CHEM 103
CHEMISTRY I

Chapter 4
REACTIONS IN AQUEOUS SOLUTIONS

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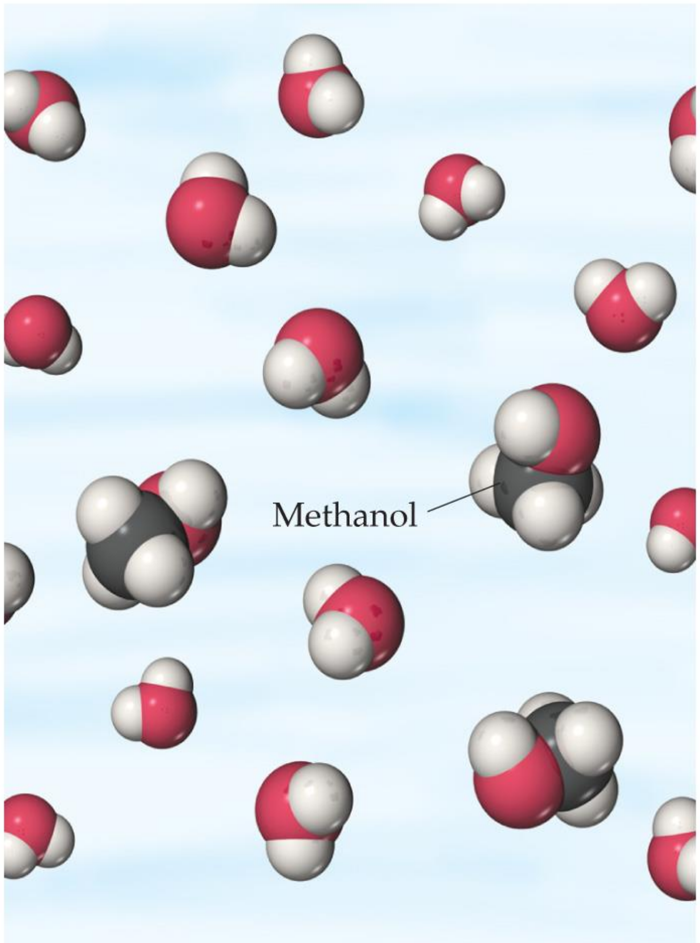
Solutions

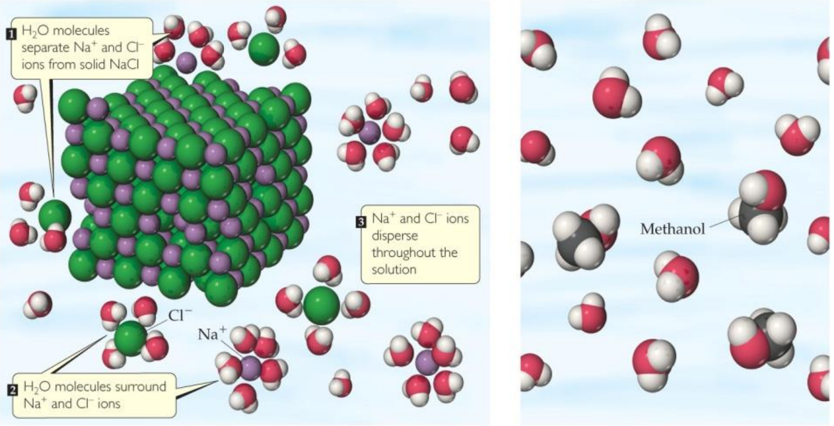


(b) Molecular substances like methanol, CH_3OH, dissolve without forming ions.

- **Solutions** are defined as homogeneous mixtures of two or more pure substances.
- The **solvent** is present in greatest abundance.
- All other substances are **solutes**.
- When water is the solvent, the solution is called an **aqueous solution**.

Aqueous Solutions

- Substances can dissolve in water by different ways:
  - Ionic Compounds dissolve by **dissociation**, where water surrounds the separated ions.
  - Molecular compounds interact with water, but most do NOT dissociate.
  - Some molecular substances react with water when they dissolve.



(a) Ionic compounds like sodium chloride, NaCl, form ions when they dissolve.

(b) Molecular substances like methanol, CH_3OH, dissolve without forming ions.

