## I. General Properties of Aqueous Solutions

• A solution is a homogeneous mixture of two or more pure substances.

• A solution is made when one substance (the **solute**) is dissolved in another (the s**olvent**).

- **Solute:** the component that is dissolved that is present in smallest amount.
- **Solvent:** the liquid when a solute is Dissolved, is usually present in greatest abundance.

Solutions in which water is the solvent are called **aqueous solutions**.

- Some solutes can *dissociate* into ions. Electric charge can be carried.

$$NaCl_{(s)} \xrightarrow{H_2O} Na_{(aq)} + Cl_{(aq)}$$

- Conduct electricity in solution are : the Cations**(+)**and the Anions (−)

### I.1. Electrolytes :

An **electrolyte**is a substance that, when dissolved in water, results in a solution that can conduct electricity.

Contains ions :

A **nonelectrolyte :** is a substance that, when dissolved, results in a solution that does not conduct electricity.

Does not contain ions :

$$C_{12}H_{22}O_{11(s)} \xrightarrow{H_2O} C_{12}H_{22}O_{11(aq)}$$

#### I.1.1. Types of electrolytes :

*Strong electrolyte* dissociates completely ⟶ Good electrical conduction.

$MgCl_2$ $\quad Mg^{2+} + 2\,Cl^-$ $\longrightarrow$

*Weak electrolyte* partially dissociates $\longrightarrow$ Fair conductor of electricity.

Examples of weak electrolytes

Exampel :

– Weak acids

$$CH_3CO_2H(aq) \rightleftharpoons CH_3CO_2^-(aq) + H^+(aq)$$

– Weak bases

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH4^+_{(aq)} + OH^-_{(aq)}$$

(Note: double arrows indicate a reaction that occurs in both directions - a state of dynamic equilibrium exists)

### II.2. Calculating Ion concentrations in a Solution of a Strong Electolyte :

What are the aluminum and sulfate ion concentrations in 0.0165 M $Al_2(SO_4)_3$.

1. Write a Balanced Chemical Equation:

   $Al_2(SO_4)_3\,(s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

2. Identify the Stoichiometric Factors :

   1mol $(Al(SO_4)_3)$ : 2 mol $(Al^{3+})$ ; 3mol $(SO_4^{2-})$

There are three main types of aqueous reactions: precipitation reactions, acid-base reactions, and oxidation-reduction (or redox) reactions.

## II. Precipitation Reactions

Precipitation reactions produce an insoluble product- the precipitate. They contain two aqueous reactants, one aqueous product, and one solid product.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

In this reaction, two soluble products, Pb(NO3)2 and KI, combine to form one soluble product, KNO3, and one insoluble product, PbI2. This is a precipitation reaction, and PbI2 is the precipitate.

*Not : In chemical equations, certain abbreviations are used to indicate the state of the substances involved. The abbreviations are as follows: s = solid; l = liquid; g = gaseous; aq = aqueous, or soluble in water.*

### II.2. Determining the Products of a Precipitation Reaction

To determine the products of a precipitation reaction, reverse the cation-anion pairs.** For example precipitation reactions follow this formula:

$$AX + BY \longrightarrow AY + BX$$

The products are just the cation-anion pairs reversed, or the “outies” (A and Y joined) and the “innies” (B and X joined).

**Determining Whether or Not a Reaction is a Precipitation Reaction**

Once you know the products of a reaction, you can use the solubility rules to see if you have an insoluble product, and thus, a precipitation reaction.

