5 Chemical Reactions
5.1 Exchange Reactions: Precipitation and Net Ionic Equations (p. 165)

An ________________ is a compound whose aqueous solution conducts electricity.

___________ electrolytes are solutes that exist in solution predominantly in the form of hydrated (solvated) ions. A good example is sodium chloride dissolved in water.

\[
\text{NaCl(s) + H}_2\text{O(l) -> NaCl(aq)}
\]

How sodium chloride exists in solution. ________________________________________

___________ electrolytes are solutes that exist in solution predominantly in the form of hydrated (solvated) molecules. A good example is oxalic acid dissolved in water.

\[
\text{H}_2\text{C}_2\text{O}_4\text{(s) + H}_2\text{O(l) -> H}_2\text{C}_2\text{O}_4\text{(aq)}
\]

How oxalic acid exists in solution. ____________________________________________

___________ are substances that do not conduct an observable amount of electricity when dissolved in water. A good example is glucose dissolved in water.

\[
\text{C}_6\text{H}_{12}\text{O}_6\text{(s) + H}_2\text{O(l) -> C}_6\text{H}_{12}\text{O}_6\text{(aq)}
\]

Glucose does not dissociate, but exists in solution exclusively as the hydrated molecule, C₆H₁₂O₆(aq).

There are three types of strong electrolytes.

1. ______________  __________  2. ______________  __________  3. ______________

Soluble salts ______________  __________ in solution!

<table>
<thead>
<tr>
<th>Solubility Rules for Ionic Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonium, Group 1A, NH₄⁺, Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺</td>
</tr>
<tr>
<td>Nitrates, NO₃⁻</td>
</tr>
<tr>
<td>Chlorides, bromides, iodides, Cl⁻, Br⁻, I⁻</td>
</tr>
<tr>
<td>Sulfates, SO₄²⁻</td>
</tr>
<tr>
<td>Chlorates, ClO₃⁻</td>
</tr>
<tr>
<td>Perchlorates, ClO₄⁻</td>
</tr>
<tr>
<td>Acetates, C₂H₃O₂⁻</td>
</tr>
<tr>
<td>Phosphates, PO₄³⁻</td>
</tr>
<tr>
<td>Carbonates, CO₃²⁻</td>
</tr>
<tr>
<td>Hydroxides, OH⁻</td>
</tr>
<tr>
<td>Oxides, O⁻</td>
</tr>
<tr>
<td>Oxalates, C₂O₄²⁻</td>
</tr>
<tr>
<td>Sulfides, S²⁻</td>
</tr>
</tbody>
</table>
Indicate whether the follow salts are soluble in water, and if so show the dissociation step with products.

<table>
<thead>
<tr>
<th>Salt</th>
<th>Soluble?</th>
<th>Dissociation products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonium sulfate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lead nitrate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Calcium sulfate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sodium oxalate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Silver iodide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Iron(III) sulfide</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Exchange or Metathesis Reaction or Double-Replacement Reaction (p. 167)**

An exchange reaction will occur when ions in solution form insoluble products, weak electrolytes, or nonelectrolytes. These are also called metathesis reactions.

**Exchange reactions have the general form** \( AB + CD \longrightarrow AD + CB. \)

Exchange reactions can be categorized as:

1. _______________________: one of the products is insoluble (e.g., \( \text{AgCl(s)} \)).
   \[
   \text{NaCl(aq)} + \text{AgNO}_3(\text{aq}) \longrightarrow
   \]

2. _______________________: most often one of the products is water, \( \text{H}_2\text{O(l)} \).
   \[
   \text{HCl(aq)} + \text{KOH(aq)} \rightarrow
   \]

3. _______________________: one of the products is a gas (e.g., \( \text{SO}_2(\text{g}) \)).
   \[
   \text{Na}_2\text{SO}_3(\text{aq}) + 2 \text{HCl(aq)} \longrightarrow
   \]


Precipitation Reactions (p. 168)

A precipitate is an insoluble solid compound formed during a chemical reaction in solution (e.g., CaCO₃(s), AgCl(s), BaSO₄(s)). Precipitation reactions are a type of chemical reaction between ions that produce a precipitate.