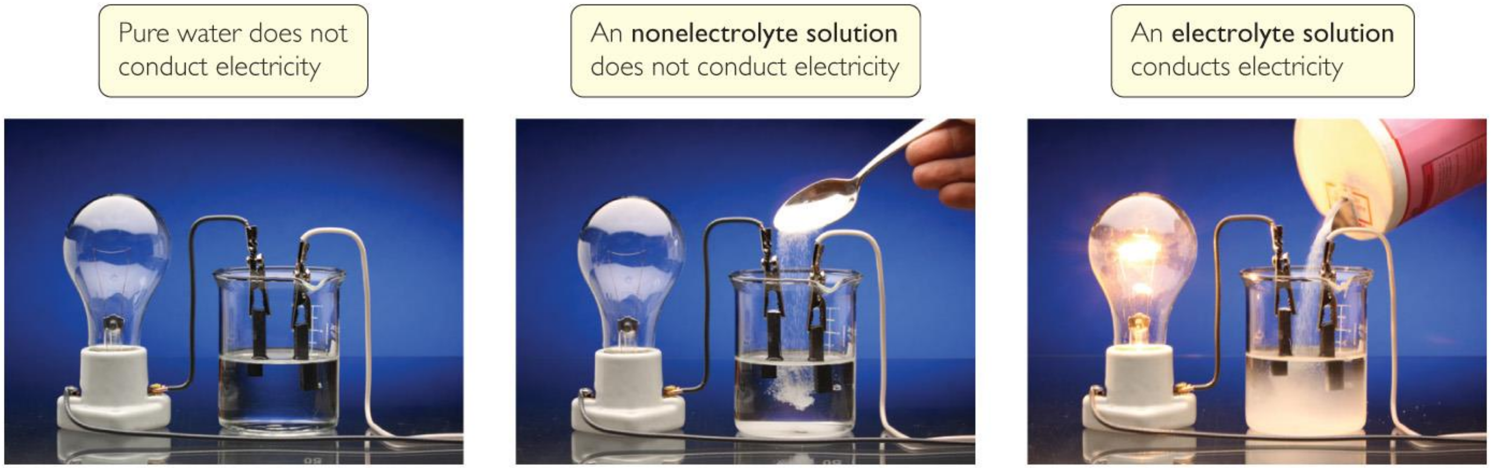
Electrolytes and Nonelectrolytes

Table 4.3 Summary of the Electrolytic Behavior of Common Soluble Ionic and Molecular Compounds

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids, weak bases	All other compounds

- An **electrolyte** is a substance that dissociates into ions when dissolved in water.
- A **nonelectrolyte** may dissolve in water, but it does not dissociate into ions when it does so.

Electrolytes



Pure water, $H_2O(l)$ | **Sucrose solution,** $C_{12}H_{22}O_{11}(aq)$ | **Sodium chloride solution,** $NaCl(aq)$

- A **strong electrolyte** dissociates completely when dissolved in water.
- A **weak electrolyte** only dissociates partially when dissolved in water.
- A **nonelectrolyte** does NOT dissociate in water.

Solubility of Ionic Compounds

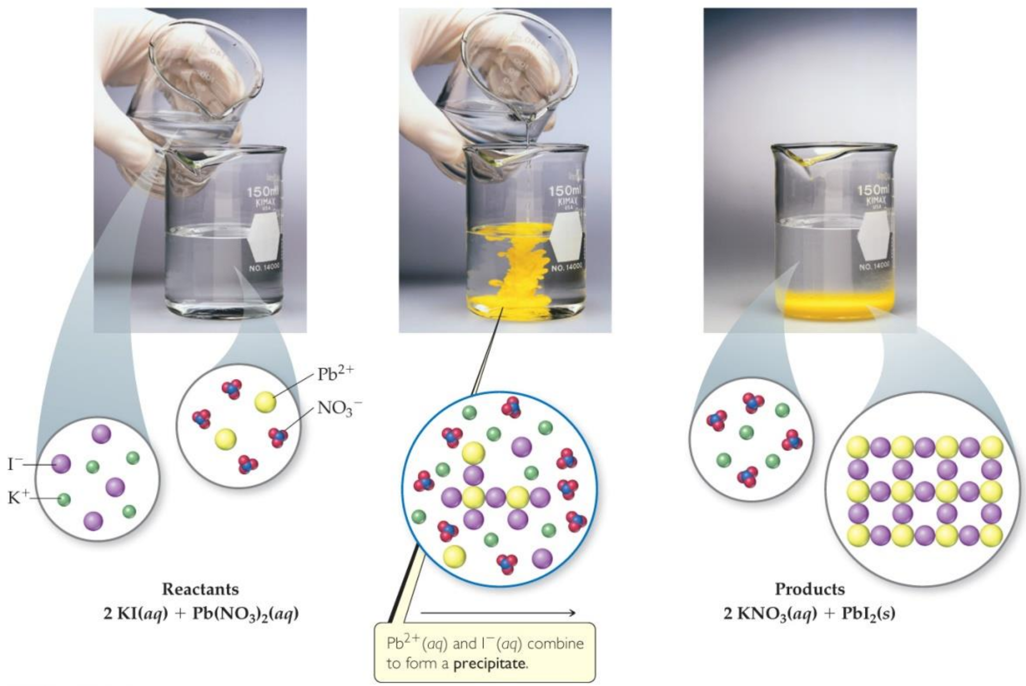
- Not all ionic compounds dissolve in water.
- A list of **solubility rules** is used to decide what combination of ions will dissolve.

Table 4.1 Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

Precipitation Reactions

When two solutions containing soluble salts are mixed, sometimes an insoluble salt will be produced. A salt "falls" out of solution, like snow out of the sky. This solid is called a **precipitate**.



Metathesis (Exchange) Reactions

- Metathesis comes from a Greek word that means “to transpose.”
- It appears as though the ions in the reactant compounds exchange, or transpose, ions, as seen in the equation below.

$$\mathbf{Ag}NO_3(aq) + K\mathbf{Cl}(aq) \longrightarrow \mathbf{AgCl}(s) + KNO_3(aq)$$

Completing and Balancing Metathesis Equations

- Steps to follow

1) Use the chemical formulas of the reactants to determine which ions are present.
2) Write formulas for the products: cation from one reactant, anion from the other. Use charges to write proper subscripts.
3) Check your solubility rules. If either product is insoluble, a precipitate forms.
4) Balance the equation.

