Chapter 7: Stoichiometry of Chemical Reactions

Chemical Equations: Interpreting and Writing

A **chemical equation** uses chemical symbols to denote what occurs in a chemical reaction.

$NH_3 + HCl \longrightarrow NH_4Cl$

Ammonia and hydrogen chloride react to produce ammonium chloride. Each chemical species that appears to the left of the arrow is called a **reactant**.

$\mathbf{NH_3} + \mathbf{HCl} \longrightarrow NH_4Cl$

Each species that appears to the right of the arrow is called a **product**.

$NH_3 + HCl \longrightarrow \mathbf{NH_4Cl}$

Labels are used to indicate the physical states in a chemical equation:

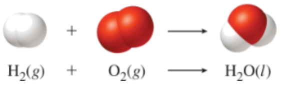
(*g*) gas, (*l*) liquid, (*s*) solid, (*aq*) ***aqueous*** [dissolved in water]

$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

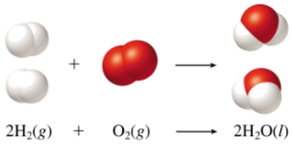
$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

Balancing Chemical Equations

Chemical equations must be **balanced** so that the law of conservation of mass is obeyed.



Balancing is achieved by writing **stoichiometric coefficients** to the left of the chemical formulas.



How to "Read" Chemical Equations:

$2Mg + O_2 \longrightarrow 2MgO$

2 atoms Mg + 1 molecule O_2 makes 2 formula units MgO

2 moles Mg + 1 mole O_2 makes 2 moles MgO

48.6 grams Mg + 32.0 grams O_2 makes 80.6 g MgO

This is NOT saying:
2 grams Mg + 1 gram O_2 makes 2 g MgO

Balancing Chemical Equations

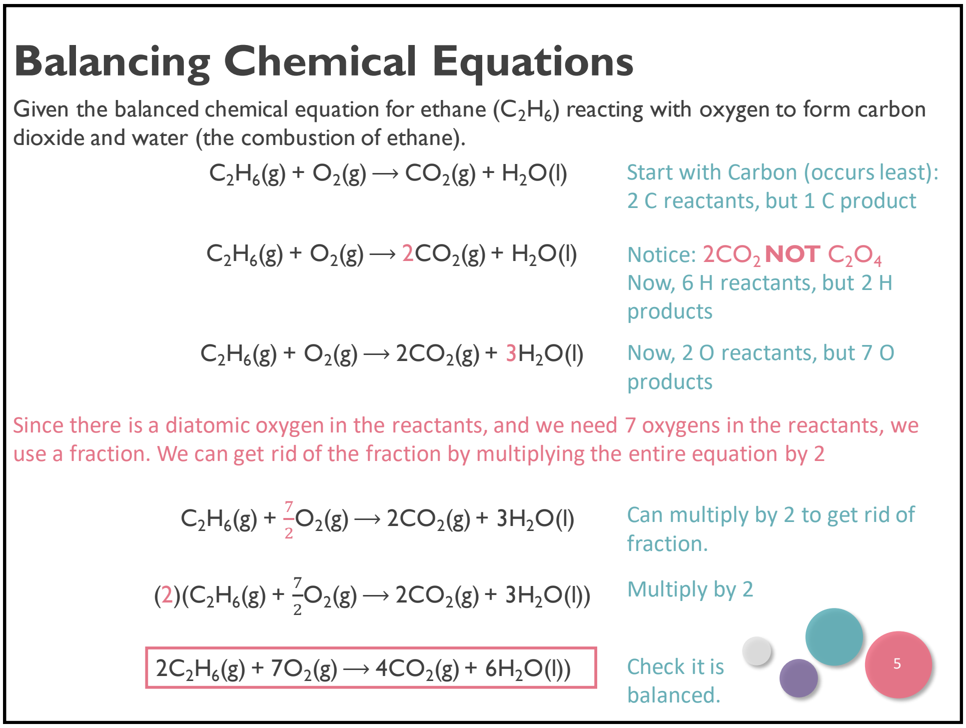
1. Write the **correct** chemical formula(s) for the reactants on the left side and the **correct** chemical formula(s) for the product(s) on the right side of the equation.

2. Change the numbers in front of the formulas (**coefficients**) to make the number of atoms of each element the same on both sides of the equation. Do not change the subscripts. (subscripts balance charge, coefficients balance equations.
   a) Start by balancing those elements that appear in only one reactant and one product. The ones that occur the least.
   b) Balance those elements that appear in two or more reactants or products.

3. Check to make sure that you have the same number of each type of atom on both sides of the equation.

Link to an Extra Tips Video

Link to Extra Practice



Balancing Chemical Equations

Given the balanced chemical equation for ethane (C_2H_6) reacting with oxygen to form carbon dioxide and water (the combustion of ethane).

$C_2H_6(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$

Start with Carbon (occurs least): 2 C reactants, but 1 C product

$C_2H_6(g) + O_2(g) \longrightarrow 2CO_2(g) + H_2O(l)$

Notice: $2CO_2$ **NOT** C_2O_4
Now, 6 H reactants, but 2 H products

$C_2H_6(g) + O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

Now, 2 O reactants, but 7 O products

Since there is a diatomic oxygen in the reactants, and we need 7 oxygens in the reactants, we use a fraction. We can get rid of the fraction by multiplying the entire equation by 2

$C_2H_6(g) + \frac{7}{2}O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

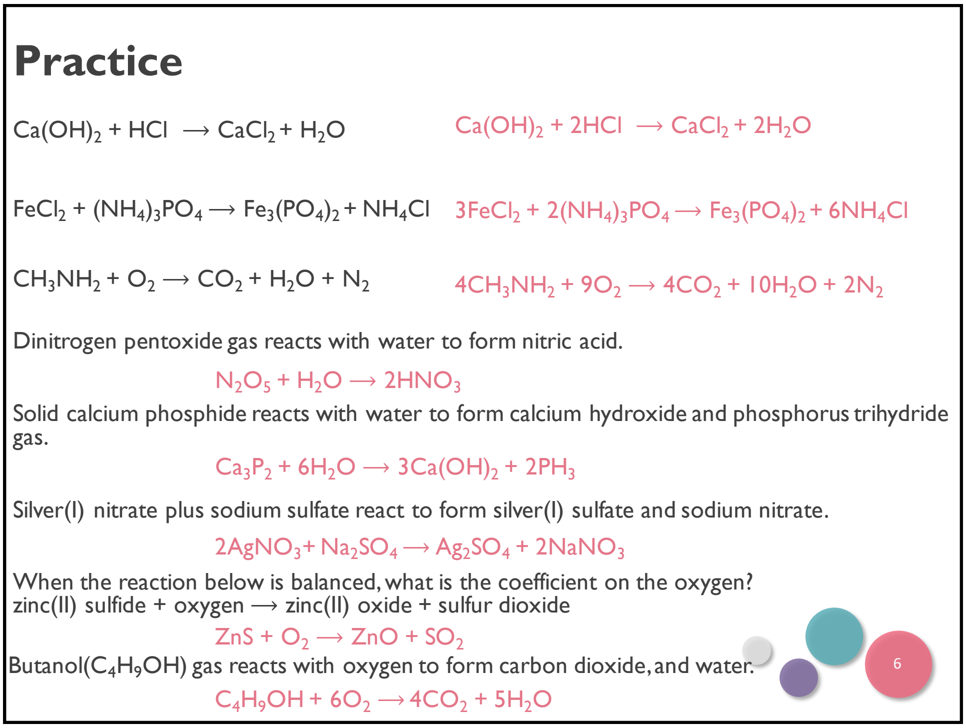
Can multiply by 2 to get rid of fraction.

$(2)(C_2H_6(g) + \frac{7}{2}O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l))$

Multiply by 2

$\boxed{2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(l))}$

Check it is balanced.



Practice

$Ca(OH)_2 + HCl \longrightarrow CaCl_2 + H_2O$ | $Ca(OH)_2 + 2HCl \longrightarrow CaCl_2 + 2H_2O$

$FeCl_2 + (NH_4)_3PO_4 \longrightarrow Fe_3(PO_4)_2 + NH_4Cl$ | $3FeCl_2 + 2(NH_4)_3PO_4 \longrightarrow Fe_3(PO_4)_2 + 6NH_4Cl$

$CH_3NH_2 + O_2 \longrightarrow CO_2 + H_2O + N_2$ | $4CH_3NH_2 + 9O_2 \longrightarrow 4CO_2 + 10H_2O + 2N_2$

Dinitrogen pentoxide gas reacts with water to form nitric acid.

$N_2O_5 + H_2O \longrightarrow 2HNO_3$

Solid calcium phosphide reacts with water to form calcium hydroxide and phosphorus trihydride gas.

$Ca_3P_2 + 6H_2O \longrightarrow 3Ca(OH)_2 + 2PH_3$

Silver(I) nitrate plus sodium sulfate react to form silver(I) sulfate and sodium nitrate.

$2AgNO_3 + Na_2SO_4 \longrightarrow Ag_2SO_4 + 2NaNO_3$

When the reaction below is balanced, what is the coefficient on the oxygen?
zinc(II) sulfide + oxygen ⟶ zinc(II) oxide + sulfur dioxide

$ZnS + O_2 \longrightarrow ZnO + SO_2$

Butanol(C_4H_9OH) gas reacts with oxygen to form carbon dioxide, and water.

