5 Chemical Reactions: An introduction

Indicators that a chemical reaction has occurred, on mixing chemicals, include the following:

- There is a colour change
- A solid forms on mixing two liquids.
- A gas evolves on mixing two substances (the mixture fizzes).
- A change of temperature occurs.

In consequence, whenever two substances are mixed, close observation for the above changes and any others are necessary.

5.1 Chemical Equations

Chemical reactions always involve breaking bonds between atoms, regrouping and rearranging the atoms and forming new bonds. Without going into specific details of the actual process, a chemical equation informs us of a net overall chemical change. It can generically be written as:

Reactants $\rightarrow$ Products

The reactants are the chemicals that are mixed at the beginning of the chemical process and are written on the left hand side of the chemical equation. The chemical bonds of reactant molecules are broken during the course of the reaction. The atoms are then regrouped and new bonds are formed. The new molecules formed through the forging of the new bonds are called the products.

e.g. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

This equation reads hydrogen gas reacts with chlorine gas to form hydrogen chloride. Note that the physical form ($g =$ gas; $s =$ solid; $l =$ liquid; $aq =$ dissolved in water) in which the molecule is found is written in brackets beside each molecule. The arrow reads “reacts to form...”. Note that the “2” in front of the HCl informs us that two molecules of HCl are produced each time one molecule of hydrogen reacts with one molecule of chlorine. In this equation, the following sequence of steps can be envisioned:
5.2 Balancing Chemical Equations

In a chemical reaction, atoms are neither created nor destroyed. The total number of each type of atom that is present at the start of the reaction must be there at the end. That is, the number of each type of atom on the reactant side of the chemical equation must equal the number of each type of atom on the product side of the equation. A balanced chemical equation reflects this condition. The act of balancing a chemical equation is to ensure that the number of each atom type is conserved.

In the above example, we see that the reactant side consists of 2 atoms of hydrogen and 2 atoms of chlorine. So does the right hand (product) side. The difference between the two sides reflects a change in the way the atoms are grouped into different molecules. Notice also that if you are asked for the formula of the compound that is formed when hydrogen atoms are bonded to chlorine, you would say HCl (H wants to obtain the noble gas count of He, which means hanging onto its electron and stealing another from Cl). Likewise, chlorine needs the noble gas count of Ar – thus it must keep all of its own electrons and acquire another. This leads to covalent bond formation (actually, there is a high ionic component to the HCl bond because H is on the metal side and Cl is on the nonmetal side, so on either count, there is only one possible formula for the combination of H and Cl atoms – HCl). Thus we are not allowed to fool around with the number of atoms within a chemical formula once the formulae of the reactants and products have been established. What we can do is to adjust the number of each type of molecule so that the number of each type of atom is the same on both sides of the chemical equation.

The steps involved in balancing chemical equations are straightforward.

1. Write out the skeleton chemical equation with the proper chemical formula of all of the molecules participating in the reaction. Once you have assured yourself that the chemical formulas of each molecule is correct, do not change them.
2. Count the number of each type of atom present on the reactant and the product sides.
3. By trial and error, adjust the number of each type of molecule so that the atoms add up. It is usually easiest to start with the most complex molecule first, then work down to adjust the simplest of the molecules. At each step, keep counting the atoms.

e.g. Hydrogen gas reacts with oxygen gas to form liquid water.

**Step 1:** Reactants: Hydrogen is a diatomic molecule, as is oxygen: \( H_2 (g) \) and \( O_2 (g) \).

Product: Water, \( H_2O (l) \)

**Skeleton equation:** \( H_2 (g) + O_2 (g) \rightarrow H_2O (l) \)

**Step 2:** Number of Atoms

Reactant Side: 2 atoms H and 2 atoms O

Product Side: 2 atoms H and 1 atoms O

**Step 3:** Adjust the number of each type of molecule until the equation is balanced. Start with balancing the water.

\[ H_2 (g) + O_2 (g) \rightarrow H_2O (l) \]

There are 2 atoms O on the left hand side but only one on the product side. To use up all the O’s, multiply the \( H_2O \) by a factor of 2 so that there are 2 molecules of water.

\[ H_2 (g) + O_2 (g) \rightarrow 2 H_2O (l) \]

Count up the atoms:

Reactant side: 2 atoms H, 2 atoms O

Product side: 4 atoms H, 2 atoms O

To balance the H’s multiply \( H_2 \) on the reactant side by a factor 2.

