

## Topic 4 – Chemical changes

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## 4.1 Reactivity of metals

### 4.1.1 Metal oxides

#### Metal oxides

$\text{Metal}_{(s)} + \text{oxygen}_{(g)} \rightarrow \text{metal oxides}_{(s)}$

#### OIL RIG

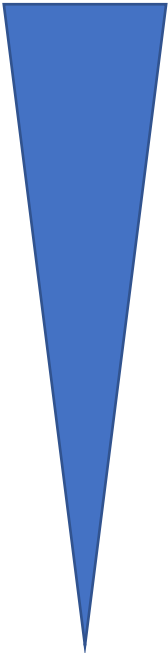
**Oxidation** is loss of electrons, gain of  $\text{O}_2$

**Reduction** is gain of electrons, loss of  $\text{O}_2$  (using carbon)

### 4.1.2 The reactivity series

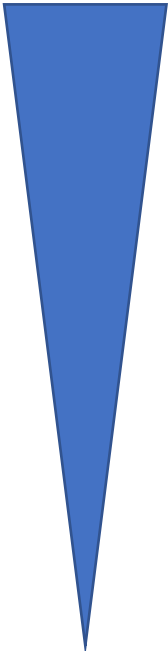
**Reactivity series** – need to memorize

Please send lions, cats, monkeys and cute zebras into the lovely, hot countries, signed general penguin

			Reaction with water	Reaction with dilute acid	
	Most reactive	Potassium	Ka	Explode	
		Sodium	Na		
		Lithium	Li		
		Calcium	Ca	<ul style="list-style-type: none"> <li>• Fizz</li> <li>• Giving off <math>\text{H}_2(\text{g})</math></li> <li>• Leaving an alkaline solution of metal hydroxide</li> </ul>	<ul style="list-style-type: none"> <li>• Fizz</li> <li>• Giving off <math>\text{H}_2(\text{g})</math></li> <li>• Forming a salt</li> </ul>
		Magnesium	Mg	Very slow reaction	
		Aluminium	Al		
		Carbon	C		
		Zinc	Zn	Very slow reaction	<ul style="list-style-type: none"> <li>• Fizz</li> <li>• Giving off <math>\text{H}_2(\text{g})</math></li> <li>• Forming a salt</li> </ul>
		Iron	Fe		
		Tin	Sn	Slight reaction with steam	React slowly with warm acid
		Lead	Pb		
		Hydrogen	H		
		Copper	Cu	Native metals so unreactive	
		Silver	Ag		
		Gold	Au		
Least reactive	Platinum	Pt			

### 4.1.3 Extraction of metals and reduction

**Ore** - a rock that contains enough metal to make it economical to extract

			Extraction of metals	
	Most reactive	Potassium	Ka	Extract by electrolysis
		Sodium	Na	
		Lithium	Li	
		Calcium	Ca	
		Magnesium	Mg	
		Aluminium	Al	
		Carbon	C	
		Zinc	Zn	<b>Metal ore</b> Extract by metal oxide by <b>reduction</b> using <b>carbon</b> <b>Why carbon?</b> C is more reactive
		Iron	Fe	
		Tin	Sn	
		Lead	Pb	
		Hydrogen	H	
		Copper	Cu	Occur native in the ground
		Silver	Ag	
	Gold	Au		
Least reactive	Platinum	Pt		

### 4.1.4 Oxidation and reduction in terms of electrons (HT only)

#### Displacement reaction

- More reactive displace less reactive elements
- **More reactive(s) + compound(aq) → less reactive(s) + compound(aq)**

Eg Magnesium + copper chloride → copper + magnesium chloride



- Magnesium loses  $2e^-$  & become  $\text{Mg}^{2+}$  → oxidation



- Copper gains  $2e^-$  & become Cu → reduction



#### Ionic equation

- Solid and liquid don't separate

**Ionic Equation**

Write the complete, total ionic and net ionic equations for the reaction of aqueous sodium carbonate with aqueous barium nitrate.

