1.1 – Formulae and equations

1. Formulae of compounds and common ions

The formula of a compound shows which elements are present, as well as the ratio for the number of atoms of each element.

e.g. $Al_2(SO_4)_3$ **aluminium sulfate**

two aluminium atoms

3 × 1 = 3 sulfur atoms

3 × 4 = 12 oxygen atoms

(The number outside the bracket multiplies everything in the bracket by that number.)

Formulae of some common substances

Name	Formula	Name	Formula
water	H_2O	hydrochloric acid	HCl
oxygen	O_2	sulfuric acid	H_2SO_4
hydrogen	H_2	nitric acid	HNO_3
nitrogen	N_2	sodium hydroxide	NaOH
carbon dioxide	CO_2	sodium carbonate	Na_2CO_3
sulfur dioxide	SO_2	ammonia	NH_3

Charges on some common ions

The charge on some ions can be deduced from the element's group in the Periodic Table.

Group	Charge	Example
1	+	Na^+
2	2+	Mg^{2+}
6	2–	O^{2-}
7	–	Cl^-

Charges on some ions are more difficult to deduce.

Cations		Anions	
hydrogen	H^+	hydroxide	OH^-
silver	Ag^+	nitrate	NO_3^-
zinc	Zn^{2+}	sulfate	SO_4^{2-}
ammonium	NH_4^+	carbonate	CO_3^{2-}

Constructing a formula from ions

- Write the symbols for the ions required.
- If the charges on the ions balance, simply write the formula without the charges.

e.g. sodium chloride Na^+ Cl^- ⟶ **formula NaCl**

- If the charges on the ions do not balance, then choose the ratio of positive to negative ions needed to balance the charges.

e.g. potassium oxide K^+ O2⁻

- The charges don't balance.
- Two K^+ ions are needed to balance the charge on one O^{2-} ion ⟶ **formula K_2O.**
- Sometimes brackets are needed for clarity when compound ions are involved.

e.g. magnesium nitrate Mg^{2+} NO_3^- ⟶ **formula $Mg(NO_3)_2$**

2. Oxidation numbers

The oxidation number of an element indicates the number of electrons that need to be lost or gained by the element to make it neutral.

Oxidation numbers increasing or decreasing during a redox reaction show which species is oxidised and which is reduced.

- The sum of the oxidation numbers in a compound is 0.
- The sum of the oxidation numbers in an ion is equal to the overall charge on the ion.
- The most electronegative element in a compound is assigned the negative oxidation number.

Rule		Species	Oxidation no.	
Uncombined		N_2	nitrogen	0
elements	0	Fe	iron	0
Group 1 metals	+1	NaCl	sodium	+1
Group 2 metals	+2	$BaSO_4$	barium	+2
oxygen	−2	MgO	oxygen	−2
except with fluorine	+2	F_2O	oxygen	+2
except in peroxides	−1	H_2O_2	oxygen	−1
hydrogen	+1	HCl	hydrogen	+1
except in				
metal hydrides	−1	KH	hydrogen	−1
fluorine	−1	F_2O	fluorine	−1

Assigning oxidation numbers

e.g. $CaCl_2$

Ca is a Group 2 metal ⟶ oxidation number +2.

There are two chlorine atoms in the formula so (2 × −1) to make it neutral ⟶ oxidation number −1.

e.g. HCO_3^-

H ⟶ oxidation number +1

O ⟶ oxidation number −2

In the formula there is one hydrogen atom ⟶ +1,

three oxygen atoms (3 × −2) ⟶ −6.

There is one carbon atom, so to obtain an overall ion charge of −1 ⟶ oxidation number +4.

3. Balanced chemical and ionic equations

Chemical equations show us what happens during a chemical reaction. Chemical equations need to be balanced, i.e. they must have the same number of atoms of each element on each side. This is achieved by putting a number in front of the formula to add more units of that substance.

State symbols give information about the states of the species in the equation.

(s) solid	(l) liquid
(g) gas	(aq) solution in water

They should always be included in ionic equations and equations showing enthalpy changes.

e.g. $Na + O_2 \longrightarrow Na_2O$

Reactants ⟶ 1 Na atom and 2 O atoms

Products ⟶ 2 Na atoms and 1 O atom

Balance the O atoms by having two units of Na_2O and then balance the Na atoms by having four units of Na.

$$4Na + O_2 \longrightarrow 2Na_2O$$

Ionic equations show only the ions that take part in a chemical reaction. Ions that do not change during the reaction – 'spectator ions' – are left out.

e.g.

$$CuSO_4(aq) + 2NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

The SO_4^{2-} and Na^+ ions do not take part. The important ions are the Cu^{2+} and OH^- ions which have reacted to form a pale, blue precipitate of copper(II) hydroxide, $Cu(OH)_2$.

$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$$