Chapter 7
Chemical Reactions

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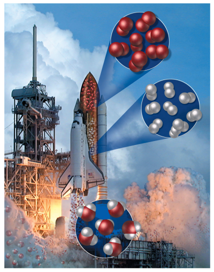
Chem 118
Introductory Chemistry



Map: Introductory Chemistry (Tro) https://chem.libretexts.org/@go/page/45050 (accessed Mar 25, 2022).

Evidence of Chemical Reactions

Chemical Reactions

- Reactions involve chemical changes in matter resulting in new substances.
- Reactions involve rearrangement and exchange of atoms to produce new molecules.
  - ✓Elements are not transmuted during a reaction.
  - ✓Atoms of different elements can combine to make new compounds.
  - ✓Molecules can combine to make bigger molecules.
  - ✓Molecules can decompose into smaller molecules or atoms.
  - ✓Atoms can be exchanged between molecules or transferred to another molecule.
  - ✓Atoms can gain or lose electrons, turning them into ions.
    - ➢Or changing the charge on ions that are already there.

Evidence of Chemical Reactions

- Look for evidence of a new substance.
- Permanent Visual clues.
  - ✓Color change.
  - ✓Precipitate formation.
    - ➢Solid that forms when liquid solutions are mixed.
  - ✓Gas bubbles.
  - ✓Large energy changes.
    - ➢Large Temperature Changes
    - ➢Emission of light.
- Other clues.
  - ✓New odor.
  - ✓Whooshing sound from a tube.
  - ✓Permanent new state.

Evidence is Not Proof

- In order to be absolutely sure that a chemical reaction has taken place, you need to go down to the molecular level and analyze the structures of the molecules at the beginning and end.

Is boiling water
a chemical change?

Practice—Decide Whether Each of the Following Involve a Chemical Reaction.

- Photosynthesis Yes, CO_2 and H_2O combine into carbohydrates
- Heating sugar until it turns black Yes, sugar decomposing
- Heating ice until it turns liquid No, molecules still same
- Digestion of food Yes, food decomposing and combining with stomach acid
- Dissolving sugar in water No, molecules still same
- Burning of alcohol in a flambé dessert

  Yes, alcohol combining with O_2 to make CO_2 and H_2O

Chemical Equations

Chemical Equations

- Short-hand way of describing a reaction.
- Provides information about the reaction.
  - ✓Formulas of reactants and products.
  - ✓States of reactants and products.
  - ✓Relative numbers of reactant and product molecules that are required.
  - ✓Can be used to determine masses of reactants used and products that can be made.

Conservation of Mass

- Matter cannot be created or destroyed.
  - ✓Therefore, the total mass cannot change.
  - ✓And the total mass of the reactants will be the same as the total mass of the products.
- In a chemical reaction, all the atoms present at the beginning are still present at the end.
  - ✓If all the atoms are still there, then the mass will not change.

The Combustion of Methane

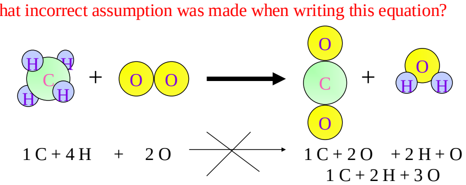
- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - ✓Whenever something burns it combines with $O_2(g)$.

Combustion of Methane

- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - ✓Whenever something burns it combines with $O_2(g)$.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

What incorrect assumption was made when writing this equation?

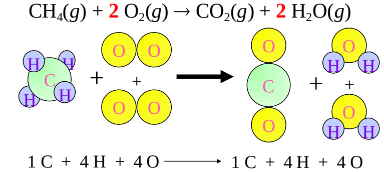


We are assuming that all reactants combine 1 molecule : 1 molecule; and that 1 molecule of each product is made – an incorrect assumption

Combustion of Methane, Balanced

- To show the reaction obeys the Law of Conservation of Mass the equation must be **balanced**.
  - ✓We adjust the numbers of molecules so there are equal numbers of atoms of each element on both sides of the arrow.

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$



Chemical Equations, Continued

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

- This equation is balanced, meaning that there are equal numbers of atoms of each element on the reactant and product sides.
  - ✓To obtain the number of atoms of an element, multiply the subscript by the coefficient.

$$1 \leftarrow C \rightarrow 1$$
$$4 \leftarrow H \rightarrow 4$$
$$4 \leftarrow O \rightarrow 2 + 2$$

Chemical Equations

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

- CH_4 and O_2 are the reactants, and CO_2 and H_2O are the products.
- The (*g*) after the formulas tells us the state of the chemical.
- The number in front of each substance tells us the numbers of those molecules in the reaction.
  - ✓Called the **coefficients**.

Symbols Used in Equations

- Symbols used to indicate state after chemical.
  - ✓(*g*) = gas; (*l*) = liquid; (*s*) = solid.
  - ✓(*aq*) = aqueous = dissolved in water.
- Energy symbols used above the arrow.
  - ✓ Δ = heat.
  - ✓ hν = light.
  - ✓shock = mechanical.
  - ✓elec = electrical.

