

Acids & Bases | Topic Notes

Examples of acid base reactions include - when a cake is baked, when we use hair conditioner, when we take antacid to relieve stomach upset, when lime is spread on soil.

Acids

- Sour taste
- turn moist blue litmus paper red
- react with metals to form hydrogen gas
- react with carbonates and hydrogen carbonates to liberate carbon dioxide
- Arrhenius definition is that an acid is a substance that dissociates in water to produce H⁺ ions.
- Bronsted Lowry defined an acid as a proton donor
- acids that donate 1 H⁺ ion are monobasic HCl, 2 H⁺ ions are dibasic H₂SO₄ and 3 H⁺ ions are tribasic, H₃PO₄.
- Strong acids fully dissociate in water. i.e. almost every molecule breaks up to give H⁺ ions.
e.g. ethanoic acid, acetic acid, methanoic acid
- Weak acids do not fully dissociate in water.
- Note that the H⁺ ion reacts with water molecules to form the H₃O⁺ ion or hydronium ion.
- **HA + H₂O ----> H₃O⁺ + A⁻**, eg. HCl + H₂O ----> H₃O⁺ + Cl⁻

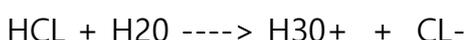
Bases

- Soapy feeling
- Turn red litmus blue (blue for Base)
- Arrhenius said that Bases were substances that fully dissociate in water to produce OH⁻ ions
- Caustic soda, to clean ovens and unblock drains
- Magnesium Hydroxide or Milk of Magnesia
- Calcium hydroxide or limewater detects presence of CO₂
- Many household cleaners contain soluble bases as degreasing agents. These convert oil and grease into soluble soaps that wash away easily.
- Strong bases - sodium hydroxide
- weak bases - calcium hydroxide and magnesium hydroxide
- A base that dissolves in water is an alkali
- Bronsted Lowry said that bases were proton acceptors

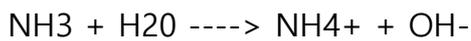
Limitations of Arrhenius theories:

- It is hydronium ions not the bare H⁺ ions suggested by Arrhenius
- Arrhenius restricted his definitions to aqueous solutions only and excluded solvents like liquid ammonia, toluene etc...
- Not all acid base reactions require water, when ammonia gas and HCl gas come into contact white fumes of ammonium chloride are formed.

An amphoteric substance can act as either an acid or a base. e.g. water



H₂O accepts a proton here



H₂O donates a proton here.

An acid changes into a conjugate base when it donates a proton.

A base changes into a conjugate acid when it accepts a proton.

A **conjugate acid base pair** consists of an acid and a base that differ by 1 proton

Neutralisation

Neutralisation is the reaction between an acid and a base to form salt and water.

Salt is a general term to describe the substance formed when the hydrogen in an acid is replaced by a metal or ammonium ion.

- **Medicine** - sometimes the stomach produces too much hydrochloric acid and the lining of the stomach may be damaged and an ulcer may form. Antacid contains a base which neutralises the HCl in the stomach and can be taken in tablet form. Bisodol, alka setzer - sodium hydrogencarbonate. ($\text{HCl} + \text{NaHCO}_3 \rightarrow \text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$)
- Milk of magnesia contains magnesium hydroxide ($\text{Mg(OH)}_2 + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O}$)
- **Agriculture** - if soil is too acidic crop yields tend to be low so farmers tend to spread lime on the soil to neutralise soil acidity.
- **Environmental protection** - In areas that suffer acid rain limestone is often added to lakes to neutralise acidity. Limestone is also used in tall chimneys in coal burning plants to neutralise the acidic emissions or substances formed from burning coal.

- **Miscellaneous** - toothpaste neutralises acids in food to reduce tooth decay, baking soda neutralises acidic bee stings, vinegar neutralises alkaline wasp stings,

Volumetric Analysis - Acid/Base

(involves reacting together volumes of solutions it is a form of quantitative analysis).

