PHYSICAL SCIENCE

MSCE NOTES

CHEMICAL REACTIONS I

- Chemical Reaction is the rearrangement of atoms to form new substance.

DEDUCING AND BALANCING CHEMICAL EQUATIONS

Balancing equation

To balance an equation, do the following;

1. Write the correct formulae for **reactants** and **products.**

2. Balance the atoms on both sides of the equation by multiplying the molecules of reactants and products.
3. Check that the balanced equation has **equal number of atoms** on both sides

- Chemical equation is the statement about a reaction

Example 1.
Carbon reacted with oxygen gas to produce carbon dioxide. Deduce a chemical equation from the statement above.

$C + O_2$ --------------- CO_2
(Balanced)

Example 2.

Sodium metal reacts with water to give off hydrogen gas and leave a solution of sodium hydroxide. Balance the equation.

$Na + H_2O$ ------------------ $NaOH + H_2$ (Unbalanced H)

There is more H on the right than on the left. Multiply the molecule where there is H on the left. (Try to do like that until that the equation is balance)

$2Na\ (s) + 2\ H_2O\ (l)$ ------------- $2NaOH\ (aq) + H_2\ (g)$ (Balanced)

THE MOLE

Mole is the RFM or RAM of substance expressed in grams.

Examples

1 mole of sodium Na is 23g
1 mole of calcium is 40g
1 mole of sulphur, S is 32
1 mole of carbon dioxide, CO2 is 44g
1 mole of sulphuric acid is 98g
2 moles of oxygen gas is 32g

MOLE CALCULATION

Mole can be calculated using the following formula;

1. Number of moles = $\frac{\text{Mass of element}}{\text{RAM}}$

N = $\frac{m}{\text{RAM}}$

2. Number of moles = $\frac{\text{Mass}}{\text{RFM (RMM)}}$

3. Number of moles of particles = Number of particles

Avogadro Number

*Avogadro number contains 6.02×10^{23} particles

Examples.

Calculate the number of moles in

(a) 108g of aluminium; Al
(b) 16g of sodium hydroxide, NaOH
(c) 19.6g of sulphuric acid

Working:

(a) # of moles = $\frac{m}{RAM}$ = 108 = 4 moles

27

(b) # of moles = $\frac{m}{RFM}$ = $\frac{16}{40}$ = $\underline{0.4 \text{ moles}}$

(c) # of moles = $\frac{m}{RFM}$ = $\frac{19.6}{}$ = $\underline{0.2 \text{ moles}}$

98

Example 2

Calculate the mass of

(a) 2 moles of iron
(b) 0.2 moles of carbon dioxide, CO_2
(c) 3 moles of oxygen gas, O_2

Working;

(a) Mass = # of moles x RAM

= 2 x 56
= <u>112g</u>

(b) Mass = # of moles x RMM
= 0.2 x 44
= <u>8.8g</u>

(c) Mass = # of moles x RMM

= 3 x 32
= <u>96g</u>

<u>**EMPIRICAL AND MOLECULAR FORMULARS**</u>

Empirical formula is a formula showing the simplest ratio of atoms present.

Example

Calculate the empirical formula of an organic compound containing 92.3 % carbon and 7.7 % hydrogen by mass. The RMM of the organic compound is 78. What is its molecular formular.

Working;

	C	**H**
% mass :	92.3	7.7
In 100g :	92.3	7.7
Moles:	$\frac{92.3}{12}$	$\frac{7.7}{1}$

Ratio of moles: 7.7 7.7

Simplest ratios of moles: $\dfrac{7.7}{7.7}$ $\dfrac{7.7}{7.7}$

1 1

Therefore the empirical formular is $\underline{CH}$

To work out the molecular formular,

$$\# \text{ of molecules} = \frac{\text{RMM of compound}}{\text{RMM of empirical formula}}$$

$$= \frac{78}{13}$$

$$= \underline{6}$$

Therefore molecular formular is 6x(CH) = $\underline{C_6H_6}$ (Benzene)

Example 2

Calculate the empirical formula of an organic compound containing 40% carbon, 6.67% hydrogen and 53.33% oxygen. The relative molecular mass of the compound is 90. What is its molecular formula?

