \boxed{6NaOH(aq)} \longrightarrow \boxed{2Al(OH)_3(s)} + \boxed{3Na_2SO_4(aq)}$$

Ionic equation

$$\boxed{2Al^{3+}(aq)} \boxed{+ \cancel{3SO_4^{2-}(aq)}} + \boxed{\cancel{6Na^+(aq)} +} \boxed{6OH^-(aq)} \longrightarrow \boxed{2Al(OH)_3(s)} + \boxed{\cancel{6Na^+(aq)} + \cancel{3SO_4^{2-}(aq)}}$$

Net ionic equation

$$\boxed{Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)}$$

Determining whether a solution is an acid or a base

Acid solutions have a sour taste.

Basic solutions are bitter to taste and soapy to touch.

A dye is used to distinguish between acid and base by the different color changes it undergoes in these solutions.

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An **acid** produces hydrogen ions when dissolved in water.

$$HNO_3(aq) \xrightarrow{H_2O} H^+(aq) + NO_3^-(aq)$$

A **base** produces hydroxide ions when dissolved in water.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Definitions of acid and base according to Brønsted and Lowry

An acid is the species (molecule or ion) that donates a proton to another species in a proton−transfer reaction.

A base is the species (molecule or ion) that accepts a proton from another species in a proton−transfer reaction.

Strong Acid

It ionizes completely in water.

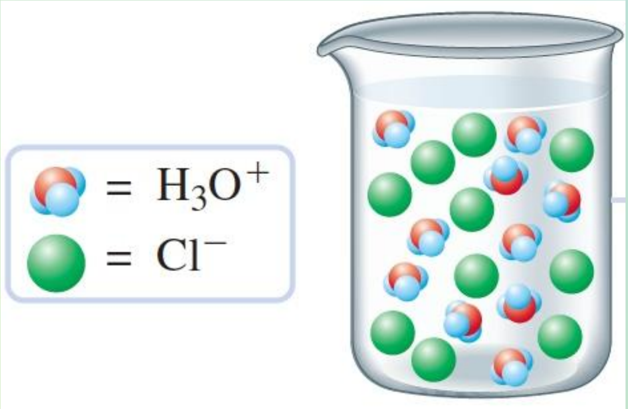
Examples - Hydrochloric acid and nitric acid

Weak Acid

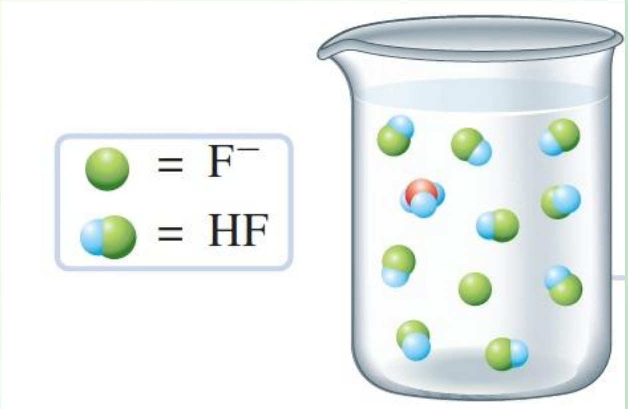
It partly ionizes in water.

Examples - Hydrocyanic acid and hydrofluoric acid

In Figure A, a solution of HCl (a strong acid) illustrated on a molecular/ionic level shows the acid in its ionic form.



In Figure B, a solution of HF (a weak acid), also illustrated on a molecular/ionic level, shows mostly molecules with very few ions.



Strong Base

It is present in an aqueous solution entirely as ions, one of which is OH^-. It is a strong electrolyte, including groups 1A and 2A elements and excluding beryllium hydroxide.

Weak Base

It undergoes partial ionization. It is a weak electrolyte.

Example - Ammonia

Identify each of the following as a strong acid or a weak base:

a. LiOH

b. $HC_2H_3O_2$

c. HBr

d. HNO_2

a. LiOH is a strong base.

b. $HC_2H_3O_2$ is a weak acid.

c. HBr is a strong acid.

d. HNO_2 is a weak acid.

Neutralization Reaction

A reaction of an acid and a base that results in an ionic compound and possibly water.

