Reaction Stoichiometry: How Much Carbon Dioxide?

- The balanced chemical equations for fossil-fuel combustion reactions provide the exact relationships between the amount of fossil fuel burned and the amount of carbon dioxide emitted.

$$2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g)$$

- 16 CO₂ molecules are produced for every 2 molecules of octane burned.
Quantities in Chemical Reactions

- The amount of every substance used and made in a chemical reaction is related to the amounts of all the other substances in the reaction.
  - Law of conservation of mass
  - Balancing equations by balancing atoms
- The study of the numerical relationship between chemical quantities in a chemical reaction is called **stoichiometry**.

Reaction Stoichiometry

- The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.
  
  \[ 2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]
  - 2 molecules of C\textsubscript{8}H\textsubscript{18} react with 25 molecules of O\textsubscript{2} to form 16 molecules of CO\textsubscript{2} and 18 molecules of H\textsubscript{2}O.
  - 2 moles of C\textsubscript{8}H\textsubscript{18} react with 25 moles of O\textsubscript{2} to form 16 moles of CO\textsubscript{2} and 18 moles of H\textsubscript{2}O.

  \[ 2 \text{ mol C}_8\text{H}_{18} : 25 \text{ mol O}_2 : 16 \text{ mol CO}_2 : 18 \text{ mol H}_2\text{O} \]
Making Pizza

• The number of pizzas you can make depends on the amount of the ingredients you use.

1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

This relationship can be expressed mathematically.

1 crust : 5 oz sauce : 2 cups cheese : 1 pizza

• We can compare the amount of pizza that can be made from 10 cups of cheese:

2 cups cheese : 1 pizza, then,

$$\frac{10 \text{ cups cheese}}{2 \text{ cups cheese}} \times \frac{1 \text{ pizza}}{2 \text{ cups cheese}} = 5 \text{ pizzas}$$

Making Molecules: Mole-to-Mole Conversions

• We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe.

The ratio of the coefficients acts as a conversion factor between the amount in moles of the reactants and products.

$$2 \text{ C}_8\text{H}_{18}(l) + 25 \text{ O}_2(g) \rightarrow 16 \text{ CO}_2(g) + 18 \text{ H}_2\text{O}(g)$$

stoichiometric ratio: 2 moles C\textsubscript{8}H\textsubscript{18} : 16 moles CO\textsubscript{2}

The ratio acts as a conversion factor between the amount in moles of the reactant, C\textsubscript{8}H\textsubscript{18}, and the amount in moles of the product, CO\textsubscript{2}. 
Mole-to-Mole Conversions

• Suppose we burn 22.0 moles of C₈H₁₈; how many moles of CO₂ form?

\[ 2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]

stoichiometric ratio: 2 moles C₈H₁₈ : 16 moles CO₂

\[
22.0 \text{ mol C}_8\text{H}_{18} \times \frac{16 \text{ mol CO}_2}{2 \text{ mol C}_8\text{H}_{18}} = 176 \text{ mol CO}_2
\]

• The combustion of 22.0 moles of C₈H₁₈ adds 176 moles of CO₂ to the atmosphere.

Making Molecules: Mass-to-Mass Conversions

• The world burned the equivalent of 3.7×10¹⁵ g of gasoline (octane) in 2013. We can estimate the mass of CO₂ produced based on the flow chart below.

- We use molar mass as a conversion factor between the mass given and amount in moles.
- We use coefficients as the conversion factor between the reactant, C₈H₁₈, and the amount in moles of the product, CO₂, and then molar mass as the conversion factor to get the mass of CO₂ produced.
Mass-to-Mass Conversions

- If we burn $3.7 \times 10^{15}$ g C$_8$H$_{18}$, how many grams of CO$_2$ form?

$$2\text{ C}_8\text{H}_{18}(l) + 25\text{ O}_2(g) \rightarrow 16\text{ CO}_2(g) + 18\text{ H}_2\text{O}(g)$$

**molar mass:** C$_8$H$_{18}$ 114.22 g/mol, CO$_2$ 44.01 g/mol

**stoichiometric ratio:** 2 moles C$_8$H$_{18}$ : 16 moles CO$_2$

$$22.0 \text{ mol C}_8\text{H}_{18} \times \frac{16 \text{ mol CO}_2}{2 \text{ mol C}_8\text{H}_{18}} = 176 \text{ mol CO}_2$$

- The combustion $3.7 \times 10^{15}$ g C$_8$H$_{18}$ adds $1.1 \times 10^{16}$ g CO$_2$ to the atmosphere.