## II.3. Solubility Rules:

| **Soluble:** All ionic compounds containing: | **Exceptions:** |
|---|---|
| 1. Alkali metals (group 1A) | 1. None |
| 2. Ammonium ($NH_4^+$) | 2. None |
| 3. Nitrate ($NO_3^-$) | 3. None |
| 4. Acetate ($C_2H_3O_2^-$) | 4. None |
| 5. Chloride ($Cl^-$) | 5. AgCl, $PbCl_2$, $Hg_2Cl_2$, $Cul_2$ |
| 6. Bromide ($Br^-$) | 6. AgBr, $PbBr_2$, $Hg_2Br_2$, $CuBr_2$ |
| 7. Iodide ($I^-$) | 7. AgI, $PbI_2$, $Hg_2I_2$, $CuI_2$ |
| 8. Sulfate ($SO_4^{2-}$) | 8. $SrSO_4$, BaSO4, $Hg_2SO_4$, $PbSO_4$, $CaSO_4$ |
| **Insoluble:** Compounds containing: | **Exceptions:** |
| 1. $S^{2-}$ | 1. When bonded to ammonium, alkali metals, $Ca^{2+}$, $Sr^{2+}$, or $Ba^{2+}$ |
| 2. $CO_3^{2-}$ | 2. When bonded to ammonium or alkalis |
| 3. $PO_4^{3-}$ | 3. Same as above |
| 4. $OH^-$ | 4. When bonded to alkali metals, $Ca^{2+}$, $Sr^{2+}$, or $Ba^{2+}$ |

Example: Predict the products formed by the aqueous reaction below and determine whether or not the reaction is a precipitation reaction.

$$BaCl_{2(aq)} + K_2Cl_{4(aq)} \longrightarrow$$

The first step is to predict the products, which we do by reversing the pairs, giving us $BaSO_4$, and KCl. *Remember to balance the equation*.

$$BaCl_{2(aq)} + K_2SO_{4(aq)} \longrightarrow BaSO_4 + 2KCl$$

Next, we use the solubility rules to determine if this is a precipitation reaction. The table tells us that compounds containing alkali metals, such as potassium, are soluble–thus KCl is soluble. We also see that sulfate is soluble *except when bonded to barium*! Thus, BaSO4 is insoluble, and this is a precipitation reaction. The whole balanced equation is:

$$BaCl_{2(aq)} + K_2SO_{4(aq)} \longrightarrow BaSO_{4(s)} + 2KCl_{(aq)}$$

**Ionic Equations**

Something that is useful when dealing with precipitation reactions is the ability to write ionic equations, which show the compounds as individual ions. Until now, you have been writing chemical equations in this form:

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

Equations written this way are known as **molecular equations**. They have a variation known as a **complete ionic equation**, in which all soluble strong electrolytes are written as individual ions. Thus, the above reaction becomes:

$$Pb^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)} + 2K^+{}_{(aq)} + I^-{}_{(aq)} \longrightarrow PbI_{2(s)} + 2K^+{}_{(aq)} + 2NO_3^-{}_{(aq)}$$

When writing a complete ionic equation, remember that only soluble strong electrolytes are written as individual ions. You already have the guidelines for determining if something is soluble; below is a table which can be used to determine if a substance is a strong, weak, or nonelectrolyte.

$$Pb^{2+}_{(aq)} + I^{-}_{(aq)} \longrightarrow PbI_{2(s)}$$

We simply took out the spectator ions- the potassium and nitrate- and ended with the net ionic equation.

To summarize the series of steps to get from one form of an equation to another:

1. Write a balanced molecular equation, just like you've been doing.
2. Rewrite the equation, showing all soluble, strong electrolytes as individual ions, to get the complete ionic equation. Keep it balanced.
3. Eliminate all spectator ions to get the net ionic equation.

**Acid-Base Reactions**

Acids are substances that release H+ ions in water. Bases accept these H+ ions, and produce OH- in water (occasionally a base such as ammonia, NH3, won't contain OH-. Most bases, though, contain hydroxide). Like other electrolytes, there are both strong and weak acids and bases. It is important to know the strong acids and bases from the weak:

| **Strong Acids** | **Common Strong Bases** |
|---|---|
| 1. Hydrochloric acid (HCl) | 1. LiOH |
| 2. Hydrobromic acid (HBr) | 2. NaOH |
| 3. Hydroiodic acid (HI) | 3. KOH |
| 4. Chloric acid ($HClO_3$) | 4. RbOH |
| 5. Perchloric acid ($HClO_4$) | 5. CsOH |
| 6. Nitric acid ($HNO_3$) | 6. $Ca(OH)_2$ |
| 7. Sulfuric acid ($H_2SO_4$) | 7. $Sr(OH)_2$ |
| | 8. $Ba(OH)_2$ |

All other acids and almost all other bases you will encounter are weak.

