

3. Chemical Reaction

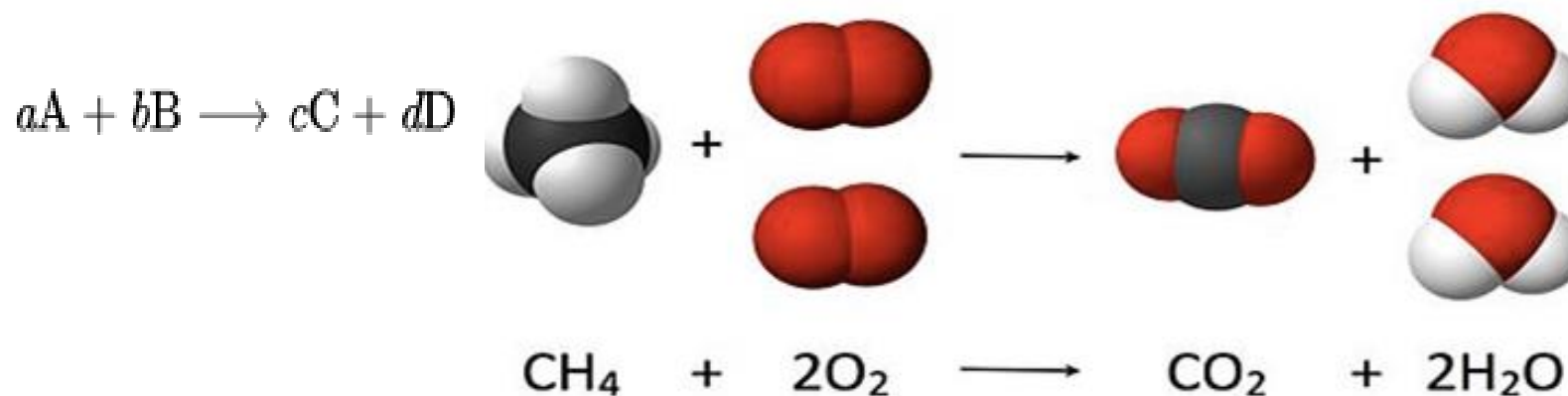
A **chemical reaction** is a process that leads to the chemical transformation of one set of chemical substances to another.

The substance (or substances) initially involved in a chemical reaction are called reactants or reagents. Chemical reactions are usually characterized by a chemical change, and they yield one or more products, which usually have properties different from the reactants. Chemical reactions are described with chemical equations, which symbolically present the starting materials, end products, and sometimes intermediate products and reaction conditions.

Reactions may proceed in the forward or reverse direction until they go to completion or reach equilibrium. Reactions that proceed in the forward direction to approach equilibrium are often described as spontaneous, requiring no input of free energy to go forward. Non-spontaneous reactions require input of free energy to go forward (examples include charging a battery by applying an external electrical power source).

Chemical equations

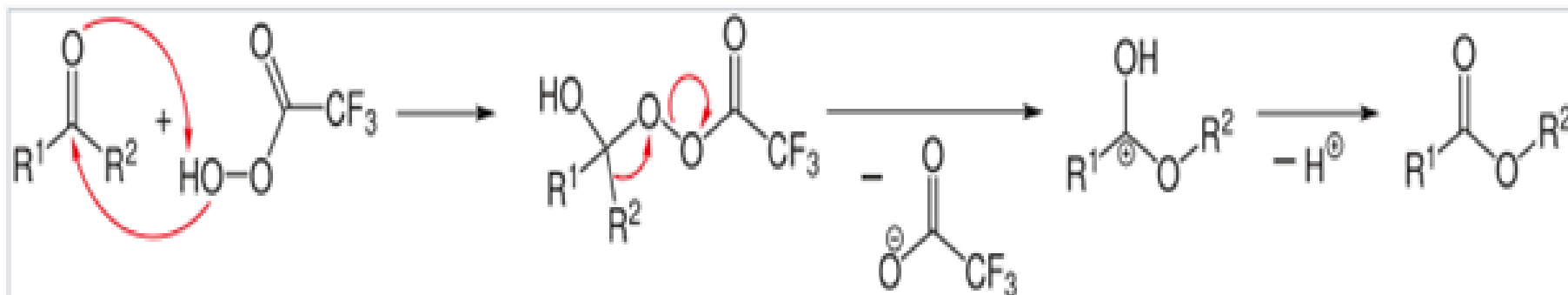
Chemical equations are used to graphically illustrate chemical reactions. They consist of chemical or structural formulas of the reactants on the left and those of the products on the right. They are separated by an arrow ($\rightarrow$) which indicates the direction and type of the reaction; the arrow is read as the word "yields". The tip of the arrow points in the direction in which the reaction proceeds. A double arrow ($\rightleftharpoons$) pointing in opposite directions is used for equilibrium reactions. Equations should be balanced according to the stoichiometry, the number of atoms of each species should be the same on both sides of the equation. This is achieved by scaling the number of involved molecules (A, B, C and D) in a schematic example below) by the appropriate integers a , b , c and d .