Ways to Write Metathesis Reactions

1) Molecular equation
2) Complete ionic equation
3) Net ionic equation

Molecular Equation

The **molecular equation** lists the reactants and products without indicating the ionic nature of the compounds.

$$AgNO_3(\textit{aq}) + KCl(\textit{aq}) \longrightarrow AgCl(\textit{s}) + KNO_3(\textit{aq})$$

Complete Ionic Equation

- In the **complete ionic equation** all strong electrolytes (strong acids, strong bases, and soluble ionic salts) are dissociated into their ions.
- This more accurately reflects the species that are found in the reaction mixture.

$$Ag^+(aq) + NO_3^-(aq) + K^+(aq) + Cl^-(aq) \longrightarrow$$
$$AgCl(s) + K^+(aq) + NO_3^-(aq)$$

Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.
- The ions crossed out are called **spectator ions**, K+ and NO_3^-, in this example.
- The remaining ions are the reactants that form the product—an insoluble salt in a precipitation reaction, as in this example.

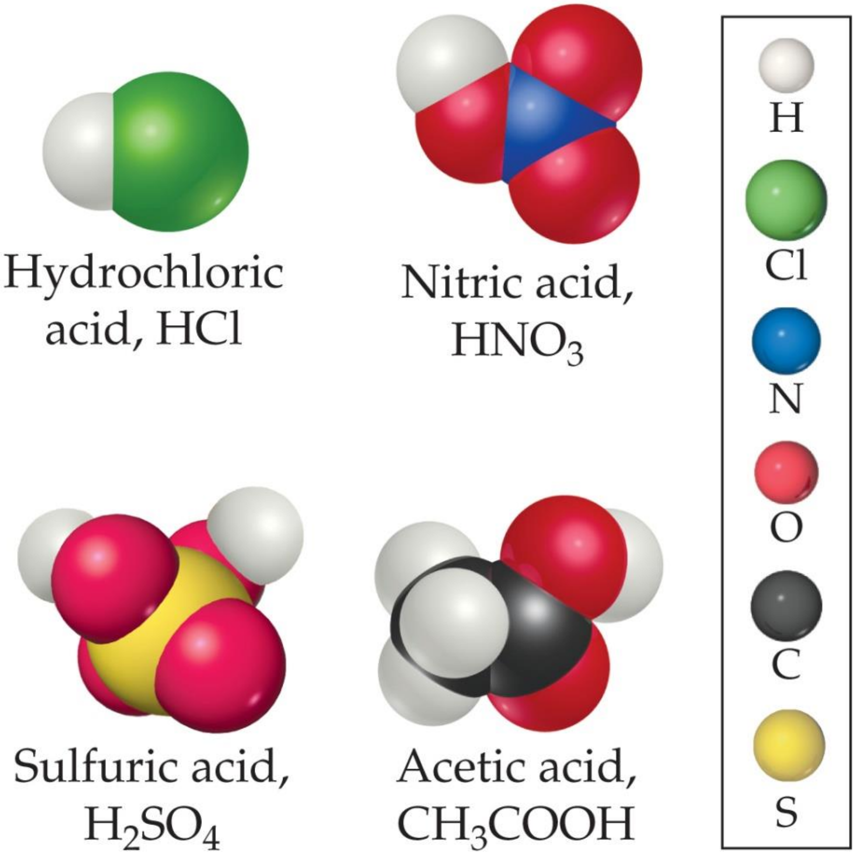
$$Ag^+(aq) + \cancel{NO_3^-}(aq) + \cancel{K^+}(aq) + Cl^-(aq) \longrightarrow$$

$$AgCl(s) + \cancel{K^+}(aq) + \cancel{NO_3^-}(aq)$$

Writing Net Ionic Equations

1. Write a balanced molecular equation.
2. Dissociate all strong electrolytes.
3. Cross out anything that remains unchanged from the left side to the right side of the equation.
4. Write the net ionic equation with the species that remain.

