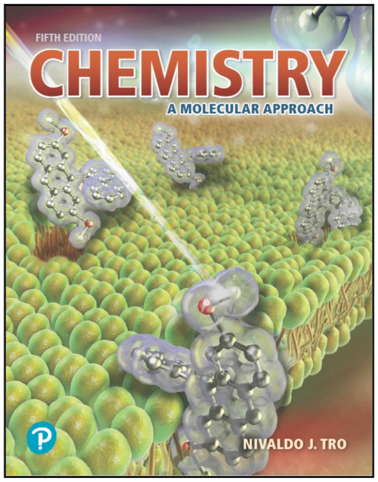
Chemistry: A Molecular Approach

Fifth Edition



Chapter 5

Introduction to Solutions and Aqueous Reactions

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Molecular Gastronomy

- **Molecular gastronomy** is a way of preparing food that involves chemistry.
- A common chemical reaction in molecular gastronomy is **precipitation**.
  - In a precipitation reaction, two solutions—homogeneous mixtures often containing a solid dissolved in a liquid—are mixed.
  - Upon mixing, a solid (or precipitate) forms.
- Chefs use a similar precipitation reaction—called spherification—to encapsulate liquids.

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Solution Concentration (1 of 3)

- Table salt and water form a homogeneous mixture.
- 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**.

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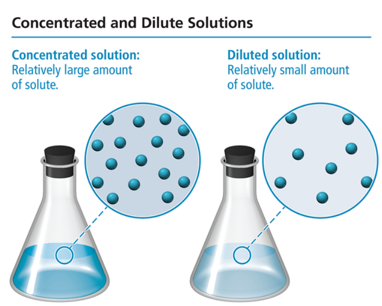
Solution Concentration (2 of 3)

- 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.



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Solution Concentration (3 of 3)

- Because solutions are mixtures, the composition can vary from one sample to another.
- We quantify the amount of solute relative to solvent, or **concentration of solution**.

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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).

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

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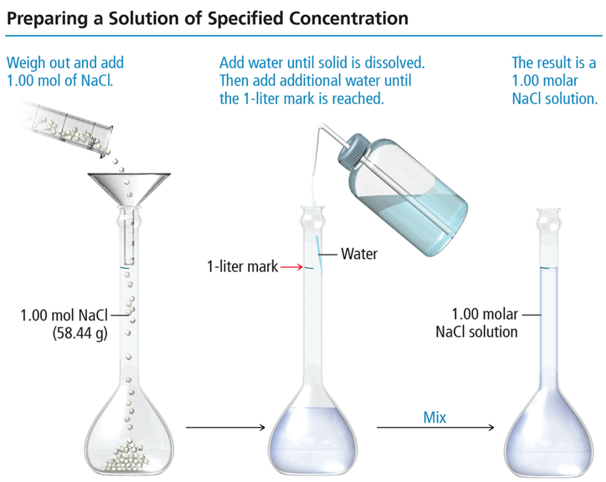
Preparing 1 L of a 1.00 M NaCl Solution



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Conceptual Connection 5.1 (1 of 2)

How many moles of solute are required to make 3.0 L of a 2.0 M solution?

a. 2.0 mol solute

b. 3.0 mol solute

c. 4.0 mol solute

d. 6.0 mol solute

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Conceptual Connection 5.1 (2 of 2)

How many moles of solute are required to make 3.0 L of a 2.0 M solution?

a. 2.0 mol solute

b. 3.0 mol solute

c. 4.0 mol solute

d. 6.0 mol solute

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Using Molarity in Calculations

- 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.

$\dfrac{0.500 \text{ mol NaCl}}{\text{L solution}}$ *converts* **L solution** → **mol NaCl**

$\dfrac{\text{L solution}}{0.500 \text{ mol NaCl}}$ *converts* **mol NaCl** → **L solution**

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Conceptual Connection 5.2 (1 of 2)

If we dissolve 25 g of salt in 251 g of water, what is the mass of the resulting solution?

a. 251 g

b. 276 g

c. 226 g

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Conceptual Connection 5.2 (2 of 2)

If we dissolve 25 g of salt in 251 g of water, what is the mass of the resulting solution?

a. 251 g

b. 276 g

c. 226 g

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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 does not change, just the volume of solution.

$$M_1 \cdot V_1 = M_2 \cdot V_2$$

$$\text{moles solute in solution}_1 = \text{moles solute in solution}_2$$

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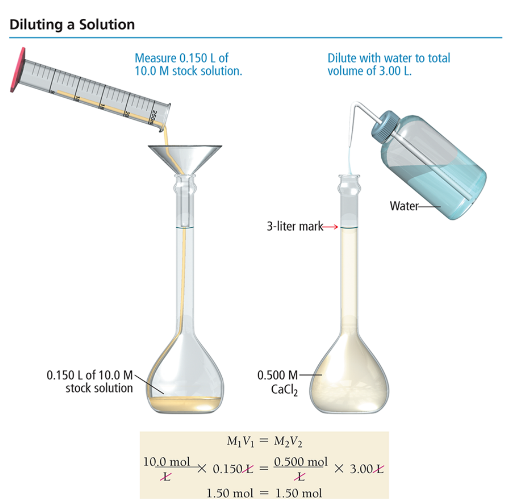
Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution



