AfL 10: Acids, Bases and Indicators:

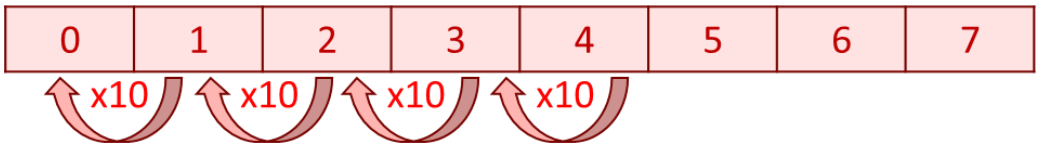
A. Ions and the effect of indicators on acids and bases:

Substance:	pH range:	Ions:	Colour in phenolphthalein:	Colour in Methyl orange:	Colour in Litmus:
Acid	1-6	H^+	Colourless	Red	Red
Alkali/Base	8-12	OH^-	Pink	Yellow	Blue
Neutral	7	None	Colourless	Orange	Purple

- An alkali is any base that is soluble. All alkalis are bases, but **not** all bases are alkalis.

B. Ions and pH:

- The concentration of an acid is linked to the pH of the substance. Nice and simply, the more H^+ ions there are, the lower the pH.
- Every time the pH decreases by one, the concentration of H^+ ions **increases** by 10 times.
- For example:
  - Calculate how much more concentrated hydrochloric acid is, with a pH of 0, than ethanoic acid, with a pH of 4.

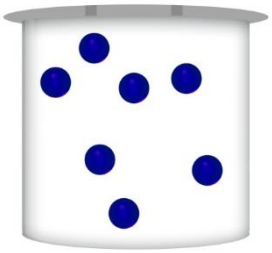


  - As you can see, the pH has gone down by 4, so the it is 10 x 10 x 10 x 10 times more concentrated = 10,000 times more concentrated.

C. Dilute, Concentrated, Strong and Weak Acids:

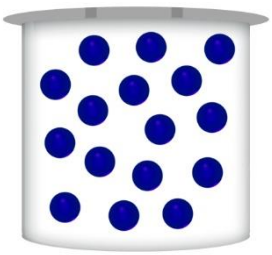
- The concentration of an acid shows you the amount of hydrogen ions there are in a litre (1 dm^3) of water.
- The strength of an acid tells you how easily they ionise – split up into H^+ ions.

Dilute



The more dilute an acid is, the less H^+ ions there are per dm^3

Concentrated



The more concentrated an acid is, the more H^+ ions there are per dm^3

Strong Acids

HCl – Hydrochloric acid
H_2SO_4 - Sulphuric acid
HNO_3 – Nitric acid

Strong acids dissociate / ionise completely – producing more H^+ ions. They have **lower** pH's (0-2)

Weak Acids

CH_3COOH - Ethanoic acid
$C_6H_8O_7$ - Citric acid
H_2CO_3 - Carbonic acid

Weak acids do not fully ionise – only a small amount dissociate – fewer H^+ ions. pH = 2-6

D. What happens to the ions during neutralisation?

When any neutralisation reaction occurs, the same basic reaction occurs. Hydrogen ions react with hydroxide ions to form water:

$$H^+ + OH^- \rightarrow H_2O$$

E. Testing for Gases:

Hydrogen: Hydrogen is a flammable gas. If you take a **lit splint** and add it to a test tube containing hydrogen, you will hear a **squeaky pop**.

Carbon Dioxide: Carbon dioxide turns limewater cloudy. **Bubble** the gas through **limewater** and if it goes **cloudy/milky**, carbon dioxide is present.

F. Naming Salts:

When an acid reacts with a base, a salt is formed. You need to know how to name the salts.

Step 1: Naming the Salt

a. If you have hydrochloric acid, HCl, you get a chloride salt
b. If you have nitric acid, HNO_3, you get a nitrate salt
c. If you have sulphuric acid, H_2SO_4, you get a sulphate salt.

