What is a Chemical Equation?

A _________________________ is a written representation of the process that occurs in a chemical reaction. A chemical equation is written with the _____________ (starting materials) on the left side of an arrow and the _______________ (resulting substance) of the chemical reaction on the right side of the equation. The head of the arrow typically points toward the right or toward the product side of the equation, although reactions may indicate equilibrium with the reaction proceeding in __________ directions simultaneously.

The elements in an equation are denoted using their symbols. _____________________ next to the symbols indicate the __________________________ numbers. Subscripts are used to indicate the number of atoms of an element present in a chemical species.

An example of a chemical equation may be seen in the combustion of methane:

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

Balancing Equations Notes

An equation for a chemical reaction in which the number of atoms for each element in the reaction and the total charge are the same for both the reactants and the products. In other words, the mass and the charge are balanced on both sides of the reaction.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
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<tbody>
<tr>
<td>+</td>
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<td>\rightarrow</td>
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<td>(s)</td>
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<td>(aq)</td>
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<td>(l)</td>
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<td>△</td>
<td></td>
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</tbody>
</table>
LAW OF CONSERVATION OF MASS

In all chemical equations the _____________________________ must be met. Matter can NOT be created nor destroyed in a chemical reaction.

Remember, in a chemical reaction, the atoms/ions are simply ____________________ to form new substances.

Therefore, chemical equations ____________________________.

WHAT IS A “BALANCED” CHEMICAL EQUATION?
A balanced chemical equation is one in which __________________ of the equation has the ____________________________ of atoms/ions of each element.

Example: \[ \text{Al (s)} + \text{O}_2 (\text{g}) \rightarrow \text{Al}_2\text{O}_3 (\text{s}) \]

Not Balanced

\[
\begin{array}{cc}
1 & \text{Al} \\
2 & \text{O} \\
2 & \text{Al} \\
3 & \text{O} \\
\end{array}
\]

Balanced

\[
\begin{array}{ccc}
4 & \text{Al} & \text{Al} \\
6 & \text{O} & \text{O} \\
\end{array}
\]

RULES FOR BALANCING CHEMICAL EQUATIONS

1. Write the _____________ chemical formulas for ____ of the reactants and the products.

2. Write the formulas of the _____________ on the ________ of the reaction arrow; write the formulas of the ________________ on the _____________ of the reaction arrow.

3. COUNT the total number of atoms/ions of each element in the reactants and the total number of atoms/ions of each element in the products.

** A ______________________ that appears ______________________ on both sides of the equation is counted as a single unit.

4. Balance the elements one at a time using ________________.

A coefficient is a small WHOLE number that is written ________________ of a chemical formula in a chemical equation.

When _____________ coefficient is written, the coefficient is assumed to be ___________.

It is best to begin with elements ________________ hydrogen and oxygen. These elements often occur more than twice in equations.
** You must __________ attempt to balance the equation by changing subscripts in chemical formulas!!!!!

5. Check each atom/ion, or polyatomic ion to be sure that the equation is ________________________.

6. Finally, make sure that all of the coefficients are in the ________________ possible whole number ratios. (At least one of the coefficients must be a prime number!)

Use coefficients to make sure the number of atoms is the same on both sides of the equation.

1. ___ H₂ + ___ O₂ → ___ H₂O

2. ___ HCl + ___ Zn → ___ ZnCl₂ + ___ H₂

3. ___ Al + ___ CaS → ___ Al₂S₃ + ___ Ca

4. Write the skeleton equation for the reaction of solid Iron and gaseous chlorine react to produce a solid iron (III) chloride

Diatomic Elements

<table>
<thead>
<tr>
<th>Diatomic Elements are always diatomic (written with a subscribe of 2) when they are in their elemental form</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. 5.</td>
</tr>
<tr>
<td>2. 6.</td>
</tr>
<tr>
<td>3. 7.</td>
</tr>
<tr>
<td>4.</td>
</tr>
</tbody>
</table>
Types of Chemical Reactions Notes

- **Synthesis** - _________________ elements or compounds combine to form one compound.
- **Decomposition** - a _________________ compound decomposes into two or more elements or smaller compounds.
- **Single Replacement** - a metal will _________________ a less active metal in an ionic compound OR a nonmetal will replace a less active nonmetal.
- **Double Replacement** - the metals in ionic compounds _________________ places.
- **Combustion** - an _________________ compound containing carbon, hydrogen and sometimes oxygen reacts with oxygen gas to form carbon dioxide and water.