---

Limiting Reactant, Theoretical Yield, Percent Yield

- Recall our pizza recipe:

  1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

- If we have 4 crusts, 10 cups of cheese, and 15 oz tomato sauce, how many pizzas can we make?

**We have enough crusts to make**

$$4 \text{ crusts} \times \frac{1 \text{ pizza}}{1 \text{ crust}} = 4 \text{ pizzas}$$

**We have enough cheese to make**

$$10 \text{ cups cheese} \times \frac{1 \text{ pizza}}{2 \text{ cups cheese}} = 5 \text{ pizzas}$$

**We have enough tomato sauce to make**

$$15 \text{ ounces tomato sauce} \times \frac{1 \text{ pizza}}{5 \text{ ounces tomato sauce}} = 3 \text{ pizzas}$$

**Limiting reactant**

**Smallest number of pizzas**
Limiting Reactant

- We have enough crusts for four pizzas, enough cheese for five pizzas, but enough tomato sauce for only three pizzas.
  - We can make only three pizzas. The tomato sauce *limits* how many pizzas we can make.

Theoretical Yield

- Tomato sauce is the **limiting reactant**, the reactant that makes *the least amount of product*.
  - The limiting reactant is also known as the *limiting reagent*.
- The maximum number of pizzas we can make depends on this ingredient. In chemical reactions, we call this the **theoretical yield**.
  - This is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
  - The ingredient that makes the least amount of pizza determines how many pizzas you can make (*theoretical yield*).
Percent Yield

Assume that while making pizzas, we burn a pizza, drop one on the floor, or other uncontrollable events happen so that we make only two pizzas. The actual amount of product made in a chemical reaction is called the actual yield.

We can determine the efficiency of making pizzas by calculating the percentage of the maximum number of pizzas we actually make. In chemical reactions, we call this the percent yield.

\[
\text{% yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = 67\%
\]

In a Chemical Reaction

- For reactions with multiple reactants, it is likely that one of the reactants will be completely used before the others.
- When this reactant is used up, the reaction stops and no more product is made.
- The reactant that limits the amount of product is called the limiting reactant.
  - It is sometimes called the limiting reagent.
  - The limiting reactant gets completely consumed.
- Reactants not completely consumed are called excess reactants.
- The amount of product that can be made from the limiting reactant is called the theoretical yield.
### Summarizing Limiting Reactant and Yield

- **Limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.

- **Reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.

<table>
<thead>
<tr>
<th>Summarizing Limiting Reactant and Yield</th>
</tr>
</thead>
<tbody>
<tr>
<td>• The <strong>theoretical yield</strong> is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.</td>
</tr>
<tr>
<td>• The <strong>actual yield</strong> is the amount of product actually produced by a chemical reaction.</td>
</tr>
<tr>
<td>• The <strong>percent yield</strong> is calculated as follows:</td>
</tr>
</tbody>
</table>
| \[
| \text{percent yield} = \left( \frac{\text{actual yield}}{\text{theoretical yield}} \right) \times 100\% |
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Recall our balanced equation for the combustion of methane:

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

– Our balanced equation for the combustion of methane implies that every one molecule of CH\(_4\) reacts with two molecules of O\(_2\).

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) \]

If we have five molecules of CH\(_4\) and eight molecules of O\(_2\), which is the limiting reactant?