- Concentration of a solution is the amount of solute that is dissolved in a given volume of solution.
- w/w % (g/100g)
- w/v % (g/100cm³)
- v/v % (cm³/cm³)
- ppm (1mg/l)
- 1 molar solution contains 1 mole of solute dissolved in 1 litre of solution.
- decimolar = 0.1 moles of solute in 1 litre solution
- find relative molecular mass and multiply by moles per litre to get grams per litre
- number of moles = (volume X molarity)/1000
- Number of moles of solute in DILUTE solution = number of moles of solute in CONCENTRATED solution
- **(Vol(dil) X Mol(dil))/1000 = (Vol(conc) X Mol(conc))/1000**
- **A primary standard is a substance which can be obtained in a stable, pure and soluble solid form so that it can be weighed out and dissolved in water to give a solution of accurately known concentration.**
- a standard solution is a solution of precisely known concentration
- Must be available in a highly pure state

- Must dissolve easily in water
- Must be stable in air
- Fairly high relative molecular mass so that precision in weighing out a known mass is high as possible
- When used in volumetric analysis it should undergo complete and rapid reaction
- Should not be hydrated
- e.g. sodium chloride, potassium dichromate, anhydrous sodium carbonate.

1. To prepare a standard solution of sodium carbonate.

- Here we weigh out accurately a known amount of sodium carbonate and dissolve it in water and make the solution up to a known volume. Sodium carbonate is a **primary standard** so the concentration of the solution will be known precisely.
- Place a clean dry clock glass on the pan of the laboratory balance, adjust the balance so it is tared off or reads at 0.
- Use a clean dry spatula and add about 1.3g sodium carbonate to the clock glass. Ensure the weight is known precisely, i.e. within 0.01g or better.
- Add the sodium carbonate to a beaker of 100cm³ deionized water. Use a wash bottle, wash all traces of sodium carbonate left on the clock glass into the beaker. Always use deionized water because impurities from the tap water could interfere with the solution.

- Stir the mixture with a glass rod until the sodium carbonate dissolves completely, using the rod and a funnel transfer it to a 250cm³ volumetric flask and use the wash bottle again to wash all traces of the sodium carbonate from the beaker, rod and funnel into the volumetric flask.
- Add deionised water to the flask within about 1 cm of the graduation mark. Then, using a dropper add deionized water until the bottom of the meniscus lies on the graduation mark, read at eye level.
- Place the stopper in the flask, invert 20 times to form a homogenous solution i.e. so the concentration is the same through out.

- **Calculations:**

$$\text{Na}_2\text{CO}_3 - \text{RMM} = 23(2) + 12 + 16(3) = 106\text{g}$$

i.e. **106g in 1 litre is a 1 molar solution**

1 g in 1 litre = 0.00943 molar solution

1.3g in 1 litre is 0.0122

1.3g in 250cm³ is 0.0122 X 4 = 0.0488 molar solution

To use a standard solution of sodium carbonate to standardise a given hydrochloric acid solution.

- The idea is here that by titrating the known volume of sodium carbonate solution with a known concentration, we can determine the concentration of the HCL solution. Methyl orange indicator indicates the end point of the titration.
- note: usually in **acid base titrations the acid is in the burette and the base is in the conical flask**. This is because if bases were in a burette they could cause the tap to stick, sodium hydroxide solution also absorbs carbon dioxide from the air and forms a crust around the tap - seizes up the burette taps, attacks glass.
- Wash out the pipette, burette and conical flask with deionized water.
- Pour 100 cm³ of the sodium carbonate solution into a beaker and 100cm³ of HCL into a separate beaker. Wash out the pipette with the sodium carbonate solution and then pipette 25cm³ of it into a conical flask. **It is better to pour into a beaker from the volumetric flask before using the pipette as impurities from the pipette could contaminate the entire solution.**
- Add three drops of methyl orange indicator to the conical flask and note the yellow color. Methyl orange is yellow in neutral or basic solutions and red in acidic solutions. Do not use too much indicator because intense color can make it hard to distinguish a color change at the end point.
- Using a funnel, fill the vertically clamped burette almost completely the HCL solution. Open the tap briefly and let some flow into a beaker marked waste so the space below the tap is

filled. Remove funnel and if needed, use a dropper so that the bottom of the meniscus sits on 0 mark.

