Table 2.7 lists some common covalent compounds.

Table 2.7 Covalent compounds

Covalent compound	Chemical formula
Water	H_2O
Ammonia	NH_3
Ethylene	C_2H_4
Nitrogen dioxide	NO_2
Carbon monoxide	CO

Names and formulae of compounds

The names and formulae of compounds can be determined by following some simple rules.

a. Ionic compounds

We examine only the rules for simple ionic compounds composed of one metal and one non-metal.

Naming rules

- Name the metal first.
- Name the non-metal second.
- Delete the last few letters of the non-metal's name and substitute '-ide'.

Examples

1. metal = zinc

   non-metal = sulfur

   ionic compound = zinc sulfide.

2. metal = calcium

   non-metal = chlorine

   ionic compound = calcium chloride.

Valency rules for formulae

The chemical formula of an ionic compound can be determined using **valency rules.**

- Each metal and non-metal is assigned a valency.
- The valency is the *combining power* of the element.
- The valency of a metal is equal to the charge on its cation.
- The valency of a non-metal is equal to the charge on its anion.
- In an ionic compound the total positive charge of the cations must equal the total negative charge of the anions.

Examples

1. **Q** Determine the chemical formula of zinc sulfide.

   A Zinc sulfide: valency of zinc = +2 (Zn^{2+})

   valency of sulfur = −2 (S^{2-})

   The total positive charge equals the total negative charge.

   Thus there is one zinc ion and one sulfide ion in the compound.

   Thus, the formula of zinc sulfide = ZnS.

2. **Q** Determine the chemical formula of calcium chloride

A Calcium chloride:

valency of calcium = +2 (Ca^{2+})

valency of chlorine = −1 (Cl^{-})

The positive charge is greater than the negative charge.

Thus two chloride ions are needed to balance the charges.

Thus there is one calcium ion and two chloride ions in the compound. Thus, the formula of calcium chloride $= CaCl_2$.

b. Covalent compounds

We examine only some simple examples of covalent compounds containing only two different elements.

Naming rules

- Name the non-metal with the lower Periodic Group number first.
- Name the non-metal with the higher Periodic Group number second.
- Use Greek prefixes to indicate the number of each type of atom

  (mono = 1; di = 2; tri = 3; tetra = 4; penta = 5).
- Delete the last few letters of the second non-metal's name and substitute '-ide'.

Examples

1. CO = carbon *mon*oxide
2. NO_2 = nitrogen *di*oxide
3. PCl_3 = phosphorus *tri*chloride
4. CF_4 = carbon *tetra*fluoride
5. N_2O_4 = *di*nitrogen *tetr*oxide

Valency rules for formulae

The valency rules are modified as there are no charged atoms in covalent compounds.

- Each non-metal atom is assigned a valency.
- For a simple covalent compound with two elements, the total valencies of each element must be equal.

Table 2.8 lists some of the common valencies of non-metals for Groups IV to VII. Note that some non-metals have more than one common valency.

- The normal valency of hydrogen is 1 and that of oxygen is 2.

Table 2.8 Common valencies of non-metal atoms in covalent compounds

Group IV	Group V	Group VI	Group VII
carbon 4	nitrogen 2, 3, 4, 5	oxygen 2	fluorine 1
silicon 4	phosphorus 3, 5	sulfur 2, 4	chlorine 1

Examples

1. **Q** Determine the chemical formula of silicon dioxide.

   A silicon dioxide: valency of Si = 4

   valency of O = 2

   To make the valencies of each element equal we need two atoms of oxygen ($2 \times 2 = 4$). Thus the chemical formula of silicon dioxide = SiO_2.

2. Q Determine the chemical formula of dinitrogen trioxide.

A dinitrogen trioxide: valency of N = 3

valency of O = 2

To make the valencies equal we need two atoms of N and three atoms of O ($2 \times 3 = 3 \times 2$).

Thus, the chemical formula of dinitrogen trioxide = N_2O_3.

Reactions, observations and equations

Making accurate observations is an important skill when performing chemical reactions. Some events during a reaction cannot be detected using our senses, while others are easy to observe.

Common observations

- A substance dissolves.
- Effervescence occurs (ie. gas bubbles form).
- The reaction mixture changes colour.
- The reaction mixture gets hot or cold.
- An insoluble substance (or precipitate) forms when solutions are mixed.
- Flames are produced or an explosion occurs.