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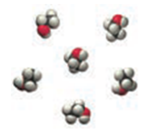
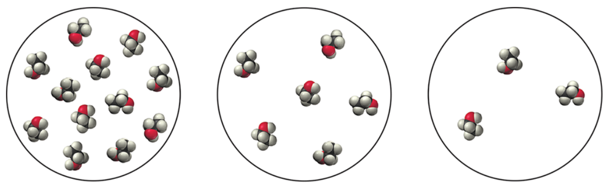
Conceptual Connection 5.3 (1 of 2)



The image shown above represents a small volume within 500 mL of aqueous ethanol (CH_3CH_2OH) solution. (The water molecules have been omitted for clarity.)

Which of the following images best represents the same volume of the solution after we add an additional 500 mL of water?



(a) (b) (c)

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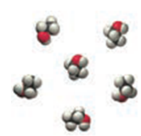
Conceptual Connection 5.3 (2 of 2)



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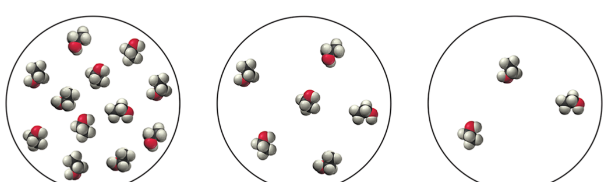

(a) (b) (c)

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Solution Stoichiometry

- In aqueous reactions, we can use the volume and concentration of a reactant or product to calculate its amount in moles.
- We can then use the stoichiometric coefficients in the chemical equation to convert to the amount of another reactant or product in moles.
  - The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product.

Volume A → Amount A (in moles) → Amount B (in moles) → Volume B

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Conceptual Connection 5.4 (1 of 2)

Consider the reaction:

$$2\ \mathrm{A}(aq) + \mathrm{B}(aq) \rightarrow \mathrm{C}(aq)$$

What is the limiting reactant if you mix equal volumes of a 1 M solution of A and a 1 M solution of B?

a. A

b. B

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Conceptual Connection 5.4 (2 of 2)

Consider the reaction:

$$2\ A(aq) + B(aq) \rightarrow C(aq)$$

What is the limiting reactant if you mix equal volumes of a 1 M solution of A and a 1 M solution of B?

a. A

b. B

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Types of Aqueous Solutions and Solubility

- Consider two familiar aqueous solutions: salt water and sugar water.
  - Salt water is a homogeneous mixture of NaCl and H_2O.
  - Sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O.
- 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?

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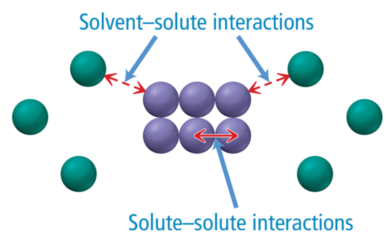
What Happens When a Solute Dissolves?

- 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.

Solute and Solvent Interactions



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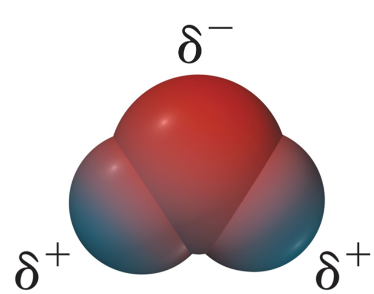
Charge Distribution in a Water Molecule

- 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 (δ^{-}) and the hydrogen side to have a partial positive charge (δ^{+}).



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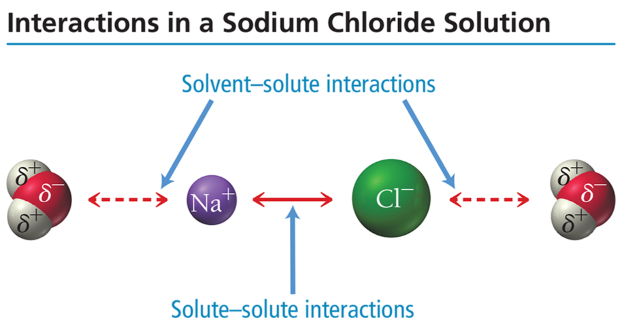
Solute and Solvent Interactions in a Sodium Chloride Solution

- 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.



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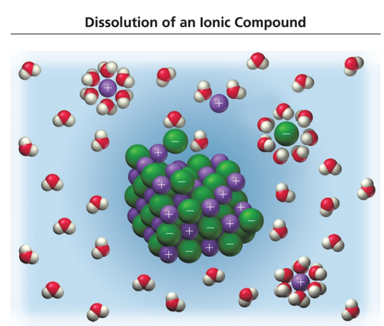
Dissolution of Ionic Compounds

- 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.



