

Factors that Affect the Rate of Dissolving and Solubility

Dissolving

One very important property of a solution is the rate of **dissolving**, or how quickly a solute dissolves in a solvent. When dissolving occurs, there is **NO CHEMICAL REACTION** involved. Therefore, the solute and solvent can be separated using **physical** properties such as **boiling points** or **filtration**.

The rate at which a solute dissolves depends on a number of factors:

i) **Temperature**

Increasing temperature increases the **kinetic energy** (energy of motion) of the molecules, which increases the frequencies of **collisions** and the rate of dissolving.

ii) **Agitation**

Stirring/shaking brings **fresh solvent** into contact with **undissolved solute**, increasing **collisions** and the rate of dissolving.

iii) **Particle Size**

Crushing solute into smaller pieces increases the **surface area** that is in contact with **solvent**, thus increasing the rate of dissolving.

The Dissolving Process

Whether or not a solute dissolves and to what extent depends on the forces of attraction between:

- Solute particles
- Solvent particles
- Solute and solvent particles

When the forces of attraction between **different** particles in a mixture are **stronger** than the forces of attraction between **like** particles in the mixture, a solution forms. The strength of each attraction influences the **solubility**, or the amount of solute that dissolves in a solvent.

The dissolving process can be broken down into three key steps:

1. The **forces** holding the **solute** together must be broken (**requires energy**)

Ionic compounds - the forces holding the ions together must be broken

Covalent molecules - the forces holding molecules together must be broken

2. The **intermolecular** forces (between particles) holding the **solvent** together must be broken (**requires energy**)
3. Solute and solvent **attract** (**release energy**) and the molecules of solute fill in the spaces between solvent molecules.

Note: *Dissolving is more likely to occur if the energy required (steps 1 and 2) is less than the energy released (step 3).*

Polar and Non-Polar Substances

In general, we can follow the rule of "**like dissolves like**" when trying to predict the solubility of different particles. **Ionic** solutes and **polar covalent** solutes dissolve in **polar solvents** and **non-polar solutes** dissolve in **non-polar solvents**.

Remember, you can use the difference in electronegativities (ΔEN) to predict if a compound is ionic, polar or non-polar.

There are a few possible forces that act between particles, which helps to explain the "like dissolves like" trend:

Dipole-Dipole Attractions - the attraction between the **opposite partial charges** on two different **polar** molecules.

Ion-Dipole Attractions - the attractive forces between an **ion** and a **polar** molecule. Ions possess a **full charge** and are therefore attracted to the **partial charge** on the polar molecules.

When ions are present in an **aqueous** solution, each ion is **hydrated**. This means that water molecules surround the ion. Hydrated ions can conduct electricity and are referred to as **electrolytes**.

Solubility

Solubility describes the **amount** of **solute** that can be dissolved in a given **amount** of **solvent** under given conditions.

A solute is described as **soluble** in a particular solvent if its solubility is **greater** than **1 g per 100 mL**.

A solute is described as **insoluble** in a particular solvent if its solubility is **less** than **0.1 g per 100 mL**.

Substances with solubility between these limits are called **slightly soluble**.

Factors affecting solubility include:

i) Molecular Size

Small molecules tend to be more soluble than **large** ones.

ii) Temperature

Affects the solubility of gases and solids in liquids.

For gases in liquids: as temp ↑ solubility ↓

For solids in liquids: as temp ↑ solubility ↑

*A **solubility curve** (graph) describes how much solute can be dissolved in a given solvent at a certain temperature.*

iii) Pressure

Affects the solubility of gases in liquids.

As pressure ↑ solubility ↑

Concentration of Solutions

Concentration is defined as the amount of **solute** per quantity of **solvent**.

The concentration of a solution can be calculated. The approach for each calculation varies, depending on the **type** of solution.

1. Calculation as Mass/Volume (m/v) Percent

Gives the mass of solute dissolved in a volume of solution, expressed as a percent.

$$\text{Mass/Volume \%} = \frac{m_{\text{solute}} (\text{g})}{V_{\text{solution}} (\text{mL})} \times 100$$

Examples:

2.00 mL of distilled water is added to 4.00 g of a powdered drug. The final volume of the solution is 3.00 mL. Calculate the percent m/v and then express the drug concentration in g/100 mL.

$$\checkmark_1 \text{ m/v \%} = \frac{m_{\text{solute}}}{V_{\text{solution}}} \times 100$$

$$= \frac{4.00 \text{ g}}{3.00 \text{ mL}} \times 100$$
$$\checkmark_2 = 133\%$$

\checkmark_3 the solution is 133% m/v. If 100 mL would have 133 g of drug.