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As seen from the equation $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$, a coefficient of 2 must be placed before the **oxygen** gas on the reactants side and before the **water** on the products side in order for, as per the law of conservation of mass, the quantity of each element does not change during the reaction

More elaborate reactions are represented by reaction schemes, which in addition to starting materials and products show important intermediates or transition states. Also, some relatively minor additions to the reaction can be indicated above the reaction arrow; examples of such additions are water, heat, illumination, a catalyst, etc. Similarly, some minor products can be placed below the arrow, often with a minus sign.

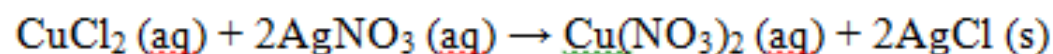


An example of organic reaction: **oxidation** of **ketones** to **esters** with a **peroxycarboxylic acid**

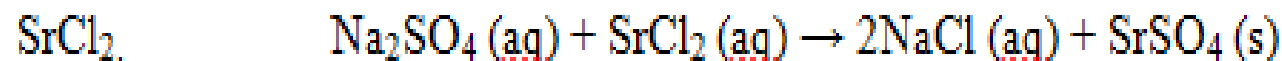
A single-replacement reaction is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products. For example,

$2\text{HCl (aq)} + \text{Zn (s)} \rightarrow \text{ZnCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$ is an example of a single-replacement reaction. The hydrogen atoms in HCl are replaced by Zn atoms, and in the process a new element—hydrogen—is formed.

A double-replacement reaction occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is



A precipitation reaction occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of solution as a solid precipitate. The formation of a solid precipitate is the driving force that makes the reaction proceeds. For example, consider the possible double-replacement reaction between Na_2SO_4 and SrCl_2 .



A net ionic equation is a chemical equation in which the dissolved ionic compounds are written as separated ions

For example, when NaCl(aq) reacts with $\text{AgNO}_3\text{(aq)}$ in a double-replacement reaction to precipitate AgCl(s) and form $\text{NaNO}_3\text{(aq)}$, the complete ionic equation includes NaCl , AgNO_3 , and NaNO_3 written as separated ions:



The reaction of an acid and a base is called a **neutralization reaction**. Although acids and bases have their own unique chemistries, the acid and base cancel each other's chemistry to produce a rather innocuous substance—water. In fact, the general reaction between an acid and a base is

$\text{acid} + \text{base} \rightarrow \text{water} + \text{salt}$ where the term salt is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. (In chemistry, the word *salt* refers to more than just table salt.) For example, the balanced chemical equation for the reaction between HCl(aq) and KOH(aq) is



Acid-base reactions

In the Brønsted–Lowry acid–base theory, an acid-base reaction involves a transfer of protons (H^+) from one species (the acid) to another (the base). When a proton is removed from an acid, the resulting species is termed that acid's conjugate base. When the proton is accepted by a base, the resulting species is termed that base's conjugate acid. In other words, acids act as proton donors and bases act as proton acceptors according to the following equation:



The reverse reaction is possible, and thus the acid/base and conjugated base/acid are always in equilibrium. The equilibrium is determined by the acid and base dissociation constants (K_a and K_b) of the involved substances. A special case of the acid-base reaction is the neutralization where an acid and a base, taken at exactly same amounts, form a neutral salt.

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Acid-base reactions can have different definitions depending on the acid-base concept employed.

