

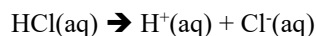
Intensive notes (Topic 4: Acids and bases)

Common acids and bases

Acid: A molecular compound which forms $\text{H}^+(\text{aq})$ as the only cation when dissolved in water

1. Hydrochloric acid (HCl)

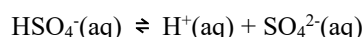
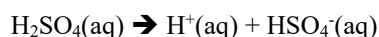
- A **strong monobasic** acid (There is 1 $\text{H}^+(\text{aq})$ can be produced by the complete ionization of one molecule of acid)



- It can be found in stomach juice. Concentrated HCl is volatile

2. Sulphuric acid (H_2SO_4)

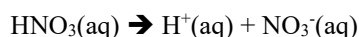
- A **strong dibasic** acid (There are 2 $\text{H}^+(\text{aq})$ can be produced by the complete ionization of one molecule of acid)



- Concentrated H_2SO_4 is hygroscopic, dehydrating and oxidizing

3. Nitric acid (HNO_3)

- A **strong monobasic** acid (There is 1 $\text{H}^+(\text{aq})$ can be produced by the complete ionization of one molecule of acid)



- Concentrated HNO_3 is volatile and oxidizing.
- Concentrated HNO_3 should be stored in brown bottle as HNO_3 can be decomposed by light.

4. Ethanoic acid (CH_3COOH)

- A **weak monobasic** acid (There is 1 $\text{H}^+(\text{aq})$ can be produced by the complete ionization of one molecule of acid)



- It can be found in vinegar. Concentrated ethanoic acid is volatile

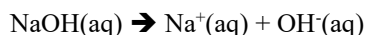
Common bases

Base: A compound which reacts with an acid forming salt and water only (Examples: Metal hydroxides, metal oxides, NH_3)

Alkalis: A base which is soluble in water (Examples: Hydroxides of Group I and some Group II metals, NH_3)

1. Sodium hydroxide (NaOH) and potassium hydroxide (KOH)

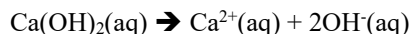
- Both are **strong** alkali



- Concentrated NaOH can be found in drain cleaner
- Concentrated NaOH and KOH are hygroscopic (absorb water) and absorb $\text{CO}_2(\text{g})$ in air readily

2. Calcium hydroxide ($\text{Ca}(\text{OH})_2$)

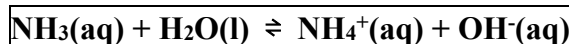
- A **strong** base, but only slightly soluble in water



- $\text{Ca}(\text{OH})_2(\text{aq})$ is also known as limewater. It can be used to neutralize acidic soils.

3. Aqueous ammonia (NH_3)

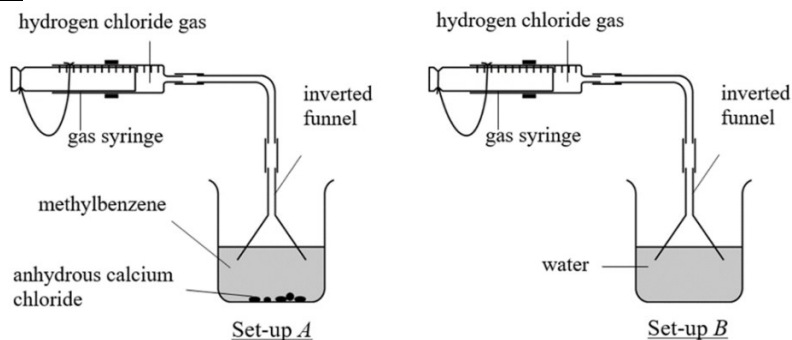
- A **weak** alkali



- NH_3 has a characteristic pungent smell. $\text{NH}_3(\text{aq})$ are volatile
- $\text{NH}_3(\text{g})$ can be tested by moist red litmus paper. It ionizes in water to give $\text{OH}^-(\text{aq})$ which turns red litmus blue.

Intensive notes (Topic 4: Acids and bases)

Water and acidic properties



Test	Set-up A (HCl dissolved in methylbenzene)	Set-up B (Hydrochloric acid)
Action on dry blue litmus paper	No observable change (H ⁺ ions are absent)	Changes from blue to red (H ⁺ ions are present)
Electrical conductivity	Does not conduct electricity (HCl remains as neutral molecules)	Conducts electricity (HCl ionizes in water to give H ⁺ and Cl ⁻)
Action on anhydrous Na ₂ CO ₃ (s)	No observable change (H ⁺ ions are absent)	Colorless gas bubbles given out (2H ⁺ (aq) + CO ₃ ²⁻ (aq) → CO ₂ (g) + H ₂ O(l))

- Use of inverted funnel can increase surface area for dissolving HCl(g), prevent sucking back of water.

Diluting concentrated acids / alkalis

- Safety precautions
 1. Wear safety goggles and protective gloves
 2. Dilution of volatile acids / alkalis (Conc. HCl, HNO₃, NH₃ etc.) should be performed in fume cupboard.
- Small amount of the concentrated acid / alkali should be added slowly to a large amount of water with stirring, but not vice versa. This is because large amount of heat will be produced during dilution. This can vaporize the water and cause splashing out of the acid /alkali

Basicity of common acids

Acid	Ionization in water	Basicity of acid
Hydrochloric acid	HCl(aq) → H ⁺ (aq) + Cl ⁻ (aq)	1 (Monobasic)
Nitric acid	HNO ₃ (aq) → H ⁺ (aq) + NO ₃ ⁻ (aq)	
Nitrous acid	HNO ₂ (aq) ⇌ H ⁺ (aq) + NO ₂ ⁻ (aq)	
Ethanoic acid	CH ₃ COOH(aq) ⇌ CH ₃ COO ⁻ (aq) + H ⁺ (aq)	
Sulphuric acid	H ₂ SO ₄ (aq) → 2H ⁺ (aq) + SO ₄ ²⁻ (aq)	2 (Dibasic)
Sulphurous acid	H ₂ SO ₃ (aq) ⇌ 2H ⁺ (aq) + SO ₃ ²⁻ (aq)	
Carbonic acid	H ₂ CO ₃ (aq) ⇌ 2H ⁺ (aq) + CO ₃ ²⁻ (aq)	
Oxalic acid	H ₂ C ₂ O ₄ (aq) ⇌ 2H ⁺ (aq) + C ₂ O ₄ ²⁻ (aq)	
Phosphoric acid	H ₃ PO ₄ (aq) ⇌ 3H ⁺ (aq) + PO ₄ ³⁻ (aq)	3 (Tribasic)
Citric acid	C ₆ H ₈ O ₇ (aq) ⇌ 3H ⁺ (aq) + C ₆ H ₅ O ₇ ³⁻ (aq)	

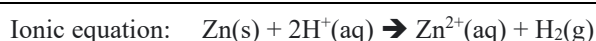
Intensive notes (Topic 4: Acids and bases)

Reactions of acids and bases

	Acid	Metal	O ²⁻ / OH ⁻	CO ₃ ²⁻ / HCO ₃ ⁻	NH ₃	NH ₄ ⁺	CO ₂ / SO ₂	Products
1.	✓	✓						Salt + H ₂
2.	✓		✓					Salt + H ₂ O
3.	✓			✓				Salt + H ₂ O + CO ₂
4.	✓				✓			NH ₄ ⁺ salt
5.			✓			✓		Salt + H ₂ O + NH ₃
6.			✓				✓	CO ₃ ²⁻ / SO ₃ ²⁻ salt + H ₂ O