$C_4H_9OH + 6O_2 \longrightarrow 4CO_2 + 5H_2O$

Reaction Types

Combination: two or more reactants combine to form a single product

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

Decomposition: two or more products form from a single reactant

$$CaCO_3(s) \rightarrow \underline{CaO}(s) + \underline{CO_2}(g)$$

Combustion: a substance burns in the presence of oxygen. Combustion of a compound that contains C and H (or C, H, and O) produces carbon dioxide gas and liquid water.

$$CH_2O(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

Double displacement (metathesis, exchange): can be precipitation reaction or molecular product (CO_2, SO_2, H_2S)

$$Na_2SO_4 + 2AgNO_3 \rightarrow 2NaNO_3 + Ag_2SO_4$$

Single displacement: one solid metal exchanges to produce a different solid metal, hydrogen, or halogen.

$$Al(s) + CuSO_4\,(aq) \rightarrow Al_2(SO_4)_3\,(aq) + Cu(s)$$

Neutralization: (specific double displacement)- acid/base to get salt + water

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Condensation: two molecules combine, and water is released.

Practice

Determine whether each of the following equations represents a combination reaction, a decomposition reaction, or a combustion reaction:

a) $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$

   2 reactants; 1 product: Combination

b) $2HCO_2H(l) + O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$

   carbon with oxygen producing CO_2 and water: Combustion

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

   1 reactant; 2 products: Decomposition

Remember:

In a **molecular equation** compounds are represented by chemical formulas as though they exist in solution as molecules or formula units. (What we have been doing)

$$Na_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow 2NaOH(aq) + BaSO_4(s)$$

Reactions in which anions in two ionic compounds exchange cations are called **double replacement reactions**.

Complete and Net Ionic Equations

In a **complete ionic equation** compounds that exist completely or predominately as ions in solution are represented as those ions.

$$Na_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow 2NaOH(aq) + BaSO_4(s)$$

In the reaction between aqueous Na_2SO_4 and $Ba(OH)_2$ the aqueous species break into pieces as follows:

$$Na_2SO_4(aq) \longrightarrow 2Na^+(aq) + SO_4^{-2}\ (aq)$$
$$Ba(OH)_2(aq) \longrightarrow Ba^{2+}(aq) + 2OH^-(aq)$$
$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

Complete Ionic Equation

$$2Na^+(aq) + SO_4^{-2}\ (aq) + Ba^{2+}(aq) + 2OH^-(aq) \longrightarrow 2Na^+(aq) + 2OH^-(aq) + BaSO_4(s)$$

An equation that includes only the species that are involved in the reaction is called a **net ionic equation**. Ions that appear on both sides of the equation are called **spectator ions**. Spectator ions do not participate in the reaction and must be crossed out to find the net ionic equation.

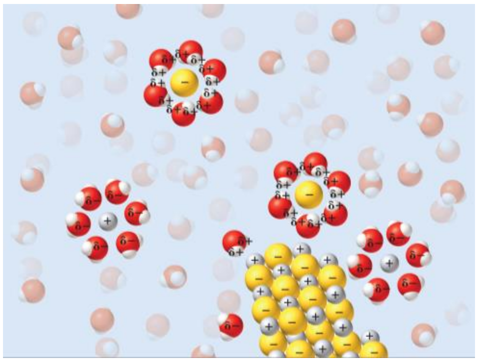
$$\cancel{2Na}^+(aq) + SO_4^{-2}\ (aq) + Ba^{2+}(aq) + \cancel{2OH}^-(aq) \longrightarrow \cancel{2Na}^+(aq) + \cancel{2OH}^-(aq) + BaSO_4(s)$$

$$\boxed{Ba^{2+}(aq) + SO_4^{-2}\ (aq) \longrightarrow BaSO_4(s)}$$

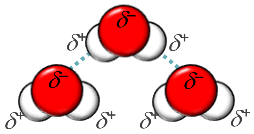
But how do you know what is aqueous (dissolves) and what doesn't?

Solubility Guidelines for Ionic Compounds in Water

Water is a good solvent for ionic compounds because it is a polar molecule. The polarity of water results from electron distributions within the molecule. The oxygen atom has an attraction for the hydrogen atoms' electrons and is therefore partially negative compared to hydrogen.



The oxygen atom is partially negative, and the hydrogen atoms are partially positive



Hydration occurs when water molecules remove the individual ions from an ionic solid surrounding them so the substances dissolves.

Solubility Guidelines for Ionic Compounds in Water

Solubility Rules (Memorize):

1. Salts of ammonium (NH_4^+) and Group IA are always soluble.
2. All chlorides (Cl^-) are soluble except $AgCl$, Hg_2Cl_2, and $PbCl_2$ which are insoluble.
3. All bromides (Br^-) are soluble except $AgBr$, Hg_2Br_2, $HgBr_2$, and $PbBr_2$ which are insoluble.
4. All iodides (I^-) are soluble except AgI, Hg_2I_2, HgI_2, and PbI_2 which are insoluble.
5. Chlorates (ClO_3^-), bicarbonates (HCO_3^-), nitrates (NO_3^-), and acetates (CH_3COO^- or $C_2H_3O_2^-$) are soluble.
6. Sulfates (SO_4^{-2}) are soluble except $CaSO_4$, $SrSO_4$, $BaSO_4$, Hg_2SO_4, $PbSO_4$, and Ag_2SO_4 which are insoluble.
7. Phosphates (PO_4^{-3}), chromates (CrO_3^{-2}), oxides (O^{-2}), sulfides (S^{-2}) and carbonates (CO_3^{-2}) are insoluble except NH_4^+ and Group IA compounds.
8. All metallic hydroxides (OH^-), are insoluble except NH_4^+, Group IA, and Ba^{2+}
9. Most silver compounds are insoluble.

The anion tells you the rule, and the cation tells you the exception

Which compound below is water soluble?

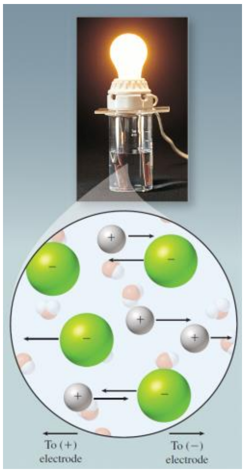
Al_2S_3 [✓ $KClO_3$] $BaSO_4$

Practice

Which compound below is water soluble?

Na_3PO_4	Soluble	Rule 1 and 7
$PbCl_2$	Insoluble	Rule 2
KI	Soluble	Rule 4
$MgCO_3$	Insoluble	Rule 7
$Ca_3(PO_4)_2$	Insoluble	Rule 7
$Mg(OH)_2$	Insoluble	Rule 8
$NaNO_3$	Soluble	Rules 1 and 5
$PbSO_4$	Insoluble	Rule 6

Electrolytes and Nonelectrolytes



An **electrolyte** is a substance that dissolves in water to yield a solution that conducts electricity. (**Ionic**) An electrolyte undergoes **dissociation** and breaks apart into its constituent ions.

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

Ionization is when a molecular compound forms ions when it dissolves

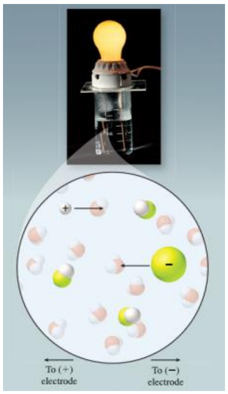
An electrolyte that dissociates completely (breaks apart completely) is known as a **strong electrolyte**.

Water soluble ionic compounds $NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$

Strong Acids $HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$

Strong Bases $NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$

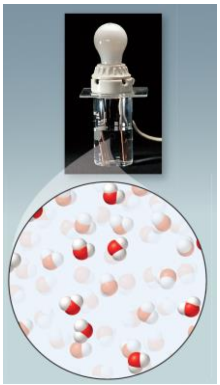
Electrolytes and Nonelectrolytes



A **weak electrolyte** is a compound that produces ions upon dissolving but exists in solution predominantly as molecules that are not ionized.

Weak Acids $HC_2H_3O_2(l) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$

Weak Bases $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$



The double arrow ($\rightleftharpoons$) denotes a reaction that occurs in both directions. When both the forward and reverse reactions occur at the same rate, the reaction is in a state of dynamic chemical equilibrium.

A **nonelectrolyte** is a substance that dissolves in water to yield a solution that does not conduct electricity.(**Molecular**)

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

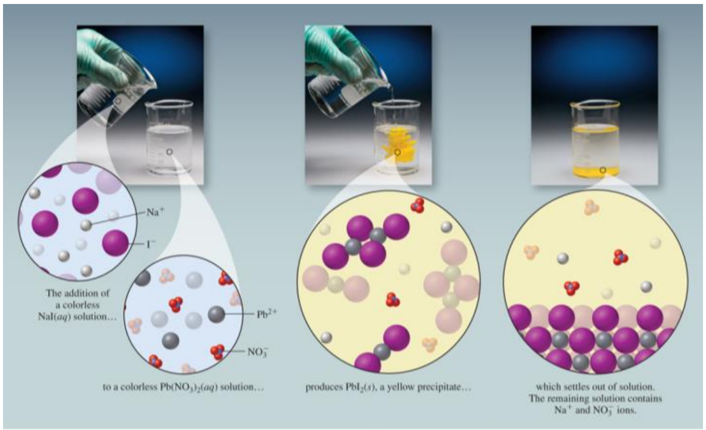
The sucrose molecules remain intact upon dissolving.

Precipitation Reactions

An insoluble product that separates from a solution is called a **precipitate**. In the example below, $PbI_2(s)$ is the precipitate.

$$2NaI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$$

A chemical reaction in which a precipitate forms is called a **precipitation reaction**.