Some of the most common are:

- Arrhenius definition: Acids dissociate in water releasing H_3O^+ ions; bases dissociate in water releasing OH^- ions.
- Brønsted-Lowry definition: Acids are proton (H^+) donors, bases are proton acceptors; this includes the Arrhenius definition.
- Lewis definition: Acids are electron-pair acceptors, bases are electron-pair donors; this includes the Brønsted-Lowry definition.
- Arrhenius concept is quite useful and explains the acid- base behavior to a good extent. However it has certain drawbacks like:
- It is limited to only aqueous solutions and require dissociation of the substance.
- It does not explain the acidic behavior of some substances which do not contain hydrogen for example, AlCl_3 . Similarly it does not explain the basic character of substances like NH_3 and Na_2CO_3 which do not have a hydroxide groups.

Brønsted and Lowry pointed out independently that acid-base reactions can be interpreted as proton-transfer reactions.

According to them, an acid is defined as a proton (H⁺) donor, and a base is defined as a proton acceptor.



The species NH₄⁺ and NH₃ are called **conjugate acid-base pair**.



• In the above example AlCl₃ is an electron deficient species. It accepts an electron pair from a molecule of NH₃ which has a lone pair of electrons on N atom.

• Thus, AlCl₃ is a lewis acid and NH₃ is a lewis base.

Balancing Chemical Equations

There is no sequence of rules that can be followed blindly to get a balanced chemical equation.

Manipulate the coefficients written in front of the formulas of the reactants and products until the number of atoms of each element on both sides of the equation are the same.

It is usually a good idea to tackle the easiest part of a problem first.

Example: Consider what happens when propane (C_3H_8) burns in air to form CO_2 and H_2O . The first thing to look for when balancing equations are relationships between the two sides of an equation.



It is easier to balance the carbon and hydrogen atoms in the equation than the oxygen atoms in this reaction because all of the carbon atoms in propane end up in CO_2 and all of the hydrogen atoms end up in H_2O , but some of the oxygen atoms end up in each compound. This means that there is no way to predict the number of O_2 molecules that are consumed in this reaction until you know how many CO_2 and H_2O molecules are produced.

We can start by noting that there are three carbon atoms in each C_3H_8 molecule. Thus, three CO_2 molecules are formed for every C_3H_8 molecule consumed.



If there are eight hydrogen atoms in each C_3H_8 molecule, there must be eight hydrogen atoms, or four H_2O molecules, on the right side of the equation.



Now that the carbon and hydrogen atoms are balanced, we can try to balance the oxygen atoms. There are six oxygen atoms in three CO₂ molecules and four oxygen atoms in four H₂O molecules. To balance the 10 oxygen atoms in the products of this reaction we need five O₂ molecules among the reactants.



There are now three carbon atoms, eight hydrogen atoms, and 10 oxygen atoms on each side of the equation. The balanced equation for this reaction is therefore written as follows.



Oxidation- Reduction Equilibria

What is oxidation?

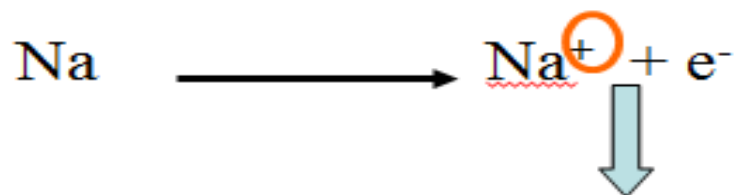
Old definition:

Combination of substance with oxygen



Current definition:

Loss of **E**lectrons is **O**xidation (**LEO**)



Positive charge represents electron deficiency
One positive charge means deficient by one electron

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What is reduction?

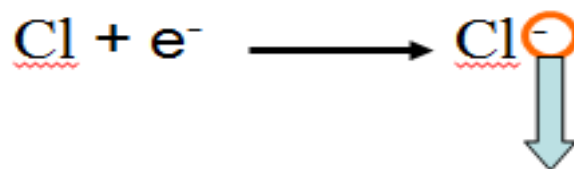
Old definition:

Removal of oxygen from a compound



Current definition:

Gain of Electrons is Reduction (GER)

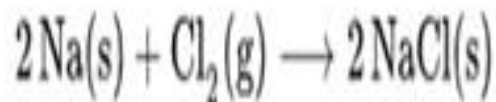


Negative charge represents electron richness
One negative charge means rich by one electron

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- Oxidation and reduction go hand in hand. In a reaction, if there is an atom undergoing oxidation, there is probably another atom undergoing reduction.
- When there is an atom that donates electrons, there is always an atom that accepts electrons.
- Rules for assigning Os (oxidation state):
 - ✓ The sum of the oxidation numbers of all of the atoms in a molecule or ion must be equal in sign and value to the charge on the molecule or ion.

