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Topic 4.6: Introduction to Titration

Revision Notes · AP Chemistry

UNIT 4 · CHEMICAL REACTIONS

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What a Titration Measures

A titration finds the concentration of a solution by reacting it with a solution of known concentration until the reaction is exactly complete.

The solution of unknown concentration is the analyte, a measured volume of it sitting in a conical flask. The solution of known concentration is the titrant, added slowly from a burette until the analyte has just finished reacting.

The volume delivered, the titre, is the measurement: with the titrant's concentration it fixes the moles of analyte in the flask.

The Apparatus: Who Does What

The burette delivers a variable volume and is read to the nearest 0.05 mL. It is rinsed with the titrant first, so water drops inside cannot dilute the standard.

The pipette transfers one exact volume of analyte. It is rinsed with the analyte: anything else inside would contaminate the sample.

The conical flask holds the analyte, and its wide base allows swirling without splashing. Swirling mixes each addition through the flask, so no pocket of unreacted titrant triggers the indicator early.

Only two or three drops of indicator are added: an indicator is itself a weak acid or base, so more would consume measurable titrant. Washing the flask's sides with distilled water is harmless, because it moves no moles of analyte.

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Equivalence Point and Endpoint: Two Different Points

The equivalence point is a point in the reaction: the moment the reactants are mixed in the exact mole ratio of the balanced equation. For a 1 : 1 reaction that means equal moles; for H_2SO_4 + 2NaOH → Na_2SO_4 + $2H_2O$ it means twice as many moles of base as of acid. Stoichiometric is the definition, not equal.

The endpoint is a point in the experiment: the indicator's colour change, at which the chemist stops the burette. It is the observable stand-in for the equivalence point.

The two nearly coincide when the indicator is well chosen, but never match exactly: a small excess of titrant is needed to flip the colour, so the endpoint sits a fraction of a drop past equivalence.

Phenolphthalein, colourless in acid and pink in base, gives the endpoint as the first permanent pale pink when base is added to acid.

Reading the Burette and Building a Reliable Titre

A titre is a difference, not a reading: titre = final burette reading − initial burette reading. Readings are taken at the bottom of the meniscus, eye level with it, to avoid parallax error.

The first titration is a rough run, adding titrant about a millilitre at a time to locate the endpoint. Its titre is never used. Accurate runs go quickly until a few millilitres short of that endpoint, then drop by drop with swirling, stopping at the first permanent colour change. An overshot run is spoiled: repeated, not rescued.

Titres within 0.10 mL of each other are concordant; the reported titre is their mean, rough run excluded.

The Calculation: From Titre to Concentration

Every titration calculation has the same three steps. Step one: moles of titrant = its concentration × the titre in litres. Step two: the equation's mole ratio converts moles of titrant into moles of analyte. Step three: concentration of analyte = moles ÷ the flask volume in litres.

The shortcut $M_1V_1 = M_2V_2$ is only those steps with the ratio already set to 1 : 1. With any other ratio it must be adjusted by that ratio, or replaced by the three steps.

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What the Titration Curve Shows

A titration curve plots pH against volume of titrant added, replacing the indicator's single colour change with the whole story of the neutralisation. For most of the run the pH changes slowly, each addition being consumed by a large excess of the other reactant. Near the equivalence point the excess left is tiny, so one more drop swings the pH sharply and the curve climbs almost vertically. The equivalence volume is read at the midpoint of that steep section.

For a strong acid titrated with a strong base, the pH at equivalence is exactly 7.0, the salt produced being neutral. That is a special case, not a rule: other combinations reach equivalence at a different pH, and their curves belong to Unit 8.

Back Titration: Titrating What Is Left

Some analytes cannot be titrated directly: an insoluble solid such as chalk reacts too slowly for a sharp endpoint. Instead, it is reacted with a known excess of reagent and the leftover is titrated.

A measured excess of standard acid is added; once the solid has reacted, the leftover acid is titrated against a standard base. Moles reacted = moles added − moles left over, and the equation's ratio converts that into moles of analyte.

