

ESAT

Chemistry

Acids, Bases and Salts

Free lesson notes

Covers

C9.1 to C9.3

Format

Lesson, worked examples, self-check

Price

Free

About this lesson

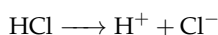
Acids are in your kitchen (vinegar, lemon juice), your stomach (hydrochloric acid) and every laboratory; bases and alkalis are in soap, bleach and indigestion tablets. When an acid meets a base they neutralise each other and make a salt and water, and this simple idea underlies a large part of ESAT Chemistry: it links to writing equations, to calculations with moles and concentrations, to rates of reaction and to energy changes (neutralisation gives out heat).

The lesson defines acids and bases in terms of the hydrogen ion, H^+ , which is the one idea that explains everything else. It then separates two pairs of words that are constantly confused, *strong/weak* and *concentrated/dilute*, and explains the pH scale (including the fact that each pH unit is a factor of ten). We then cover the reactions of acids with metals, carbonates, bases and alkalis, how neutralisation is written as an ionic equation, why it is exothermic, and how a soluble salt is made. Specification codes: C9.1, C9.2 and C9.3. Relative atomic masses are given where a calculation needs them.

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1 What is an acid?

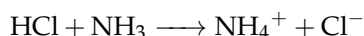
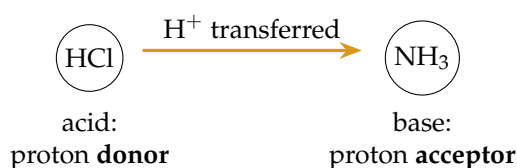
Acids give hydrogen ions. An **acid** is a substance that releases hydrogen ions, H^+ , when it dissolves in water. A hydrogen atom is one proton and one electron; removing the electron leaves a bare proton, so H^+ is a *proton*, and an acid is a **proton donor**. Hydrochloric acid, for instance, separates into ions in water:



It is the H^+ ions that give all acids their typical properties: sour taste (never taste laboratory acids), the ability to turn indicator red, and the same set of reactions with metals, carbonates and bases.

Common laboratory acids. Hydrochloric acid HCl , sulfuric acid H_2SO_4 , nitric acid HNO_3 , and ethanoic (acetic) acid CH_3COOH , found in vinegar. (The names of their salts come from the acid: hydrochloric acid gives *chlorides*, sulfuric acid gives *sulfates*, nitric acid gives *nitrates*, ethanoic acid gives *ethanoates*.)

Monoprotic, diprotic, triprotic. The number of H^+ ions that one molecule of acid can release is its **basicity**. HCl and HNO_3 release one: they are **monoprotic**. H_2SO_4 can release two: **diprotic**. Phosphoric acid, H_3PO_4 , can release three: **triprotic**. It matters because a diprotic acid needs twice as much alkali to neutralise it, mole for mole. Careful: it is the number of hydrogens that can be *released as ions* that counts. Ethanoic acid, CH_3COOH , has four hydrogens but only the one in the $-COOH$ group is released, so it is monoprotic.



Acids in water. Pure hydrogen chloride gas contains no ions. The ions form when it dissolves; in water the hydrogen ion is attached to a water molecule (H_3O^+) but we write $H^+(aq)$ for simplicity.

Worked example: basicity

How many moles of sodium hydroxide neutralise exactly 0.5 mol of sulfuric acid? $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$.

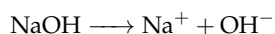
Sulfuric acid is diprotic, and the equation shows 1 : 2. So 0.5 mol of acid needs $2 \times 0.5 = 1.0$ mol of NaOH.

Common mistakes

- Thinking pure acid contains H^+ . The ions form when the acid dissolves in water (the symbol (aq)).
- Counting all hydrogen atoms: ethanoic acid is not tetraprotic; only the acidic hydrogen is released.
- Calling H^+ an "atom". It is a hydrogen *ion* (a proton).