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

– First we calculate the number of CO\(_2\) molecules that can be made from five CH\(_4\) molecules.

\[ 5 \text{ CH}_4 \times \frac{1 \text{ CO}_2}{1 \text{ CH}_4} = 5 \text{ CO}_2 \]
Combustion of Methane

- Then we calculate the number of CO\(_2\) molecules that can be made from eight O\(_2\) molecules.

\[
8 \text{O}_2 \times \frac{1 \text{CO}_2}{2 \text{O}_2} = 4 \text{CO}_2
\]

- We have enough CH\(_4\) to make five CO\(_2\) molecules and four CO\(_2\) molecules.
- Therefore, O\(_2\) is the limiting reactant, and four CO\(_2\) molecules is the theoretical yield.
- CH\(_4\) is in excess.

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

- When working in the lab, we normally measure reactant quantities in grams.
- To find the limiting reactant and theoretical yield, we must first convert grams to moles.
A reactant mixture contains 42.5 g Mg and 33.8 g O\(_2\). What is the limiting reactant and theoretical yield?

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

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses
When table salt is mixed with water, it seems to disappear or become a liquid; the mixture is homogeneous.
  – The salt is still there, as you can tell from the taste or simply boiling away the water.

Homogeneous mixtures are called **solutions**.

The majority component is the **solvent**.

The minority component is the **solute**.

A solution in which water is the solvent is an **aqueous solution**.

Because solutions are mixtures, the composition can vary from one sample to another.
  – Pure substances have constant composition.
  – Saltwater samples from different seas or lakes have different amounts of salt.

So, to describe solutions accurately, we quantify the amount of solute relative to solvent, or **concentration of solution**.
Solution Concentration

• Solutions are often described quantitatively, as dilute or concentrated.

• **Dilute solutions** have a small amount of solute compared to solvents.

• **Concentrated solutions** have a large amount of solute compared to solvents.

Solution Concentration: Molarity

• A common way to express solution concentration is **molarity** \( (M) \).
  
  – Molarity is the amount of solute (in moles) divided by the volume of solution (in liters).

\[
\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}
\]
Preparing 1 L of a 1.00 M NaCl Solution

Preparing a Solution of Specified Concentration

• We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution.
  – For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.

Using Molarity in Calculations

\[
\frac{0.500 \text{ mol NaCl}}{\text{L solution}} \quad \text{converts} \quad \text{L solution} \quad \text{mol NaCl}
\]

\[
\frac{\text{L solution}}{0.500 \text{ mol NaCl}} \quad \text{converts} \quad \text{mol NaCl} \quad \text{L solution}
\]
**Solution Dilution**

- Often, solutions are stored as concentrated **stock solutions**.
- To make solutions of lower concentrations from these stock solutions, more solvent is added.
  - The amount of solute doesn’t change, just the volume of solution:
    
    moles solute in solution 1 = moles solute in solution 2

- The concentrations and volumes of the stock and new solutions are inversely proportional:
  
  \[ M_1 \cdot V_1 = M_2 \cdot V_2 \]

**Preparing 3.00 L of 0.500 M CaCl\(_2\) from a 10.0 M Stock Solution**

\[ \frac{10.0 \text{ mol}}{L} \times 0.150 \text{ L} = 1.50 \text{ mol} \times 3.00 \text{ L} \]

\[ 1.50 \text{ mol} = 1.50 \text{ mol} \]
Solution Stoichiometry

- Because molarity relates the moles of solute to the liters of solution, it can be used to convert between amount of reactants and/or products in a chemical reaction.
  - The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product.

Types of Aqueous Solutions and Solubility

- Consider two familiar aqueous solutions: saltwater and sugar water.
  - Saltwater is a homogeneous mixture of NaCl and H₂O.
  - Sugar water is a homogeneous mixture of C₁₂H₂₂O₁₁ and H₂O.
- As you stir either of these two substances into the water, it seems to disappear.
  - How do solids such as salt and sugar dissolve in water?
What Happens When a Solute Dissolves?

- There are attractive forces between the solute particles holding them together.
- There are also attractive forces between the solvent molecules.
- When we mix the solute with the solvent, there are attractive forces between the solute particles and the solvent molecules.
- If the attractions between solute and solvent are strong enough, the solute will dissolve.