Worked Examples

Example 1: A 1 : 1 calculation from burette readings. A 25.0 mL sample of HCl is titrated with 0.100 M NaOH, with burette readings of 0.40 mL and 24.90 mL. Titre = 24.90 − 0.40 = 24.50 mL. Moles of NaOH = 0.100 × 0.02450 = 0.00245 mol. The equation HCl + NaOH → NaCl + H_2O is 1 : 1, so the flask held 0.00245 mol of HCl, and its concentration is 0.00245 ÷ 0.0250 = 0.0980 M.

Example 2: A 1 : 2 ratio. 20.0 mL of H_2SO_4 is neutralised by 32.40 mL of 0.200 M NaOH. Moles of NaOH = 0.200 × 0.03240 = 0.00648 mol. The equation H_2SO_4 + 2NaOH → Na_2SO_4 + 2H_2O consumes 2 mol of NaOH per mole of acid, so moles of acid = 0.00648 ÷ 2 = 0.00324 mol, and concentration = 0.00324 ÷ 0.0200 = 0.162 M. Skipping the ratio doubles the answer.

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Example 3: A back titration. A 0.500 g antacid tablet containing $CaCO_3$ is dissolved in 50.0 mL of 0.200 M HCl, an excess. The leftover acid needs 28.60 mL of 0.150 M NaOH. Acid added = 0.200 × 0.0500 = 0.0100 mol; NaOH used = 0.150 × 0.02860 = 0.00429 mol, so 0.00429 mol of acid was left over and 0.0100 − 0.00429 = 0.00571 mol reacted. From $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$, moles of $CaCO_3$ = 0.00571 ÷ 2 = 0.00286 mol, mass 0.00286 × 100.09 = 0.286 g. The tablet is 0.286 ÷ 0.500 × 100 = 57.2% $CaCO_3$.

Traps That Cost Marks in Topic 4.6

Treating the endpoint and the equivalence point as the same thing. The equivalence point is stoichiometric; the endpoint is the indicator's colour change. The endpoint measures the equivalence point, small overshoot included; it is not a synonym for it.

Applying $M_1V_1 = M_2V_2$ to a reaction that is not 1 : 1. When 30.0 mL of 0.200 M H_2SO_4 neutralises 25.0 mL of NaOH, the shortcut gives 0.240 M. The equation needs 2 mol of NaOH per mole of acid: the true value is 0.480 M.

Feeding millilitres into concentration × volume. Moles = concentration × volume needs litres. A titre of 24.50 mL is 0.02450 L; using 24.50 inflates the moles a thousandfold.

Rinsing the burette with water instead of titrant. Water inside dilutes the standard, so a larger titre is delivered and the moles are overcounted. The analyte's concentration comes out too high.

Averaging every titre, rough run included. The rough titre is only a scouting value. Only concordant accurate titres are averaged.

Assuming the pH at equivalence is always 7.0. That holds only for a strong acid with a strong base. Equivalence means the equation's ratio has been reached; the pH depends on what the reaction leaves behind.

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Key Relationships to Remember

- Titrant: known concentration, delivered from the burette. Analyte: unknown concentration, measured volume in the flask
- Equivalence point: reactants mixed in the exact mole ratio of the balanced equation. Endpoint: the indicator's colour change, the observed signal that stands in for it
- Titre = final burette reading − initial burette reading; reliable work averages concordant titres within 0.10 mL, excluding the rough run
- Moles of titrant = concentration × titre in litres; the equation's ratio converts to moles of analyte; concentration = moles ÷ volume in litres
- $M_1V_1 = M_2V_2$ works as it stands only for a 1 : 1 ratio; any other ratio must be built into the calculation
- On a titration curve, the equivalence volume is the midpoint of the steep section, and the pH there is 7.0 only for a strong acid with a strong base
- In a back titration, moles reacted = moles of reagent added − moles left over

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