In the following redox reaction, hazardous sodium metal reacts with toxic chlorine gas to form the ionic compound sodium chloride, or common table salt:



In the reaction, sodium metal goes from an oxidation state of 0 (as it is a pure element) to +1: in other words, the sodium lost one electron and is said to have been oxidized. On the other hand, the chlorine gas goes from an oxidation of 0 (it is also a pure element) to -1: the chlorine gains one electron and is said to have been reduced. Because the chlorine is the one reduced, it is considered the electron acceptor, or in other words, induces oxidation in the sodium – thus the chlorine gas is considered the oxidizing agent. Conversely, the sodium is oxidized or is the electron donor, and thus induces reduction in the other species and is considered the *reducing agent*.

Find the OS of N in NH_4^+

Let the OS of each N be = x

OS of H = +1 (Remember H belongs to Gp. 1A)

So we have,

$$(\text{OS of N}) + 4 (\text{OS of H}) = +1$$

$$(x) + 4 (+1) = +1$$

$$(x) + (4) = +1$$

$$(x) = -3$$

Find the OS of each C in $\text{C}_2\text{O}_4^{2-}$

Let the OS of each C be = x

OS of O = -2 (Remember O belongs to Gp. 6A)

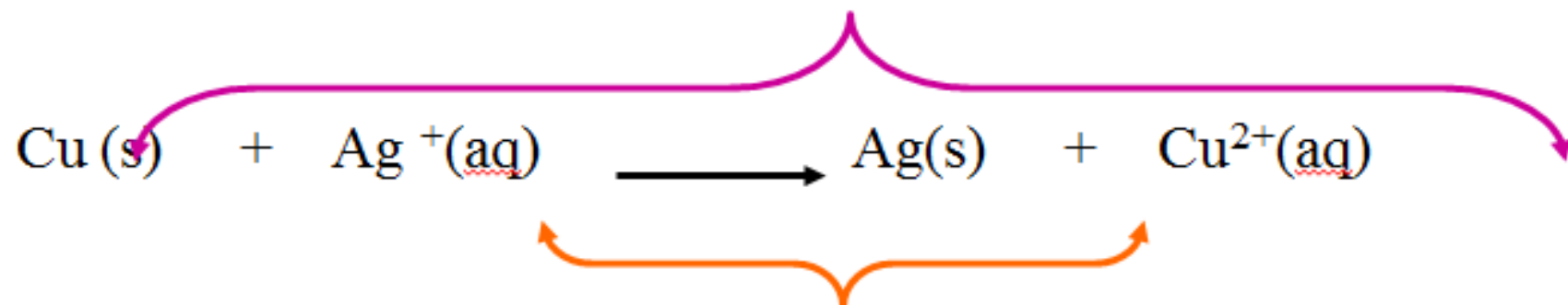
So we have,

$$2(\text{OS of C}) + 4 (\text{OS of O}) = -2$$

$$2(x) + 4 (-2) = -2$$

$$2(x) + (-8) = -2, x = +3$$

Balancing simple redox reactions



Step 1: Pick out similar species from the equation



Step 2: Balance the equations individually for charges and number of atoms



Balancing simple redox reactions



$\text{Cu}^0(\text{S})$ becomes $\text{Cu}^{2+}(\text{aq})$ by **loosing 2 electrons**
So $\text{Cu}^0(\text{S})$ getting oxidized to $\text{Cu}^{2+}(\text{aq})$ is the **oxidizing half reaction**.



$\text{Ag}^+(\text{aq})$ becomes $\text{Ag}^0(\text{S})$ by **gaining 1 electron**.
So $\text{Ag}^+(\text{aq})$ getting reduced to $\text{Ag}^0(\text{S})$ is the **reducing half reaction**.