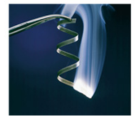
Writing Balanced Chemical Equations

1. Write a skeletal equation by writing the formula of each reactant and product.
2. Count the number of atoms of each element on each side of the equation.
   - ✓ Polyatomic ions may often be counted as if they are one "element".
3. Pick an element to balance.
   - ✓ If an element is found in only one compound on both sides, balance it first.
     - ➢ Metals before nonmetals.
   - ✓ Leave elements that are free elements somewhere in the equation until last.
     - ➢ Balance free elements by adjusting the coefficient where it is a free element.

Writing Balanced Chemical Equations, Continued

4. Find the least common multiple (LCM) of the number of atoms on each side.
   - ✓ The LCM of 3 and 2 is 6.
5. Multiply each count by a factor to make it equal to the LCM.
6. Use this factor as a coefficient in the equation.
   - ✓ If there is already a coefficient there, multiply it by the factor.
   - ✓ It must go in front of entire molecules, not between atoms within a molecule.
7. Recount and repeat until balanced.

Example



- When magnesium metal burns in air, it produces a white, powdery compound magnesium oxide.

1. Write a skeletal equation

$Mg(s) + O_2(g) \rightarrow MgO(s)$

2. Count the number of atoms on each side.

$Mg(s) + O_2(g) \rightarrow MgO(s)$

$1 \leftarrow Mg \rightarrow 1$

$2 \leftarrow O \rightarrow 1$

Example, Continued

- When magnesium metal burns in air, it produces a white, powdery compound magnesium oxide.

$Mg(s) + O_2(g) \rightarrow MgO(s)$

3. Pick an element to balance.
   - ✓ Avoid element in multiple compounds.
   - ✓ Do free elements last.
   - ✓ Since Mg already balanced, pick O.
4. Find the LCM of both sides
5. and multiply each side by factor so it equals LCM.
   - ✓ LCM of 2 and 1 is 2.

$Mg(s) + O_2(g) \rightarrow MgO(s)$

$1 \leftarrow Mg \rightarrow 1$

$1 \times 2 \leftarrow O \rightarrow 1 \times 2$

Example, Continued

6. Use factors as coefficients in ***front*** of the compound containing the element.

$Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

$1 \leftarrow Mg \rightarrow 1 \times 2$

$2 \leftarrow O \rightarrow 1 \times 2$

7. Recount—Mg not balanced now

$Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

$1 \leftarrow Mg \rightarrow 2$

$2 \leftarrow O \rightarrow 2$

Example, Continued

8. and Repeat—attacking an unbalanced element.

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

$2 \times 1 \leftarrow Mg \rightarrow 2$

$2 \leftarrow O \rightarrow 2$

9. Recount—Mg not balanced now

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

$2 \leftarrow Mg \rightarrow 2$

$2 \leftarrow O \rightarrow 2$

Another Example

- Under appropriate conditions at 1000°C, ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide and steam

1. write the skeletal equation

a) first in words
   - ✓ identify the state of each chemical

ammonia(g) + oxygen(g) → nitrogen monoxide(g) + water(g)

b) then write the equation in formulas
   - ✓ identify diatomic elements
   - ✓ identify polyatomic ions
   - ✓ determine formulas

$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$

Another Example, Continued

2) count the number of atoms of on each side

$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$

$1 \leftarrow N \rightarrow 1$

$3 \leftarrow H \rightarrow 2$

$2 \leftarrow O \rightarrow 1 + 1$

Another Example, Continued

3) pick an element to balance - H
   - ✓ avoid element in multiple compounds on same side - O
4) find least common multiple of both sides (6)
5) multiply each side by factor so it equals LCM

$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$

$1 \leftarrow N \rightarrow 1$

$2 \times 3 \leftarrow H \rightarrow 3 \times 2$

$2 \leftarrow O \rightarrow 2$

Another Example, Continued

6) use factors as coefficients in ***front*** of compound containing the element

$2\ NH_3(g) + O_2(g) \rightarrow NO(g) + 3\ H_2O(g)$

$2 \times 1 \leftarrow N \rightarrow 1$

$2 \times 3 \leftarrow H \rightarrow 3 \times 2$

$2 \leftarrow O \rightarrow (3 \times 1) + 1$

Another Example, Continued

7) Recount – N & O not balanced

$2\ NH_3(g) + O_2(g) \rightarrow NO(g) + 3\ H_2O(g)$

$2 \leftarrow N \rightarrow 1$

$6 \leftarrow H \rightarrow 6$

$2 \leftarrow O \rightarrow 1 + 3$

7) and Repeat – attack the N

$2\ NH_3(g) + O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$

$2 \leftarrow N \rightarrow 1 \times 2$

$6 \leftarrow H \rightarrow 6$

$2 \leftarrow O \rightarrow (1 \times 2) + 3$

Another Example, Continued

7) Recount Again – Still not balanced and the only element left is O!

$2\ NH_3(g) + O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$

$2 \leftarrow N \rightarrow 2$

$6 \leftarrow H \rightarrow 6$

$2 \leftarrow O \rightarrow 2 + 3$

Another Example, Continued

7) and Repeat Again
   - ✓ A trick of the trade - when you are forced to attack an element that is in 3 or more compounds – find where it is uncombined. You can find a factor to make it any amount you want, even if that factor is a fraction!

