


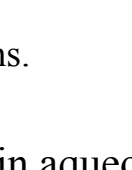
## Chapter 4

### Reactions in Aqueous Solutions



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## Preview

- Aqueous solutions and their chemistry.
- Various types of reactions.
  - Precipitation reactions.
  - Acid-base reactions.
  - Oxidation-reduction reactions.
- The concept of molarity.
- Stoichiometry of reactions in aqueous solutions.

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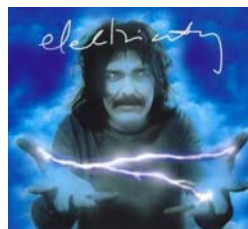
## Aqueous Solutions

- **Solution** is a **solute** (for example NaCl, NaOH or ethanol) dissolved in a **solvent**.  
(When the solvent is H<sub>2</sub>O, => **Aqueous Solutions**).

- What common examples of solutions you can think of?

- Coffee, Tea, Sea, ...

Can these types of solutions conduct electricity??



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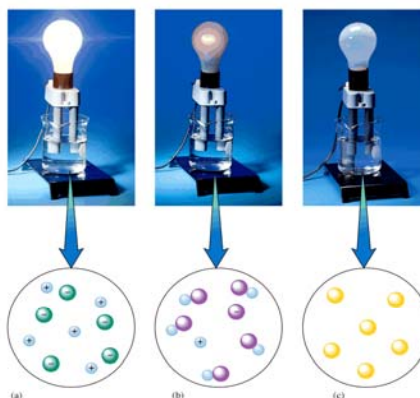
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## General Properties of Aqueous Solutions

- Aqueous solutions can conduct electric current at different efficiencies (Arrhenius postulate in 1880s).

- Strong electrolytes. **Many ions** present in solution (NaCl).
- Weak electrolytes. **Few ions** present in solution (Acetic acid).
- Nonelectrolytes. **No ions** present in solution (sugar).

- Lighting a bulb with aqueous solutions.



Free ions work as charge carriers in solutions to complete the circuit.

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## Strong Electrolytes

- **Strong electrolytes** are completely dissolved in water to yield a solution that conducts electricity efficiently.

- Salts (NaCl, KI). *Hydration process*
- Strong acids (HCl, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, HClO<sub>4</sub>).
- Strong bases (NaOH, KOH).

- NaCl salt



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## Strong Electrolytes

- **Strong Acid** : produces H<sup>+</sup> ions (protons) and it is completely ionized when dissolved in water.

TABLE 4.1

The Strong Acids

Acid	Ionization Equation
Hydrochloric acid	$\text{HCl}(aq) \longrightarrow \text{H}^+(aq) + \text{Cl}^-(aq)$
Hydrobromic acid	$\text{HBr}(aq) \longrightarrow \text{H}^+(aq) + \text{Br}^-(aq)$
Hydroiodic acid	$\text{HI}(aq) \longrightarrow \text{H}^+(aq) + \text{I}^-(aq)$
Nitric acid	$\text{HNO}_3(aq) \longrightarrow \text{H}^+(aq) + \text{NO}_3^-(aq)$
Chloric acid	$\text{HClO}_3(aq) \longrightarrow \text{H}^+(aq) + \text{ClO}_3^-(aq)$
Perchloric acid	$\text{HClO}_4(aq) \longrightarrow \text{H}^+(aq) + \text{ClO}_4^-(aq)$
Sulfuric acid*	$\text{H}_2\text{SO}_4(aq) \longrightarrow \text{H}^+(aq) + \text{HSO}_4^-(aq)$
	$\text{HSO}_4^-(aq) \rightleftharpoons \text{H}^+(aq) + \text{SO}_4^{2-}(aq)$

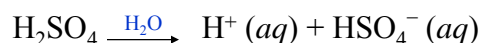
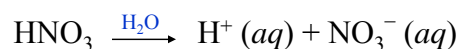
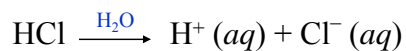
\*Note that although each sulfuric acid molecule has two ionizable hydrogen atoms, it only undergoes the first ionization completely, effectively producing one H<sup>+</sup> ion and one HSO<sub>4</sub><sup>-</sup> ion per H<sub>2</sub>SO<sub>4</sub> molecule. The second ionization happens only to a very small extent.

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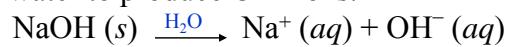
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## Strong Electrolytes

- When strong acids are put in water, they are completely ionized producing **protons** ( $\text{H}^+$  ions) and anions.



- Strong bases completely dissolve in water to produce  $\text{OH}^-$  ions.



Almost no HCl units are present.



Almost no NaOH units are present.

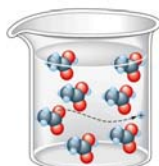


## Nonelectrolytes

- Nonelectrolytes** can dissolve in water but don't produce ions (no electrical conductivity), like ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) and sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ).
- Water is a nonelectrolyte and doesn't produce any ions.



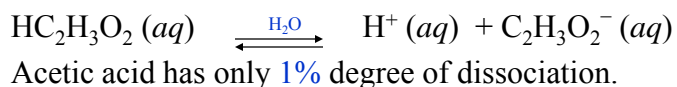
## Weak Electrolytes



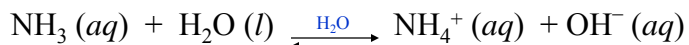
● Hydrogen  
● Oxygen  
● Carbon

- **Weak electrolytes** have a small degree of ionizations and exist *predominantly* as molecules rather than ions.