What mass of a drug is required to make a 2.0 L solution if the recommended concentration is 1.7%?

$$\checkmark_1 m_{\text{solute}} = \frac{(\text{m/v \%}) (V_{\text{solution}})}{100}$$
$$= \frac{(1.7)(2000)}{100}$$

$$\checkmark_2 = 34 \text{ g}$$

✓ 300 34g of the drug is required

2. Calculation as Mass/Mass (m/m) Percent

Gives the mass of solute divided by the mass of solution, expressed as a percent.

$$\text{Mass/Mass \%} = \frac{m_{\text{solute}}(\text{g})}{m_{\text{solution}}(\text{g})} \times 100$$

Example:

An aqueous solution of calcium chloride has a mass of 23.47 g. The solvent was evaporated and the residue has a mass of 4.58 g. Calculate the m/m % of calcium chloride in the solution. How many grams of calcium chloride would be present in a 100 g sample?

✓ 1 $m/m \% = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100$
 $= \frac{4.58\text{g}}{23.47} \times 100$
✓ 2 $= 19.5\%$

✓ 3. the solution is 19.5% m/m
✓ 4. 100 g sample would have 19.5g of CaCl_2 .

3. Concentration as Volume/Volume (v/v) Percent

Gives the volume of solute divided by the volume of solution, expressed as a percent.

$$\text{Volume/Volume \%} = \frac{V_{\text{solute}}(\text{mL})}{V_{\text{solution}}(\text{mL})} \times 100$$

Example:

Rubbing alcohol is sold as a 70% (v/v) solution. What volume of alcohol is used to make 500 mL of rubbing alcohol?

✓ 1 $V_{\text{solute}} = \frac{(v/v \%)(V_{\text{solution}})}{100}$
 $= \frac{(70)(500)}{100}$
✓ 2 $= 350$

✓ 30% 350 mL of alcohol is needed

4. Parts per Million (ppm) and Parts per Billion (ppb)

Describes the concentration of very small quantities. Usually expressed in terms of mass/mass relationships.

$$\text{ppm} = \frac{m_{\text{solute}}(\text{g})}{m_{\text{solution}}(\text{g})} \times 10^6$$

$$\text{ppb} = \frac{m_{\text{solute}}(\text{g})}{m_{\text{solution}}(\text{g})} \times 10^9$$

Note: Your final answer does not refer to the number of particles per million or billion, but rather the mass of solute compared to the mass of solution.

Example:

A shipment of oranges is returned if it contains more than 25 ppb of mould. A company received 20 000 kg of oranges. What is the maximum mass of mould allowed before the shipment should be sent back?

$$\begin{aligned} m_{\text{solute}} &= \frac{(\text{ppb})(m_{\text{solution}})}{10^9} \\ &= \frac{(25)(20\,000\,000)}{10^9} \\ &= 0.5\text{ g} \end{aligned}$$

∴ 0.5 g of mould would be present.

Molar Concentration

Molarity (C) is the number of **moles** of **solute** dissolved per **litre** of **solvent**.

The equation we use to calculate molar concentration is:



Where, C = molarity (mol/L or M)
 n = amount of solute in moles
 V = volume of solution in Litres

Examples:

What is the molar concentration of 1.20 g of NaNO_3 in 80.0 mL of solution?

$$1 \quad n = \frac{m}{M}$$

$$= \frac{1.20 \text{ g}}{84.994 \text{ g/mol}}$$

$$2 \quad = 0.014 \text{ mol}$$

$$3 \quad C = \frac{n}{V}$$

$$= \frac{0.014 \text{ mol}}{0.080 \text{ L}}$$

$$4 \quad = 0.175 \text{ M}$$

5 $\therefore [\text{NaNO}_3]$ is 0.175 M mol/L

How many grams of potassium hydroxide will be required to prepare 650 mL of 0.430 M solution? KOH

$$1 \quad n = CV$$

$$= (0.430)(0.650)$$

$$2 \quad = 0.2795 \text{ mol}$$

$$3 \quad m = nM$$

$$= (0.2795)(56.105)$$

$$4 \quad = 15.68 \text{ g}$$

5 $\therefore 15.68 \text{ g}$ are required.