$2\ NH_3(g) + ?\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$

$2 \leftarrow N \rightarrow 2$

$6 \leftarrow H \rightarrow 6$

$2 \leftarrow O \rightarrow 2 + 3$

   - ✓ We want to make the O on the left equal 5, therefore we will multiply it by 2.5

$2\ NH_3(g) + 2.5\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$

$2 \leftarrow N \rightarrow 2$

$6 \leftarrow H \rightarrow 6$

$2.5 \times 2 \leftarrow O \rightarrow 2 + 3$

Another Example, Continued

7) You can't have a coefficient that isn't a whole number. Multiply all the coefficients by a number to eliminate fractions
   - ✓ If ?.5, then multiply by 2; if ?.33, then 3; if ?.25, then 4

$\{2\ NH_3(g) + 2.5\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)\} \times 2$

$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

$4 \leftarrow N \rightarrow 4$

$12 \leftarrow H \rightarrow 12$

$10 \leftarrow O \rightarrow 4 + 6$

Practice 1

When aluminum metal reacts with air, it produces a white, powdery compound called aluminum oxide.

- ✓ Reacting with air means reacting with O_2:

Aluminum(*s*) + oxygen(*g*) → aluminum oxide(*s*)

$Al(s) + O_2(g) \rightarrow Al_2O_3(s)$

$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

Practice #2

Acetic acid reacts with the metal aluminum to make aqueous aluminum acetate and gaseous hydrogen.

- ✓ Acids are always aqueous.
- ✓ Metals are solid except for mercury.
- ✓ Acid/Metal Reactions liberate H_2 gas.

$Al(s) + HC_2H_3O_2(aq) \rightarrow Al(C_2H_3O_2)_3(aq) + H_2(g)$

$2\ Al(s) + 6\ HC_2H_3O_2(aq) \rightarrow 2\ Al(C_2H_3O_2)_3(aq) + 3\ H_2(g)$

Practice #3

Combustion of ethyl alcohol

- ✓ Combustion is burning, and therefore, reacts with O_2.
- ✓ Combustion of compounds containing C and H always make $CO_2(g)$ and $H_2O(g)$ as products.

$C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

$C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

Practice #4, Continued

Combustion of liquid butane (C_4H_{10}) in a lighter.

$C_4H_{10}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

$2\ C_4H_{10}(l) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$

Aqueous Solutions

Aqueous Solutions

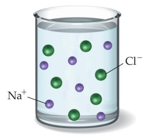
- Many times, the chemicals we are reacting together are dissolved in water.
  - ✓Mixtures of a chemical dissolved in water are called **aqueous solutions**.
- Dissolving the chemicals in water helps them to react together faster.
  - ✓The water separates the chemicals into individual molecules or ions.
  - ✓The separate, free-floating particles come in contact more frequently so the reaction speeds up.

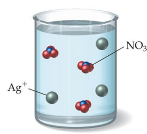
Predicting Whether a Reaction Will Occur in Aqueous Solution

- "Forces" that drive a reaction:
  - ✓ Formation of a solid.
  - ✓ Formation of water.
  - ✓ Formation of a gas.
  - ✓ Transfer of electrons.
- When chemicals (dissolved in water) are mixed and one of the above-noted forces occur, the reaction will generally happen.

Dissociation

- When ionic compounds dissolve in water, the anions and cations are separated from each other. This is called **dissociation**.
  - ✓However, not all ionic compounds are soluble in water!
- When compounds containing **polyatomic ions dissociate**, the polyatomic group stays together as one ion.

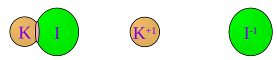




Dissociation, Continued

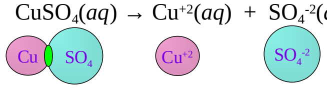
- Potassium iodide dissociates in water into potassium cations and iodide anions.

$KI(aq) \rightarrow K^{+1}(aq) + I^{-1}(aq)$



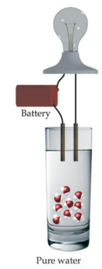
- Copper(II) sulfate dissociates in water into copper(II) cations and sulfate anions.

$CuSO_4(aq) \rightarrow Cu^{+2}(aq) + SO_4^{-2}(aq)$



Electrolytes

- **Electrolytes** are substances whose water solution is a conductor of electricity.
- All electrolytes have ions dissolved in water.



Electrolytes, Continued

- In **strong** electrolytes, **all** the electrolyte molecules or formula units are separated into ions.
  - ✓Example: Sodium Chloride
- In **non**electrolytes, **none** of the molecules are separated into ions.
  - ✓Example: Sugar
- In **weak** electrolytes, a **small percentage** of the molecules are separated into ions.
  - ✓Example: Acetic Acid

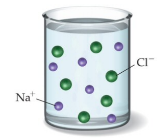
Types of Electrolytes

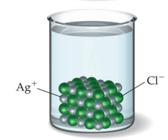
- **Salts** = Water soluble ionic compounds.
  - ✓All strong electrolytes.
- **Acids** = Form H^{+1} ions and anions in water solution.
  - ✓Strong acid = strong electrolyte, weak acid = weak electrolyte.
- **Bases** = Form HO^{-1} ions and cations in water solutions
  - ✓Strong base = strong electrolyte, weak base = weak electrolyte.

