Chapter 4

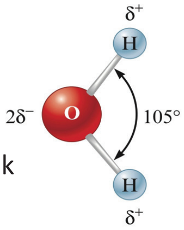
Types of Chemical Reactions and Solution Stoichiometry

Water

- One of the most important substances on Earth.
- One of the most abundant substances on Earth.
- Can dissolve many different substances.

  It's not that it is super versatile.

  It's just that it is everywhere, in every nook and cranny of the Earth
- It is absolutely essential to life directly and indirectly (in some surprising ways)
- A polar molecule because of its unequal (asymmetric, lopsided) charge distribution.



Water has been the natural choice of solvent to conduct solution chemistry because it is:

- easy to find, almost free, and cheap to purify
- easy to work with
- hugely important in the chemistry of the planet

If a solid is soluble in water, it means water molecules are able to provide favorable (attractive) interactions, competitive with the attractive interactions within the solid.

Remarkably, water molecules can often compete with the strong attractions present between the cations and anions of an ionic solid.

➢ When it does, that solid is soluble in water.

Aqueous Solutions

- **Solvent**

  liquid water
- **Solute**

  substance being dissolved
- **Electrolyte**

  substance that, when dissolved in a polar solvent like water, produces a solution that can conduct electricity

  (because it produces ions, which are charged)

Electrolytes

Strong electrolyte:

When dissolved in water, dissociates almost completely into ions

- Highly water-soluble ionic compounds (e.g. NaCl, K_2SO_4)
- Strong acids (e.g. HCl, H_2SO_4)
- Strong bases (e.g. KOH)

In this course, **we will assume that all ionic compounds dissolve as separated cations and anions**

Weak electrolyte:

When dissolved in water, only some of it produces ions

- Weak acids (e.g. $HC_2H_3O_2$)
- Weak bases (e.g. NH_3)

Non-electrolyte:

Does not dissociate into ions at all (e.g. sugars, alcohols, oils)

Electrolyte solutions conduct electricity

ions can carry the electric charge though the solution

KCl
High conductivity

Strong Electrolytes
conduct current efficiently (bulb shines brightly at high concentration)

Acetic Acid
Low conductivity

Weak Electrolytes
conduct only a small current (bulb glows dimly even at high concentration)

Sugar
No conductivity

Nonelectrolytes
conduct no current (bulb remains unlit).

Be aware:

The terms “strong electrolyte” or “weak electrolyte” refers to the **<u>substance</u>** dissolved (i.e. solute)

- A concentrated solution of a strong electrolyte is a good conductor and can make the bulb glow bright.

But …

- A <u>dilute</u> solution of a strong electrolyte won’t conduct much electricity, and the bulb will <u>not</u> glow bright.
  - ➢ Even if the dissolved “strong electrolyte” fully dissociated, if there isn’t much of it, it will still have low conductivity

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On the other hand,

Even a concentrated solution of a weak electrolyte is not a good conductor and will not make the bulb glow bright.

A molecular substance like hydrogen chloride (HCl), which is <u>not</u> ionic, but fully dissociates into ions when dissolved in water, is <u>also</u> a strong electrolyte.

- ➢ Because it’s a strong acid, fully dissociating into $H^+(aq)$ and $Cl^-(aq)$ ions

Note: Acids produce H^+ *ions in water, along with an anion*

Strong acids are strong electrolytes
-- fully ionize in water

Weak acids are weak electrolytes
-- only partially ionize in water

Almost all the bases we see in this course are **metal hydroxides** like NaOH or KOH, and are strong electrolytes (<u>because they are ionic compounds</u>).

- ➢ Remember: we will assume that all ionic compounds fully dissociate into their ions when dissolved

Note: Bases produce OH⁻ ions in water, along with a cation

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

One **weak base** we encounter is ammonia: $\mathbf{NH_3}$

- **A <u>molecular</u> compound** (as opposed to the usual metal hydroxide like NaOH)
  -- no guarantee it will fully ionize
  -- indeed it doesn’t
- A small portion reacts with water to form NH_4^+ and OH^- ← Produced by reacting with water
- Not fully converted to ions
  -- so it is a **weak electrolyte**

Chemical Reactions in Solutions

To do anything useful, we must know:

- The nature of the reaction.
- The amounts of chemicals present in the solutions.

Concentration of Solutions: **Molarity**

Molarity (M) = moles of solute per 1 liter of solution

$$M = \textbf{Molarity} = \frac{\textbf{moles of solute}}{\textbf{volume of solution (in L)}} = \frac{n}{V}$$

Unit of Molarity: “Molar” (M)

Slightly confusing:
Unit symbol M is the same as the symbol for the concept

If 2.0 liters of a solution contains 6.0 moles of solute, its concentration is:

$$\frac{6.0 \text{ mol}}{2.0 \text{ L}} = 3.0 \text{ molar} \quad \text{or} \quad 3.0\ M$$

Example

A 500.0-g sample of potassium phosphate is dissolved in enough water to make 1.50 L of solution. What is the molarity of the solution?

Steps:

1. Convert mass to moles, using molar mass
2. Use formula for molarity

Molar mass of K_3PO_4 = (3)(39.10) + (1)(30.97) + (4)(16.00) = 212.27 g/mol (m.m.)

$$\text{Number of moles} = n = 500.0 \text{ g} \times \frac{1 \text{ mol}}{212.27 \text{ g}} = 2.36 \text{ mol}$$

$$\text{Molarity} = M = \frac{n}{V} = \frac{2.36 \text{ mol}}{1.50 \text{ L}} = 1.57 \text{ M (molar)}$$

Concentration of Ions
-- given the concentration of an ionic compound

- Multiply the concentration of the compound by the number of ions in the formula.
- Applies to cations, anions, i.e. all ions

For a 0.25 M $CaCl_2$ solution:

$$CaCl_2\,(aq) \longrightarrow Ca^{2+}\,(aq) + 2Cl^-\,(aq)$$

Molarities of Ca^{2+} and Cl^-:

$$M_{Ca^{2+}} = 1 \times 0.25\,M = 0.25\,M$$

$$M_{Cl^-} = 2 \times 0.25\,M = 0.50\,M$$

$$M_{ions} = (1+2) \times 0.25\,M = 0.75\,M$$

Example

Which of the following solutions has the greatest number of ions?

a) 400.0 mL of 0.10 M NaCl

M_{ions} = (1+1)(0.10 M) = 0.20 M

n_{ions} = M_{ions} V= (0.20) (400.0x10^{-3}) = 0.080 mol

b) 300.0 mL of 0.10 M $CaCl_2$

M_{ions} = (1+2)(0.10 M) = 0.30 M

n_{ions} = M_{ions} V= (0.30) (300.0x10^{-3}) = $\boxed{0.090 \text{ mol}}$

c) 200.0 mL of 0.10 M $FeCl_3$

M_{ions} = (1+3)(0.10 M) = 0.40 M

n_{ions} = M_{ions} V= (0.40) (200.0x10^{-3}) = 0.080 mol

d) 800.0 mL of 0.10 M sucrose

M_{ions} = 0 mol Not an electrolyte!

Note that finding the greatest number of **moles** is sufficient.

