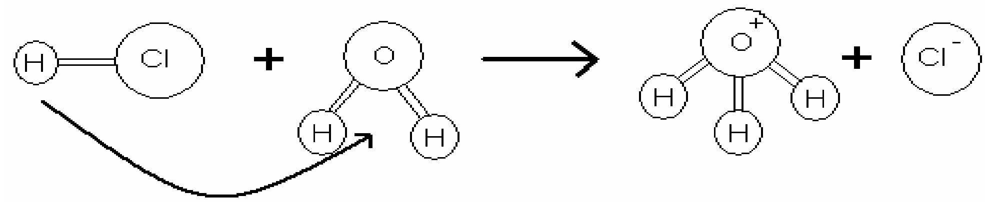
Chapter 7: Acids and Bases

1. **Acids**

Arrhenius' definition:

- ✧ An acid is a chemical substance which ionizes in water to produce hydrogen ions, H^+.
- ✧ Example: Hydrochloric acid, HCl is obtained when a hydrogen chloride molecule, HCl dissolves in water and ionizes to produce a hydrogen ion H^+.
- ✧ The H^+ combines with water molecule, H_2O to form a hydroxonium ion, H_3O^+. This ion can be written as H^+.

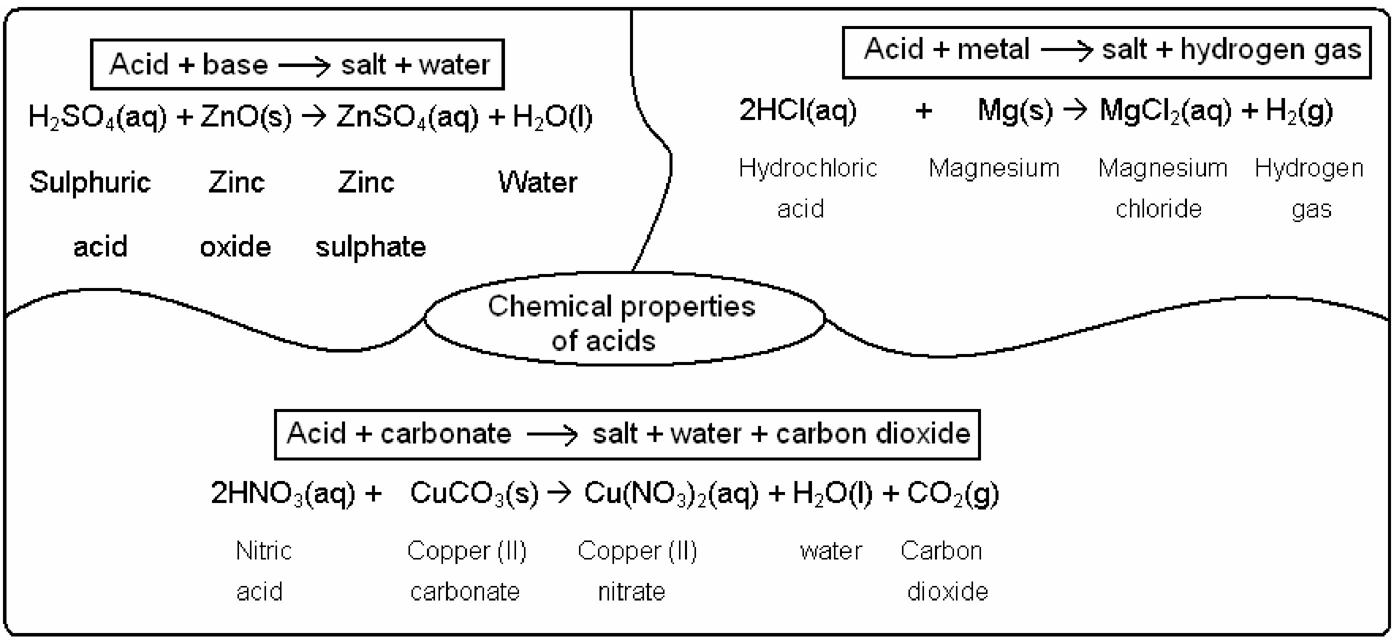


- ✧ We can classify an acid as a monoprotic acid (HCl) or a diprotic (H_2SO_4) acid based on its basicity.
- ✧ Basicity is the number of ionisable hydrogen atoms per molecule of an acid.3.
- ✧ Mineral acid: HCl, H_2SO_4, HNO_3.
- ✧ Organic acid: CH_3COOH, HCOOH, $H_2C_2O_4$