For example, if you react lithium hydroxide with nitric acid, you take the name of the metal – lithium – and add the salt ending, which in this case is a nitrate:

lithium hydroxide + nitric acid → lithium nitrate

Step 2: The by-products:

You also get by-products in the reaction, depending on what you are reacting the acid with:

a. A metal on its own will produce hydrogen gas, H_2
b. A metal oxide/hydroxide will produce water, H_2O, and
c. A carbonate will produce water **and** carbon dioxide, $H_2O + CO_2$

Step 3: Putting it all together:

Question: Copper carbonate, $CuCO_3$ reacts with sulphuric acid. Write the word equation. (2)

i. Work out the salt. From my reactants, the metal is copper, and the acid gives me a sulphate – therefore my salt is copper sulphate
ii. Work out the by-products. I have a carbonate; therefore, I have water and carbon dioxide.

Copper carbonate + sulphuric acid → Copper sulphate + water + carbon dioxide

G. What is seen when acids and bases react together?

Look at the chemical equation below. What would you be able to **see** during this reaction?

$$CuCO_{3\,(s)} + HNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$$

The state symbols show that a solid has disappeared and that a gas has formed. Therefore, we will see:

1. The solid has disappeared/dissolved (1)
2. There will be bubbling/fizzing/effervescence proving there is a gas (1)

Don't say a gas is formed – you can't **see** that – but you can see bubbles, which **prove** there is a gas.

AfL 11: Preparing Salts:

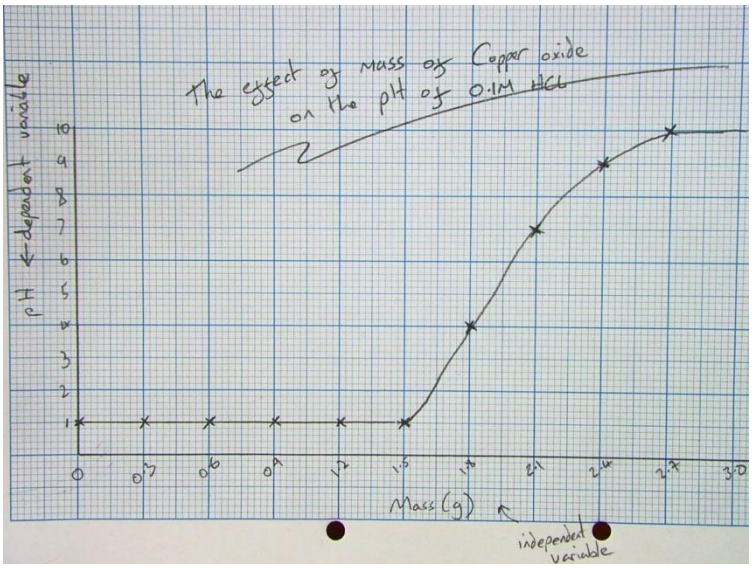
A. Core Practical: Investigating the change in pH when calcium hydroxide is added to an acid:

Calcium hydroxide	+	Hydrochloric Acid	→	Calcium chloride	+	Water
$Ca(OH)_2$	+	2HCl	→	$CaCl_2$	+	$2H_2O$

This is one of the big practicals that you need to know for the exam. You could be asked how to carry it out; to analyse the results; to explain the results or to evaluate the risks.

To investigate the pH of a substance, you will need to be able to do the following:

1. Measure out 50cm^3 of hydrochloric acid into a beaker.
2. Record the pH of the solution by putting some Universal Indicator paper onto a white tile and place a drop of the acid onto it.
3. Leave it for 30 seconds to make sure the colour change is complete and then record the pH.
4. Measure out 0.3g of calcium hydroxide and add it to the beaker.
5. Record the pH of the substance and repeat until 2.4g of the solid is added.
6. Plot a graph as seen on the right.