2 Bases and alkalis

Bases accept hydrogen ions. A **base** is a substance that neutralises an acid by accepting H^+ ions: a **proton acceptor**. Typical bases are metal oxides and hydroxides (for example CuO , MgO , NaOH , Ca(OH)_2), metal carbonates and ammonia, NH_3 . An **alkali** is a base that is *soluble in water*, and it releases hydroxide ions, OH^- , in solution:



So every alkali is a base, but not every base is an alkali: copper(II) oxide is a base (it neutralises acids) but is insoluble, so it is not an alkali.

How a base accepts H^+ . The hydroxide ion takes the hydrogen ion to make water: $\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$. The oxide ion does the same: $\text{O}^{2-} + 2\text{H}^+ \longrightarrow \text{H}_2\text{O}$. The carbonate ion gives water and carbon dioxide: $\text{CO}_3^{2-} + 2\text{H}^+ \longrightarrow \text{H}_2\text{O} + \text{CO}_2$. Ammonia accepts an H^+ directly to make the ammonium ion, $\text{NH}_3 + \text{H}^+ \longrightarrow \text{NH}_4^+$; in water it also produces a small amount of OH^- , so it is an alkali, but a weak one (see below).

Worked example: base or alkali

Classify each as base only, or base and alkali: (a) MgO (insoluble), (b) KOH (soluble), (c) NH_3 (soluble). All three accept H^+ , so all are bases. (b) and (c) are soluble and give OH^- in water, so they are alkalis; (a) is insoluble, so base only.

Exam tip

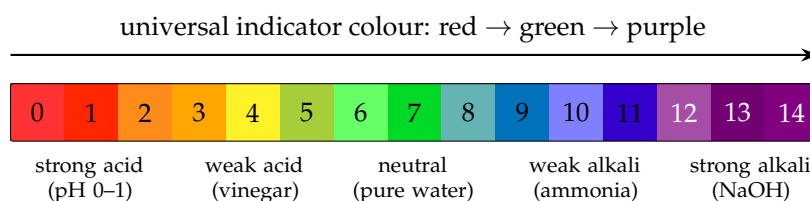
Remember the pair: *acid = proton donor, base = proton acceptor*. Many ESAT questions ask which species is acting as the acid or the base in a reaction: it is the one that gives, or takes, H^+ .

Common mistakes

- Writing that all bases dissolve in water. Only alkalis do.
- Thinking that alkalis contain H^+ . They contain (or produce) OH^- ; low H^+ is why their pH is high.
- Saying that ammonia "contains hydroxide". It becomes an alkali by taking H^+ from water, leaving OH^- .

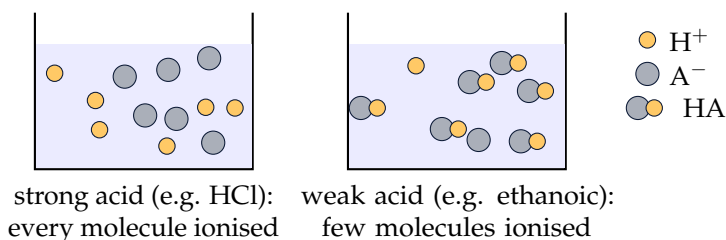
3 Strength, concentration and the pH scale

The pH scale. The pH scale measures how acidic or alkaline a solution is, usually from 0 to 14. Below 7 is acidic, 7 is neutral (pure water) and above 7 is alkaline. A low pH means a high concentration of H^+ . Universal indicator shows the pH by its colour; a pH meter reads it directly.

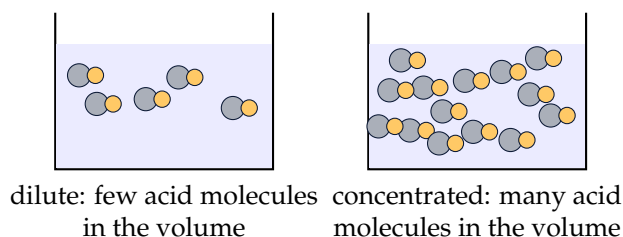


A factor of ten. The scale is logarithmic: *each step of 1 in pH changes the H^+ concentration by a factor of ten*. A solution of pH 3 has ten times the concentration of H^+ of one at pH 4, and 100 times that of pH 5. Going from pH 1 to pH 4 means the H^+ concentration falls by $10 \times 10 \times 10 = 1000$.