Based upon the solubility rules, a reaction can be predicted as to whether it will take place or not.

When solutions of electrolytes are mixed, a reaction will occur if one of the products is insoluble.

\[ \text{NaOH(aq)} + \text{ZnSO}_4(aq) \rightarrow \]

When solutions of electrolytes are mixed, a reaction does not occur if all of the potential products are soluble.

\[ \text{Na}_2\text{SO}_4(aq) + \text{KNO}_3(aq) \rightarrow \]

or

\[ \text{Na}_2\text{SO}_4(aq) + \text{KNO}_3(aq) \rightarrow \]

Complete and balance the following reactions and determine if a precipitation reaction would occur.

1) \((\text{NH}_4)_2\text{CO}_3(aq) + \text{CaCl}_2(aq) \rightarrow \)

2) \(\text{Na}_2\text{S(aq)} + \text{CoCl}_2(aq) \rightarrow \)

3) \(\text{K}_3\text{PO}_4(aq) + \text{NaNO}_3(aq) \rightarrow \)

4) \(\text{BaCl}_2(aq) + \text{ZnSO}_4(aq) \rightarrow \)

5) \(\text{KCl(aq)} + \text{AgNO}_3(aq) \rightarrow \)

6) \(\text{FeCl}_3(aq) + \text{NaOH(aq)} \rightarrow \)

7) \(\text{MgCl}_2(aq) + \text{Li}_3\text{PO}_4(aq) \rightarrow \)
Net Ionic Equations (p. 169)
To help get a better grasp of what is actually occurring in a precipitation reaction, the reaction equation can be written in three different forms.

First, the ___________________ __________________ equation shows compounds as if they exist in solution in the form of molecules, even though they are soluble ionic compounds and dissociate in solution.

\[(\text{NH}_4)_2\text{CO}_3(aq) + \text{CaCl}_2(aq) \rightarrow\]

Second, the ___________________ __________ equation shows the soluble ionic compounds as ions in solution and molecules are shown as nondissociated molecules in solution.

Finally, the _____ __________ equation shows only the species that take part in a reaction.

__________________ ions are the ions that do not take part in a reaction and therefore remain in the same form and state on either side of the equation arrow. In the reaction just cited the ammonium cation, \(\text{NH}_4^{1+}(aq)\), and the chloride anion, \(\text{Cl}^{1-}(aq)\), are spectator ions.

Write the balanced molecular equation, the complete ionic equation, and the net ionic equations for the following reactions. (See page 169-170 for a set of general steps for writing these type of equations.)

1) \((\text{NH}_4)_2\text{CO}_3(aq) + \text{CaCl}_2(aq) \rightarrow\)

2) \(\text{Na}_2\text{S}(aq) + \text{CoCl}_2(aq) \rightarrow\)

3) \(\text{BaCl}_2(aq) + \text{ZnSO}_4(aq) \rightarrow\)

4) \(\text{KCl}(aq) + \text{AgNO}_3(aq) \rightarrow\)

5) \(\text{FeCl}_3(aq) + \text{NaOH}(aq) \rightarrow\)

6) \(\text{MgCl}_2(aq) + \text{Li}_3\text{PO}_4(aq) \rightarrow\)
5.2 Acids, Bases, and Exchange Reactions (p. 172)

An acid is a compound that:

1. ______________________ the effects of a base.
2. increases the concentration of the _____________ ion (____) when dissolved in water.
3. can donate a ________________ (i.e., hydrogen ion, H\(^{+}\)).

The hydrogen cation, H\(^{+}\), is a hydrogen atom that has lost its one electron. As such, the hydrogen ion is simply a proton. However, the hydrogen ion cannot exist in water and in aqueous solution, the hydrogen ion combines with water to form the hydronium ion.

The formula for the hydronium ion is ________.

The equation showing the reaction between the hydrogen ion and water can be represented as follows.

\[ \text{H}^{+} + \text{H}_2\text{O} \longrightarrow \]

\[ \text{H}^{+} \approx \text{________} \]

For convenience, the hydronium ion, H\(_3\)O\(^{+}\), will be indicated by H\(^{+}\).