Working;

	C	H	O
% mass :	40	6.67	53.33

In 100g : 40
6.67 53.33

Moles: $\underline{40}$
$\underline{6.67}$ $\underline{53.33}$
12 1
16

Ratio of moles: 3.33
6.67 3.33

Simplest ratios of moles: $\underline{3.33}$
$\underline{6.67}$ $\underline{3.33}$
3.33
3.33 3.33

1
2 1

Empirical formula $\underline{CH_2O}$

Molecular formula = RMM of compound

RMM of empirical formula

$$= \frac{90}{30} = 3$$

$= 3x(CH_2O)$

$= \underline{C_3H_6O_3}$

PERCENTAGE COMPOSITION OF ELEMENT IN A COMPOUND

To work out the percentage composition of element, follow the following steps;

1. From the molecular formular work out the RFM of the compound
2. Work out the mass of the element in the compound.
3. Calculate the % mass using the formular;

$$\% \text{ mass} = \frac{\text{mass of element}}{\text{RFM of the compound}} \times 100$$

Example 1.
Work out the % of C in propene

Working;

RFM of propene C3H6 = 3C + 6H

(3 x 12) + (6 x 1)

= 36 + 6 = 42

Mass of C in propene = 3 x 12 = 36

% mass = mass of C / RMM = 36 x 100 / 42 = 85.7%

Example 2.

What is the % composition of potassium in potassium nitrate, KNO_3?

Working;

RFM of KNO_3 = 39 + 14 + 48 = 101

Mass of K = 39

$$\% \text{ mass of K} = \frac{39 \times 100}{101} = \underline{38.62\%}$$

MOLARITY / CONCETRATION

- Molarity or concentration is measured in moles per cubic decimeter (mol dm^{-3})
- When one mole of a substance is dissolved in water and the solution is made up to 1 dm^3 ($1000cm^3$), 1 molar (1M) solution is produced.

$$\textbf{Molarity / Concentration} = \frac{\textbf{Number of moles}}{\textbf{Volume (in } \mathbf{dm^3}\textbf{)}}$$

Example;

Calculate the concentration of a solution of sodium hydroxide , NaOH, which was made by dissolving 10g of solid sodium hydroxide in water making up to $250cm^3$.
(Ar Na = 23; O = 16; H = 1)

Working.

$$\textbf{Concetration} = \frac{\text{\# of moles of NaOH}}{\text{Volume (in dm}^3\text{)}}$$

$$\text{\# of moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{RFM of NaOH}} = \frac{10}{40} = \underline{0.25 \text{ moles}}$$

$$\textbf{Concetration} = \frac{\text{\# of moles of NaOH}}{\text{Volume (in dm}^3\text{)}} = \frac{0.25 \text{ moles}}{250 \text{ dm}^3}$$

1000

= $\underline{0.25 \text{ x } 1000}$

250

$\underline{= 1M}$

Example 2

Calculate the concentration of solution containing 9.8g of sulphuric acid dissolved in water and made up to 500 cm^3.

Working.

# of moles of H_2SO_4 in 9.8g = $\underline{\text{Mass}}$

= $\underline{0.98}$ = $\underline{0.1 \text{ moles}}$

RFM

98

Molarity = $\frac{0.1\text{moles} \times 1000}{500}$ = **0.2M**

CALCULATING THE MASS OF SOLUTE TO DISSOLVE TO PREPARE SOLUTION OF KNOWN VOLUME AND CONC ETRFATION

Example 1.

Calculate the mass of potassium hydroxide, KOH, that needs to be used to prepare 500cm^3 of a
2 M solution in water. **(Ar, H = 1; O = 16; K = 39)**

Working

- Mass can be calculated from the **number of moles in a solution**

Mass = # of moles x RFM
* (we do not have number of moles, therefore calculate number of **moles in a solution**)

# of moles = Molarity x Volume (dm^3)

$$= 2 \times \frac{500}{1000} dm^3$$

$$= \underline{1 \text{ mole}}$$

Mass = # of moles x RFM of KOH

= 1 x 56

= <u>56g</u>

Example 2

Calculate the mass of potassium nitrate, KNO_3, which needs to be used to prepare $200cm^3$ of a 4M solution.

Working

# of moles = Molarity x Volume (dm^3)

$$= 4 \times \frac{200}{1000} = \underline{0.8 \text{ moles}}$$

Mass = # of moles x RFM of KNO_3

= 0.8 x 101

= 80.8g

STANDARD SOLUTION

Standard solution is a solution whose concetration is known.