Charge Distribution in a Water Molecule

- There is an uneven distribution of electrons within the water molecule.
  - This causes the oxygen side of the molecule to have a partial negative charge ($\delta^-$) and the hydrogen side to have a partial positive charge ($\delta^+$).
Solute and Solvent Interactions in a Sodium Chloride Solution

- When sodium chloride is put into water, the attraction of Na\(^+\) and Cl\(^-\) ions to water molecules competes with the attraction among the oppositely charged ions themselves.

Dissolution of Ionic Compounds

- Each ion is attracted to the surrounding water molecules and pulled off and away from the crystal.

- When it enters the solution, the ion is surrounded by water molecules, insulating it from other ions.

- The result is a solution with free-moving, charged particles able to conduct electricity.
Electrolyte and Nonelectrolyte Solutions

- Materials that dissolve in water to form a solution containing ions will conduct electricity. These are called electrolytes.
- Materials that dissolve in water to form a solution with no ions will not conduct electricity. These are called nonelectrolytes.
- A solution of salt (an electrolyte) conducts electrical current. A solution of sugar (a nonelectrolyte) does not.

Electrolyte and Nonelectrolyte Solutions

- Ionic substances, such as sodium chloride, that completely dissociate into ions when they dissolve in water, are strong electrolytes.
- Except for acids, most molecular compounds, for example sugar, dissolve in water as intact molecules, or nonelectrolytes.
- Acids ionize to varying degrees in water. Those that completely ionize are strong acids. Those that don’t are weak acids.
Sugar Dissolution in Water

Acids

• Acids are molecular compounds that **ionize** when they dissolve in water.
  – The molecules are pulled apart by their attraction for the water.
  – When acids ionize, they form $H^+$ cations and also anions.
• The percentage of molecules that ionize varies from one acid to another.
• Acids that ionize virtually 100% are called **strong acids**.
  \[ \text{HCl}(aq) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) \]
• Acids that only ionize a small percentage are called **weak acids**.
  \[ \text{HF}(aq) \rightleftharpoons \text{H}^+(aq) + \text{F}^-(aq) \]
## Strong and Weak Electrolytes

- **Strong electrolytes** are materials that dissolve completely as ions.
  - Ionic compounds and strong acids.
  - Solutions are good conductors of electricity.
- **Weak electrolytes** are materials that dissolve mostly as molecules but partially as ions.
  - Weak acids.
  - Solutions conduct electricity, but not well.

## Dissociation and Ionization

- When ionic compounds dissolve in water, the anions and cations are separated from each other. This is called **dissociation**.
  \[ \text{Na}_2\text{S}(aq) \rightarrow 2 \text{Na}^+(aq) + \text{S}^{2-}(aq) \]
- When compounds containing polyatomic ions dissociate, the polyatomic group stays together as one ion.
  \[ \text{Na}_2\text{SO}_4(aq) \rightarrow 2 \text{Na}^+(aq) + \text{SO}_4^{2-}(aq) \]
- When strong acids dissolve in water, the molecule **ionizes** into $\text{H}^+$ and anions.
  \[ \text{H}_2\text{SO}_4(aq) \rightarrow 2 \text{H}^+(aq) + \text{SO}_4^{2-}(aq) \]
When an ionic compound dissolves in water, the resulting solution contains not the intact ionic compound itself but its component ions dissolved in water.

However, not all ionic compounds dissolve in water. For example, AgCl remains solid and appears as a white powder at the bottom of the water.

In general, a compound is termed **soluble** if it dissolves in water and **insoluble** if it does not.
Solubility of Salts

• If we mix solid AgNO\(_3\) with water, it dissolves and forms a strong electrolyte solution.

• Silver chloride, on the other hand, is almost completely insoluble.
  – If we mix solid AgCl with water, virtually all of it remains as a solid within the liquid water.

When Will a Salt Dissolve?

• Whether a particular compound is soluble or insoluble depends on several factors.

• Predicting whether a compound will dissolve in water is not easy.