The ionic compound that is a product of a neutralization reaction is called a **salt**.

$$2HCl(aq) + Ca(OH)_2(aq) \longrightarrow CaCl_2(aq) + 2H_2O(l)$$

Write the molecular and net ionic equations for the neutralization of HNO_2(*aq*) by NaOH(*aq*).

Molecular Equation

(Balance the reaction and include state symbols)

$$\boxed{HNO_2(aq)} + \boxed{NaOH(aq)} \longrightarrow \boxed{NaNO_2(aq)} + \boxed{H_2O(l)}$$

Complete Ionic Equation

$$\boxed{HNO_2(aq)} + \boxed{\cancel{Na^+(aq)} + OH^-(aq)} \longrightarrow \boxed{\cancel{Na^+(aq)} + NO_2^-(aq)} + \boxed{H_2O(l)}$$

Net Ionic Equation

$$\boxed{HNO_2(aq) + OH^-(aq) \longrightarrow NO_2^-(aq) + H_2O(l)}$$

Polyprotic Acid

It is an acid that results in two or more acidic hydrogens per molecule.

For example:

$$\underset{\text{Carbonate}}{Na_2CO_3\,(aq)} + \underset{\text{Acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Salt}}{2NaCl(aq)} + H_2O(l) + CO_2(g)$$

Acid−base reactions with gas formation

Sulfides, carbonates, and sulfites react with acid to form a gas.

$$\underset{\text{Carbonate}}{Na_2CO_3}(aq) + \underset{\text{Acid}}{2HCl}(aq) \longrightarrow$$

$$\underset{\text{Salt}}{2NaCl}(aq) + H_2O(l) + CO_2(g)$$

Interchanging the carbonate ion with the chloride ion gives the following equation:

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2CO_3(aq)$$


Food Club
VINEGAR
ARM & HAMMER
Baking Soda

Write the molecular, ionic, and net ionic equations for the reaction of zinc sulfide with hydrochloric acid.

Molecular Equation

(Balance the reaction and include state symbols.)

$$\boxed{ZnS(s)} + \boxed{2HCl(aq)} \longrightarrow \boxed{ZnCl_2(aq)} + \boxed{H_2S\ g}$$

Ionic Equation

$$\boxed{ZnS(s)} + \boxed{2H^+(aq)} + \boxed{\cancel{2Cl^-(aq)}} \longrightarrow$$

$$\boxed{Zn^{2+}(aq)} + \boxed{\cancel{2Cl^-(aq)}} + \boxed{H_2S(g)}$$

Net Ionic Equation

$$\boxed{ZnS(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2S(g)}$$

Oxidation Number

This is the actual charge of the atom if it exists as a monoatomic atom, or a hypothetical charge assigned to the atom in the substance using simple rules.

Rules for Assigning Oxidation Numbers

1. Elements - The oxidation number of an atom in an element is zero.
2. Monoatomic ions - The oxidation number of an atom in a monatomic ion equals the charge on the ion.
3. Oxygen - The oxidation number of oxygen is −2 in most of its compounds. (An exception is O in H_2O_2 and other peroxides, where the oxidation number is −1.)
4. Hydrogen - The oxidation number of hydrogen is +1 in most of its compounds. (The oxidation number of hydrogen is −1 in binary compounds with a metal such as CaH_2.)

5. Halogens - The oxidation number of fluorine is −1. Each of the other halogens (Cl, Br, I) has an oxidation number of −1 in binary compounds, except when the other element is another halogen above it in the periodic table or the other element is oxygen.
6. Compounds and ions - The sum of the oxidation numbers of the atoms in a compound is zero. The sum of the oxidation numbers of the atoms in a polyatomic ion equals the charge on the ion.

Obtain the oxidation number of the chlorine atom in ClO_3^-.

Using Rule 6:

(Oxidation number of Cl) +3 ×(oxidation number of O) = –1

Using rule 3:

(Oxidation number of Cl) +3 ×(–2) = –1

Oxidation number of Cl (in ClO_3^-) = –1 –3 ×(–2) = +5

$\boxed{\text{In } ClO_3^-\text{, the oxidation number of Cl is 5.}}$

Half-reaction

It is one of two parts of an oxidation–reduction reaction, one part of which involves a loss of electrons (or increase in oxidation number) and the other part of which involves a gain of electrons (or decrease in oxidation number).