**Examples**

\[ \text{Cu}(s) + \text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag}(s) \]

Single Replacement

\[ \text{C}_6\text{H}_{12}\text{O}_6(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

Combustion

\[ \text{Fe}(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \]

Synthesis

\[ \text{CuCl}_2(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{PbCl}_2(s) \]

Double Replacement

\[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]

Decomposition

- **Combination** - Definition - two or more substances react to form 1 product. Usually releases energy, _________________.

Combustion reactions that contain oxygen as a reactant can also be considered combustion.

\[ \text{A} + \text{X} \rightarrow \text{AX} \]

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

\[ \text{CaO} (s) + \text{H}_2\text{O} (l) \rightarrow \text{Ca(OH)}_2 (s) \]

One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:

\[ 8 \text{Fe} + \text{S}_8 \rightarrow 8 \text{FeS} \]

- **Decomposition** - Definition - A single compound breaks down into 2 or more elements or compounds

\[ \text{AX} \rightarrow \text{A} + \text{X} \]

\[ 2\text{NaN}_3(s) \rightarrow 2\text{Na}(s) + 3\text{N}_2 (g) \]

\[ 2\text{KClO}_3 (s) \rightarrow 2\text{KCl} (s) + 3\text{O}_2 (g) \]

\[ \text{CaCO}_3 (s) \rightarrow \text{CaO} (s) + \text{CO}_2 (g) \]
These reactions often require an energy source as an initiator. Energy sources can be heat, light, or electricity. They are usually _____________________.

One example of a decomposition reaction is the electrolysis of water to make oxygen and hydrogen gas:

\[ 2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2 \]

Definition - Oxygen gas combines with a substance and releases energy in the form of light or heat. So combustion reactions are usually exothermic. Combination reactions that contain oxygen as a reactant can also be considered combustion.

\[ A + \text{O}_2 \rightarrow \]
\[ \text{C(s)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{energy} \]
\[ 4 \text{Fe (s)} + 3\text{O}_2 (g) \rightarrow 2\text{Fe}_2\text{O}_3 (s) + \text{energy} \]

For hydrocarbons:
\[ C_x\text{H}_y + [x + (y/4)] \text{O}_2 \rightarrow x\text{CO}_2 + (y/2)\text{H}_2\text{O} \]
\[ C_3\text{H}_8(g) + \text{O}_2(g) \rightarrow \text{CO}_2 (g) + \text{H}_2\text{O} (g) + \text{light} + \text{heat} \]

Definition - one ion replaces another in a compound.
\[ \text{AB} + \text{C} \rightarrow \text{AC} + \text{B} \]

One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

\[ \text{Mg} + 2 \text{H}_2\text{O} \rightarrow \text{Mg(OH)}_2 + \text{H}_2 \]

Definition - two ions replace each other or switch places in compounds.
\[ \text{AB} + \text{CD} \rightarrow \text{AC} + \text{BD} \]

One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

\[ \text{Pb(NO}_3)_2 + 2 \text{KI} \rightarrow \text{PbI}_2 + 2 \text{KNO}_3 \]
## Sample Problems (the solutions are in the next section)

<table>
<thead>
<tr>
<th>List the type of the following reactions.</th>
<th>Solutions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) NaOH + KNO₃ --&gt; NaNO₃ + KOH</td>
<td>1)</td>
</tr>
<tr>
<td>2) CH₄ + 2 O₂ --&gt; CO₂ + 2 H₂O</td>
<td>2)</td>
</tr>
<tr>
<td>3) 2 Fe + 6 NaBr --&gt; 2 FeBr₃ + 6 Na</td>
<td>3)</td>
</tr>
<tr>
<td>4) CaSO₄ + Mg(OH)₂ --&gt; Ca(OH)₂ + MgSO₄</td>
<td>4)</td>
</tr>
<tr>
<td>5) NH₄OH + HBr --&gt; H₂O + NH₄Br</td>
<td>5)</td>
</tr>
<tr>
<td>6) Pb + O₂ --&gt; PbO₂</td>
<td>6)</td>
</tr>
<tr>
<td>7) Na₂CO₃ --&gt; Na₂O + CO₂</td>
<td>7)</td>
</tr>
</tbody>
</table>

