

## Chemical Reactions

Chemical reactions are the heart of chemistry. Some reactions, such as those accompanying a forest fire or the explosion of dynamite, are quite dramatic. Others are much less obvious, although all chemical reactions must involve detectable change. A chemical reaction involves a change from reactant substances to product substances, and the product substances will have physical and chemical properties different from those of the reactants.

### Types of Chemical Reactions:

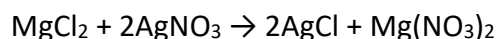
Among the several million known substances, many millions of chemical reactions are possible. How can we know whether two substances will react whether they are mixed? How can we predict the products? Although it is not possible to give completely general answers to these questions, it is possible to make sense of chemical reactions. Most of the reactions we will study belong to one of three types:

1. **Precipitation reactions.** In these reactions, mixing solutions of two ionic substances, and a solid ionic substance (a precipitate) forms.
2. **Acid–base reactions.** An acid substance reacts with a base substance. Such reactions involve the transfer of a proton between reactants.
3. **Oxidation–reduction reactions.** These involve the transfer of electrons between reactants.

### Precipitation Reactions

A precipitation reaction occurs in aqueous solution because one product is insoluble. A **precipitate** is *an insoluble solid compound formed during a chemical reaction in solution*.

When we write a precipitation reaction as a molecular equation, the reaction has the form of an exchange reaction. In a precipitation reaction, the anions exchange between the two cations (or vice versa). Let us discuss a precipitation reaction between magnesium chloride,  $\text{MgCl}_2$ , and silver nitrate,  $\text{AgNO}_3$ . The balanced reaction is:



Let us verify that  $\text{MgCl}_2$  and  $\text{AgNO}_3$  are soluble and then check the solubilities of the products. Rule 3 in Table 1 says that chlorides are soluble, with certain exceptions, which do not include magnesium chloride. Thus, we predict that magnesium chloride is soluble. Rule 2 indicates that nitrates are soluble, so  $\text{AgNO}_3$  is soluble as well. The potential products are silver chloride and magnesium nitrate. According to Rule 3, silver chloride is one of the exceptions to the general solubility of chlorides. Therefore, we predict that the silver chloride is insoluble. Magnesium nitrate is soluble according to Rule 2.

**Table 1. Solubility Rules for Ionic Compounds**

<b>Rule Applies to</b>	<b>Statement</b>	<b>Exceptions</b>
1. $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$	Group IA and ammonium compounds are soluble.	—
2. $\text{C}_2\text{H}_3\text{O}_2^-$ , $\text{NO}_3^-$	Acetates and nitrates are soluble.	—
3. $\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$	Most chlorides, bromides, and iodides are soluble.	$\text{AgCl}$ , $\text{Hg}_2\text{Cl}_2$ , $\text{PbCl}_2$ , $\text{AgBr}$ , $\text{HgBr}_2$ , $\text{PbBr}_2$ , $\text{AgI}$ , $\text{HgI}_2$ , $\text{Hg}_2\text{I}_2$ , $\text{PbI}_2$
4. $\text{SO}_4^{2-}$	Most sulfates are soluble.	$\text{CaSO}_4$ , $\text{SrSO}_4$ , $\text{BaSO}_4$ , $\text{Ag}_2\text{SO}_4$ , $\text{Hg}_2\text{SO}_4$ , $\text{PbSO}_4$
5. $\text{CO}_3^{2-}$	Most carbonates are insoluble.	Group IA carbonates, $(\text{NH}_4)_2\text{CO}_3$
6. $\text{PO}_4^{3-}$	Most phosphates are insoluble.	Group IA phosphates, $(\text{NH}_4)_3\text{PO}_4$
7. $\text{S}^{2-}$	Most sulfides are insoluble.	Group IA sulfides, $(\text{NH}_4)_2\text{S}$
8. $\text{OH}^-$	Most hydroxides are insoluble.	Group IA hydroxides, $\text{Ca}(\text{OH})_2$ , $\text{Sr}(\text{OH})_2$ , $\text{Ba}(\text{OH})_2$

### Acid-Base Reactions

Acids and bases are some of the most important electrolytes\*. **Acids** have a sour taste, change litmus from blue to red, pH value 0 to below 7 and evolve  $\text{H}_2$  gas upon reaction with active metals (such as alkali metals, alkaline earth metals, Zn, Al). Solutions of **bases**, on the other hand, have a bitter taste, change red litmus back to blue, pH value above 7 to 14, and a soapy feel. Acid react with bases to form salt and water. Their aqueous (water) solutions conduct electric current. Some examples of acids are acetic acid, present in vinegar; citric acid, a constituent of lemon juice; and hydrochloric acid, found in the digestive fluid of the stomach. An example of a base is aqueous ammonia, often used as a household cleaner (Table 2).