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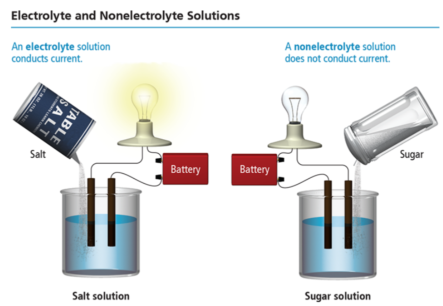
Electrolyte and Nonelectrolyte Solutions (1 of 2)

- 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.



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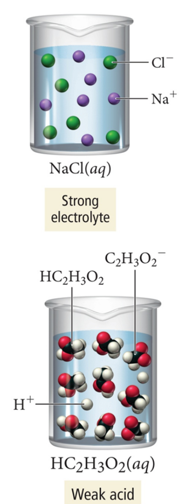
Electrolyte and Nonelectrolyte Solutions (2 of 2)

- 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**.



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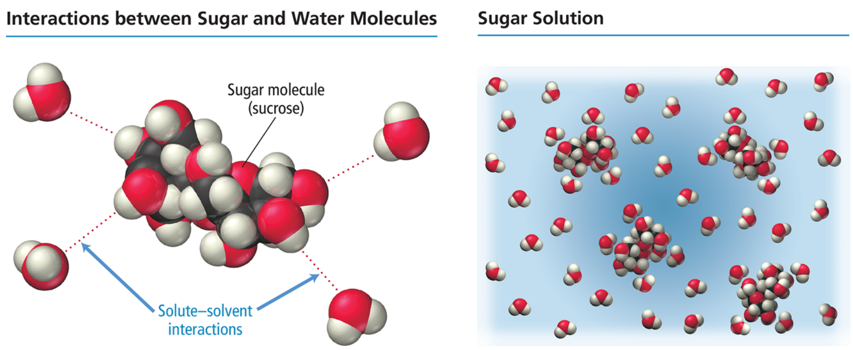
Sugar Dissolution in Water



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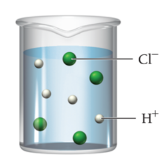
Acids

- Acids are molecular compounds that form ions when dissolved in water.
- Acids that completely ionize in water are called **strong acids**.

$$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

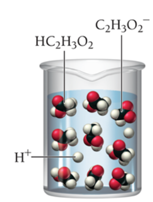
- Acids that do not completely ionize in water are called **weak acids**.

$$HF(aq) \quad H^{+}(aq) + F^{-}(aq)$$



$HCl(aq)$

Strong acid



$HC_2H_3O_2(aq)$

Weak acid

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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.

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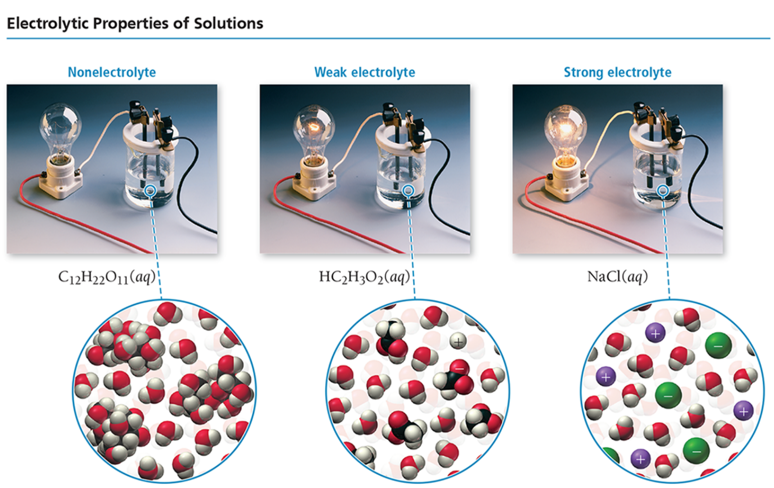
Classes of Dissolved Materials



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Conceptual Connection 5.5 (1 of 2)

Which aqueous solution conducts electricity?

a. 1.0 M KBr

b. 1.0 M $C_6H_{12}O_6$

c. 1.0 M CH_3OH

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Conceptual Connection 5.5 (2 of 2)

Which aqueous solution conducts electricity?

a. 1.0 M KBr

b. 1.0 M $C_6H_{12}O_6$

c. 1.0 M CH_3OH

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The Solubility of Ionic Compounds

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

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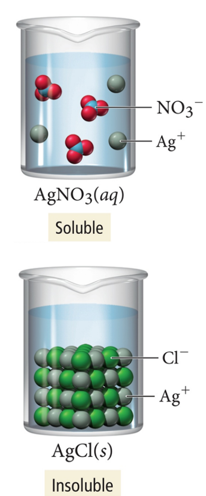
Solubility of Salts

- 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.



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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**.

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Solubility Rules (1 of 2)

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li^+, Na^+, K^+, and NH_4^+	None
NO_3^- and $C_2H_3O_2^-$	None
Cl^-, Br^-, and I^-	When these ions pair with Ag^+, Hg_4^{2+} or Pb^{2+} the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, Ag^+, or Ca^{2+}, the resulting compound is insoluble.