Multiplying each mol by 6.022 × 10^{23} wouldn’t affect the ranking

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Number of moles of solute is the product of molarity (M) and volume (V)

$$\underset{\uparrow}{n} = M \cdot V$$

Moles of solute

Dilution

The process of adding solvent to a concentrated solution to achieve the molarity desired for a particular solution.

- Dilution with water does not alter the numbers of moles of solute present.

Moles of solute **before** dilution = $n_1 = M_1 \cdot V_1$

Moles of solute **after** dilution = $n_2 = M_2 \cdot V_2$

$$n_1 = n_2 \Rightarrow \boxed{M_1 \cdot V_1 = M_2 \cdot V_2}$$

Concept Practice

A 0.50 *M* solution of sodium chloride in an open beaker sits on a lab bench. Which of the following would decrease the concentration of the salt solution?

a) Pour some of the solution down the sink drain.

b) Add more sodium chloride to the solution.

c) Let the solution sit in open air for a couple of days.

d) Add water to the solution.

e) At least two of the above would decrease the concentration of the salt solution.

Practice

What is the volume $\overset{V_1}{\downarrow}$ of a $\overset{M_1}{\downarrow}$ 2.00 *M* NaOH solution needed to make $\underset{V_2}{\uparrow}$ 150.0 mL of a $\underset{M_2}{\uparrow}$ 0.800 *M* NaOH solution?

$$M_1 V_1 = M_2 V_2$$

If in mL units, then V_1 will also be in mL

$$(2.00\ M)\ V_1 = (0.800\ M)\ (150.0\ \text{mL})$$

$$V_1 = \frac{(0.800\ M)(150.0\ \text{mL})}{(2.00\ M)} = 60.0\ \text{mL}$$

Types of Chemical Reactions

- Precipitation
- Acid–Base

Double-displacement reactions

Also called:

Double-replacement

Exchange

Methathesis

- Oxidation–Reduction

  A wide variety of reactions, including single-replacement reactions

- Decomposition
- Combination (Synthesis)

Precipitation Reactions

A **<u>double displacement</u>** reaction in which a <u>solid</u> forms and separates from the solution.

Precipitate (*verb*): to drop out of solution

Precipitate (*noun*): the solid that forms

We will consider precipitation of <u>ionic compounds</u>

Remember: When ionic compounds dissolve in water, the resulting solution contains the **separated ions**.

If ions from two soluble ionic compounds

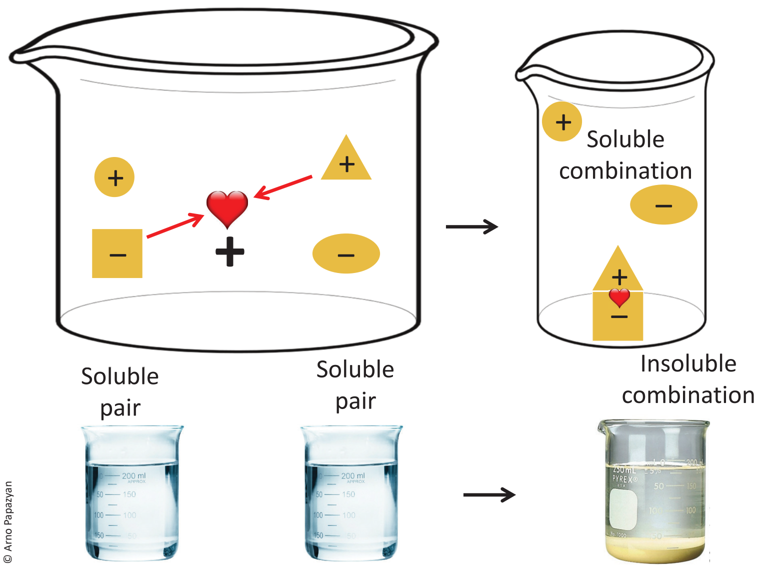
<u>cation</u> from one, ***<u>anion</u>*** from the other

correspond to an **<u>insoluble</u>** compound

they combine and <u>precipitate</u>

if they are in the same solution

- ➢ it doesn't matter if there are soluble combinations
- ➢ **insoluble combination wins**
  - ➢ **Because these ions are "happier" together in solid form**
  - ➢ **That's why that combination is insoluble**



Terms related to precipitation

- *Soluble:* solid dissolves in solution; (*aq*) is used in the reaction equation.
- *Insoluble:* solid does not dissolve in solution; (*s*) is used in reaction equation.
- *Insoluble* and *slightly soluble* are often used interchangeably. We will do that too.
  - What's called "soluble" or "insoluble" may differ slightly, depending on the industry, discipline, or preferences of an instructor
- *Marginally soluble:* considered "soluble", but barely

  -- cannot have concentrated solutions

Solubility Rules for cation-anion pairs

Before you get scared by a bunch of rules to memorize, observe that the solubility rules are ordered in decreasing solubility:

Always soluble

Less soluble as we go down the list of rules

"Always" **in**soluble

Always **soluble** cations (*with any anion!*)
- They never precipitate!

Always **soluble** anions (*with any cation!*)
- They never precipitate!

Generally **soluble** anions
-- with exceptions

Generally **insoluble** anions
-- with exceptions

"Others": **insoluble** anions
-- except with "always soluble" cations

Solubility Rules for cation-anion pairs

1. Alkali metal (**Group 1**) cations and **NH_4^+** are always soluble
2. **Nitrate** (& **perchlorate, chlorate, acetate**) salts are always soluble.
3. Most **Cl^-**, **Br^-**, and **I^-** salts are soluble (except with Ag^+, Pb^{2+}, Hg_2^{2+}).
   halides
4. Most **sulfates** are soluble (except with Sr^{2+}, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+}; and Ag_2SO_4 is only "marginally soluble").
5. Most **hydroxides** are insoluble or only slightly soluble.
   Exceptions: see Rule 1, and
   $Ba(OH)_2$, $Sr(OH)_2$, $Ca(OH)_2$ are somewhat soluble
6. Salts of **"other"*** anions (CO_3^{2-}, CrO_4^{2-}, PO_4^{3-}, $C_2O_4^{2-}$ etc. we see in General Chemistry**) are insoluble (exceptions: see Rule 1)

* *anions not mentioned explicitly in the rules*

** ***not literally*** *all other anions known to humankind!*

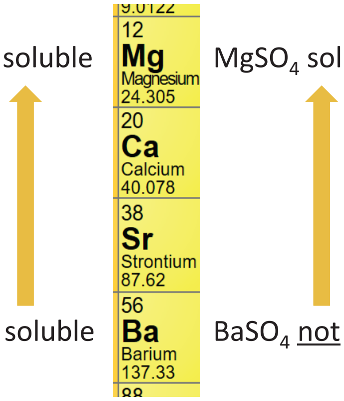
Additional observations about alkaline earth ionic compound solubilities

For Group 2 metals, hydroxide & sulfate solubility trends are opposite of each other

$Mg(OH)_2$ not soluble

decreasing solubility of hydroxides

$Ba(OH)_2$ soluble



$MgSO_4$ soluble

increasing solubility of sulfates

$BaSO_4$ not soluble

On sulfides:

- Sulfide ion, S^{2-}, is generally **in**soluble (except, as always, with alkali metals and ammonium)
- Sulfides of Group 2 metals (alkaline earth) **are** soluble.

But:
- Sulfide ion in water is unstable in the long term.
- It slowly reacts with water to produce H_2S gas.