Oxidation

It is the half-reaction in which there is a loss of electrons by a species (or an increase of oxidation number).

Reduction

It is the half-reaction in which there is a gain of electrons by a species (or a decrease of oxidation number).

Oxidizing Agent

It is a species that oxidizes another species; it is itself reduced.

Reducing Agent

It is a species that reduces another species; it is itself oxidized.

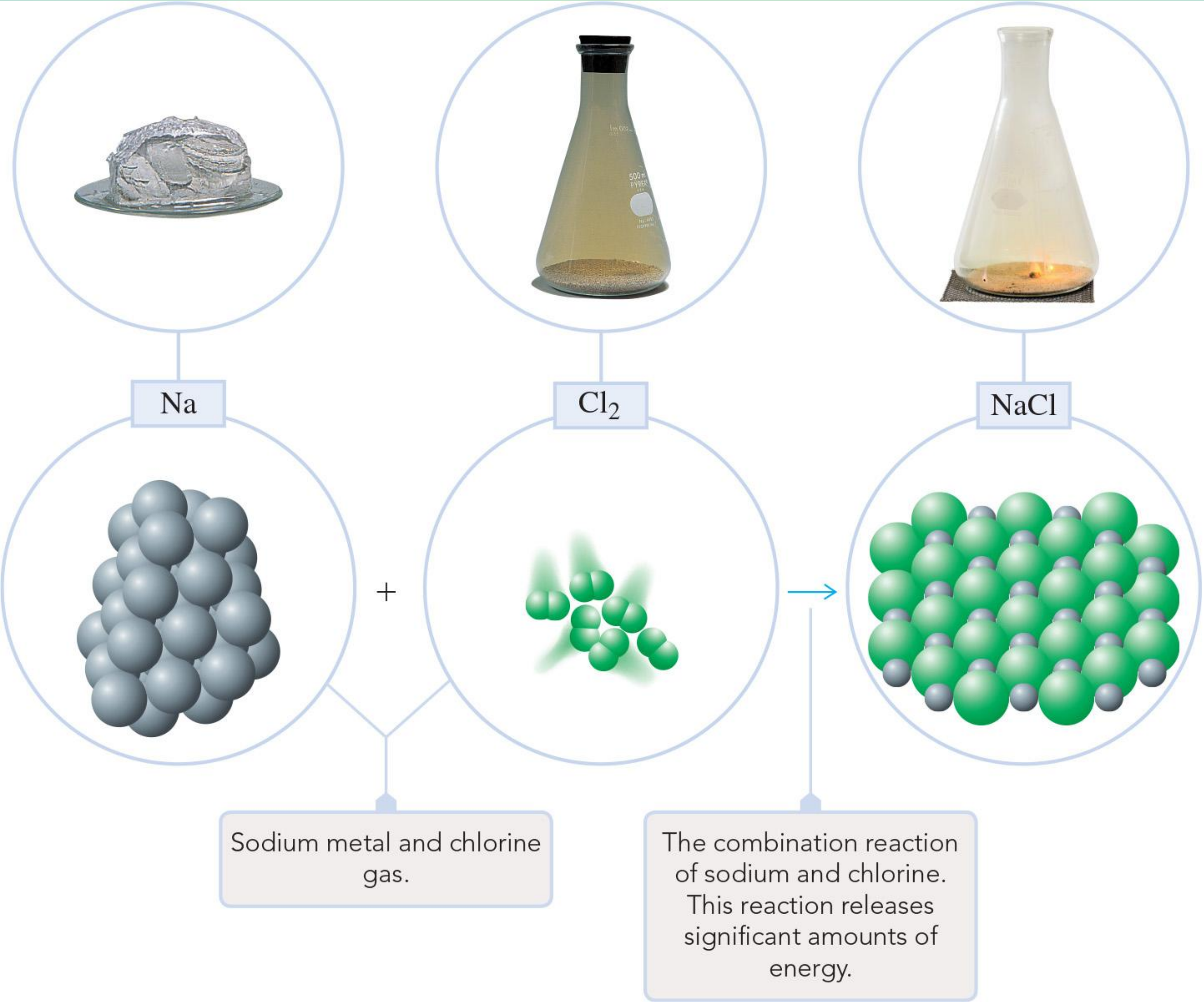
Common Oxidation–Reduction Reactions

- Combination reaction
- Decomposition reaction
- Displacement reaction
- Combustion reaction