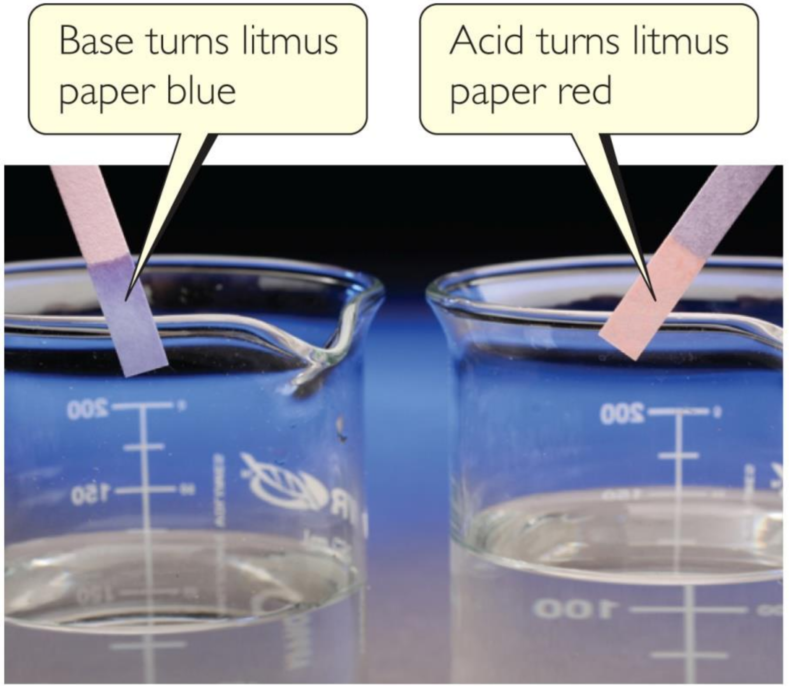
Acids



- The Swedish physicist and chemist S. A. Arrhenius defined acids as substances that increase the concentration of H^+ when dissolved in water.
- Both the Danish chemist J. N. Brønsted and the British chemist T. M. Lowry defined them as proton donors.

Bases

- Arrhenius defined bases as substances that increase the concentration of OH^- when dissolved in water.
- Brønsted and Lowry defined them as proton acceptors.



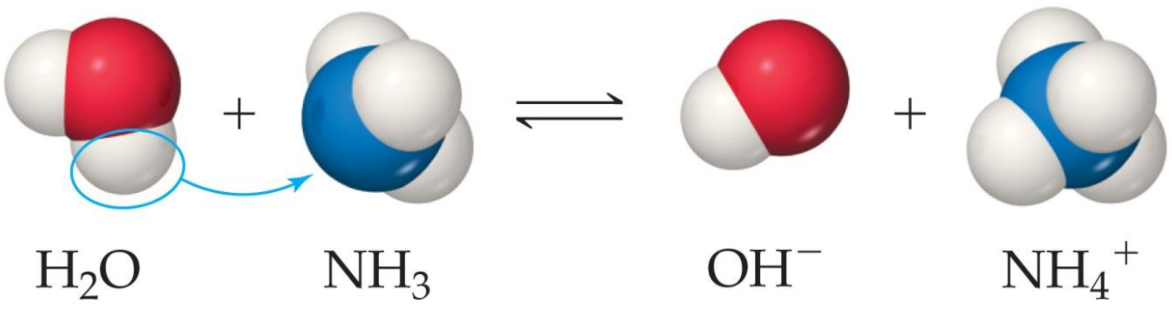
Strong or Weak?

- Strong acids completely dissociate in water; weak acids only partially dissociate.
- Strong bases dissociate to metal cations and hydroxide anions in water; weak bases only partially react to produce hydroxide anions.

Table 4.2 Common Strong Acids and Bases

Strong Acids	Strong Bases
Hydrochloric acid, HCl	Group 1A metal hydroxides [$LiOH$, $NaOH$, KOH, $RbOH$, $CsOH$]
Hydrobromic acid, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic acid, HI	
Chloric acid, $HClO_3$	
Perchloric acid, $HClO_4$	
Nitric acid, HNO_3	
Sulfuric acid (first proton), H_2SO_4	

Acid-Base Reactions



- In an acid–base reaction, the acid (H_2O above) donates a proton (H^+) to the base (NH_3 above).
- Reactions between an acid and a base are called **neutralization reactions**.
- When the base is a metal hydroxide, water and a **salt** (an ionic compound) are produced.

Neutralization Reactions

When a strong acid (like HCl) reacts with a strong base (like NaOH), the net ionic equation is circled below:

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

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{Na^+}(aq) + OH^-(aq) \longrightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$$

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

Gas-Forming Reactions

- Some metathesis reactions do not give the product expected.
- When a carbonate or bicarbonate reacts with an acid, the products are a salt, carbon dioxide, and water.