• The best way to do it is to conduct experiments to test whether a compound will dissolve in water, and then develop some rules based on those experimental results.
  – We call this method the empirical method.
Solubility Rules

**TABLE 4.1 Solubility Rules for Ionic Compounds in Water**

<table>
<thead>
<tr>
<th>Compounds Containing the Following Ions Are Generally Soluble</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺, Na⁺, K⁺, and NH₄⁺</td>
<td>None</td>
</tr>
<tr>
<td>NO₃⁻ and C₂H₅O₂⁻</td>
<td>None</td>
</tr>
<tr>
<td>Cl⁻, Br⁻, and I⁻</td>
<td>When these ions pair with Ag⁺, Hg₂⁺, or Pb²⁺, the resulting compounds are insoluble.</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>When SO₄²⁻ pairs with Sr²⁺, Ba²⁺, Pb²⁺, Ag⁺, or Ca²⁺, the resulting compound is insoluble.</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Compounds Containing the Following Ions Are Generally Insoluble</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>OH⁻ and S²⁻</td>
<td>When these ions pair with Li⁺, Na⁺, K⁺, or NH₄⁺, the resulting compounds are soluble.</td>
</tr>
<tr>
<td></td>
<td>When S²⁻ pairs with Ca²⁺, Sr²⁺, or Ba²⁺, the resulting compound is soluble.</td>
</tr>
<tr>
<td></td>
<td>When OH⁻ pairs with Ca²⁺, Sr²⁺, or Ba²⁺, the resulting compound is slightly soluble.</td>
</tr>
<tr>
<td>CO₃²⁻ and PO₄³⁻</td>
<td>When these ions pair with Li⁺, Na⁺, K⁺, or NH₄⁺, the resulting compounds are soluble.</td>
</tr>
</tbody>
</table>

Precipitation Reactions

- **Precipitation reactions** are reactions in which a solid forms when we mix two solutions.
  - Reactions between aqueous solutions of ionic compounds produce an ionic compound that is insoluble in water.
  - The insoluble product is called a **precipitate**.
No Precipitation Means No Reaction

- Precipitation reactions do not always occur when two aqueous solutions are mixed.
  - Nothing happens when combining solutions of KI and NaCl.

  \[ \text{KI}(aq) + \text{NaCl}(aq) \rightarrow \text{No Reaction} \]
Predicting Precipitation Reactions

1. Determine what ions each aqueous reactant has.
2. Determine formulas of possible products.
   - Exchange ions.
     • (+) ion from one reactant with (–) ion from other
   - Balance charges of combined ions to get the formula of each product.
3. Determine solubility of each product in water.
   - Use the solubility rules.
   - If the product is insoluble or slightly soluble, it will precipitate.
4. If neither product will precipitate, write no reaction after the arrow.

Predicting Precipitation Reactions

5. If any of the possible products are insoluble, write their formulas as the products of the reaction using (s) after the formula to indicate solid. Write any soluble products with (aq) after the formula to indicate aqueous.
6. Balance the equation.
   - Remember to change only coefficients, not subscripts.
Predicting Precipitation Reactions

Representing Aqueous Reactions

- An equation showing the complete neutral formulas for each compound in the aqueous reaction as if they existed as molecules is called a **molecular equation**.

\[ 2 \text{KOH}(aq) + \text{Mg(NO}_3\text{)}_2(aq) \rightarrow 2 \text{KNO}_3(aq) + \text{Mg(OH)}_2(s) \]

- In actual solutions of soluble ionic compounds, dissolved substances are present as ions. Equations that describe the nature of the dissolved species in solution are called **complete ionic equations**.
Aqueous strong electrolytes (soluble salts, strong acids, strong bases) are written as ions.

Insoluble substances, weak electrolytes, and nonelectrolytes are written in molecule form.

- Solids, liquids, and gases are not dissolved, hence molecule form

\[
2 K^+_{(aq)} + 2 OH^-_{(aq)} + Mg^{2+}_{(aq)} + 2 NO_3^-_{(aq)} \rightarrow 2 K^+_{(aq)} + 2 NO_3^-_{(aq)} + Mg(OH)_2(s)
\]

Notice that in the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation.