[\*An **electrolyte** is a substance that dissolves in water to give an electrically conducting solution. Sodium chloride, table salt, is an example of an electrolyte. A **non-electrolyte** is a substance that dissolves in water to give a non-conducting or very poorly conducting solution. A common example is sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ , which is ordinary table sugar. Another example is methanol,  $\text{CH}_3\text{OH}$ , a compound used in car window washer solution. A **strong electrolyte** is an electrolyte that exists in solution almost entirely as ions. A **weak electrolyte** is an electrolyte that dissolves in water to give a relatively small percentage of ions.]

**TABLE 2. Common Acids and Bases**

<b>Name</b>	<b>Formula</b>	<b>Remarks</b>
<b>Acids</b>		
Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$	Found in vinegar
Acetylsalicylic acid	$\text{HC}_9\text{H}_7\text{O}_4$	Aspirin
Ascorbic acid	$\text{H}_2\text{C}_6\text{H}_6\text{O}_6$	Vitamin C
Citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	Found in lemon juice
Hydrochloric acid	$\text{HCl}$	Found in gastric juice (digestive fluid in stomach)
Sulfuric acid	$\text{H}_2\text{SO}_4$	Battery acid
<b>Bases</b>		
Ammonia	$\text{NH}_3$	Aqueous solution used as a household cleaner
Calcium hydroxide	$\text{Ca}(\text{OH})_2$	Slaked lime (used in mortar for building construction)
Magnesium hydroxide	$\text{Mg}(\text{OH})_2$	Milk of magnesia (antacid and laxative)
Sodium hydroxide	$\text{NaOH}$	Drain cleaners, oven cleaners

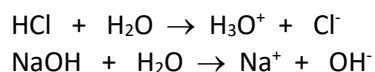
Another simple property of acids and bases is their ability to cause color changes in certain dyes. An acid–base indicator is a dye used to distinguish between acidic and basic solutions by means of the color changes it undergoes in these solutions. Such dyes are common in natural materials. The amber color of tea, for example, is lightened by the addition of lemon juice (citric acid). Red cabbage juice changes to green and then yellow when a base is added. The green and yellow colors change back to red when an acid is added. Litmus is a common laboratory acid–base indicator. This dye, produced from certain species of lichens, turns red in acidic solution and blue in basic solution. Phenolphthalein, another laboratory acid–base indicator, is colorless in acidic solution and pink in basic solution.

An important chemical characteristic of acids and bases is the way they react with one another. To understand these acid–base reactions, we need to have precise definitions of the terms acid and base.

**Definition of Acids and Bases:** Three concepts acid and base are available.

- (1) Arrhenius concept
- (2) Bronsted-Lowry concept
- (3) Lewis concept

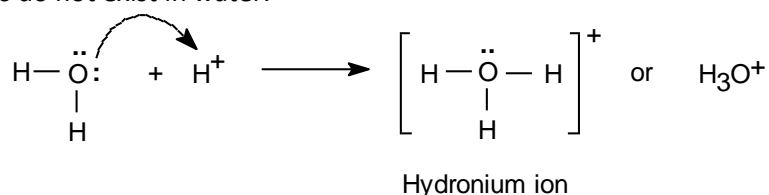
**Arrhenius concept** (1884): It is based on ionic dissociation of compound in water. An **acid** is a compound that releases  $H^+$  ions in  $H_2O$ . For example,  $HCl$  is an acid since it releases  $H^+$  ion when dissolved in  $H_2O$ . A **base** is a compound that releases  $OH^-$  ions in  $H_2O$ . For example,  $NaOH$  is a base since it releases  $OH^-$  ion when dissolved in  $H_2O$ .



*Usefulness:* This concept is useful in the study of chemical reactions.