$$CaCO_3(s) + 2\ HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

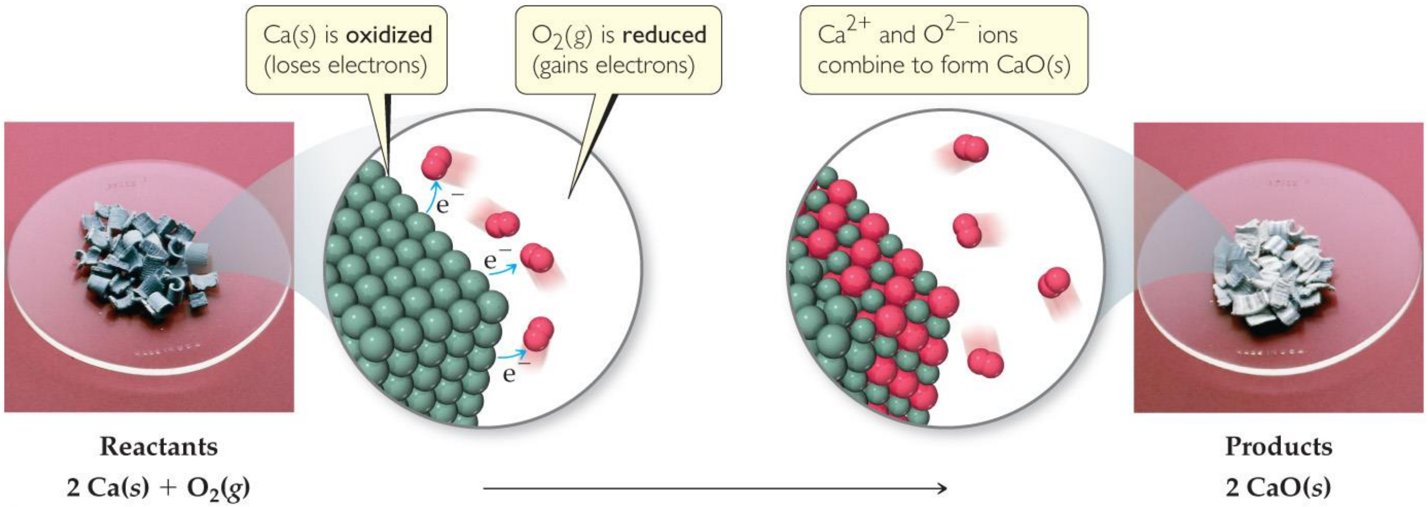
$$NaHCO_3(aq) + HBr(aq) \longrightarrow NaBr(aq) + CO_2(g) + H_2O(l)$$

Gas-Forming Reactions

This reaction gives the predicted product, but you had better carry it out in the hood—the gas produced has a foul odor!

$$Na_2S(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2S(g)$$

Oxidation-Reduction Reactions



- Loss of electrons is **oxidation**.
- Gain of electrons is **reduction**.
- One cannot occur without the other.
- The reactions are often called **redox reactions**.

Oxidation Numbers

To determine if an oxidation–reduction reaction has occurred, we assign an **oxidation number** to each element in a neutral compound or charged entity.

Rules to Assign Oxidation Numbers

- Elements in their elemental form have an oxidation number of zero.
- The oxidation number of a monatomic ion is the same as its charge.

Rules to Assign Oxidation Numbers

- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.
  - Oxygen has an oxidation number of −2, except in the peroxide ion, in which it has an oxidation number of −1.
  - Hydrogen is −1 when bonded to a metal, +1 when bonded to a nonmetal.

Rules to Assign Oxidation Numbers

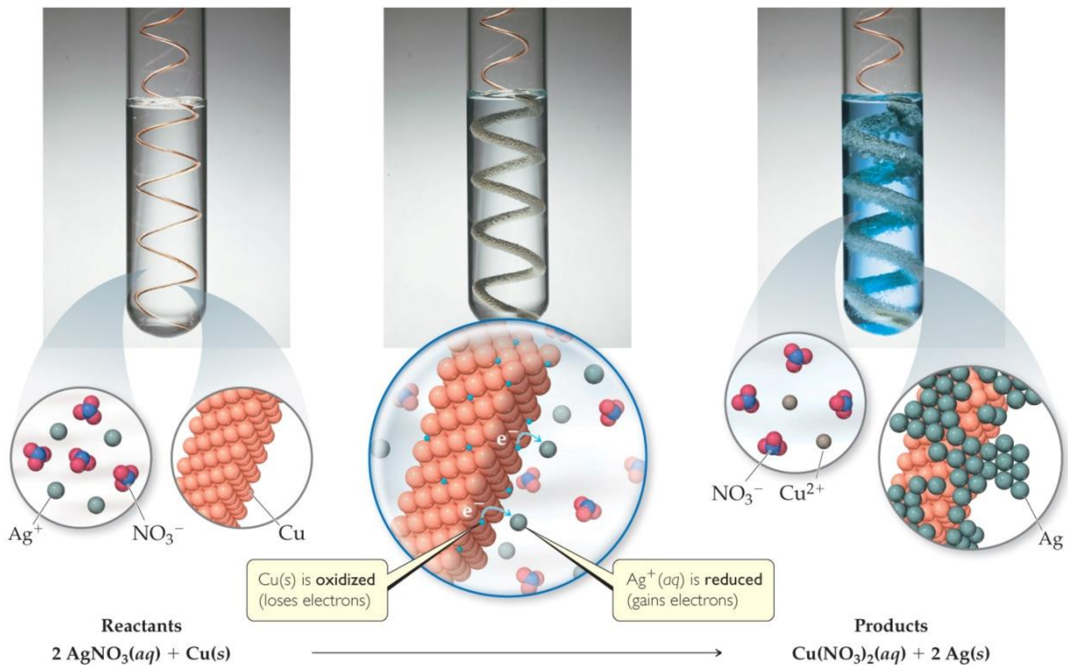
- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.
  - Fluorine always has an oxidation number of −1.
  - The other halogens have an oxidation number of −1 when they are negative; they can have positive oxidation numbers, most notably in oxyanions.

Rules to Assign Oxidation Numbers

- The sum of the oxidation numbers in a neutral compound is zero.
- The sum of the oxidation numbers in a polyatomic ion is the charge on the ion.

Displacement Reactions

In **displacement reactions**, ions oxidize an element.
In this reaction, silver ions oxidize copper metal:



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

The reverse reaction does NOT occur. Why not?