- Place a white tile on base of conical flask so a color change is indicated more easily.
- Using the thumb and first 2 fingers of the left hand open the tap and allow the acid to run fairly quickly into the conical flask. Swirl and wash the sides down of the flask with deionized water to ensure no drops of acid are stuck to sides of flask. It is important that all acid added reacts. Note that adding **water (deionized) does NOT affect the number of moles of sodium carbonate in the flask.**
- When it turns pink stop the titration and note the volume of acid added. This rough titration tells you approximately where the end point is.
- Perform 2 accurate titrations and add the acid drop by drop as you approach the end point.
- The titre is the titration result taking the average of the 2 readings which agree within 0.1 cm³ of each other. This is the **mean titre.**

Calculations:

We made a 0.0488 molar solution of sodium carbonate and titrated it against HCL in 25 cm³ portions. Mean titre was 19.55 (mean of 2 accurate titrations). The equation is



- For every mole of base there is 2 moles of acid so the ratio is 2:1

$$\frac{\text{Volume of acid} \times \text{molarity of acid}}{\text{No. of moles of acid}} = \frac{\text{Volume of base} \times \text{molarity base}}{\text{no of moles of base}}$$

$$V_a = 19.55 \text{ cm}^3$$

$$M_a = ?$$

$$N_a = 2$$

$$V_b = 25 \text{ cm}^3$$

$$M_b = 0.0488$$

$$N_b = 1$$

$$19.55x/2 = 1.22$$

$$19.55x = 2.44$$

$$x = 0.1248 \text{ moles per litre}$$

$$0.1248 \times 36.5 \text{ g (Mr of HCL)} = 4.55 \text{ g/l}$$

To make up an approximate 0.1M solution of sodium hydroxide, to standardize it (obtain exact concentration) with a standard HCL solution and hence prepare a sample of sodium chloride.

- The standardized HCL solution prepared previously is a SECONDARY STANDARD. It can now be used to standardize other solutions. Sodium hydroxide absorbs water vapour from air, it cannot be weighed in fixed amounts to give a solution of accurately known concentration.
- **HCL + NaOH ----> NaCl + H₂O**
- Weigh out about 1g of sodium hydroxide in a clock glass, the mass need not be accurately known.
- Transfer to a beaker of 100cm³ deionized water and transfer washings from clock glass and spatula. Stir until dissolved and then transfer using a funnel into a 250 cm³ volumetric flask. Add the washings from the glass rod, funnel and beaker. Add deionized water to within about 1 cm of the graduation mark and using a dropper, ensure that the bottom of the meniscus sits on the graduation mark. Invert 20 times and stopper to form a homogenous solution.
- Pour about 100cm³ of this into a clean dry beaker and 100cm³ of the HCl into another clean dry beaker.
- Rinse the pipette with deionized water and then some of the sodium hydroxide solution. Pipette about 25cm³ of it into a conical flask and add three drops of methyl orange indicator to this.
- Using a small funnel fill the burette with dilute HCl, open tap briefly to fill the space below it and use a dropper to let bottom of meniscus sit on graduation mark.

- Use a white tile and titrate, the color change is from **yellow to red**.
- 1 rough and 2 accurate titrations.
- Repeat above titration and leave out indicator, add the mean titre and stop. Pour the contents of the conical flask into an evaporating dish and evaporate almost to dryness over a hotplate, white crystals of sodium chloride formed.

To determine the percentage of ethanoic acid in vinegar.