*Limitations:*

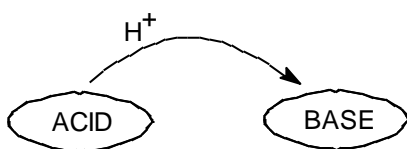
- (a) Free  $H^+$  ions do not exist in water.

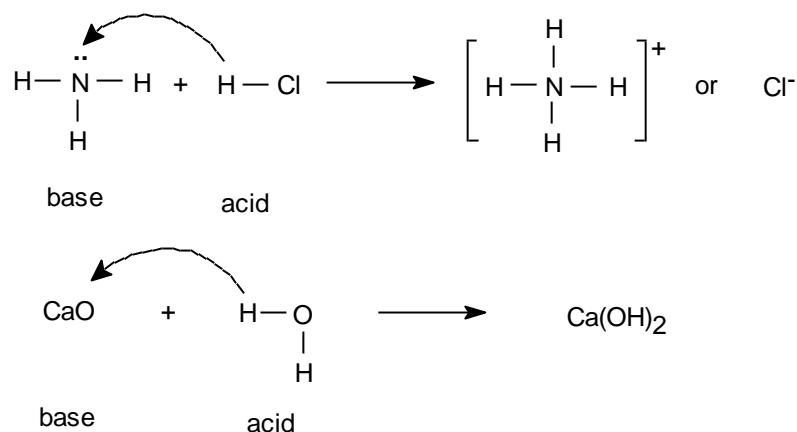


- (b) Limited to water only: These definitions are applicable to water only.  
(c) Some bases do not contain  $OH^-$ . Example:  $NH_3$ ,  $CaO$

**Bronsted-Lowry concept** (1923):

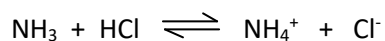
An acid is any molecule or ion that can donate a proton ( $H^+$ ). A base is any molecule or ion that can accept a proton ( $H^+$ ).



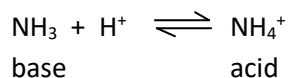


*Bronsted-Lowry concept is superior to Arrhenius concept:*

- (a) Much wider scope.
- (b) Not limited to aqueous solutions.

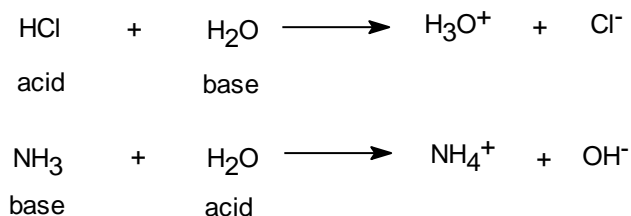


- (c) Release of  $\text{OH}^-$  not necessary to qualify as a base.



"An acid is a proton donor, while a base is a proton acceptor."

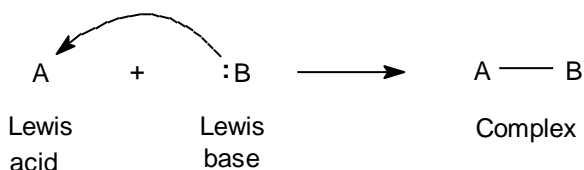
**Amphiprotic substances:** Molecule or ions that behave both as Bronsted acid and Bronsted base is said to be amphiprotic. For example,  $\text{H}_2\text{O}$ .



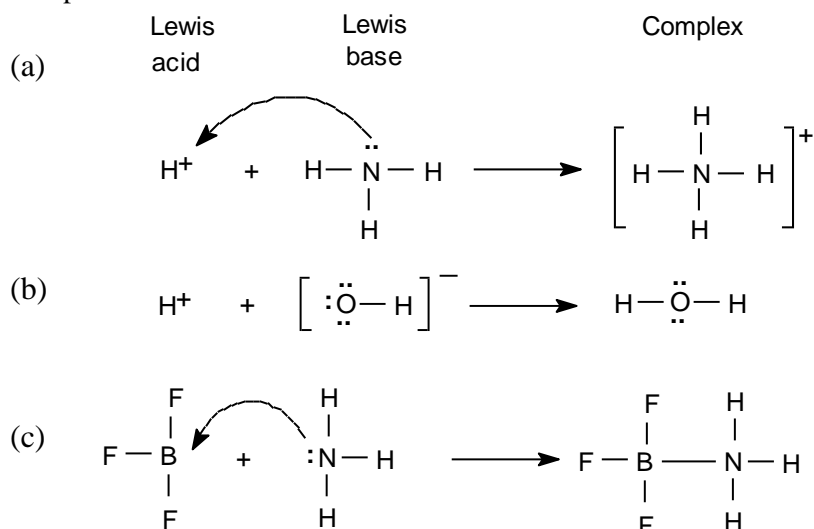
Other Examples:  $\text{HS}^-$ ,  $\text{Al}(\text{H}_2\text{O})_3(\text{OH})_3$  etc. (both with  $\text{OH}^-$  as acid, with  $\text{H}_3\text{O}^+$  as base)