✧ Malic acid is found in apples ✧ Citric acid is found in citrus fruits such as oranges. ✧ Tartaric acid is found in grapes	✧ Ethanoic acid is found in vinegar ✧ Latic acid is found in sour milk ✧ Tannic acid is found in tea leaves. ✧ Ascorbic acid is vitamin C

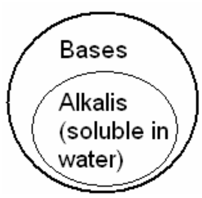
Chemical properties of acids:

i. Acids react with bases to form salts and water.
ii. Acids react with reactive metals to produce salts and hydrogen gas, H_2.
iii. Acids react with carbonates to produce salts, water and carbon dioxide, CO_2.

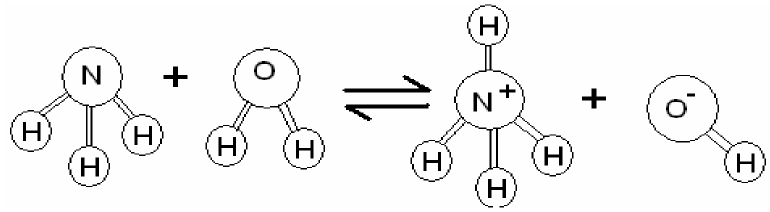


2. **Bases and alkalis**

Arrhenius' definition:

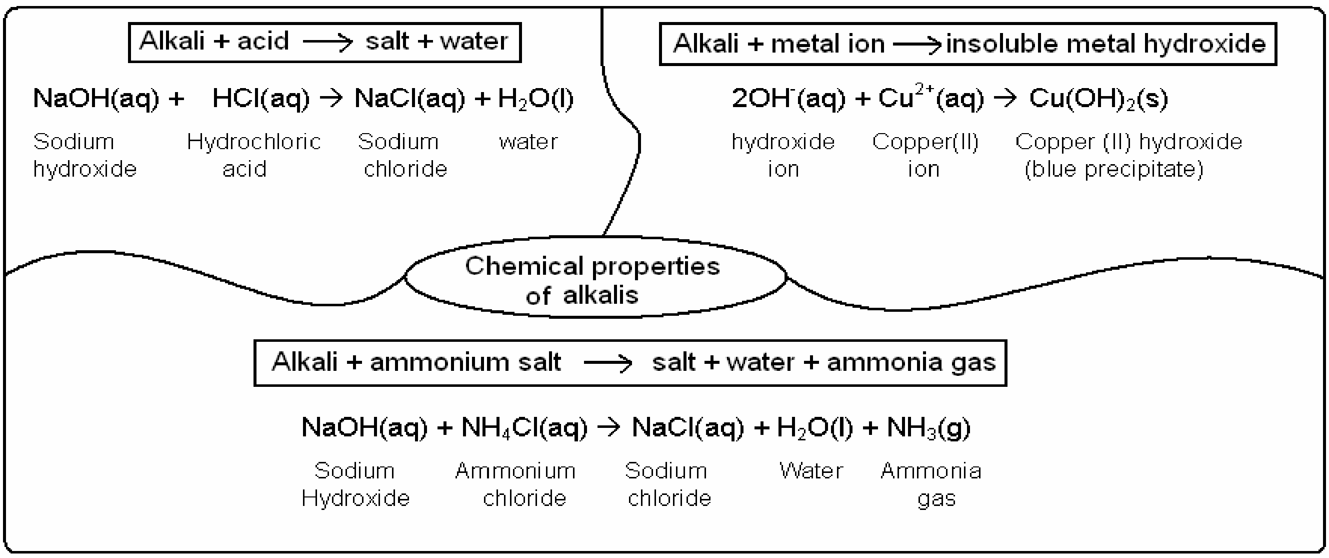


- A base is a chemical substance which ionizes in water to produce hydroxid
- Soluble bases are known as alkalis. Examples: NaOH, KOH, NH_3.
- $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
- $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$



Chemical properties of alkalis:

i. Alkalis react with acids to form salts and water.
ii. When a mixture of an alkali and an ammonium salt is heated, ammonia gas, NH_3 is liberated.
iii. Most metal hydroxides are insoluble in water. Hence, adding an alkali to most metal ion solutions will give a precipitate of an insoluble metal hydroxide.