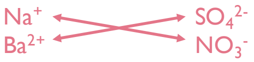


Precipitation Reactions

$Na_2SO_4(aq)$ and $Ba(NO_3)_2(aq)$ are mixed. Will they react (is there a product present that isn't aqueous?

$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow$

How do you find the products?



You get $NaNO_3$ and $BaSO_4$. Are either of these insoluble (solubility rules)?

Molecular Equation:

$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow$ **$BaSO_4(s)$** $+ 2\ NaNO_3(aq)$

Full Ionic Equation:

$\cancel{2Na^+}(aq) + SO_4^{-2}(aq) + Ba^{2+}(aq) + \cancel{2NO_3^-}(aq) \longrightarrow BaSO_4(s) + \cancel{2Na^+}(aq) + \cancel{2NO_3^-}(aq)$

Net Ionic Equation:

$SO_4^{-2}(aq) + Ba^{2+}(aq) \longrightarrow BaSO_4(s)$

Yes, a reaction occurs.

Practice

Add the physical states to the two products:

$ZnBr_2(aq) + 2AgNO_3(aq) \longrightarrow Zn(NO_3)_2 + 2AgBr$

What is the precipitate in this reaction?

$ZnBr_2(aq) + 2AgNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2AgBr(s)$

Net Ionic: $2Br^-(aq) + 2Ag^+(aq) \longrightarrow 2AgBr(s)$

Add the physical states to the two products:

$2NaI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2 + 2NaNO_3$

What is the precipitate in this reaction?

$2NaI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$

Net Ionic: $2I^-(aq) + Pb^{+2}(aq) \longrightarrow PbI_2(s)$

Practice (Don’t forget to balance)

Which combination of reactants will produce a precipitate?

A. $Ni(NO_3)_2(aq) + BaCl_2\ (aq) \longrightarrow Ba(NO_3)_2(aq) + NiCl_2(aq)$

B. $Na_2CO_3\ (aq) + 2NH_4Br(aq) \longrightarrow 2NaBr(aq) + (NH_4)_2CO_3(aq)$

C. $CuSO_4(aq) + 2KNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + K_2SO_4(aq)$

[✓ D.] $Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$

E. $MgSO_4(aq) + 2NH_4Cl(aq) \longrightarrow (NH_4)_2SO_4(aq) + MgCl_2(aq)$

Acid-Base Reactions

Acids can be either strong or weak. A **strong acid** = **strong electrolyte** and **strong bases = strong electrolytes** (dissociate completely). Strong bases are the hydroxides of Group 1A and heavy Group 2A.

Strong Acids (Memorize)

HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HNO_3	Nitric acid
H_2SO_4	Sulfuric acid
$HClO_4$	Perchloric acid

Strong Bases (Memorize)

LiOH	Lithium hydroxide
NaOH	Sodium hydroxide
KOH	Potassium hydroxide
RbOH	Rubidium hydroxide
CsOH	Cesium hydroxide
$Ca(OH)_2$	Calcium hydroxide
$Ba(OH)_2$	Barium hydroxide
$Sr(OH)_2$	Strontium hydroxide

Common Weak Acid

HF	Hydrofluoric acid
H_3PO_4	Phosphoric acid
CH_3COOH	Acetic acid
H_2CO_3	Carbonic acid
HCN	Hydrocyanic acid

Common Weak Base

NH_3	Ammonia
CH_3NH_2	Methylamine

Acid-Base Reactions

An **Arrhenius acid** is one that ionizes in water to produce protons (H^+).

$$\mathbf{HCl}(g) \xrightarrow{H_2O} \mathbf{H^+}(aq) + \mathbf{Cl^-}(aq)$$

An **Arrhenius base** is one that dissociates in water to produce OH^- ions.

$$\mathbf{NaOH}(s) \xrightarrow{H_2O} \mathbf{Na^+}(aq) + \mathbf{OH^-}(aq)$$

A **Brønsted acid** is a proton donor (loses a proton) and a **Brønsted base** is a proton acceptor (gains a proton).

In these definitions, a proton refers to a hydrogen atom that has lost its electron—also known as a hydrogen ion (H^+). If an acid is in water, the acid will donate a proton to water and the product is known as a hydronium ion (H_3O^+).

$$\left.\begin{array}{l}\text{hydrogen ion } (H^+) \\ \text{proton} \\ \text{hydronium ion } (H_3O^+)\end{array}\right\} \begin{array}{l}\text{All referencing the} \\ \text{same aqueous species}\end{array}$$

$\mathbf{HF}(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + F^-(aq)$

HF is acid, lost a H^+ (product is one proton less than reactant)

$\mathbf{NH_3}(aq) + H_2O(l) \rightleftharpoons \mathbf{NH_4^+}(aq) + OH^-(aq)$

NH_3 is base, gained a H^+ (product is one proton more than reactant.

Acid-Base Neutralization

A **monoprotic acid** has one proton to donate. Hydrochloric acid is an example:

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

A **polyprotic acid** has more than one acidic hydrogen atom. Sulfuric acid, H_2SO_4, is an example of a **diprotic acid**; there are two acidic hydrogen atoms.

Polyprotic acids lose protons in a stepwise fashion:

Step 1: $H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$

In H_2SO_4, the first ionization is strong.

Step 2: $HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$

In H_2SO_4, the second ionization occurs only to a very small extent (equilibrium established).

A **neutralization reaction** is a reaction between an acid and a base. Generally, a neutralization reaction produces **water** and a **salt**.

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

The net ionic equation of a strong acid– strong base reactions is:

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

Strong Acid-Strong Base Net Ionic

Strong Acid reacting with a Strong Base (remember that strong acids and bases completely ionize; they are strong electrolytes). Below is an example of a general strong acid/strong base reaction (M is metal):

Molecular Equation:

$HX(aq) + MOH(aq) \longrightarrow MX(aq) + H_2O(l)$

Full Ionic:

$H^+(aq) + \cancel{X^-}(aq) + \cancel{M^+}(aq) + OH^-(aq) \longrightarrow \cancel{M^+}(aq) + \cancel{X^-}(aq) + H_2O(l)$

Net Ionic:

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

Give the net ionic equation for the reaction between hydrochloric acid and sodium hydroxide.

$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{Na^+}(aq) + OH^-(aq) \longrightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$

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

Weak Acid-Strong Base Net Ionic

Weak Acid reacting with a Strong Base (remember that weak acids don't completely ionize, they will make the product, but we can't break them apart in full ionic equation. Below is an example of a general weak acid/strong base reaction:

Molecular Equation:

$HX(aq) + MOH(aq) \longrightarrow MX(aq) + H_2O(l)$

Full Ionic:

$\boxed{HX(aq)} + M^+(aq) + OH^-(aq) \longrightarrow M^+(aq) + X^-(aq) + H_2O(l)$

Weak electrolyte (weak acid) won't break apart in full ionic equation but it does make the product (MX).

Net Ionic:

$HX(aq) + OH^-(aq) \rightarrow X^-(aq) + H_2O(l)$

Acid/Base Net Ionic Equation Practice

Give the Net Ionic Equation for the following reactions:

$HCOOH + NaOH$ → **$HCOOH + OH^- \longrightarrow HCOO^- + H_2O$**

$Sr(OH)_2 + HCl$ → **$H^+ + OH^- \longrightarrow H_2O$**

$HF + LiOH$ → **$HF + OH^- \longrightarrow F^- + H_2O$**

$CH_3COOH + Ba(OH)_2$ → **$CH_3COOH + OH^- \longrightarrow CH_3COO^- + H_2O$**

$HBr + NaOH$ → **$H^+ + OH^- \longrightarrow H_2O$**

$HNO_3 + NH_3$ → **$H^+ + NH_3 \longrightarrow NH_4^+$**

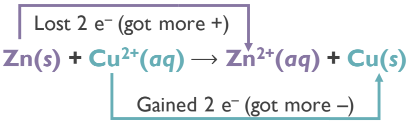
What is the net ionic equation for the acid-base reaction that occurs when acetic acid and potassium hydroxide solutions are mixed?

$CH_3COOH + OH^- \longrightarrow CH_3COO^- + H_2O$

Oxidation-Reduction Reactions

An **oxidation-reduction** (**RedOx**) reaction is a chemical reaction in which electrons are transferred from one reactant to another.

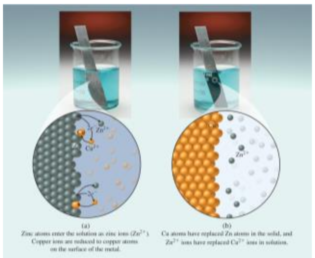
Oxidation is the loss of electrons. **Reduction** is the gain of electrons.



Lost 2 e^- (got more +)

$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Gained 2 e^- (got more –)



Zinc solid loses 2 e^- and becomes aqueous zinc (meaning it is no longer a solid in solution but is dissolved). Zinc was oxidized to Zn^{2+}. The compound that contains the elements/ion that is oxidized is the **Reducing Agent** (the compound that causes reduction), therefore Zn(s) is the reducing Agent.

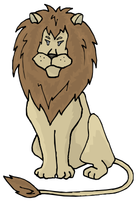
Copper(II) gained 2 e^- to become copper solid. Cu^{2+} was reduced to Cu(s). The compound that contains the elements/ion that is reduced is the **Oxidizing Agent** (the compound that causes the oxidation), therefore Cu^{2+} is the Oxidizing Agent.

