

3 Major types of reactions in aqueous solutions:

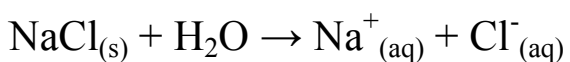
-Acid-Base Reactions

-Precipitation Reactions (double displacement)

-Oxidation-Reduction (not limited to aqueous reactions)

Solutions:

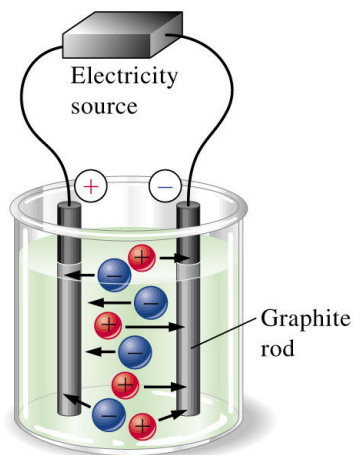
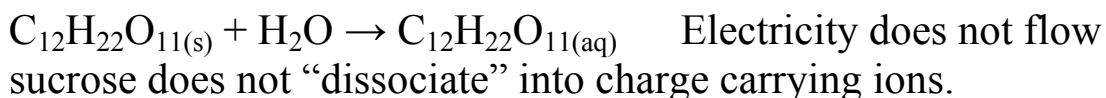
Electrolyte:



Electricity Flows

Ions are charge carriers

Non-electrolyte:

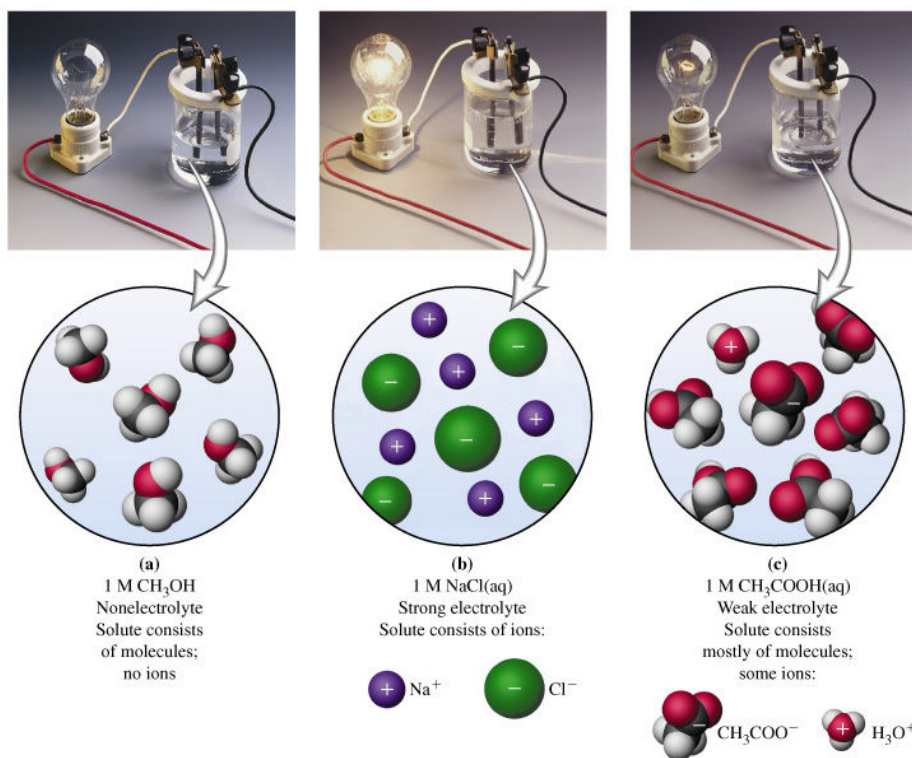


[Electrolyte vs nonelectrolyte video clip](#)

Arrhenius's Theory of Electrolytic Dissociation

Certain substances dissociate into cations (+) and anions (-) when they dissolve in water.