Reaction 1

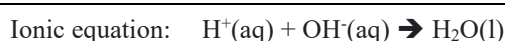
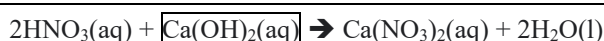
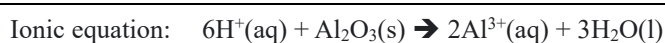
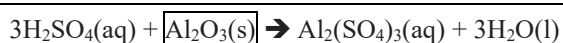
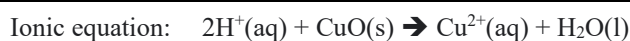
Acid + Metal → Salt + Hydrogen



Reaction 2

Acid + Metal oxide → Salt + Water

Acid + Metal hydroxide → Salt + Water



Reaction 3

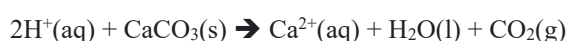
Acid + Carbonate salt → Salt + Water + Carbon dioxide

Acid + Hydrogencarbonate salt → Salt + Water + Carbon dioxide

Acid + Solid carbonate salt



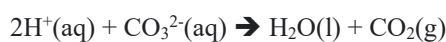
Ionic equation:



Acid + Aqueous carbonate salt



Ionic equation:



Acid + Solid hydrogencarbonate salt



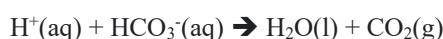
Ionic equation:



Acid + Aqueous hydrogencarbonate salt

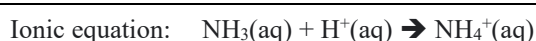
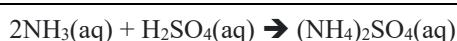


Ionic equation:



Reaction 4

Acid + Ammonia → Ammonium salt



Reaction 5

Ammonium salt + Metal hydroxide → Salt + Ammonia + Water

Ammonium salt + Metal oxide → Salt + Ammonia + Water

Solid ammonium salt + Base



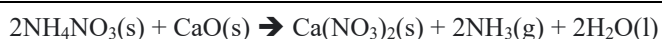
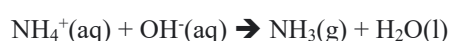
Ionic equation:



Aqueous ammonium salt + Base



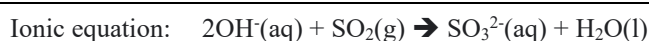
Ionic equation:



Ionic equation: ---

Reaction 6

Carbon dioxide (Sulphur dioxide) + Metal hydroxide → Metal carbonate (Metal sulphite) + Water



Intensive notes (Topic 4: Acids and bases)

Precipitation (Salt + alkali)

Cation	Color	Addition of NaOH(aq) / KOH(aq)	Addition of NH ₃ (aq)
K ⁺	Colorless	No observable change	No observable change
Na ⁺	Colorless	No observable change	No observable change
Ca ²⁺	Colorless	White precipitate forms $\text{Ca}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Ca}(\text{OH})_2(\text{s})$	No observable change
Mg ²⁺	Colorless	White precipitate forms $\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$	White precipitate forms $\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$
Al ³⁺	Colorless	White precipitate forms and it dissolves in excess NaOH(aq) to give colorless solution $\text{Al}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$ $\text{Al}(\text{OH})_3(\text{s}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{Al}(\text{OH})_4^{-}(\text{aq})$	White precipitate forms $\text{Al}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$
Zn ²⁺	Colorless	White precipitate forms and it dissolves in excess NaOH(aq) to give colorless solution $\text{Zn}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s})$ $\text{Zn}(\text{OH})_2(\text{s}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_4^{2-}(\text{aq})$	White precipitate forms and it dissolves in excess NH ₃ (aq) to give colorless solution $\text{Zn}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s})$ $\text{Zn}(\text{OH})_2(\text{s}) + 4\text{NH}_3(\text{aq}) \rightarrow \text{Zn}(\text{NH}_3)_4^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$
Fe ²⁺	Green	Green precipitate forms $\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$	Green precipitate forms $\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$
Fe ³⁺	Yellow	Brown precipitate forms $\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$	Brown precipitate forms $\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$
Pb ²⁺	Colorless	White precipitate forms and it dissolves in excess NaOH(aq) to give colorless solution $\text{Pb}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Pb}(\text{OH})_2(\text{s})$ $\text{Pb}(\text{OH})_2(\text{s}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Pb}(\text{OH})_4^{2-}(\text{aq})$	White precipitate forms $\text{Pb}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Pb}(\text{OH})_2(\text{s})$
Cu ²⁺	Blue	Blue precipitate forms $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$	Blue precipitate forms and it dissolves in excess NH ₃ (aq) to give deep blue solution $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$ $\text{Cu}(\text{OH})_2(\text{s}) + 4\text{NH}_3(\text{aq}) \rightarrow \text{Cu}(\text{NH}_3)_4^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$
Ag ⁺	Colorless	Dark brown precipitate forms $2\text{Ag}^{+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Ag}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l})$	Dark brown precipitate forms and it dissolves in excess NH ₃ (aq) to give a colorless solution $2\text{Ag}^{+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Ag}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l})$ $\text{Ag}_2\text{O}(\text{s}) + 4\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{Ag}(\text{NH}_3)_2^{+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$
NH ₄ ⁺	Colorless	Colorless pungent smell gas forms which can turn moist red litmus paper blue $\text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})$	No observable change

Intensive notes (Topic 4: Acids and bases)

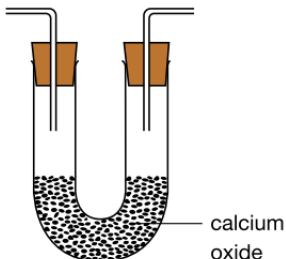
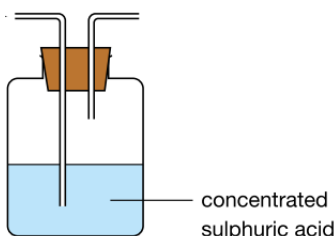
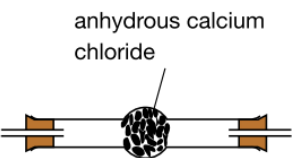
Precipitation (Salt + Salt / Acid)

Compound of	Solubility in water
NH_4^+ , Na^+ , K^+ , NO_3^- , HCO_3^-	All are soluble
Cl^- , Br^- , I^-	All are soluble (Except AgCl, AgBr, AgI, PbCl₂, PbBr₂, PbI₂)
SO_4^{2-}	All are soluble (Except BaSO₄, CaSO₄, PbSO₄)
CO_3^{2-}	All are insoluble (Except Na_2CO_3 , K_2CO_3 , $(\text{NH}_4)_2\text{CO}_3$)
O^{2-} / OH^-	All are insoluble (Except Na_2O , K_2O , NaOH , KOH) CaO , MgO , $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ are slightly soluble.


