Topic 2a – Equations
Revision Notes

1) Formulae

a) Elements

- For most elements the formula is just the symbol e.g. Na for sodium, S for sulphur
- The exceptions are the seven diatomic elements – H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2

b) Ionic compounds

- Compounds of a metal and a non-metal are made of ions
- Metal ions have a positive charge
- Ions of Group 1 elements have a +1 charge, ions of Group 2 elements have a +2 charge
- For transition elements, like copper and iron, the number after the name gives the charge on the ion e.g. copper(II) oxide contains Cu^{2+} ions, iron(III) oxide contains Fe^{3+} ions
- Non-metal ions have a negative charge
- Ions of Group 7 elements have a -1 charge, ions of Group 6 elements have a -2 charge
- You need to learn the formulas of five ions: nitrate, NO_3^-, carbonate, CO_3^{2-}, sulphate, SO_4^{2-}, hydroxide, OH^-, and ammonium, NH_4^+
- To work out the formula of an ionic compound
    - Write the formulae of the ions
    - Adjust the number of each ion so that there is no overall charge

Example 1 – magnesium bromide	Example 2 – aluminium nitrate
Ions are Mg^{2+} and Br^-	Ions are Al^{3+} and NO_3^-
Need 2 x Br^- to balance Mg^{2+}	Need 3 x NO_3^- to balance Al^{3+}
Formula is $MgBr_2$	Formula is $Al(NO_3)_3$

c) Covalent compounds

- Some formulae for covalent compounds can be worked out from the name.
- The prefix mono- means one, di- means two and tri- means three.
- Therefore, carbon monoxide is CO, silicon dioxide is SiO_2 and sulphur trioxide is SO_3
- Other formulae have to be learnt e.g. ammonia is NH_3 and methane is CH_4

2) Equations

- **There are no word equations at A-level. An equation means a balanced symbol equation.**
- To write a balanced symbol equation:
  - Identify the reactants and products
  - Write a word equation
  - Write down the formula for each substance
  - Balance the equation by putting numbers in front of formulae
  - Add state symbols (s), (l), (g) or (aq)

Example – marble chips and hydrochloric acid

Reactants are calcium carbonate and hydrochloric acid
Products are calcium chloride, carbon dioxide and water

Calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water
$CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Ca	1	1
C	1	1
O	3	3
H	1	2
Cl	1	2

2 in front of HCl balances the equation

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Adding state symbols

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

4) Ionic equations

- Ionic equations leave out ions that are unchanged in a reaction. They give a clearer picture of what is happening in a reaction
- To go from a symbol equation to an ionic equation:
  - Split up anything that is (aq) and ionic (acids, alkalis and salts)
  - Cancel ions that are on both sides

Example

$$\mathbf{HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)}$$

$$\mathbf{H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow}$$
$$\mathbf{Na^+(aq) + Cl^-(aq) + H_2O(l)}$$

$$\mathbf{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$$

Topic 2b – Acids
Revision Notes

1) Acids

- Common acids are hydrochloric acid, HCl, sulphuric acid, H_2SO_4, and nitric acid, HNO_3
- An acid releases H^+ ions in aqueous solution
- An acid is a proton donor (or H^+ donor)

a) Effect of solvent on acidity

- Hydrogen chloride, HCl, is a gas
- When dissolved in water in splits up into ions and becomes hydrochloric acid

$$\mathbf{HCl(g) + aq \rightarrow H^+(aq) + Cl^-(aq)}$$

- It is the presence of $H^+(aq)$ ions that allows hydrochloric acid to behave as an acid
- Hydrogen chloride gas will also dissolve in organic solvents, such as hexane. It stays as HCl molecules in these solvents. It does not split up into ions and, so, does not behave as an acid e.g. does not release CO_2 from a carbonate

2) Bases

- Bases are metal oxides, metal hydroxides and ammonia solution. Bases destroy acidity in neutralisation reactions
- A base readily accepts H^+ ions from an acid e.g.

$$\mathbf{OH^- + H^+ \quad H_2O}$$
$$\mathbf{NH_3 + H^+ \quad NH_4^+}$$

- A base is a proton acceptor (or H^+ acceptor)

3) Alkalis

- An alkali is a soluble base. Alkalis releases OH^- ions in aqueous solution
- Common alkalis are sodium hydroxide, NaOH, potassium hydroxide, KOH, and aqueous ammonia, $NH_3(aq)$
- Alkalis are a sub-set of bases. Sodium hydroxide and sodium oxide are bases **and** alkalis because they both produce OH^- ions when dissolved in water.
- Copper(II) oxide is a base but not an alkali because it is insoluble

4) Salts

- A salt is formed when the H^+ of an acid is replaced by a metal ion or NH_4^+
- Salts are formed when acids react with bases and carbonates

Acid + carbonate → salt + water + carbon dioxide

e.g. $\mathbf{2HCl(aq) + MgCO_3(s) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)}$

Observations – effervescence (fizzing), carbonate dissolves/disappears

Acid + base → salt + water

e.g. **$2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$**

Observations – base dissolves/disappears

Acid + alkali → salt + water

e.g. **$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$**

Observations – no visible change unless indicator added

e.g. **Sulphuric acid + ammonia → ammonium sulphate**

$H_2SO_4(aq) + 2NH_3(aq) \rightarrow (NH_4)_2SO_4(aq)$