These ions are called **spectator ions** because they do not participate in the reaction (soluble salts, strong acids, and strong bases).
Net Ionic Equation

- An ionic equation in which the spectator ions are removed is called a net ionic equation.

\[
2\text{K}^+(aq) + 2\text{OH}^-(aq) + \text{Mg}^{2+}(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{K}^+(aq) + 2\text{NO}_3^-(aq) + \text{Mg(OH)}_2(s)
\]

- The net ionic equation is

\[
2\text{OH}^-(aq) + \text{Mg}^{2+}(aq) \rightarrow \text{Mg(OH)}_2(s)
\]

Examples

- Write the ionic and net ionic equation for each of the following:

1. \(\text{K}_2\text{SO}_4(aq) + 2\text{AgNO}_3(aq) \rightarrow 2\text{KNO}_3(aq) + \text{Ag}_2\text{SO}_4(s)\)

   \[
   2\text{K}^+(aq) + \text{SO}_4^{2-}(aq) + 2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{K}^+(aq) + 2\text{NO}_3^-(aq) + \text{Ag}_2\text{SO}_4(s)
   \]

   \[
   2\text{Ag}^+(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{Ag}_2\text{SO}_4(s)
   \]

2. \(\text{Na}_2\text{CO}_3(aq) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)\)

   \[
   2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq) + 2\text{H}^+(aq) + 2\text{Cl}^-(aq) \rightarrow 2\text{Na}^+(aq) + 2\text{Cl}^-(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]

   \[
   \text{CO}_3^{2-}(aq) + 2\text{H}^+(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]
Two other important classes of reactions that occur in aqueous solution are 1. acid–base reactions and 2. gas-evolution reactions.

Acid–base reaction:

– An acid–base reaction is also called a neutralization reaction.
– An acid reacts with a base, and the two neutralize each other, producing water (or in some cases a weak electrolyte).

In a gas-evolution reaction, a gas is produced, resulting in bubbling.

In both acid–base and gas-evolution reactions, as in precipitation reactions, the reactions occur when the anion from one reactant combines with the cation of the other.

Many gas-evolution reactions are also acid–base reactions.
**Arrhenius Definitions:**

- **Acid:** Substance that produces $H^+$ in aqueous solution. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
- In solution, $H^+$ bonds with water to produce the **hydronium ion**, $H_3O^+$.
- **Polyprotic acids** contain more than one ionizable proton and release them sequentially.
- The first ionizable proton is strong while subsequent ionizable protons are weak.
- **Base:** Substance that produces $OH^-$ ions in aqueous solution.
  $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

**Acid–Base Reactions**

- These reactions are called **neutralization reactions** because the acid and base neutralize each other's properties.
  $2 \text{HNO}_3(aq) + \text{Ca(OH)}_2(aq) \rightarrow \text{Ca(NO}_3)_2(aq) + 2 \text{H}_2\text{O}(l)$
- The net ionic equation for an acid–base reaction is
  $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
  - As long as the salt that forms is soluble in water.
Acids and Bases in Solution

- Acids ionize in water to form H\(^+\) ions.
  - More precisely, the H\(^+\) from the acid molecule is donated to a water molecule to form hydronium ion, H\(_3\)O\(^+\).
  - Most chemists use H\(^+\) and H\(_3\)O\(^+\) interchangeably.
- Bases dissociate in water to form OH\(^-\) ions.
  - Bases, such as NH\(_3\), that do not contain OH\(^-\) ions, produce OH\(^-\) by pulling H\(^+\) off water molecules.
- In the reaction of an acid with a base, the H\(^+\) from the acid combines with the OH\(^-\) from the base to make water.
- The cation from the base combines with the anion from the acid to make the salt.

Acid–Base Reaction

- HCl\((aq)\) + NaOH\((aq)\) → NaCl\((aq)\) + H\(_2\)O\((l)\)
Some Common Acids and Bases

<table>
<thead>
<tr>
<th>Name of Acid</th>
<th>Formula</th>
<th>Name of Base</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>Sodium hydroxide</td>
<td>NaOH</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
<td>HBr</td>
<td>Lithium hydroxide</td>
<td>LiOH</td>
</tr>
<tr>
<td>Hydroiodic acid</td>
<td>HI</td>
<td>Potassium hydroxide</td>
<td>KOH</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>HNO₃</td>
<td>Calcium hydroxide</td>
<td>Ca(OH)₂</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H₂SO₄</td>
<td>Barium hydroxide</td>
<td>Ba(OH)₂</td>
</tr>
<tr>
<td>Perchloric acid</td>
<td>HClO₄</td>
<td>Ammonia*</td>
<td>NH₃ (weak base)</td>
</tr>
</tbody>
</table>