Solubility

When Will a Salt Dissolve?

- A compound is **soluble** in a liquid if it dissolves in that liquid.
  - ✓NaCl is soluble in water, but AgCl is not.
- A compound is **insoluble** if a significant amount does not dissolve in that liquid.
  - ✓AgCl is insoluble in water.
    - ➢Though there is a very small amount dissolved, but not enough to be significant.





When Will a Salt Dissolve?, Continued

- Predicting whether a compound will dissolve in water is not easy.
- A convenient way to do it is to do some experiments to test whether a compound will dissolve in water, then develop some rules based on those experimental results.
  - ✓We call this method the **empirical method**.

Solubility Rules: Compounds that Are Generally Soluble in Water

Compounds containing the following ions are generally soluble	Exceptions (when combined with ions on the left the compound is insoluble)
Li^+, Na^+, K^+, NH_4^+	none
NO_3^-, $C_2H_3O_2^-$	none
Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Solubility Rules: Compounds that Are Generally Insoluble

Compounds containing the following ions are generally insoluble	Exceptions (when combined with ions on the left the compound is soluble or slightly soluble)
OH^-	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
S^{2-}	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
CO_3^{2-}, PO_4^{3-}	Li^+, Na^+, K^+, NH_4^+

Using the Solubility Rules to Predict an Ionic Compound's Solubility in Water

- First check the cation: If it is Li^+, Na^+, K^+, or NH_4^+, then the compound will be soluble in water.
  - ✓Regardless of the anion.
- If the cation is not Li^+, Na^+, K^+, or NH_4^+, then follow the rule for the anion.
- If a rule says the compounds are mostly soluble, then the exceptions are insoluble.
- If a rule says the compounds are mostly insoluble, then the exceptions are soluble.
  - ✓Note: slightly soluble ≈ insoluble.

Determine if Each of the Following Is Soluble in Water

- KOH
- AgBr
- $CaCl_2$
- $Pb(NO_3)_2$
- $PbSO_4$

Determine if Each of the Following Is Soluble in Water, Continued

- KOH **Soluble, because the cation is K^+.**
- AgBr **Insoluble, this is an exception!**
- $CaCl_2$ **Soluble, most compounds with Cl^- are soluble.**
- $Pb(NO_3)_2$ **Soluble, because the anion is NO_3^-.**
- $PbSO_4$ **Insoluble, even though most compounds with SO_4^{2-} are soluble, this is an exception.**

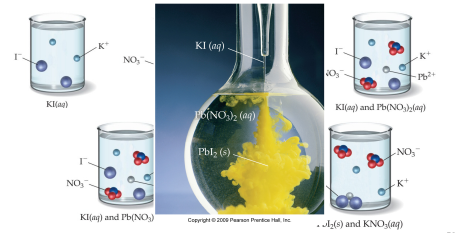
Precipitation Reactions

Precipitation Reactions

- Many reactions are done by mixing aqueous solutions of electrolytes together.
- When this is done, often a reaction will take place from the cations and anions in the two solutions that are exchanging.
- If the ion exchange results in forming a compound that is insoluble in water, it will come out of solution as a **precipitate**.

Precipitation Reactions, Continued

$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + PbI_2(s)$



No Precipitate Formation = No Reaction

$KI(aq) + NaCl(aq) \rightarrow KCl(aq) + NaI(aq)$

All ions still present, ∴ no reaction.

Process for Predicting the Products of a Precipitation Reaction

1. Write the formula for the reactants and Determine what ions each aqueous reactant has.
2. Exchange ions.
   - ✓ (+) ion from one reactant with (-) ion from the other.
3. Balance charges of combined ions to get formula of each product.
4. Balance the equation.
   - ✓ Count atoms.
5. Determine solubility of each product in water.
   - ✓ Use the solubility rules.
   - ✓ If product is insoluble or slightly soluble, it will precipitate.
   - ✓ If neither product will precipitate, **no reaction.**

Example 7.7—When an Aqueous Solution of Sodium Carbonate Is Added to an Aqueous Solution of Copper(II) Chloride, a White Solid Forms.

1. Write the formulas of the reactants and Determine the ions present when each reactant dissociates.

$Na_2CO_3(aq) + CuCl_2(aq) \rightarrow$

$(Na^+ + CO_3^{2-}) + (Cu^{+2} + Cl^-) \rightarrow$

2. Exchange the ions.

$(Na^+ + CO_3^{2-}) + (Cu^{+2} + Cl^-) \rightarrow (Na^+ + Cl^-) + (Cu^{+2} + CO_3^{2-})$

Example 7.7—When an Aqueous Solution of Sodium Carbonate Is Added to an Aqueous Solution of Copper(II) Chloride, a White Solid Forms, Continued.

3. Write the formulas of the products.
   - ✓ Cross charges and reduce.

$Na_2CO_3(aq) + CuCl_2(aq) \rightarrow NaCl + CuCO_3$

4. Balance the equation.

$Na_2CO_3(aq) + CuCl_2(aq) \rightarrow 2\ NaCl + CuCO_3$

Example 7.7—When an Aqueous Solution of Sodium Carbonate Is Added to an Aqueous Solution of Copper(II) Chloride, a White Solid Forms, Continued.