Preparing Solutions and Dilutions

A **standard (stock)** solution is a solution with **known concentration**.

There are 2 ways to prepare a solution:

- Dissolve a desired amount of solute in a set amount of solvent
- Dilute a standard solution

A useful tool in preparing solutions is a **volumetric flask** → a pear-shaped glass with a flat bottom and a long neck. Volumetric flasks provide are very accurate tools for measuring volumes.

To prepare a solution you should perform the following steps:

- ~~X~~ 1. Determine the **mass (g) of solute** required to make the desired **volume (L)** and **concentration (mol/L)** of solution.
$$\textcircled{A} \quad n = CV \quad \textcircled{B} \quad m = nM$$
- ~~X~~ 2. Measure out and dissolve the **solute** in approximately $\frac{1}{2}$ the **total volume of solvent**.
- ~~X~~ 3. Raise the **volume of solution** to the desired total volume by adding more **solvent**.

Diluting is a process that makes a solution that is less concentrated. This can be done by:

- Adding more solvent (more common)
- Removing solute

Dilution Calculations:

Step 1: Find the number of **moles** you need

Step 2: Find the **volume of stock solution** you need

Step 3: Top up with **distilled water**

in diluted solution

$$n = CV$$
$$V = \frac{n}{C}$$

$$V_{H_2O} = V_{solution} - V_{stock}$$

Example #1

How do you make a 1.50 L solutions of NaCl with a concentration of 6.00 M from a stock solution with a concentration of 15.0 M?

$$\begin{aligned} \textcircled{1} n_{\text{diluted}} &= C V \\ &= (6.00 \text{ M})(1.5 \text{ L}) \\ &= 9.0 \text{ mol} \end{aligned}$$

$$\textcircled{2} V = \frac{n}{C} = \frac{9.00 \text{ mol}}{15.0 \text{ M}}$$

$$\begin{aligned} \textcircled{3} V_{\text{H}_2\text{O}} &= V_{\text{solution}} - V_{\text{stock}} \\ &= 1.5 - 0.6 \end{aligned}$$

$$\begin{aligned} &= 0.60 \text{ L} \\ &= 0.6 \text{ L of stock} \\ &0.9 \text{ L of H}_2\text{O} \end{aligned}$$

Alternatively we can perform dilution calculations using the following equation:

$$= 0.9 \text{ L}$$

Where,
 C_1 = Molarity of stock solution
 V_1 = Volume of stock solution
 C_2 = Molarity of "new" solution
 V_2 = Volume of "new" solution

$$C_1 V_1 = C_2 V_2$$

Lets try this equation to solve the previous example!!!

$$V_1 = \frac{C_2 V_2}{C_1}$$

$$= \frac{(6)(1.5)}{(15)}$$

$$\rightarrow V_1 = 0.60 \text{ L}$$

only ↓ to be used
for dilutions!!
m m
mm

Example #2

If 85.0 mL of 0.950 M sodium sulfate solution was used to prepare 200 mL of a dilute sodium sulfate solution, what is the new concentration made?

Reactions in an Aqueous Solution- Ionic Equations

When an ionic compound is placed in water, most will **dissolve**, which means they are **soluble** in water. Some ionic compounds will remain as a **solid** and are **insoluble**.

If an ionic compound dissolves in water, it means that the compound is temporarily splitting apart into its **ions**. This process is referred to as an ionic compound **dissociating**. This is NOT a **chemical reaction** and the ionic compound will readily **re-crystallize** when removed from the water source.

Double displacement reactions occur in water, and are a direct result of ionic compounds dissociating into their ions. Recall that a double displacement reaction will only occur if **water**, **gas** or a **precipitate** forms.

We can show the step-by-step process of a double displacement reaction by writing out an **ionic equation**. There are several different components to an ionic equation.

Term	Definition
Total Ionic Equation	Equation that illustrates all soluble ionic compounds in their ionic form <i>(aq) → break into ions</i>
Net Ionic Equation	Equation that depicts only the ions that are involved in a chemical reaction <i>ions that make (s)</i>
Spectator Ion	An ion that is present during a chemical reaction but does not participate in the reaction.