Because its solutions are unstable, I don't include sulfide in the "normal" rules of solubility.

A silly way to remember the Solubility Rules

-- mnemonics are about the exceptions

-- exceptions are insoluble in rules 3 & 4, and soluble in later rules

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Some silly mnemonics for the Solubility Rules

1. Group **I** and **am**monia **"I am" always soluble**
2. **pe**rchlorate, **c**hlorate, **a**cetate, **n**itrate **"pecan" always soluble**
3. Halides (Cl^-, Br^-, I^-) generally **soluble**
   - except with Ag^+, Hg_2^{2+}, Pb^{2+}

     "He lost his halo at the Silver Hog Pub"
4. Sulfate (SO_4^{2-}) generally **soluble**
   - except with Sr^{2+}, Ba^{2+}, Ca^{2+}, Ag^+, Hg_2^{2+}, Pb^{2+}

     "So 4 his fun, Sr. (Señor) Baca goes to the Silver Hog Pub"
5. Hydroxide (OH^-) generally **insoluble**
   - except with Sr^{2+}, Ba^{2+}, Ca^{2+}

     "Oh, except for Sr. Baca!"
6. Compounds of other anions are **insoluble** (exception: Rule 1)

   Others are powders, except where I am

Which of the following ions forms a compound with Pb^{2+} that is soluble in water?

a) S^{2-}

b) Cl^-

c) NO_3^-

d) SO_4^{2-}

e) Na^+

Precipitation reactions can be represented in three different forms:

- Molecular equation
- Complete ionic equation
- Net ionic equation

"Molecular" Equation

A better term would have been "Formula Equation" since there are no molecules around.

- All ions or ionic compounds

- Reactants and products shown as compounds

  -- instead of the dissociated ions

- Use solubility rules to determine which compounds are aqueous and which compounds are solids.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Ag^+ and Cl^- are not soluble together

Complete Ionic Equation

All soluble substances that are ionized in aqueous solution are represented as ions.

Only the precipitate is written as a compound

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \longrightarrow \boxed{AgCl(s)} + Na^+(aq) + NO_3^-(aq)$$

Precipitate is not written as separate ions!

Spectator ion

Spectator ion

Spectator ions don't participate in the reaction.

They appear on both sides of the reaction in the same state.

Net Ionic Equation

- Includes only the ions that do something.

  i.e. the ions that precipitated

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

- No spectator ions

  Na^+ and NO_3^- are spectator ions. They remain in solution. So they don't show up in the net ionic equation

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Concept Practice

Write the correct "molecular" equation, complete ionic equation, and net ionic equation for the reaction between cobalt(II) chloride and sodium hydroxide.

Formula (Molecular) Equation:

$$CoCl_2(aq) + 2NaOH(aq) \rightarrow Co(OH)_2(s) + 2NaCl(aq)$$

Complete Ionic Equation:

$$Co^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + 2OH^-(aq) \rightarrow Co(OH)_2(s) + 2Na^+(aq) + 2Cl^-(aq)$$

Net Ionic Equation:

$$Co^{2+}(aq) + 2OH^-(aq) \rightarrow Co(OH)_2(s)$$

Solving Stoichiometry Problems for Reactions in Solution

1. Identify the species present in the combined solution, and determine what reaction occurs.
2. Write the balanced equation (net ionic or "molecular", depending on which seems more convenient) for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product(s), as required.
6. Convert to grams or other units, as required.

Not really different from general stoichiomery questions!

Practice

10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume volumes add up perfectly). Will there be a reaction? If so, what is the product?

- **Two solutions of ionic compounds**, typically we think of a double-displacement reaction of the "**precipitation**" kind.
- Will there be a precipitate? **If not**, the ions would still be floating around just as before the two solutions were mixed, and there would be **no reaction**.

sodium phosphate: Na_3PO_4 (aq) exists as $3\ Na^+ (aq) + \boxed{PO_4^{3-} (aq)}$

insoluble together

lead (II) nitrate: $Pb(NO_3)_2$ (aq) exists as $\boxed{Pb^{2+} (aq)} + 2\ NO_3^- (aq)$

"Molecular" (or "formula") equation:

$$2\ Na_3PO_4\ (aq) + 3\ Pb(NO_3)_2\ (aq) \rightarrow \underbrace{\mathbf{Pb_3(PO_4)_2\ (s)}}_{\text{precipitate}} + 6\ NaNO_3\ (aq)$$

Practice (continuing the same question)

10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume volumes add up perfectly). "What mass of precipitate will form?"

$$\textcircled{2}\ Na_3PO_4\ (aq) + \textcircled{3}\ Pb(NO_3)_2\ (aq) \rightarrow \underbrace{Pb_3(PO_4)_2\ (s)}_{\text{precipitate}} + 6\ NaNO_3\ (aq)$$

$\underset{M}{(0.30)}\underset{V}{(10.0x10^{-3})}$ $\underset{M}{(0.20)}\underset{V}{(20.0x10^{-3})}$

$3.0x10^{-3}$ mol **$4.0x10^{-3}$ mol**

Calculate number of moles from:
n = M · V

$\div \textcircled{2}$ $\div \textcircled{3}$

$1.5x10^{-3} > 1.3x10^{-3}$

Smaller; limiting

To find the limiting reactant:
Divide reactant moles by their coefficients in the equation

$$4.0x10^{-3}\ mol\ Pb(NO_3)_2 \times \frac{1\ mol\ Pb_3(PO_4)_2}{\underset{\text{3 mol Pb(NO}_3)_2\text{ gives 1 mol Pb}_3\text{(PO}_4)_2}{3\ mol\ Pb(NO_3)_2}} \times \frac{811.54\ g\ Pb_3(PO_4)_2}{1\ mol\ Pb_3(PO_4)_2} = \mathbf{1.1\ g\ Pb_3(PO_4)_2}$$

Practice (continuing the same question)

Initial amounts:

What is the concentration of **phosphate ions remaining** in solution, after precipitation?"

For a reacting ion, it's better to use the net ionic equation, with moles of ions

$$\mathbf{2\ PO_4^{3-}\ (aq) + 3\ Pb^{2+}\ (aq) \rightarrow Pb_3(PO_4)_2\ (s)}$$

We calculated earlier the moles of reactant compounds

$n_{Na_3PO_4} = 3.0x10^{-3}$ mol before reaction

$n_{Pb(NO_3)_2} = 4.0x10^{-3}$ mol before reaction

$$n_{PO_4^{3-}} = \frac{1\ mol\ PO_4^{3-}}{1\ mol\ Na_3PO_4}\ 3.0x10^{-3}\ mol\ Na_3PO_4 = 3.0x10^{-3}\ mol\ PO_4^{3-}$$

$$n_{Pb^{2+}} = \frac{1\ mol\ Pb^{2+}}{1\ mol\ Pb(NO_3)_2}\ 4.0x10^{-3}\ mol\ Pb(NO_3)_2 = 4.0x10^{-3}\ mol\ Pb^{2+}$$