RedOx Reactions and Electron Transfer

Loss of electrons is oxidation

Gain of electrons is reduction

LeO says GeR

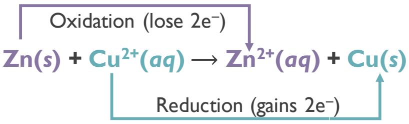


Oxidation is loss

Reduction is gain

OiL RiG

A redox reaction is the sum of an oxidation **half-reaction** and a reduction **half-reaction**.



Oxidation (lose $2e^-$)

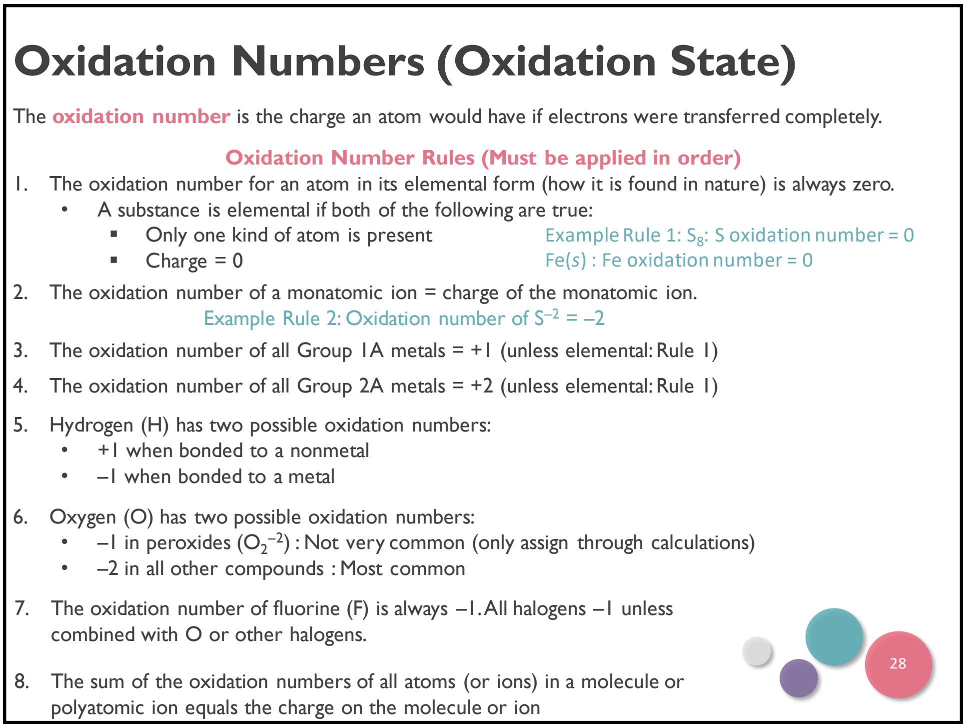
$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Reduction (gains $2e^-$)

Oxidation half-reaction: $Zn(s) \longrightarrow Zn^{2+}(aq) + \cancel{2e^-}$

Reduction half-reaction: $Cu^{2+}(aq) + \cancel{2e^-} \longrightarrow Cu(s)$

Overall redox reaction: $Cu^{2+}(aq) + Zn(s) \longrightarrow Zn^{2+}(aq) + Cu(s)$



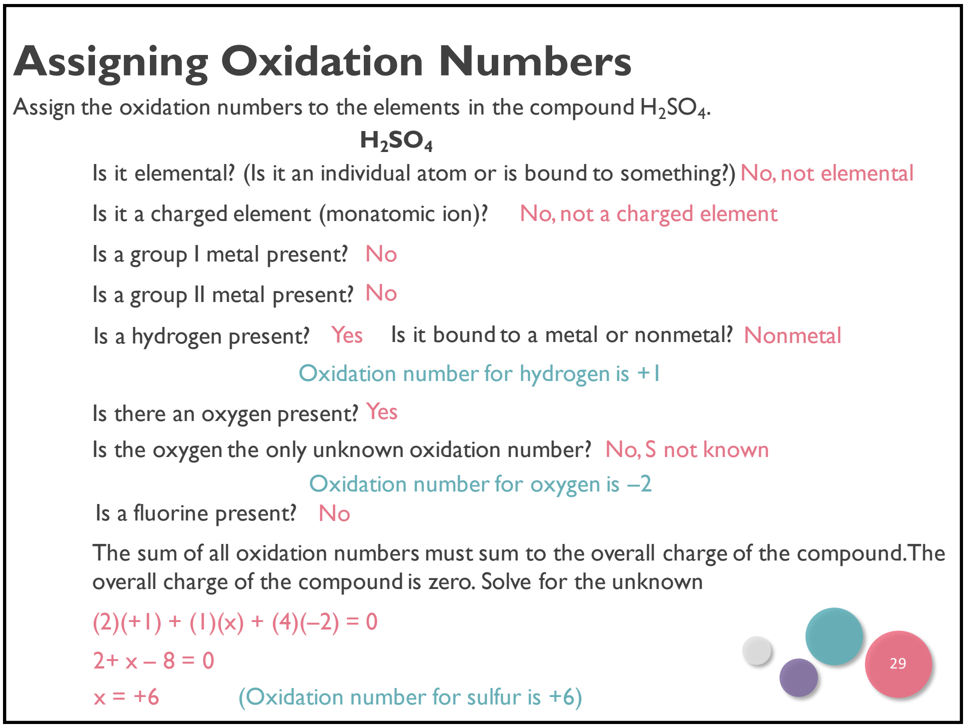
Oxidation Numbers (Oxidation State)

The **oxidation number** is the charge an atom would have if electrons were transferred completely.

Oxidation Number Rules (Must be applied in order)

1. The oxidation number for an atom in its elemental form (how it is found in nature) is always zero.
    - A substance is elemental if both of the following are true:
        - Only one kind of atom is present
        - Charge = 0

    Example Rule 1: S_8: S oxidation number = 0
    Fe(*s*) : Fe oxidation number = 0
2. The oxidation number of a monatomic ion = charge of the monatomic ion.
    Example Rule 2: Oxidation number of $S^{-2} = -2$
3. The oxidation number of all Group 1A metals = +1 (unless elemental: Rule 1)
4. The oxidation number of all Group 2A metals = +2 (unless elemental: Rule 1)
5. Hydrogen (H) has two possible oxidation numbers:
    - +1 when bonded to a nonmetal
    - –1 when bonded to a metal
6. Oxygen (O) has two possible oxidation numbers:
    - –1 in peroxides (O_2^{-2}) : Not very common (only assign through calculations)
    - –2 in all other compounds : Most common
7. The oxidation number of fluorine (F) is always –1. All halogens –1 unless combined with O or other halogens.
8. The sum of the oxidation numbers of all atoms (or ions) in a molecule or polyatomic ion equals the charge on the molecule or ion



Assigning Oxidation Numbers

Assign the oxidation numbers to the elements in the compound H_2SO_4.

H_2SO_4

Is it elemental? (Is it an individual atom or is bound to something?) No, not elemental

Is it a charged element (monatomic ion)? No, not a charged element

Is a group I metal present? No

Is a group II metal present? No

Is a hydrogen present? Yes Is it bound to a metal or nonmetal? Nonmetal

Oxidation number for hydrogen is +1

Is there an oxygen present? Yes

Is the oxygen the only unknown oxidation number? No, S not known

Oxidation number for oxygen is –2

Is a fluorine present? No

The sum of all oxidation numbers must sum to the overall charge of the compound. The overall charge of the compound is zero. Solve for the unknown

$(2)(+1) + (1)(x) + (4)(-2) = 0$

$2 + x - 8 = 0$

$x = +6$ (Oxidation number for sulfur is +6)

Assigning Oxidation Numbers

ClO_3^-

Is it elemental? (Is it an individual atom or is bound to something?) No, not elemental

Is it a charged element (monatomic ion)? No, not a charged element

Is a group I metal present? No

Is a group II metal present? No

Is a hydrogen present? No

Is there an oxygen present? Yes

Is the oxygen the only unknown oxidation number? No, Cl not known

Oxidation number for oxygen is –2

Is a fluorine present? No

The sum of all oxidation numbers must sum to the overall charge of the compound. The overall charge of the compound is zero. Solve for the unknown

$(1)(x) + (3)(-2) = -1$

$X - 6 = -1$

$x = +5$ (Oxidation number for chlorine is +5)

Oxidation Numbers

The oxidation number is sometimes called the ***oxidation state***.

	$H_2(g)$	+	$F_2(g)$	→	$2HF(g)$
Oxidation number:	0		0		+1 –1
Total contribution to charge:	0		0		+1 –1

	$N_2(g)$	+	$3H_2(g)$	→	$2NH_3(g)$
Oxidation number:	0		0		–3 +1
Total contribution to charge:	0		0		–3 +3

Practice

Assign the oxidation number for the underlined element.

$K\underline{\mathbf{N}}O_2$	N: +3
$\underline{\mathbf{Cs}}_2O$	Cs: +1
$\underline{\mathbf{C}}_2O_4^{2-}$	C: +3
$K\underline{\mathbf{Mn}}O_4$	Mn: +7
$Mg(\underline{\mathbf{N}}O_3)_2$	N: +5
$\underline{\mathbf{V}}O_3$	V: +6
$\underline{P}Cl_5$	P: +5
$H_3\underline{\mathbf{P}}O_3$	P: +3
$\underline{\mathbf{I}}F_7$	I: +7

Oxidation Numbers and RedOx Reactions

We use oxidation numbers to identify RedOx reactions. If there is an oxidation number change, then electrons (e^-) are transferred. If electrons are transferred it is a RedOx reaction.