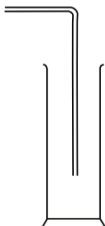
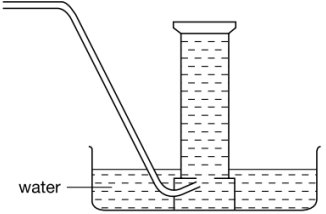
Consider the following examples:

Solution A	Solution B	Chemical equation (A is added to B)
$\text{AgNO}_3(\text{aq})$	$\text{Cl}^-(\text{aq})$	$\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
$\text{Pb}(\text{NO}_3)_2(\text{aq})$	$\text{Cl}^-(\text{aq})$	$\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$
$\text{Pb}(\text{NO}_3)_2(\text{aq})$	$\text{SO}_4^{2-}(\text{aq})$	$\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$
$\text{Ca}(\text{NO}_3)_2(\text{aq})$ / $\text{CaCl}_2(\text{aq})$ / $\text{Ca}(\text{OH})_2(\text{aq})$	$\text{SO}_4^{2-}(\text{aq})$	$\text{Ca}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{CaSO}_4(\text{s})$
$\text{Ba}(\text{NO}_3)_2(\text{aq})$ / $\text{BaCl}_2(\text{aq})$ / $\text{Ba}(\text{OH})_2(\text{aq})$	$\text{SO}_4^{2-}(\text{aq})$	$\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
$\text{Na}_2\text{CO}_3(\text{aq})$ / $\text{K}_2\text{CO}_3(\text{aq})$ / $(\text{NH}_4)_2\text{CO}_3(\text{aq})$	Aqueous salt [Except $\text{Na}^+(\text{aq})$, $\text{K}^+(\text{aq})$, $\text{NH}_4^+(\text{aq})$]	$2\text{M}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{M}_2\text{CO}_3(\text{s})$ $\text{M}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{MCO}_3(\text{s})$

Drying gas

 <p style="text-align: center;">calcium oxide</p>	 <p style="text-align: center;">concentrated sulphuric acid</p>	 <p style="text-align: center;">anhydrous calcium chloride</p>
Anhydrous CaO (Drying NH_3)	Concentrated H_2SO_4 (Drying acidic gases)	Anhydrous CaCl_2 (Drying acidic gases)

Collecting gas

		 <p style="text-align: center;">water</p>
Upward delivery / Downward displacement (Gases that are less dense than air)	Downward delivery / Upward displacement (Gases that are denser than air)	Displacement of water (Insoluble gases)

Intensive notes (Topic 4: Acids and bases)

Concentration calculation

$$\text{Molarity (mol dm}^{-3}\text{)} = \frac{\text{Mole (mol)}}{\text{Volume (dm}^3\text{)}}$$

$$\text{Concentration (g dm}^{-3}\text{)} = \frac{\text{Mass (g)}}{\text{Volume (dm}^3\text{)}}$$

Example 1

Solution X is prepared by dissolving 25.0 g of $\text{Fe}_2(\text{SO}_4)_3(\text{s})$ in 500.0 cm^3 water. (Formula mass of $\text{Fe}_2(\text{SO}_4)_3 = 399.9$)

- (a) Calculate the concentration of the solution, in g dm^{-3} .

Answer: Concentration (in g dm^{-3}) = $25 / 500 = 0.05 \text{ g dm}^{-3}$

- (b) Calculate the molarity of the solution.

Answer: Mole of $\text{Fe}_2(\text{SO}_4)_3 = 25 / 399.9 = 0.0625$

Molarity of $\text{Fe}_2(\text{SO}_4)_3 = 0.0625 / (500/1000) = 0.125 \text{ M}$

- (c) Calculate the concentration of $\text{Fe}^{3+}(\text{aq})$ in solution X.

Answer: Since 1 mole of $\text{Fe}_2(\text{SO}_4)_3$ can dissociate to give 2 moles of Fe^{3+}

Mole of $\text{Fe}^{3+}(\text{aq}) = 0.0625 \times 2 = 0.125$

Concentration of $\text{Fe}^{3+}(\text{aq}) = 0.125 / (500/1000) = 0.250 \text{ M}$

Example 2

Solution X contains 100 cm^3 of 0.5 M $\text{CaCl}_2(\text{aq})$

- (a) Express the concentration of CaCl_2 in g cm^{-3} . (Formula mass of $\text{CaCl}_2 = 111.1$)

Answer:

$$\frac{\text{mol}}{\text{dm}^3} \xrightarrow[\text{x 1000}]{\text{x Molar mass}} \frac{\text{g}}{\text{cm}^3}$$

Concentration in $\text{g cm}^{-3} = (0.5 \text{ mol} \times 111.1 \text{ g mol}^{-1}) / (1 \text{ dm}^3 \times 1000 \text{ cm}^3 \text{ dm}^{-3}) = 0.05555 \text{ g cm}^{-3}$

- (b) Calculate the mass of CaCl_2 in solution X.

Answer: Mole of $\text{CaCl}_2 = (0.5)(100/1000) = 0.05 \text{ mol}$

Mass of $\text{CaCl}_2 = 0.05 \times 111.1 = 5.555 \text{ g}$

Example 3

Solution X contains 100 cm^3 of 8 M $\text{HCl}(\text{aq})$

- (a) 400 cm^3 of distilled water is added to X. What is the concentration of HCl in the resulting mixture?

Answer: $(M_1)(V_1) = (M_2)(V_2)$

$(8)(100) = (M_2)(100 + 400)$

$M_2 = 1.6 \text{ M}$

- (b) What volume of water is required to dilute 100 cm^3 of 8 M HCl to a concentration of 2 M?

Answer: $(M_1)(V_1) = (M_2)(V_2)$

$(8)(100) = (2)(100 + v)$

$v = 300 \text{ cm}^3$

Intensive notes (Topic 4: Acids and bases)

Example 4

Solution X contains 200.0 cm³ of 1.0 M Na₂SO₄(aq).

- (a) 300.0 cm³ of 1.5 M NaCl(aq) is added to X. Calculate the concentration of Na⁺(aq) in the resulting mixture.

Answer: Concentration of Na⁺(aq) = Total mole of Na⁺(aq) / Total volume of solution

$$= [(200/1000)(1.0)(2) + (300/1000)(1.5)] / [(200/1000) + (300/1000)] = 1.7 \text{ M}$$

- (b) What volume of 0.1 M Na₃PO₄(aq) is required to prepare a mixture containing 0.64 M Na⁺(aq)?

Answer: Let v be the volume of 0.1 M Na₃PO₄(aq) required
 Concentration of Na⁺(aq) = Total mole of Na⁺(aq) / Total volume of solution

$$0.64 = [(200/1000)(1.0)(2) + (v)(0.1)(3)] / [(200/1000) + v]$$

$$v = 0.8 \text{ dm}^3$$

Example 5

A sample of HCl has a density of 1.18 g cm⁻³ and the percentage by mass of HCl in this sample is 36%. Calculate the concentration of HCl in mol dm⁻³. (Molecular mass of HCl = 36.5)

Answer: Take 1000 cm ³ for calculation Mass of HCl in 1000 cm ³ sample $= 1000 \times 1.18 \times 36\% = 424.8 \text{ g}$ Mole of HCl in 1000 cm ³ sample = $424.8/36.5 = 11.6$ Concentration of HCl = $11.6 / (1000/1000) = 11.6 \text{ M}$	In v cm ³ solution: $\text{Molarity} = \frac{(\text{volume in cm}^3)(\text{density in g cm}^3)(\%)}{\frac{(\text{Molar mass})(\text{volume in cm}^3)}{1000}}$
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Color of indicators

1. Methyl orange

Color	Red	Orange	Yellow
pH range	< 3.1	3.1 – 4.4	> 4.4

2. Phenolphthalein

Color	Colorless	Very pale pink	Pink
pH range	< 8.3	8.3 – 10	> 10

3. Litmus

Color	Red	Purple	Blue
pH range	< 5	5 – 8	> 8

Note: The above indicators do not show the pH value of the solution. To measure the pH of the solution, we can use universal indicator or pH meter (more accurate)

Acidity and alkalinity of an aqueous solution

Acidic solution	Neutral solution	Alkaline solution
[H ⁺ (aq)] > [OH ⁻ (aq)]	[H ⁺ (aq)] = [OH ⁻ (aq)]	[H ⁺ (aq)] < [OH ⁻ (aq)]

Note: [H⁺(aq)] and [OH⁻(aq)] are always present in aqueous solution. That is, there are still H⁺(aq) ions in a strongly alkaline solution and there are still OH⁻(aq) ions in a strongly acidic solution.

