# Metals

## Metal Extraction

**Unreactive** metals, such as gold, can be found in the earth as **elements**. However, more reactive metals are often found as **compounds** and need to be **extracted** by chemical processes. The rocks that metals are extracted from are called **ores**.

The most common compounds are **oxides**. Metals **less reactive than carbon** can be extracted from oxides by **reduction with carbon**. This involves heating the ore to high temperatures in the presence of carbon.

EXAMPLE!

$$2CuO + C \rightarrow 2Cu + CO_2$$

Most reactive → Least reactive

| Metal | Extraction |
|---|---|
| Potassium, Sodium, Lithium, Calcium, Magnesium | These metals are **too reactive** for carbon extraction and must be extracted via **electrolysis** |
| **Carbon** | |
| Zinc, Iron, Copper, Silver | These metals can be extracted by **reduction with carbon** |
| Gold | Found as an **element** in the earth |

## THE REACTIVITY SERIES

Most reactive → Least reactive

| Metal | Reacts with oxygen (in air) | Reacts with dilute acid | Reacts with water |
|---|---|---|---|
| Potassium | Fast | Explosive | Catches fire |
| Sodium | | | Fizzes violently |
| Lithium | | | |
| Calcium | | Fast bubbling | Fizzes |
| *Carbon goes here* | | | |
| Magnesium | | | No reaction |
| Zinc | | | |
| Iron | Needs to be heated | Slow bubbling | |
| Copper | | | |
| Silver | | | |
| Gold | No reaction | | |

### WHAT DETERMINES REACTIVITY?

- Reactivity is determined by **how easy** it is for the metal **to lose its outer electrons**.
- **Group 1** metals are at the top of the series because they only have **1** outer electron and it is **easy** to lose.
- **Transition metals** are at the bottom because they have **more** outer electrons which are much **harder** to lose.

## COMMON REACTIONS

### REACTIONS WITH OXYGEN

metal + oxygen → **metal oxide**

EXAMPLE! $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

### REACTIONS WITH WATER

metal + water → **metal hydroxide** + hydrogen

$Zn + 2H_2O \rightarrow Zn(OH)_2 + H_2$

EXAMPLE!

| Metal | Observations when added to water |
|---|---|
| Potassium | Fizzes, melts, moves on surface, purple flame |
| Sodium | Fizzes, melts, moves on surface, sometimes yellow flame |
| Lithium | Fizzes, moves slowly on surface |
| Calcium | Fizzes, white solid forms |

### REACTIONS WITH DILUTE ACID

EXAMPLE!

metal + hydrochloric acid → **metal chloride** + hydrogen

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

metal + sulfuric acid → **metal sulfate** + hydrogen

$Ca + H_2SO_4 \rightarrow CaSO_4 + H_2$

metal + nitric acid → **metal nitrate** + hydrogen

$Zn + 2HNO_3 \rightarrow Zn(NO_3)_2 + H_2$

**Exam Tip!** Always make sure your symbol equations are balanced!

## Oxidation and Reduction

**OXIDATION** = gain of oxygen
**REDUCTION** = loss of oxygen

Metals extracted from their ores by burning with carbon get **REDUCED**.

EXAMPLE!

Loses oxygen = **REDUCTION**

$$ZnO + C \rightarrow Zn + CO$$

Gains oxygen = **OXIDATION**

HT ONLY

**Exam Tip!** If you struggle to remember the definitions, try memorising OIL RIG – Oxidation Is Loss, Reduction Is Gain.

Oxidation and reduction can also be described in terms of **electrons**.

- If a species is **oxidised**, it **loses** electrons.
- If a species is **reduced**, it **gains** electrons.

If one species loses electrons they must be picked up by another species. Reactions where both oxidation and reduction occur are called **redox reactions**.