Example #1

Word Equation	Silver nitrate reacts with sodium chloride
Balanced Equation	$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
Total Ionic Equation	$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$
Net Ionic Equation	$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
Spectator Ions	$Na^+(aq)$ $NO_3^-(aq)$

Example #2

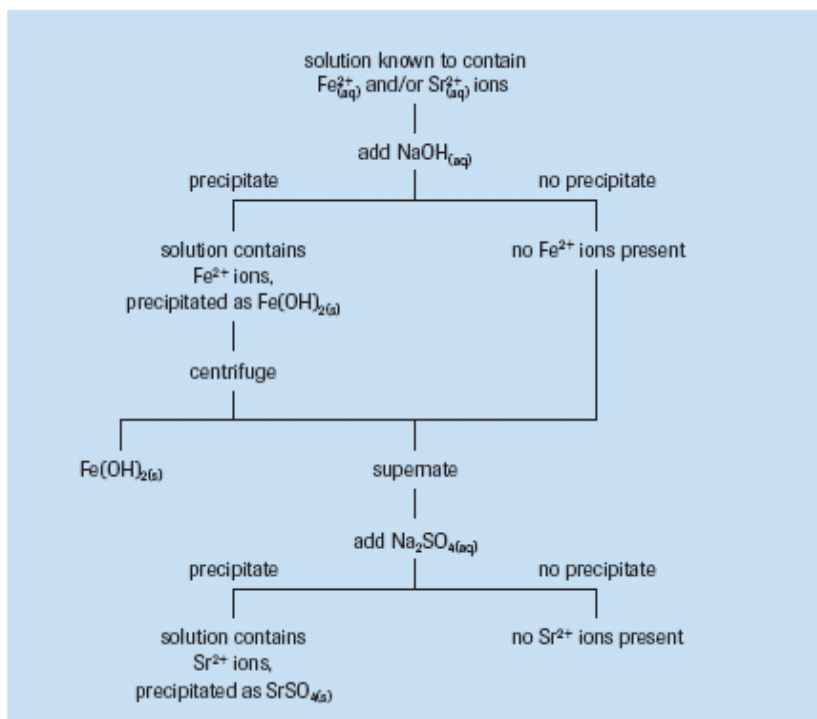
Word Equation	Calcium bromide reacts with lithium chlorate
Balanced Equation	$\text{CaBr}_2(\text{aq}) + 2\text{LiClO}_3(\text{aq}) \rightarrow \text{Ca}(\text{ClO}_3)_2(\text{s}) + 2\text{LiBr}(\text{aq})$
Total Ionic Equation	$\text{Ca}^{2+}(\text{aq}) + 2\text{Br}^{-1}(\text{aq}) + 2\text{Li}^{+1}(\text{aq}) + 2\text{ClO}_3^{-1}(\text{aq}) \rightarrow \text{Ca}(\text{ClO}_3)_2(\text{s}) + 2\text{Li}^{+1}(\text{aq}) + 2\text{Br}^{-1}(\text{aq})$
Net Ionic Equation	$\text{Ca}^{2+}(\text{aq}) + 2\text{ClO}_3^{-1}(\text{aq}) \rightarrow \text{Ca}(\text{ClO}_3)_2(\text{s})$
Spectator Ions	$\text{Br}^{-1}(\text{aq})$ and $\text{Li}^{+1}(\text{aq})$

Precipitate reactions can be used to generate a precipitation profile for known ions, which can be used to identify ions in solution.

	CO_3^{-2}	OH^{-1}	SO_4^{-2}
Ca^{+2}			
Ba^{+2}			
Unknown	ppt	ppt	ppt

The unknown precipitation profile matches that of _____.

Flow charts can also be used to illustrate what ions may be added to a solution containing many ions to allow for individual separation of ions by precipitation.



When attempting to separate ions from a solution a piece of lab equipment called a **centrifuge** can be used. A centrifuge **spins** at very high speeds to separate different particles from each other based on their **densities**.