**Acid formulas are most often written with the donatable proton at the beginning of the formula.** HCl; H\(_2\)SO\(_3\); HC\(_7\)H\(_5\)O\(_2\)

Acids can be categorized base upon (1) the number of donatable protons, (2) the presence or absence of oxygen, and (3) strength.

1) In terms of the number of donatable protons, acids can be monoprotic, diprotic, or even triprotic.

<table>
<thead>
<tr>
<th></th>
<th>HCl</th>
<th>H(_2)SO(_4)</th>
<th>H(_3)PO(_4)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2) Acids which contain oxygen are referred to as oxoacids (or oxyacids). This term stands in contradistinction to ‘hydracids’ (e.g. HCl) lacking oxygen. More strictly, the term oxoacid designates a compound which contains oxygen, at least one other element, and at least one hydrogen bound to oxygen, and which produces a conjugate base by loss of one or more hydrogen ions.

<table>
<thead>
<tr>
<th></th>
<th>H(_2)SO(_4)</th>
<th>HCl</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
3) Based upon strength, acids can be divided into two broad groups, strong or weak. This designation is dependent upon the degree to which an acid dissociates aqueous solution.

**Strong acids** completely dissociate (100% dissociation) in solution and are therefore strong electrolytes. There are only six strong acids. An example of a strong acid is hydrochloric acid, HCl.

\[
\text{HCl(aq)} \rightarrow \nabla
\]

The six strong acids, Table 5.2, p. 174

<table>
<thead>
<tr>
<th>Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
</tr>
<tr>
<td>Hydroiodic acid</td>
</tr>
<tr>
<td>Nitric acid</td>
</tr>
<tr>
<td>Sulfuric acid</td>
</tr>
<tr>
<td>Perchloric acid</td>
</tr>
</tbody>
</table>

**Weak acids** dissociate very little (i.e., exist primarily in the form of a molecular species) and are therefore weak electrolytes. The vast majority of acids are weak acids. An example of a weak acid is acetic acid, H\(_2\)C\(_2\)O\(_2\). A 1.0 M acetic acid is only about 0.5% dissociated.

\[
\text{H}\text{C}_2\text{H}_3\text{O}_2\text{(aq)} \rightleftharpoons \nabla
\]

Sulfuric acid is a diprotic acid and is a strong acid in its first ionization and a weak acid in its second ionization.
- First ionization: \(\text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{HSO}_4^-\text{(aq)}\)
- Second ionization: \(\text{HSO}_4^-\text{(aq)} \rightleftharpoons \text{H}^+(\text{aq}) + \text{SO}_4^{2-}\text{(aq)}\)

Phosphoric acid is triprotic acid and is weak in all three of its ionization steps.
- First ionization: \(\text{H}_3\text{PO}_4\text{(aq)} \rightleftharpoons \text{H}^+\text{(aq)} + \text{H}_2\text{PO}_4^-\text{(aq)}\)
- Second ionization: \(\text{H}_2\text{PO}_4^-\text{(aq)} \rightleftharpoons \text{H}^+\text{(aq)} + \text{HPO}_4^{2-}\text{(aq)}\)
- Third ionization: \(\text{HPO}_4^{2-}\text{(aq)} \rightleftharpoons \text{H}^+\text{(aq)} + \text{PO}_4^{3-}\text{(aq)}\)

**Show the dissociation process for the following acids.**

1) \(\text{HNO}_3\text{(aq)}\)

2) \(\text{H}_2\text{SO}_4\text{ (aq)}\)

3) \(\text{HClO}_4\text{(aq)}\)

4) \(\text{H}_2\text{C}_2\text{O}_4\text{(aq)}\)
A base is a compound that:

1. ______________________ the effects of an acid.

2. increases the concentration of the ________________ ion (____) when dissolved in water.

3. can ____________ a proton (i.e., hydrogen ion, H\(^{+}\)).