EXAMPLE!

$$Fe + 3AgNO_3 \rightarrow Fe(NO_3)_3 + 3Ag$$

$$Fe(s) + 3Ag^+(aq) \rightarrow Fe^{3+}(aq) + 3Ag(s)$$

Fe loses electrons to become $Fe^{3+}$ = **OXIDATION**
$Ag^+$ gains electrons to become Ag = **REDUCTION**

## Displacement Reactions

A more reactive metal can **displace** a less reactive metal from a compound. The **reactivity series** can be used to determine whether such a reaction will be successful.

EXAMPLE!

$$Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$$

✔ Zinc is **more reactive** than copper. This displacement reaction **will** occur.

$$Zn(s) + MgCl_2(aq) \rightarrow ZnCl_2(aq) + Mg(s)$$

✘ Zinc is **less reactive** than magnesium. This displacement reaction **will not** occur.

# Acids I





## DEFINITIONS

**ACID** = a substance that produces **$H^+$ ions** in an aqueous solution, which has a pH **less than 7**

**BASE** = a substance that has a pH of **more than 7**

**ALKALI** = a substance that produces **$OH^-$ ions** in an aqueous solution, which has a pH **more than** 7 (a soluble base)

**INDICATOR** = a compound that **changes colour** at a fixed pH

## Common Reactions

EXAMPLE!

### REACTIONS WITH HYDROXIDES

**metal hydroxide + acid → metal salt + water**
$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

### REACTIONS WITH OXIDES

**metal oxide + acid → metal salt + water**
$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$

### REACTIONS WITH CARBONATES

**metal carbonate + acid → metal salt + water + carbon dioxide**
$CuCO_3 + 2HNO_3 \rightarrow Cu(NO_3)_2 + H_2O + CO_2$

These are **neutralisation** reactions – when an acid reacts with an alkali or a base, $H^+$ ions react with $OH^-$ ions and form water.

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

**Exam Tip!**
Acid plus metal is NOT a neutralisation reaction – no water is produced.

### REACTIONS WITH METALS

metal + acid → metal salt + hydrogen
$Fe + 2HCl \rightarrow FeCl_2 + H_2$

### PREDICTING PRODUCTS OF REACTIONS

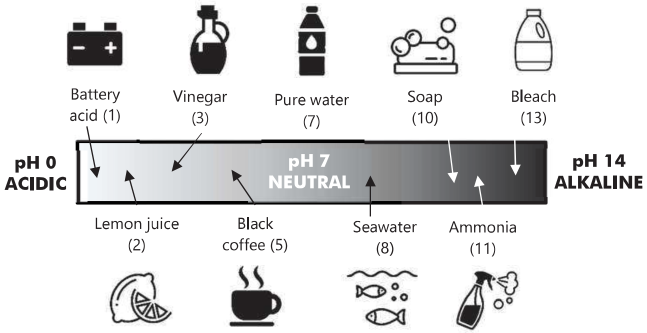
A metal salt is an ionic compound containing a **positive ion** (cation) and a **negative ion** (anion).

- The cation always comes from the **metal**.
- The anion always comes from the **acid**.

Hydrochloric acid → metal **chlorides**
Nitric acid → metal **nitrates**
Sulfuric acid → metal **sulfates**

## pH

- pH is a numerical scale that indicates how **acidic** or **alkaline** a solution is.
- A **neutral** solution is at pH **7**.
- An **acidic** solution has a pH **less than 7**.
- An **alkaline** solution has a pH **more than 7**.
- pH can be measured using an **indicator**, which changes colour at a certain pH value.




Most indicators only have one colour change but **universal indicator** has a spectrum of colour changes, allowing it to estimate the pH of a substance between values of **0 and 14**.

- Acidic solutions are **red/orange**
- Neutral solutions are **green**
- Alkaline solutions are **blue/purple**

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## REDOX REACTIONS WITH ACIDS

The reaction between a metal and an acid is a **redox reaction** – the metal gets **oxidised** and the $H^+$ ions in the acid get **reduced**.

EXAMPLE!