As you can see from the experiment:

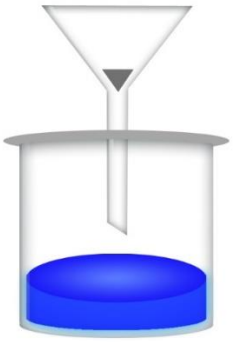
- The pH stays at 1 until 1.5g of Calcium Hydroxide is added.
- After 1.5g, the pH starts to increase rapidly and continues until about 2.7g of the solid is added, at which point it stays at pH 10.
- The reaction is **neutral** at 2.1g.

Risks and management:	Improvements:	Variables:
• Calcium oxide = corrosive/ Calcium hydroxide = irritant – weak goggles and if you get it on your hands, wash it off.	• Use a pH meter instead of Universal Indicator paper. • It is more accurate (1 dp.) than UI paper.	• Independent: Mass of $Ca(OH)_2$ • Dependent: pH of substance • Control: Volume of acid, concentration of acid, type of tablet, etc.

B. Core Practical: How to prepare soluble salts from an acid and an insoluble reactant:

Example: How can you produce soluble copper sulphate from sulphuric acid and insoluble copper oxide?

1. **Heat the acid in a water bath** to speed up the reaction. (In a fume cupboard to reduce acidic fumes)
2. **Add the copper oxide to the acid**. A reaction will occur, producing copper sulphate – your soluble salt
3. **Continue adding copper oxide until no more reacts**. The acid is now neutral, and you have your salt, water and excess copper oxide.
4. **Use a filter paper to remove the excess copper oxide**, leaving you with the copper sulphate and water
5. **Heat the solution gently**. To evaporate off some of the water.
6. **Leave the solution to cool and crystallise the salt** – then leave to dry.



C: How to prepare soluble salts from an acid and a soluble reactant:

Example: How can you produce soluble sodium chloride from hydrochloric acid and soluble sodium hydroxide?

1. **Measure out a certain amount of alkali (sodium hydroxide) into a conical flask.**
2. **Add an indicator – such as phenolphthalein**. Phenolphthalein will be pink in an alkali/base.
3. **Add the hydrochloric acid slowly until the end point**. For phenolphthalein, the end point will be when it turns colourless.
4. **Repeat the experiment without the indicator**, this gives you just sodium chloride and water.
5. **Heat the solution gently**. To evaporate off some of the water.
6. **Leave the solution to cool and crystallise the salt** – then leave to dry.

D. How to predict insoluble salts:

When it comes to preparing insoluble salts from soluble salts, first you need to know the solubility rules:

Salts:	Soluble:	Insoluble:
Sodium, potassium, ammonium	All soluble	-
Nitrates	All soluble	-
Chlorides	Mostly soluble	Silver chloride and lead chloride
Sulphates	Mostly soluble	Lead, barium and calcium sulphate
Carbonates/hydroxides	Sodium, potassium and ammonium	Mostly insoluble

E. How to prepare insoluble salts:

If you wanted to prepare a pure dry precipitate of lead chloride from the above reaction, there are 4 steps you need to follow:

1. **Dissolve** the solid copper chloride and lead nitrate
2. **Mix** them together to produce your copper nitrate and lead chloride
3. **Filter** the solution to give you lead chloride in the filter paper
4. **Wash** the filter paper to remove the soluble copper nitrate
5. **Dry** the filter paper in a desiccator of an oven to leave you with your pure, dry precipitate.