Observations should be consistent with the chemical equation for the reaction.

Chemical equations

A chemical equation summarises the events of a chemical reaction.

- The reacting chemicals (reactants) are written first.
- The new substances that form (products) are written last.
- An arrow separates the reactants and products. The arrow stands for the phrase 'reacts to form'.

- The equation shows that atoms are conserved in the reaction (ie. the number of atoms of reactants equals the number of atoms of products).
- The equation is consistent with the law of matter conservation.

Chemical equations can be written as a **word equation** or a **symbolic equation**.

Let us examine some simple cases.

Example 1. Burning magnesium in air

Observation: Silvery magnesium is heated in a Bunsen burner flame. The magnesium burns with a dazzling bright white light and a crumbly white powder is formed.

Explanation: The hot magnesium atoms react with oxygen molecules in the air to form a white ionic compound called magnesium oxide.

Word equation

magnesium + oxygen → magnesium oxide

Symbolic equation

Step 1. Write the correct symbols and formula for each substance.

magnesium = Mg

oxygen = O_2 (Remember that oxygen gas is a diatomic molecule.)

magnesium oxide = MgO (Remember that the valency of Mg = +2 and of oxygen = −2.)

Step 2. Replace the words in the word equation by the chemical formula.

$$Mg + O_2 \rightarrow MgO$$

Step 3. Check that the atoms of each element balance on each side of the equation. (Atoms cannot disappear! All they do is get rearranged.)

Reactants: 1 magnesium atom

2 oxygen atoms (There are two atoms in the one molecule.)

Products: 1 magnesium atom (ion)

1 oxygen atom (oxide ion)

Conclusion: The magnesium atoms balance but the oxygen atoms do not.

Step 4. Place integers in front of the formula to achieve an atom balance.

In this case place a 2 in front of Mg and a 2 in front of MgO.

$$2Mg + O_2 \rightarrow 2MgO$$

Step 5. Re-check that the atoms now balance.

There are two magnesium atoms on each side and two oxygen atoms on each side.

Figure 2.12 shows a particle model of the reaction equation for magnesium plus oxygen.

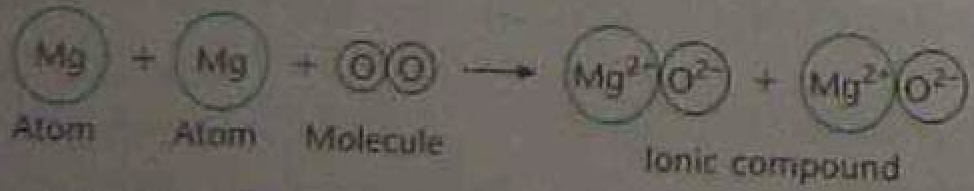

Figure 2.12 Particle model of magnesium + oxygen reaction

Example 2. Reaction of sodium and chlorine

Observation: Silvery sodium is heated until it forms a pool of molten sodium. The hot sodium is placed in a gas jar of yellow-green chlorine gas. A white smoke of very fine crystals is seen to form.

Explanation: The hot atoms of sodium reac with chlorine molecules to form white crystals of the ionic compound called sodium chloride.

Word equation

sodium + chlorine → sodium chloride

Symbolic equation

Step 1. Write the correct symbols and formulae for each substance.

sodium = Na

chlorine = Cl_2 (Remember that chlorine gas is a diatomic molecule.)

sodium chloride = NaCl. (Remember that the valency of Na = +1 and of chlorine = –1.)

Step 2. Replace the words in the word equation by the chemical formula.

$$Na + Cl_2 \rightarrow NaCl$$

Step 3. Check that the atoms of each element balance on each side of the equation. (Atoms cannot disappear! All they do is get rearranged.)

Reactants: 1 sodium atom

2 chlorine atoms (There are 2 atoms in the one molecule.)

Products: 1 sodium atom (ion)

1 chlorine atom (oxide ion)

Conclusion: The sodium atoms balance but the chlorine atoms do not.

Step 4. Place integers in front of the formula to achieve an atom balance.

In this case place a 2 in front of Na and a 2 in front of NaCl.

$$2Na + Cl_2 \rightarrow 2NaCl$$

Step 5. Re-check that the atoms now balance.

There are 2 sodium atoms on each side and 2 chlorine atoms on each side.

Combustion reactions

In a combustion reaction a **fuel** reacts with an **oxidiser** (eg. oxygen in the air) to produce **energy**, usually in the form of heat.