Combination Reaction

A reaction in which two substances combine to form a third substance.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$



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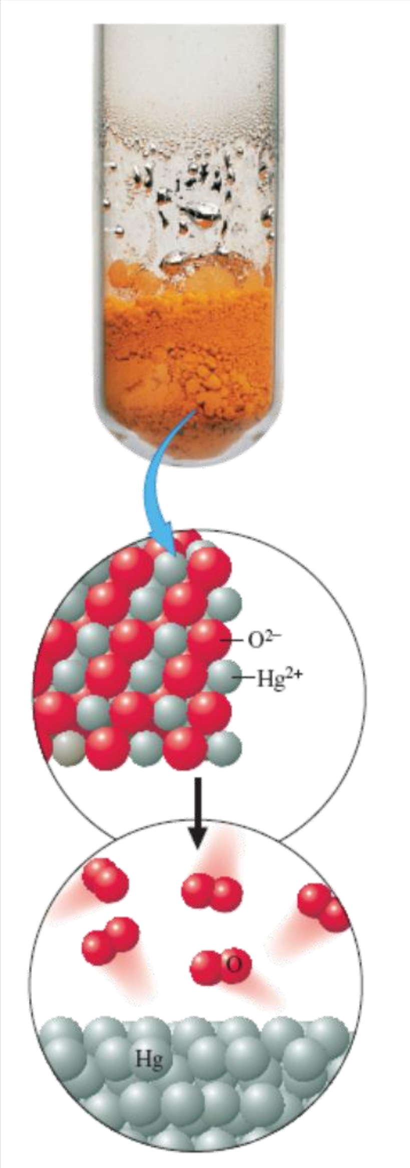
Decomposition Reaction

It is a reaction in which a single compound reacts to give two or more substances.

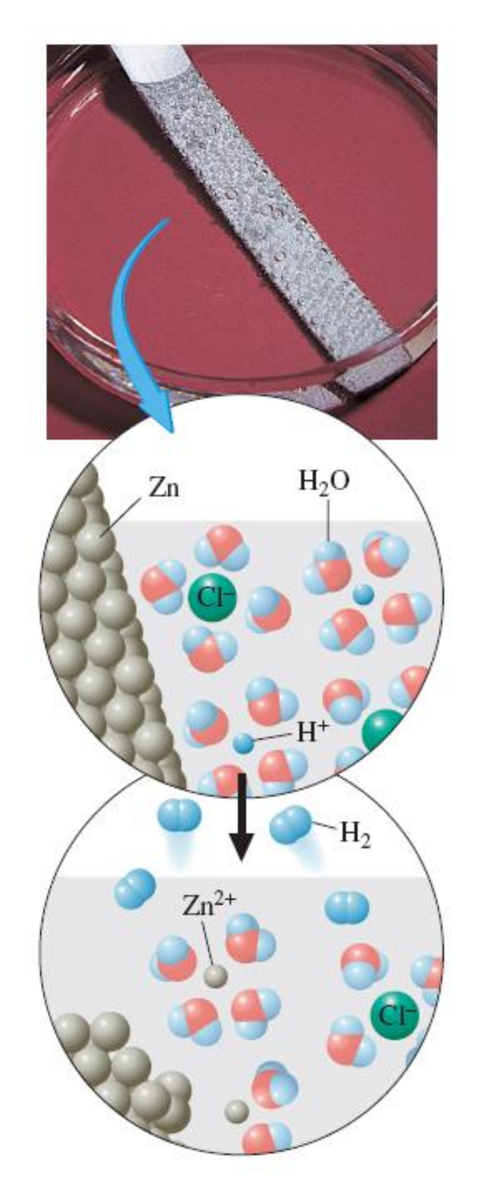
$$2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$$


O2–
Hg2+
O
Hg

Displacement Reaction

It is a reaction in which an element reacts with a compound, displacing another element from it.