Eg.1M solution, 0.5 M solution

Preparing standard solution

- Calculate the mass of the solute to be dissolved

- Measure the mass on a triple beam balance
- Put the solute in a beaker or conical flask
- Pour some water in the beaker / flask with solute to dissolve the solute
- Transfer the solution into the conical flask of known volume.
- Add some more water to the mark on the neck and shake the flask.
- The solution prepared is the standard solution

DETERMINING THE CONCETRATION OF A SOLUTION OF UNKNOWN CONCETRATION USING TITRATION

- Measure the accurate volume of the standard solution, **eg 25cm^3**
- Pour the **standard solution** in a conical flask or beaker.
- Add indicator to the standard solution. If you add phenolphlein to base it will turn pink.
- Pour the **unknown solution** in the burette and record the volume, **eg 50 cm^3**
.
- Gradually titrate the unknown solution into the flask of standard solution.
- Shake the solution regularly until the colour of the solution changes to colourless.
- Stop titration when the colour changes and record the volume, **eg 30 cm^3.**
- Deduce the balanced equation of the reactants eg

$NaOH (aq) + HCl (aq) \rightarrow NaCl (aq) + H_2O$

- Calculate the Molarity / concetration of the unknown using the formular:

$$\mathbf{M_1 V_1 \text{ (base)} = M_2 V_2 \text{ (acid)}}$$

(where $\mathbf{M_1}$ and $\mathbf{V_1}$ are Molarity and volume of **base** and $\mathbf{M_2}$ and $\mathbf{M_2}$ are Molarity and volume of **acid**)

Eg If $M_1 = 4M$; $V_1 = 25\ cm^3$; $V_2 = 20\ cm^3$; Then

$$\mathbf{M_2 = \underline{M_1 V_1 \text{ (base)}} = \underline{4M \times 25\ cm^3 \times 1000} = \underline{5M}}$$

V_2 (acid)

20 cm^3 x 1000

MOLAR VOLUME

- Molar volume is the volume occupied by 1 mole of gas.
- 1 mole of gas occupies a volume of 24 dm^3 at room temperature and pressure.
- Gases are easily calculated in moles rather than in masses.

Number of moles of gas = <u>Volume of gas</u>

24 dm^3

Volume of gas = # of moles x 24dm^3

Example 1.

Calculate the number of moles of ammonia , NH3, in a volume of 72 dm^3 of the gas measured at rtp.

Working

# of moles = $\frac{\text{Volume of gas}}{24\ dm^3} = \frac{72}{24} = \underline{3\ \text{moles}}$

Example 2.

Calculate the volume of carbon dioxide gas, CO_2, occupied by 0.5 mole of gas at rtp.

Working.

Volume = # of moles x 24 dm^3 = 0.5 mole x 24 dm^3 = $\underline{12\ dm^3}$

REACTING MASSES

- When substances **react**, the **first thing** to do in order to find the reacting masses is to **deduce chemical equation** as shown in examples.

Example1.
What mass of water is produced when 4g of copper oxide is reduced by hydrogen?
(Ar, Cu = 64 ; O = 16 ; H = 1)

Working

Equation: $CuO + H_2$ ------------ $Cu + H_2O$

1 mole of CuO ---------- 1 mole of H_2O

RMM 80g CuO ---------- 18g H_2O

1g CuO ---------- $\frac{18}{80}$ g H_2O

4g CuO ---------- $\frac{18}{80}$ x 4g H_2O = 0.9g

Example 2

Calculate the mass of NaOH needed to neutralize 7,3g HCl.

Working.

Equation : NaOH + HCl -------------- NaCl + H_2O

1 mole NaOH -------------- 1 mole HCl

40g NaOH -------------- 36.5g HCl

$$\frac{40g}{36.5} \times 7.3g \text{ NaOH}$$

= 8g NaOH

Example 3.
What volume of Carbon dioxide at rtp is produced by burning 12g of Carbon?

Working.
Equation: $C(s) + O_2(g) \text{ ------------ } CO_2\,(g)$

1 mole of C produce 1 mole of CO_2

12g of carbon produce 24 dm^3 of CO_2

Example 4

HEATS OF REACTION

Classification of Reactions

Reactions are classified as;

1. Exothermic
2. Endothermic

1. Exothermic reaction is the reaction in which heat is given out to the surrounding.

Eg. Reaction of sulphuric acid with water.

