



Chapter 3: Solution Chemistry

Section 3.1: Solubility Rules (For Ionic Compounds in Water)

Section 3.1.1: Introduction

Solubility is defined as the ability of a compound to dissolve in a solvent, here, in water. Different compounds have different solubility in water. Here, we will only deal with ionic compounds since a number of simple rules exist to decide whether specific ions impart solubility or not. For example, NaCl dissolves very easily in water while CaCO_3 dissolves only very sparingly. Molecular compounds can also exhibit a wide range of solubility in water (for example sugar is very soluble in water, while oil and gasoline are not). The discussion of solubility of molecular compounds is however postponed until [Chapter 10](#) where we study “Intermolecular Forces and Liquid Properties”.

Section 3.1.2: Solubility Rules for Chlorides, Bromides and Iodides

All chloride, bromide and iodide ionic compounds are soluble in water except those containing silver(I), mercury(I), lead(II) and copper(I).

Section 3.1.3: Dissociation of Chlorides, Bromides and Iodides

Practice making chloride, bromide and iodide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 3.1.4: Solubility Rules for Acetate, Chlorate, Perchlorate, Nitrate and Hydroxide Compounds

All acetate, chlorate, perchlorate and nitrate ionic compounds are soluble in water.

All group 1 hydroxide compounds are soluble in water. Magnesium, calcium, strontium and barium hydroxides are only slightly soluble. All other hydroxides are insoluble in water.

Section 3.1.5: Dissociation of Acetate, Chlorate, Perchlorate, Nitrate and Hydroxide Compounds

Practice making acetate, chlorate, perchlorate, nitrate and hydroxide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 3.1.6: Solubility Rules for Sulfides

Group 1 sulfide ionic compounds are water soluble while all other sulfide compounds are not.

Section 3.1.7: Dissociation of Sulfide Compounds

Practice making sulfide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 3.1.8: Solubility Rules for Carbonates, Chromates and Sulfates

Group 1 carbonate and chromate ionic compounds are water soluble while all other carbonate and chromate compounds are not.

Most sulfate ionic compounds are soluble except ionic compounds made with Mg, Ca, Sr, Ba, Ra, Hg(I) and Pb(II).

Section 3.1.9: Dissociation of Carbonate, Chromate and Sulfate Compounds

Practice making carbonate, chromate and sulfate compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 3.1.10: Solubility Rules and Dissociation of Ammonium Compounds

All ammonium compounds are water soluble.

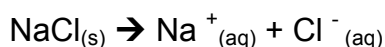
Part 3.2: Ionic Equations and Precipitation Reactions

Section 3.2.1: Writing Ionic Equations

Ionic equations represent chemical reactions between ions. Such reactions require at least two ionic compounds (4 ions). To illustrate precipitation reactions, we will consider three examples.

3.2.1.a Reaction of $\text{NaCl}_{(aq)}$ with $\text{AgNO}_3_{(aq)}$

$\text{NaCl}_{(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $\text{Na}^+_{(aq)}$ and $\text{Cl}^-_{(aq)}$ ions, according to the ionic equation:



$\text{AgNO}_3_{(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $\text{Ag}^+_{(aq)}$ and $\text{NO}_3^-_{(aq)}$ ions, according to the ionic equation:



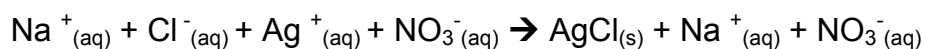
Let us mix these two aqueous solutions. Does a reaction occur and if so, what reaction is it?

Obviously, $\text{Cl}^-_{(aq)}$ will not react with $\text{Na}^+_{(aq)}$. Does $\text{Cl}^-_{(aq)}$ form a compound with $\text{Ag}^+_{(aq)}$? Whether these ions form a compound in the presence of water depends on whether the resulting compound is soluble in water or not. The solubility rules for chlorides, bromides and iodides state that silver chloride is not soluble in water. Hence, the silver and chloride ions form an ionic compound. We say that a **precipitate** of solid silver chloride (white creamy solid) has formed by reaction of sodium chloride with silver nitrate.

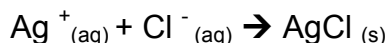
Now, we must ask whether another precipitation reaction takes place between the sodium and the nitrate ions. Solubility rules for nitrate ions tell us that all nitrate compounds are soluble in water. Hence, sodium and nitrate ions remain as free ions in solution and do not form a precipitate.

The sodium and nitrate ions do not participate in the reaction. They are present and actually watch the silver and chloride ions react to form a solid ionic compound. We say that the sodium and nitrate ions are **spectator ions**.