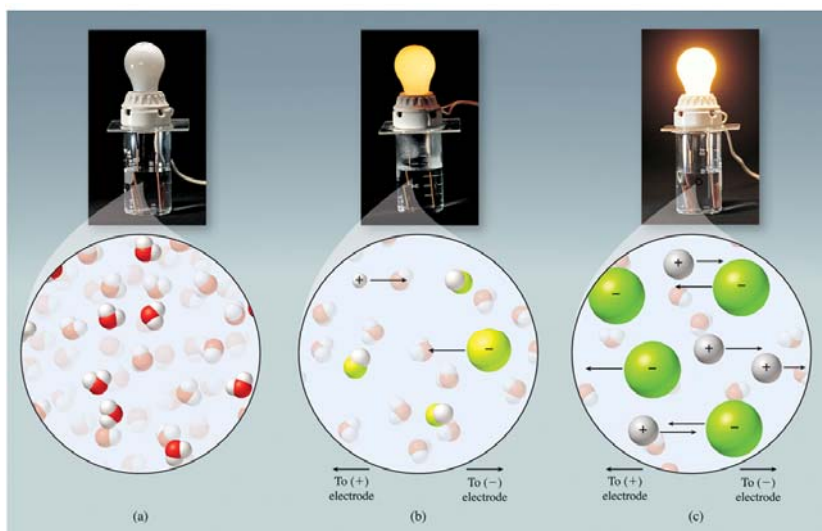
- **Weak Acids:** they are *very slightly* ionized in water producing a few number of protons ( $\text{H}^+$ ).



- **Weak Bases:** they *very slightly* dissolve in water producing a few number of hydroxide ions ( $\text{OH}^-$ ).



## Electrolytes and Nonelectrolytes



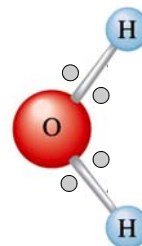
## Types of Chemical Reactions in Solutions

- Types of chemical reactions in solutions are generally:
  - Precipitation reactions.
  - Acid-base reactions.
  - Oxidation-reduction reactions.



## The Hydration Process in Aqueous Solutions

- **Why do salt, sugar, and other solid dissolve in water??**
- Water is a bent molecule (not linear).
- O-H bonds are covalent (O and H atoms share electrons).
- Because the oxygen atom has a greater attraction for electrons, shared electrons tend to spend more time closer to the oxygen atom than to either of the hydrogen atoms.
- In  $\text{H}_2\text{O}$ , oxygen is partially negative ( $\delta^-$ ) and hydrogens are partially positive ( $\delta^+$ ), giving rise to a *polar molecule*.
- $\delta$  means less than one unit of charge.



Chapter 4 Section 2

## The Hydration Process

When ionic substances dissolve in water, they break up (*dissociate*) into individual cations and anions.

$$\text{NaCl}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$$

Hydrations causes the salt to dissociate (fall apart).

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Chapter 4 Section 2

## Precipitation Reactions

The abbreviation (*aq*) means that the ions of the compound are separate and moving around independently in water.

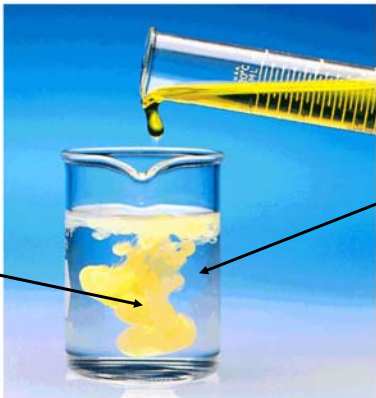
$\begin{array}{c} \text{K}^+ \\ \text{CrO}_4^{--} \\ \text{K}^+ \\ \text{K}^+ \\ \text{CrO}_4^{--} \\ \text{K}^+ \\ \text{K}^+ \text{K}^+ \\ \text{CrO}_4^{--} \end{array}$	+	$\begin{array}{c} \text{Ba}^{++} \\ \text{NO}_3^- \\ \text{NO}_3^- \\ \text{Ba}^{++} \\ \text{NO}_3^- \\ \text{NO}_3^- \\ \text{Ba}^{++} \\ \text{NO}_3^- \\ \text{NO}_3^- \end{array}$	→	$\begin{array}{c} \text{CrO}_4^{--} \\ \text{Ba}^{++} \quad \text{NO}_3^- \\ \text{NO}_3^- \quad \text{K}^+ \\ \text{K}^+ \quad \text{Ba}^{++} \quad \text{NO}_3^- \\ \text{NO}_3^- \quad \text{Ba}^{++} \\ \text{CrO}_4^{--} \\ \text{NO}_3^- \quad \text{K}^+ \quad \text{K}^+ \\ \text{NO}_3^- \quad \text{K}^+ \quad \text{CrO}_4^{--} \end{array}$
$\text{K}_2\text{CrO}_4(aq)$		$\text{Ba}(\text{NO}_3)_2(aq)$		

Do these four types of ions remain as ions or some new compound "precipitation" could form?

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## Precipitation Reactions

- When an insoluble substance is produced from mixing two solutions, the reaction is said to be a **precipitation reaction** and the insoluble substance is called **precipitate**.