**Lewis concept (1930):**

An acid is a substance (molecule or ions) which can accept a pair of electrons. A base is a substance (molecule or ions) which can donate a pair of electrons.



The combination of Lewis acid and Lewis base is called a complex. All cations or molecules short of an electron pair act as Lewis acids; and all anions or molecules having a lone pair of electron act as Lewis bases. Examples:



### Strong and Weak Acids and Bases

Acids and bases are classified as strong or weak, depending on whether they are strong or weak electrolytes. A **strong acid** is *an acid that ionizes completely in water; it is a strong electrolyte*. Hydrochloric acid,  $\text{HCl}(\text{aq})$ , and nitric acid,  $\text{HNO}_3(\text{aq})$ , are examples of strong acids.

A **weak acid** is *an acid that only partly ionizes in water; it is a weak electrolyte*. Examples of weak acids are hydrocyanic acid,  $\text{HCN}(\text{aq})$ , and hydrofluoric acid,  $\text{HF}(\text{aq})$ .

A **strong base** is *a base that is present in aqueous solution entirely as ions, one of which is  $\text{OH}^-$ ; it is a strong electrolyte*. The ionic compound sodium hydroxide,  $\text{NaOH}$ , is an example of a strong base. It dissolves in water as  $\text{Na}^+$  and  $\text{OH}^-$ . The hydroxides of Groups IA and IIA elements, except for beryllium hydroxide, are strong bases (see Table 3).

A **weak base** is *a base that is only partly ionized in water; it is a weak electrolyte*. Ammonia,  $\text{NH}_3$ , is an example.

**TABLE 3. Common Strong and Weak Acids and Bases**

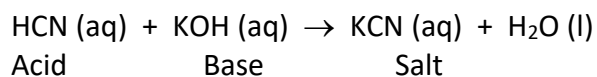
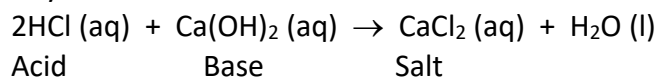
Strong Acids	Strong Bases	Weak Acids	Weak Bases
$\text{HClO}_4$	$\text{LiOH}$	$\text{HCN}$	$\text{NH}_3$
$\text{H}_2\text{SO}_4$	$\text{NaOH}$	$\text{HF}$	$\text{Mg}(\text{OH})_2$
$\text{HI}$	$\text{KOH}$	$\text{CH}_3\text{COOH}$	$\text{Al}(\text{OH})_3$
$\text{HBr}$	$\text{Ca}(\text{OH})_2$	$\text{H}_2\text{CO}_3$	$\text{Na}_2\text{CO}_3$
$\text{HCl}$	$\text{Sr}(\text{OH})_2$	$\text{H}_3\text{PO}_4$	$\text{NaHCO}_3$
$\text{HNO}_3$	$\text{Ba}(\text{OH})_2$	$\text{HCOOH}$	$\text{CH}_2\text{NH}_2$
		$\text{HNO}_2$	$\text{C}_5\text{H}_5\text{N}$

Non-metal oxides are acidic. Example:  $\text{SO}_3 + \text{H}_2\text{O} = \text{H}_2\text{SO}_4$

Metal oxides are basic. Example:  $\text{CaO} + \text{H}_2\text{O} = \text{Ca}(\text{OH})_2$

## Neutralization Reactions

One of the chemical properties of acids and bases is that they neutralize one another. A **neutralization reaction** is a reaction of an acid and a base that results in an ionic compound and possibly water. When a base is added to an acid solution, the acid is said to be neutralized. The ionic compound that is a product of a neutralization reaction is called a **salt**. Most ionic compounds other than hydroxides and oxides are salts.