Intensive notes (Topic 4: Acids and bases)

pH calculation

$$\text{pH} = -\log [\text{H}^+(\text{aq})]$$

$$[\text{H}^+(\text{aq})] = 10^{-\text{pH}}$$

Example 1

- (a) Calculate the pH of 0.1 M HCl(aq)

Answer: $\text{pH} = -\log[\text{H}^+(\text{aq})] = -\log(0.1) = 1$

- (b) Calculate the pH of 0.1 M H₂SO₄(aq)

Answer: $\text{pH} = -\log[\text{H}^+(\text{aq})] = -\log(0.1 \times 2) = 0.699$

Example 2

- (a) The pH of a sample of HCl(aq) is 1.5. Calculate the concentration of HCl(aq)

Answer: $[\text{H}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-1.5} = 0.0316 \text{ M}$

Since $[\text{H}^+(\text{aq})] : [\text{HCl}(\text{aq})] = 1 : 1$, $[\text{HCl}(\text{aq})] = 0.0316 \text{ M}$

- (b) The pH of a sample of H₂SO₄(aq) is 1.5. Calculate the concentration of H₂SO₄(aq)

Answer: $[\text{H}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-1.5} = 0.0316 \text{ M}$

Since $[\text{H}^+(\text{aq})] : [\text{H}_2\text{SO}_4(\text{aq})] = 2 : 1$, $[\text{H}_2\text{SO}_4(\text{aq})] = 0.0316 \times 1/2 = 0.0158 \text{ M}$

Example 3

It is given that at 25°C, $[\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$.

- (a) Calculate the pH of 0.01 M NaOH(aq)

Answer: $[\text{H}^+(\text{aq})](0.01) = 1.0 \times 10^{-14}$

$$[\text{H}^+(\text{aq})] = 1.0 \times 10^{-12}$$

$$\text{pH} = -\log[\text{H}^+(\text{aq})] = -\log(1.0 \times 10^{-12}) = 12$$

- (b) Calculate the pH of 0.01 M Ca(OH)₂(aq)

Answer: $[\text{H}^+(\text{aq})](0.01 \times 2) = 1.0 \times 10^{-14}$

$$[\text{H}^+(\text{aq})] = 5.0 \times 10^{-13}$$

$$\text{pH} = -\log[\text{H}^+(\text{aq})] = -\log(5.0 \times 10^{-13}) = 12.3$$

Example 4

It is given that at 25°C, $[\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$.

- (a) The pH of a sample of KOH(aq) is 12.2. Calculate the concentration of KOH(aq).

Answer: $[\text{H}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-12.2} = 6.31 \times 10^{-13} \text{ M}$

$$(6.31 \times 10^{-13})[\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14}$$

$$[\text{OH}^-(\text{aq})] = 0.0158 \text{ M}$$

Since $[\text{OH}^-(\text{aq})] : [\text{KOH}(\text{aq})] = 1 : 1$, $[\text{KOH}(\text{aq})] = 0.0158 \text{ M}$

- (b) The pH of a sample of Ca(OH)₂ is 11.8. Calculate the concentration of Ca(OH)₂(aq)

Answer: $[\text{H}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-11.8} = 1.58 \times 10^{-12} \text{ M}$

$$(1.58 \times 10^{-12})[\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14}$$

$$[\text{OH}^-(\text{aq})] = 6.31 \times 10^{-3} \text{ M}$$

Since $[\text{OH}^-(\text{aq})] : [\text{Ca}(\text{OH})_2(\text{aq})] = 2 : 1$, $[\text{Ca}(\text{OH})_2(\text{aq})] = 6.31 \times 10^{-3} \times 1/2 = 3.15 \times 10^{-3} \text{ M}$

Intensive notes (Topic 4: Acids and bases)

Strength of acid and alkali

Strong acid: An acid which ionizes completely in water HCl, HNO ₃ , H ₂ SO ₄	Strong alkali: An alkali which dissociates completely in water *NaOH, *KOH, *Ca(OH) ₂
Weak acid: An acid which ionizes slightly in water #CH ₃ COOH, #H ₂ C ₂ O ₄ , H ₂ CO ₃ , H ₂ SO ₃ , H ₃ PO ₄	Weak alkali: An alkali which ionizes slightly in water NH ₃

*Ions are present in NaOH, KOH and Ca(OH)₂ already. They dissociate (not ionize) in water.

#Most organic acids are weak acids.

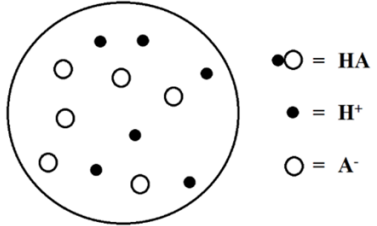
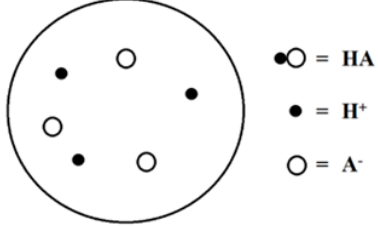
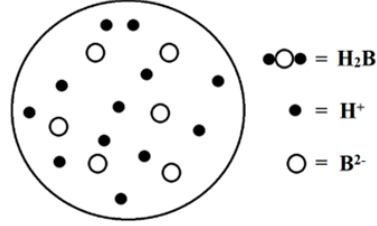
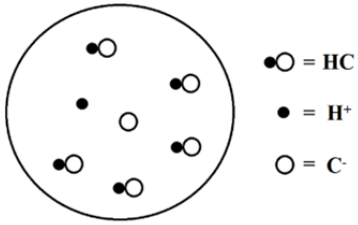
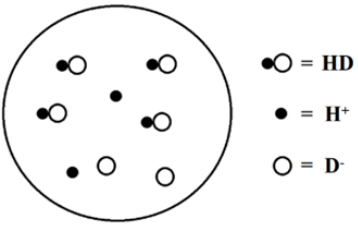
Ionization of strong acid HCl: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Ionization of weak acid CH₃COOH: $\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$

Ionization of weak alkali NH₃: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

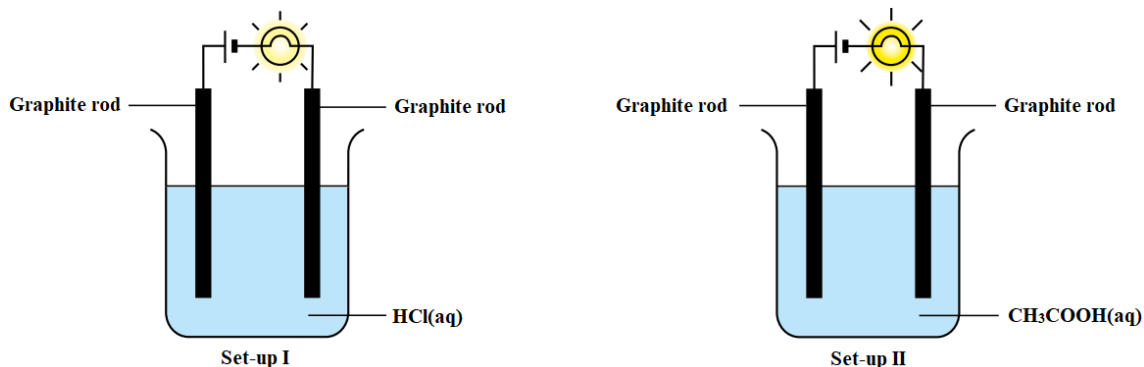
	100.0 cm ³ of 0.1 M HCl(aq)	100.0 cm ³ of 0.1 M CH ₃ COOH(aq)
Conductivity	Higher (as it contains higher concentration of mobile ions)	Lower (as it contains lower concentration of mobile ions)
pH	= 1.0 (As [H ⁺ (aq)] = [HCl(aq)])	> 1.0 (As [H ⁺ (aq)] < [CH ₃ COOH(aq)])
Mole of NaOH(aq) required for complete neutralization	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ 0.01 mol of NaOH required (As it contains 0.01 mole HCl)	$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$ 0.01 mol of NaOH required (As it contains 0.01 mole CH ₃ COOH)