Base formulas are most often written with the hydroxide ion at the end of the formula.
NaOH; Ca(OH)\(_2\); Al(OH)\(_3\)

Bases can be either strong or weak. Strong bases are completely dissociated in solution and are therefore strong electrolytes. Most strong bases are inorganic.

$$\text{NaOH(aq)} \longrightarrow$$

The six strong bases, Table 5.2, p. 174

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium Hydroxide</td>
<td></td>
</tr>
<tr>
<td>Sodium Hydroxide</td>
<td></td>
</tr>
<tr>
<td>Potassium Hydroxide</td>
<td></td>
</tr>
<tr>
<td>Calcium Hydroxide</td>
<td></td>
</tr>
<tr>
<td>Strontium Hydroxide</td>
<td></td>
</tr>
<tr>
<td>Barium Hydroxide</td>
<td></td>
</tr>
</tbody>
</table>

Some bases such as Ca(OH)\(_2\) are not very soluble in water, but the part that does dissolve completely dissociates. Ca(OH)\(_2\) can be considered to be a strong electrolyte although it is not very soluble.

Ammonia is a weak base. Gaseous ammonia dissolves in water and then the dissolved ammonia reacts with water to produce hydroxide ions, but most remains as molecules. Ammonia is less than 5% ionized.

$$\text{NH}_3(g) + \text{H}_2\text{O}(l) \longrightarrow \text{NH}_3(\text{aq})$$

$$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$$

Amines are organic derivatives of ammonia and like ammonia are weak bases.

$$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{NH}_3^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$$
Acid-Base Reactions: Neutralization (p. 176)

A **neutralization rxn** is the reaction of an acid and a base to produce a _______.

A **salt** is a compound:
1. composed of metal cations or ammonium ion \((\text{NH}_4^1+)\), and nonmetal anions or polyatomic anions other than hydroxide.
2. formed when an acid and a base react. \text{ACID} + \text{BASE} \rightarrow \text{SALT}

Examples: NaCl; MgBr\(_2\); K\(_2\)SO\(_4\); NH\(_4\)NO\(_3\)

In a neutralization rxn, the identifying characteristics of the acid and base cancel out, or in other words, are neutralized by each other. If the base is a hydroxide base (e.g., NaOH, Ca(OH)\(_2\)) water is also formed along with the salt.

\[
\text{ACID} + \text{BASE} \rightarrow \text{SALT} + \text{WATER}
\]

**Complete and balance the following neutralization reactions.**

1) \(\text{HClO}_4(aq) + \text{LiOH}(aq) \rightarrow\)

2) \(\text{H}_2\text{SO}_4(aq) + \text{NaOH}(aq) \rightarrow\)

3) \(\text{KOH}(aq) + \text{HI}(aq) + \rightarrow\)

4) \(\text{Ca(OH)}_2(aq) + \text{HBr}(aq) \rightarrow\)

5) \(\text{H}_3\text{PO}_4(aq) + \text{Ca(OH)}_2(aq) \rightarrow\)

6) \(\text{NH}_3(aq) + \text{HCl}(aq) \rightarrow\)

7) \(\text{H}_2\text{SO}_4(aq) + \text{NH}_3(aq) \rightarrow\)
Neutralization reactions can be treated in the same manner as precipitation reaction and be written as balanced molecular equations, complete ionic equations, and net ionic equations.

**Balanced Molecular Equation**
\[ \text{HClO}_4(aq) + \text{LiOH}(aq) \rightarrow \text{LiClO}_4(aq) + \text{H}_2\text{O}(l) \]

**Complete Ionic Equation**
\[ \text{H}^+ (aq) + \text{ClO}_4^{-1}(aq) + \text{Li}^{+}(aq) + \text{OH}^{-1}(aq) \rightarrow \text{Li}^{+}(aq) + \text{ClO}_4^{-1}(aq) + \text{H}_2\text{O}(l) \]

**Net Ionic Equation**
\[ \text{H}^+ (aq) + \text{OH}^{-1}(aq) \rightarrow \text{H}_2\text{O}(l) \]

The net ionic equation for the reaction between most strong acids and most strong bases are identical.

In the net ionic equation of a weak acid and a strong base, the acid is written in its undissociated form.