Nonelectrolytes, strong electrolytes and weak electrolytes are determined by their ability to dissociate in solution.



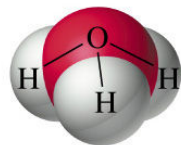
It is mostly soluble ionic compounds and “strong” acids that constitute strong electrolytes (see solubility table and strong acid chart).

[Strong vs Weak electrolyte video clip](#)

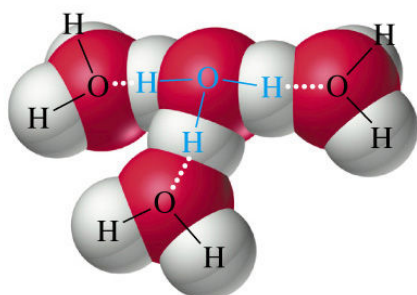
Example (4.1a): Seawater is essentially 0.438M NaCl and 0.0512M MgCl_2 , together with several other minor solutes. What are $[\text{Na}^+]$, $[\text{Mg}^{2+}]$ and $[\text{Cl}^-]$ in water?

Strong vs Weak Acids and bases:

Acids produce hydronium, H_3O^+ (sometimes represented with just H^+), ions in solution



Hydronium ion
 H_3O^+



A hydrated proton
 H_9O_4^+

TABLE 5.1 Common Strong Acids and Strong Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HClO_4	RbOH
HNO_3	CsOH
$\text{H}_2\text{SO}_4^{\text{a}}$	$\text{Ca}(\text{OH})_2$
	$\text{Sr}(\text{OH})_2$
	$\text{Ba}(\text{OH})_2$

^a H_2SO_4 ionizes in two distinct steps. It is a strong acid only in its first ionization step (see Section 17-6).

Some acids will go through multiple ionization phases, but they get progressively weaker with each proton lost. Ex. H_2SO_4

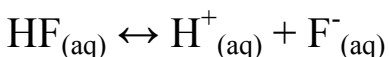
[Acid nature video clip](#)

Strong bases tend to be alkali metal hydroxides and some of the alkaline earth metal hydroxides. Weak bases tend to be less soluble hydroxides and organic bases including the amine functional group.

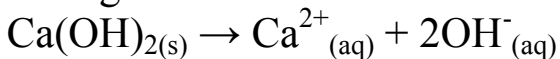
Strong Acid



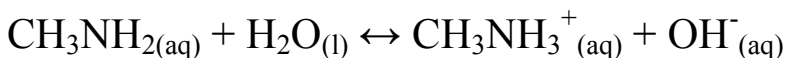
Weak Acid



Strong Base

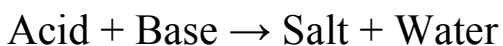


Weak Base



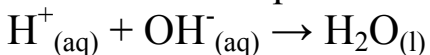
Neutralization reactions between acids and bases

General formula:



The “net” ionic equation only lists the components of the reaction that take part in the reaction itself (i.e. it excludes spectator ions)

The net ionic equation for all strong acid/base neutralizations is



Exercise 4.2:

Calcium hydroxide is used to neutralize a waste stream of hydrochloric acid. Write a) a complete formula equation, b) and ionic equation and c) a net ionic equation for this neutralization reaction.

Acid-Base Titrations

Analytical method of determining the unknown concentration of an acid or base from a neutralization reaction using a solution of precisely known concentration (standardized solution)

The solution being titrated is known as the “titrant”

A pH indicator is generally used to visually monitor the progress of the reaction.

The **endpoint** of a reaction is that point when a pH indicator changes color.

The **equivalence point** of a reaction is that point when equal moles of acid and base have reacted in their exact stoichiometric proportions.

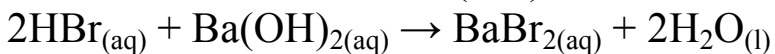
The endpoint and the equivalence point are usually, but not always, set up to occur at the same time.



Examples:

Exercise 4.3A

What volume of 0.01060M $\text{HBr}_{(\text{aq})}$ is required to neutralize 25.00mL of 0.01580M $\text{Ba}(\text{OH})_2$?