Summary on strength, concentration and basicity

 <p>●○ = HA ● = H⁺ ○ = A⁻</p>	 <p>●○ = HA ● = H⁺ ○ = A⁻</p>	 <p>●○○ = H₂B ● = H⁺ ○○ = B²⁻</p>
1.0 M HA(aq) (A strong monobasic acid)	0.5 M HA(aq) (A strong monobasic acid)	1.0 M H ₂ B(aq) (A strong dibasic acid)
 <p>●○ = HC ● = H⁺ ○ = C⁻</p>	 <p>●○○ = HD ● = H⁺ ○ = D⁻</p>	
1.0 M HC(aq) (A weak monobasic acid)	1.0 M HD(aq) (A weak monobasic acid but stronger than HC)	

Experiment comparing strengths of acids and alkalis

Experiment 1



Observation: Set-up I has a higher electrical conductivity

Explanation: HCl is a stronger acid than CH₃COOH(aq). HCl ionizes completely in water while CH₃COOH(aq) ionizes only slightly in water. Solution in set-up I has a higher concentration of mobile ions.

Fair comparison: Both HCl(aq) and CH₃COOH(aq) should have same volume, concentration and temperature.

Experiment 2

(a) HA is a monobasic acid. At room conditions, the pH of a 0.05 M of HA(aq) is 1.30. Based on this information and appropriate calculation, comment whether HA is completely ionized in water.

Answer: Assume HA is a strong acid. $HA \rightarrow H^+(aq) + A^-(aq)$

$$[HA(aq)] : [H^+(aq)] = 1 : 1$$

$$[H^+(aq)] = 0.05 \text{ M}$$

$$pH = -\log[H^+(aq)] = -\log(0.05) = 1.3$$

As the pH of the solution agrees with the theoretical value, it is a strong acid.

(b) H₂B is a dibasic acid. At room conditions, the pH of a 0.005 M of H₂B(aq) is 4.1. Based on this information and appropriate calculation, comment whether H₂B is completely ionized in water.

Answer: Assume H₂B is a strong acid. $H_2B \rightarrow 2H^+(aq) + B^{2-}(aq)$

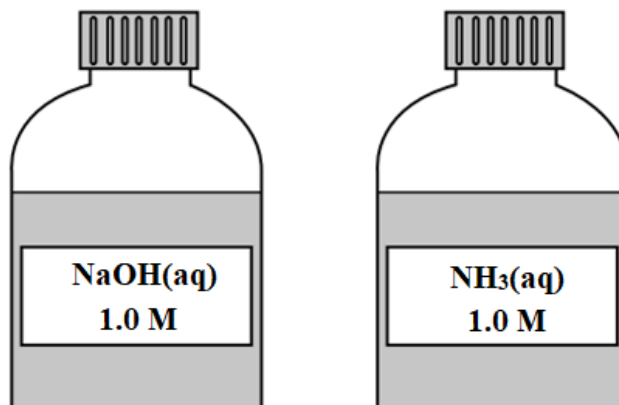
$$[H_2B(aq)] : [H^+(aq)] = 1 : 2$$

$$[H^+(aq)] = 0.005 \times 2 = 0.01 \text{ M}$$

$$pH = -\log[H^+(aq)] = -\log(0.01) = 2 < 4.1$$

As the pH of the solution is higher than the theoretical value, it is a weak acid.

Experiment 3

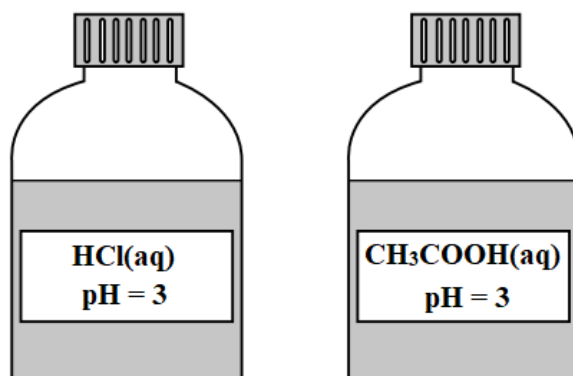


Deduce whether NaOH(aq) or NH₃(aq) has a lower pH.

Answer: NH₃(aq) has a lower pH. NaOH(aq) is a stronger alkali than NH₃(aq). NaOH dissociates completely in water while NH₃(aq) ionizes only slightly in water. NaOH(aq) has a higher concentration of OH⁻(aq).

Experiment 4

Consider the following two acids



Deduce whether HCl(aq) or CH₃COOH(aq) has a higher molarity.

Answer: For HCl(aq), [H⁺(aq)] = [Cl⁻(aq)] = 0.001 M. As HCl ionizes completely in water, [HCl(aq)] should be 0.001 M. For CH₃COOH(aq), [H⁺(aq)] = [CH₃COO⁻(aq)] = 0.001 M. As CH₃COOH ionizes slightly in water, [CH₃COOH(aq)] must be higher than [H⁺(aq)] (0.001 M). Therefore CH₃COOH should have a higher molarity.

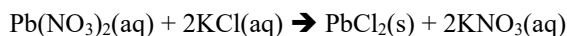
Intensive notes (Topic 4: Acids and bases)

Insoluble salt preparation

[Cation] nitrate solution + Potassium / sodium [anion] solution → Insoluble salt + Potassium / sodium nitrate solution

Example: Preparation of lead(II) chloride

Step 1. Mix lead(II) nitrate solution and potassium chloride solution



Step 2. Filter to obtain the precipitate (as residue)

Step 3. Wash the precipitate with large amount of distilled water.

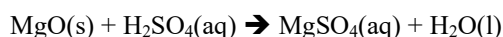
Step 4. Dry the precipitate with filter paper

Soluble salt preparation

Excess metal oxide (or metal hydroxide / metal carbonate / metal) + acid → Salt + Water

Example: Preparation of magnesium sulphate crystals

Step 1*. Add excess magnesium oxide to dilute sulphuric acid, then filter the mixture to obtain magnesium sulphate solution



Step 2^ . Heat the magnesium sulphate solution until it becomes saturated

Step 3# . Cool the solution slowly to obtain the crystals. Filter to obtain the crystals.

Step 4@ . Wash the crystals with small amount of cold distilled water

Step 5% . Dry the crystals with filter paper

* Excess MgO can ensure that all the sulphuric acid has reacted and the product is not contaminated by the acid.

^ To test whether the solution is saturated, we can dip a glass rod into the solution and take it out. If crystals are formed on it upon cooling, the solution is saturated

When the temperature of the solution drops, the solubility will decrease. Water cannot hold all the dissolved solute and some of the salt crystals will be formed

@ Washing the crystals can remove the water-soluble impurities. Using a small amount can reduce loss of the salt.

% It is not suitable to dry the crystals by heating. Otherwise water of crystallization may be removed / Salt may be decomposed upon heating.