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Solubility Rules (2 of 2)

Compounds Containing the Following Ions Are Generally Insoluble	Exceptions
OH^- and S^{2-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.
	When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble.
	When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is slightly soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.

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Conceptual Connection 5.6 (1 of 2)

The presence of one of the following ions within a compound indicates that a compound is soluble with no exceptions. Which ion?

a. OH^-

b. SO_4^{2-}

c. NO_3^-

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Conceptual Connection 5.6 (2 of 2)

The presence of one of the following ions within a compound indicates that a compound is soluble with no exceptions. Which ion?

a. OH^-

b. SO_4^{2-}

c. $\mathbf{NO_3^-}$

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Precipitation Reactions

- **Precipitation reactions** are ones 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**.

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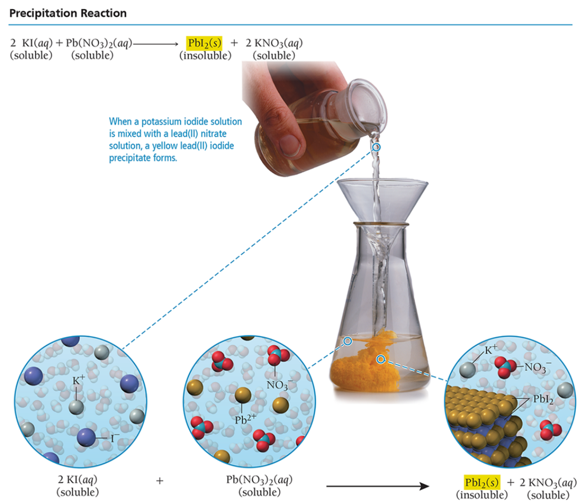
Precipitation of Lead (II) Iodide



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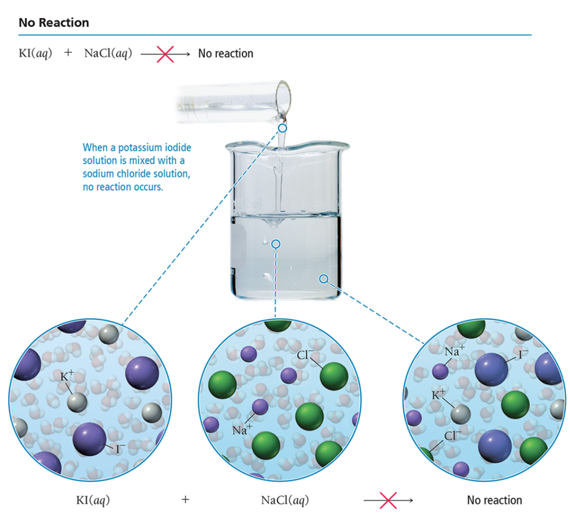
No Precipitation Means No Reaction



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Predicting Precipitation Reactions (1 of 3)

1. Determine what ions constitute each aqueous reactant.
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.

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Predicting Precipitation Reactions (2 of 3)

4. If neither product will precipitate, write **no reaction** after the arrow.
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.

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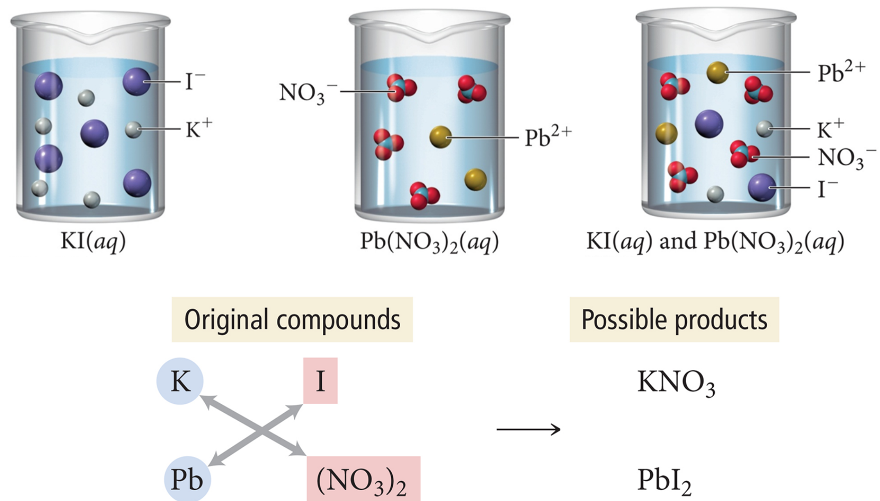
Predicting Precipitation Reactions (3 of 3)



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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**.

$$2\ KOH(aq) + Mg(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + Mg(OH)_2(s)$$

- 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**.