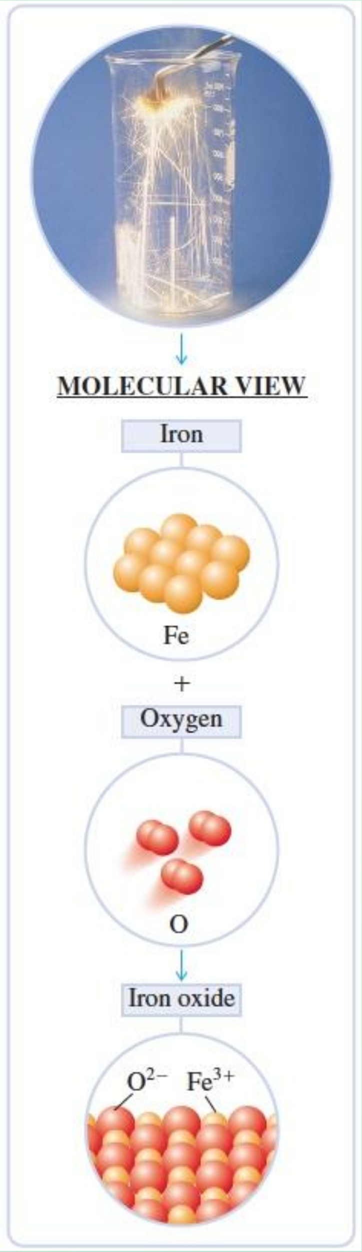
$$Zn(\mathit{s}) + 2HCl(\mathit{aq}) \longrightarrow ZnCl_2(\mathit{aq}) + H_2(g)$$


Zn
H_2O
Cl^-
H^+
H_2
Zn^{2+}
Cl^-

Combustion Reaction

It is a reaction in which a substance reacts with oxygen, usually with the rapid release of heat to produce a flame.

$$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$$


MOLECULAR VIEW
Iron
Fe
+
Oxygen
O
Iron oxide
O^{2-}
Fe^{3+}

Balancing Simple Oxidation−Reduction Reactions: Half-Reaction Method

First, identify what is oxidized and what is reduced by determining the oxidation numbers.

Balance each half-reaction and combine them to obtain a balanced oxidation reaction.

For the reaction

$$\underset{0}{Zn(s)} + \underset{+1}{Ag^{+}(aq)} \longrightarrow \underset{+2}{Zn^{2+}(aq)} + \underset{0}{Ag(s)}$$

Zn is oxidized, resulting in Zn^{2}.

Ag^{+} is reduced, resulting in Ag.

Next, balance the charge in each half-reaction by adding electrons to the more positive side.

$$Zn \longrightarrow Zn^{2+}(aq) + 2e^{-} \quad \text{oxidation half-reaction}$$

$$Ag^{+}(aq) + e^{-} \longrightarrow Ag \quad \text{reduction half-reaction}$$

Zinc loses more electrons during oxidation than silver gains during reduction. Since Ag can gain only one electron, the amount of Ag must be doubled for it to accept all the electrons produced by Zn during oxidation. This can be accomplished by multiplying each half-reaction by a factor.

$$1 \times (\mathrm{Zn} \longrightarrow \mathrm{Zn}^{2+} + 2e^{-})$$
$$2 \times (\mathrm{Ag}^{+} + e^{-} \longrightarrow \mathrm{Ag})$$

The electrons cancel, resulting in the balanced oxidation reaction

$$\mathrm{Zn}(s) + 2\mathrm{Ag}^{+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + 2\mathrm{Ag}(s)$$

Use the half-reaction method and balance the following reaction:

$$\mathrm{Mg}(s) + \mathrm{N_2}(g) \longrightarrow \mathrm{Mg_3N_2}(s)$$

Identifying the oxidation states of the elements

$$\overset{0}{\mathrm{Mg}}(s) + \overset{0}{\mathrm{N2}}(g) \longrightarrow \overset{+2}{\mathrm{Mg_3}}\overset{-3}{\mathrm{N_2}}(s)$$

Balancing oxidation and reduction reactions

$\mathrm{Mg} \longrightarrow \mathrm{Mg^{2+}} + 2e^{-}$ balanced oxidation half-reaction

$\mathrm{N_2} + 6e^{-} \longrightarrow 2\mathrm{N^{3-}}$ balanced oxidation half-reaction