Oxidation numbers also allow us to track what loses electrons and what gains electrons. If negative charge is gained, electrons are gained. If negative charge is lost, electrons are lost.

If an oxidation number gets more positive, it lost electrons, the element was oxidized. If an oxidation number get more negative, it gained electrons, the element was reduced.

Which species is reduced in the following reaction?

$2Al(s) + 3Br_2(l) \longrightarrow 2AlBr_3(s)$

Al

[✓ B. Br_2]

C. $AlBr_3$

D. None of them

$$\underset{0}{2Al(s)} + \underset{0}{3Br_2(l)} \longrightarrow 6\underset{+3}{Al}\underset{-1}{Br_3}(s)$$

Oxidation Number: 0 0 +3 −1

Al lost negative charge, got more positive, it was oxidized.

Br gained negative charge, got more negative, it was reduced.

What is the oxidation number of the chromium atom in K_2CrO_4?

A. 2 B. 4 [✓ C. 6] D. 8

Oxidation of Metals in Aqueous Solutions

In a single displacement reaction, an atom or an ion in a compound is replaced by an atom of another element.

$$\mathbf{Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)}$$

Oxidation Number: 0 +2 –1 +2 –1 0

Zinc displaces or replaces copper in the dissolved salt. Zn is oxidized to Zn^{2+}. Cu^{2+} is reduced to Cu. When a metal is oxidized by an aqueous solution, it becomes an aqueous ion.

The **activity series** is a list of metals (and hydrogen) arranged from top to bottom in order of decreasing ease of oxidation.

Increasing ease of oxidation

Oxidation Half Reaction
$Zn \longrightarrow Zn^{2+} + 2e^-$
$Fe \longrightarrow Fe^{2+} + 2e^-$
$Ni \longrightarrow Ni^{2+} + 2e^-$
$H_2 \longrightarrow 2H^+ + 2e^-$
$Cu \longrightarrow Cu^{2+} + 2e^-$
$Ag \longrightarrow Ag^+ + e^-$
$Au \longrightarrow Au^{3+} + 3e^-$

Metals listed at the **top** are called **active metals**.
Metals listed at the **bottom** are called **noble metals**.
An element in the series will be oxidized by the ions of any element that appears below it in the table.

Reaction 1: $\mathbf{Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)}$

Reaction 2: $\mathbf{Cu(s) + ZnCl_2(aq) \rightarrow}$ **no reaction**

Reaction 1, Zn is oxidized by Cu. This is possible because Zn(s) is above Cu on the activity series.

Oxidation of Metals in Aqueous Solutions

Which of the following reactions will occur?

- $Co(s) + BaI_2(aq) \longrightarrow$ No Reaction
  - Is Cobalt below Barium? Yes
  - Is Cobalt easier to oxidize than Barium? No; Ba above Co
- $Sn(s) + CuBr_2(aq) \longrightarrow SnBr_2(aq) + Cu(s)$
  - Is Tin below Copper? No
  - Is Tin easier to oxidize than Copper? Yes
- $Ag(s) + NaCl(aq) \longrightarrow$ No Reaction
  - Is Silver below Sodium? Yes
  - Is Silver easier to oxidize than Sodium? No; Na above Ag
- $Fe(s) + PtCl_2(aq) \longrightarrow Pt(s) + FeCl_2(aq)$
  - Is Iron below Platinum? No
  - Is Iron easier to oxidize than Platinum? Yes; Fe above Pt
- $Cr(s) + AuCl_3(aq) \longrightarrow Au(s) + CrCl_3$
  - Is Chromium below Gold? No
  - Is Chromium easier to oxidize than Gold? Yes; Cr above Au
- $Pb(s) + Zn(NO_3)_{2(aq)} \longrightarrow$ No Reaction
  - Is Lead below Zinc? Yes
  - Is Lead easier to oxidize than Zinc? No; Zn above Pb

Oxidation Half Reaction
$Ba \longrightarrow Ba^{2+} + 2e^-$
$Na \longrightarrow Na^+ + e^-$
$Zn \longrightarrow Zn^{2+} + 2e^-$
$Cr \longrightarrow Cr^{3+} + 3e^-$
$Fe \longrightarrow Fe^{2+} + 2e^-$
$Co \longrightarrow Co^{2+} + 2e^-$
$Sn \longrightarrow Sn^{2+} + 2e^-$
$Pb \longrightarrow Pb^{2+} + 2e^-$
$Cu \longrightarrow Cu^{2+} + 2e^-$
$Ag \longrightarrow Ag^+ + e^-$
$Pt \longrightarrow Pt^{2+} + 2e^-$
$Au \longrightarrow Au^{3+} + 3e^-$

Oxidation of Metals in Aqueous Solutions

Which of the following reactions will occur?

Is the solid above the aqueous?

$Cr(s) + Au^{3+}(aq) \longrightarrow$ Yes

$K(s) + CuCl_2(aq) \longrightarrow$ Yes

$Co(s) + Sn(NO_3)_2(aq) \longrightarrow$ Yes

$Pb(s) + Cr(OH)_3(aq) \longrightarrow$ No

$H_2(g) + ZnCl_2(aq) \longrightarrow$ No

$Zn(s) + HCl(aq) \longrightarrow$ Yes

Oxidation Half Reaction
$Li \longrightarrow Li^+ + e^-$
$K \longrightarrow K^+ + e^-$
$Ba \longrightarrow Ba^{2+} + 2e^-$
$Na \longrightarrow Na^+ + e^-$
$Zn \longrightarrow Zn^{2+} + 2e^-$
$Cr \longrightarrow Cr^{3+} + 3e^-$
$Fe \longrightarrow Fe^{2+} + 2e^-$
$Co \longrightarrow Co^{2+} + 2e^-$
$Sn \longrightarrow Sn^{2+} + 2e^-$
$Pb \longrightarrow Pb^{2+} + 2e^-$
$H_2 \longrightarrow 2H^+ + 2e^-$
$Cu \longrightarrow Cu^{2+} + 2e^-$
$Ag \longrightarrow Ag^+ + e^-$
$Pt \longrightarrow Pt^{2+} + 2e^-$
$Au \longrightarrow Au^{3+} + 3e^-$

Balancing Simple RedOx Equations

Redox reactions must have both mass balance and charge balance.

$$\mathbf{Cr(s) + Ni^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Ni(s)}$$

Oxidation half-reaction: $Cr(s) \longrightarrow Cr^{3+}(aq) + \boxed{3e^-}$

Reduction half-reaction: $Ni^{2+}(aq) + \boxed{2e^-} \longrightarrow Ni(s)$

Before adding half-reactions, the electrons must balance.

Step 1: Balance the Electrons: Multiply the oxidation half-reaction by 2

Oxidation half-reaction: $2\left[Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^- \right]$

Step 2: Balance the Electrons: Multiply the reduction half-reaction by 3

Reduction half-reaction: $3\left[Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s) \right]$

Oxidation half-reaction: $2Cr(s) \longrightarrow 2Cr^{3+}(aq) + \cancel{6e^-}$

Reduction half-reaction: $3Ni^{2+}(aq) + \cancel{6e^-} \longrightarrow 3Ni(s)$

$$3Ni^{2+}(aq) + 2Cr(s) \longrightarrow 3Ni(s) + 2Cr^{3+}(aq)$$

This is known as the ***half-reaction method***.

Example

Predict which of the following reactions will occur, and for those that will occur, balance the equation and indicate which element is oxidized and which is reduced: (a) $Al(s) + CaCl_2(aq) \rightarrow$?
(b) $Cr(s) + Pb(C_2H_3O_2)_2(aq) \rightarrow$ (c) $Sn(s) + HI(aq) \rightarrow$?

Solution

(a)No reaction. Calcium is above aluminum on the activity series table.

(b)The two half-reactions are represented by the following:

Oxidation: $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$
Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

In order to balance the charges, we must multiply the oxidation half-reaction by 2 and the reduction half-reaction by 3:

$(2)[Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-] = 2Cr(s) \rightarrow 2Cr^{3+}(aq) + 6e^-$
$(3)[Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)] = 3Pb^{2+}(aq) + 6e^- \rightarrow 3Pb(s)$

We can then add the two half-reactions, canceling the electrons on both sides to get

$$2Cr(s) + 3Pb^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Pb(s)$$

The overall balanced molecular equation is

$$2Cr(s) + 3Pb(C_2H_3O_2)_2(aq) \rightarrow 2Cr(C_2H_3O_2)_3(aq) + 3Pb(s)$$

Chromium is oxidized (0 to +3) and lead is reduced (+2 to 0).

Example Continued (Part C)

Predict which of the following reactions will occur, and for those that will occur, balance the equation and indicate which element is oxidized and which is reduced: (a) $Al(s) + CaCl_2(aq) \rightarrow$?
(b) $Cr(s) + Pb(C_2H_3O_2)_2(aq) \rightarrow$ (c) $Sn(s) + HI(aq) \rightarrow$?

Solution

(c)The two half-reactions are as follows:

Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
Reduction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Adding the two half-reactions and canceling the electrons on both sides yields

$$Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$$

The overall balanced molecular equation is

$$Sn(s) + 2HI(aq) \rightarrow SnI_2(aq) + H_2(g)$$

Tin is oxidized (0 to +2) and hydrogen is reduced (+1 to 0). Reactions in which hydrogen ion is reduced to hydrogen gas are known as **hydrogen displacement** reactions.