The salt formed in a neutralization reaction consists of cations obtained from the base and anions obtained from the acid. In the first example, the base is  $\text{Ca(OH)}_2$ , which supplies  $\text{Ca}^{2+}$  cations; the acid is  $\text{HCl}$ , which supplies  $\text{Cl}^-$  anions. The salt contains  $\text{Ca}^{2+}$  and  $\text{Cl}^-$  ions ( $\text{CaCl}_2$ ).

## Oxidation-Reduction Reactions

**Redox** reactions include all chemical reactions in which atoms have their oxidation state changed; in general, redox reactions involve the transfer of electrons between species. This can be either a simple process, such as the oxidation of carbon to yield carbon dioxide ( $\text{CO}_2$ ) or the reduction of carbon by hydrogen to yield methane ( $\text{CH}_4$ ), or a complex process such as the oxidation of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) in the human body through a series of complex electron transfer processes.

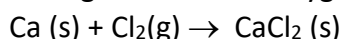
[**Oxidation state** (or **oxidation number**) of an atom in a substance is the actual charge of the atom if it exists as a monatomic ion, or a hypothetical charge assigned to the atom in the substance by simple rules. An oxidation–reduction reaction is one in which one or more atoms change oxidation number, implying that there has been a transfer of electrons.]

The term "redox" comes from two concepts involved with electron transfer: reduction and oxidation. It can be explained in simple terms:

**Oxidation** is the *loss* of electrons or an *increase* in oxidation state by a molecule, atom, or ion.

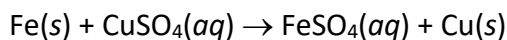
**Reduction** is the *gain* of electrons or a *decrease* in oxidation state by a molecule, atom, or ion.

Formerly, the term *oxidation* meant "reaction with oxygen." The current definition greatly enlarges the meaning of this term. Consider the reaction of calcium metal with chlorine gas; the reaction looks similar to the burning of calcium in oxygen. The chemical equation is:

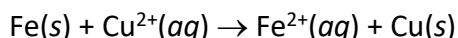


In this reaction, the calcium atom is oxidized, because it increases in oxidation number (from 0 to +2). Chlorine is reduced; it decreases in oxidation number from 0 to –1. This is clearly an oxidation–reduction reaction that does not involve oxygen.

As a simple example of an oxidation–reduction reaction, let us look at what happens when you dip an iron nail into a blue solution of copper(II) sulfate. What you see is that the iron nail becomes coated with a reddish-brown tinge of metallic copper. The molecular equation for this reaction is



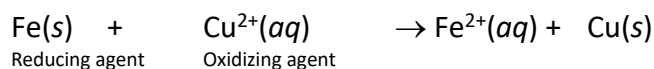
The net ionic equation is



The electron-transfer aspect of the reaction is apparent from this equation. Note that each iron atom in the metal loses two electrons to form an iron(II) ion, and each copper(II) ion gains two electrons to form a copper atom in the metal. The net effect is that two electrons are transferred from each iron atom in the metal to each copper(II) ion.

**Oxidation** is the half-reaction in which there is a loss of electrons by a species (or an increase of oxidation number of an atom). **Reduction** is the half-reaction in which there is a gain of electrons by a species (or a decrease in the oxidation number of an atom). Thus, the equation  $\text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2e^-$  represents the oxidation half-reaction, and the equation  $\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s)$  represents the reduction half-reaction.

A species that is oxidized (loses electrons or contains an atom that increases in oxidation number) and a species that is reduced (gains electrons or contains an atom that decreases in oxidation number). An **oxidizing agent** is a species that gain electrons and oxidizes another species; it is itself reduced. Similarly, a **reducing agent** is a species that loss electrons and reduces another species; it is itself oxidized.



**Importance of Oxidation-Reduction:** Many natural processes do not happen without oxidation-reduction reactions:

- ❖ Carbohydrate (food) is produced with oxygen by *reducing* carbon dioxide in presence of water and sunlight energy (**Photosynthesis**).
- ❖ Nutrients are oxidized in the cells (plants & animals) to produce energy (**Respiration**),
- ❖ Fossil fuels such as oil, coal and gasoline are oxidized to produce energy in the internal combustion engines to run vehicles,
- ❖ Metal oxides are reduced into metals, for example, iron ores are reduced to produce iron metals from its ores.