The **complete ionic equation** between these compounds is written as:

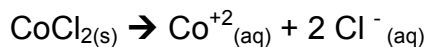


Since the $\text{Na}^+_{(\text{aq})}$ and $\text{NO}_3^-_{(\text{aq})}$ ions are spectator ions (i.e. they do not participate in the reaction), they can be cancelled on both sides of the equation. The resulting equation is called the **net ionic equation** and is written as:



3.2.1.b Reaction of $\text{CoCl}_2_{(\text{aq})}$ with $\text{Na}_2\text{S}_{(\text{aq})}$

$\text{CoCl}_2_{(\text{s})}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $\text{Co}^{+2}_{(\text{aq})}$ and $\text{Cl}^-_{(\text{aq})}$ ions according to the ionic equation:



$\text{Na}_2\text{S}_{(\text{s})}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $\text{Na}^+_{(\text{aq})}$ and $\text{S}^{-2}_{(\text{aq})}$ ions according to the ionic equation:

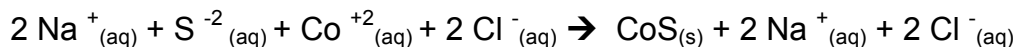


Mixing of these two aqueous solutions then makes it possible for a reaction between these ions to take place. Will a reaction occur and if so, how is it represented?

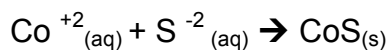
Obviously, $\text{Na}^+_{(\text{aq})}$ will not react with $\text{Cl}^-_{(\text{aq})}$ since NaCl is water soluble but $\text{Co}^{+2}_{(\text{aq})}$ will form a compound with $\text{S}^{-2}_{(\text{aq})}$ since CoS is not water soluble (see solubility rules for sulfides). Mixing the two solutions lead to a dark **precipitate** of solid cobalt(II) sulfide.

Since the sodium and chloride ions remain as free ions in solution and do not form a precipitate, they are **spectator ions** during the precipitation of CoS .

The **complete ionic equation** between these compounds is written as:



Since the $\text{Na}^+_{(\text{aq})}$ and $\text{Cl}^-_{(\text{aq})}$ ions are spectator ions (i.e. they do not participate in the reaction), they can be cancelled on both side of the equation. The resulting equation is called the **net ionic equation** and is written as:



3.2.1.c Reaction of $\text{NaCl}_{(\text{aq})}$ and $\text{Ca}(\text{NO}_3)_2_{(\text{aq})}$

$\text{NaCl}_{(s)}$ and $\text{Ca}(\text{NO}_3)_{2(s)}$ are both soluble in water. However, $\text{NaNO}_{3(s)}$ and $\text{CaCl}_{2(s)}$ are also soluble in water. Hence, mixing aqueous solutions of sodium chloride and calcium nitrate does not lead to precipitation. All ions remain free in solution and there is NO reaction between them.

Sections 3.2.2 – 3.2.3: Interactive Ionic Reactions

Practice, practice, practice

Part 3.3: Electrolytes and Nonelectrolytes

Section 3.3.1: Electrolytes and Nonelectrolytes

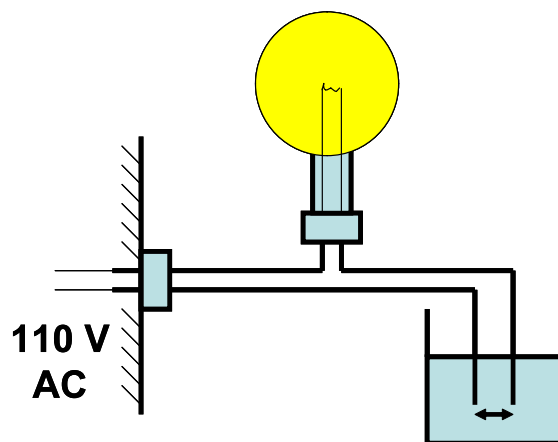
A substance which, when dissolved in water, makes an electrically conducting solution is called an **electrolyte**. Electrolytes are substances which dissociate into ions when dissolved in water.

Examples of electrolytes are ionic compounds which are water soluble, acids and bases. (Note that ionic compounds which are not soluble in water will not dissociate into ions when mixed with water, hence, they will not conduct electricity).

A substance which, when dissolved in water, makes an electrically non-conducting solution is called a **nonelectrolyte**. Nonelectrolytes do not dissociate into ions when dissolved in water. Examples of nonelectrolytes are molecular compounds other than acids and bases.

To decide whether a compound is an electrolyte or a nonelectrolyte, we can use the following experiment: (see video).