Think About It Check your conclusions by working each problem backward. Write each equation in reverse and compare the positions of the elements in the activity series.

Other Types of Redox Reactions

Combination reactions can involve oxidation and reduction. Hydrogen is oxidized from 0 to +1 and nitrogen is reduced from 0 to –3.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Decomposition can also be a redox reaction. Na^+ is reduced to Na and H^- is oxidized to H_2.

$$NaH(s) \longrightarrow 2Na(s) + 3H_2(g)$$

Disproportionation reactions occur when one element undergoes both oxidation and reduction. Oxygen in H_2O_2 (and other peroxides) has an oxidation number of –1.

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

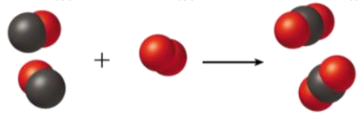
Combustion is also a redox process.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

The Mole and Chemical Reactions

Consider the complete reaction of 3.82 moles of CO to form CO_2. Calculate the number of moles of O_2 needed.

$$2\,CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$



The balanced chemical equation tells us that for every 2 CO moles we get 2 CO_2 moles, and for every 2CO moles we need 1O_2 mole; and for every 1O_2 mole we get 2CO_2 moles. These are the mole : mole conversion factors (they come from a **balanced** chemical equation).

$$\left(\frac{3.82\ \cancel{\text{mol CO}}}{}\right)\left(\frac{1\text{ mol O}_2}{2\ \cancel{\text{mol CO}}}\right) = 1.91\text{ mol O}_2$$

The Mole and Chemical Reactions

What **mass** of O_2 and Br_2 is produced by the reaction of 25.0 g of TiO_2 with excess BrF_3?

$$3\,TiO_2(s) + 4\,BrF_3(l) \longrightarrow 3\,TiF_4(s) + 2\,Br_2(l) + 3\,O_2(g)$$

Make sure equation is balance to find Stoichiometric ratios: $3TiO_2$ gives $3O_2$; $3TiO_2$ gives $2Br_2$; along with others…… everything in equation is related through coefficients

Excess BrF_3 means that there is enough BrF_3 to react all the TiO_2.

$$\left(\frac{25.0\ \cancel{g\ TiO_2}}{}\right)\left(\frac{1\ \cancel{mol\ TiO_2}}{79.88\ \cancel{g\ TiO_2}}\right)\left(\frac{3\ \cancel{mol\ O_2}}{3\ \cancel{mol\ TiO_2}}\right)\left(\frac{32.0\ g\ O_2}{1\ \cancel{mol\ O_2}}\right) = 10.0\ g\ O_2$$

$$\left(\frac{25.0\ \cancel{g\ TiO_2}}{}\right)\left(\frac{1\ \cancel{mol\ TiO_2}}{79.88\ \cancel{g\ TiO_2}}\right)\left(\frac{2\ \cancel{mol\ Br_2}}{3\ \cancel{mol\ TiO_2}}\right)\left(\frac{159.81\ g\ Br_2}{1\ \cancel{mol\ Br_2}}\right) = 33.4\ g\ Br_2$$

The Mole and Chemical Reactions

Mole Reactant → Mole Product

Mass Reactant → Mole Reactant → Mole Product

Mass Reactant → Mole Reactant → Mole Product → Mass Product

Write the balanced chemical equation. Convert quantities of known substances into moles, then use coefficients in balanced equation to calculate the number of moles of the sought quantity, and convert moles of sought quantity into desired units

Methanol burns in air according to the equation

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

If 209 g of methanol are used up in the combustion, what mass of water is produced?

grams CH_3OH → moles CH_3OH → moles H_2O → grams H_2O

$$\left(\frac{209\ \cancel{g\ CH_3OH}}{}\right)\left(\frac{1\ \cancel{mol\ CH_3OH}}{32.0\ \cancel{g\ CH_3OH}}\right)\left(\frac{4\ \cancel{mol\ H_2O}}{2\ \cancel{mol\ CH_3OH}}\right)\left(\frac{18.0\ g\ H_2O}{1\ \cancel{mol\ H_2O}}\right) = 235\ g\ H_2O$$

Limiting Reactants (Limiting Reagents)

The reactant used up first in a reaction is called the **limiting reactant**.

Excess reactants are those present in quantities greater than necessary to react with the quantity of the limiting reactant.

Calculate how much product you can generate with both reactants. The one that produces less product is the limiting reagent!

Consider the reaction between **5 moles of CO** and **8 moles of H_2** to produce methanol via the following reaction:

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(l)$$

How many moles of CH_3OH can be made from 5 mol CO?

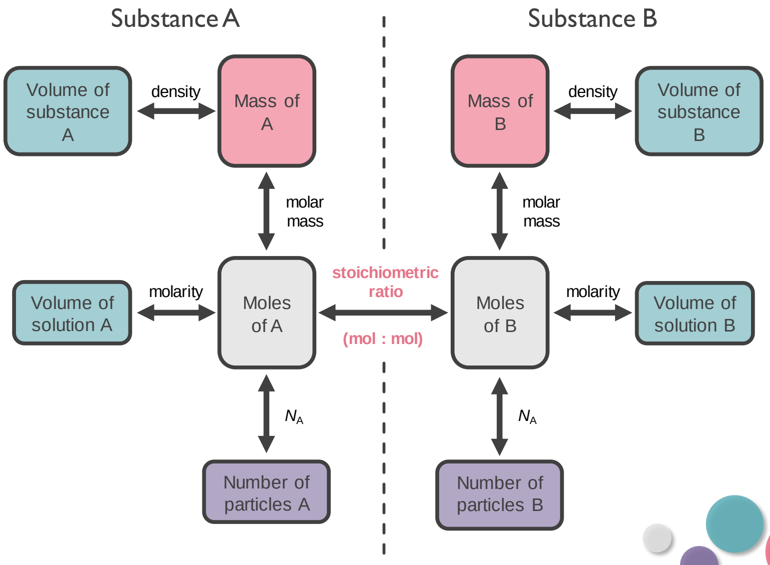
$$\left(\frac{5\ \cancel{\text{mol CO}}}{}\right)\left(\frac{1\ \text{mol CH}_3\text{OH}}{1\ \cancel{\text{mol CO}}}\right) = 5\ \text{mol CH}_3\text{OH}$$

How many moles of CH_3OH can be made from 8 mol H_2?

$$\left(\frac{8\ \cancel{\text{mol H}_2}}{}\right)\left(\frac{1\ \text{mol CH}_3\text{OH}}{2\ \cancel{\text{mol H}_2}}\right) = 4\ \text{mol CH}_3\text{OH}$$

H_2 will be used up first - **limiting reactant**.
CO is still available - **excess reactant**.

Limiting Reactants (Limiting Reagents)



Substance A

Substance B

Volume of substance A ⟷ density ⟷ Mass of A

Mass of A ⟷ molar mass ⟷ Moles of A

Volume of solution A ⟷ molarity ⟷ Moles of A

Moles of A ⟷ N_A ⟷ Number of particles A

Moles of A ⟷ stoichiometric ratio (mol : mol) ⟷ Moles of B

Mass of B ⟷ density ⟷ Volume of substance B

Mass of B ⟷ molar mass ⟷ Moles of B

Moles of B ⟷ molarity ⟷ Volume of solution B

Moles of B ⟷ N_A ⟷ Number of particles B

Practice

In one process, 124 g of Al are reacted with 601 g of Fe_2O_3, calculate the mass of Al_2O_3 formed. Calculate the mass of excess reactant left after the reaction is complete (assuming the reaction finishes completely)

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

Maximum Amount Possible: Theoretical Yield

$$\left(\frac{124\ \cancel{g\ Al}}{}\right)\left(\frac{1\ \cancel{mol}\ Al}{27.0\ \cancel{g\ Al}}\right)\left(\frac{1\ \cancel{mol\ Al_20_3}}{2\ \cancel{mol}\ Al}\right)\left(\frac{102.\ g\ Al_20_3}{1\ \cancel{mol\ Al_20_3}}\right) = \boxed{234\ g\ Al_20_3\ possible}$$

$$\left(\frac{601\ \cancel{g\ Fe_20_3}}{}\right)\left(\frac{1\ \cancel{mol\ Fe_20_3}}{158.\ \cancel{g\ Fe_20_3}}\right)\left(\frac{1\ \cancel{mol\ Al_20_3}}{1\ \cancel{mol\ Fe_20_3}}\right)\left(\frac{102.\ g\ Al_20_3}{1\ \cancel{mol\ Al_20_3}}\right) = 388\ g\ Al_20_3\ possible$$

Al is the limiting reagent and Fe_2O_3 is the excess reagent. To find the amount of excess reagent left over, you find how much excess reagent it requires to make the maximum amount of product possible from the limiting reagent.

$$\left(\frac{234\ \cancel{g\ Al_20_3}}{}\right)\left(\frac{1\ \cancel{mol\ Al_20_3}}{102.\ \cancel{g\ Al_20_3}}\right)\left(\frac{1\ \cancel{mol\ Fe_20_3}}{1\ \cancel{mol\ Al_20_3}}\right)\left(\frac{158.\ g\ Fe_20_3}{1\ \cancel{mol\ Fe_20_3}}\right) = 362\ g\ Fe_20_3\ USED$$