Strong and weak: a statement about ionisation. A **strong acid** ionises completely in water: every molecule releases its H^+ (hydrochloric, sulfuric and nitric acids). A **weak acid** is only partially ionised: most of its molecules stay intact and only a small fraction release H^+ (ethanoic acid, citric acid). The ionisation of a weak acid is a reversible reaction: $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$.



Concentrated and dilute: a statement about how much acid. **Concentrated** means a large amount of acid in a given volume; **dilute** means a small amount. These words say nothing about how far the acid ionises.



So you can have a dilute strong acid (0.001 mol/dm^3 HCl, pH 3) and a concentrated weak acid (several mol/dm^3 ethanoic acid). The two ideas are independent. At the *same concentration*, a strong acid has the lower pH, because it releases more H^+ , and it reacts faster, but equal volumes and concentrations of two monoprotic acids will neutralise the *same* amount of alkali, because the weak acid releases more H^+ as the reaction proceeds (the equilibrium shifts).

Worked example: diluting by a factor of ten

10 cm^3 of hydrochloric acid of pH 2 is added to water to make 1000 cm^3 . What is the pH of the new solution? The volume is increased 100 times, so the concentration of H^+ falls by a factor of $100 = 10 \times 10$. Each factor of 10 raises the pH by 1, so the pH rises by 2: the new pH is 4. (This works because HCl is a strong acid, fully ionised before and after.)

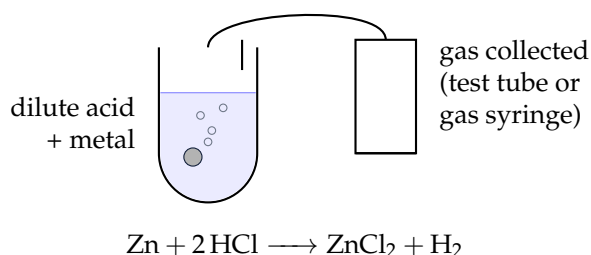
Common mistakes

- Using "strong" and "concentrated" as synonyms. Strength is about ionisation; concentration is about amount per volume.
- Thinking pH 2 is twice as acidic as pH 4. It has a hundred times more H^+ .
- Saying a weak acid will not neutralise as much alkali as a strong acid of the same concentration and volume. The amount of alkali needed depends on the number of moles of acid.

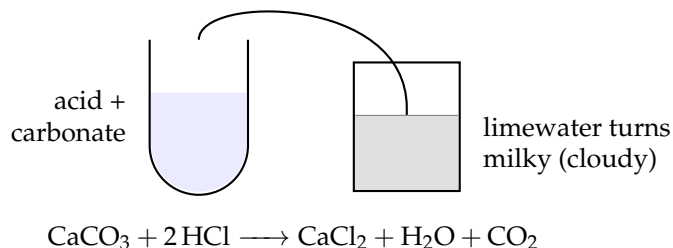
4 Reactions of acids

Acids give three families of reaction, each making a **salt**. A salt is a compound in which the hydrogen of an acid has been replaced by a metal (or ammonium) ion. The name has two parts: the metal, then the acid's stem (chloride, sulfate, nitrate).

(a) Acid + metal $\rightarrow$ salt + hydrogen. Only metals above hydrogen in the reactivity series react: $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$, $\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$. The metal atoms lose electrons to become ions and the H^+ ions gain electrons to become H_2 .



(b) **Acid + metal carbonate** → **salt + water + carbon dioxide**. $\text{CaCO}_3 + 2\text{HCl} \longrightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$. The carbon dioxide causes fizzing (effervescence) and turns limewater milky.



(c) **Acid + base (metal oxide or hydroxide)** → **salt + water**. For example $\text{CuO} + \text{H}_2\text{SO}_4 \longrightarrow \text{CuSO}_4 + \text{H}_2\text{O}$ and $\text{NaOH} + \text{HCl} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$. This is **neutralisation**. The acid and base cancel each other.