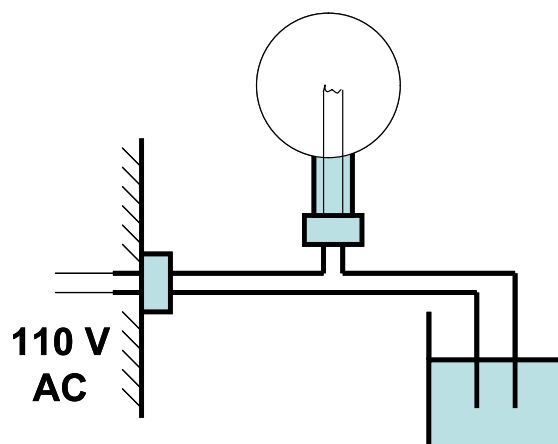
A compound is dissolved in water and the solution placed in a beaker. The two leads of the electrical circuit are dipped in the beaker. If dissolution of the compound is accompanied by dissociation of the compound into ions, the ions will be attracted to the oppositely charged leads. Motion of the ions between the leads (where the arrow is) closes the electrical circuit and current flows through the light bulb (see figure below).



Solution contains an Electrolyte
The bulb is lit !!!

Note that electrical current is characterized by the motion of electrons in the metallic wires and the motion of ions (not electrons) in the solution.

If the solution contains no ions (or a really low concentration of ions as in distilled water), then the circuit is not closed and electrical current does not flow through the light bulb.



Solution contains a
Non-electrolyte
The bulb is NOT lit!!!

Section 3.3.2: Strong and Weak Electrolytes

Electrolytes dissolved in water produce ions to various extents.

Some electrolytes when dissolved in water **dissociate completely** into cations and anions. Such compounds are called **strong electrolytes**.

Examples of strong electrolytes are soluble ionic compounds, strong acids and strong bases (see part 2.5). For example, when $\text{NaCl}_{(s)}$ is dissolved in water all NaCl dissociates into $\text{Na}^+_{(aq)}$ and $\text{Cl}^-_{(aq)}$. NaCl is a strong electrolyte. Similarly, when dissolving HCl gas (hydrogen chloride) in water, we form hydrochloric acid. Hydrochloric acid is written as $\text{HCl}_{(aq)}$ and exists as $\text{H}^+_{(aq)}$ and $\text{Cl}^-_{(aq)}$ ions. Using the above set-up, a strong electrolyte is characterized by a very brightly lit light bulb.

Some electrolytes, when dissolved in water **dissociate partially** into cations and anions. Partial dissociation means that some of the original compound exists in solution in the **undissociated** (molecular) form. Substances that dissociate partially in solution are called **weak electrolytes**. Examples of weak electrolytes are weak acids and weak bases (see part 2.5). For instance, acetic acid, $\text{HC}_2\text{H}_3\text{O}_2(l)$, the main component of vinegar dissolves in water and dissociates partially into $\text{H}^+_{(aq)}$ and $\text{C}_2\text{H}_3\text{O}_2^-_{(aq)}$ ions. However, most of the molecules of $\text{HC}_2\text{H}_3\text{O}_2(aq)$ remain undissociated in solution. Using the above set-up, a weak electrolyte is characterized by a dimly lit light bulb.

Part 3.4: Acids and Bases

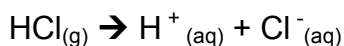
Section 3.4.1: Properties of Acids and Bases

Acids and bases taste different (sour vs. soapy) but do not try this at home unless it is vinegar, lemon juice (acids) or milk of magnesia (base).

Section 3.4.2: Arrhenius Acids and Bases

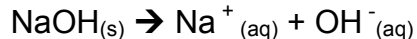
Svante Arrhenius proposed the first definition for acids and bases. An acid according to Arrhenius is a substance which, when dissolved in water, produces hydrogen ions (protons), $\text{H}^+_{(aq)}$.

Consider HCl gas. The dissolution reaction is written as:



A base according to Arrhenius is a substance which, when dissolved in water, produces hydroxide ions, $\text{OH}^-_{(aq)}$.

Consider NaOH solid. The dissolution reaction is written as:



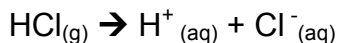
Section 3.4.3: Brønsted - Lowry Acids and Bases

Acids according to **Brønsted-Lowry** are substances that donate protons (H^+) to another species.

Consider the formation of hydrochloric acid:



$\text{H}_3\text{O}^+_{(aq)}$ is called the hydronium ion. This reaction is also written as:



as $\text{H}^+_{(aq)}$ is equivalent to $\text{H}_3\text{O}^+_{(aq)}$

Bases according to **Brønsted-Lowry** are substances that accept a proton (H^+) from another species.