Activity Series

Table 4.5 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^{+}(aq) + e^{-}$
Potassium	$K(s) \longrightarrow K^{+}(aq) + e^{-}$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^{-}$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^{-}$
Sodium	$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^{-}$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^{-}$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^{-}$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^{-}$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^{-}$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^{-}$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^{-}$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^{-}$
Hydrogen	$H_2(g) \longrightarrow 2\,H^{+}(aq) + 2e^{-}$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
Silver	$Ag(s) \longrightarrow Ag^{+}(aq) + e^{-}$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^{-}$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^{-}$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$

- Elements higher on the activity series are more reactive.
- They are more likely to exist as ions.

Metal/Acid Displacement Reactions

- The elements above hydrogen will react with acids to produce hydrogen gas.
- The metal is oxidized to a cation.

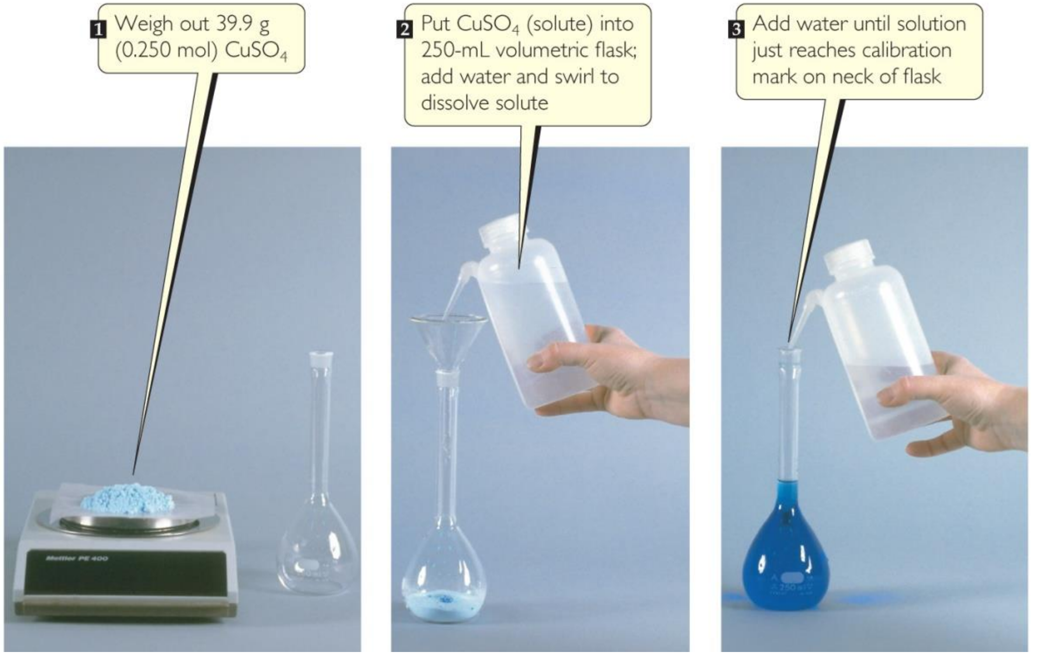
Molarity

- The quantity of solute in a solution can matter to a chemist.
- We call the amount dissolved its **concentration**.
- **Molarity** is one way to measure the concentration of a solution:

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

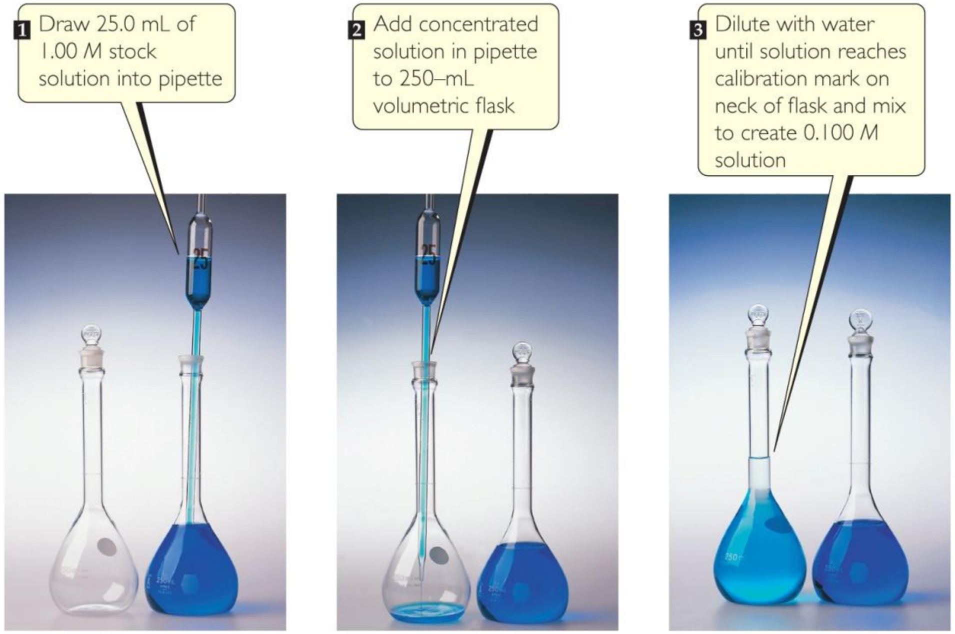
Mixing a Solution

- To create a solution of a known molarity, weigh out a known mass (and, therefore, number of moles) of the solute.
- Then add solute to a volumetric flask, and add solvent to the line on the neck of the flask.