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Writing a Complete Ionic Equation

- 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 written in molecule form

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

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Ionic Equation

- 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).

$$Pb^{2+}(aq) + 2\ NO_3^{-}(aq) + 2\ K^{+}(aq) + 2\ Cl^{-}(aq) \longrightarrow PbCl_2(s) + 2\ K^{+}(aq) + 2\ NO_3^{-}(aq)$$

Spectator ions

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Net Ionic Equation

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

$$Pb^{2+}(aq) + \cancel{2\ NO_3^-}(aq) + \cancel{2\ K^+}(aq) + 2\ Cl^-(aq) \rightarrow$$
$$PbCl_2(s) + \cancel{2\ K^+}(aq) + \cancel{2\ NO_3^-}(aq)$$

- The net ionic equation is

$$Pb^{2+}(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2(s)$$

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Summarizing Aqueous Equations

- A **molecular equation** is a chemical equation showing the complete, neutral formulas for every compound in a reaction.
- A **complete ionic equation** is a chemical equation showing all of the species as they are actually present in solution: strong electrolytes are therefore represented as their component ions.
- A **net ionic equation** is an equation showing only the species that actually change during the reaction.

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Conceptual Connection 5.7 (1 of 2)

Which of the ions listed below is a spectator ion in the complete ionic equation shown here?

$$Ag^{+}(aq) + NO_3^{-}(aq) + Na^{+}(aq) + I^{-}(aq) \rightarrow AgI(s) + NO_3^{-}(aq) + Na^{+}(aq)$$

a. $Ag^{+}(aq)$

b. $NO_3^{-}(aq)$

c. $I^{-}(aq)$

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Conceptual Connection 5.7 (2 of 2)

Which of the ions listed below is a spectator ion in the complete ionic equation shown here?

$$Ag^{+}(aq) + NO_3^{-}(aq) + Na^{+}(aq) + I^{-}(aq) \rightarrow AgI(s) + NO_3^{-}(aq) + Na^{+}(aq)$$

a. $Ag^{+}(aq)$

b. $NO_3^{-}(aq)$

c. $I^{-}(aq)$

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Acid–Base and Gas-Evolution Reactions (1 of 2)

- Two other important classes of reactions that occur in aqueous solution are
  1. acid–base reactions, and
  2. gas-evolution reactions.
- Acid–base reaction:
  - 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).

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Acid–Base and Gas-Evolution Reactions (2 of 2)

- 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.

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Acid–Base Reactions (1 of 2)

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.

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

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Acid–Base Reactions (2 of 2)

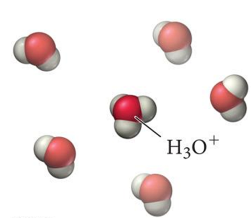
- These reactions are called **neutralization reactions** because the acid and base neutralize each other's properties.

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$$

- 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.



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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**, $\mathbf{H_3O^+}$.
    - Most chemists use H^+ and H_3O^+ 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.

$$\underbrace{HCl(aq)}_{\text{Acid}} + \underbrace{NaOH(aq)}_{\text{Base}} \longrightarrow \underbrace{H_2O(l)}_{\text{Water}} + \underbrace{NaCl(aq)}_{\text{Salt}}$$

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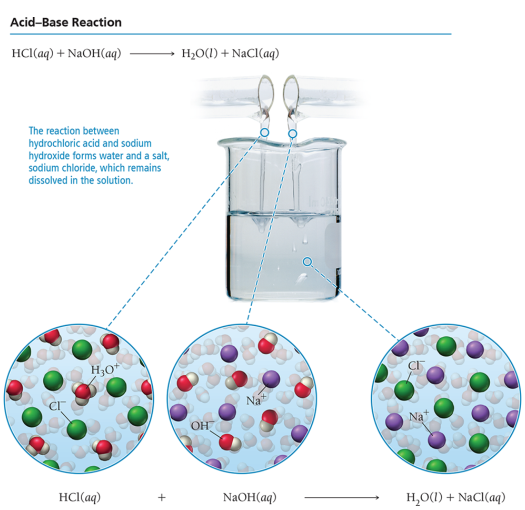
Acid–Base Reaction



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Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Formic acid	$HCHO_2$ (weak acid)		
Acetic acid	$HC_2H_3O_2$ (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:

$NH_3(aq) + H_2O(l)$ É $NH_4^+(aq) + OH^+(aq)$.

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Predict the Product of the Reactions

1. $HCl(aq) + Ba(OH)_2(aq) \rightarrow$

$$(H^+ + Cl^-) + (Ba^{2+} + OH^-) \rightarrow (H^+ + OH^-) + (Ba^{2+} + Cl^-)$$

$$HCl(aq) + Ba(OH)_2(aq) \rightarrow H_2O(l) + BaCl_2$$

$$2\ HCl(aq) + Ba(OH)_2(aq) \rightarrow 2\ H_2O(l) + BaCl_2(aq)$$

2. $HC_2H_3O_2(aq) + KOH(aq) \rightarrow$

$$(H^+ + C_2H_3O_2^-) + (K^+ + OH^-) \rightarrow (H^+ + OH^-) + (K^+ + C_2H_3O_2^-)$$

$$HC_2H_3O_2(aq) + KOH(aq) \rightarrow H_2O(l) + KC_2H_3O_2$$

$$HC_2H_3O_2(aq) + 2\ KOH(aq) \rightarrow 2\ H_2O(l) + KC_2H_3O_2$$

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Acid–Base Titrations

- A **titration** is a laboratory procedure where a substance in a solution of known concentration (**titrant**) is reacted with another substance in a solution of unknown concentration (**analyte**).
- The **equivalence point** is the point in the titration when the H^+ and 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.