Consider the dissolution of ammonia in water:



$\text{NH}_4^+_{(aq)}$ is the ammonium ion.

Section 3.4.4: Strong Acids and Weak Acids

Strong acids completely ionize when dissolved in water. The following list of strong acids is a list you need to memorize. All acids in this list are strong acids. All other acids are weak acids.

Strong Acids:

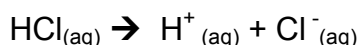
HCl _(aq)	hydrochloric acid
HBr _(aq)	hydrobromic acid
HI _(aq)	hydroiodic acid
HNO _{3(aq)}	nitric acid
HClO _{4(aq)}	perchloric acid
HClO _{3(aq)}	chloric acid
H ₂ SO _{4(aq)}	sulfuric acid

MEMORIZE THE ABOVE LIST

Weak Acids

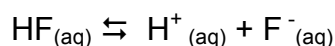
HClO _{2(aq)}	chlorous acid
HClO _(aq)	hypochlorous acid
HBrO _{4(aq)}	perbromic acid
HBrO _{3(aq)}	bromic acid
HBrO _{2(aq)}	bromous acid
HBrO _(aq)	hypobromous acid
HF _(aq)	hydrofluoric acid
HC ₂ H ₃ O _{2(aq)}	acetic acid
HNO _{2(aq)}	nitrous acid
H ₂ SO _{3(aq)}	sulfurous acid
H ₃ PO _{4(aq)}	phosphoric acid
H ₃ PO _{3(aq)}	phosphorous acid
etc...	

The complete dissociation of a **strong acid** is represented by the following chemical equation:



The **single arrow** indicates that the ionization of HCl goes to completion. All HCl dissociates to H⁺_(aq) and Cl⁻_(aq) ions. Strong acids are strong electrolytes.

The partial dissociation of a **weak acid** is represented by the following chemical equation:



The **double arrow** indicates that the ionization of HF does not go to completion. Some HF dissociates to H⁺_(aq) and F⁻_(aq) ions and some HF remains in molecular form. We say that the dissociation (or ionization) of HF reaches an equilibrium where HF molecules, H⁺_(aq) and F⁻_(aq) ions coexist in solution. Weak acids are weak electrolytes.

Section 3.4.5: Strong Bases and Weak Bases

Strong bases completely ionize when dissolved in water. The following list of strong bases is a list you need to memorize. All bases in this list are strong bases. All other bases are weak bases.

Strong Bases:

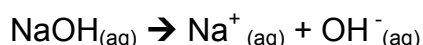
$\text{LiOH}_{(aq)}$	lithium hydroxide
$\text{NaOH}_{(aq)}$	sodium hydroxide
$\text{KOH}_{(aq)}$	potassium hydroxide
$\text{CsOH}_{(aq)}$	cesium hydroxide
$\text{Ca(OH)}_2_{(aq)}$	calcium hydroxide
$\text{Sr(OH)}_2_{(aq)}$	strontium hydroxide
$\text{Ba(OH)}_2_{(aq)}$	barium hydroxide

Weak Bases

$\text{Mg(OH)}_2_{(aq)}$	magnesium hydroxide
$\text{NH}_3_{(aq)}$	ammonia
$\text{F}^-_{(aq)}$	fluoride ion
$\text{C}_2\text{H}_3\text{O}_2^-_{(aq)}$	acetate ion
$\text{ClO}_2^-_{(aq)}$	chlorite ion

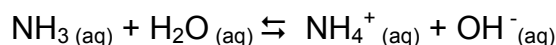
MEMORIZE THE ABOVE LIST

Note that strong bases are groups I and II hydroxides except beryllium and magnesium hydroxides. The complete dissociation of a **strong base** is represented by the following chemical equation (in the case of NaOH):



The **single arrow** (\rightarrow) indicates that the dissociation of NaOH goes to completion. All NaOH dissociate to $\text{Na}^+_{(aq)}$ and $\text{OH}^-_{(aq)}$ ions. Strong bases are strong electrolytes.

The partial dissociation of a **weak base** is represented by the following chemical equation (in the case of NH_3):



The **double arrow** (\rightleftharpoons) indicates that the ionization of $\text{NH}_3_{(aq)}$ does not go to completion. Some $\text{NH}_3_{(aq)}$ ionizes to $\text{NH}_4^+_{(aq)}$ and some NH_3 remains in molecular form. We say that the ionization of NH_3 reaches an equilibrium where NH_3 molecules, $\text{NH}_4^+_{(aq)}$ and $\text{OH}^-_{(aq)}$ ions coexist in solution. Weak bases are weak electrolytes.