**Word equation:**

sodium carbonate + barium nitrate → sodium nitrate + barium carbonate

**Complete chemical equation:**

$$\text{Na}_2\text{CO}_3 + \text{Ba}(\text{NO}_3)_2 \rightarrow 2\text{NaNO}_3 + \text{BaCO}_3(\text{s})$$
**Total ionic equation:**

$$2\text{Na}^{+1} + \text{CO}_3^{-2} + \text{Ba}^{+2} + 2\text{NO}_3^{-1} \rightarrow 2\text{Na}^{+1} + 2\text{NO}_3^{-1} + \text{BaCO}_3(\text{s})$$
**Net ionic equation:**

$$2\cancel{\text{Na}^{+1}} + \text{CO}_3^{-2} + \text{Ba}^{+2} + 2\cancel{\text{NO}_3^{-1}} \rightarrow 2\cancel{\text{Na}^{+1}} + 2\cancel{\text{NO}_3^{-1}} + \text{BaCO}_3(\text{s})$$

$$\text{Ba}^{+2} + \text{CO}_3^{-2} \rightarrow \text{BaCO}_3(\text{s})$$

## 4.2 Reactions of acids

**Important formulae**

Acid + alkali / base → salt + water

Acid + metal → salt + hydrogen

Acid + metal oxide → salt + water

Acid + metal hydroxide → salt + water

Acid + metal carbonate → salt + water + carbon dioxide

### 4.2.1 Reactions of acids with metals

(Reactivity: see table in 4.1.2 The reactivity series)

**Acid + metal → salt + hydrogen**

Acid	Hydrochloric acid	Sulfuric acid
Metal	Metal + hydrochloric acid → metal chloride + hydrogen $\text{M} + 2\text{HCl} \rightarrow \text{MCl}_2 + \text{H}_2$	Metal + sulfuric acid → metal sulfate + hydrogen $\text{M} + \text{H}_2\text{SO}_4 \rightarrow \text{MSO}_4 + \text{H}_2$

Knowledge of reactions limited to those of magnesium, zinc and iron with hydrochloric and sulfuric acids

Acid	Hydrochloric acid	Sulfuric acid
Magnesium	$\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$	$\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$
Zinc	$\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$	$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$
Iron	$\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$	$\text{Fe} + \text{H}_2\text{SO}_4 \rightarrow \text{FeSO}_4 + \text{H}_2$

### 4.2.2 Neutralisation of acids and salt production

(See 4.2.1 Reactions of acids with metals – Important formulae)

(See 4.2.4 The pH scale and neutralization)

**Redox reaction - reduction and oxidation happens at the same time**

Zinc displaces copper from a solution of copper(II) sulfate. Using ionic equations, determine which species undergoes oxidation and which species undergoes reduction. (2)

- Equation:  $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$
- Ionic equation:  $\text{Zn} + \text{Cu}^{2+} + \text{SO}_4^{2-} \rightarrow \text{Zn}^{2+} + \text{SO}_4^{2-} + \text{Cu}$

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- $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ 
  - Zn oxidised as lost electron
- $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$  = reduction
  - H reduced as gained electron

### 4.2.3 Soluble salts

#### Acid + metal oxide $\rightarrow$ salt + water

For example

Copper oxide + sulfuric acid  $\rightarrow$  copper sulfate + water



Notes

- Metal oxides are often insoluble in water – reaction mixture needs warming
- Ionic equation is always  $\text{H}^+(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l})$
- This is a **neutralisation reaction** as  $\text{H}^+$  are used up

#### Acid + metal hydroxide $\rightarrow$ salt water

For example

Sodium hydroxide + nitric acid  $\rightarrow$  sodium nitrate + water



Notes

- Ionic equation is always  $\text{H}^+(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l})$
- This is a **neutralisation reaction** as  $\text{H}^+$  are used up

#### Acid + metal carbonate $\rightarrow$ salt + water + carbon dioxide

For example

Calcium carbonate + hydrochloric acid  $\rightarrow$  calcium chloride + water + carbon dioxide



Notes

- When acids react with carbonates, there is fizzing as  $\text{CO}_2(\text{g})$  is formed
- Ionic equation is always  $\text{H}^+(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l})$
- This is a **neutralisation reaction** as  $\text{H}^+$  are used up

Describe how to make pure, dry sample of named salt from information provided.