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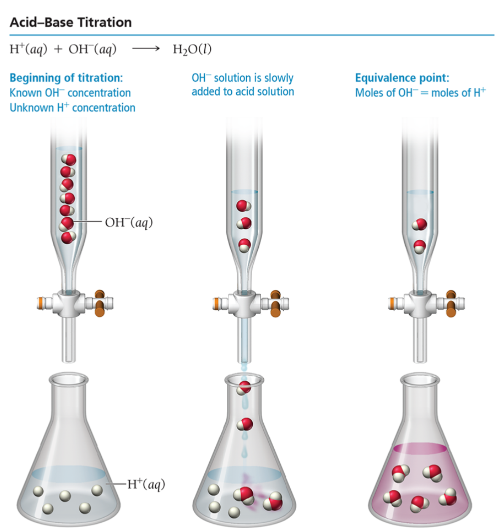
Acid–Base Titration



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Indicator in Titration

In the previous titration figure, 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.

Indicator in Titration

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Conceptual Connection 5.8 (1 of 2)

A 10.0 mL sample of 0.20 M HBr solution is titrated with 0.10 M NaOH. What volume of NaOH is required to reach the equivalence point?

a. 10.0 mL

b. 20.0 mL

c. 40.0 mL

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Conceptual Connection 5.8 (2 of 2)

A 10.0 mL sample of 0.20 M HBr solution is titrated with 0.10 M NaOH. What volume of NaOH is required to reach the equivalence point?

a. 10.0 mL

b. 20.0 mL

c. 40.0 mL

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Gas-Evolving Reactions

- Some reactions form a gas directly from the ion exchange.

$$H_2SO_4(aq) + Li_2S(aq) \rightarrow Li_2SO_4(aq) + H_2S(g)$$

- Other reactions form a gas by the subsequent decomposition of one of the ion exchange products into a gas and water.

$$HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2CO_3(aq)$$

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

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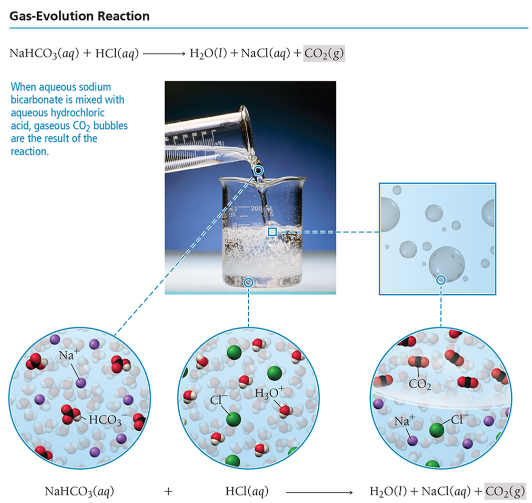
Gas-Evolution Reaction



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Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2\ HCl(aq) + K_2S(aq) \rightarrow H_2S(g) + 2\ KCl(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2\ HCl(aq) + K_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + 2\ KCl(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2\ HCl(aq) + K_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g) + 2\ KCl(aq)$
Ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \rightarrow H_2O(l) + NH_3(g) + KCl(aq)$

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Oxidation–Reduction Reactions

- The reactions in which electrons are transferred from one reactant to the other are called **oxidation–reduction reactions,** or **redox reactions**.
- Many redox reactions involve the reaction of a substance with oxygen.

$$4\ \text{Fe}(s) + 3\ \text{O}_2(g) \rightarrow 2\ \text{Fe}_2\text{O}_3(s)(\text{rusting})$$

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

$$2\ \text{H}_2(g) + \text{O}_2(g) \rightarrow 2\ \text{H}_2\text{O}(g)(\text{combustion})$$

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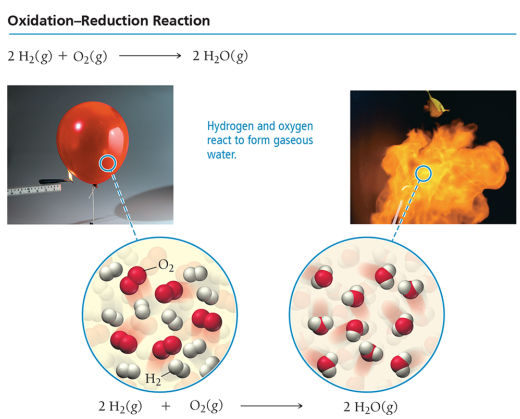
Combustion as Redox

$$2\ \text{H}_2(g) + \text{O}_2(g) \rightarrow 2\ \text{H}_2\text{O}(g)$$



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Oxidation and Reduction

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

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{Na^+Cl^-}(s)$$

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + 1\ \mathrm{e^-}\ \ (\text{oxidation})$$

$$\mathrm{Cl_2} + 2\ \mathrm{e^-} \rightarrow 2\ \mathrm{Cl^-}\ \ (\text{reduction})$$