Excess Reactant Used + Excess Reactant Left Over = Total Excess Reactant

Excess Reactant Left Over = 601 g – 362 g = 239 g Fe_2O_3 left over

Practice

$$3\ NH_4NO_3 + Na_3PO_4 \longrightarrow (NH_4)_3PO_4 + 3\ NaNO_3$$

If you start with 30 g ammonium nitrate and 50 g sodium phosphate: a. what is the limiting reactant? b. how many grams of $NaNO_3$ would you make? c. How much of the excess reactant was used up? d. how much of the excess reactant was left over?

$$\left(\frac{30\ \cancel{g\ NH_4}NO_3}{}\right)\left(\frac{1\ \cancel{mol\ NH_4}NO_3}{80.06\ \cancel{g\ NH_4}NO_3}\right)\left(\frac{3\ \cancel{mol\ Na}NO_3}{3\ \cancel{mol\ NH_4}NO_3}\right)\left(\frac{85.\ g\ NaNO_3}{1\ \cancel{mol\ NaNO_3}}\right) = 31.85\ g\ NaNO_3\ possible$$

$$\left(\frac{50\ \cancel{g\ Na_3}PO_4}{}\right)\left(\frac{1\ \cancel{mol}\ Na_3PO_4}{163.94\ \cancel{g\ Na_3}PO_4}\right)\left(\frac{3\ \cancel{mol}\ NaNO_3}{1\ \cancel{mol\ Na_3}PO_4}\right)\left(\frac{85.\ g\ NaNO_3}{1\ \cancel{mol}\ NaNO_3}\right) = 77.77\ g\ NaNO_3\ possible$$

a. NH_4NO_3 Limiting Reactant (so Na_3PO_4 excess reactant)

b. 31.85 g $NaNO_3$ maximum possible to make

c. $$\left(\frac{31.85\ \cancel{g\ NaN}O_3}{}\right)\left(\frac{1\ \cancel{mol}\ NaNO_3}{85.\ \cancel{g\ NaNO_3}}\right)\left(\frac{1\ \cancel{mol\ Na_3}PO_4}{3\ \cancel{mol}\ NaNO_3}\right)\left(\frac{163.94\ g\ Na_3PO_4}{1\ \cancel{mol\ Na_3}PO_4}\right)$$

$$= 20.48\ g\ Na_3PO_4\ used\ to\ make\ max\ amount\ of\ NaNO_3$$

d. $50\ gNa_3PO_4\ - 20.48\ g\ Na_3PO_4 = 29.52\ g\ Na_3PO_4$ left over

Reaction Yield

The **theoretical yield** is the amount of product that forms when all the limiting reactant reacts to form the desired product. This is the maximum amount of product that can form from the limiting reactant (based on calculations)

The **actual yield** is the amount of product actually obtained from a reaction.

The **percent yield** tells what percentage the actual yield is of the theoretical yield.

$$\%\ \text{Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}}\ x\ 100\%$$

Practice

Consider the following reaction:

$$C_4H_9OH + NaBr + H_2SO_4 \longrightarrow C_4H_9Br + NaHSO_4 + H_2O$$

If 15.0 g of C_4H_9OH react with 22.4 g of NaBr and 32.7 g of H_2SO_4 to yield 17.1 g of C_4H_9Br, what is the percent yield of this reaction?

$$\left(\frac{15.0\ \cancel{g\ C_4H_9}OH}{}\right)\left(\frac{1\ \cancel{mol\ C_4}H_9OH}{74.14\ \cancel{g\ C_4H_9}OH}\right)\left(\frac{1\ \cancel{mol\ C_4}H_9Br}{1\ \cancel{mol\ C_4}H_9OH}\right)\left(\frac{137.03\ g\ C_4H_9Br}{1\ \cancel{mol\ C_4}H_9Br}\right) = \begin{matrix}27.72\ g\ C_4H_9Br\\ \text{possible}\end{matrix}$$

$$\left(\frac{22.4\ \cancel{g\ NaBr}}{}\right)\left(\frac{1\ \cancel{mol\ NaBr}}{102.89\ \cancel{g\ NaBr}}\right)\left(\frac{1\ \cancel{mol\ C_4}H_9Br}{1\ \cancel{mol\ NaBr}}\right)\left(\frac{137.03\ g\ C_4H_9Br}{1\ \cancel{mol\ C_4}H_9Br}\right) = \begin{matrix}29.83\ g\ C_4H_9Br\\ \text{possible}\end{matrix}$$

$$\left(\frac{32.7\ \cancel{g\ H_2SO_4}}{}\right)\left(\frac{1\ \cancel{mol\ H_2}SO_4}{98.09\ \cancel{g\ H_2SO_4}}\right)\left(\frac{1\ \cancel{mol\ C_4}H_9Br}{1\ \cancel{mol\ H_2}SO_4}\right)\left(\frac{137.03\ g\ C_4H_9Br}{1\ \cancel{mol\ C_4}H_9Br}\right) = \begin{matrix}45.68\ g\ C_4H_9Br\\ \text{possible}\end{matrix}$$

$$\%\ \text{Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}}\ x\ 100\%$$

$$\%\ \text{Yield} = \frac{17.1}{27.72}\ x\ 100\%$$

$$\%\ \text{Yield} = 61.7\%$$

Aqueous Reactions and Chemical Analysis

Gravimetric analysis is an analytical technique based on the measurement of mass. Gravimetric analysis is highly accurate.Applicable only to reactions that go to completion or have nearly 100 % yield.

A 0.8633-g sample of an ionic compound containing chloride ions and an unknown metal cation is dissolved in water and treated with an excess of $AgNO_3$.If 1.5615 g of AgCl precipitate forms, what is the percent by mass of Cl in the original sample?

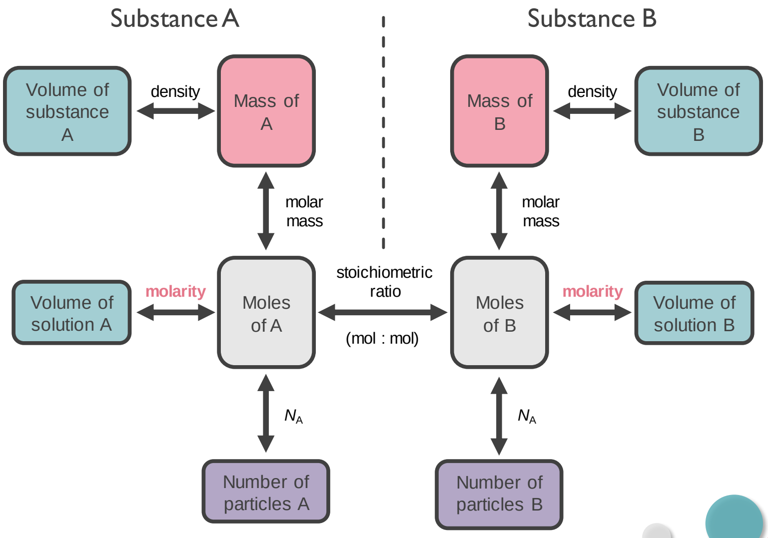
Strategy Using the mass of AgCl precipitate and the percent composition of AgCl, determine what mass of chloride the precipitate contains. The chloride in the precipitate was originally in the unknown compound. Using the mass of chloride and the mass of the original sample, determine the percent Cl in the compound.

$$\frac{35.45 \text{ g Cl}}{(35.45 \text{ g} + 107.9 \text{ g AgCl})} \; x \; 100\% = 24.72\% \text{ Cl in AgCl}$$

$$(0.2472)(1.5615 \text{ g AgCl}) = 0.3860 \text{ g Cl in AgCl precipitate formed}$$

$$\frac{0.3860 \text{ g Cl in AgCl}}{0.8633 \text{ g unknown sample}} \text{ x } 100\% = 44.71\% \text{ Cl in Unknown Sample}$$

Molarity and Reaction in Aqueous Solution



Molarity and Reaction in Aqueous Solution

What volume, in mL, of 0.0875 M H_2SO_4 is required to neutralize 25.0 mL of 0.234 M NaOH?

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$\text{Volume (L)} \xrightarrow{\text{Molarity}} \text{Moles} \xrightarrow[\text{Mol : Mol}]{} \text{Moles of Sought Compound} \xrightarrow{\text{Molarity}} \text{Volume of Sought Compound}$$

$$\left(\frac{0.025\ \cancel{\text{L NaOH}}}{}\right)\left(\frac{0.234\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{L NaOH}}}\right)\left(\frac{1\ \cancel{\text{mol } H_2SO_4}}{2\ \cancel{\text{mol NaOH}}}\right)\left(\frac{1\ \cancel{\text{L } H_2SO_4}}{0.0875\ \cancel{\text{mol } H_2SO_4}}\right)\left(\frac{1000\ \text{mL}}{\cancel{1\ \text{L}}}\right) = 33.4\ \text{mL}$$

Molarity and Reaction in Aqueous Solution

25.0 mL of 0.234 M $FeCl_3$ and 50.0 mL of 0.453 M NaOH are mixed. Which reactant is limiting? How many moles of $Fe(OH)_3$ will form?

$$FeCl_3(aq) + 3\ NaOH(aq) \longrightarrow 3\ NaCl(aq) + Fe(OH)_3(s)$$

$$\text{Volume (L)} \xrightarrow{\text{Molarity}} \text{Moles} \xrightarrow[\text{Mol : Mol}]{} \text{Moles of Sought Compound}$$

$$\left(\frac{0.025\ \cancel{\text{L } FeCl_3}}{}\right)\left(\frac{0.234\ \cancel{\text{mol } FeCl_3}}{1\ \cancel{\text{L } FeCl_3}}\right)\left(\frac{1\ \text{mol } Fe(OH)_3}{1\ \cancel{\text{mol } FeCl_3}}\right) = 0.00585\ \text{mol } Fe(OH)_3$$

$$\left(\frac{0.050\ \cancel{\text{L NaOH}}}{}\right)\left(\frac{0.453\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{L NaOH}}}\right)\left(\frac{1\ \text{mol } Fe(OH)_3}{3\ \cancel{\text{mol NaOH}}}\right) = 0.00755\ \text{mol } Fe(OH)_3$$

Acid-Base Titrations

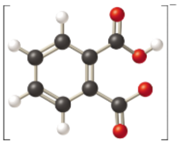
Quantitative studies of acid-base neutralization reactions are most conveniently carried out using a technique known as a **titration**.