Multiplying each half-reaction by a factor that will cancel the electrons:

$$3 \times (Mg \longrightarrow Mg^{2+}(aq) + 2e^-)$$
$$1 \times (N_2 + 6e^- \longrightarrow 2N^{3-})$$

This gives:

$$3Mg + N_2 + \cancel{6e^-} \longrightarrow 3Mg^{2+} + 2N^{3-} + \cancel{6e^-}$$

Therefore, the balanced combination (oxidation-reduction) reaction is:

$$3Mg + N_2 \longrightarrow 3Mg^{2+} + 2N^{3-}$$

This can also be written:

$$3Mg\ (s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

Molar Concentration, Molarity, (*M*)

Moles of solute per liter of solution

$$\text{Molarity}\ (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

To prepare a solution, add the measured amount of solute to a volumetric flask and then add water to bring the solution to the mark on the flask.

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Diluting Solutions

Attaining a solution with specific aqueous content requires diluting it with a predetermined amount of water. This requires knowledge of the relationship between the molarity of the solution before dilution, and that after dilution.

It is known that:

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

This can be rearranged to give

$$\text{moles of solute} = \text{molarity} \times \text{liters of solution}$$

This can be rearranged to give:

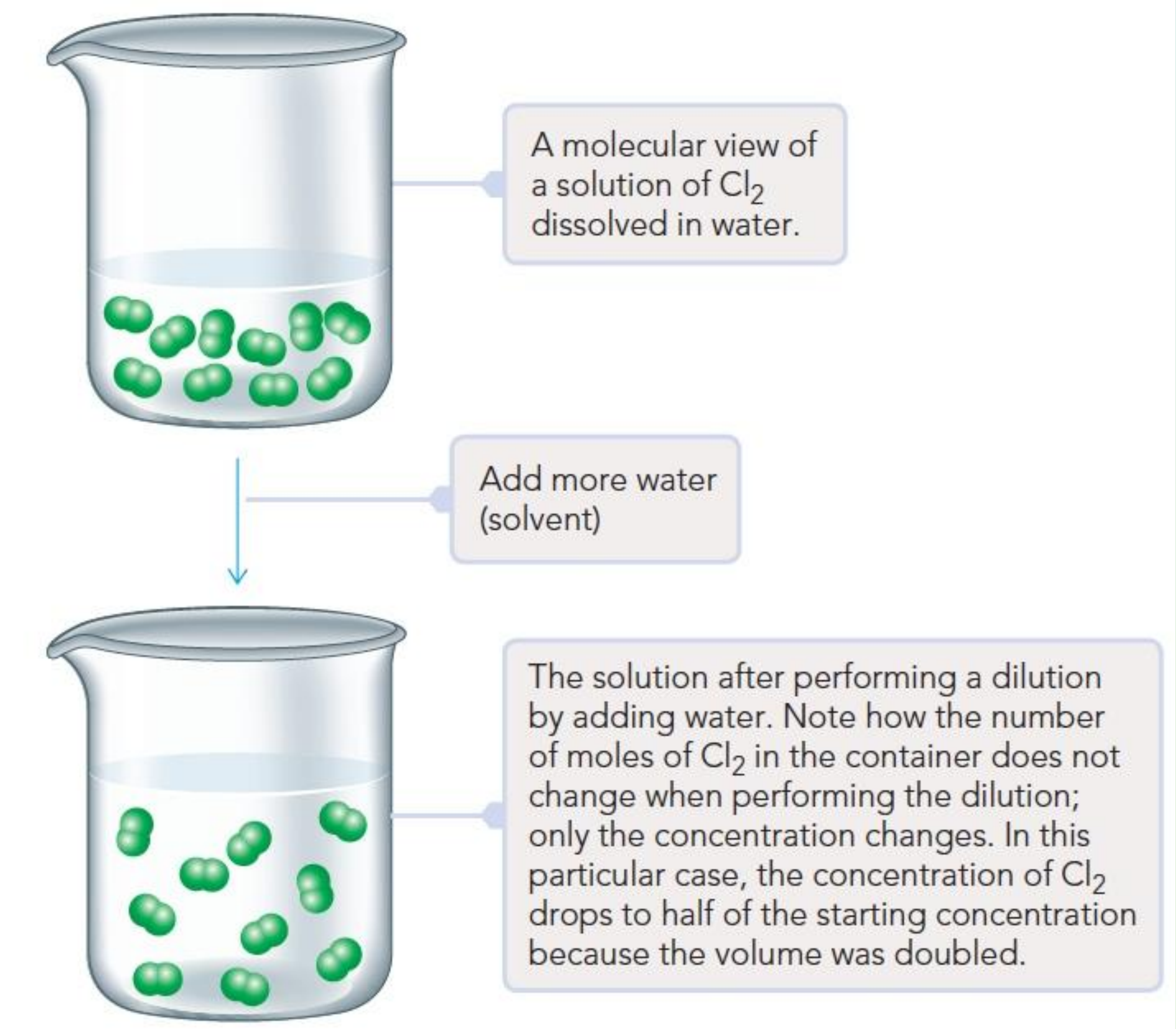
$$\text{Moles of solute} = M_i \times V_i$$