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

$Zn \rightarrow Zn^{2+}$ **OXIDATION**
$2H^+ \rightarrow H_2$ **REDUCTION**

$Cl^-$ ions don't change during the reaction so they aren't shown in the ionic equation – they are called **spectator ions**

**Exam Tip!**
Remember OIL RIG – Oxidation Is Loss, Reduction Is Gain (of electrons).

## Strong Acids and Weak Acids

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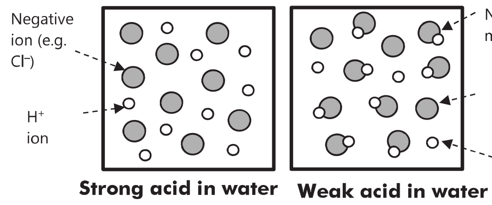
**STRONG ACIDS completely ionise** in water – all the molecules break down into ions.
**WEAK ACIDS partially ionise** – only some molecules break into ions.




If you made solutions of a strong acid and a weak acid of **equal concentration**, the strong acid solution would contain **more $H^+$** ions and would, therefore, have a **lower pH**.

EXAMPLE!

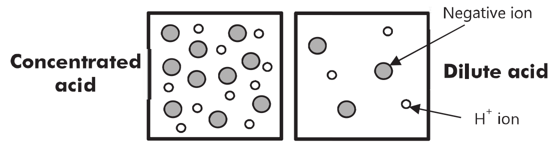
| Strong acid | Weak acid |
|---|---|
| Sulfuric acid | Ethanoic acid (vinegar) |
| Nitric acid | Citric acid (lemon juice) |
| Hydrochloric acid | Carbonic acid (fizzy drinks) |

If you decrease the pH of a solution by **one unit**, the concentration of $H^+$ increases by a **factor of 10**.

- A solution at pH 2 has **10 times** more $H^+$ ions than a solution at pH 3
- A solution at pH 1 has **100 times** more $H^+$ ions than a solution at pH 3

In a **neutral** solution, the concentration of $H^+$ ions is **equal** to the concentration of $OH^-$ ions.

**DILUTE ACIDS** only have a **small number** of molecules dissolved in the solvent.
**CONCENTRATED ACIDS** have a **large number** of molecules dissolved in the solvent.




**Exam Tip!**
Dilute/concentrated refers to the **amount of substance** in a solution, whereas weak/strong refers to the **degree of ionisation**. Don't get the terms mixed up!

# Acids II

## PRACTICAL SKILLS

## PREPARING SOLUBLE SALTS

**Soluble salts** can be made by reacting an **acid** with an **insoluble base**.
Pure dry salts can be isolated from solutions by **crystallisation**.

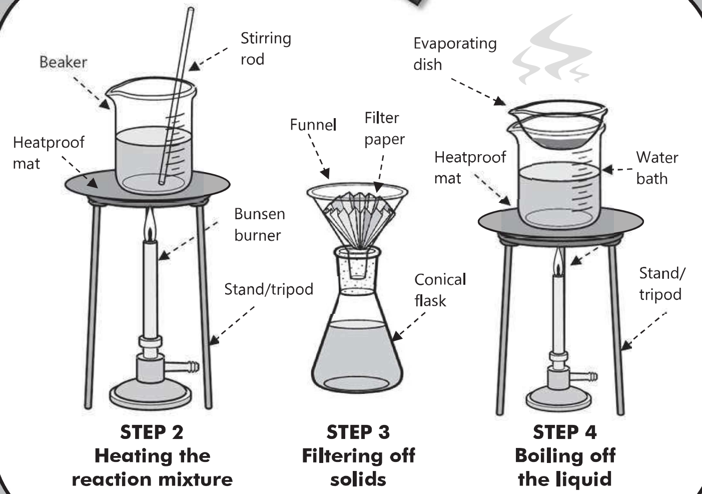
### Method

1. Measure out the **acid** into a beaker.
2. Add the **insoluble base** (e.g. metal oxide or carbonate) and **heat**. Keep adding, while heating and stirring, until **no more dissolves**.
3. Allow to **cool** and **filter** the solution to remove any unreacted solid.
4. **Heat** the solution using a **water bath** to **boil off** some of the liquid.
5. Allow to **cool** and **crystallise** in an evaporating dish.
6. **Filter** the crystals from any remaining solution.
7. Leave to **dry** on a paper towel.