Balancing and naming. Work out the formula of the salt from the ion charges (as in the bonding lesson), then balance. With nitric acid and calcium hydroxide: calcium is Ca^{2+} and nitrate is NO_3^- , so the salt is $\text{Ca}(\text{NO}_3)_2$:



Ammonia is a base without a metal: $\text{NH}_3 + \text{HCl} \longrightarrow \text{NH}_4\text{Cl}$ makes ammonium chloride only (no water), because the base simply accepts H^+ .

Worked example: predict the products

Write the equation for magnesium carbonate and dilute sulfuric acid.

Carbonate plus acid gives a salt, water and carbon dioxide. The salt is magnesium sulfate (Mg^{2+} , SO_4^{2-} , formula MgSO_4): $\text{MgCO}_3 + \text{H}_2\text{SO}_4 \longrightarrow \text{MgSO}_4 + \text{H}_2\text{O} + \text{CO}_2$. Balanced: Mg 1, C 1, O: $3 + 4 = 7$ on the left and $4 + 1 + 2 = 7$ on the right; H 2 on both sides.

Worked example: a volume of gas

What volume of carbon dioxide at room temperature and pressure is produced when 5 g of calcium carbonate reacts with excess hydrochloric acid? (A_r : Ca = 40, C = 12, O = 16; molar gas volume 24 dm^3 .)

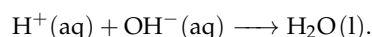
$M_r(\text{CaCO}_3) = 40 + 12 + 48 = 100$. Moles = $5/100 = 0.05$. The equation shows 1 : 1, so 0.05 mol of CO_2 . Volume = $0.05 \times 24 = 1.2\text{ dm}^3$.

Common mistakes

- Forgetting that carbonates also give water. The three products are salt, water and carbon dioxide.
- Predicting that copper reacts with dilute acid to give hydrogen. Copper is below hydrogen.
- Writing the sulfate of sodium as NaSO_4 . It is Na_2SO_4 (SO_4^{2-} needs two Na^+).

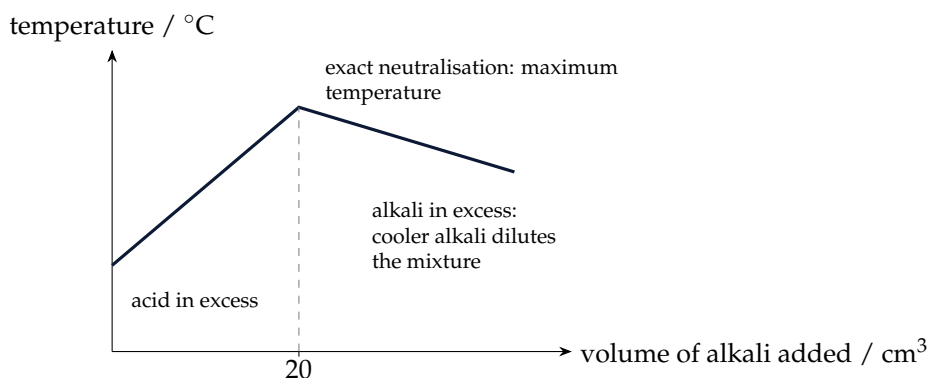
5 Neutralisation, ionic equations and energy

The ionic equation. In a neutralisation of a strong acid by an alkali, the spectator ions are unchanged, and the only reaction is



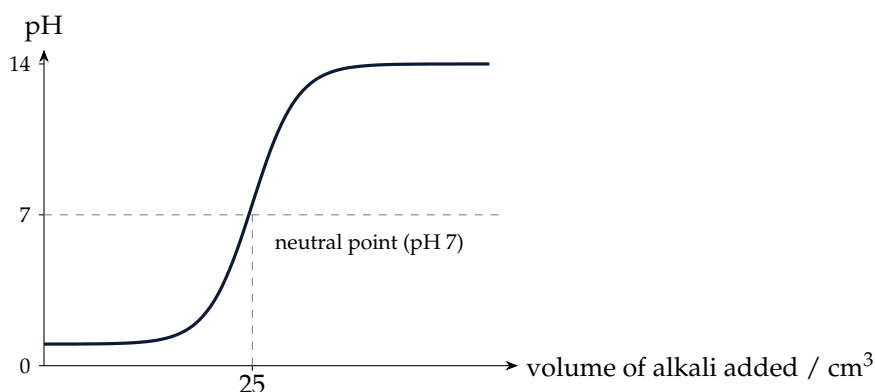
For example, in $\text{HCl} + \text{NaOH} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$, the Na^+ and Cl^- ions are in solution both before and after (spectator ions). The ionic equation shows what actually happens: hydrogen ions and hydroxide ions combine to make water. This is why *all* strong acid–strong alkali neutralisations are alike.