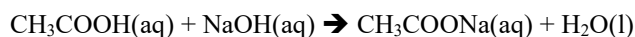
Group I / ammonium salt preparation

Acid + Alkali → Salt + Water

Acid + Metal carbonate → Salt + Water + Carbon dioxide

Example: Preparation of sodium ethanoate crystals

Step 1 Place a fixed volume of sodium hydroxide solution in a conical flask. Add a few drops of phenolphthalein. Run the ethanoic acid from the burette into the conical flask until the solution in the flask turns from pink to colorless



Step 2 Repeat step 1 with exactly the same volumes of acid and alkali used. However DO NOT add phenolphthalein this time

Step 3. Heat the sodium ethanoate solution until it becomes saturated

Step 4. Cool the solution slowly to obtain the crystals. Filter to obtain the crystals.

Step 5. Wash the crystals with small amount of cold distilled water

Step 6. Dry the crystals with filter paper

Applications of neutralization

1. Adjusting soil pH

Acidic soil can be neutralized by slaked lime ($\text{Ca}(\text{OH})_2$) or quicklime (CaO)

Alkaline soil can be neutralized by ammonium sulphate (which is also a fertilizer)

2. Neutralizing excess acid in stomach (antacids)

$\text{Mg}(\text{OH})_2$ and $\text{Al}(\text{OH})_3$ are weak bases in antacids. (Strong base such as NaOH are not used as it is too corrosive)

CaCO_3 and NaHCO_3 can also be used. However, $\text{CO}_2(\text{g})$ produced can lead to uncomfortable feeling in stomach

3. Treating industrial waste

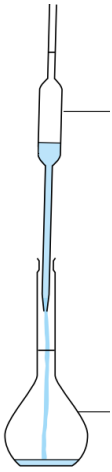
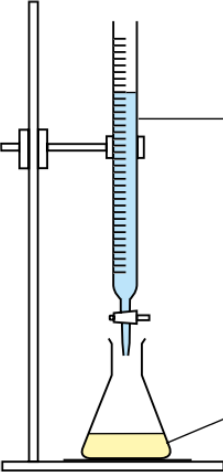
Acidic liquid waste should be treated with slaked lime ($\text{Ca}(\text{OH})_2$) or sodium carbonate

Alkaline liquid waste should be treated with dilute sulphuric acid

Note: Neutralization is NOT an application in acid / alkali burns. This is because neutralization is an exothermic reaction (a process which releases heat). In case of acid / alkali burns, put the affected area under cool running water.

Intensive notes (Topic 4: Acids and bases)

Titration

Dilution set-up		Titration set-up	
	<p>Pipette</p> <ul style="list-style-type: none"> - Rinse with distilled water and then the solution it is going to deliver - It is used to deliver a specific volume of a solution accurately. 		<p>*Burette</p> <ul style="list-style-type: none"> - Rinse with distilled water and then the solution it is going to deliver - It is used to deliver variable volumes (ranging from 0.05 to 50.00 cm³) of a solution accurately
	<p>Volumetric flask</p> <ul style="list-style-type: none"> - Rinse with distilled water - It is used to make up a solution to a specific volume accurately. 		<p>Conical flask</p> <ul style="list-style-type: none"> - Rinse with distilled water - It is used to hold the solution to be titrated.

* While preparing a burette filled with solution, make sure that:

1. The space between the tap and the tip is filled with the solution
2. No air bubbles inside the burette
3. The meniscus is not above the zero mark
4. The burette is in a vertical position

Standard solution preparation

Standard solution is a solution with accurately known concentration

Primary standard is a substance which can be used to prepare a standard solution directly. It can be weighed out accurately and used to prepare a solution of accurate concentration.

Standard solution can be used to determine the concentration / molar mass / percentage by mass of another reagent etc.

Properties of primary standard:

- Readily available in high purity
- Chemically stable
- Highly soluble in water
- Does not absorb water vapor
- Does not lose water of crystallization when hydrated
- Non-volatile and non-toxic

Example 1: Preparation of 250.0 cm³ standard 1.0 M Na₂CO₃(aq)

Mass of Na₂CO₃(s) needed = (1.0)(250/1000)(106) = 26.5 g

- Step 1 Dissolve 26.5 g of Na₂CO₃(s) in a sufficient amount of distilled water in a beaker
- Step 2 Transfer the solution into a 250.0 cm³ volumetric flask
- Step 3 Rinse the beaker with distilled water several times and transfer all the washings into the volumetric flask
- Step 4 Make up to the graduation mark with distilled water

Example 2: Preparation of standard 100.0 cm³ standard 0.3 M NaOH(aq) from standard 1.2 M NaOH(aq)

Volume of 1.2 M NaOH(aq) needed = (0.3 / 1.2)(100) = 25 cm³

- Step 1 Use a 25.0 cm³ pipette to transfer 25.0 cm³ of the 1.2 M NaOH(aq) to the 100.0 cm³ volumetric flask.
- Step 2 Make up to the graduation mark with distilled water

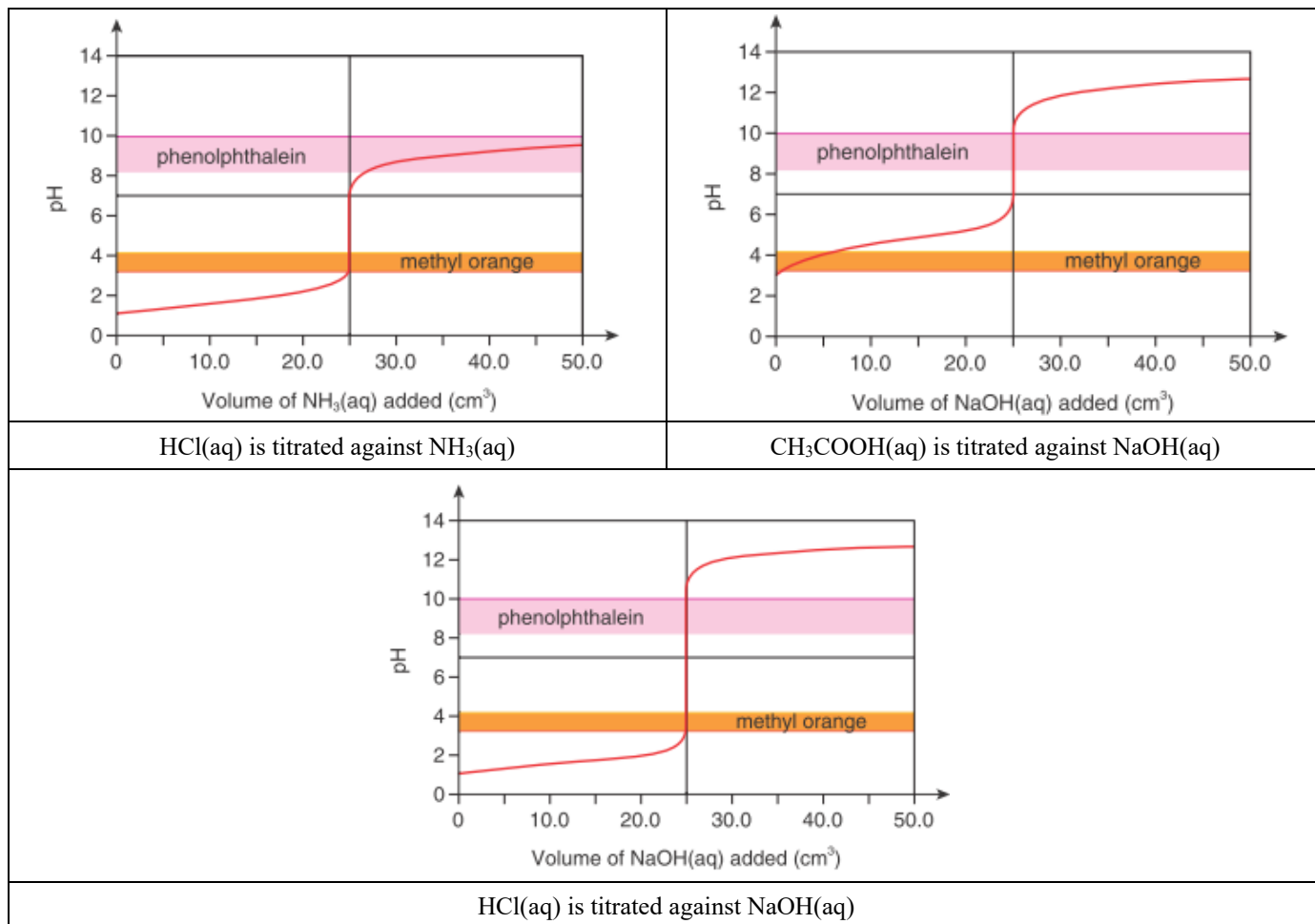
Intensive notes (Topic 4: Acids and bases)