A titration is a volumetric technique that uses burettes. The point in the titration where the acid has been neutralized is called the **equivalence point**. The equivalence point is usually signalled by a colour change. (A link to help with understanding titrations: titration curve link)

The colour change is brought about by the use of an **indicator**. Indicators have distinctly different colours in acidic and basic media. The indicator is chosen so that the colour change, or **endpoint**, is very close to the equivalence point. Phenolphthalein is a common indicator.

Sodium hydroxide solutions are commonly used in titrations. NaOH solutions must be **standardized** as the concentrations change over time. (NaOH reacts with CO_2 that slowly dissolves into the solution forming carbonic acid.)

The acid potassium hydrogen phthalate (KHP) is frequently used to standardize NaOH solutions. It only has 1 ionizable proton.



$HC_8H_4O_4^-$

Practice

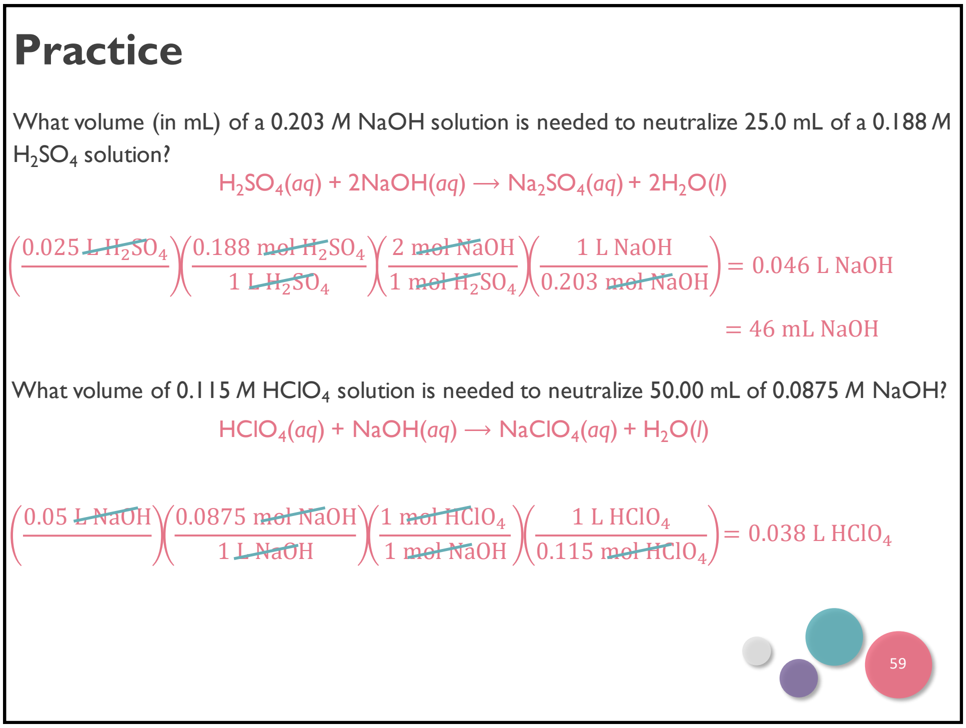
In a titration experiment, a student finds that 25.49 mL of an NaOH solution is needed to neutralize 0.7137 g of KHP (a monoprotic acid with a MM=204.2 g mol⁻¹). What is the concentration (in *M*) of the NaOH solution?

$$\text{NaOH} + \text{HX} \longrightarrow \text{NaX} + H_2O(l)$$

Strategy Using the mass given and the molar mass of KHP, determine the number of moles of KHP. Recognize that the number of moles of NaOH in the volume given is equal to the number of moles of KHP. Divide moles of NaOH by volume (in liters) to get molarity.

$$\left(\frac{0.7137\ \cancel{\text{g KHP}}}{}\right)\left(\frac{1\ \cancel{\text{mol KHP}}}{204.2\ \cancel{\text{g KHP}}}\right)\left(\frac{1\ \text{mol NaOH}}{1\ \cancel{\text{mol KHP}}}\right) = 0.003495\ \text{mol NaOH}$$

$$\frac{0.003495\ \text{mol NaOH}}{0.02549\ \text{L NaOH}} = 0.1371\ M$$



Practice

What volume (in mL) of a 0.203 *M* NaOH solution is needed to neutralize 25.0 mL of a 0.188 *M* H_2SO_4 solution?

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$\left(\frac{0.025\ \cancel{\text{L}\ H_2SO_4}}{}\right)\left(\frac{0.188\ \cancel{\text{mol}\ H_2SO_4}}{1\ \cancel{\text{L}\ H_2SO_4}}\right)\left(\frac{2\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol}\ H_2SO_4}}\right)\left(\frac{1\ \text{L NaOH}}{0.203\ \cancel{\text{mol NaOH}}}\right) = 0.046\ \text{L NaOH}$$

$$= 46\ \text{mL NaOH}$$

What volume of 0.115 *M* $HClO_4$ solution is needed to neutralize 50.00 mL of 0.0875 *M* NaOH?

$$HClO_4(aq) + NaOH(aq) \longrightarrow NaClO_4(aq) + H_2O(l)$$

$$\left(\frac{0.05\ \cancel{\text{L NaOH}}}{}\right)\left(\frac{0.0875\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{L NaOH}}}\right)\left(\frac{1\ \cancel{\text{mol}\ HClO_4}}{1\ \cancel{\text{mol NaOH}}}\right)\left(\frac{1\ \text{L}\ HClO_4}{0.115\ \cancel{\text{mol}\ HClO_4}}\right) = 0.038\ \text{L}\ HClO_4$$

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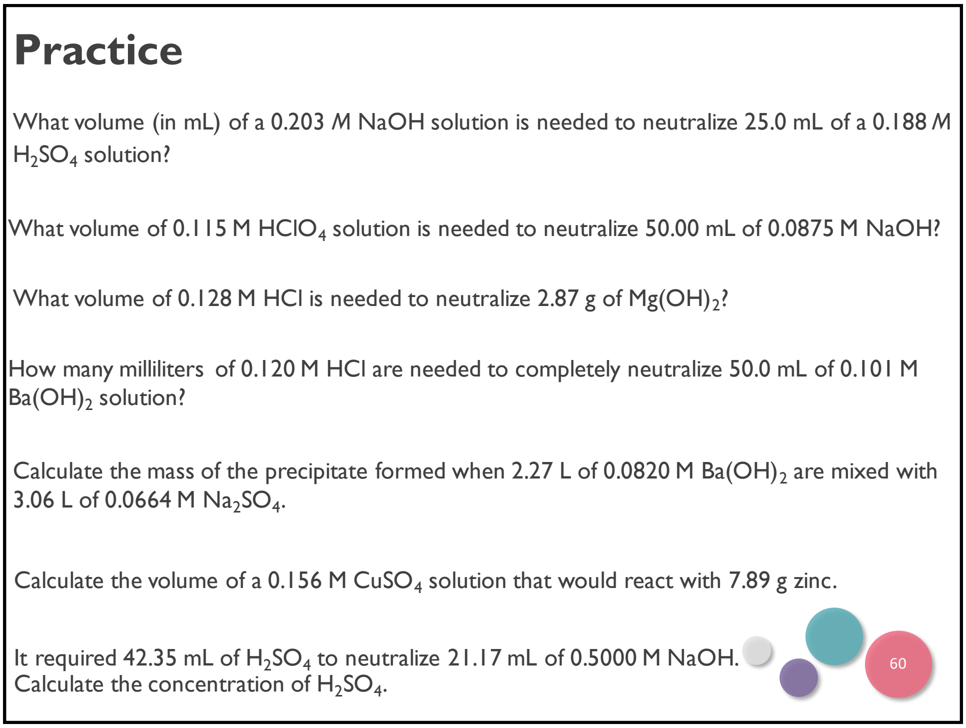

Practice

What volume (in mL) of a 0.203 *M* NaOH solution is needed to neutralize 25.0 mL of a 0.188 *M* H_2SO_4 solution?

What volume of 0.115 M $HClO_4$ solution is needed to neutralize 50.00 mL of 0.0875 M NaOH?

What volume of 0.128 M HCl is needed to neutralize 2.87 g of $Mg(OH)_2$?

How many milliliters of 0.120 M HCl are needed to completely neutralize 50.0 mL of 0.101 M $Ba(OH)_2$ solution?

Calculate the mass of the precipitate formed when 2.27 L of 0.0820 M $Ba(OH)_2$ are mixed with 3.06 L of 0.0664 M Na_2SO_4.

Calculate the volume of a 0.156 M $CuSO_4$ solution that would react with 7.89 g zinc.

It required 42.35 mL of H_2SO_4 to neutralize 21.17 mL of 0.5000 M NaOH. Calculate the concentration of H_2SO_4.

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