**Balanced Molecular Equation**
\[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaOH}(aq) \rightarrow \text{Na} \text{C}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{O}(l) \]

**Complete Ionic Equation**
\[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{Na}^{+}(aq) + \text{OH}^{-1}(aq) \rightarrow \text{Na}^{+} + \text{C}_2\text{H}_3\text{O}_2^{-1}(aq) + \text{H}_2\text{O}(l) \]

**Net Ionic Equation**
\[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{OH}^{-1}(aq) \rightarrow \text{C}_2\text{H}_3\text{O}_2^{-1}(aq) + \text{H}_2\text{O}(l) \]

In the net ionic eqn of a strong acid and a weak base, the base is written in its undissociated form.
\[ \text{Mg(OH)}_2(s) + 2 \text{H}^+ (aq) \rightarrow \text{Mg}^{2+}(aq) + 2 \text{H}_2\text{O}(l) \]

The net ionic equation for the reaction of ammonia with hydrochloric acid is as follows.
\[ \text{H}^+ (aq) + \text{NH}_3(aq) \rightarrow \text{NH}_4^{+}(aq) \]
Gas Formation (p. 179)

Certain salts, notably carbonates, sulfites, and sulfides, react with acids to form a gaseous product. These anions first form an unstable acid, which then decomposes to form the gas.

<table>
<thead>
<tr>
<th>Anion</th>
<th>Reaction with H(^{1+})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bicarbonate</td>
<td>(\text{HCO}_3^{-}) + H(^{1+}) --(\rightarrow) (\text{CO}_2(g) + \text{H}_2\text{O(l)})</td>
</tr>
<tr>
<td>Carbonate</td>
<td>(\text{CO}_3^{2-} + 2 \text{H}^{1+} --(\rightarrow) (\text{CO}_2(g) + \text{H}_2\text{O(l)})</td>
</tr>
<tr>
<td>Bisulfite</td>
<td>(\text{HSO}_3^{-} + \text{H}^{1+} --(\rightarrow) (\text{SO}_2(g) + \text{H}_2\text{O(l)})</td>
</tr>
<tr>
<td>Sulfite</td>
<td>(\text{SO}_3^{2-} + 2 \text{H}^{1+} --(\rightarrow) (\text{SO}_2(g) + \text{H}_2\text{O(l)})</td>
</tr>
<tr>
<td>Bisulfide</td>
<td>(\text{HS}^{-} + \text{H}^{1+} --(\rightarrow) (\text{H}_2\text{S(g)})</td>
</tr>
<tr>
<td>Sulfide</td>
<td>(\text{S}^{2-} + 2 \text{H}^{1+} --(\rightarrow) (\text{H}_2\text{S(g)})</td>
</tr>
</tbody>
</table>

**Carbonates**
\[\text{Na}_2\text{CO}_3(aq) + 2 \text{HCl(aq)} ----\(\rightarrow\) 2 \text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2(g)\]

**Sulfites**
\[\text{Na}_2\text{SO}_3(aq) + 2 \text{HCl(aq)} ----\(\rightarrow\) 2 \text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{SO}_2(g)\]

**Sulfides**
\[\text{Na}_2\text{S(aq)} + 2 \text{HCl(aq)} ----\(\rightarrow\) 2 \text{NaCl(aq)} + \text{H}_2\text{S(g)}\]

**Ammonium salts** react with strong bases to yield ammonia gas.
\[\text{NaOH(aq)} + \text{NH}_4\text{Cl(aq)} ----\(\rightarrow\) \text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{NH}_3(g)\]
5.3 Oxidation-Reduction (p. 181)

Among the first reactions studied by early scientists were those that involved oxygen. The removal of oxygen from metal oxides to give the metals in their elemental forms was described as reduction.

Oxidation-reduction reactions are also referred to as **redox reactions**.

**Redox reactions involve the transfer of _________________ from one reactant to another.**

The **reducing agent** is the species that is oxidized. It _________ electrons.

The **oxidizing agent** is the species being reduced. It _________ electrons.