- Vinegar is a dilute solution of ethanoic acid, the amount present can be determined by titrating the acid against a solution standard of sodium hydroxide.
- Phenolphthalein indicator is used to titrate a weak acid against a strong base. The color change is from pink to colorless.
- **$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$**
- Dilute the sample of vinegar 5 times, i.e. pipette two 25 cm^3 samples of the vinegar into a 250 cm^3 volumetric flask, make up to mark with deionized water.
- Wash out pipette with sodium hydroxide solution after having washed it out with water. Pipette 25 cm^3 sodium hydroxide solution into a conical flask and add 3 drops of phenolphthalein indicator.
- Rinse the burette with water and using a small funnel, fill it with the diluted vinegar.
- Use a white tile and perform 1 rough and 2 accurate titrations.

• Calculations:

average titre was 14.2 cm^3 . Calculate the concentration of ethanoic acid in mol/l, g/l, w/v

$$V_a = 14.2$$

$$M_a = X$$

$$N_a = 1 \text{ (1:1 ratio)}$$

$$V_b = 25$$

$$M_b = 0.12$$

$$N_b = 1$$

$$14.2 X = 25(0.12)$$

$X = 0.211$ moles per litre

multiply this by 5 as it was diluted by a factor of 5

$X = 1.055$ moles per litre

$\text{CH}_3\text{COOH} = 12 + 3 + 12 + 16 + 16 + 1 = 60\text{g rmm}$

$1.055 \times 60 = 63.3\text{g/l}$

If there's 63 g in 1000 cm^3 (1L) then there is 6.33g in 100 cm^3 - 6.33% w/v

To determine the percentage of crystallization in washing soda or hydrated sodium carbonate.

- Water of crystallization is the water that makes up the crystal structure of the compound.
Crystals that contain water of crystallization are hydrated.
- $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$
- Here we make up a solution of washing soda and titrate it against a HCL solution of known concentration.
- Methyl Orange goes from **Yellow to Pink**
- 5g crystals taking the usual precautions > 100cm³ water > 250cm³ volumetric flask.
- $\text{Na}_2\text{CO}_3 + 2\text{HCL} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$

• Calculations:

$$V_a = 28.7$$

$$M_a = 0.21$$

$$N_a = 2$$

$$V_b = 25\text{cm}^3$$

$$M_b = X$$

$$N_b = 1$$

$$28.7(0.21)/2 = 25X$$

$$3.0135/25 = X$$

$$X = 0.12054 \text{ moles per litre}$$

If there is 0.121 moles per litre there must be $0.121/4 = 0.03$ moles per 250 cm³.

1 mole of $\text{Na}_2\text{CO}_3 = 106\text{g}$

0.03 moles = 3.18g

We weighed out 5 g of it so $5 - 3.18 = 1.82$ g of water.

$1.82/5 \times 100 =$ percentage of water of crystallization **(actual grams of water/total grams of washing soda X 100) = 36.4%**

Calculate the **value of X - get the relative molecular mass**

there is 0.03 moles in 5 g of sodium carbonate - $5/0.03 = 167$ g

$\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 167$

$106 + 18x = 167$

$18x = 61$

$x = 3$

Instruments for volumetric analysis:

I. Graduated Cylinder

- Only used for measuring out APPROXIMATE volumes of liquids, i.e. correct within 0.5 cm^3 , always rinse out with deionized water before use.

II. Volumetric Flask

- Rinse out with deionized water before use, designed to contain a DEFINITE volume of solution.
- The last few cm^3 must be added carefully so the bottom of meniscus lies on the engraved graduation mark, use a small dropper for this.

- After making the solution the contents must be well mixed to form a homogenous solution. This is a solution that is uniform in concentration throughout, otherwise solution in neck may have a different concentration than the body. stopper and invert 20 times.