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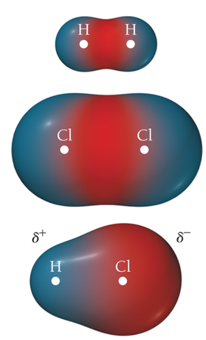
Redox Reaction

- Electron transfer does not need to be a **complete** transfer for the reaction to qualify as oxidation–reduction. Example:

$$\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{HCl}(g)$$

- There is uneven sharing of electrons when hydrogen bonds to chlorine, resulting in an increase of electron density (reduction) for chlorine and a decrease in electron density (oxidation) for hydrogen.



Hydrogen loses electron density (oxidation) and chlorine gains electron density (reduction).

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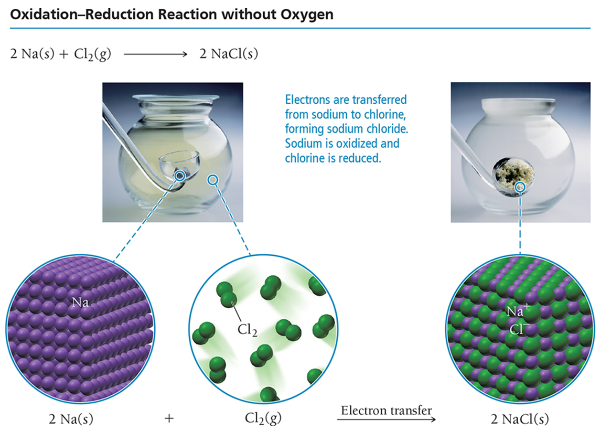
Redox without Combustion

$$2\ \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\ \text{NaCl}(s)$$



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Reactions of Metals with Nonmetals

- Consider the following reaction:

$$2\ \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\ \text{NaCl}(s)$$

- The reaction involves a metal reacting with a nonmetal.
- In addition, the reaction involves the conversion of free elements into ions.

$$2\ \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\ \text{Na}^+\text{Cl}^-(s)$$

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Oxidation States (1 of 2)

- We need a method for determining how the electrons are transferred in reactions that don't involve metal-nonmetal or combustion.
- Chemists assign a number to each element in a reaction called an **oxidation state** or **oxidation number** allowing them to determine the electron flow in the reaction, like electron bookkeeping.

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Oxidation States (2 of 2)

- The oxidation number of an atom in a compound is the "charge" it would have if all shared electrons were assigned to the atom with the greatest attraction for those electrons.
- Oxidation states are imaginary charges assigned based on a set of rules.
  - They are written with magnitude before charge, for example +1 or −1.
- Ion charges are real, measurable charges.
  - They are written with charge before magnitude, for example 1+ or 1−.

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Rules for Assigning Oxidation States (1 of 4)

The following rules are in order of priority:

1. Free elements have an oxidation state = 0.

$$\mathbf{Cu = 0 \text{ and } Cl_2 = 0}$$

2. Monatomic ions have an oxidation state equal to their charge.

$$\mathbf{Ca^{2+} = +2 \quad Cl^{-} = -1}$$

3. (a) The sum of the oxidation states of all the atoms in a compound is 0.

$$\mathbf{\text{For } H_2O:\ H = +1 \text{ and } O = -2}$$
$$\mathbf{2(+1) + (-2) = 0}$$

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Rules for Assigning Oxidation States (2 of 4)

(b) The sum of the oxidation states of all the atoms in a polyatomic ion equals the charge on the ion.

$$NO_3^{-}: N = +5 \text{ and } O = -2 \quad (+5) + 3(-2) = -1$$

4. (a) Group I metals have an oxidation state of +1 in all of their compounds.

$$Na = +1 \text{ in } NaCl$$

(b) Group II metals have an oxidation state of +2 in all of their compounds.

$$Mg = +2 \text{ in } MgCl_2$$

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Rules for Assigning Oxidation States (3 of 4)

5. In their compounds, nonmetals have oxidation states according to the table below.

   Nonmetals higher on the table take priority.

Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 −1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 −2 ox state
Group 7A	−1	CCl_4 −1 ox state
Group 6A	−2	H_2S −2 ox state
Group 5A	−3	NH_3 −3 ox state

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Rules for Assigning Oxidation States (4 of 4)

- When assigning oxidation states, keep these points in mind:
  - The oxidation state of any given element generally depends on what other elements are present in the compound (except groups 1A and 2A metals)
  - When following the hierarchy shown in rule 5, give priority to the element(s) highest on the list and then assign the oxidation state of the element lowest on the list using rule 3.
  - When assigning oxidation states to elements that are not covered by rules 4 and 5 (such as carbon), use rule 3 to deduce their oxidation state once all other oxidation states have been assigned.