Remaining amount = (Initial amount) – (Consumed amount)

$$2\ PO_4^{3-}\,(aq) + 3\ Pb^{2+}\,(aq) \longrightarrow Pb_3(PO_4)_2\,(s)$$

limiting

Because $Pb(NO_3)_2$ is limiting

	PO_4^{3-}	Pb^{2+}
Initial:	$3.\underline{0}0\text{x}10^{-3}$ mol	$4.0\text{x}10^{-3}$ mol
Consumed:	$2.\underline{6}7\text{x}10^{-3}$ mol	$4.0\text{x}10^{-3}$ mol
Remaining:	$\underline{3}.3\text{x}10^{-4}$ mol	0 mol

$$\text{mol } PO_4^{3-} \text{ consumed} = 4.0\text{x}10^{-3} \text{ mol } Pb^{2+} \text{x} \frac{2 \text{ mol } PO_4^{3-}}{3 \text{ mol } Pb^{2+}} = 2.\underline{6}7 \text{ x}10^{-3} \text{ mol } PO_4^{3-} \text{ consumed}$$

[3 mol Pb^{2+} consumes 2 mol PO_4^{3-}]

$$\text{mol } PO_4^{3-} \text{ remaining} = 3.\underline{0}0\text{x}10^{-3} \text{ mol} - 2.\underline{6}7 \text{ x}10^{-3} \text{ mol} = \underline{3}.3 \text{ x}10^{-4} \text{ mol } PO_4^{3-} \text{ remaining}$$

V_{total} = 10.0mL + 20.0mL = $30.0 \text{ x } 10^{-3}$ L

$$M_{PO_4^{3-}} = \frac{\underline{3}.3 \text{ x}10^{-4} \text{ mol}}{30.0 \text{ x } 10^{-3} \text{ L}} = 0.01\ M$$

Practice (continuing the same question)

What is the **concentration of nitrate ions remaining** in solution after the reaction is complete?

- NO_3^- is a spectator ion.
- It remains in solution.
- So, as far as NO_3^- is concerned, it's a **dilution** problem.
- The other solution didn't have any nitrate, so the nitrate got diluted when the two solutions combined.

We find the moles of NO_3^- was brought in by 4.0×10^{-3} mol $Pb(NO_3)_2$

$$n_{NO_3^-} = 4.0 \times 10^{-3} \text{ mol } Pb(NO_3)_2 \ \frac{2 \text{ mol } NO_3^-}{1 \text{ mol } Pb(NO_3)_2} = 8.0\text{x}10^{-3} \text{ mol } NO_3^-$$

before and after reaction

$$M_{NO_3^-} = \frac{8.0 \text{ x}10^{-3} \text{ mol}}{30.0 \text{ x } 10^{-3} \text{ L}} = 0.27\ M$$

Practice (continuing the same question)

We could also use

$$M_1V_1 = M_2V_2$$

if we start with the initial concentration of NO_3^- instead of the moles of NO_3^-

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Acids, Bases, and Their Reactions

Acid-base "neutralization" reactions are also "double-displacement"

We touched on acids and bases earlier

- When we discussed inorganic nomenclature
- And earlier in this chapter when we considered electrolytes

We will now look at them at more depth

Acids & bases as defined by Arrhenius

- Oldest acid-base concept
- Applies in aqueous solutions
- **Acids provide H^+ in aqueous solution**
- **Bases provide OH^- in aqueous solution**
- When an acid and a base "neutralize" each other
  - H^+ and OH^- combine to form H_2O
  - Remaining parts of the acid and the base form a "salt"

Acids as defined by Arrhenius ("Arrhenius Acids")

Acid: provides H^+

$$H\{\text{Anion}\}\,(aq) \longrightarrow H^+\,(aq) + \{\text{Anion}\}\,(aq)$$

$$HNO_3\,(aq) \longrightarrow H^+\,(aq) + NO_3^-\,(aq)$$

"Acidic" hydrogens are attached to an anion

H^+(aq) produced actually attaches to a water molecule, forming H_3O^+

H_3O^+ (aq) is called "hydronium"

- In the formula of an Arrhenius acid, the acidic hydrogen is written first.
- Any other hydrogens later in the formula are not acidic.

  $HC_2H_3O_2$

  acidic   not acidic
- An acid with one acidic hydrogen is "monoprotic"

  HNO_3
- An acid with two acidic hydrogens is "diprotic"

  H_2SO_4
- An acid with three acidic hydrogens is "triprotic"

  H_3PO_4

Weak acids:

Only some of the weak acid molecules dissociate into ions

$$HC_2H_3O_2\ (aq) \longrightarrow H^+\ (aq) + C_2H_3O_2^-\ (aq)$$

Most acetic acid molecules remain undissociated

Only some of the molecules dissociate into ions

The format of a reaction equation gives the impression that all reactions are always complete. That's not true

Strong acids:

Practically all of the strong acid molecules dissociate into H^+ and the anion of the acid.

$$HBr\ (aq) \longrightarrow H^+\ (aq) + Br^-\ (aq)$$

All of the acid molecules dissociate into ions

List of "Strong Acids" we need to know:

$$HI(aq) \longrightarrow H^+(aq) + I^-(aq)$$

$$HBr(aq) \longrightarrow H^+(aq) + Br^-(aq)$$

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Binary hydrogen compounds of halogens, except HF*(aq)*

$$HClO_4(aq) \longrightarrow H^+(aq) + ClO_4^-(aq)$$

$$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$$

Even for a strong acid like H_2SO_4, if there is more than one acidic H, only the 1st dissociation is complete. The 2nd H doesn't dissociate easily.

"Weak Acids" we need to know:

- Any acid not listed among the "strong"

  -- Some may be fairly strong, but just don't dissociate fully
- Organic acids such as acetic acid ($HC_2H_3O_2$)
- HF*(aq)*
- H_3PO_4
- The anion produced after diprotic or triprotic acids, even if they are strong in releasing their first acidic H

  $HSO_4^-\ (aq) \rightarrow H^+(aq) + SO_4^{2-}\ (aq)$

  Incomplete dissociation

Extracurricular knowledge, if you enjoy obscure info

The strange case of Hydrogen Fluoride (HF)

- HF is normally regarded as a weak acid.
- In dilute aqueous solution, after HF dissociates fully, F^- forms a tight ion pair with the H_3O^+ just created
  - ➢ makes H_3O^+ unavailable to function as free $H^+(aq)$
- Actually when there isn't much water around, HF is a stronger acid than our usual "strong" acids like H_2SO_4, HCl, etc.

Nevertheless:

In the world of General Chemistry where aqueous (and not very concentrated) solutions rule, **we will regard HF as a "weak acid"**

Bases as defined by Arrhenius ("Arrhenius bases")

Base: provides OH^-

$$\{Cation\}\mathbf{OH}\ (aq) \longrightarrow \{Cation\}^+\ (aq) + \mathbf{OH^-}\ (aq)$$

$$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$$

- Bases in the Arrhenius definition act by producing OH^- ions
- Most inorganic bases are metal hydroxides, such as NaOH, KOH, $Ca(OH)_2$.
- In General Chemistry it's common to assume all metal hydroxides to be strong bases.
- Ammonia (NH_3) is the only weak base we will need to know in this course.

NH_3 , a molecular compound and a weak base, creates the hydroxide ion by reacting with water:

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Since it is a weak base, the reaction is not complete, and most of the ammonia molecules dissolved in water remain as NH_3 molecules.