3. **Role of water**

Acids	**Alkalis**
✓ Acid only shows its acidic properties when dissolved in water. ✓ This is because without water, an acid still exists as covalent molecules and there are no H^+ present. ✓ In the presence of water, the acid ionizes to form H^+. ✓ This causes the acid to show its acidic properties. ✓ Dry acids do not conduct electricity. This is because there are no freely moving ions. Dry acid exists as covalent molecules. ✓ Only aqueous solution of acid conducts electricity.	✓ Alkali only shows its alkaline properties when dissolved in water. ✓ Water helps bases to dissociate to produce hydroxide ions, OH^- that are responsible for the alkaline properties. ✓ Dry alkalis do not conduct electricity. ✓ An aqueous alkali solution can conduct electricity showing the presence of freely moving ions.

4. **The pH scale**

- The pH is used to indicate the degree of acidity or alkalinity of a solution.
- It consists of pH values that range from 0 to 14.
- pH < 7 ➔ an acidic solution
- pH = 0 ➔ a neutral solution
- pH > 7 ➔ an alkaline solution

← Increasingly acidic Neutral ↓ Increasingly alkaline →

0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

- The pH value is actually a measure of the concentration of hydrogen ions, H^+ and hydroxide ions, OH^-.
- The lower the pH value, the higher the concentration of hydrogen ions, H^+
- The higher the pH value, the higher the concentration of hydroxide ions, OH^-.
- A pH meter, pH paper or Universal indicator can be used to determine the pH value of a solution.

5.	**Strong**	**Weak**
Acids	- Strong acids are fully ionized in water. - Example: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ - $HNO_{3+} \rightarrow H^+ + NO_3^-$ - $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ - The one-way arrow → indicates complete dissociation - Complete dissociation (100%) in water by a strong acid produces a high concentration of H^+ ions and hence a low pH.	- Weak acids are only partially ionized in water. - Most of the organic acids such as ethanoic acid, ethanedioic acid, methanoic acid, citric acid and tartaric acid are weak acids. - Example:$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$ - $H_2C_2O_4 \rightleftharpoons 2H^+ + C_2O_4^{2-}$ - The incomplete ionization is indicated by a reversible sign. - In a weak solution, a big portion of the weak acid exists as molecules and only a small portion dissociates to ions.
Alkalis	- A strong alkali is completely ionized in water. - NaOH is a strong alkali, ionizes completely when dissolved in water. - $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$ - $KOH \rightarrow K^+ + OH^-$	- A weak alkali ionizes partially in water. - NH_3 is an example of a weak alkali. - $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ - $Ca(OH_2) \rightleftharpoons Ca^{2+} + 2OH^-$

6. **Concentration**

- The concentration of a solution refers to the quantity of solute in a given volume of solution, which is usually 1 dm^3.
- The quantity of solute can be measured in grams or moles.

$$\text{Concentration } (g/dm^3) = \frac{\text{Mass of solute (g)}}{\text{Volume of solution } (dm^3)}$$

$$\text{Concentration } (mol/dm^3) = \frac{\text{Number of moles of solute (mol)}}{\text{Volume of solution } (dm^3)}$$

- Concentration of a solution can be expressed in g dm^{-3} or mol dm^{-3}.
- Unit: Molarity (mol dm^{-3}), molar concentration (M)

7. **Molarity**

- Molarity is the number of moles of solute that are present in 1 dm^3 of solution.

Molarity (mol dm^{-3}) —— x Molar mass ——→ Concentration (g dm^{-3})

Molarity (mol dm^{-3}) ←—— ÷ Molar Mass —— Concentration (g dm^{-3})

$$\text{Molarity} = \frac{\text{Number of moles of solute (mol)}}{\text{Volume of solution } (dm^3)}$$

- Number of moles of solute = Molarity x volume of solution

$$n = MV$$

- n = Number of moles of solute (mol)
- M = Molarity of solution (mol dm^{-3})
- V = Volume of solution (dm^3)

Example 1

5.00 g of copper (II) sulphate is dissolved in water to form 500 cm^3 solution. Calculate the concentration of copper (II) sulphate solution in g dm^{-3}.

Solution:

Mass of copper (II) sulphate = 5.00 g

Volume of solution = 500 cm^3

$= \frac{500}{1000} dm^3$.