$K_2CrO_4 (aq)$   
 $Ba(NO_3)_2 (aq)$

$2K^+(aq) + CrO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq)$

Products

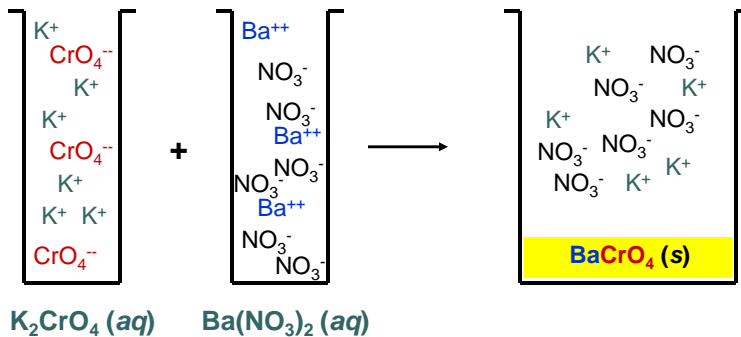
$BaCrO_4$  OR  $2KNO_3$

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## Precipitation Reactions



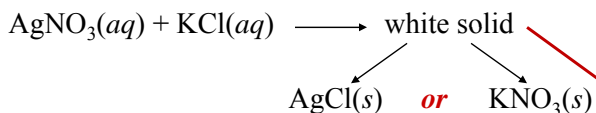
How can you know which one will precipitate and which one will not?





## Precipitation Reactions

Another example of precipitation reactions:



How can you know which one will precipitate and which one will not?

We need to make use of solubility rules of salts in water:

- Soluble.
- Insoluble (Not soluble).

## Solubility Rules for Salts in Water

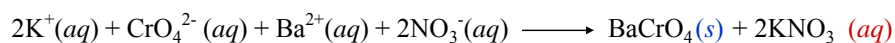
**TABLE 4.2** Solubility Guidelines: Soluble Compounds

	Water-Soluble Compounds	Insoluble Exceptions
1	Compounds containing an alkali metal cation ( $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ ) or the ammonium ion ( $\text{NH}_4^+$ )	
2	Compounds containing the nitrate ion ( $\text{NO}_3^-$ ), acetate ion ( $\text{C}_2\text{H}_3\text{O}_2^-$ ), or chlorate ion ( $\text{ClO}_3^-$ )	
3	Compounds containing the chloride ion ( $\text{Cl}^-$ ), bromide ion ( $\text{Br}^-$ ), or iodide ion ( $\text{I}^-$ )	Compounds containing $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , or $\text{Pb}^{2+}$
4	Compounds containing the sulfate ion ( $\text{SO}_4^{2-}$ )	Compounds containing $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , $\text{Pb}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , or $\text{Ba}^{2+}$

**TABLE 4.3** Solubility Guidelines: Insoluble Compounds

	Water-Insoluble Compounds	Soluble Exceptions
5	Compounds containing the carbonate ion ( $\text{CO}_3^{2-}$ ), phosphate ion ( $\text{PO}_4^{3-}$ ), chromate ion ( $\text{CrO}_4^{2-}$ ), or sulfide ion ( $\text{S}^{2-}$ )	Compounds containing $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ , or $\text{NH}_4^+$
6	Compounds containing the hydroxide ion ( $\text{OH}^-$ )	Compounds containing $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ , or $\text{Ba}^{2+}$

## Solubility Rules for Salts in Water

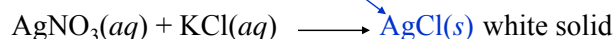


Rule 5 indicates that it is not soluble

Rules 1 & 2 indicate that it is soluble

Rule 3 (exception) indicates that  $\text{AgCl}$  is not soluble

$\text{K}^+$  and  $\text{NO}_3^-$  are called spectator ions



### Sample Exercise

Using the solubility rules, predict what will happen when the following pairs of solutions are mixed.

- $\text{KNO}_3(\text{aq})$  &  $\text{BaCl}_2(\text{aq})$
- $\text{Na}_2\text{SO}_4(\text{aq})$  &  $\text{Pb}(\text{NO}_3)_2(\text{aq})$
- $\text{KOH}(\text{aq})$  &  $\text{Fe}(\text{NO}_3)_3(\text{aq})$

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## Solubility Rules (Exercises)

Soluble Compounds

1

Compounds containing an alkali metal cation ( $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ ) or the ammonium ion ( $\text{NH}_4^+$ )

Exceptions

2

Compounds containing the nitrate ion ( $\text{NO}_3^-$ ), acetate ion ( $\text{C}_2\text{H}_3\text{O}_2^-$ ), or chlorate ion ( $\text{ClO}_3^-$ )

3

Compounds containing the chloride ion ( $\text{Cl}^-$ ), bromide ion ( $\text{Br}^-$ ), or iodide ion ( $\text{I}^-$ )

Compounds containing  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$ , or  $\text{Pb}^{2+}$

4

Compounds containing the sulfate ion ( $\text{SO}_4^{2-}$ )

Compounds containing  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ , or  $\text{Ba}^{2+}$

Insoluble Compounds

5

Compounds containing the carbonate ion ( $\text{CO}_3^{2-}$ ), phosphate ion ( $\text{PO}_4^{3-}$ ), chromate ion ( $\text{CrO}_4^{2-}$ ), or sulfide ion ( $\text{S}^{2-}$ )

Compounds containing  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ , or  $\text{NH}_4^+$

6

Compounds containing the hydroxide ion ( $\text{OH}^-$ )

Compounds containing  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ , or  $\text{Ba}^{2+}$

### Sample Exercise

Using the solubility rules above, predict what will happen when the following pairs of solutions are mixed.