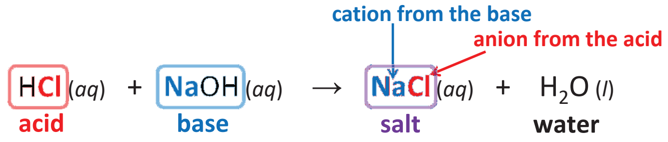
When you see a reagent bottle labeled "NH_4OH" or "ammonium hydroxide", it is mostly dissolved ammonia rather than the "compound" NH_4OH.

Arrhenius acid-base reactions (Neutralization)

When we think about acid-base neutralization reactions, we often think in Arrhenius terms

acid base salt water

$$\boxed{H\{Anion\}}(aq) + \boxed{\{Cation\}OH}(aq) \longrightarrow \boxed{\{Cation\}\{Anion\}}(aq) + H_2O\ (l)$$



cation from the base

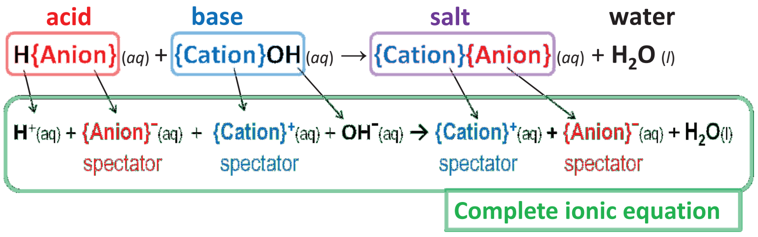
anion from the acid

$$\underset{\text{acid}}{\boxed{HCl}}(aq) + \underset{\text{base}}{\boxed{NaOH}}(aq) \longrightarrow \underset{\text{salt}}{\boxed{NaCl}}(aq) + \underset{\text{water}}{H_2O\ (l)}$$

- Ions are shown only as part of compounds above.
- So it is the "molecular equation" representation of an acid-base reaction.

Complete ionic equation for Arrhenius acid-base reactions

Acid, base, and the produced salt written as ions



acid base salt water

$$\boxed{H\{Anion\}}(aq) + \boxed{\{Cation\}OH}(aq) \longrightarrow \boxed{\{Cation\}\{Anion\}}(aq) + H_2O\ (l)$$

$$H^+(aq) + \underset{\text{spectator}}{\{Anion\}^-(aq)} + \underset{\text{spectator}}{\{Cation\}^+(aq)} + OH^-(aq) \rightarrow \underset{\text{spectator}}{\{Cation\}^+(aq)} + \underset{\text{spectator}}{\{Anion\}^-(aq)} + H_2O(l)$$

Complete ionic equation

Example:

acid HBr (aq) base NaOH (aq) salt NaBr (aq)

$$\boxed{H^+(aq) + \underset{\text{spectator}}{Br^-(aq)}} + \boxed{\underset{\text{spectator}}{Na^+(aq)} + OH^-(aq)} \longrightarrow \boxed{\underset{\text{spectator}}{Na^+(aq)} + \underset{\text{spectator}}{Br^-(aq)}} + H_2O\ (l)$$

Complete ionic equation for an Arrhenius acid-base reaction is written the same way even if we have a weak acid and/or weak base.

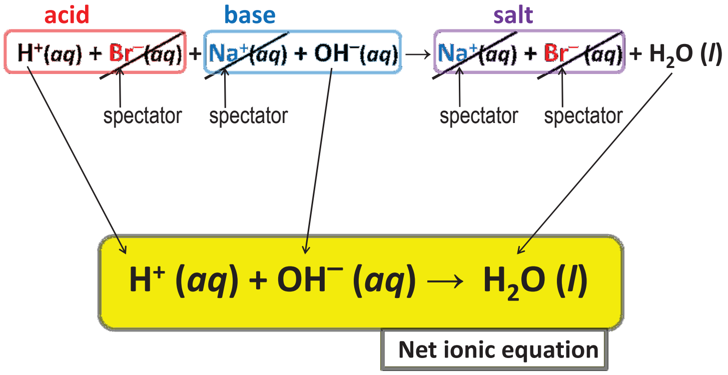
- Partial dissociation of a weak acid or base into ions only applies when the acid or base is alone in solution
- In an acid/base reaction, even weak acids and bases are consumed completely (if they are limiting)

acid $HC_2H_3O_2$ (aq) base NaOH (aq) salt $NaC_2H_3O_2$ (aq)

$$\boxed{H^+(aq) + \underset{\text{spectator}}{C_2H_3O_2^-(aq)}} + \boxed{\underset{\text{spectator}}{Na^+(aq)} + OH^-(aq)} \longrightarrow \boxed{\underset{\text{spectator}}{Na^+(aq)} + \underset{\text{spectator}}{C_2H_3O_2^-(aq)}} + H_2O(l)$$

Net ionic equation for an Arrhenius acid-base reaction

Obtained when we eliminate spectator ions from the complete ionic equation



acid base salt

$$\boxed{H^+(aq) + \underset{\text{spectator}}{\cancel{Br^-(aq)}}} + \boxed{\underset{\text{spectator}}{\cancel{Na^+(aq)}} + OH^-(aq)} \longrightarrow \boxed{\underset{\text{spectator}}{\cancel{Na^+(aq)}} + \underset{\text{spectator}}{\cancel{Br^-(aq)}}} + H_2O\ (l)$$

$$\boxed{H^+\ (aq) + OH^-\ (aq) \longrightarrow H_2O\ (l)}$$

Net ionic equation

Performing Calculations for Acid–Base Reactions

- No different from other reactions, really.
- Often we do “titrations”
- In a titration there is no limiting reactant, by definition
- Balancing is easy because **H^+ and OH^- must come in equal moles**
  - **we can even do the stoichiometry without writing the reaction**
  - **but don’t feel you must be able to**

Acid–Base Titrations

Titration: delivery of a measured volume of a solution of known concentration (the titrant) into a solution containing the substance being analyzed (the analyte).

Equivalence point: exactly enough titrant added to react completely with the analyte.

- The ideal we try to achieve at the end of titration

Endpoint: the indicator changes color so you can tell the equivalence point has been reached. It’s an experimental approximation of the equivalence point.

- What we can achieve in practice, with our indicator

There is no limiting or excess reactant in titrations -- by design

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Practice

For the titration of sulfuric acid (H_2SO_4) with sodium hydroxide (NaOH), how many mL of 0.1054 *M* sodium hydroxide would be required to react with 1.00 L of 0.500 *M* sulfuric acid to reach the endpoint?

$$H_2SO_4\,(aq) + 2\,NaOH\,(aq) \rightarrow Na_2SO_4\,(aq) + 2\,H_2O\,(aq)$$

$$n_{H_2SO_4} = M_{H_2SO_4} V_{H_2SO_4} = (0.500)(1.00) = 0.500 \text{ mol } H_2SO_4$$

$$n_{NaOH} = 0.500 \text{ mol } H_2SO_4 \frac{2 \text{ mol NaOH}}{1 \text{ mol } H_2SO_4} = 1.00 \text{ mol NaOH}$$

$$n_{NaOH} = M_{NaOH} V_{NaOH}$$

$$1.00 \text{ mol} = (0.1054 \text{ mol/L})\, V_{NaOH}$$

$$\Rightarrow \quad V_{NaOH} = 9.49 \text{ L} = 9.49\text{x}10^3 \text{ mL}$$

When the acid and base are used in stoichiometric amounts, such as during titrations, you can use the following formula:

$$\underbrace{\alpha\, M_{acid} V_{acid}}_{\textbf{Moles of } H^+ \textbf{ from acid}} = \underbrace{\beta\, M_{base} V_{base}}_{\textbf{Moles of } OH^- \textbf{ from base}}$$

Where α = number of acidic H’s in the acid formula
β = number of OH’s in the base formula

α = 1 for HCl
α = 2 for H_2SO_4
α = 3 for H_3PO_4

β = 1 for KOH
β = 2 for $Ca(OH)_2$
β = 3 for $Al(OH)_3$

Don’t confuse it with the dilution formula!