Apply Task 3: Acids, Bases & Indicators

Question:		Marks:
1:	An acid can be reacted with a base in a titration experiment. i. The general equation for the reaction of an acid with a base is (1) ☐ **A** acid + base → alkali + water ☐ **B** acid + base → salt + carbon dioxide ☐ **C** acid + base → salt + water + hydrogen ☐ **D** acid + base → salt + water ii. Write the ionic equation to show what happens to the ions during neutralisation (1) iii. Describe the test for hydrogen (2) iv. Describe the test for carbon dioxide (2)	**6**
2:	A titration is used to determine the exact volumes of sulfuric acid and sodium hydroxide solution that neutralise each other. i. State the name of an indicator that could be used, and the colour change seen in this titration. (3) • Indicator: _____ • Colour in sodium hydroxide solution: _____ • Colour when neutral: _____ ii. As the pH of the solution of sulfuric acid and sodium hydroxide changes from pH 3 to pH 5, the concentration of H^+ ions changes. Calculate how much **less** concentrated the solution is at pH 5. (2)	**5**
3:	Acids can be strong, dilute, concentrated or weak. i. Describe the difference between a strong and a weak acid (2) ii. Describe the difference between a dilute and concentrated acid (2) iii. Explain how a weak, dilute acid can be altered to increase the speed of a reaction. (2)	**6**
4:	When sodium carbonate, Na_2CO_3, is added to nitric acid, HNO_3, a neutralisation reaction occurs and sodium nitrate, $Na_2(NO_3)_2$, is produced. i. Write the word equation for the reaction. (2) ii. Describe what you would **see** when solid sodium carbonate is added to dilute nitric acid. (1) iii. Write the balanced equation for the reaction. (3)	**6**

Red:	Amber:	Green:	Blue:
1-6	7-12	13-18	19-23

Apply Task 4: Preparing Salts

Question:		Marks:
1:	When calcium hydroxide is added to hydrochloric acid, calcium chloride and water are formed. i. Describe an experiment to investigate how changing the mass of copper hydroxide added to hydrochloric acid affects the pH of the solution. (4) ii. Describe a hazard from the experiment and suggest how to minimise the risk (2)	6
2:	Copper sulfate is a soluble salt. Copper sulfate can be prepared by reacting copper oxide with dilute sulfuric acid. Copper oxide is an insoluble solid. Describe how you would prepare some pure, dry crystals of copper sulfate by reacting excess copper hydroxide with dilute sulfuric acid.	6
3:	Sodium nitrate is a soluble salt. Sodium nitrate can be prepared by reacting soluble sodium hydroxide with nitric acid. Describe how you would prepare some pure, dry crystals of sodium nitrate by reacting sodium hydroxide with nitric acid.	6
4:	When sodium hydroxide and copper nitrate react together, a precipitate forms. i. Write the word equation for the reaction. (1) ii. Name the precipitate formed. (1) iii. Explain how to produce a pure, dry precipitate from soluble copper nitrate and sodium hydroxide. (4)	6

Red:	Amber:	Green:	Blue:
1-6	7-12	13-18	19-24

Check Task 3: Acids, Bases & Indicators

Question:		
1:	An acid can be reacted with a base in a titration experiment. i. The general equation for the reaction of an acid with a base is (1) ☒ **D acid + base → salt + water** ii. Write the ionic equation to show what happens to the ions during neutralisation (1) ❄ $H^+ + OH^- \rightarrow H_2O$ iii. Describe the test for hydrogen (2) ❄ Lit splint (1)…squeaky pop (1) iv. Describe the test for carbon dioxide (2) ❄ (Bubble through) limewater (1) which goes cloudy (1)	**6**
2:	A titration is used to determine the exact volumes of sulfuric acid and sodium hydroxide solution that neutralise each other. i. State the name of an indicator that could be used, and the colour change seen in this titration. (3) • Indicator: Methyl orange / phenolphthalein / litmus (1) • Colour in sodium hydroxide solution: Yellow / pink / blue (1) • Colour when neutral: Orange / colourless / purple (1) ii. As the pH of the solution of sulfuric acid and sodium hydroxide changes from pH 3 to pH 5, the concentration of H^+ ions changes. Calculate how much **less** concentrated the solution is at pH 5. (2) ❄ pH change: 2 = 10 x 10 (1) ❄ = 100 times less concentrated (1) ❄ 100 times gets both marks (1)	**5**
3:	Acids can be strong, dilute, concentrated or weak. i. Describe the difference between a strong and a weak acid (2) ❄ Strong acids **ionise fully** into H^+ ions (1) ❄ Weak acids only **ionise partially (1)** ii. Describe the difference between a dilute and concentrated acid (2) ❄ Concentrated = lots of H^+ ions ❄ Dilute = less H^+ ions iii. Explain how a weak, dilute acid can be altered to increase the speed of a reaction. (2) ❄ Make it more concentrated (1) ❄ Evaporate some of the water (1) ❄ Add more H^+ ions in the same volume (1)	**6**
4:	When sodium carbonate, Na_2CO_3, is added to nitric acid, HNO_3, a neutralisation reaction occurs and sodium nitrate, $NaNO_3$, is produced. i. Write the word equation for the reaction. (2) ❄ Sodium carbonate + nitric acid → sodium nitrate (1) + carbon dioxide + water (1) ii. Describe what you would **see** when solid sodium carbonate is added to dilute nitric acid. (1). Any one from: ❄ Bubbles (1) ❄ Solid {dissolving / disappearing} (1) iii. Write the balanced equation for the reaction. (3) ❄ $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + CO_2 + H_2O$ ❄ LHS (1); RHS (1); Balancing (1)	**6**