## III. Acid-Base Reactions: Neutralization Reactions

When acids and bases react, a **neutralization reaction** occurs. In this reaction, the acid donates an $H^+$ ion. This joins with the hydroxide ion from the base to form water, while the anion from the acid and the cation from the base join to form an ionic compound. Here is a typical acid base reaction:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

In fact, these neutralization reactions have the same form as the precipitation reactions we looked at earlier:

$$AX + BY \longrightarrow AY + BX$$

In general, any neutralization reaction between an acid and base produces water and a **salt**

Now, write the complete ionic form of the above reaction. You should get:

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

Next, write the net ionic equation. It should look like this:

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

This equation summarizes the essential feature of the reaction between any strong acid and strong base- a hydrogen ion and a hydroxide ion forming water. Any reaction between a strong acid and a strong base can be brought back to this. Only strong acid-strong base reactions can be reduced this way, however. With weak acids or bases, it's a bit different, as we see in this reaction involving acetic acid:

$$2CH_3COOH_{(aq)} + Ba(OH)_{2(aq)} \longrightarrow Ba(CH_3COO)_{2\,(aq)} + H_2O_{(l)}$$

When writing the complete ionic equation for this reaction, remember that acetic acid is a weak electrolyte. Thus, the complete ionic equation is:

$$2CH_3COOH_{(aq)} + Ba^{2+}_{(aq)} + 2OH^-_{(aq)} \longrightarrow Ba^{2+}_{(aq)} + 2CH_3COO^-_{(aq)} + 2H_2O_{(l)}$$

If we write the net ionic equation, it comes to:

$$2CH_3COOH_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2CH_3COO^-_{(aq)} + 2H_2O_{(l)}$$

Because this reaction involves a weak acid, we cannot reduce it to the same short equation that we could in our strong acid-strong base reaction.

## IV. Oxidation-Reduction Reactions

Oxidation-reduction, or **redox reactions**, take place between metals and non-metals. Below is a good example of a redox reaction:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2\,Fe(l) + 3\,CO_2(g)$$

- **Oxidation:** Loss of electrons by a substance.An oxidation occurs when an atom or ion loses electrons.

- **Reduction:** Gain of electronsby a substance.A reduction occurs when an atom or ion gains electrons.

•*One cannot occur without the other*: **Redox-**reaction Reaction in which electrons are transferred from one reactant to another.

Oxidation and reduction always occur together :

$Fe^{3+}$ is reduced to metallic iron.

CO(g) is oxidized to carbon dioxide.

Assign oxidation states:

$$\overset{3+\ \ 2-}{Fe_2O_3(s)} + 3\overset{2+\ 2-}{CO(g)} \rightarrow 2\ \overset{0}{Fe(l)} + 3\ \overset{4+\ 2-}{CO_2(g)}$$

Oxidation number (O.N.)

- Oxidation :

O.N. of some element *increases* in the reaction. Electrons are on the right of the equation

- Reduction

O.N. of some element *decreases* in the reaction. Electrons are on the left of the equation.

## IV.1. Oxidation numbers :

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

For neutral molecules and polyatomic ions, the oxidation numberof a given atom is a hypothetical charge.This charge is assigned by artificially dividing up the electrons among the atoms in the molecule or ion.

Rulesfor assigning oxidation numbers:

- Elements in their **elemental form**have an oxidation number of 0.

  $O_2$, Fe, Mg, C

- The oxidation number of a **monatomic ion**is the same as its charge.

$O^{2-}$ (ON = -II) ; $Fe^{2+}$ (ON = +II)

- **Nonmetals** tend to have negative oxidation numbers, although some are positivein certain compounds or ions.
- **Oxygen** has an oxidation number of −II, except in the peroxide ion, in which it has an oxidation number of –I, and in $F_2O$ (ON = +I).
- **Hydrogen**is −I when bonded to a metal, +I when bonded to a nonmetal.
- **Fluorine**always has an oxidation number of −1.
- The **other halogens** have an oxidation number of –I when they are negative; they can have positive oxidation numbers, however, most notably in oxyanions.