Sections 3.5 – 3.6: Introduction to Molarity

Molarity refers to the concentration of species in solution. The species is called **solute**. The solute dissolves in a liquid called the **solvent**. Hence, a **solution** is composed of solute and solvent. Qualitatively, a dilute solution has a low concentration of solute and a concentrated solution has a high concentration of solute. The terms dilute and concentrated are used in a comparative sense.

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

$$\text{moles (n)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$\text{Molarity (M)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)} \times \text{liters of solution (L)}}$$

Thus, $\text{mass (g)} = \text{Molarity (M)} \times \text{molar mass (g/mol)} \times \text{liters of solution (L)}$

The molarity (M) of a solution can be used to calculate:

- (a) the amount of solute (moles or mass) in a given volume of solution, and
- (b) the volume of solution containing a given amount (moles) of solute.

Example 1: Calculate the mass of KMnO_4 needed to prepare 700. mL of a 0.583 M solution of KMnO_4 .

$\text{Molar Mass (KMnO}_4) = 39.1 + 54.9 + 4 \times 16.0 = 158.0 \text{ g/mol (KMnO}_4\text{)}$.

$$\text{Molarity} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

Hence, $\text{Mass (g)} = \text{Molarity (mol/L)} \times \text{molar mass (g/mol)} \times \text{volume (L)}$

Note that the volume must be expressed in liter (L) and $1 \text{ L} = 1000 \text{ mL}$

$$\text{Mass (g)} = 0.583 \text{ mol/L} \times 158.04 \text{ g/mol} \times 0.700 \text{ L} = 64.5 \text{ g}$$

Therefore, 64.5 g of KMnO_4 dissolved in 700.mL of water produces a 0.583 M KMnO_4 solution.

Example 2: Calculate the volume in mL of a 0.215 M H_2SO_4 solution containing 0.948 g of H_2SO_4 .

$\text{Molar Mass (H}_2\text{SO}_4) = 2 \times 1.0 + 32.1 + 4 \times 16.0 = 98.1 \text{ g/mol (H}_2\text{SO}_4\text{)}$.

$$\text{Molarity (mol/L)} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

Hence,

$$\text{Volume of solution (L)} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{molarity (mol/L)}}$$

$$\text{Volume of solution (L)} = \frac{0.948 \text{ g}}{98.1 \text{ g/mol} \times 0.215 \text{ mol/L}} = 0.0449 \text{ L}$$

Hence, the volume of solution is 44.9 mL.

Example 3: A 75.0 mL solution of AgNO₃ contains 0.0385 moles of AgNO₃. Calculate the molarity of the solution.

$$\text{Molarity (M)} = \frac{\text{moles of solute (n)}}{\text{liters of solution (L)}} = \frac{0.0385 \text{ mol}}{0.075 \text{ L}}$$

$$\text{Molarity} = 0.513 \text{ M}$$

In Section 3.6, practice the Interactive Problems.

Sections 3.7 – 3.8: Molarity of Ionic Solutions

Aqueous solutions of salts conduct electricity, showing that salts break into ions in solution. Salts break down into cations and anions. The ratio of cations and anions produced is given by the ion's chemical formula.

Polyatomic ions remain intact as ions in solution.

For example: Sodium sulfate (Na₂SO₄) dissolved in water breaks down into Na⁺_(aq) and SO₄⁻²_(aq) ions. SO₄⁻² is a polyatomic ion. From the chemical formula Na₂SO₄, the ratio of ions in solution is: 2 Na⁺_(aq) to 1 SO₄⁻²_(aq)

For example: Consider an aqueous solution of 1 M NaCl. In solution, NaCl breaks down into 1 Na⁺_(aq) and 1 Cl⁻_(aq) ions. The concentration of Na⁺_(aq) is 1 M and that of Cl⁻_(aq) is 1 M

For example: Now, consider a 1 M aqueous solution of Na₃PO₄. In solution, Na₃PO₄ breaks down into 3 Na⁺_(aq) and 1 PO₄⁻³_(aq). Hence, in 1 M Na₃PO₄ solution, the concentration of Na⁺_(aq) is 3 M and that of PO₄⁻³_(aq) is 1 M.

Thus, an ionic solution contains a mixture of ions, each with its own molarity.

Example: Calculate the molarity of ionic species present in 250 mL of an aqueous solution containing 1.85 g of ammonium sulfate.