Formic acid         | HCHO₂ (weak acid) |
Acetic acid         | CH₃CO₂H (weak acid) |
Hydrofluoric acid   | HF (weak acid) |

*Ammonia does not contain OH⁻, but it produces OH⁻ in a reaction with water that occurs only to a small extent: NH₃(aq) + H₂O(l) ⇄ NH₄⁺(aq) + OH⁻(aq).

Predict the Product of the Reactions

1. HCl(aq) + Ba(OH)₂(aq) →

   (H⁺ + Cl⁻) + (Ba²⁺ + OH⁻) → (H²⁺ + OH⁻) + (Ba²⁺ + Cl⁻)

   HCl(aq) + Ba(OH)₂(aq) → H₂O(l) + BaCl₂

   2 HCl(aq) + Ba(OH)₂(aq) → 2 H₂O(l) + BaCl₂(aq)

2. H₂SO₄(aq) + LiOH(aq) →

   (H⁺ + SO₄²⁻) + (Li⁺ + OH⁻) → (H⁺ + OH⁻) + (Li⁺ + SO₄²⁻)

   H₂SO₄(aq) + LiOH(aq) → H₂O(l) + Li₂SO₄

   H₂SO₄(aq) + 2 LiOH(aq) → 2 H₂O(l) + Li₂SO₄(aq)
**Acid–Base Titrations**

- A **titration** is a laboratory procedure where a substance in a solution of known concentration (titration) is reacted with another substance in a solution of unknown concentration (analyte).
- The **equivalence point** is the point in the titration when the $\text{H}^+$ and $\text{OH}^-$ from reactants are in their stoichiometric ratio and are completely reacted.
- An **indicator** is a dye whose color depends on the acidity or basicity of solution.
**Titration**

In this titration, NaOH is added to a dilute HCl solution. When the NaOH and HCl reach stoichiometric proportions (the equivalence point), the phenolphthalein indicator changes color to pink.

**Gas-Evolving Reactions**

- Some reactions form a gas directly from the ion exchange.
  \[ K_2S(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2S(g) \]

- Other reactions form a gas by the subsequent decomposition of one of the ion exchange products into a gas and water.
  \[ \text{NaHCO}_3(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{CO}_3(aq) \]
  \[ \text{H}_2\text{CO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) \]
Gas-Evolution Reaction

When aqueous sodium bicarbonate is mixed with aqueous hydrochloric acid, gaseous CO₂ bubbles are the result of the reaction.

NaHCO₃(aq) + HCl(aq) → H₂O(l) + NaCl(aq) + CO₂(g)

Types of Compounds That Undergo Gas-Evolution Reactions

<table>
<thead>
<tr>
<th>Reactant Type</th>
<th>Intermediate Product</th>
<th>Gas Evolved</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulfides</td>
<td>None</td>
<td>H₂S(g)</td>
<td>2 HCl(aq) → H₂S(g) + 2 HCl(aq)</td>
</tr>
<tr>
<td>Carbonates and bicarbonates</td>
<td>H₂CO₃</td>
<td>CO₂(g)</td>
<td>2 HCl(aq) → K₂CO₃(aq) + H₂O(l) + CO₂(g) + 2 KCl(aq)</td>
</tr>
<tr>
<td>Sulfites and bisulfites</td>
<td>H₂SO₃</td>
<td>SO₂(g)</td>
<td>2 HCl(aq) → K₂SO₃(aq) + H₂O(l) + SO₂(g) + 2 KCl(aq)</td>
</tr>
<tr>
<td>Ammonium</td>
<td>NH₄OH</td>
<td>NH₃(g)</td>
<td>NH₄Cl(aq) + KOH(aq) → H₂O(l) + NH₃(g) + KOH(aq)</td>
</tr>
</tbody>
</table>