- $\text{Na}_2\text{SO}_4(\text{aq})$  &  $\text{Pb}(\text{NO}_3)_2(\text{aq})$   $\text{PbSO}_4$  solid forms *Rule # 4*
- $\text{KNO}_3(\text{aq})$  &  $\text{BaCl}_2(\text{aq})$  No precipitation forms.
- $\text{KOH}(\text{aq})$  &  $\text{Fe}(\text{NO}_3)_3(\text{aq})$   $\text{Fe}(\text{OH})_3$  solid forms *Rule # 6*

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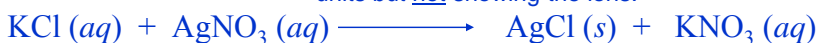
## Describing Reactions in Solution

### Example:

Aqueous potassium chloride is added to aqueous silver nitrate.

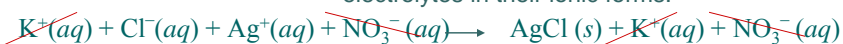
#### Molecular equation:

It shows reactants and products as formula units but not showing the ions.



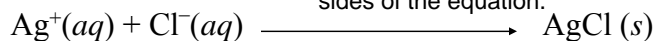
#### Ionic equation:

It shows all substances that are strong electrolytes in their ionic forms.



#### Net ionic equation:

It excludes the spectator ions from the two sides of the equation.



## Acid-Base Reactions: Introduction

### o Acids

- Have sour (acidic) taste.
  - Acetic acid in vinegar
  - Citric acid in fruits.
  - Hydrochloric acid of stomach reflux.
  - Carbonic acid in soft drinks.
  - Ascorbic acid is vitamin C.
- Concentrated acids are very dangerous:
  - can dissolve metals and form hydrogen gas ( $\text{H}_2$ ).
  - react with carbonate slat (limestone) to produce carbon dioxide gas ( $\text{CO}_2$ ).



## Acid-Base Reactions: Introduction

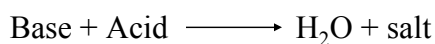
### Bases

- Have bitter taste.



- Many soaps, detergents, bleaches, and toothpaste contain NaOH (caustic soda). It can dissolve grease, oil, and fat. (feel slippery).
- Antacids have  $\text{Al}(\text{OH})_3$  or  $\text{Mg}(\text{OH})_2$ . They neutralize the gastric acid in the stomach.

- In general, strong bases react with strong acids to give water and salt.



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## Definitions of Acids and Bases

### Strong acids and strong bases:

- They ionize (dissociate) completely when dissolved in water. They are strong electrolytes.

### Arrhenius:

- **Acids** are substances that produce  $\text{H}^+$  when dissolved in water.
- **Bases** are substances that produce  $\text{OH}^-$  when dissolved in water.

### Brønsted:

- **Acids** are proton donors.
- **Bases** are proton acceptors.

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## Acid-Base Reactions

- **Arrhenius:**
  - **Acids** are substances that produce  $H^+$  when dissolved in water.
  - **Bases** are substances that produce  $OH^-$  when dissolved in water.
- **Brønsted:**
  - **Acids** are proton donors.
  - **Bases** are proton acceptors.

$NH_3$  is a base in the Arrhenius sense and in the Brønsted sense.

$H_2O$  is an acid in the Brønsted sense, but not in the Arrhenius sense

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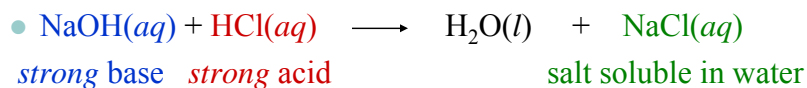
## Types of Acids

- **Monoprotic acid:**
  - The acid has one proton to donate.  $HCl$  and  $HNO_3$
  - Most of the strong acids are monoprotic acids.
- **Diprotic acid:**
  - The acid has two protons to donate.  $H_2SO_4$  and  $H_2CO_3$
  - Only  $H_2SO_4$  among the polyprotic acids is a strong acid.
- **Triprotic acid:**
  - The acid has three protons to donate.  $H_3PO_4$  and  $H_3C_6H_5O_7$
- **Bases** can also be monobasic, dibasic, tribasic.
  - $NaOH$   $Ba(OH)_2$   $Al(OH)_3$

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## Acid-Base Neutralization Reactions

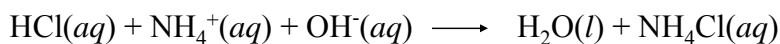
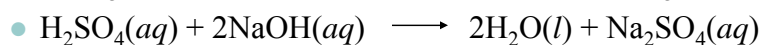
- In **neutralization reactions** an aqueous acid and base produce water and salt.



- Net ionic equation:



- Other examples of acid-base neutralization reactions:



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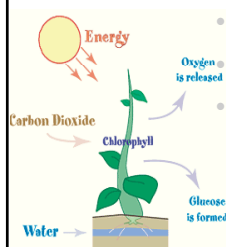
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## Oxidation-Reduction Reactions

- Oxidation-reduction reactions** (sometimes called **redox reactions**) are reactions involving the transfer of one electron or more from one reactant to another.

Redox reaction also involves the change in **oxidation states** for molecules.

- These reactions are very common in life:



- Photosynthesis. (conversion of  $\text{CO}_2$  and  $\text{H}_2\text{O}$  into sugar)
- Oxidation of sugar and fat in our bodies to produce energy.
- Combustion that provides humanity with power.



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## Oxidation-Reduction Reactions

**Oxidation of zinc in a solution of copper sulfate**

(a) Zinc atoms enter the solution as zinc ions ( $\text{Zn}^{2+}$ ). Copper ions are reduced to copper atoms on the surface of the metal.

(b) Cu atoms have replaced Zn atoms in the solid, and  $\text{Zn}^{2+}$  ions have replaced  $\text{Cu}^{2+}$  ions in solution.