III. Pipette

- Designed to deliver an EXACT volume of solution, usually 20/25 cm³.
- Use a pipette filler to draw the solution above the graduation mark. This suction it.
- Liquid is allowed to run slowly from the pipette into a waste beaker until the bottom of the meniscus sits on the graduation mark. The tip is touched against the side of the beaker to remove any drops adhering to the tips.
- Liquid is allowed to discharge slowly into the conical flask, tip of pipette in contact with the side of the flask to avoid splash of liquid.
- Small amount of liquid will remain in the tip DO NOT BLOW, it has been taken into account when the pipette was calibrated at manufacturing.
- Rinse with deionized water then with the solution it is to contain. This removes drops of water in the pipette that would further dilute the solution.

IV. Burette

- Always check tap to see it turns freely.
- Rinse with deionized water and then with the solution it is to contain to prevent further dilution.
- Open tap briefly over a waste beaker to ensure the space below the tap is filled, check to ensure that there are no air bubbles trapped in it.

- Should be clamped vertically, remember to fill with a funnel and dropper if necessary, piece of white paper used, read from **TOP** of meniscus.

V. Conical Flask

- Designed to be swirled without the splashing out of liquid.
- The shape allows droplets of solution adhering to sides to be washed down easily.
- It should NOT be rinsed with solution to be contained, the pipette places the EXACT volume in the flask, if drops were already there it would increase volume and hinder results.

A titration is a laboratory procedure whereby a measured volume of one solution is added to a known volume of another solution until the reaction is complete. The stage when both solutions react completely with each other is the equivalence point, when this is shown by an indicator, it is the END point.

Acid-Base titrations

Concentration of Solutions

Molarity = moles per litre of solution (mol l^{-1})

$$\text{molarity} = \frac{\text{mass per litre}}{M_r (\text{g mol}^{-1})}$$

grams of solute per litre (g l^{-1})

parts per million (ppm) = mg l^{-1} ($1\text{g} = 1000\text{mg}$)

$$\text{moles} = \frac{\text{volume (cm}^3\text{)} \times \text{concentration (M)}}{1000}$$

$$\text{moles} = \frac{\text{mass (g)}}{M_r (\text{g mol}^{-1})}$$

can be expressed as percentages:

$$\% \text{ w/v} = \text{g}/100\text{cm}^3$$

$$\% \text{ w/w} = \text{g}/100\text{g}$$

$$\% \text{ v/v} = \text{cm}^3/100\text{cm}^3$$

diluted solutions: $\text{moles}_{\text{dil}} = \text{moles}_{\text{conc}}$

$$V_{\text{dilute}} \times M_{\text{dilute}} = V_{\text{conc}} \times M_{\text{conc}}$$

Preparation of Standard Solutions and Titration Procedure

A standard solution is a solution of accurately known concentration (made using a primary standard solute which is 100% pure, stable and soluble).

A primary standard solute is a pure compound with a high molecular mass, from which solutions of a known concentration can be made. It must be 100% pure, stable and soluble.

Preparation:

- Weigh the solute accurately
- Transfer the solute to a clean beaker and add a small volume of pure water (less than 100cm³)
- Rinse the clock glass using the wash bottle filled with pure water and add the rinsings to the beaker
- Stir the mixture until the solute is dissolved completely
- Rinse the stirring rod into the beaker using the wash bottle
- Using a funnel, transfer the solution from the beaker into the volumetric flask
- Rinse the beaker several times with the wash bottle and add all of the rinsings to the volumetric flask
- Rinse and remove the funnel
- Fill the volumetric flask with pure water to within 1cm of the calibration mark
- Add pure water dropwise until the bottom of the meniscus rests on the calibration mark when read at eye level
- Seal the flask and invert 20+ times to ensure a homogeneous solution

Sources of Error:

- Volumetric flask must be used as when filled to the calibration mark at the temperature stated on the flask it contains a specific known volume
- Analytical balance must be used to measure masses of solute as it is extremely accurate
- Ensure no solution is lost when the solid is being dissolved in a beaker or when the solution is transferred from the beaker to the volumetric flask by using a wash bottle to transfer all rinsings