Ammonium sulfate has a chemical formula: $(\text{NH}_4)_2\text{SO}_4$

The molar mass of $(\text{NH}_4)_2\text{SO}_4$ is: $2 \times 14.0 + 8 \times 1.0 + 32.1 + 4 \times 16.0 = 132.1 \text{ g/mol}$

$$\text{Molarity } ((\text{NH}_4)_2\text{SO}_4) = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

$$\text{Molarity } ((\text{NH}_4)_2\text{SO}_4) = \frac{1.85 \text{ g}}{132.1 \text{ g/mol} \times 0.25 \text{ L}} = 0.056 \text{ mol/L}$$

Hence, the concentration of $(\text{NH}_4)_2\text{SO}_4$ is 0.056 M.

In solution, $(\text{NH}_4)_2\text{SO}_4$ breaks down as $2 \text{ NH}_4^+_{(\text{aq})}$ and $1 \text{ SO}_4^{2-}_{(\text{aq})}$

Thus, the concentration of $\text{NH}_4^+_{(\text{aq})} = 2 \times 0.056 = 0.11 \text{ M}$ and
the concentration of $\text{SO}_4^{2-}_{(\text{aq})} = 0.056 \text{ M}$

In Section 3.8, practice Interactive Problems.

Sections 3.9 – 3.10: Diluting Solutions

When solutions of acids, bases and soluble ionic compounds are manufactured and transported, they are often prepared as **stock solutions**. These stock solutions are highly concentrated solutions. Different usages of these solutions often require different (typically lower) concentrations of solutes in these solutions. To prepare a solution of a substance from a stock solution we will need to dilute the stock solution.

Example: Commercially available concentrated HCl is usually a 12.0 M solution. How does one prepare 100. mL of 1.0 M HCl solution from 12.0 M HCl solution?

The obvious answer is to dilute the concentrated HCl (i.e. 12.0 M HCl) solution. But, what volume of 12.0 M HCl should be diluted? To answer this question, recall the definition of molarity.

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

Hence, moles of solute (n) = Molarity (M) \times liters of solution (L)

This equation can be written as: moles of solute (n) = $M_i \times V_i$. The subscript “i” stands for “initial”.

Thus, M_i = initial molarity
and V_i = initial volume

Now, add water to the stock solution. This changes the concentration from M_i to M_f and the volume from V_i to V_f . The subscript “f” stands for “final”.

Thus, M_f = final molarity
and V_f = final volume

Because the number of moles of solute has not changed during dilution, moles of solute (n) = $M_f \times V_f$, and,

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

In the above example: M_i = 12.0 M
 V_i = To be calculated
 M_f = 1.0 M
 V_f = 100. mL

$$12.0 \text{ M} \times V_i = 1.0 \text{ M} \times 100. \text{ mL}$$

$$V_i = \frac{1.0 \text{ M} \times 100. \text{ mL}}{12.0 \text{ M}} = 8.3 \text{ mL}$$

(in the correct number of significant figures)

Note: You can use any volume units as long as V_i and V_f have the same units.

Hence, to prepare the diluted solution, take a clean 100 mL volumetric flask, add approximately 25 mL distilled water, then, carefully and slowly, add 8.3 mL 12.0 M HCl solution. Add more water to bring the total volume to 100. mL.

Example: A stock solution of 15.8 M HNO_3 is given. How many mL of the stock solution are required to make 0.250 liters of 0.12 M HNO_3 ?

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

M_i = 15.8 M
 V_i = To be calculated
 M_f = 0.12 M
 V_f = 0.250 L

$$15.8 \text{ M} \times V_i = 0.12 \text{ M} \times 0.250 \text{ L}$$

$$V_i = \frac{0.12 \text{ M} \times 0.250 \text{ L}}{15.8 \text{ M}}$$

$$V_i = 0.0019 \text{ L} = 1.9 \text{ mL}$$

In Section 3.10, practice Interactive Problems.

Section 3.11: Volumetric Analysis

An acid – base reaction is often called a neutralization reaction. When just enough base is added to the acid, the acid is neutralized.

Acids and bases are classified as strong and weak. Hence, the nature of neutralization depends upon the strengths of acids and bases.

The acid – base reactions are classified into three categories:

Category 1: Strong acid – Strong base

Category 2: Weak acid – Strong base

Category 3: Strong acid – Weak base

In all three categories, there are a few common steps that should be applied to perform a volumetric analysis.

Category 1: Strong acid – Strong base

Step 1: List the ions present in a combined solution.
For example: In the reaction of HCl and NaOH,
The ions present in a combined solution will be H^+ , Cl^- , Na^+ ,
 OH^-

Step 2: Figure out what reactions would occur with these ions and determine the net ionic equation. In the example, there are two possibilities.
 H^+ combines with OH^- to form H_2O
 Na^+ combines with Cl^- to form NaCl

The net ionic equation is: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

This is the net ionic equation, in general, for a reaction between a strong acid and a strong base.