$= 0.5\ dm^3$

Hence, concentration of copper (II) sulphate solution

$= \frac{5.00}{0.5} = 10.0\ g\ dm^{-3}$.

$$\text{Concentration } (g/dm^3) = \frac{\text{Mass of solute (g)}}{\text{Volume of solution } (dm^3)}$$

Example 2:

What is the mass of sodium carbonate required to dissolve in water to prepare a 200 cm3 solution that contains 50 g dm^{-3}?

Solution:

Volume of solution = 200 cm^3

$= \frac{200}{1000} dm^3 = 0.2\ dm^3$

$$\text{Concentration } (g\ dm^{-3}) = \frac{\text{mass of } Na_2CO_3 \text{ dissolved (g)}}{\text{volume of solution } (dm^3)}$$

Mass of Na_2CO_3 required

$= 50\ g\ dm^{-3} \times 0.2\ dm^3 = 10\ g$

Mass = Concentration (g dm^{-3}) x Volume of solution (dm^3)

Example 3	**Example 4**
Calculate the number of moles of ammonia in 150 cm^3 of 2 mol dm^{-3} aqueous ammonia. *Solution:* Number of moles = $\frac{MV}{1000}$ Number of moles of ammonia = $2 \times \frac{150}{1000}$ = 0.3	A 250 cm^3 solution contains 0.4 moles of nitric acid. Calculate the molarity of the nitric acid. *Solution:* Number of moles = $\frac{MV}{1000}$ Molarity of nitric acid, M = $\frac{0.4 \times 1000}{250}$ = 1.6 mol dm^{-3}

8. **Dilution method**

- Adding water to a concentrated solution changes the concentration of the solution but does not change the amount of solute present in the solution.

$$M_1 \times V_1 = M_2 \times V_2$$

- M_1 = the molarity of the solution before water is added.
- V_1 = the volume of the solution before water is added.
- M_2 = the molarity of the solution after water is added.
- V_2 = volume of the solution after water is added.

9. **Neutralisation**

- During neutralization, an acid reacts with an alkali to form a salt and water only.
- The actual reaction that occurred is between one hydrogen ion, H^+, from the acid and one hydroxide ion, OH^-, from the alkali to produce one molecule of water, H_2O.
- Chemical equation: HCl + NaOH → NaCl + H_2O
- **Ionic equation : $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$** / $H_3O^+ + OH^- \rightarrow 2H_2O$

10. **Acid-base titration**

- Titration is a very useful laboratory technique in which one solution is used to analyse another solution.
- In acid-base titration, an acid of know concentration is carefully delivered from a burette to completely neutralize a known volume of an alkali in a conical flask.
- The end point is a point in titration where the indicator changes colour.

Indicator	Colour in acidic solution	Colour in neutral solution	Colour in alkaline solution
Litmus solution	Red	Purple	Blue
Phenolphatalein	**Colourless**	**Colourless**	**Pink**
Methyl orange	Red	Orange	Yellow
Methyl red	Red	Orange	Yellow
Universal indicator	Red/Orange/yellow	Green	Purple

Preparation $100cm^3$ of 0.1 mol dm^{-3} aqueous sodium hydroxide solution

Materials: NaOH solid and distilled water

Apparatus: Electronic balance, weighing bottle, $100cm^3$ volumetric flask, filter funnel and washing bottle

Procedure:

1. The mass od NaOH required to prepare $100cm^3$ of 0.1 mol dm^{-3} awueous NaOH is calculated as follows:

   Mass of NaOH required

   = mol NaOH x RMM NaOH

   = $\frac{MV}{1000}$ x (23 + 16 + 1) = $\frac{0.1 \times 100}{1000} \times 40 = 0.4$ g
2. 0.4g of NaOH solid is weighed accurately in a weighing bottle using an electronic balance.
3. NaOH solid is transferred to a small beaker. Sufficient distilled water is added to dissolved all the solid NaOH.
4. Using a filter funnel and glass rod, the NaOH solution is transferred to a $100cm^3$ volumetric flask.
5. The small beaker, the weighing bottle and the filter funnel are all rinsed with distilled water using a washing bottle, and the contents are transferred into the volumetric flask.
6. Distilled water is then added slowly until the water level mark of the volumetric flask.
7. Close the volumetric flask with the stopper. The volumetric flask is then shaken and inverted to mix the solution completely. The solution prepared is $100cm^3$ of 0.1 mol dm^{-3} aqueous sodium hydroxide.