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## Oxidation-Reduction Reactions

- Oxidation is losing electrons:  

$$\text{Zn(s)} \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-}$$
- Reduction is gaining electrons:  

$$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu(s)}$$

Half-reactions

- Redox (oxidation-reduction) reactions  
 If something in solution gets oxidized, then something else must be reduced (and vice versa).  

$$\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)} + 2\text{e}^{-}$$

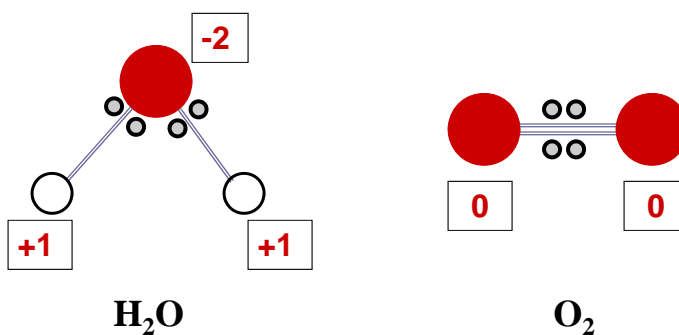

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$$\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$$

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## Oxidation States (Oxidation Numbers)

- Oxidation state** is an imaginary charge on an atom if the electrons were transferred completely to that atom. Normally, the shared electrons are completely assigned to the atoms the have stronger attraction for the electrons.



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## Oxidation State Rules

**TABLE 4.5** Elements with Reliable Oxidation Numbers in Compounds or Polyatomic Ions

Element	Oxidation Number	Exceptions
Fluorine	-1	
Group 1A or 2A metal	+1 or +2, respectively	
Hydrogen	+1	Any combination with a Group 1A or 2A metal to form a metal hydride. Examples: LiH and CaH <sub>2</sub> —the oxidation number of H is -1 in both examples.
Oxygen	-2	Any combination with something higher on the list that necessitates its having a different oxidation number (see rule 2 for assigning oxidation numbers). Examples: H <sub>2</sub> O <sub>2</sub> and KO <sub>2</sub> —the oxidation number of O for H <sub>2</sub> O <sub>2</sub> is 1 and for KO <sub>2</sub> is - $\frac{1}{2}$ .
Group 7A (other than fluorine)	-1	Any combination with something higher on the list that necessitates its having a different oxidation number (see rule 2 for assigning oxidation numbers). Examples: ClF, BrO <sub>3</sub> <sup>-</sup> , and IO <sub>3</sub> <sup>-</sup> —the oxidation numbers of Cl, Br, and I are +1, +7, and +5, respectively.

- The oxidation number for any element in its *elemental form* is zero (O<sub>2</sub>, F<sub>2</sub>).
- The oxidation number in any chemical species must *sum to the overall charge* on the species.
- The oxidation states in *ionic compounds* are the same as the charge each atom has by its own (PbS, NaCl)

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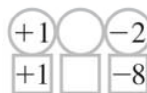
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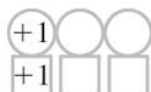
## Assigning Oxidation States



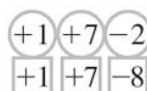
(a)



(c)



(b)



(d)

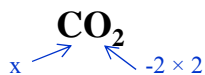
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## Assigning Oxidation States

### Exercise:

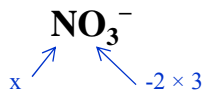
Assign oxidation states for all atoms in the following:

a)  $\text{CO}_2$       b)  $\text{SF}_6$       c)  $\text{NO}_3^-$ 

$$\begin{aligned} \text{Total charge} &= 0 = -2(2) + x \\ x &= +4 \end{aligned}$$



$$\begin{aligned} \text{Total charge} &= 0 = -1(6) + x \\ x &= +6 \end{aligned}$$



$$\begin{aligned} \text{Total charge} &= -1 = -2(3) + x \\ x &= +5 \end{aligned}$$

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## Oxidation-Reduction Process

$$\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$$

**Zn**

$-2e^-$  ↓

**Zn<sup>2+</sup>**

- Oxidized
- Losing electron(s)
- Oxidation state increases
- Reducing agent

**Cu<sup>2+</sup>**

↓  $+2e^-$

**Cu**

- Reduced
- Gaining electron(s)
- Oxidation state decreases
- Oxidizing agent

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## Oxidation States in Redox Reactions

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

$\begin{matrix} 0 & & 0 \\ \nearrow & & \searrow \\ 0 & & 0 \end{matrix}$ 
 $\begin{matrix} +1 & & -1 \\ \nearrow & & \searrow \\ +1 & & -1 \end{matrix}$

⇒  $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$

$\begin{matrix} -4 & & +1 \times 4 & & 0 \\ \nearrow & & \uparrow & & \searrow \\ -4 & & +1 & & 0 \end{matrix}$ 
 $\begin{matrix} +4 & & -2 \times 2 & & +1 \times 2 & & -2 \\ \nearrow & & \uparrow & & \uparrow & & \searrow \\ +4 & & -2 & & +1 & & -2 \end{matrix}$

$\begin{matrix} -4 & & & & +4 \\ \nearrow & & & & \uparrow \\ -4 & & & & +4 \end{matrix}$ 
 $\text{CH}_4 \longrightarrow \text{CO}_2 + 8e^-$

CH<sub>4</sub> is a reducing agent

$\begin{matrix} 0 & & & & -2 \times 2 & & -2 \\ \nearrow & & & & \uparrow & & \searrow \\ 0 & & & & -2 & & -2 \end{matrix}$ 
 $2\text{O}_2 + 8e^- \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

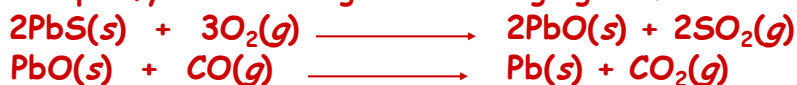
O<sub>2</sub> is an oxidizing agent

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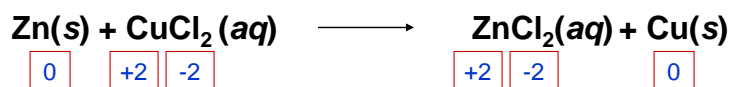
## Exercise

For the following two reactions:

- \* determine the oxidation states,
- \* identify the atoms that are oxidized and reduced, and
- \* specify the oxidizing and reducing agents.