Exercise 4.3B

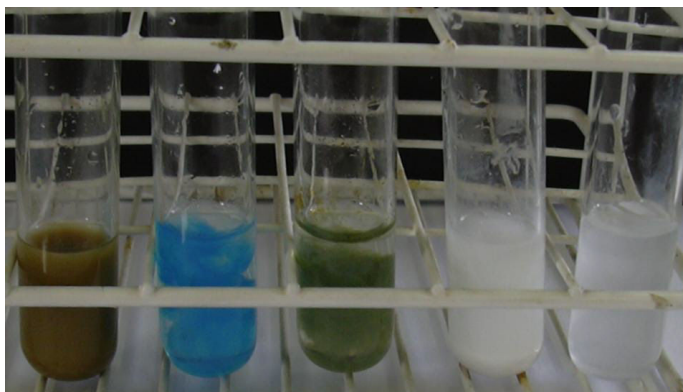
A 2.000g sample of sulfuric acid solution that is 96.5% H_2SO_4 by mass is dissolved in a quantity of water and titrated. What volume of 0.3580M $\text{KOH}_{(\text{aq})}$ is required for the titration? Assume that at the equivalence point of the titration the solution is $\text{K}_2\text{SO}_{4(\text{aq})}$

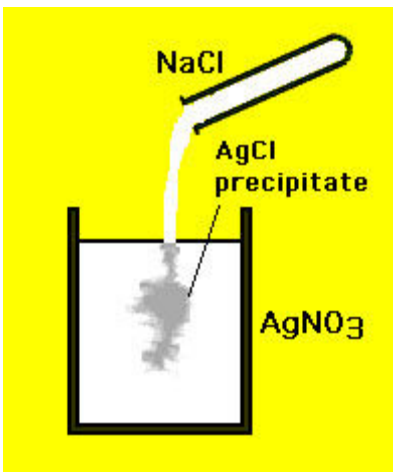
(Note that even though the second ionizing hydrogen on the H_2SO_4 is weak, it will be completely reacted at the end of the titration.)

Precipitations Reactions

a.k.a. Double displacement reactions or metathesis reactions

Reactions in which two solutions containing soluble compounds are combined to produce at least one insoluble product.



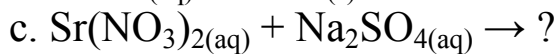
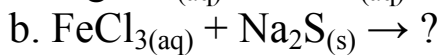
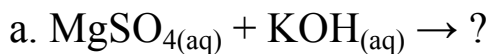


There must be a net ionic equation possible (i.e. everything can't end up being spectator ions)

Examples:

Exercise 4.6

Predict whether a reaction will occur in each of the following cases. If so, write a net ionic equation for the reaction.



Oxidation-Reduction Reactions

Oxidation-reduction (redox) reactions differ from the aqueous reactions studied so far (precipitation and neutralization) in that there must be a transfer of electrons that takes place (i.e. a change in the oxidation number at least some of the substances)

More generally, redox reactions can be classified as those in which:

Oxidation	Reduction
Loses electrons	Gains electrons
Gains oxygen	Loses oxygen
Loses hydrogen	Gains hydrogen

Oxidation numbers are assigned to all the elements in compounds and ions to keep track of the change in oxidation number as the reaction progresses from reactants to products.

The redox number is often times the true expected charge on a species, but other times it represents more of a “virtual charge” that is assigned by convention to the elements in a substance.

Assigning oxidation numbers requires you to know the rules for assigning oxidation numbers (see handout).

Redox reactions must be balanced in order to ensure that the number of electrons lost is equal to the number of electrons gained.

Often, a redox reaction will be split into two “half-reactions”, and the oxidation and reduction halves will be balanced independent of one another. This often times occurs in acidic or basic medium

and there is a specific set of rules for dealing with each situation (see handout).