## Practice Balancing Chemical Equations Worksheet

**Balance the following chemical equations using coefficients**

1. Al(OH)₃(s) + HCl (aq) \rightarrow AlCl₃ (aq) + H₂O (l)
2. Fe₂O₃ (s) + CO (g) \rightarrow Fe₃O₄(s) + CO₂ (g)
3. FeO (s) + O₂ (g) \rightarrow Fe₂O₃ (s)
4. C₆H₆ (l) + O₂ (g) \rightarrow CO₂ (g) + H₂O (g)
5. Ca(OH)₂ (aq) + H₃PO₄ (aq) \rightarrow H₂O (l) + Ca₃(PO₄)₂ (s)
6. I₄O₉ (s) \rightarrow I₂O₆(s) + I₂ (s) + O₂ (g)
7. Eu (s) + HF (g) \rightarrow EuF₃ (s) + H₂ (g)
8. NaHCO₃ (aq) + C₆H₈O₇ (aq) \rightarrow CO₂ (g) + H₂O (l) + Na₃C₆H₅O₇ (aq)
9. Ni (s) + CO (g) \rightarrow Ni(CO)₄ (g)
10. K₂PtCl₄ (aq) + NH₃ (aq) \rightarrow Pt(NH₃)₂Cl₂ (s) + KCl (aq)
Practice
Write the following chemical equations and balance using coefficients.

1. Liquid mercury reacts with liquid bromine to produce solid mercury (II) Bromide

2. Solid calcium carbonate decomposes upon heating to produce solid calcium oxide and carbon dioxide gas.

3. Solid calcium will react with liquid water to produce aqueous calcium hydroxide and hydrogen gas.

4. Butane gas (C\textsubscript{4}H\textsubscript{10}) will react with oxygen gas to produce carbon dioxide gas and water vapor.

5. Solid aluminum will react with oxygen gas to produce solid aluminum oxide.

6. Aluminum metal is oxidized by oxygen (from the air) to form aluminum oxide.

7. Sodium oxide reacts with carbon dioxide to form sodium carbonate.

8. Calcium metal reacts with water to form calcium hydroxide and hydrogen gas.

9. Potassium nitrate decomposes to form potassium nitrite and oxygen.

10. Barium metal reacts with Iron (III) sulfate to produce barium sulfate and iron metal.

11. Barium chloride reacts with sodium sulfate to produce barium sulfate and sodium chloride.
Practice

TYPES OF CHEMICAL REACTIONS

Directions
(a) Write and balance the given equation.
(b) Indicate the type of chemical reaction represented.

1. Iron reacts with oxygen gas to produce Iron (III) oxide.
   (a)
   (b)

2. Propane (C₃H₈) reacts with oxygen gas to produce carbon dioxide and water.
   (a)
   (b)

3. Bromine gas reacts with potassium iodide to produce potassium bromide and iodine gas.
   (a)
   (b)

4. Hydrogen peroxide will produce water and oxygen gas if left in sunlight.
   (a)
   (b)

5. Phosphorous reacts with oxygen gas to produce tetraphosphorous decoxide.
   (a)
   (b)

6. Iron (III) Chloride reacts with sodium hydroxide to produce Iron (III) hydroxide and sodium chloride.
   (a)
   (b)

7. Iron (III) oxide reacts with hydrogen gas to produce iron and water.
   (a)
   (b)
8. Octane ($C_8H_{18}$) reacts with oxygen gas to produce carbon dioxide and water.

(a)

(b)

9. Calcium carbonate reacts with aluminum phosphate to produce calcium phosphate and aluminum carbonate.

(a)

(b)

10. Aluminum hydroxide decomposes to produce aluminum oxide and water.

(a)

(b)

11. Zinc reacts with silver nitrate to produce zinc nitrate and silver.

(a)

(b)

12. Glucose ($C_6H_{12}O_6$) reacts with oxygen gas to produce carbon dioxide and water.

(a)

(b)

13. Potassium oxide reacts with water to produce potassium hydroxide.

(a)

(b)

14. Lead (IV) oxide decomposes into lead (II) oxide and oxygen gas.

(a)

(b)

15. Hydrochloric acid (hydrogen chloride) reacts with barium hydroxide to produce water and barium chloride.

(a)

(b)
More Practice
Change the coefficients to make the number of atoms of each element equal on both sides of the equation

1. Calcium metal reacts with water to form solid calcium hydroxide and hydrogen gas.

2. Zinc hydroxide solution reacts with lithium to form lithium hydroxide solution and zinc metal.

3. Liquid propanol (C₃H₇OH) reacts with oxygen gas to form carbon dioxide gas and water vapor.

4. Aluminum metal reacts with oxygen gas to form solid aluminum oxide.

5. Liquid carbonic acid (hydrogen carbonate) decomposes into carbon dioxide gas and water.

6. Lead (II) nitrate solution reacts with iron (III) chloride solution to form solid lead (II) chloride and Iron (III) nitrate solution.