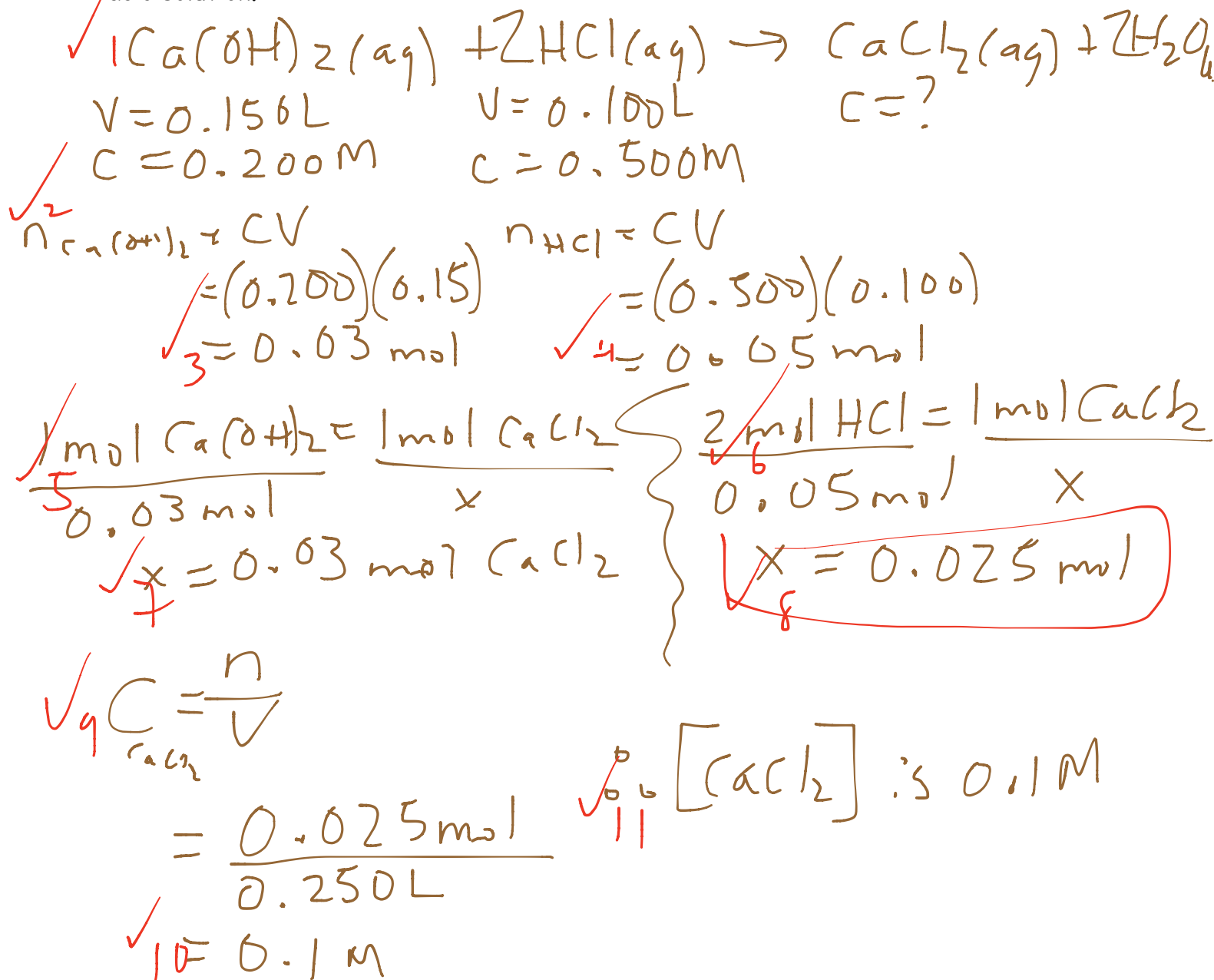
The part of the solution that does not settle to the bottom of the centrifuge is called the **supernatant**.

Solution Stoichiometry

Recall that stoichiometry involves calculating the amounts of reactants and products in chemical reactions using a balanced chemical equation. Previously you learned how to calculate the amount of atoms, particles or mass of a compound using the stoichiometry strategies. You can apply these same skills when approaching calculations involving solutions, with the addition of a few additional steps.

Example #1

Calculate the concentration of calcium chloride in a solution made by mixing 150 mL of a 0.200 M calcium hydroxide solution with 100 mL of a 0.500 M hydrochloric acid solution.

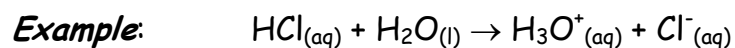


Example #2

Suppose you want to remove the barium ions from 120 mL of 0.050 00 M aqueous barium nitrate solution. What is the minimum mass of sodium carbonate that you should add?