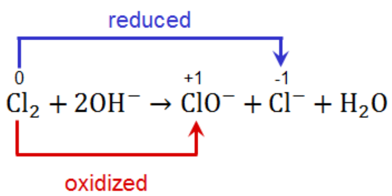
- The **sum of the oxidation numbers** in a **neutral compound**is 0.
- The **sum of the oxidation numbers** in a **polyatomic ion**is the charge

$CCl_4$ (Cl O.N.= -1) ; N.O.(C) + 4(-I) = 0 C O.N.= +4

**Disproportionation Reaction :**

The same element is simultaneously oxidized and reduced.

Example:




$$\overset{0}{Cl_2} + 2OH^- \rightarrow \overset{+1}{Cl}O^- + \overset{-1}{Cl^-} + H_2O$$

## IV.2. Oxidation and reduction half reaction

- A reaction represented by two half-reactions.

*Oxidation:* $Zn(s) \rightarrow Zn^{2+}_{(aq)} + 2\,e$

*Reduction:* $Cu^{2+}_{(aq)} + 2\,e \rightarrow Cu_{(s)}$

*Overall:* $Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$

Balancing redox reaction :

Divide reaction into half-reaction

$MnO_4^-{}_{(aq)} + Br^-_{(aq)} \rightarrow Mn^{2+}_{(aq)} + Br_{2(g)}$

$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)}$

$Br^-_{(aq)} \rightarrow Br_{2(g)}$

2- assign an ON to each atom in the eaquation and determine whether it is a redox reaction :

+VII +II

$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)}$

-I 0

$Br^-{}_{(aq)} \rightarrow Br_{2(g)}$

3- balance the atoms of element and the ON by adding electrons

+VII +II

$MnO_4^-{}_{(aq)} + 5e \rightarrow Mn^{2+}{}_{(aq)}$

-I 0

$2\ Br^-{}_{(aq)} \rightarrow Br_{2(g)} + 2e$

3- balance charge by adding:

$H^+$ ions in acidic solution :

$MnO_4^-{}_{(aq)} + 5e + 8\ H^+ \rightarrow Mn^{2+}{}_{(aq)}$

$2\ Br^-{}_{(aq)} \rightarrow Br_{2(g)} + 2e$

$OH^-$ ions in basic solution

$MnO_4^-{}_{(aq)} + 5e \rightarrow Mn^{2+}{}_{(aq)} + 8\ OH^-$

$2\ Br^-{}_{(aq)} \rightarrow Br_{2(g)} + 2e$

5- balance H and O atoms by adding $H_2O$

In acidic medium :

$MnO_4^-{}_{(aq)} + 5e + 8\ H^+ \rightarrow Mn^{2+}{}_{(aq)} + 4\ H_2O$

$2\ Br^-{}_{(aq)} \rightarrow Br_{2(g)} + 2e$

In basic medium :

$MnO_4^-{}_{(aq)} + 5e + 4\ H_2O \rightarrow Mn^{2+}{}_{(aq)} + 8\ OH^-$

$2\ Br^-{}_{(aq)} \rightarrow Br_{2(g)} + 2e$

6- combine the two half-equation in such a way as eleminte electrons

In acidic medium

$2MnO_4^-{}_{(aq)} + \cancel{10e} + 16\ H^+ \rightarrow 2Mn^{2+}{}_{(aq)} + 8\ H_2O$

$10Br^-{}_{(aq)} \rightarrow 5Br_{2(g)} + \cancel{10e}$

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$2MnO_4^-{}_{(aq)} + 10Br^-{}_{(aq)} + 16\ H^+ \rightarrow 2Mn^{2+}{}_{(aq)} + 5Br_{2(g)} + 8\ H_2O$

In basic medium

$2MnO_4^-{}_{(aq)} + \cancel{10e} + 8\ H_2O \rightarrow 2Mn^{2+}{}_{(aq)} + 16\ OH^-$

$10Br^-{}_{(aq)} \rightarrow 5Br_{2(g)} + \cancel{10e}$

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$2MnO_4^-{}_{(aq)} + 10Br^-{}_{(aq)} + 8\ H_2O \rightarrow 2Mn^{2+}{}_{(aq)} + 5Br_{2(g)} + 16\ OH^-$