The **oxidizing agent** brings about the reduction of another element and the **reducing agent** brings about the oxidation of another element.

The reduction of tin(IV) oxide to elemental tin by reacting with elemental carbon is an example of a redox reaction.

\[
\text{SnO}_2(s) + C(s) \rightarrow \text{Sn}(s) + 2 \text{CO}(g)
\]

In this reaction carbon brings about the reduction of the tin ore to tin metal so it is referred to as the **reducing agent**.

**Redox reactions can often be identified by presence of an element in its elemental form on one side of the reaction arrow and as a part of a compound on the other side of the reaction arrow.**

In the example cited above, tin is in a compound on the reactant side and exists as the element on the product side. Also, carbon is in its elemental form on the reactant side and on the product side it is in a compound.

The combustion of fuels and the reactions of metals with oxygen to give oxides were described by the word oxidation. The burning of magnesium in air is a redox reaction.  

\[
2 \text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s)
\]

In this instance, oxygen is the oxidizing agent because it brings about the oxidation of the metal.

In \(\text{ZnSO}_4 + \text{Mg} \rightarrow \text{MgSO}_4 + \text{Zn}\), the \(\text{Zn}^{2+}\) cation is reduced so it is the oxidizing agent. The Mg is oxidized so it is the reducing agent.

In \(2 \text{HCl} + \text{Fe} \rightarrow \text{FeCl}_2 + \text{H}_2\), the \(\text{H}^{1+}\) cation is reduced so it is the oxidizing agent. The Fe is oxidized so it is the reducing agent.
5.4 Oxidation Numbers (p. 185)
The oxidation number is the charge of the atom if it exists as a monatomic ion or a hypothetical charge assigned to the atom in the substance by simple rules.

Rules for Assigning Oxidation Numbers (p. 186)
1. The oxidation number of an atom of a pure element is 0.
2. The oxidation number of a monatomic ion equals its charge.
3. Some elements have the same oxidation number in the majority of compounds they form and can be used as a reference.
   a. Hydrogen is +1 unless combined with a metal, in this case it is –1.
   b. Fluorine is -1.
   c. Oxygen is –2, except in peroxides it is –1.
   d. In binary compounds with metals, Group 7A (halogens) elements have an oxidation number of –1.
   e. In binary compounds, Group 6A elements have an oxidation number of –2
   f. In binary compounds, Group 5A elements have an oxidation number of -3.
   g. Group 1A metal cations all have an oxidation number of +1
   h. Group 2A metal cations all have an oxidation number of +2.
4. For the atoms in a neutral species (an isolated atom, a molecule, or a formula unit) the total of all the oxidation numbers is 0; the sum of the oxidation numbers in a polyatomic ion equals the charge on the ion.

Oxidation Number Problem #1
What is the oxidation number of nitrogen in ammonia? [-3]

Oxidation Number Problem #2
What is the oxidation number of nitrogen in sodium nitrate, NaNO₃? [+5]

Oxidation Number Problem #3
What is the oxidation number of nitrogen in sodium nitrite? [+3]

Oxidation Number Problem #4
What is the oxidation number of chlorine in calcium hypochlorite, Ca(ClO)₂? [+1]

Oxidation Number Problem #5
What is the oxidation number of carbon in carbon dioxide? [+4]

Oxidation Number Problem #6
What is the oxidation number of carbon in methane? [-4]
Redox reactions can also be identified by a change in oxidation number. If the oxidation number of a given atom changes, then a redox reaction has occurred.

<table>
<thead>
<tr>
<th>OXIDATION</th>
<th>REDUCTION</th>
<th>&quot;OIL RIG&quot; acronym</th>
</tr>
</thead>
<tbody>
<tr>
<td>Involves the ______ of electrons.</td>
<td>Involves the _________ of electrons.</td>
<td>Oxidation Involves Loss of electrons.</td>
</tr>
<tr>
<td>Is characterized by an _____________ in oxidation number.</td>
<td>Is characterized by a _____________ in oxidation number.</td>
<td>Reduction Involves Gain of electrons.</td>
</tr>
</tbody>
</table>

Assign oxidation numbers to each atom in the following reactions.