**Describe a safe method for making pure crystals of copper sulfate from copper carbonate and dilute sulfuric acid. Use the information in the figure above to help you. In your method you should name all of the apparatus you will use. (6)**

1. Pour sulfuric acid in the beaker
2. Add copper carbonate one spatula at a time until copper carbonate is in excess or until no more effervescence occurs
3. Filter excess copper carbonate using filter paper and funnel
4. Pour solution into evaporating basin
5. Heat using Bunsen burner
6. Leave to crystallise / for water to evaporate
7. Decant solution
8. Pat dry using filter paper
9. Wear safety goggles

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**Level 3 (5 - 6 marks):** A coherent method is described with relevant detail, and in correct sequence which demonstrates a broad understanding of the relevant scientific techniques and procedures. The steps in the method are logically ordered. The method would lead to the production of valid results.

#### 4.2.4 The pH scale and neutralization

Acid	Alkalis
<ul style="list-style-type: none"><li>• Produce H<sup>+</sup> in aq solution</li><li>• Proton donors</li></ul>	<ul style="list-style-type: none"><li>• When dissolved in water, produce OH<sup>-</sup></li><li>• Proton acceptors</li></ul>

##### pH scale (0-14)

- A measure of the acidity / alkalinity of a solution using universal indicator or pH probe
- pH 7 = neutral solution
- pH less than 7 = acids(aq)
- pH more than 7 = alkali(aq)
- Stronger the acid, lower the pH
- As pH decreases by 1 unit, H<sup>+</sup> concentration of solution increases by factor of 10

##### Neutralisation reaction

**Acid + alkali → salt + water**

- Ionic equation is always  $\text{H}^+(\text{aq}) + \text{OH}^- \rightarrow \text{H}_2\text{O}(\text{l})$
- H<sup>+</sup> are used up

#### 4.2.5 Titrations (chemistry only)

**Describe how to carry out titrations using strong acids and strong alkalis (limited to sulfuric, hydrochloric and nitric acids only) to find the reacting volumes accurately (6)**

1. Use pipette to measure 25cm<sup>3</sup> of alkali into conical flask.
2. Fill burette with acid using funnel
3. Add 3 drops of indicator eg phenolphthalein to conical flask & swirl
4. Place conical flask on white tile directly below burette
5. Record initial reading of acid in burette
6. Open the burette tap & add acid dropwise from burette with swirling towards endpoint
7. Close burette when colour changes
  - Phenolphthalein - pink → colourless
  - Methyl orange - yellow → red
  - Litmus - blue → red
8. Record final reading of acid in burette & calculate titre.
9. Repeat 1-8 till obtain concordant result (within 0.1cm<sup>3</sup> of each other)

**The student used phenolphthalein as an indicator rather than universal indicator. Explain why a mixed indicator should not be used for an acid-base titration (2)**

- Mixed indicators change colour gradually over a pH range
- During an acid-base titration, you want to see a sudden colour change at the endpoint, so you need to use a single indicator

**Why is pipette used to measure HCl? (1)**

- Pipette can measure out a known volume accurately

### Why is burette used to measure sodium carbonate? (1)

- Burette can measure out an unknown volume accurately

## 4.2.6 Strong and weak acids (HT only)

**Strong acid** - completely ionised in aqueous solution

- Eg hydrochloric, nitric and sulfuric acids

**Weak acid** - partially ionised in aqueous solution

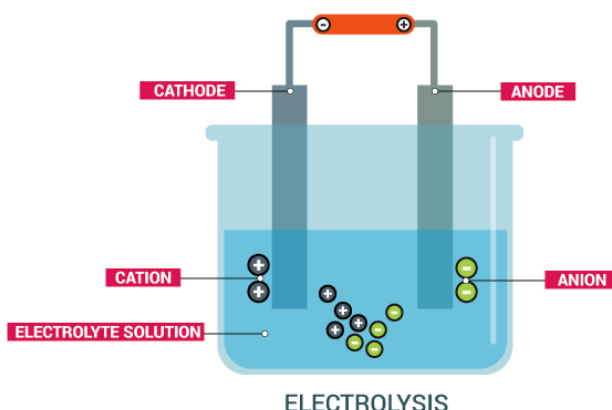
- Eg ethanoic, citric and carbonic acids