Fuel + Oxidiser → Products + Energy

Examples of combustion reactions are:

1. Burning a wax candle
2. Burning natural gas in a Bunsen burner
3. Combustion of petrol in a car's engine
4. A bushfire (wood cellulose and other fuels burn in oxygen)
5. Explosion of hydrogen in air
6. Burning sulfur in air

Each of these combustion reactions does not happen spontaneously when the reactants are combined. An **ignition source** is needed. We often use a spark or burning match to ignite other fuels. Once the combustion starts, the heat energy produced is sufficient to keep the reaction going. If the oxygen concentration is low, less heat is produced in the combustion and products such as black soot (carbon) or poisonous carbon monoxide are formed.

Worked example. Burning natural gas in a Bunsen burner

In this example we assume that natural gas is methane (CH_4).

Observations: With the valve (or hole) in the burner open, the gas burns with a blue flame. Heat and light energy are produced.

Explanation: The hot molecules of methane and oxygen combine to form carbon dioxide molecules and water molecules.

Word equation

methane + oxygen → carbon dioxide + water

Balanced symbolic equation

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

The balanced symbolic equation shows that one molecule of methane reacts with two molecules of oxygen to form one molecule of carbon dioxide and two molecules of water. A scientist can prove that carbon dioxide and water are formed by collecting the colourless gases coming off the flame and cooling them. The water vapour condenses to liquid water and the carbon dioxide can be confirmed because it turns limewater milky white.

Figure 2.13 shows a particle diagram of the combustion reaction of methane.

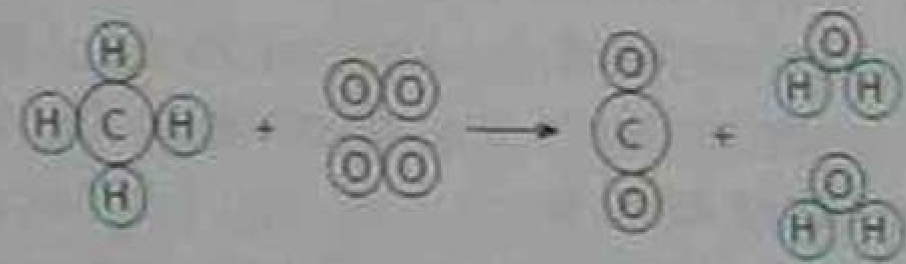

Figure 2.13 Particle diagram of methane combustion

Investigating combustion

Investigate the combustion of a wax candle in air and in a beaker containing oxygen (prepared by the catalytic decomposition of hydrogen peroxide with manganese dioxide). This experiment should show that the rate of combustion of the candle increases, as the candle burns faster when more oxygen is present.

Corrosion reactions

We are all familiar with the **rusting** of iron and steel objects. Steel bridges have to be regularly maintained to prevent them rusting. Rusted steel is weak and the steel structure will collapse if the rust penetrates deeply into the structure. In the home, wet steel wool left on the sink will quickly turn brown because it rusts. Rusting is an example of the process of corrosion.

- Corrosion is the process in which a metal **degrades** or wears away on exposure to the environment.

Iron corrodes because it reacts with oxygen and water in the environment to form a surface layer of a crumbly, brown iron compound that is called rust.

iron + oxygen + water → rust

Iron will not corrode if either oxygen or water is absent. Therefore, to protect iron from corrosion its surface must be coated with a material that prevents oxygen and/or water making contact. Paints and thin layers of metal (eg. zinc, tin) are commonly used to protect iron from corrosion. Coating of iron with zinc is called **galvanising**.

Aluminium is a common structural metal that has significant corrosion resistance. This is due to the natural layer of aluminium oxide that forms on the surface of aluminium when it is exposed to air. The oxide layer protects the aluminium below from further attack.

Investigating corrosion

The conditions under which corrosion occurs can be investigated in a series of controlled experiments. Equal quantities of fresh samples of steel wool can be placed in test tubes under different conditions. These conditions could include:

- dry air (use drying crystals in the stoppered test tube)
- moist air (place a drop of water in the stoppered tube)
- air and water present (cover half the steel wool with water)
- water only (use boiled water and immerse the steel wool; add a layer of oil to prevent air re-entering the water)
- a control tube with steel wool in an open tube.

These experiments should confirm that most rusting occurs in the presence of air and moisture.