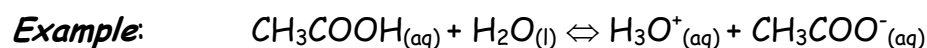
Strong and Weak Acids and Bases

Strong acid - an acid that dissociates completely into ions in water.



When hydrogen chloride molecules enter an aqueous solution, 100% of the hydrogen chloride molecules dissociate. As a result the solution contains the same percent of H^+ ions (in the form of H_3O^+) and Cl^- ions: 100%

Weak acid - an acid that dissociates very slightly in a water solution.



On average, only about 1% of the acetic acid molecules dissociate at any given moment.

Notice that the arrow used in the dissociation of a weak acid points in both directions. This indicates that the reaction is **reversible**. The products of the reaction will also react to produce the original reactants.

Some useful terms:

Term	Definition	Example
Monoprotic acid	Contains only a single hydrogen ion that can dissociate	HCl
Diprotic acid	Contains two hydrogen ions that dissociate to form two anions	H_2SO_4
Triprotic acid	Contains three hydrogen ions that dissociate to form three anions	H_3PO_4

In both diprotic and triprotic acids, the dissociation of the first hydrogen ion will result in a stronger acid than the acid formed by the second and third dissociation.

Strong base - a base that dissociates completely into ions in water.

Examples: NaOH, MgO

Weak base - a base that dissociates very slightly in a water solution.

Example: NH_3

Concentration of an Acid or Base

Recall that when in solution, acids and bases dissociate into ions. When you determine the concentration of hydrogen ions in solution (amount of H^+ ions/ total solution volume) you are determining the pH of that particular solution. pH stands for, "the power of hydrogen". The pH of a substance can be determined a number of different ways, such as with the use of pH paper, an electronic pH meter or mathematically using the following formulas:

Square brackets [] around a chemical formula represents, "the concentration of"

Examples:

What is the pH of a solution with a $[\text{H}_3\text{O}^+]$ of 1.0×10^{-5} ?

Gastric juice has a pH of 1.5, what is the $[\text{H}_3\text{O}^+]$?

The relative concentration of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ ions are as follows:

Acidic
Neutral
Basic

A pH scale is a convenient way to relate the pH of a solution to its degree of acidity/alkalinity.

The pH scale ranges from 1 to 14 and each pH unit represents a factor of 10.

Examples:

A change in pH from 3 to 8 is a(n) _____ increase/decrease in $[\text{H}_3\text{O}^+]$

A change in pH from 11 to 2 is a(n) _____ increase/decrease in $[\text{H}_3\text{O}^+]$

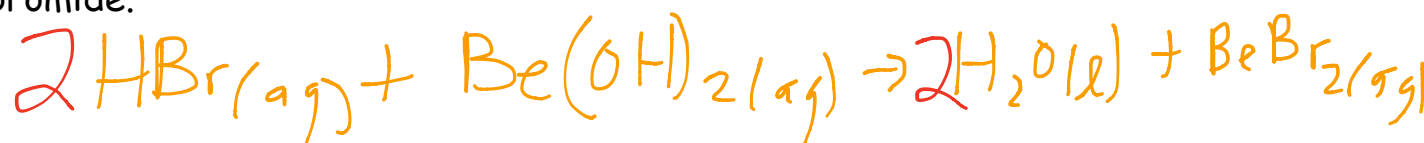
Neutralization Reactions

Neutralization occurs when **hydroxide ions** (Arrhenius base) and **hydrogen ions** (acid) are mixed to make **water** and a **salt**. The general word equation is:

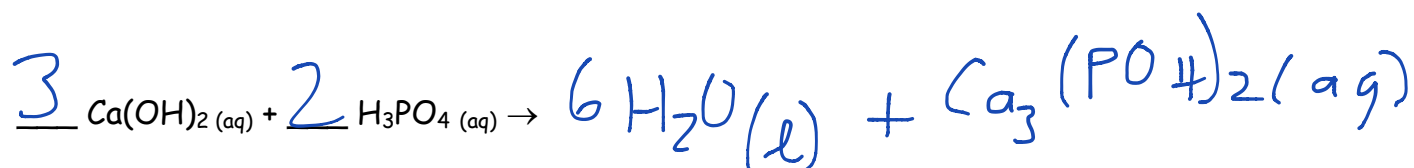
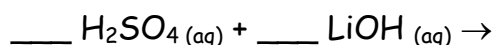


Example:

Aqueous solutions of hydrobromic acid and beryllium hydroxide undergo a neutralization reaction to produce water and beryllium bromide.