**Dilute** - has less acid molecules in a give volume of solution

**Concentrated** - has more acid molecules in a give volume of solution

## 4.3 Electrolysis

### 4.3.1 The process of electrolysis



#### Electrolysis

- When ionic substance is molten (electrolyte)
- Ions move freely & pass electric current
  - +ve ions to move to -ve electrode (**cathode**)
  - -ve ions to move to +ve electrode (**anode**)
- Ions are discharged at electrodes producing elements

#### Why is electrolysis used?

- To extract metals from molten compounds

#### Why is electrolysis expensive? (2)

- Uses large amount of energy
- To melt compounds
- To produce electric current

### 4.3.2 Electrolysis of molten ionic compounds

When a simple ionic compound (eg lead bromide) is electrolysed in molten state using inert electrodes

- Metal (lead) is produced at cathode
- Non-metal (bromine) is produced at anode

Students should be able to predict the products of electrolysis of binary ionic compounds in the molten state

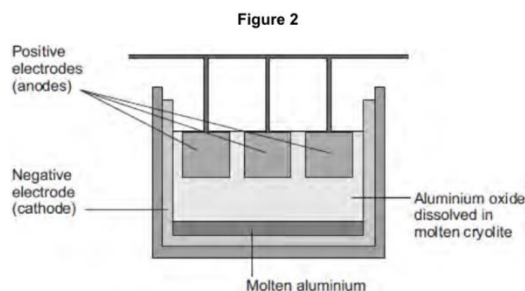
### 4.3.3 Using electrolysis to extract metals

#### Extraction of aluminium

- Aluminium is manufactured by electrolysis of a molten mixture of aluminium oxide & cryolite using carbon as +ve electrode.

<b>Anode</b>	$2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$ (oxidation)
<b>Cathode</b>	$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$ (reduction)
<b>Overall</b>	$2\text{Al}_2\text{O}_3(\text{l}) \rightarrow 4\text{Al}(\text{s}) + 3\text{O}_2(\text{g})$

(b) Aluminium is extracted by electrolysis, as shown in Figure 2.



**Why can aluminium not be extracted by heating aluminium oxide with carbon? (1)**

- Al is more reactive than C so it displaced C

**Explain why aluminium forms at the negative electrode during electrolysis (1)**

- $\text{Al}^{3+}$  is +ve charged so it will attract to -ve charged electrode, where they gain  $\text{e}^-$
- $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$

**Explain why +ve electrode is used up during the process (3)**

- Electrode is made of carbon
- $\text{O}_2$  is produced at +ve electrode
- $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$

**Why does the carbon anode used in the electrolysis cell need to be continually replaced (3)**

- $\text{O}_2$  is produced at +ve electrode
- $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$
- Carbon anode wears away

**Why cryolite is used? (1)**

- Lower melting pt of aluminium oxide



### 4.3.4 Electrolysis of aqueous solutions

In aq solutions

<b>Anode</b>	Halide present → given off If not → O <sub>2</sub> released $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$ (oxidation)
<b>Cathode</b>	Cation more reactive than H <sub>2</sub> → H <sub>2</sub> produced If not → metal produced $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ (reduction)

In liquid - ions split up eg NaF → Na + F<sub>2</sub>

### 4.3.5 Representation of reactions at electrodes as half equations (HT only)

Common electrolytes

Aq solution (ions present)	Products at <b>Anode</b>	Products at <b>Cathode</b>
Concentrated NaCl	$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ (oxidation) Cl <sub>2</sub> released	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ (reduction) H <sub>2</sub> released
Concentrated CuSO <sub>4</sub>	$4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$ (oxidation) O <sub>2</sub> released	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (reduction) Cu less reactive than H <sub>2</sub> so Cu produced