Neutralisation is exothermic. Forming the new bonds in water releases energy to the surroundings, so the temperature rises. We can follow it by adding alkali to acid in a polystyrene cup and recording the temperature after each addition.



The temperature rises until the acid is exactly neutralised and then falls, because additional (cooler) alkali only dilutes and cools the mixture. The peak shows the volume needed for neutralisation.

Following the pH. If we add alkali to acid and measure the pH, the pH rises slowly, then jumps from about 3 to about 11 over a very small volume of alkali (the neutralisation point, pH 7 for a strong acid and strong alkali), then levels off. This sharp change is why indicators can show the end point in a titration to the nearest drop.



Calculating from the equation. A neutralisation calculation uses moles. Moles of solute = concentration $\times$ volume in dm^3 , and then the ratio in the balanced equation links the two solutions.

Worked example: how much alkali?

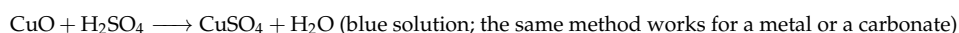
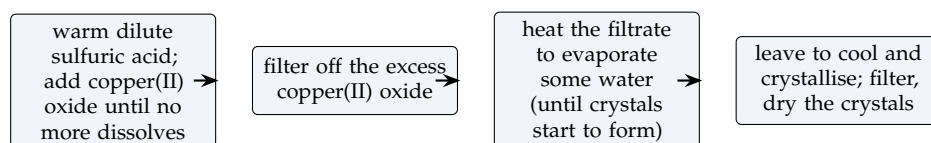
25 cm^3 of sulfuric acid, 0.10 mol/dm^3 , is neutralised by potassium hydroxide solution of concentration 0.20 mol/dm^3 . What volume of alkali is needed? $\text{H}_2\text{SO}_4 + 2\text{KOH} \longrightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$.
Moles acid = $0.10 \times 0.025 = 0.0025$. The ratio is 1 : 2, so moles KOH = 0.0050 . Volume = $0.0050/0.20 = 0.025 \text{ dm}^3 = 25 \text{ cm}^3$.

Common mistakes

- Writing the ionic equation with the spectator ions included.
- Thinking neutralisation is endothermic because the temperature falls after the peak. The fall is due to extra unreacted solution cooling the mixture.
- Forgetting the 1 : 2 ratio when the acid is diprotic.

6 Making a soluble salt

To make a pure, dry sample of a soluble salt such as copper(II) sulfate we use an acid and an insoluble reactant (a metal, an insoluble base or an insoluble carbonate), because the unreacted excess can be filtered off.



Why this order? (1) Warm the acid to speed up the reaction, and add the solid in excess: that ensures all the acid has reacted, so the final solution contains only the salt. (2) Filter to remove the excess solid; the *filtrate* is the salt solution. (3) Heat gently to evaporate some of the water until the solution is saturated (when a drop on a cold glass rod forms crystals); do not evaporate to dryness, because that would leave poor crystals and might decompose a hydrated salt. (4) Cool slowly to give crystals, filter them off, wash with a little cold distilled water and dry between filter papers.

If both reactants are soluble, for example making sodium chloride from sodium hydroxide and hydrochloric acid, the excess cannot be filtered off. You find the exact volumes by titration with an indicator, then repeat without the indicator and evaporate the solution.

Worked example: choosing reagents

Name a suitable pair of reagents to make zinc sulfate crystals, and say why the zinc compound should be in excess. Zinc oxide (or zinc carbonate, or zinc metal) with dilute sulfuric acid: $\text{ZnO} + \text{H}_2\text{SO}_4 \longrightarrow \text{ZnSO}_4 + \text{H}_2\text{O}$. The solid is added in excess so all the acid is used up; the leftover solid is insoluble and can be filtered off, leaving only zinc sulfate solution.