Dilution

- One can also dilute a more concentrated solution by
  - using a pipet to deliver a volume of the solution to a new volumetric flask, and
  - adding solvent to the line on the neck of the new flask.



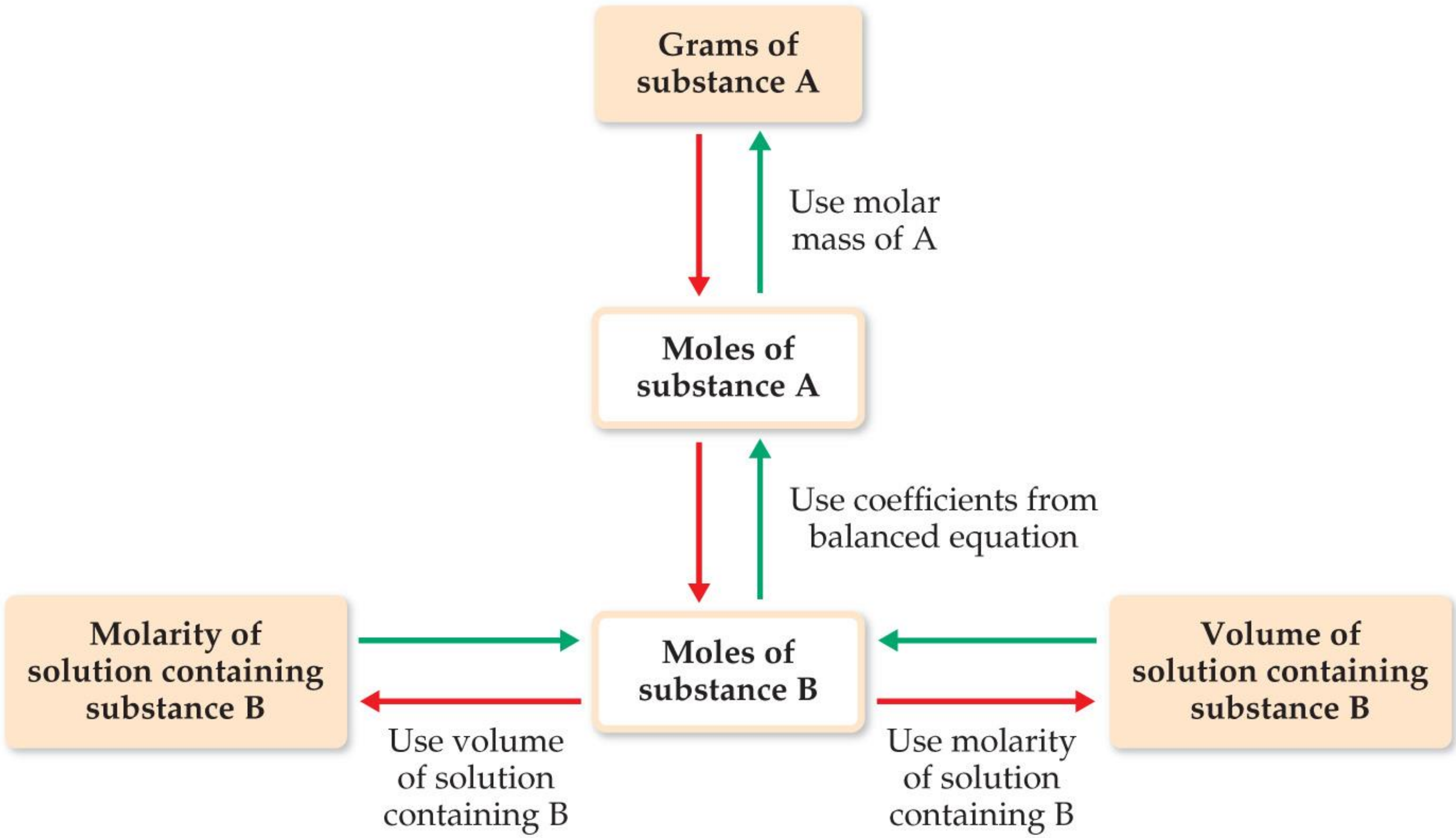
Dilution

The molarity of the new solution can be determined from the equation

$$M_c \times V_c = M_d \times V_d,$$

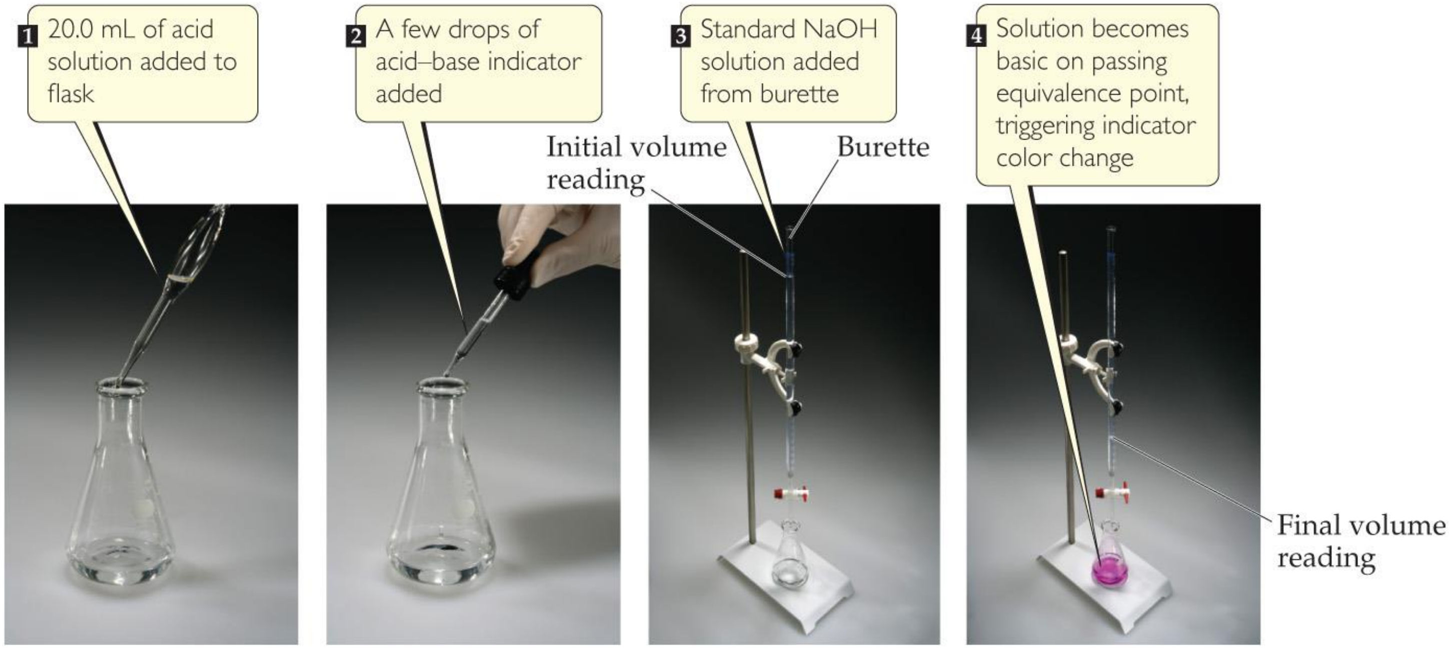
where M_c and M_d are the molarity of the concentrated and dilute solutions, respectively, and V_c and V_d are the volumes of the two solutions.