Then the previous example:

For the titration of sulfuric acid (H_2SO_4) with sodium hydroxide (NaOH), how many mL of 0.1054 *M* sodium hydroxide would be required to react with 1.00 L of 0.500 *M* sulfuric acid to reach the endpoint?

could be solved this way:

$$\alpha\, M_{acid} V_{acid} = \beta\, M_{base} V_{base}$$

$$(2)\,(0.500)\,(1.00) = (1)\,(0.1054)\,V_{NaOH}$$

$$V_{NaOH} = 9.49 \text{ L} = 9.49 \text{ x } 10^3 \text{mL}$$

A common ingredient in “antacids” that neutralize the “stomach acid” (HCl(aq)) is $CaCO_3$, a “salt”. How come?

$CaCO_3$ is a “salt”, and not a “base” in the Arrhenius sense. How does it neutralize an acid?

Strong acids can “push” weak acids out of their salts, to form their own salt instead.

salt of weak acid | strong acid | salt of strong acid | weak acid

$$CaCO_3(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2CO_3\,(aq)$$

$$H_2CO_3\,(aq) \longrightarrow H_2O(l) + CO_2\,(g)$$

The weak carbonic acid happens to be unstable

Another example:

salt of weak acid | strong acid | salt of strong acid | weak acid

$$NaC_2H_3O_2(s) + HNO_3(aq) \longrightarrow NaNO_3\,(aq) + HC_2H_3O_2\,(aq)$$

The solution starts smelling like vinegar (acetic acid; weak)

"Strong acids pushing weak acids out of their salts" can be understood as an acid/base reaction under a more general definition: "Bronsted-Lowry" definition of acids and bases

The anion (e.g. CO_3^{2-}) of a weak acid (e.g. H_2CO_3) "accepting" an H^+ from a stronger acid (e.g. HCl) is acting as a "Bronsted-Lowry" base.

$$CO_3^{2-}(aq) + 2HCl(aq) \rightarrow 2Cl^-(aq) + H_2CO_3(aq)$$

We will study the more general "Bronsted-Lowry" definition of acids and bases in a later chapter.

Redox Reactions

Oxidation-**Red**uction reactions

- One or more electrons are transferred.
- Electrons are not created or destroyed in chemical reactions:

number of **electrons gained** = number of **electrons lost**

Reduction
(yes, it's confusing, but there is a reason for the name)

Oxidation

Remember by:

Oxidation **I**s **L**oss of e^-

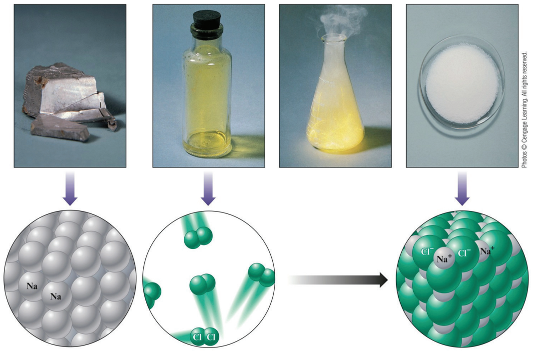
Reduction **I**s **G**ain of e^-

OIL RIG

Originally "oxidation" referred literally to combining with oxygen

- The more oxygens combined with an atom, the more "oxidized" it was
- Oxygen was thought of as "oxide" in its compounds (with its –2 charge)
- So, the more oxygens bonded to an atom, the more positive it was
  - To have a neutral compound
- Oxidation was then generalized to mean an increase in "charge" (and not necessarily bonding to oxygen)
- And charge itself was generalized to the more abstract "oxidation state" (which is often not an actual charge)

A simple redox reaction: Reaction of Sodium and Chlorine



$$2\,Na\,(s) + Cl_2\,(g) \rightarrow 2\,NaCl\,(s)$$

Charge: 0 0 +1 –1

We keep track of electron gain and loss using Oxidation numbers (Oxidation states)

- They are like charges, but not exactly
- Oxidation numbers are the charges that would exist if the atoms always formed ionic bonds, with the electron ending up on the atom that likes them more.
  e.g. oxygen likes electrons more than hydrogen
- Oxidation numbers are often not actual charges
  -- when the bonds the atom makes is not ionic
  -- But charges on ions are real
- Even when they are not actual charges, the net change in oxidation states does determine the net number of electrons transferred

---

Hierarchy of Rules for Assigning Oxidation States

(1) **The oxidation state of an atom in a free element is 0.**

(2) The oxidation state of a monoatomic ion is equal to its charge.

(3) The sum of the oxidation states of all atoms in
  (3a) a neutral molecule or formula unit is 0
  (3b) a polyatomic ion is equal to the charge of the ion

Sum of oxidation states in an entity = net charge

(4) In their compounds,
  (4a) **Group I** metals have an oxidation state of **+1**
  (4b) **Group II** metals have an oxidation state of **+2**

(5) In their compounds, nonmetals are assigned oxidation states according to a hierarchical table. **Elements higher in the table have priority over lower entries.**

"Oxygen is –2" can be overridden by a "higher" rule like "Fluorine is –1"

Fluorine	–1
Hydrogen	+1
Oxygen	–2
Group 7A	–1
Group 6A	–2
Group 5A	–3

Examples on monatomic species

Cu(s)	Oxidation state = 0
H_2(g)	Oxidation state = 0
H_2(l)	Oxidation state = 0
H (g)	Oxidation state = 0
H^+(aq)	Oxidation state = +1
S_8(s)	Oxidation state = 0
S^{2-} (s)	Oxidation state = –2
Fe^{3+}(aq)	Oxidation state = +3

Finding the oxidation states of elements in a compound or polyatomic ion

1. If you recognize a polyatomic ion, treat it separately first
2. Find the lowest priority element in compound or ion
   - Deal with it last
   - Its oxidation state will be determined by higher-priority rules
3. Assign other elements their preferred oxidations states
4. Determine the oxidation state of the last (lowest priority) element using the oxidation states of other elements and Rule 3
   - It might not agree with its "default" oxidation state (e.g. oxygen might end up with –1 or +2 instead of the "default" –2). That's completely ok.

