Types of Equations

It is most important for a chemist to be able to write correctly balanced equations and to interpret equations written by others. It is also very helpful if he/she knows how to predict the products of certain specific types of reactions.

Purpose: This document is intended to help you, the chemistry student, learn the basics of writing and balancing equations, how to predict the products of five general types of reactions and including how to write and balance equations for the combustion of hydrocarbons.

General Information | Combination Reactions | Decomposition Reactions | Single Replacement Reactions | Double Replacement Reactions | Combustion Reactions | Oxidation-Reduction Reactions

I. Formulas show chemistry at a standstill. Equations show chemistry in action.

A. Equations show:

1. the reactants which enter into a reaction.
2. the products which are formed by the reaction.
3. the amounts of each substance used and each substance produced.

B. Two important principles to remember:

1. Every chemical compound has a formula which cannot be altered.
2. A chemical reaction must account for every atom that is used. This is an application of the Law of Conservation of Matter which states that in a chemical reaction atoms are neither created nor destroyed.

C. Some things to remember about writing equations:

1. The diatomic elements when they stand alone are always written H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2
2. The sign, →? , means "yields" and shows the direction of the action.
3. A small delta, (Δ), above the arrow shows that heat has been added.
4. A double arrow, ↔ , shows that the reaction is reversible and can go in both directions.
5. Before beginning to balance an equation, check each formula to see that it is correct. **NEVER** change a formula during the balancing of an equation.

6. Balancing is done by placing coefficients in front of the formulas to insure the same number of atoms of each element on both sides of the arrow.

   Practice Balancing Equations

7. Always consult the **Activity Series** of metals and nonmetals before attempting to write equations for replacement reactions.
8. If a reactant or product is a solid, (s) is placed after the formula.
9. If a reactant or product is a gas, (g) is placed after it.
10. If a reactant or product is in water solution, (aq) is placed after it.
11. Some products are unstable and break down (decompose) as they are produced during the reaction. You need to be able to recognize these products when they occur and write the decomposition products in their places.

Examples:

- $H_2CO_{3(aq)} \rightarrow ?$ $H_2O_{(l)} + CO_{2(g)}$

  Carbonic acid, as in soft drinks, decomposes when it is formed.
- $H_2SO_{3(aq)} \rightarrow ?$ $H_2O_{(l)} + SO_{2(g)}$

  Sulfurous acid also decomposes as it is formed.
- $NH_4OH_{(aq)} \rightarrow ?$ $NH_{3(g)} + H_2O_{(l)}$

  You can definitely smell the odor of ammonia gas because whenever "ammonium hydroxide" is formed it decomposes into ammonia and water.

D. Rules for writing equations.

1. Write down the formula(s) for any substance entering into the reaction. Place a plus (+) sign between the formulas as needed and put the yield arrow after the last one.
2. Examine the formulas carefully and decide which of the four types of equations applies to the reaction you are considering. On the basis of your decision, write down the correct formulas for all products formed, placing them to the right of the arrow.

II. Five basic types of chemical reactions:

A. Combination (Synthesis):

- two or more elements or compounds may combine to form a more complex compound.
- **Basic form: A + X →? AX**

Examples of combination reactions:

1. Metal + oxygen →? metal oxide

   EX. $2Mg_{(s)} + O_{2(g)}$ →? $2MgO_{(s)}$

2. Nonmetal + oxygen →? nonmetallic oxide

   EX. $C_{(s)} + O_{2(g)}$ →? $CO_{2(g)}$

3. Metal oxide + water →? metallic hydroxide

   EX. $MgO_{(s)} + H_2O_{(l)}$ →? $Mg(OH)_{2(s)}$

4. Nonmetallic oxide + water →? acid

   EX. $CO_{2(g)} + H_2O_{(l)}$ →? $H_2CO_{3(aq)}$

- Metal + nonmetal →? salt

  EX. 2 $Na_{(s)} + Cl_{2(g)}$ →? $2NaCl_{(s)}$

- A few nonmetals combine with each other.

  EX. $2P_{(s)} + 3Cl_{2(g)}$ →? $2PCl_{3(g)}$

These two reactions must be remembered:

- $N_{2(g)} + 3H_{2(g)}$ →? $2NH_{3(g)}$
- $NH_{3(g)} + H_2O_{(l)}$ →? $NH_4OH_{(aq)}$

B. Decomposition:

A single compound breaks down into its component parts or simpler compounds.

Basic form: AX →? A + X

Examples of decomposition reactions:

Metallic carbonates, when heated, form metallic oxides and $CO_{2(g)}$.

EX. $CaCO_{3(s)}$ →? $CaO_{(s)} + CO_{2(g)}$

- Most metallic hydroxides, when heated, decompose into metallic oxides and water.

  EX. $Ca(OH)_{2(s)}$ →? $CaO_{(s)} + H_2O_{(g)}$

- Metallic chlorates, when heated, decompose into metallic chlorides and oxygen.

  EX. $2KClO_{3(s)}$ →? $2KCl_{(s)} + 3O_{2(g)}$

- Some acids, when heated, decompose into nonmetallic oxides and water.

EX. $H_2SO_4 \rightarrow$? $H_2O_{(l)} + SO_{3(g)}$

- Some oxides, when heated, decompose.