Since the number of electrons lost and gained must always be the same in a redox reaction, so you may need to multiply through each half reaction by a number so that if you add the half reactions back together to get the *net ionic equation*, the electrons will cancel out.

Things to remember:

Redox reactions always occur in pairs

The charge on any element is zero.

The species being reduced is the oxidizing agent: The species being oxidized is the reducing agent.

Remember the saying: **LEO** goes **GER**

Loses **E**lectrons **O**xidized

Gains **E**lectrons **R**educed

Oxidation and reduction must occur in pairs. (You can't have an oxidation without a reduction and vice versa)

Charge must be conserved in the chemical reaction.

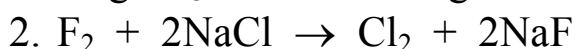
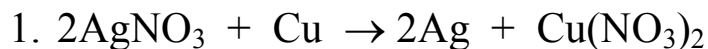
In a *disproportionation* reaction, one substance undergoes both oxidation and reduction.

A **reducing agent** is a substance that reduces another substance. It itself is oxidized.

Conversely, an **oxidizing agent** oxidizes another species and is itself reduced.

When referring to an "agent", usually the entire compound is identified, not just the component oxidized or reduced.

Example: Determine which element is oxidized and which is reduced in the following reactions:



Answers: Silver is reduced ($1^+ \rightarrow 0$), copper is oxidized ($0 \rightarrow 2^+$); Fluorine is reduced ($0 \rightarrow 1^-$), chlorine is oxidized ($1^- \rightarrow 0$).

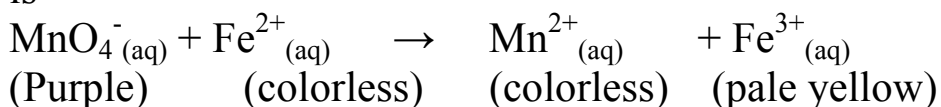
Examples: Assign oxidation numbers to each element in the reaction and balance.



(Note: O is a 1^- in hydrogen peroxide)

Example:

We wish to analyze an iron ore for its iron content. The iron in the sample can be converted quantitatively to the iron(II) ion, Fe^{2+} , in aqueous solution, and this solution can then be titrated with aqueous potassium permanganate, KMnO_4 . The equation for the reaction occurring in the course of this titration (in acidic solution) is

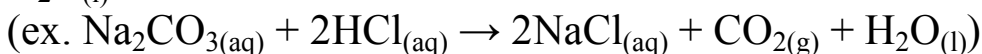


A 1.026g sample of iron containing ore requires 24.35mL of 0.0195M KMnO_4 to reach the equivalence point. What is the mass percent of iron in the ore?

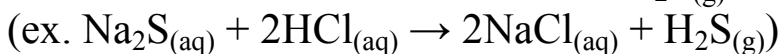
Other reactions:

Gas-forming reactions

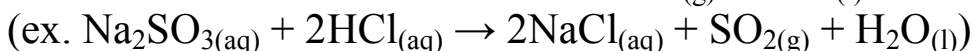
Metal carbonate or bicarbonate + acid \rightarrow metal salt + $\text{CO}_{2(g)}$ + $\text{H}_2\text{O}_{(l)}$



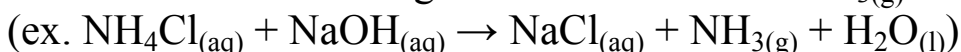
Metal sulfide + acid \rightarrow metal salt + $\text{H}_2\text{S}_{(g)}$



Metal sulfite + acid \rightarrow metal salt + $\text{SO}_{2(g)}$ + $\text{H}_2\text{O}_{(l)}$



Ammonium salt + strong base \rightarrow metal salt + $\text{NH}_{3(g)}$ + $\text{H}_2\text{O}_{(l)}$



Driving Forces:

Reaction Type:

Driving Force

Precipitation

Formation of an insoluble compound

Acid-Base Neutralization

Formation of a salt and water (proton transfer)

Gas-Forming

Evolution of water-insoluble gas such as CO_2

Oxidation-Reduction

Electron transfer