5. Determine the solubility of each product. Write an (*s*) after the insoluble products and a (*aq*) after the soluble products

NaCl is soluble.

$CuCO_3$ is insoluble.

$Na_2CO_3(aq) + CuCl_2(aq) \rightarrow 2\ NaCl(aq) + CuCO_3(s)$

Practice–Predict the Products and Balance the Equation

- $KCl(aq) + AgNO_3(aq) \rightarrow$
- $KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s)$
- $Na_2S(aq) + CaCl_2(aq) \rightarrow$
- $Na_2S(aq) + CaCl_2(aq) \rightarrow 2\ NaCl(aq) + CaS(aq)$
- No reaction.

Ionic Equations

- Equations that describe the chemicals put into the water and the product molecules are called **molecular equations**.

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

- Equations that describe the actual dissolved species are called **complete ionic equations**.
  - ✓ Aqueous electrolytes are written as ions.
    - ➢ Soluble salts, strong acids, strong bases.
  - ✓ Insoluble substances and nonelectrolytes written in molecule form.
    - ➢ Solids, liquids, and gases are not dissolved, therefore, molecule form.

$2K^{+1}{}_{(aq)} + 2OH^{-1}{}_{(aq)} + Mg^{+2}{}_{(aq)} + 2NO_3^{-1}{}_{(aq)} \rightarrow 2K^{+1}{}_{(aq)} + 2NO_3^{-1}{}_{(aq)} + Mg(OH)_{2(s)}$

Writing Complete Ionic Equations

- Rewrite the molecular equation, but dissociate strong electrolytes into individual ions.
  - ✓ Strong electrolytes must be aqueous.
    - ➢ Solids, liquids, or gases cannot be electrolytes.
  - ✓ All soluble ionic compounds are strong electrolytes.
  - ✓ Strong acids are strong electrolytes.
    - ➢ HCl, HNO_3, H_2SO_4.
    - ➢ Weak acids are not written in the dissociated ion form.
  - ✓ Molecular compounds do not have ions, leave in the molecular form.

Ionic Equations

- Ions that are both reactants and products are called **spectator ions.**

$2K^{+1}{}_{(aq)} + 2OH^{-1}{}_{(aq)} + Mg^{+2}{}_{(aq)} + 2NO_3^{-1}{}_{(aq)} \rightarrow 2K^{+1}{}_{(aq)} + 2NO_3^{-1}{}_{(aq)} + Mg(OH)_{2(s)}$

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

$2OH^{-1}{}_{(aq)} + Mg^{+2}{}_{(aq)} \rightarrow Mg(OH)_{2(s)}$

Writing Net Ionic Equations

- First, identify the spectator ions in the complete ionic equation.
  - ✓ Identical ions on both sides of the equation.
- Cancel out the spectator ions—the result is the net ionic equation.

Summary

- 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.
- A **net ionic equation** is an equation showing only the species that actually participate in the reaction.

Practice–Write the Ionic and Net Ionic Equation.

$K_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + BaSO_4(s)$

$2K^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow 2K^+(aq) + 2NO_3^-(aq) + BaSO_4(s)$

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

$Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + CO_2(g) + H_2O(l)$

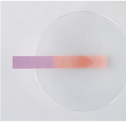
$2Na^+(aq) + CO_3^{2-}(aq) + 2H^+(aq) + 2\ Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + CO_2(g) + H_2O(l)$

$CO_3^{-2}(aq) + 2\ H^{+1}(aq) \rightarrow CO_2(g) + H_2O(l)$

Acid/Base Reactions

Properties of Acids

- Sour taste..
- React with "active" metals, not noble metals.
  - ✓ I.e., Al, Zn, Fe, but not Cu, Ag or Au.
    $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$
  - ✓ Corrosive.
- React with carbonates, producing CO_2.
  - ✓ Marble, baking soda, chalk, limestone.
    $CaCO_3 + 2\ HCl \rightarrow CaCl_2 + CO_2 + H_2O$
- React with bases to form ionic salts.
  - ✓ And often water.

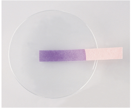


Common Acids

Chemical name	Formula	Old name	Strength
Nitric acid	HNO_3	Aqua fortis	Strong
Sulfuric acid	H_2SO_4	Vitriolic acid	Strong
Hydrochloric acid	HCl	Muriatic acid	Strong
Phosphoric acid	H_3PO_4		Moderate
Chloric acid	$HClO_3$		Moderate
Acetic acid	$HC_2H_3O_2$	Vinegar	Weak
Hydrofluoric acid	HF		Weak
Carbonic acid	H_2CO_3	Soda water	Weak
Boric acid	H_3BO_3		Weak

Properties of Bases

- Taste bitter
- Caustic
- Feel slippery
- Red Litmus = blue.
- React with acids to form ionic salts.
  - ✓ And often water.