Check Task 4: Preparing Salts

Question:		
1:	When calcium hydroxide is added to hydrochloric acid, calcium chloride and water are formed. i. Describe an experiment to investigate how changing the mass of copper hydroxide added to hydrochloric acid affects the pH of the solution. (4) ❄ Measure out hydrochloric acid (1) ❄ Measure the pH (with pH paper) (1) ❄ Add {0.3g/specific mass} of copper hydroxide (1) ❄ Remeasure the pH (1) ❄ Keep adding more solid and measuring the pH (1) ii. Describe a hazard from the experiment and suggest how to minimise the risk (2) ❄ Hazard: Acid is an irritant (1) ❄ Risk: {If on hands wash off / wear goggles to protect eyes / wear gloves} (1)	6
2:	Copper sulfate is a soluble salt. Copper sulfate can be prepared by reacting copper oxide with dilute sulfuric acid. Copper oxide is an insoluble solid. Describe how you would prepare some pure, dry crystals of copper sulfate by reacting excess copper hydroxide with dilute sulfuric acid. ❄ Step 1: Add copper hydroxide (1) until the solution is neutral (1) ❄ Step 2: Filter the mixture (1) ❄ Step 3: Heat the solution / evaporate the water (1) ❄ Step 4: Stop heating when about half of the solution has evaporated (1) ❄ Step 5: Leave the solution to cool (1) ❄ Step 6: Dry the crystals using filter paper (1)	6
3:	Sodium nitrate is a soluble salt. Sodium nitrate can be prepared by reacting soluble sodium hydroxide with nitric acid. Describe how you would prepare some pure, dry crystals of sodium nitrate by reacting sodium hydroxide with nitric acid. ❄ Step 1: Add sodium hydroxide to a conical flask (1) ❄ Step 2: Add phenolphthalein indicator – it will go pink (1) ❄ Step 3: Add hydrochloric acid to a burette (1) ❄ Step 4: Add the acid until the indicator goes colourless (1) ❄ Step 5: Repeat without the indicator (1) so that it is pure ❄ Step 6: Heat the solution / evaporate the water / crystallise (1) ❄ Step 7: Leave the solution to cool (1) ❄ Step 6: Dry the crystals using filter paper (1)	6
4:	When sodium hydroxide and copper nitrate react together, a precipitate forms. i. Write the word equation for the reaction. (1) ❄ Sodium hydroxide + copper nitrate → sodium nitrate + copper hydroxide (1) ii. Name the precipitate formed. (1) ❄ Copper hydroxide (1) iii. Explain how to produce a pure, dry precipitate from soluble copper nitrate and sodium hydroxide. (4) ❄ Dissolve (1) ❄ Mix (1) ❄ Filter (1) ❄ Wash (1) ❄ Dry (1)	6