7. Aluminum metal reacts with silver sulfate solution to form aluminum sulfate solution and silver metal.
8. Methane gas (CH₄) reacts with oxygen gas to form carbon dioxide gas and water vapor.

9. Iron metal reacts with bromine gas to form iron (III) bromide solid.


Rules for Predicting Products of Chemical Reactions

1. Hydrocarbon + O₂ ➝ CO₂ + H₂O (Combustion)
   a. 2C₄H₁₀ + 13 O₂ ➝ 8 CO₂ + 10 H₂O

2. Metal Carbonate ➝ Metal Oxide + CO₂ (Decomposition)
   a. MgCO₃ ➝ MgO + CO₂
   b. Synthesis: Metal Oxide + CO₂ ➝ Metal Carbonate

3. Metal Sulfites ➝ Metal Oxide + SO₂ (Decomposition)
   a. CaSO₃ ➝ CaO + SO₂
   b. Synthesis: Metal Oxide + SO₂ ➝ Metal Sulfite

4. Metal Hydride + H₂O ➝ Metal Hydroxide + H₂ (Double Replacement)
   a. KH + H₂O ➝ KOH + H₂

5. Metal + H₂O ➝ Metal Hydroxide + H₂ (Single Replacement)
   a. 2Na + 2H₂O ➝ 2NaOH + H₂

6. Metal Oxide + H₂O ➝ Metal Hydroxide (Synthesis)
   a. MgO + H₂O ➝ Mg(OH)₂

7. Non-metal oxide + H₂O ➝ ternary acid (Synthesis)
   a. N₂O₃ + H₂O ➝ 2 HNO₂
   b. N₂O₅ + H₂O ➝ 2 HNO₃
   c. CO₂ + H₂O ➝ H₂CO₃
Predicting Products: Write the COMPLETE balanced equation

1. Hydrochloric acid (HCl) reacts with sodium hydroxide.

2. Sodium reacts with oxygen

3. Mercury (II) oxide

4. Zinc reacts with lead (II) Nitrate

5. Silver nitrate reacts with calcium chloride

6. \( \text{C}_7\text{H}_{16} \) reacts with oxygen

7. \( \text{CH}_3\text{OH} \) reacts with oxygen

8. Magnesium reacts with Fluorine

9. Copper (II) Chloride

10. Aluminum reacts with Calcium Sulfide

11. Potassium Hydroxide reacts with Zinc Chloride

12. \( \text{C}_2\text{H}_2 \) reacts with oxygen

13. Sodium Iodide reacts with chlorine

14. Aluminum reacts with sulfur
Balance the following equations:

\[ _____ \text{HCl} + _____ \text{Al} \rightarrow _____ \text{AlCl}_3 + _____ \text{H}_2 \]

\[ _____ \text{HNO}_3 + _____ \text{Zn} \rightarrow _____ \text{Zn(NO}_3)_2 + _____ \text{NO}_2 + _____ \text{H}_2\text{O} \]

\[ _____ \text{HNO}_3 + _____ \text{Sn} \rightarrow _____ \text{SnO}_2 + _____ \text{NO}_2 + _____ \text{H}_2\text{O} \]

Write the balanced equation for the following chemical reaction AND the type of reaction that has occurred.

Sodium metal is added to sulfuric acid (\(\text{H}_2\text{SO}_4\)). Hydrogen gas is produced, along with sodium sulfate.

White phosphorus (\(\text{P}_4\)) reacts with chlorine to make phosphorus trichloride.

Magnesium chlorate, \(\text{Mg(ClO}_3)_2\), is heated strongly until it decomposes into magnesium chloride and oxygen gas.

Write the balanced equation for the following chemical reaction (you must predict the product) AND the type of reaction that has occurred.

Butane gas (\(\text{C}_4\text{H}_{10}\)) is burned completely in air.

Iron metal is added to bromine- Iron (III) product.

Calcium metal is added to phosphoric acid (\(\text{H}_3\text{PO}_4\)).

Aluminum metal is added to a solution of iron (III) chloride.

A piece of iron metal is exposed to oxygen gas (\(\text{Fe}^{3+}\) product would form).

Calcium metal is added to water.