Common Bases

Chemical name	Formula	Common name	Strength
Sodium hydroxide	NaOH	Lye, caustic soda	Strong
Potassium hydroxide	KOH	Caustic potash	Strong
Calcium hydroxide	$Ca(OH)_2$	Slaked lime	Strong
Magnesium hydroxide	$Mg(OH)_2$	Milk of magnesia	Weak
Ammonium hydroxide	NH_4OH, $\{NH_3(aq)\}$	Ammonia water, aqueous ammonia	Weak

Acid–Base Reactions

- Also called **neutralization reactions** because the acid and base neutralize each other's properties.
- In the reaction of an acid with a base, the H^{+1} from the acid combines with the OH^{-1} from the base to make water.
- The cation from the base combines with the anion from the acid to make the salt.

acid + base → salt + water

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

- The net ionic equation for an acid-base reaction often is:
  $H^{+1}(aq) + OH^{-1}(aq) \rightarrow H_2O(l)$
  - ✓ As long as the salt that forms is soluble in water.

Process for Predicting the Products of an Acid–Base Reaction

1. Determine what ions each aqueous reactant has.
2. Exchange ions.
   - ✓ (+) ion from one reactant with (-) ion from the other.
   - ✓ H^+ combines with OH^- to make water.
3. Balance charges of combined ions to get formula of the salt.
4. Balance the equation.
   - ✓ Count atoms.
5. Determine solubility of the salt.
   - ✓ Use the solubility rules.
   - ✓ If the salt is insoluble or slightly soluble, it will precipitate.

Example 7.11—Write the Molecular, Ionic, and Net-Ionic Equation for the Reaction of Aqueous Nitric Acid with Aqueous Calcium Hydroxide.

1. Write the formulas of the reactants.
   $HNO_3(aq) + Ca(OH)_2(aq) \rightarrow$
   a. Determine the ions present when each reactant dissociates.
   $(H^+ + NO_3^-) + (Ca^{2+} + OH^-) \rightarrow$
2. Exchange the ions, H^{+1} combines with OH^- to make $H_2O(l)$.
   $(H^+ + NO_3^-) + (Ca^{2+} + OH^-) \rightarrow (Ca^{+2} + NO_3^-) + H_2O(l)$

Example 7.11—Write the Molecular, Ionic, and Net-Ionic Equation for the Reaction of Aqueous Nitric Acid with Aqueous Calcium Hydroxide, Continued.

3. Write the formula of the salt product.
   - ✓ Cross charges and reduce.
   $HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2 + H_2O(l)$
4. Balance the equation.
   - ✓ May be quickly balanced by matching the numbers of H and OH to make H_2O.
   - ✓ Coefficient of the salt is always 1.
   $2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2 + 2\ H_2O(l)$

Example 7.11—Write the Molecular, Ionic, and Net-Ionic Equation for the Reaction of Aqueous Nitric Acid with Aqueous Calcium Hydroxide, Continued.

5. Determine the solubility of the salt.
   $Ca(NO_3)_2$ is soluble.
   a. Write an (*s*) after an insoluble salt and an (*aq*) after a soluble salt.

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

Example 7.11—Write the Molecular, Ionic, and Net-Ionic Equation for the Reaction of Aqueous Nitric Acid with Aqueous Calcium Hydroxide, Continued.

6. Dissociate all strong electrolytes to get complete ionic equation.
   - ✓ HNO_3 is a strong acid, $Ca(OH)_2$ and $Ca(NO_3)_2$ are ionic.
   - ✓ Not H_2O.

$2H^+_{(aq)} + 2NO_3^-{}_{(aq)} + Ca^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Ca^{2+}_{(aq)} + 2NO_{3(aq)} + H_2O_{(l)}$

7. Eliminate spectator ions to get net-ionic equation.
   $2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O\ (l)$
   $H^+(aq) + OH^-(aq) \rightarrow H_2O\ (l)$

Practice—Complete and Balance These Acid–Base Reactions.

$NH_4OH(aq) + H_2SO_4(aq) \rightarrow$

$2\ NH_4OH(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + 2\ H_2O(l)$

$Al(OH)_3(aq) + H_2SO_3(aq) \rightarrow$

$2\ Al(OH)_3(aq) + 3\ H_2SO_3(aq) \rightarrow Al_2(SO_3)_3(s) + 6\ H_2O(l)$

$Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow$

$Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ H_2O(l)$

Gas Evolution Reactions

Gas Evolution Reactions

- Reactions in which the driving force is the production of a material that escapes as a gas are called **gas evolution 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 decomposition of one of the ion exchange products into a gas and water.
  $K_2SO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2SO_3(aq)$
  $H_2SO_3 \rightarrow H_2O(l) + SO_2(g)$

Gas evolution reaction

Compounds that Undergo Gas Evolving Reactions

Reactant type	Reacting with	Ion exchange product	Decom-pose?	Gas formed	Example
$Metal_nS$, metal HS	Acid	H_2S	No	H_2S	$K_2S(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2S(g)$
$Metal_nCO_3$, metal HCO_3	Acid	H_2CO_3	Yes	CO_2	$K_2CO_3(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + CO_2(g) + H_2O(l)$
$Metal_nSO_3$, metal HSO_3	Acid	H_2SO_3	Yes	SO_2	$K_2SO_3(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + SO_2(g) + H_2O(l)$
$(NH_4)_n$anion	Base	NH_4OH	Yes	NH_3	$KOH(aq) + NH_4Cl(aq) \rightarrow KCl(aq) + NH_3(g) + H_2O(l)$

Process for Predicting the Products of a Gas-Evolving Reaction

1. Determine what ions each aqueous reactant has.
2. Exchange ions.
   - ✓ (+) ion from one reactant with (-) ion from the other.
3. Balance charges of combined ions to get formula of each product.
4. Check to see if either product is H_2S.
5. Check to see if either product decomposes. If so, rewrite as $H_2O(l)$ and a gas.
   - ✓ See Table 7.4
6. Balance the equation.
7. Determine solubility of other product in water.