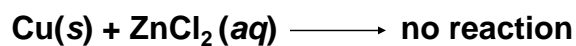
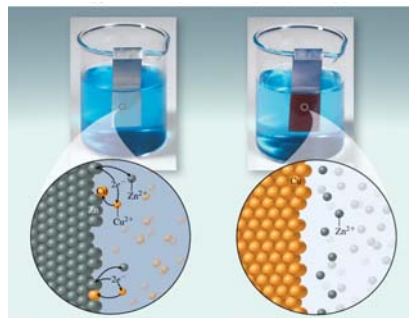


## Redox Reactions in Aqueous Solutions



“Displacement reaction”

- o What would happen if you place copper metal into a solution of  $\text{ZnCl}_2$ ?
- Would  $\text{Cu}(s)$  be oxidized by  $\text{Zn}^{2+}(aq)$  ions the way  $\text{Zn}(s)$  is oxidized by  $\text{Cu}^{2+}(aq)$  ions?



## The Activity Series

- The **activity series** shows the order of ease the metal is to be oxidized.
- Metals at the top of the list are called the **active metals**.  
Metals at the bottom of the list are known as **noble metals**.

TABLE 4.6 Activity Series

Element	Oxidation Half-Reaction
Lithium	$\text{Li} \longrightarrow \text{Li}^+ + e^-$
Potassium	$\text{K} \longrightarrow \text{K}^+ + e^-$
Barium	$\text{Ba} \longrightarrow \text{Ba}^{2+} + 2e^-$
Calcium	$\text{Ca} \longrightarrow \text{Ca}^{2+} + 2e^-$
Sodium	$\text{Na} \longrightarrow \text{Na}^+ + e^-$
Magnesium	$\text{Mg} \longrightarrow \text{Mg}^{2+} + 2e^-$
Aluminum	$\text{Al} \longrightarrow \text{Al}^{3+} + 3e^-$
Manganese	$\text{Mn} \longrightarrow \text{Mn}^{2+} + 2e^-$
Zinc	$\text{Zn} \longrightarrow \text{Zn}^{2+} + 2e^-$
Chromium	$\text{Cr} \longrightarrow \text{Cr}^{3+} + 3e^-$
Iron	$\text{Fe} \longrightarrow \text{Fe}^{2+} + 2e^-$
Cadmium	$\text{Cd} \longrightarrow \text{Cd}^{2+} + 2e^-$
Cobalt	$\text{Co} \longrightarrow \text{Co}^{2+} + 2e^-$
Nickel	$\text{Ni} \longrightarrow \text{Ni}^{2+} + 2e^-$
Tin	$\text{Sn} \longrightarrow \text{Sn}^{2+} + 2e^-$
Lead	$\text{Pb} \longrightarrow \text{Pb}^{2+} + 2e^-$
Hydrogen	$\text{H}_2 \longrightarrow 2\text{H}^+ + 2e^-$
Copper	$\text{Cu} \longrightarrow \text{Cu}^{2+} + 2e^-$
Silver	$\text{Ag} \longrightarrow \text{Ag}^+ + e^-$
Mercury	$\text{Hg} \longrightarrow \text{Hg}^{2+} + 2e^-$
Platinum	$\text{Pt} \longrightarrow \text{Pt}^{2+} + 2e^-$
Gold	$\text{Au} \longrightarrow \text{Au}^{3+} + 3e^-$

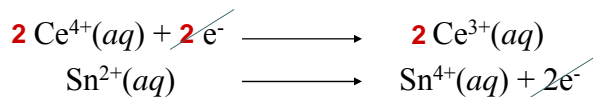
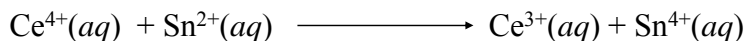
Increasing ease of oxidation

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## Balancing Oxidation-Reduction Equations

- The Half-Reaction Method:**

A half reaction is that reaction that involves either oxidation or reduction.



Atoms and charges (electrons) must be all balanced.

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## Concentration of Solutions

- **Concentration** is the amount of chemicals (solutes) present (dissolved) in the solution.

$$\text{Molarity} = M = \frac{\text{Moles of solute}}{\text{Liters of solution}} = \frac{n}{V}$$

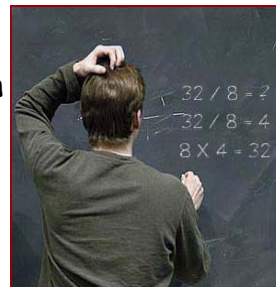
*"Molar Concentration"*

Molarity has the unit of mol/L , mol·L<sup>-1</sup> , *M*.