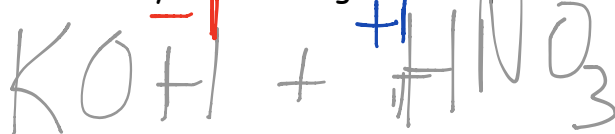
Complete the following equations:



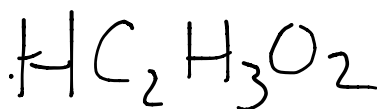
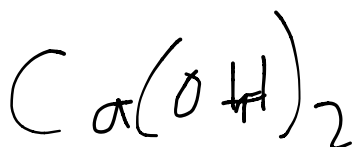
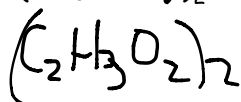
Which acid and base would you react together to produce the following salts:

⁺¹

i) KNO_3



ii) $\text{Ca}(\text{CH}_3\text{COO})_2$



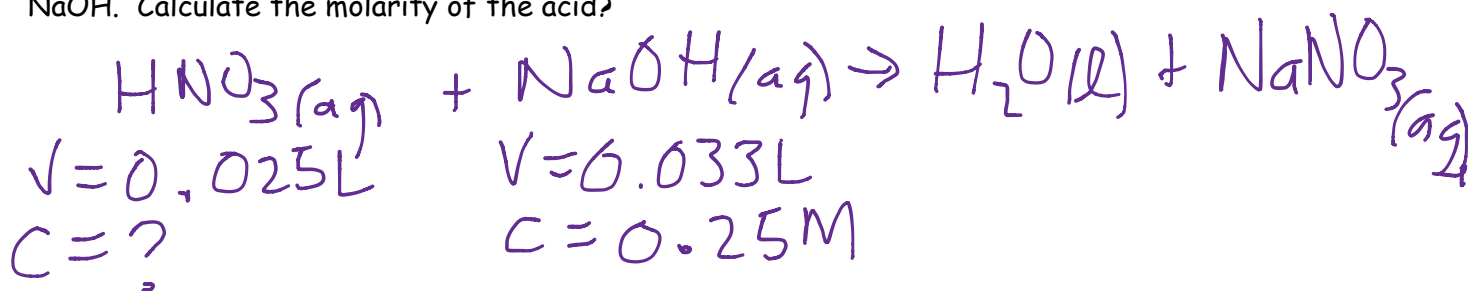
Acid-Base Titrations

A "titration" refers to a technique that involves the careful measuring of the **volume** of one solution required to completely react with a **known quantity** of another.

In an acid-base titration, measuring the volume of a **base** (of **known molarity**) allows us to determine the **molarity** of the **acid**. In this case an **acid-base indicator** is used to indicate when the neutralization reaction is complete. **Phenolphthalein** is the most common indicator used. It will be **clear** when added to the **acid**; neutralization occurs at the first signs of the solution **turning** and **remaining** a **faint pink colour**.

Example #1

In an acid-base titration, 25.00 mL of HNO_3 is required to neutralize 33.00 mL of 0.25 M NaOH . Calculate the molarity of the acid?



$$\begin{aligned} n_{\text{NaOH}} &= CV \\ &= (0.25)(0.033) \\ &= 0.00825 \text{ mol} \end{aligned}$$

$$\frac{1 \text{ mol NaOH}}{0.00825 \text{ mol}} = \frac{1 \text{ mol HNO}_3}{X}$$

$$X = 0.00825 \text{ mol of HNO}_3$$

$$\begin{aligned} C_{\text{HNO}_3} &= \frac{n}{V} \\ &= \frac{(0.00825)}{0.025} \\ &= 0.33 \text{ M} \end{aligned}$$

$\therefore [\text{HNO}_3]$ is 0.33M

Example #2

In an acid-base titration, 43.00 mL of 0.30 M KOH is required to neutralize 10.00 mL of H_2SO_4 . Calculate the molarity of the acid?