Example—When an Aqueous Solution of Sodium Sulfite Is Added to an Aqueous Solution of Nitric Acid, a Gas Evolves.

1. Write the formulas of the reactants.

$$Na_2SO_3(aq) + HNO_3(aq) \rightarrow$$

   a. Determine the ions present when each reactant dissociates.

$$(Na^{+1} + SO_3^{-2}) + (H^{+1} + NO_3^{-1}) \rightarrow$$

2. Exchange the ions.

$$(Na^{+1} + SO_3^{-2}) + (H^{+1} + NO_3^{-1}) \rightarrow (Na^{+1} + NO_3^{-1}) + (H^{+1} + SO_3^{-2})$$

Example—When an Aqueous Solution of Sodium Sulfite Is Added to an Aqueous Solution of Nitric Acid, a Gas Evolves, Continued.

3. Write the formulas of the products.
   - ✓ Cross charges and reduce.

$$Na_2SO_3(aq) + HNO_3(aq) \rightarrow NaNO_3 + H_2SO_3$$

4. Check to see if either product is H_2S. No.
5. Check to see of either product decomposes. Yes.
   - ✓ H_2SO_3 decomposes into $SO_2(g) + H_2O(l)$.

$$Na_2SO_3(aq) + HNO_3(aq) \rightarrow NaNO_3 + SO_2(g) + H_2O(l)$$

Example—When an Aqueous Solution of Sodium Sulfite Is Added to an Aqueous Solution of Nitric Acid, a Gas Evolves, Continued.

6. Balance the equation.

$$Na_2SO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ NaNO_3 + SO_2(g) + H_2O(l)$$

7. Determine the solubility of other product.

$NaNO_3$ is soluble.

   a. Write an (*s*) after the insoluble products and an (*aq*) after the soluble products.

$$Na_2SO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ NaNO_3(aq) + SO_2(g) + H_2O(l)$$

Practice—Complete the Following Reactions.

$PbS(s) + H_2SO_4(aq) \rightarrow$

$PbS(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2S(g)$

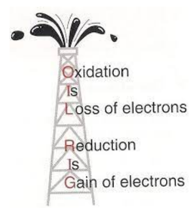
$HNO_3(aq) + NaHCO_3(aq) \rightarrow$

$HNO_3(aq) + NaHCO_3(aq) \rightarrow NaNO_3(aq) + CO_2(g) + H_2O(l)$

Oxidation Reduction Reactions

Oxidation–Reduction Reactions

- Redox reactions occur when one chemical species loses one or more electrons to another.
- We say that the element that loses electrons in the reaction is **oxidized**.
- And the substance that gains electrons in the reaction is **reduced**.
- You cannot have one without the other.
- In combustion, the O atoms in O_2 are reduced, and the non-O atoms in the other material are oxidized.



Reactions of Metals with Nonmetals (Oxidation–Reduction)

- Metals react with nonmetals to form ionic compounds.
- The metal loses electrons and becomes a cation.
  - ✓ The metal undergoes **oxidation**.
- The nonmetal gains electrons and becomes an anion.
  - ✓ The nonmetal undergoes **reduction**.
- In the reaction, electrons are transferred from the metal to the nonmetal.

Ionic Compound Formation as Redox

- In the reaction:

$$Mg + Cl_2 \rightarrow MgCl_2$$

- The magnesium atoms are oxidized.

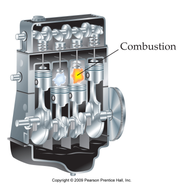
$$Mg \rightarrow Mg^{2+} + 2\ e^-$$

- The chlorine atoms are reduced.

$$Cl + 1\ e^- \rightarrow Cl^-$$

Combustion Reactions

- Reactions in which $O_2(g)$ is a reactant are called **combustion reactions**.
- Combustion reactions release lots of energy. They are **exothermic.**
- Combustion reactions are a subclass of oxidation–reduction reactions.



$$2\ C_8H_{18}(g) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

Combustion as Redox

- In the following reaction:

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

- The magnesium atoms are oxidized.

$$Mg \rightarrow Mg^{2+} + 2\ e^-$$

- The oxygen atoms are reduced.

$$O + 2\ e^- \rightarrow O^{2-}$$

Combustion as Redox

- Even though the following reaction does not involve ion formation, electrons are still transferred.