- If you have 0.1 moles of NaOH present in 1L aqueous solution, the solution has 0.1*M* concentration.
- Useful web links:
  - <http://dbhs.wvusd.k12.ca.us/webdocs/Solutions/Molarity.html>.
  - <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson64.htm>.
  - <http://www.iun.edu/~cpanhd/C101webnotes/aqueoussolns/molarity.html>.

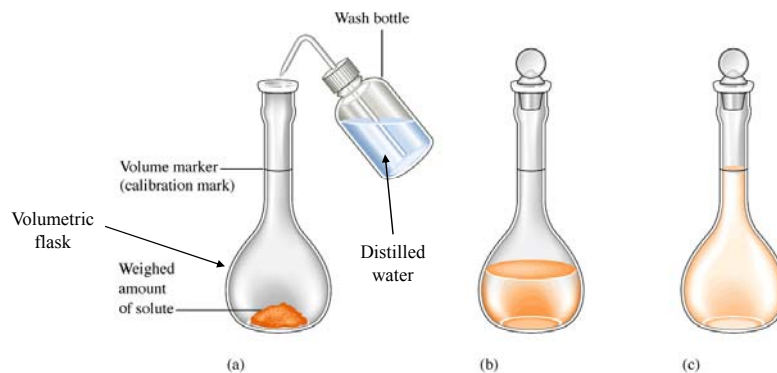
## Exercise on Molarity Calculations

Calculate the molarity of solution prepared by dissolving 1.56g of HCl in water to make 26.8 mL solution.



## Preparing a Solution from a Solid

- A **standard solution** is a solution whose concentration is accurately known.
- Steps of preparing a standard solution:

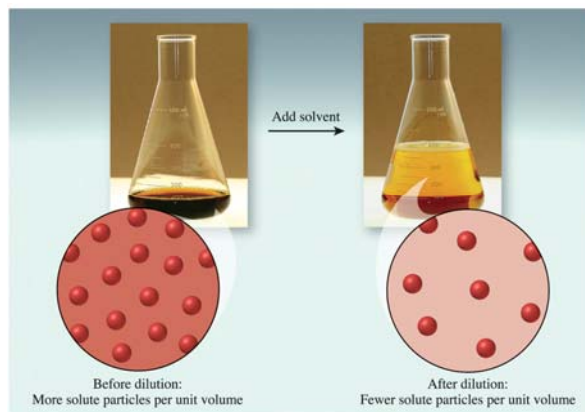


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## Dilution

- Dilution is the procedure of adding water to *stock solutions* “often are concentrated solutions and kept in the laboratory” to achieve the desired concentration.



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## Dilution

- Dilution is the procedure of adding water to *stock solutions* “often are concentrated solutions and kept in the laboratory” to achieve the desired concentration.
- It is always true that:

Moles of solute **before** dilution = Moles of solutes **after** dilution

$$\text{Since no. of moles} = M \times V = \frac{\text{moles}}{\text{liters}} \times \text{liters}$$

Then,  $M \times V$  (before dilution) =  $M \times V$  (after dilution)

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

## Dilution

- What volume of 1.00M  $\text{KMnO}_4$  is needed to prepare 1.00 L of a 0.400M  $\text{KMnO}_4$  solution?

Moles of solute **before** dilution = Moles of solutes **after** dilution

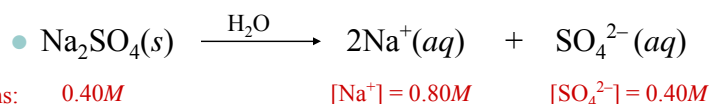
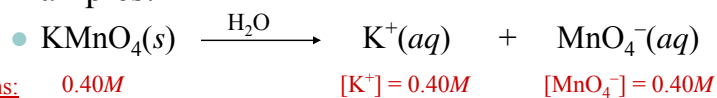
$$M_d \times V_d = M_c \times V_c$$

$$\begin{aligned} V_c &= M_d / M_c \times V_d \\ &= 0.400M / 1.00M \times 1.00L \\ &= 0.400 L \end{aligned}$$

Answer is 400 mL of the 1.00M  $\text{KMnO}_4$  stock solution.

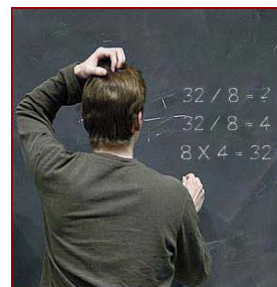
## Solution Stoichiometry

- Soluble ionic compounds are strong electrolytes, i.e. they dissociate completely and exist as ions in aqueous solutions.
- Examples:



## Exercises on Solution Stoichiometry Calculations

Give the concentration of  $\text{ClO}_4^-$  ions in  $1M \text{Fe}(\text{ClO}_4)_3$  solution.

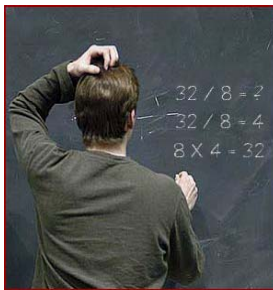




Chapter 4 Section 5

Exercises on Solution Stoichiometry  
Calculations

Calculate the number of moles of  $\text{Cl}^-$  ions in 1.75L of  $1.0 \times 10^{-3} \text{ M}$   $\text{ZnCl}_2$  solution.



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Exercises on Solution Stoichiometry  
Calculations

28.0 mL of 0.250M  $\text{HNO}_3$  and 53.0 mL of 0.320M  $\text{KOH}$  are mixed.

- Calculate the amount of water formed in the resulting reaction.
- What are the concentrations of  $\text{H}^+$  and  $\text{OH}^-$  ions in excess after the reaction goes to completion?