Final Balancing act:

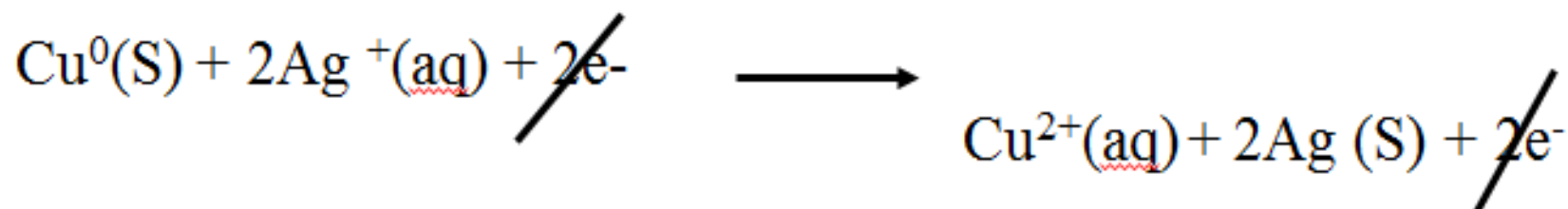
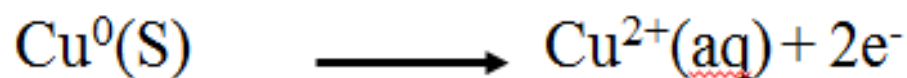
Making the number of electrons equal in both half reactions



So we have,



Balancing simple redox reactions

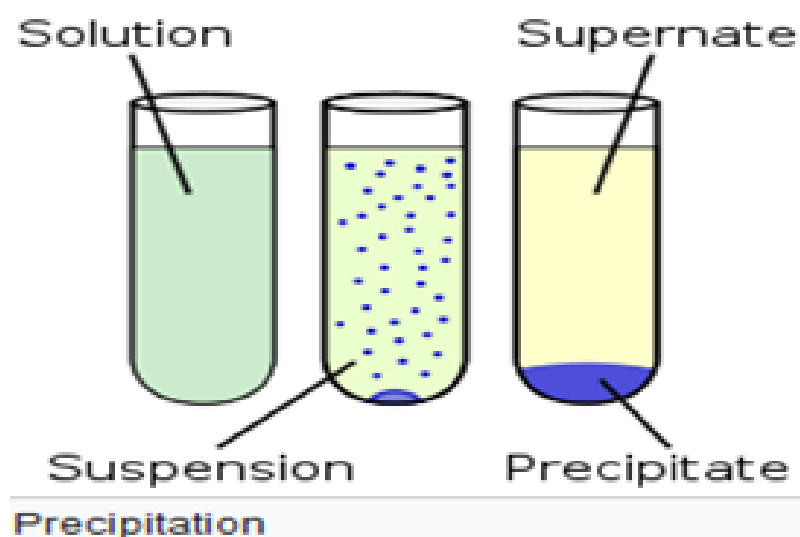


Number of e^- s involved in the overall reaction is 2

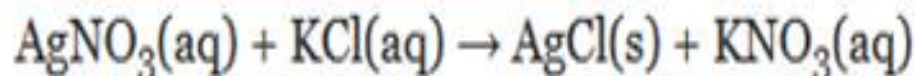
Precipitation Reactions

Precipitation reactions occur when cations and anions in aqueous solution combine to form an insoluble ionic solid called a **precipitate**. Whether or not such a reaction occurs can be determined by using the **solubility rules** for common ionic solids. Because not all aqueous reactions form precipitates, one must consult the solubility rules before determining the state of the products and writing a **net ionic equation**.

It usually takes place when the concentration of dissolved ions exceeds the solubility limit and forms an insoluble salt. This process can be assisted by adding a precipitating agent or by removal of the solvent.

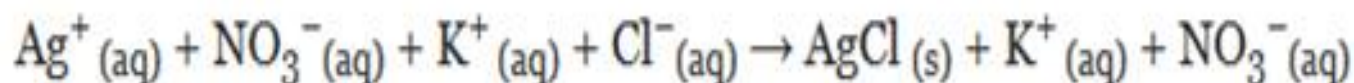


The reaction between silver nitrate and potassium chloride is a precipitation reaction because solid silver chloride is formed as a product.



The reaction may be recognized as a precipitation because two ionic aqueous solutions (aq) react to yield a solid product (s).

It's common to write precipitation reactions in terms of the ions in the solution. This is called a complete ionic equation:



Net ionic equation for this reaction =?

Thank You All!!