Find the oxidation states for each of the elements in:

Lowest priority (doesn't even have a preferred oxidation state)

Oxygen is –2 (from Rule 5); and not overridden by a higher rule

$$\overset{+4}{C}\,\overset{-2}{O}_3{}^{2-}$$

No specific rule; determined from oxygen's, using the rule "sum of oxidation numbers equals the charge of the ion"

$(1)x + 3(-2) = -2 \Rightarrow x = +4$

Find the oxidation states for each of the elements in $K_2Cr_2O_7$

Suppose we didn't recognize the Cr_2O_7 part as the $Cr_2O_7{}^{2-}$ ion and we looked at the entire formula

The oxidation state of a monatomic ion is equal to its charge (Rule 2); Group 1A metals are always +1 in compounds

Lowest priority (doesn't even have a preferred oxidation state)

Oxygen is –2 (from Rule 5); and not overridden by a higher rule

$$\overset{+1}{K}_2\overset{+6}{Cr}_2\overset{-2}{O}_7$$

No specific rule; determined from others, using the rule "sum of oxidation numbers in a compound equals zero"

$(2)(+1) + (2)x + 7(-2) = 0 \Rightarrow x = +6$

We got lucky because K has a high-priority oxidation state

Find the oxidation states for each of the elements in $K_2Cr_2O_7$

Suppose we recognized the Cr_2O_7 part as the $Cr_2O_7{}^{2-}$ ion

Oxygen is –2 (from Rule 5); and not overridden by a higher rule

Lowest priority (doesn't even have a preferred oxidation state)

determined from oxygen and "sum of oxidation numbers in an ion equals ion charge"

$(2)x + 7(-2) = -2 \Rightarrow x = +6$

$$\overset{+6}{Cr}_2\overset{-2}{O}_7{}^{2-}$$

Then we would have determined the oxidation state of the metal as we always did. Its charge needs to neutralize the anion(s) in the formula. But we already know K is always +1.

$$\overset{+1}{K}_2\overset{+6}{Cr}_2\overset{-2}{O}_7$$

Find the oxidation states for each of the elements in $CuSO_4$

Here we must recognize SO_4 as $SO_4{}^{2-}$. Otherwise we either can't determine both Cu and S, or worse, we assign -2 to S as "higher priority" than Cu (since there is a rule for S, but there isn't one for Cu), and end up assigning +10 to Cu!

Oxygen = –2 as highest priority element here

"sum of oxidation numbers in an ion equals ion charge"

$(2)x + 7(-2) = -2 \Rightarrow x = +6$

$$\overset{+6}{S}\overset{-2}{O}_4{}^{2-}$$

Because it needs to neutralize the 2- charge of the sulfate.

We could do this before learning about oxidation numbers!

$$\overset{+2}{Cu}\,\overset{+6}{S}\overset{-2}{O}_4$$

Find the oxidation states for each of the elements in:

Fluorine is –1 (from Rule 5); and not overridden by a higher rule

$\overset{+2}{O}\overset{-1}{F_2}$

Rule 5 value of -2 overruled by higher priority Fluorine; oxidation state of O is then determined from "sum of oxidation numbers in a compound equals zero"

$(1)x + 2(-1) = 0 \Rightarrow x = +2$

Find the oxidation states for each of the elements in:

Group 1 metals = +1 in their compounds

$\overset{+1}{Li}\overset{-1}{H}$

Rule 5 value of +1 overruled by higher priority Rule 2; determined from "sum of oxidation numbers in a compound equals zero"

$(1)(+1) + (1)x = 0 \Rightarrow x = -1$

Re-stating "oxidation" and "reduction" in terms of oxidation state (instead of electrons)

Oxidation: **increase in oxidation state** (loss of electrons)

Reduction: **decrease in oxidation state** (gain of electrons)

Oxidation and reduction are coupled

- Electrons lost by one atom must be gained by another.
- Electrons are not lost; they are transferred

Number of electrons lost = Number of electrons gained

- If there is an atom getting oxidized, there must be at least one other atom getting reduced
- If there is an atom getting reduced, there must be at least one other atom getting oxidized

---

Oxidation and reduction are coupled

Oxidized substance: reducing agent

- If something is **oxidized**, it must be **reducing** something else

Reduced substance: oxidizing agent

- If something is **reduced**, it must be **oxidizing** something else

And remember:

- "oxidizing agent" and "reducing agent" are always **on the reactant side**!
- They **cannot be products**!

In a redox reaction:

- What we said about electrons applies equally to oxidation states.
- An increase in the oxidation state of an atom is coupled with a decrease in the oxidation state of another atom

It's a redox reaction if …

The oxidation state of any element changes in the reaction

Remember: If there is an element that changed its oxidation state, there must be another that also changed in the opposite direction

---

It's <u>probably</u> a redox reaction if there is an elemental substance such as O_2(g), Cu(s), etc., anywhere in the reaction

Oxidation state is zero for an elemental substance

In a <u>compound</u> atoms normally have nonzero oxidation states.

If the elemental substance like O_2(g), H_2(g), S_8(s), Fe(s) appears on either side of a reaction, it must occur in a compound on the other side of the reaction.

$$\overset{0}{C}(s) + \overset{0}{O_2}(g) \rightarrow \overset{+4}{C}\overset{-2}{O_2}(g)$$

$$\overset{0}{Fe}(s) + \overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O_4}(aq) \rightarrow \overset{+2}{Fe}\overset{+6}{S}\overset{-2}{O_4}(aq) + \overset{0}{Cu}(s)$$

There are exceptions to:

"It's <u>probably</u> a redox reaction if there is an elemental substance such as O_2(g), Cu(s), etc., anywhere in the reaction"

<u>Exception</u>:
When the element is in an "elemental state" on both sides of the reaction. For example, there is no redox here:

$O_2(g) + O(g) \rightarrow O_3(g)$

<u>Exception</u>:
When the element happens to have an average of zero oxidation state in a compound. For example, there is no redox here:

$C_6H_{12}O_6\,(s) \rightarrow C(s) + H_2O\,(g)$

Example

Which of the following are oxidation-reduction reactions?
Identify the oxidizing agent and the reducing agent.
Remember: they are always <u>reactants, never products</u>

oxidized

reduced

a) $\overset{0}{Zn}(s) + 2\,\overset{+1}{H}\overset{-1}{Cl}(aq) \rightarrow ZnCl\overset{-1}{_2}(aq) + \overset{0}{H_2}(g)$

reducing agent

oxidizing agent

Redox reaction

2x+(7)(-2) = -2

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

b) $\overset{+6}{Cr_2}\overset{-2}{O_7^{2-}}(aq) + 2\,\overset{-2}{O}\overset{+1}{H^-}(aq) \rightarrow 2\,\overset{+6}{Cr}\overset{-2}{O_4^{2-}}(aq) + \overset{+1}{H_2}\overset{-2}{O}(l)$

Not a Redox reaction

Single-displacement reactions

Type of <u>redox</u> reactions where:

- **the more "active" element (metal or hydrogen) replaces the less active one.**

The **more active element gets oxidized** into a <u>cation</u>

- Reduces the cation of the less active element

The **less active element gets reduced** <u>out of its cation form in to its elemental form</u>

In a single-displacement reaction:

The **more active element gets oxidized** into a cation

The **less active element gets reduced** out of its cation form into its elemental form

$A(s) + B^+(aq) \rightarrow A^+(aq) + B(s)$

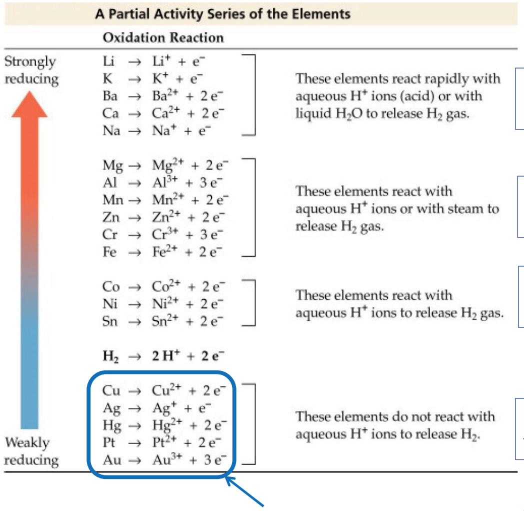
More active than B

Less active than A

$A(s) + H^+(aq) \rightarrow A^+(aq) + H_2(g)$

More active than H_2(g)