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Conceptual Connection 5.9 (1 of 2)

Which statement best describes the difference between the charge of a polyatomic ion and the oxidation states of its constituent atoms? (For example, the charge of NO_3^- is 1−, and the oxidation states of its atoms are +5 for the nitrogen atom and −2 for each oxygen atom.)

a. The charge of a polyatomic ion is a property of the entire ion, while the oxidation states are assigned to each individual atom.

b. The oxidation state of the ion is the same as its charge.

c. The charge of a polyatomic ion is not a real physical property, while the oxidation states of atoms are actual physical properties.

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Conceptual Connection 5.9 (2 of 2)

Which statement best describes the difference between the charge of a polyatomic ion and the oxidation states of its constituent atoms? (For example, the charge of NO_3^- is 1−, and the oxidation states of its atoms are +5 for the nitrogen atom and −2 for each oxygen atom.)

a. The charge of a polyatomic ion is a property of the entire ion, while the oxidation states are assigned to each individual atom.

b. The oxidation state of the ion is the same as its charge.

c. The charge of a polyatomic ion is not a real physical property, while the oxidation states of atoms are actual physical properties.

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Identifying Redox Reactions

- Oxidation: An increase in oxidation state
- Reduction: A decrease in oxidation state

$$\mathrm{C} + 2\,\mathrm{S} \longrightarrow \mathrm{CS_2}$$

Oxidation states: 0 0 +4 −2

Reduction

Oxidation

  - Carbon changes from an oxidation state of 0 to an oxidation state of +4.
    - Carbon **loses electrons** and is **oxidized**.
  - Sulfur changes from an oxidation state of 0 to an oxidation state of −2.
    - Sulfur **gains electrons** and is **reduced**.

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Redox Reactions

- Oxidation and reduction must occur simultaneously.
  - If an atom loses electrons another atom must take them.
- The reactant that causes reduction in another reactant is called the **reducing agent**.
  - The reducing agent contains the element that is oxidized.
- The reactant that causes oxidation in another reactant is called the **oxidizing agent**.
  - The oxidizing agent contains the element that is reduced.

$$2\ \mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{Na^+Cl^-}(s)$$

Na is oxidized, while Cl is reduced.

Na is the reducing agent, and Cl_2 is the oxidizing agent.

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Guidelines for Oxidizing and Reducing Agents

Redox reactions:

- Any reaction in which there is a change in the oxidation states of atoms in going from reactants to products.

In a redox reaction:

- The oxidizing agent oxidizes another substance (and is itself reduced).
- The reducing agent reduces another substance (and is itself oxidized).

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Conceptual Connection 5.10 (1 of 2)

Which statement is true?

a. A redox reaction involves either the transfer of an electron or a change in the oxidation state of an element.

b. If any of the reactants or products in a reaction contain oxygen, the reaction is a redox reaction.

c. In a reaction, oxidation can occur independently of reduction.

d. In a redox reaction, any increase in the oxidation state of a reactant must be accompanied by a decrease in the oxidation state of another reactant.

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Conceptual Connection 5.10 (2 of 2)

Which statement is true?

a. A redox reaction involves either the transfer of an electron or a change in the oxidation state of an element.

b. If any of the reactants or products in a reaction contain oxygen, the reaction is a redox reaction.

c. In a reaction, oxidation can occur independently of reduction.

d. In a redox reaction, any increase in the oxidation state of a reactant must be accompanied by a decrease in the oxidation state of another reactant.

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Activity Series to Determine Spontaneity

- **Activity series of metals** is a table listing metals in order of decreasing tendency to lose electrons.
  - Metals at the top have higher tendency to undergo oxidation, most reactive.
  - Metals at the bottom have the least tendency to undergo oxidation, least reactive.
- Each reaction in the activity series is an oxidation half-reaction.
- The half-reactions at the top are most likely to occur in the forward direction.
- The half-reactions at the bottom are most likely to occur in the reverse direction.

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Activity Series

$Li(s) \longrightarrow Li^{+}(aq) + e^{-}$	Most reactive
$K(s) \longrightarrow K^{+}(aq) + e^{-}$	Most easily oxidized
$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^{-}$	Strongest tendency to lose electrons
$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$	
$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^{-}$	
$Al(s) \longrightarrow Al^{3+}(aq) + 3e^{-}$	
$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^{-}$	
$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$	
$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^{-}$	
$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$	
$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^{-}$	
$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^{-}$	
$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^{-}$	
$H_2(g) \longrightarrow 2H^{+}(aq) + 2e^{-}$	Least reactive
$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$	Most difficult to oxidize
$Ag(s) \longrightarrow Ag^{+}(aq) + e^{-}$	
$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$	Least tendency to lose electrons

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Conceptual Connection 5.11 (1 of 2)

Which metal is most easily oxidized?

a. Na

b. Cr

c. Au

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Conceptual Connection 5.11 (2 of 2)

Which metal is most easily oxidized?

a. Na

b. Cr

c. Au

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