pH curve and indicator

Equivalence point is the point at which an acid and an alkali have just completely reacted with each other.

End point is the point at which the indicator changes colour sharply.

The following titration curve



Choice of indicator

Titration between:	Strong acid	Weak acid
Strong base	Phenolphthalein and methyl orange	Phenolphthalein
Weak base	Methyl orange	No suitable indicator

*Methyl orange changes its color at around pH 3-5. Phenolphthalein changes its color at around pH 8-10.

Color change at end point

	Acid in conical flask	Acid in burette
Methyl Orange	From red to orange	From yellow to orange
Phenolphthalein	From colorless to pale pink	From pink to very pale pink

*Litmus solution is not used as indicator in titration as its colour change at end point is not sharp

Intensive notes (Topic 4: Acids and bases)

Titration calculations

$\text{Concentration} = \frac{\text{Mole}}{\text{Volume}}$	$\text{Basicity} = \frac{\text{Mole of alkali}}{\text{Mole of acid}}$
$\text{Molar mass} = \frac{\text{Mass}}{\text{Mole}}$	$\text{Percentage by mass} = \frac{(\text{Mole})(\text{Molar mass})}{\text{Total mass of sample}} \times 100\%$

Example 1A (Calculate the concentration of solution)

25.0 cm³ of NH₃(aq) sample is diluted to 100.0 cm³ with distilled water. 25.00 cm³ of 0.012 M H₂SO₄ was titrated against diluted NH₃(aq) in a burette. The table below shows the results of the titrations:

	Trial	1	2	3	4
Final burette reading / cm ³	15.00	12.80	13.90	14.20	12.40
Initial burette reading / cm ³	2.00	1.50	1.00	3.00	1.00

(a) Calculate the reasonable average of the diluted NH₃(aq) used

Answer: $[(12.8 - 1.5) + (14.2 - 3) + (12.4 - 1)] / 3 = 11.30 \text{ cm}^3$ (2 decimal places + unit!!!)
(Trial 2 is not reasonable, should not be included)

(b) Suggest a suitable indicator and state the color change at end point

Answer: Methyl orange. (Strong acid vs weak base, only methyl orange can be used)
From red to orange (11.30 cm³ NH₃ → burette contains NH₃ → conical flask contains H₂SO₄)

(c) Calculate the concentration of NH₃ sample

Answer: $2\text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$
Mole of H₂SO₄ = $(25/1000) \times 0.012 = 3 \times 10^{-4}$
Mole of NH₃ in the 25.0 cm³ sample = $3 \times 10^{-4} \times 2 \times 100/11.3 = 5.31 \times 10^{-3}$
Concentration of NH₃ = $5.31 \times 10^{-3} / (25/1000) = 0.212 \text{ M}$

(d) Why was the titration repeated several times?

Answer: To reduce errors of the titration results.

Example 1B (Calculate the concentration of solution)

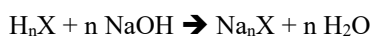
To find out the concentration of HCl, a laboratory technician first drew from the bottle a sample of 10.00 cm³ of the HCl and diluted it to 100.0 cm³ in a volumetric flask. The diluted acid sample was then used to titrate a standard Na₂CO₃(aq) placed in a conical flask. 10.00 cm³ of 1.06 mol dm⁻³ Na₂CO₃(aq) required 20.30 cm³ of the diluted acid sample to reach the end point. Using the titration result, calculate the concentration, in mol dm⁻³, of the HCl sample.

Answer: $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$
Mole of Na₂CO₃ = $(10/1000) \times 1.06 = 0.0106$
Mole of HCl in the 10.00 cm³ sample = $0.0106 \times 2 \times 100/20.3 = 0.1044$
Concentration of HCl = $0.1044 / (10/1000) = 10.44 \text{ M}$

Example 2 (Calculate the basicity of acid)

24.62 cm³ of 0.207 M NaOH(aq) can completely neutralize 25.00 cm³ of 0.102 M tartronic acid. Calculate the basicity of tartronic acid.

Answer: Let H_nX be the formula of tartronic acid where n is the basicity of acid.



Mole of alkali : Mole of acid = n : 1

$$[(24.62/1000) \times 0.207] : [(25/1000) \times 0.102] = n : 1$$

$$n = 2 \quad (\text{Tartronic acid is a dibasic acid})$$

Example 3 (Calculate the molar mass)

A student was given a sample of a water-soluble metal carbonate, M₂CO₃(s). In order to deduce what M was, the student prepared a 100.0 cm³ aqueous solution of the carbonate using 1.14 g of the sample. The student then withdrew several 10.0 cm³ portions of the solution and titrated each portion with 0.085 mol dm³ HCl(aq) using methyl orange as indicator. The mean titre was 25.30 cm³. Based on the experimental results, deduce what M is. (Relative atomic masses: C = 12.0, O = 16.0)



$$\text{Mole of HCl} = (25.3/1000) \times 0.085 = 2.15 \times 10^{-3}$$

$$\text{Mole of M}_2\text{CO}_3 \text{ in the sample} = 2.15 \times 10^{-3} \times 1/2 \times 100/10 = 0.01075$$

Molar mass = Mass / Mole

$$(2\text{M} + 12 + 16 \times 3) = 1.14 / 0.01075$$

$$\text{M} = 23 \quad (\text{Relative atomic mass of M} = 23. \text{ M should be sodium})$$

Example 4 (Calculate the water of crystallization from molar mass)

An experiment was performed to determine the number of water of crystallization, n, in a sample of hydrated sodium carbonate (Na₂CO₃·nH₂O). 2.86 g of the sample was dissolved completely in about 50 cm³ of deionized water in a conical flask. The solution obtained was alkaline and was immediately titrated with 0.58 M HCl(aq). It required 34.50 cm³ of the acid to reach the end point. Calculate the number of water of crystallization, n. (Relative atomic masses: H = 1.0, C = 12.0, O = 16.0, Na = 23.0)



$$\text{Mole of HCl(aq) used} = (34.5/1000) \times 0.58 = 0.02$$

$$\text{Mole of Na}_2\text{CO}_3 = 0.02 \times 1/2 = 0.01 = \text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$$

Molar mass = Mass / Mole

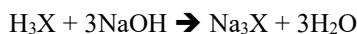
$$[23 \times 2 + 12 + 16 \times 3 + n(1 \times 2 + 16)] = 2.86 / 0.01$$

$$n = 10$$

Example 5 (Calculate the percentage by mass of a sample)