Exam tip

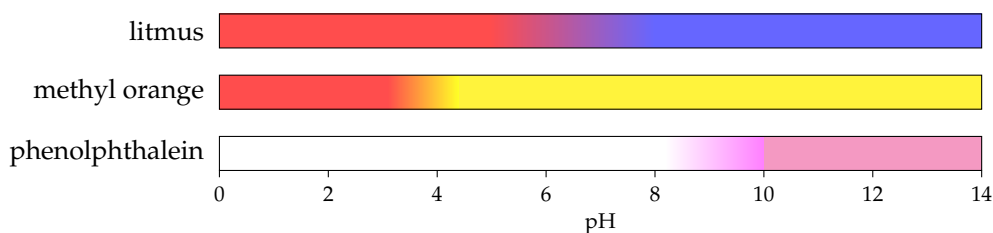
For questions on making a salt, the three keywords are *excess*, *filter* and *crystallise*. If the question asks which reactant should be in excess, it is always the insoluble one.

Common mistakes

- Using sodium or potassium metal (too dangerous) or a soluble base such as NaOH when a filter step is needed.
- Saying the mixture is filtered to remove the product. The salt stays dissolved in the filtrate.
- Evaporating all the water. Stop at the crystallisation point.

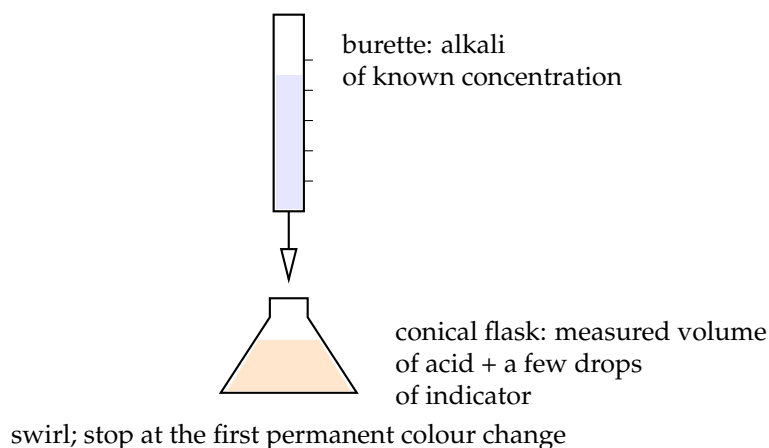
7 Indicators and titration

Indicators. An **indicator** is a dye whose colour depends on the pH. Different indicators change colour over different pH ranges.



Litmus is red in acid and blue in alkali (changing around pH 7); methyl orange is red in acid and yellow in alkali (changing around pH 3–4); phenolphthalein is colourless in acid and pink in alkali (changing around pH 8–10). For a strong acid with a strong alkali, any of these shows the end point, because the pH jumps so far on the last drop. Universal indicator mixes several dyes so it shows a whole range of colours, but it is not used to find an end point because the colour change is gradual.

Titration. A **titration** measures the exact volume of one solution needed to react with a measured volume of another. A pipette measures 25.0 mL of acid into a conical flask, a few drops of indicator are added, and alkali is run in from a burette until the indicator just changes colour (the *end point*).



The first run is rough (a large volume per step), then you repeat, adding drop by drop near the end point, until two or more results agree within 0.1 cm^3 (*concordant* results), and take their mean. Use the mean in the calculation.

Worked example: a titration calculation

A student finds that 25.0 cm^3 of sodium hydroxide of unknown concentration is neutralised by 20.0 cm^3 of 0.10 mol/dm^3 hydrochloric acid. Find the concentration of the sodium hydroxide.

Moles $\text{HCl} = 0.10 \times 0.0200 = 0.0020$. The ratio $\text{HCl} + \text{NaOH} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$ is $1 : 1$, so there are 0.0020 mol of NaOH in 0.0250 dm^3 . Concentration $= 0.0020 / 0.0250 = 0.080 \text{ mol/dm}^3$.

Worked example: choosing the mean

Titres: $23.4, 22.4, 22.3, 22.4 \text{ cm}^3$. What mean should be used?