**Safety Tip!**

Wear goggles when carrying out the experiment. Quickly wipe up any spills with a damp paper towel. If you get any acid on your skin or in your eyes, wash thoroughly with lots of cold water.

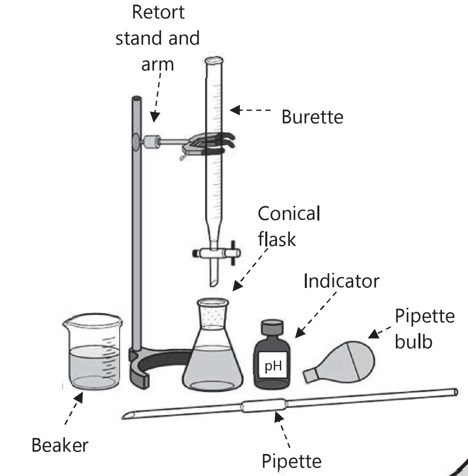
### Equipment




## ACID–BASE TITRATIONS

This example is for an **acid of unknown concentration** reacting with an **alkali of known concentration** but also works the other way round.

### Equipment




### Method

1. Measure a known volume of alkali (known concentration) into a conical flask using a pipette.
2. Add a few drops of a suitable indicator and mix well.
3. Fill up the burette with acid using beaker (unknown concentration). Make sure the section under the tap is also filled.
4. Add acid slowly to the conical flask by opening the tap.
5. When approaching the end point, add the acid dropwise until the indicator changes colour.
6. Record the volume used from the burette.
7. Repeat the experiment until concordant results are achieved (at least two results within 0.1 $cm^3$ of each other).

### Results

Results of a titration can be reported like this:

**EXAMPLE!**

| Repeat | Rough | 1 | 2 | 3 |
|---|---|---|---|---|
| **Initial reading** | **0.00** | **0.40** | **1.30** | **0.20** |
| **Final reading** | **22.10** | **21.90** | **26.10** | **21.80** |
| **Titre volume ($cm^3$)** | 22.10 | 21.50 | 24.80 | 21.60 |

The 'rough' value shouldn't be included in any calculations.

These results are within 0.1 $cm^3$ of each other so they are **concordant**.

This result is significantly different from the others so is called an **anomaly** and is not counted.

The **mean titre** is the average of all the concordant titre values – do not include your rough titre in this.

$$\text{Mean titre} = \frac{21.50 + 21.60}{2} = 21.55\ cm^3$$

**Exam Tip!**
Remember to convert between $cm^3$ and $dm^3$ by dividing by 1000.

### Calculations

HT ONLY

If you know the concentration of one of the solutions used, you can use the **reacting volumes** to calculate the other concentration.

**STEP 1:** calculate the **number of moles** of A (known concentration)
**STEP 2:** use the **balanced equation** to find the **number of moles** of B (unknown concentration)
**STEP 3:** calculate the **concentration** of B in mol/$dm^3$

**EXAMPLE!**

25 $cm^3$ of a 0.5 mol/$dm^3$ solution of KOH reacted with 19.5 $cm^3$ of HCl.
**Find the concentration of HCl.**

$$KOH + HCl \rightarrow KCl + H_2O$$

**STEP 1:** no. moles KOH = volume × concentration = 0.025 × 0.5 = 0.0125 mol
**STEP 2:** ratio of KOH : HCl is 1 : 1 so 0.0125 mol HCl reacted
**STEP 3:** concentration of HCl = no. moles ÷ volume = 0.0125 ÷ 0.0195 = **0.64 mol/$dm^3$**