EX. $2HgO_{(s)} \rightarrow$? $2Hg_{(l)} + O_{2(g)}$

- Some decomposition reactions are produced by electricity.

EX. $2H_2O_{(l)} \rightarrow$? $2H_{2(g)} + O_{2(g)}$

EX. $2NaCl_{(l)} \rightarrow$? $2Na_{(s)} + Cl_{2(g)}$

C. Single Replacement:

a more active element takes the place of another element in a compound and sets the less active one free.

Basic form: A + BX →? AX + B or AX + Y →? AY + X

Examples of single replacement reactions:

Replacement of a metal in a compound by a more active metal.

EX. $Fe_{(s)} + CuSO_{4(aq)} \rightarrow$? $FeSO_{4(aq)} + Cu_{(s)}$

- Replacement of hydrogen in water by an active metal.

EX. $2Na_{(s)} + 2H_2O_{(l)} \rightarrow$? $2NaOH_{(aq)} + H_{2(g)}$

EX. $Mg_{(s)} + H_2O_{(g)} \rightarrow$? $MgO_{(s)} + H_{2(g)}$

- Replacement of hydrogen in acids by active metals.

EX. $Zn_{(s)} + 2HCl_{(aq)} \rightarrow$? $ZnCl_{2(aq)} + H_{2(g)}$

- Replacement of nonmetals by more active nonmetals.

EX. $Cl_{2(g)} + 2NaBr_{(aq)} \rightarrow$? $2NaCl_{(aq)} + Br_{2(l)}$

NOTE: Refer to the **activity series for metals and hydrogen**, and nonmetals (the halogens, group 7A) to predict products of replacement reactions. If the free element is above the element to be replaced in the compound, then the reaction will occur. If it is below, then no reaction occurs.

D. Double Replacement (Ionic):

- occurs between ions in aqueous solution. A reaction will occur when a pair of ions come together to produce at least one of the following:
  1. a precipitate
  2. a gas
  3. water or some other non-ionized substance.
- Basic form: AX + BY →? AY + BX

Examples of double replacement reactions:

1. Formation of precipitate.

   EX. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow$? $NaNO_{3(aq)} + AgCl_{(s)}$

   EX. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow$? $2NaCl_{(aq)} + BaSO_{4(s)}$

2. Formation of a gas.

   EX. $HCl_{(aq)} + FeS_{(s)} \rightarrow$? $FeCl_{2(aq)} + H_2S_{(g)}$

- Formation of water. (If the reaction is between an acid and a base it is called a neutralization reaction.)

  EX. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow ?$ $NaCl_{(aq)} + H_2O_{(l)}$

- Formation of a product which decomposes.

  EX. $CaCO_{3(s)} + HCl_{(aq)} \rightarrow ?$ $CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

NOTE: Use the **solubility rules or the solubility table** to decide whether a product of an ionic reaction is insoluble in water and will thus form a precipitate. If a compound is soluble in water then it should be shown as being in aqueous solution, or left as separate ions. It is, in fact, often more desirable to show only those ions that are actually taking part in the actual reaction. Equations of this type are called **net ionic equations**.

E. Combustion of Hydrocarbons:

Another important type of reaction, in addition to the four types above, is that of the combustion of a hydrocarbon. When a hydrocarbon is burned with sufficient oxygen supply, the products are always carbon dioxide and water vapor. If the supply of oxygen is low or restricted, then carbon monoxide will be produced. This is why it is so dangerous to have an automobile engine running inside a closed garage or to use a charcoal grill indoors.

Hydrocarbon (C_xH_y) + $O_{2(g)} \rightarrow ?$ $CO_{2(g)} + H_2O_{(g)}$

EX. $CH_{4(g)} + 2O_{2(g)} \rightarrow ?$ $CO_{2(g)} + 2H_2O_{(g)}$

EX. $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow ?$ $8CO_{2(g)} + 10H_2O_{(g)}$

NOTE:

Complete combustion means the higher oxidation number is attained, ie., CO_2 is formed.

Incomplete combustion means the lower oxidation number is attained, ie., CO is formed.

The phrase "To burn" means to add oxygen unless told otherwise.

Combustion of Other Things:

Remember: Anytime you have a reaction with oxygen it might also be considered a combustion reaction. Therefore, some reactions can be classified under two types. (ie., some of the examples under Combination Reactions above)

III. Oxidation - Reduction Reactions

When a metal reacts with a nonmetal, an ionic compound is formed. The ions are formed when the metal transfers one or more electrons to the nonmetal, the metal becoming a cation and the nonmetal atom becoming an anion. Therefore, a metal-nonmetal reaction can often be assumed to be an oxidation-reduction reaction, which involves electron transfer.

Two metals can also undergo an oxidation-reduction reaction. At this point we can recognize these reactions only by looking for O_2 as a reactant or product. When two nonmetals react, the compound formed is not ionic.

EX. $Na + Cl_2 \rightarrow$? $NaCl$

EX. $Al + Fe_2O_3 \rightarrow$? $Fe + Al_2O_3$

EX. $SO_2 + O_2 \rightarrow$? SO_3

It should be obvious that some reactions can be considered as more than one kind: for example, a combination reaction can also be an oxidation-reduction reaction. Some reactions could even be considered to be three kinds of reaction.