Step 3: Perform volumetric calculation using the relationship: $M_i \times V_i = M_f \times V_f$
This equation can be written as

$$M_{H^+} \times V_{H^+} = M_{OH^-} \times V_{OH^-}$$

since the neutralization reaction is carried out under stoichiometric conditions only when the number of moles of OH^- is equal to the number of moles of H^+ . (Remember moles = molarity \times volume)

Example: Calculate the volume in mL of 0.100 M HCl needed to neutralize 50.0 mL of 0.250 M NaOH.

$$M_{H^+} = 0.100 \text{ M}$$

$$V_{H^+} = \text{To be calculated}$$

$$M_{OH^-} = 0.250 \text{ M}$$

$$V_{OH^-} = 50.0 \text{ mL}$$

$$0.100 \text{ M} \times V_{OH^-} = 0.250 \text{ M} \times 50.0 \text{ mL}$$

$$V_{OH^-} = \frac{0.250 \text{ M} \times 50.0 \text{ mL}}{0.100 \text{ M}}$$

$$V_{OH^-} = 125 \text{ mL}$$

Hence, 125 mL of 0.100 M HCl are needed to neutralize 50.0 mL of 0.250 M NaOH.

A titration is an analytical technique used to determine the volume or concentration of an acid or base. Volumetric analysis is a method involving calculations based on titration. A titration can be performed and the volumetric analysis can be done between

Category 1: Strong acid – Strong base

Category 2: Weak acid – Strong base

Category 3: Strong acid – Weak base

Considering the scope of the topic only category 1 is discussed in this chapter. For a more thorough discussion of titration reactions, see [Chapter 16](#).

Note: The example below is not discussed in depth on the DVD.

Example: Let us determine the volume of a 0.250 M Ba(OH)₂ aqueous solution needed to neutralize 15.0 mL of a 0.175 M HCl solution.

We will solve this problem by two methods. The first method is similar to the one used in the example. The second method is identical to that used in stoichiometric problems in Chapter 2.

Method A:

The balanced chemical reaction is written as: $\text{Ba(OH)}_2 + 2 \text{HCl} \rightarrow \text{BaCl}_2 + 2 \text{H}_2\text{O}$

The net ionic reaction is written as: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

At the stoichiometric point, moles (H⁺) = moles (OH⁻)

or using molarities,

Volume (HCl) x Molarity (H⁺) = Volume (Ba(OH)₂) x Molarity (OH⁻)

Hence, Volume (Ba(OH)₂) = Volume (HCl) x $\frac{\text{Molarity (H}^+)}{\text{Molarity (OH}^-)}$

Volume (HCl) = 0.015 L

Molarity (H⁺) = Molarity (HCl) = 0.175 M (since 1 mol H⁺ per 1 mol HCl)

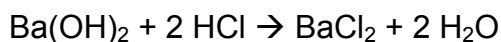
Molarity (OH⁻) = $\frac{\text{moles (OH}^-)}{1\text{L}} = \frac{\text{moles (Ba(OH)}_2)}{1\text{L}} \times \frac{2 \text{ mol (OH}^-)}{1 \text{ mol (Ba(OH)}_2)}$

Molarity (OH⁻) = 2 x Molarity (Ba(OH)₂) = 2 x 0.250 M = 0.500 M

Hence, Volume (Ba(OH)₂) = 0.015 L x $\frac{0.175 \text{ M}}{0.500 \text{ M}}$ = 0.00525 L = 5.25 mL

Method B:

Solve this problem as a normal stoichiometry problem. The balanced chemical reaction is written as:



We want to calculate V(Ba(OH)₂), the volume of Ba(OH)₂ solution necessary to completely react 15.0 mL of 0.175 M HCl. Use the method of conversion factors.

$$V(\text{Ba}(\text{OH})_2) = 0.0150 \text{ L (HCl)} \times \frac{0.175 \text{ mol (HCl)}}{1 \text{ L}} \times \frac{1 \text{ mol (Ba}(\text{OH})_2)}{2 \text{ mol (HCl)}} \\ \times \frac{1 \text{ L}}{0.250 \text{ mol (Ba}(\text{OH})_2)} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

Hence, $V(\text{Ba}(\text{OH})_2) = 0.0150 \times 0.175 \times (1/2) \times (1/0.250) \times (1000) \text{ mL}$

$$V(\text{Ba}(\text{OH})_2) = 5.25 \text{ mL}$$

Note what conversion factors are used:

First, we write the given volume of HCl (0.0150 L)

Second, we transform L of HCl into mol (HCl) using HCl molarity (0.175 M)

Third, we use the mole ratio from the balanced reaction : $\frac{1 \text{ mol (Ba}(\text{OH})_2)}{2 \text{ mol (HCl)}}$

to convert moles (HCl) to moles (Ba(OH)₂).