Less active than A



A Partial Activity Series of the Elements

	Oxidation Reaction	
Strongly reducing	$Li \rightarrow Li^+ + e^-$ $K \rightarrow K^+ + e^-$ $Ba \rightarrow Ba^{2+} + 2e^-$ $Ca \rightarrow Ca^{2+} + 2e^-$ $Na \rightarrow Na^+ + e^-$	These elements react rapidly with aqueous H^+ ions (acid) or with liquid H_2O to release H_2 gas.
	$Mg \rightarrow Mg^{2+} + 2e^-$ $Al \rightarrow Al^{3+} + 3e^-$ $Mn \rightarrow Mn^{2+} + 2e^-$ $Zn \rightarrow Zn^{2+} + 2e^-$ $Cr \rightarrow Cr^{3+} + 3e^-$ $Fe \rightarrow Fe^{2+} + 2e^-$	These elements react with aqueous H^+ ions or with steam to release H_2 gas.
	$Co \rightarrow Co^{2+} + 2e^-$ $Ni \rightarrow Ni^{2+} + 2e^-$ $Sn \rightarrow Sn^{2+} + 2e^-$	These elements react with aqueous H^+ ions to release H_2 gas.
	$H_2 \rightarrow 2H^+ + 2e^-$	
Weakly reducing	$Cu \rightarrow Cu^{2+} + 2e^-$ $Ag \rightarrow Ag^+ + e^-$ $Hg \rightarrow Hg^{2+} + 2e^-$ $Pt \rightarrow Pt^{2+} + 2e^-$ $Au \rightarrow Au^{3+} + 3e^-$	These elements do not react with aqueous H^+ ions to release H_2.

Doesn't need an acid to produce $H_2(g)$

Needs acidic solution or steam to produce $H_2(g)$

Needs acidic solution to produce $H_2(g)$

Cannot reduce H in any form to produce $H_2(g)$

Need to remember only the "noble metals" that don't react with acids

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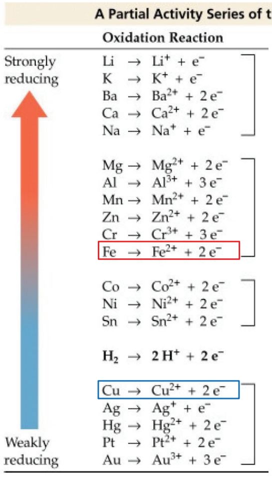

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	Oxidation Reaction
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	$Mg \rightarrow Mg^{2+} + 2e^-$ $Al \rightarrow Al^{3+} + 3e^-$ $Mn \rightarrow Mn^{2+} + 2e^-$ $Zn \rightarrow Zn^{2+} + 2e^-$ $Cr \rightarrow Cr^{3+} + 3e^-$ $Fe \rightarrow Fe^{2+} + 2e^-$
	$Co \rightarrow Co^{2+} + 2e^-$ $Ni \rightarrow Ni^{2+} + 2e^-$ $Sn \rightarrow Sn^{2+} + 2e^-$
	$H_2 \rightarrow 2H^+ + 2e^-$
Weakly reducing	$Cu \rightarrow Cu^{2+} + 2e^-$ $Ag \rightarrow Ag^+ + e^-$ $Hg \rightarrow Hg^{2+} + 2e^-$ $Pt \rightarrow Pt^{2+} + 2e^-$ $Au \rightarrow Au^{3+} + 3e^-$

- Metals tend to lose electrons
- The stronger the tendency, the more reactive the metal
- The more active metals are more readily oxidized (stronger reducing agents)
- More active metal reduces the less active metal

More reactive (strongly reducing) element will replace the less reactive element in a compound:

$$\underset{\text{more active}}{Fe(s)} + CuSO_4(aq) \longrightarrow \underset{\text{less active}}{Cu(s)} + FeSO_4(aq)$$

Predicting single-replacement reactions from the activity series

You don't need to "memorize" the activity series, other than the listed "noble metals" (less active than $H_2(g)$)

If you know "noble metals" Cu, Ag, Hg, Pt, Au, you can predict:

- **Noble metals don't replace $H^+_{(aq)}$** to produce $H_{2(g)}$
  - No reaction with acids
- **Non-noble metals replace $H^+_{(aq)}$** to produce $H_2(g)$
- **Non-noble metals replace noble metals** from compounds by reducing them to elemental form

Consult the activity series table to predict the "competition" between noble metals or between non-noble metals

For example, Cu(s) **can** reduce Ag^+ ion to Ag(s)
Ag(s) **cannot** reduce Cu^+ ion to Cu(s)
Zn(s) **can** reduce Fe^{2+} to Fe(s)
Fe(s) **cannot** reduce Zn^{2+} to Zn(s)

- Activity series includes Hydrogen, $H_2(g)$, because its cation is what acids produce.
- If a metal is more active than hydrogen, it "displaces" $H^+(aq)$ by reducing it to hydrogen molecules $H_2(g)$.
- Metals that are higher than hydrogen in the activity series react with (and are corroded by) acids.
  - While we "feel" that the acid is attacking the metal, actually it is the metal attacking the acid and displacing hydrogen
- "Noble metals" such as Cu, Hg, Ag, Au, Pt are less active than hydrogen, do not displace $H^+(aq)$, and are resistant to corrosion by acids.

Practice

$Mg(s) + ZnCl_2(aq) \longrightarrow Zn(s) + MgCl_2(aq)$

$Ag(s) + SnCl_2(aq) \longrightarrow$ No reaction

$Au(s) + AgCl\ (aq) \longrightarrow$ No reaction

$Zn(s) + NiCl_2\ (aq) \longrightarrow Ni(s) + ZnCl_2(aq)$

$2Al(s) + 3CuCl_2\ (aq) \longrightarrow 3Cu(s) + AlCl_3(aq)$

$Zn(s) + HCl(aq) \longrightarrow H_2(g) + ZnCl_2(aq)$

$Cu(s) + HBr(aq) \longrightarrow$ No reaction

$Fe(s) + ZnCl_2(aq) \longrightarrow$ No reaction

Decomposition Reactions

- A compound breaks down into other compounds or elements
- It may involve oxidation-reduction, but not necessarily
- **One reactant, more than one product**

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$
$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$
$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$$
$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$
$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Synthesis (Combination) Reactions

- Opposite of Decomposition
- More than one substance (compounds or elements) combine to give <u>one product</u>
- It may involve oxidation-reduction, but not necessarily
- **More than one reactant, one product**

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

$$2H_2O(l) + O_2(g) \longrightarrow 2H_2O_2(aq)$$

$$H_2O(l) + CO_2(g) \longrightarrow H_2CO_3(aq)$$

$$CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$$

$$2KCl(s) + 3O_2(g) \longrightarrow 2KClO_3(s)$$