The first (rough) titre 23.4 is not concordant, so it is discarded. The others agree within 0.1 : mean $= (22.4 + 22.3 + 22.4) / 3 = 22.37 \approx 22.4 \text{ cm}^3$.

Common mistakes

- Averaging in the rough titre. Only concordant results are averaged.
- Using a measuring cylinder for the exact volume: a pipette (or burette) gives accurate volumes.
- Mixing up end point and neutral point for weak acids: the indicator must change at the pH where the reaction is complete.

8 Putting it together: reading acid questions

Almost every ESAT acid question reduces to a short list of steps, which you can follow in order:

1. **Identify the type of reaction.** Acid with a metal, a carbonate, a base (oxide or hydroxide) or an alkali? That fixes the products: salt + hydrogen; salt + water + carbon dioxide; salt + water.
2. **Write the formula of the salt** from the ion charges, then balance the equation.
3. **Mole ratio.** Convert the data to moles (mass / M_r or concentration $\times$ volume), then use the ratio in the equation. For gases at room temperature and pressure, one mole occupies 24 dm^3 .
4. **Strong or weak? Concentrated or dilute?** Decide which of the two ideas the question is testing. Strength changes the H^+ concentration and pH *at the same concentration*; it does not change the amount of alkali needed.
5. **pH changes.** Dilution by 10 raises the pH of a strong acid by 1; dilution by 100 raises it by 2.

Worked example: a full problem

1.2 g of magnesium (relative atomic mass 24) reacts with excess dilute hydrochloric acid. Find (a) the moles of magnesium, (b) the volume of hydrogen at room temperature and pressure, (c) the moles of HCl used.

(a) $1.2 / 24 = 0.050 \text{ mol}$. (b) $\text{Mg} + 2 \text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$: 1 mol Mg gives 1 mol H_2 , so $0.050 \times 24 = 1.2 \text{ dm}^3$. (c) 2 mol HCl per mol Mg : 0.10 mol .

Worked example: limiting reagent

0.20 mol of CaCO_3 is added to 0.30 mol of HCl . Which is the limiting reagent and how many moles of CO_2 form?
 $\text{CaCO}_3 + 2\text{HCl} \longrightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$
 0.20 mol CaCO_3 would need 0.40 mol HCl but only 0.30 is present, so HCl is limiting. CO_2 made = $0.30/2 = 0.15$ mol.

Exam tip

If two acids are compared, write down which of strength, concentration and volume is the same in the question; whatever is the same decides which things are the same in the answer (for example, same moles of monoprotic acid means same amount of alkali needed).

Common mistakes

- Using the wrong mole ratio. Always read it from the balanced equation, not from the number of hydrogens in the acid.
- Forgetting to convert cm^3 to dm^3 when using concentration $\times$ volume.
- Calling the product of an acid and a hydroxide just "salt". Water is made too.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. Write the equation for the dissociation of nitric acid in water.
2. Is CuO an alkali? Explain.
3. By what factor does the H^+ concentration change from pH 5 to pH 2?
4. Explain the difference between a strong acid and a concentrated acid.
5. Name the products of zinc carbonate and nitric acid.
6. Write the ionic equation for the neutralisation of hydrochloric acid by sodium hydroxide.
7. Why is excess insoluble solid used when making a salt?
8. How many moles of H^+ can be released by one mole of phosphoric acid, H_3PO_4 ?

Answers to the self-check

1. $\text{HNO}_3 \longrightarrow \text{H}^+ + \text{NO}_3^-$.
2. No: it is a base (it neutralises acids) but is insoluble, so it does not produce OH^- in solution.
3. It increases by a factor of $10^3 = 1000$.
4. Strong: fully ionised. Concentrated: lots of acid per unit volume.
5. Zinc nitrate, water and carbon dioxide.
6. $\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$.
7. So all the acid reacts; the excess is filtered off.
8. Three (triprotic).

Next step

Heurisco's *ESAT Chemistry: Acids, Bases and Salts Practice Set* has twenty original exam-style questions on this lesson, with particle diagrams, pH curves and temperature graphs, in order of difficulty, with a full worked solution for each.