Fourth, we convert moles (Ba(OH)₂) to liters (Ba(OH)₂) using the molarity of the Ba(OH)₂ solution.

Use whatever method (A or B) works best for you!!

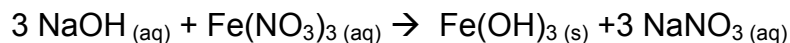
Interactive Problems are included in this section for your practice and understanding.

Sections 3.12 - 3.13: Gravimetric Analysis

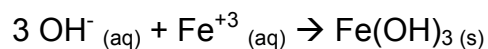
Gravimetric analysis is a chemical analysis method based on the measurement of masses. It can be used, for instance, in combination with precipitation reactions to determine the amount of a species present in solution. We first isolate the precipitate by filtration and drying and subsequently weigh it. Using this mass and stoichiometry relationships we can determine the amount (mass or moles) of species present in solution.

Example 1: Calculate the mass of Fe(OH)₃ formed when 50.0 mL of a 0.300 M NaOH_(aq) is added to 30.0 mL of 0.250 M Fe(NO₃)_{3(aq)} solution.

Step 1: Write the balanced chemical reaction:



Step 2: Write the net ionic equation:



Step 3: Determine the number of moles of each reactant:

$$\text{Moles (OH}^-) = \text{Moles (NaOH)} = 0.300 \text{ M} \times 0.050 \text{ L} = 0.0150 \text{ mol}$$

$$\text{Moles (Fe}^{+3}) = \text{Moles (Fe(OH)}_3) = 0.250 \text{ M} \times 0.030 \text{ L} = 0.00750 \text{ mol}$$

Step 4: Determine the limiting reactant. To do so calculate the number of moles of product formed from the given number of moles of each reactant.

From the hydroxide ion reactant.

$$\text{Moles (Fe(OH)}_3) = 0.0150 \text{ mol (OH}^-) \times \frac{1 \text{ mol (Fe(OH)}_3)}{3 \text{ mol (OH}^-)} = 0.00500 \text{ mol}$$

From the iron(III) ion reactant.

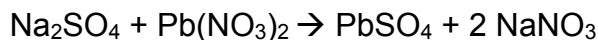
$$\text{Moles (Fe(OH)}_3) = 0.00750 \text{ mol (Fe}^{+3}) \times \frac{1 \text{ mol (Fe(OH)}_3)}{1 \text{ mol (Fe}^{+3})} = 0.00750 \text{ mol}$$

Hence, the limiting reactant is the hydroxide and the theoretical product is 0.00500 mol (Fe(OH)₃). The molar mass of iron(III) hydroxide is 106.9 g/mol. Hence, the mass of iron hydroxide collected should be 0.535 g.

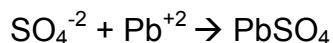
Example 2: A liter sample of polluted water is analyzed for lead, believed to be in the lead(II) ionic form. This is done by adding excess sodium sulfate to the polluted water sample. The mass of PbSO₄ produced is 300.0 mg. Calculate the mass of Pb in one liter solution.

All the lead in water reacts with the sulfate ions to form a PbSO₄ precipitate.

The balanced chemical reaction is written as:



The net ionic reaction is written as:



$$\text{Molar mass (PbSO}_4) = 303.3 \text{ g/mol}$$

$$\text{Molar mass (Pb)} = 207.2 \text{ g/mol}$$

Using the method of conversion factors, we transform the mass of PbSO_4 into the mass of Pb.

$$\text{Mass (Pb)} = 300.0 \text{ mg PbSO}_4 \times \frac{1 \text{ g (PbSO}_4\text{)}}{1000 \text{ mg (PbSO}_4\text{)}} \times \frac{1 \text{ mol (PbSO}_4\text{)}}{303.3 \text{ g (PbSO}_4\text{)}} \\ \times \frac{1 \text{ mol (Pb)}}{1 \text{ mol (PbSO}_4\text{)}} \times \frac{207.2 \text{ g (Pb)}}{1 \text{ mol (Pb)}}$$

$$\text{Mass (Pb)} = 0.2049 \text{ g Pb} = 204.9 \text{ mg Pb}$$

In Section 3.13, practice Interactive Problems.

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