Upon dilution, the concentration and volume changes.

$$\text{Moles of solute} = M_f \times V_f$$

As the moles of solute have not changed during the dilution, it can be observed that:

$$M_i \times V_i = M_f \times V_f$$


A molecular view of a solution of Cl_2 dissolved in water.
Add more water (solvent)
The solution after performing a dilution by adding water. Note how the number of moles of Cl_2 in the container does not change when performing the dilution; only the concentration changes. In this particular case, the concentration of Cl_2 drops to half of the starting concentration because the volume was doubled.

The photo on the right shows a volumetric flask used in dilution.

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Determine the milliliters of a solution of 14.8 *M* NH_3 required to produce 100 mL of NH_3 with a molarity of 1.00 *M.*

The initial volume can be determined by rearranging the dilution formula.

$$M_i V_i = M_f V_f$$

$$V_i = \frac{M_f V_f}{M_i}$$

Substituting the known values:

$$V_i = \frac{1.00\,\cancel{M} \times 100.0\text{ mL}}{14.8\,\cancel{M}} = \boxed{6.76\text{ mL}}$$

Two concepts that can help verify the result:

1. The volume of the more concentrated solution should always be less than the volume of the final solution.
2. The initial concentration of a solution is greater than the final concentration after dilution.

Quantitative Analysis

It is the determination of the amount of a substance or species present in a material.

Gravimetric Analysis

It is a type of quantitative analysis in which the amount of a species in a material is determined by converting the species to a product that can be isolated completely and weighed.

The figure on the right shows the reaction between an unknown amount of barium ion and K_2CrO_4 (yellow), forming the yellow $BaCrO_4$ precipitate.

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The figure to the right shows $BaCrO_4$ precipitate being filtered. It can then be dried and weighed to determine the mass of the precipitate.

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A 1.000-L sample of polluted water was analyzed for lead (II) ion, Pb^{2+}, by adding an excess of sodium sulfate to it. The mass of lead (II) sulfate that precipitated was 229.8 mg. What is the mass of lead in a liter of water? Give the answer as milligrams of lead per liter of solution.

Obtaining the mass percentage of Pb in $PbSO_4$:

$$\% \text{ Pb} = \frac{207.2 \text{ g/mol}}{303.3 \text{ g/mol}} \times 100\% = 68.32\%$$

Therefore, a sample of water contains:

$$229.8 \text{ mg } PbSO_4 \times 0.6832 = 157.0 \text{ mg Pb}$$

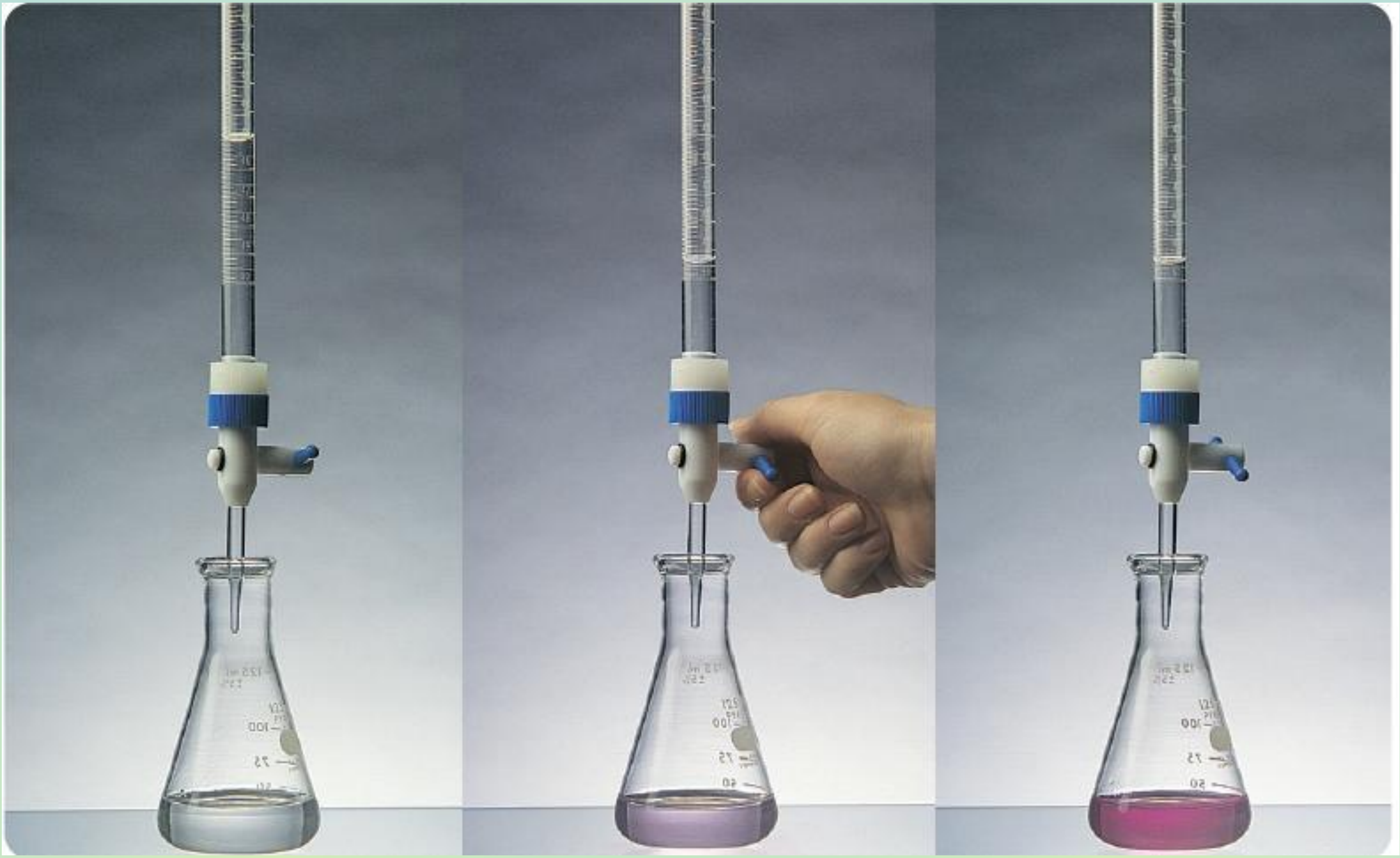
$\boxed{\text{The water sample contains 157.0 mg Pb per liter.}}$

Volumetric Analysis

It is a type of quantitative analysis based on titration.

Titration

It is a procedure for determining the amount of substance A by adding a carefully measured volume with a known concentration of B until the reaction of A and B is just complete.



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In the titration shown above, the indicator changes color to indicate when the reaction is just complete.

A flask contains an unknown amount of HCl. This solution is titrated with 0.207 *M* NaOH. It takes 4.47 mL of the NaOH solution to complete the reaction. Determine the mass of HCl.

Determine the stoichiometry of the reaction by writing the balanced equation.

$$\mathrm{NaOH}(aq) + \mathrm{HCl}(aq) \longrightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

Determining the mass of HCl:

$$4.47 \times 10^{-3}\,\mathrm{L\ NaOH\ soln} \times \frac{0.207\ \mathrm{mol\ NaOH}}{1\ \mathrm{L\ NaOH\ soln}} \times \frac{1\ \mathrm{mol\ HCl}}{1\ \mathrm{mol\ NaOH}} \times \frac{36.5\ \mathrm{HCl}}{1\ \mathrm{mol\ HCl}} = \boxed{0.0338\ \mathrm{g\ HCl}}$$