⇒ Net ionic equation:  

$$\text{H}^+ (aq) + \text{OH}^- (aq) \longrightarrow \text{H}_2\text{O} (l)$$

⇒ From volume and conc. find the moles for  $\text{H}^+$  and  $\text{OH}^-$ .  
 moles of  $\text{H}^+$  =  $7.00 \times 10^{-3}$  mol  
 moles of  $\text{OH}^-$  =  $1.70 \times 10^{-2}$  mol

⇒ Determine which reactant is the limiting one. Then find the amount of  $\text{H}_2\text{O}$  formed.

⇒ Conc. of excess  $\text{OH}^-$  =  $\frac{\text{Moles of unreacted OH}^-}{\text{Total volume of solution}}$

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## Exercises on Solution Stoichiometry Calculations

When aqueous solutions of  $\text{Na}_2\text{SO}_4$  and  $\text{Pb}(\text{NO}_3)_2$  are mixed,  $\text{PbSO}_4$  precipitates. Calculate the mass of  $\text{PbSO}_4$  formed when 1.25L of 0.0500M  $\text{Pb}(\text{NO}_3)_2$  and 2.00L of 0.0250M  $\text{Na}_2\text{SO}_4$  are mixed. How many ions of  $\text{Pb}^{2+}$  will remain unreacted in the solution?

1. Identify the ions and possible solid product.
2. Give net ionic equation.
3. Find numbers of moles for  $\text{Pb}^{2+}$  and  $\text{SO}_4^{2-}$ .
4. Which one is limiting?
5. Calculate moles (then grams) of  $\text{PbSO}_4$  based on limiting reactant.

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Chapter 4 Section 5

## Exercises on Solution Stoichiometry Calculations

What mass of  $\text{Na}_2\text{CrO}_4$  is required to precipitate all of the silver ions from 75.0 mL of a 0.100 M solution of  $\text{AgNO}_3$ ?



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## Aqueous Reactions and Chemical Analysis

Many aqueous reactions are very useful for determining how much of a particular substance is present in a sample.

- Gravimetric analysis.
- Acid-base titration.



## Gravimetric Analysis

- It is an analytical technique that is based on the measurement of mass. The precipitate formed out of a precipitation reaction is isolated and measured.
  - The reaction must have 100% yield.
  - The precipitate must be completely insoluble.

## Gravimetric Analysis

A 0.8633-g sample of an ionic compound  $MCl_x$  is dissolved in water and treated with an excess  $AgNO_3$ . If 1.5615 g of  $AgCl$  precipitate forms, what is the % by mass of Cl in  $MCl_x$ ?

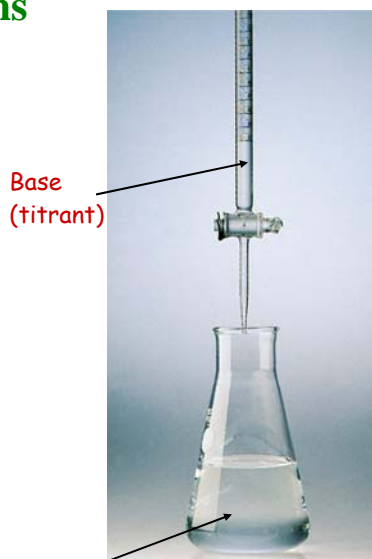
% mass of Cl in  $AgCl$  =

mass of Cl in  $AgCl$  ppt =

% mass of Cl in  $MCl_x$  =

## Acid-Base Titrations

- *Titration* (or standardization) is used to characterize aqueous solutions (*acidic or basic*) of an unknown concentration.
- It is done by gradually adding a strong acid (or strong base) solution (*titrant*) of known concentration to a base (or acid) solution (*analyte*) for which the concentration is needed to be determined, with the presence of an *indicator*.



Base  
(titrant)

Acid (analyte) + indicator

Chapter 4 Section 6

## Acid-Base Titrations

Titrant  
(Known concentration)

Analyte  
(known volume,  
Unknown concentration)

# mol OH<sup>-</sup> = VxM  
(dispensed)

Equivalence  
point: Change in  
color (indicator)  
#mol OH<sup>-</sup> = # mol H<sup>+</sup>

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

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## Acid-Base Titrations

- In this example, when the reaction is completed, the base titrant neutralizes the acid analyte. The point of neutralization (*end point* or *equivalence point*)
- The end point can be visually located by using *indicators*, that change their colors when an excess of the titrant is present in the solution.
- This experiment has to be done very carefully.

Known volume and concentration

Known volume

<http://real.video.ufl.edu:8080/ramgen/chm2040/demos/A15-2-20.rm>

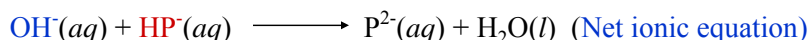
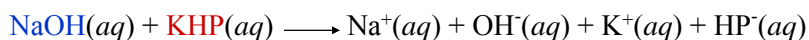
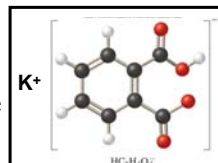
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## Standardization of NaOH Solution Using KHP

- **Standardization** is the process to accurately determine the concentration of a solution before using it as a titrant in a titration experiment.

- **KHP** is potassium hydrogen phthalate.

It is a monoprotic acid that is used to standardize NaOH solutions of unknown concentrations.



- In this case, solution of **KHP** which has a **known mass** is titrated with NaOH of **unknown concentration**.

## Standardization of NaOH Solution Using KHP

**In a titration experiment, it was found that 25.49 ml of NaOH solution was needed to neutralize 0.7137 g of KHP. What is the concentration of the NaOH solution?**