Citric acid is a tribasic acid found in lemon. A solid sample contained citric acid and other soluble inert substances. 1.65 g of the sample was dissolved in deionized water and diluted to 250.0 cm³. After that, 25.00 cm³ of the diluted solution was withdrawn and titrated with 0.123 M NaOH(aq) using phenolphthalein as an indicator. 18.45 cm³ of the NaOH(aq) was required to reach the end point. Calculate the percentage by mass of citric acid in the solid sample. (Molar mass of citric acid = 192.0 g)

Answer: Since citric acid is a tribasic acid, let H₃X be the formula of citric acid.



$$\text{Mole of NaOH} = (18.45/1000) \times 0.123 = 2.27 \times 10^{-3}$$

$$\text{Mole of citric acid in 1.65 g sample} = 2.27 \times 10^{-3} \times 1/3 \times 250/25 = 7.56 \times 10^{-3}$$

$$\text{Percentage by mass} = [(7.56 \times 10^{-3})(192) / 1.65] \times 100\% = 88.0\%$$

Example 6 (Back titration)

Eggshells mainly contain calcium carbonate and a small amount of organic substances. The percentage by mass of calcium carbonate in a sample of eggshell was determined. 0.204 g of the eggshell sample was put into a conical flask. After that, 25.00 cm³ of 0.200 M HCl(aq) was added. The mixture was heated for 15 minutes. After cooling down, the mixture was titrated with 0.102 M NaOH(aq). 16.85 cm³ of NaOH(aq) was needed to reach the end point of titration.

(a) Calculate the percentage by mass of calcium carbonate in the sample.

(Relative atomic masses: C = 12.0, O = 16.0, Ca = 40.1)



$$\text{Mole of NaOH} = (16.85/1000) \times 0.102 = 1.72 \times 10^{-3}$$

$$\text{Mole of excess HCl} = 1.72 \times 10^{-3} \times 1/1 = 1.72 \times 10^{-3}$$



$$\text{Mole of all HCl used} = (25/1000) \times 0.2 = 5 \times 10^{-3}$$

$$\text{Mole of HCl reacted with CaCO}_3 = 5 \times 10^{-3} - 1.72 \times 10^{-3} = 3.28 \times 10^{-3}$$

$$\text{Mole of CaCO}_3 \text{ in the sample} = 3.28 \times 10^{-3} \times 1/2 = 1.64 \times 10^{-3}$$

$$\text{Percentage by mass} = [1.64 \times 10^{-3} \times (40.1 + 12 + 16 \times 3)] / 0.204 \times 100\% = 80.5\%$$

(b) State ONE assumption in your calculation in (a).

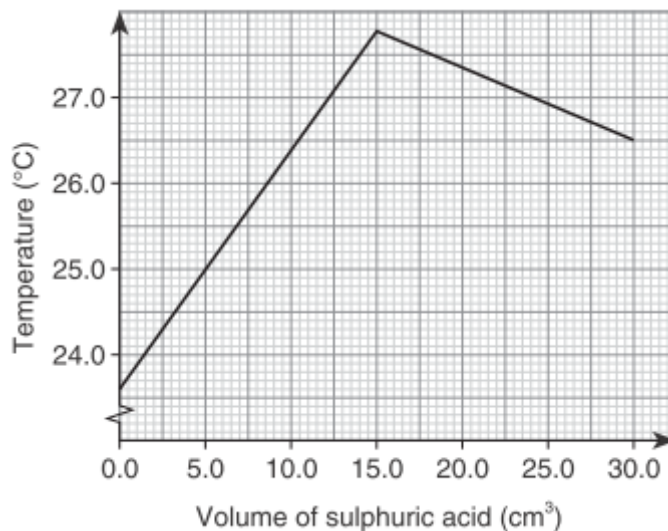
Answer: Only CaCO₃ reacted with HCl / The reaction between CaCO₃ and HCl was complete

(c) Explain why back titration, but not direct titration, is used in the experiment

Answer: Calcium carbonate is insoluble in water.

Example 7 (Temperature change in a titration curve)

In an experiment, 10.0 cm³ of NaOH(aq) is placed in an expanded polystyrene cup, and the initial temperature of the alkali is measured using a thermometer. After that, 0.10 M of H₂SO₄(aq) is added to the alkali, 1.00 cm³ at a time. The temperature of the reaction mixture is recorded immediately after each addition of the acid. The following graph is obtained.



- (a) Explain why NaOH(aq) should be placed in an expanded polystyrene cup.

Answer: To reduce the heat loss to the surrounding

- (b) Explain the temperature change in the above titration curve.

Answer: The temperature of the mixture rises as the reaction between NaOH(aq) and H₂SO₄(aq) is exothermic.

The temperature of the mixture reaches maximum at equivalence point (i.e. 15.0 cm³).

The temperature of the mixture then falls due to the addition of excess H₂SO₄.

- (c) Calculate the molarity of NaOH(aq) used.

Answer: $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

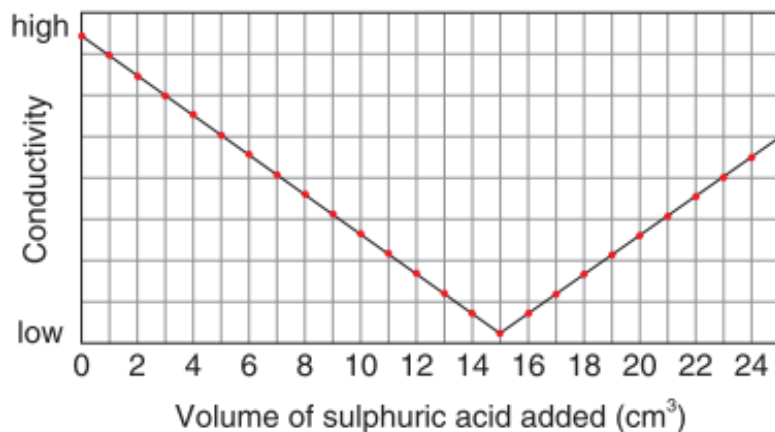
Mole of H₂SO₄ = (15/1000) × 0.1 = 0.0015

Mole of NaOH = 0.0015 × 2 = 0.003

Concentration of NaOH = 0.003 / (10/1000) = 0.3 M

Example 8 (Conductivity change in a titration curve)

In an experiment, 25.0 cm³ of Ba(OH)₂(aq) is placed in an expanded polystyrene cup, and the initial electrical conductivity of the alkali is measured. After that, 0.10 M of H₂SO₄(aq) is added to the alkali, 1.00 cm³ at a time. The conductivity of the reaction mixture is recorded immediately after each addition of the acid. The following graph is obtained.



- (a) Explain the electrical conductivity change in the above titration curve.

Answer: The conductivity of the mixture decreases. It is because:

1. Ba²⁺(aq) undergoes precipitation with SO₄²⁻(aq) to give BaSO₄(s)
2. OH⁻(aq) undergoes neutralization with H⁺(aq) to give H₂O(l)

Concentration of mobile ions in Ba(OH)₂(aq) decreases.

The conductivity of the mixture reaches minimum at equivalence point (i.e. 15.0 cm³).

The conductivity of the mixture then increases due to the addition of excess H₂SO₄(aq) which contains H⁺(aq) and SO₄²⁻(aq) for conducting electricity

- (b) Calculate the molarity of NaOH(aq) used.

Answer: Ba(OH)₂ + H₂SO₄ → BaSO₄ + 2H₂O

Mole of H₂SO₄ = (15/1000) × 0.1 = 0.0015

Mole of Ba(OH)₂ = 0.0015

Concentration of Ba(OH)₂ = 0.0015 / (25/1000) = 0.06 M