1) \( \text{ZnSO}_4(\text{aq}) + \text{Mg}(s) \rightarrow \text{MgSO}_4(\text{aq}) + \text{Zn}(s) \)

2) \( 2 \text{ HCl}(\text{aq}) + \text{Fe}(s) \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2(\text{g}) \)

3) \( \text{CH}_4(\text{g}) + 2 \text{ O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{ H}_2\text{O}(\text{g}) \)
5.6 Solution Concentration (p. 192)
Molarity is defined as moles of solute per liter of solution.
\[ M = \frac{\text{mole of solute}}{\text{liter of solution}} \]

**Molarity Problem #1**
If 400.0 mL of a solution contains \(5.00 \times 10^{-3}\) moles of AgNO₃, what is the molarity of this solution? [0.0125 M AgNO₃]

**Molarity Problem #2**
Calculate the molarity of an HCl solution containing 18.23 g of HCl in 355.0 mL of solution. [1.408 M HCl]

**Molarity Problem #3**
How many moles of NaCl are contained in a 27.49 mL sample of a 0.350 M solution of NaCl? [9.62 \times 10^{-3} \text{ mol NaCl}]

**Molarity Problem #4**
What volume of 0.274 M NaOH is needed to provide a sample of solution that contains 0.620 grams of NaOH? [56.57 mL]

Dilution of Solutions (p. 195)
\[ M_1V_1 = M_2V_2 \]

**Dilution Problem #1**
What volume of 18.0 M H₂SO₄ must is required to make 100 mL of a 5.0 M solution of H₂SO₄? (28 mL)

**Dilution Problem #2**
How many milliliters of water are required to dilute 6 M H₂SO₄ to 175 ml of 0.3 M H₂SO₄? (166 mL)
5.7 Molarity and Reactions in Aqueous Solutions (p. 198)

Mass or Volume of A $\leftrightarrow$ moles A $\leftrightarrow$ moles of B $\leftrightarrow$ Volume or Mass of B

Solution Stoichiometry Problem #1
How many mL of 6.0 M HCl are needed to digest 3.78 g of Mg? [52 mL]

\[ \text{Mg}(s) + 2 \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g) \]

Solution Stoichiometry Problem #2
What mass of calcium carbonate will react with 25.00 mL of 6.0 M HCl?

\[ \text{CaCO}_3(s) + 2 \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \]

Solution Stoichiometry Problem #3
How many mL of 0.0487 M Ba(OH)$_2$ are needed to react with 35.67 mL of 0.0748 M HCl?

Solution Stoichiometry Problem #4
What volume of 0.591 M silver nitrate is needed to completely react with 50.00 mL of 0.0748 M NaCl?

5.8 Aqueous Solution Titrations (p. 199)

A ________________ is as an analytical procedure of determining the concentration of one substance in solution by reacting it with a solution of another substance whose concentration is known, called a titrant (or standard solution). To carry out the process, we add the titrant, using a buret, to a known volume of the other solution until the reaction between the two substances is just complete, that is until chemically equivalent amounts of the two reactants are present. We often tell when we have reached this point, called the endpoint or equivalence point, by a change in color of an added chemical substance called an indicator. The volume of the titrant is used to calculate the number of moles delivered from the buret. The moles of the other reactant are obtained from the coefficients in the balanced reaction equation, and this value is used with the volume of the solution of unknown concentration to calculate its molarity.

Volume of A $\leftrightarrow$ moles A $\leftrightarrow$ moles of B $\leftrightarrow$ Molarity of B

Titration Problem #1
In a titration 25.00 mL of sodium hydroxide solution was neutralized by 32.72 mL of hydrochloric acid. The HCl had a concentration of 0.129 M. Find the concentration of the NaOH solution. [0.169 M NaOH]

Titration Problem #2
During a titration, a 20.00 mL portion of a 0.100 M H$_2$SO$_4$ solution was carefully measured into a flask and phenolphthalein was added to it. The solution was titrated with 18.47 mL of NaOH to reach the endpoint. What is the concentration of the NaOH? [0.217 M NaOH]