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

- The carbon atoms are oxidized.

$$C^{-4} \rightarrow C^{+4} + 8\ e^-$$

  - ✓ These are not charges, they are called **oxidation numbers**, but they help us see the electron transfer.
- The oxygen atoms are reduced.

$$O + 2\ e^- \rightarrow O^{-2}$$

Recognizing Redox Reactions

- Any reaction where O_2 is a reactant or a product is a redox reaction.
- Any reaction between a metal and a nonmetal is redox.
- Any reaction where electrons are transferred is redox.
  - ✓ When a free element gets combined into a compound, it will be either oxidized or reduced.

$$N_2(g) + H_2(g) \rightarrow NH_3(g)$$

  - ✓ When a metal cation changes its charge, it will be either oxidized if its charge increases or reduced if its charge decreases.

$$CuCl(aq) + FeCl_3(aq) \rightarrow FeCl_2(aq) + CuCl_2(aq)$$

Practice—Decide Whether Each of the Following Reactions Is a Redox Reaction.

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

Yes, metal + nonmetal.

$CaSO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + SO_2(g) + H_2O(l)$

No, this is a gas evolving reaction.

$Fe_2O_3(s) + C(s) \rightarrow 2\ Fe(s) + 3\ CO(g)$

Yes, the Fe is reduced and the C gets combined.

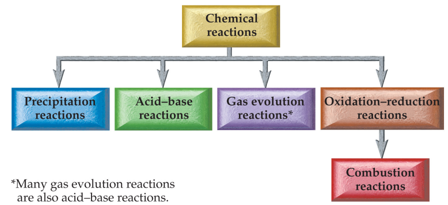
$SO_2(g) + O_2(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

Yes, O_2 reactant.

Classifying Reactions

Classifying Reactions

- One way is based on the process that happens.
  - ✓ Precipitation, neutralization, formation of a gas, or transfer of electrons.

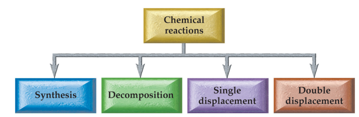


*Many gas evolution reactions are also acid–base reactions.

Classifying Reactions, Continued

- Another scheme classifies reactions by what the atoms do.

Type of reaction	General equation
Synthesis	$A + B \rightarrow AB$
Decomposition	$AB \rightarrow A + B$
Displacement	$A + BC \rightarrow AC + B$
Double displacement	$AB + CD \rightarrow AD + CB$



Synthesis Reactions

- Also known as **composition** or **combination** reactions.
- Two (or more) reactants combine together to make **one product**.
  - ✓Simpler substances combining together.

$$2\,CO + O_2 \rightarrow 2\,CO_2$$
$$2\,Mg + O_2 \rightarrow 2\,MgO$$
$$HgI_2 + 2\,KI \rightarrow K_2HgI_4$$

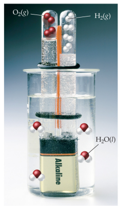
Decomposition Reactions

- A large molecule is broken apart into smaller molecules or its elements.
  - ✓Caused by addition of energy into the molecule.
- **Have only one reactant**, make 2 or more products.

$$2\,FeCl_3(s) \xrightarrow{elec} 2\,FeCl_2(l) + Cl_2$$
$$2\,HgO(s) \xrightarrow{\Delta} 2\,Hg(l) + O_2$$
$$2\,O_3 \xrightarrow{h\upsilon} 3\,O_2$$

Decomposition of Water

$$2\,H_2O(l) \xrightarrow{elec} 2\,H_2(g) + O_2(g)$$



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

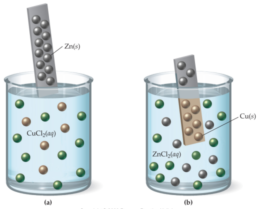
Single Displacement Reactions

- Reactions that involve one atom displacing another and replacing it in a compound.
- In the reaction $Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$, the atom Zn displaces H from the compound.
- Other examples of displacement reactions are:

$$Fe_2O_3(s) + Al(s) \rightarrow Fe(s) + Al_2O_3(s)$$
$$2\,Na(s) + 2\,H_2O(aq) \rightarrow 2\,NaOH(aq) + H_2(g)$$

Displacement of Copper by Zinc

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



Double Displacement Reactions

- Two ionic compounds exchange ions.
  - $X^{\oplus}Y^{\ominus}(aq) + A^{\oplus}B^{\ominus}(aq) \rightarrow X^{\oplus}B^{\ominus} + A^{\oplus}Y^{\ominus}$
- May be followed by decomposition of one of the products to make a gas.
- Precipitation, acid–base, and gas evolving reactions are also double displacement reactions.

Examples of Double Displacement

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\,NaCl(aq)$$

$$Ba(OH)_2(s) + 2\,HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)$$

$$Li_2CO_3(aq) + 2\,HCl(aq) \rightarrow 2\,LiCl(aq) + CO_2(g) + H_2O(l)$$

Practice—Classify the Following Reactions as Synthesis, Decomposition, Single Displacement, or Double Displacement.

$$3\,Mg(s) + 2\,FeCl_3(aq) \rightarrow 3\,MgCl_2(aq) + 2\,Fe(s)$$

Single Displacement

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

Synthesis

$$3\,KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3\,H_2O(l)$$

Double Displacement

$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

Decomposition