\[ 2 H_2 (g) + O_2 (g) \rightarrow 2 H_2O (l) \]
Count up the atoms:
Reactant side: 4 atoms H, 2 atoms O
Product side: 4 atoms H, 2 atoms O

The equation is now balanced. The balanced chemical equation now reads “two molecules of hydrogen react with one molecule of oxygen to form 2 molecules of water”.

e.g. Sodium reacts with water to form hydrogen gas and sodium hydroxide
Step 1: Reactants: Sodium is monatomic: \( Na (s) \), \( H_2O \ (l) \)
Product: Hydrogen is a diatomic molecule, \( H_2 \ (g) \) and \( NaOH \ (aq) \).

**Skeleton equation:** \( Na (s) + H_2O \ (l) \rightarrow NaOH \ (aq) + H_2 \ (g) \)

Step 2: Number of Atoms
Reactant Side: 1 atom Na, 2 atoms H and 1 atom O
Product Side: 1 atom Na, 3 atoms H and 1 atom O

Step 3: Adjust the number of each type of molecule until the equation is balanced.

\[ Na (s) + H_2O \ (l) \rightarrow NaOH \ (aq) + H_2 \ (g) \]

Note that there are an even number of H’s on the reactant side and an odd number of H’s on the product side. There is no real way to make the even number odd, but we can make the odd number even by multiplying by a factor of 2 molecules. The molecule that makes the number of H’s odd on the right hand side is NaOH – multiply NaOH by a factor of 2.

\[ Na (s) + H_2O \ (l) \rightarrow 2NaOH \ (aq) + H_2 \ (g) \]

Count up the atoms:
Reactant Side: 1 atom Na, 4 atoms H and 2 atom O
Product Side: 2 atom Na, 4 atoms H and 2 atom O

To balance the H’s multiply \( H_2O \) on the reactant side by a factor 2.

\[ Na (s) + 2H_2O \ (l) \rightarrow 2NaOH \ (aq) + H_2 \ (g) \]

Reactant Side: 1 atom Na, 4 atoms H and 2 atom O
Product Side: 2 atom Na, 4 atoms H and 2 atom O

Now balance the number of Na’s by multiplying Na by a factor 2 on the reactant side.

\[ 2Na (s) + 2H_2O \ (l) \rightarrow 2NaOH \ (aq) + H_2 \ (g) \]

Count up the atoms:
Reactant Side: 2 atom Na, 4 atoms H and 2 atoms O
Product Side: 2 atom Na, 4 atoms H and 2 atom O

The equation is now balanced. The balanced chemical equation now reads “two molecules of hydrogen react with one molecule of oxygen to form 2 molecules of water”.
Ammonia gas reacts with oxygen under the appropriate conditions at 1000°C to form nitrogen monoxide and water. Write the balanced equation.

\[ \text{NH}_3 (g) + O_2 (g) \rightarrow NO (g) + H_2O (g) \]

Odd number of H’s on the LHS so multiply NH\textsubscript{3} by factor 2.

\[ 2 \text{NH}_3 (g) + O_2 (g) \rightarrow 2NO (g) + H_2O (g) \]

Balance the N’s and H’s by multiplying NO by factor 2 and H\textsubscript{2}O by factor 3.

\[ 2 \text{NH}_3 (g) + O_2 (g) \rightarrow 2NO (g) + 3H_2O (g) \]

This gives us 2 O’s on the LHS and 5 O’s on the RHS. We can fix this by multiplying O\textsubscript{2} on LHS by 5/2.

\[ 2 \text{NH}_3 (g) + \left(\frac{5}{2}\right)O_2 (g) \rightarrow 2NO (g) + 3H_2O (g) \]

We can remove the fraction by multiplying the entire equation by factor 2.

\[ 2 \times \left[ 2 \text{NH}_3 (g) + \left(\frac{5}{2}\right)O_2 (g) \rightarrow 2NO (g) + 3H_2O (g) \right] \]

\[ 4 \text{NH}_3 (g) + 5O_2 (g) \rightarrow 4NO (g) + 6H_2O (g) \]

**Exercises:**

e.g. iron reacts with oxygen to form iron (III) oxide

e.g. Combustion of glucose to form carbon dioxide and water.

e.g. zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas.

e.g. Propane burns in oxygen to form carbon dioxide and water.