Using Molarities in Stoichiometric Calculations

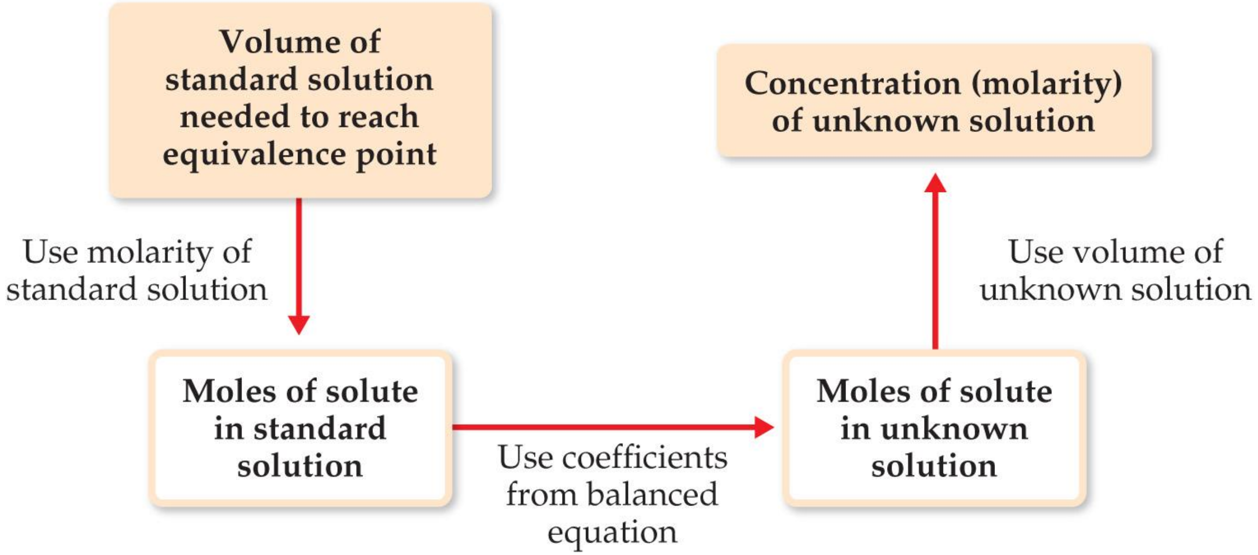


Titration

A **titration** is an analytical technique in which one can calculate the concentration of a solute in a solution.



Titration



- A solution of known concentration, called a **standard solution**, is used to determine the unknown concentration of another solution.
- The reaction is complete at the **equivalence point**.