1. H_2SO_4 + H_2O ------------ 2 H^+ (aq) + SO_4^{2-} (aq) + heat energy

Burning fuels:

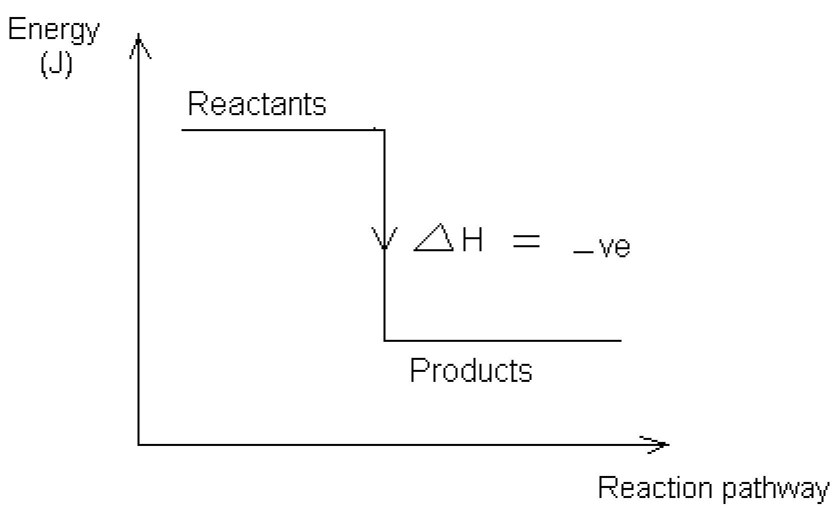
2. CH_4 + $2O_2$ ------------- CO_2 + $2H_2O$ + heat energy

3. Addition of water to anhydrous copper sulphate to form hydrous copper sulphate

$CuSO_4 + H_2O$ ----------- $CuSO_4.5H_2O$ + heat

- The heat which is given out rises the temperature of the surrounding.
- The products are at lower energy level than the reactants.
- The difference in height represents the energy given out.
- The change in energy (ΔH) is negative (-) because initial energy is subtracted from the final ; ΔH = (Energy of products – Energy of reactants)

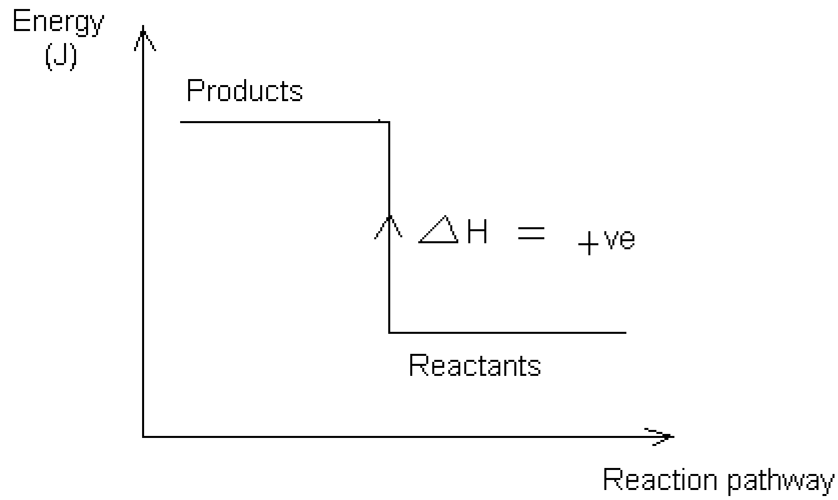
Energy Level Diagram for Exothermic Reaction



2. **Endothermic** reaction is the reaction in which heat is taken in from surrounding.

- This is shown by a decrease in temperature of the surrounding .
- The products are at a higher energy level than reactants.
- The change in energy (ΔH is positive (+) because the initial energy is subtracted from the final energy; ΔH = (Final energy - Initial energy).

Energy Level Diagram for Endothermic Reaction



Bond Formation and Bond Breaking

- Bond breaking is endothermic because energy must be taken in (supplied) to break the existing bond
- Bond formation is exothermic because energy is released (given out) when new bonds form.

CHEMICAL REACTIONS II

<u>OXIDATION AND REDUCTION</u>

Oxidation is the loss of electrons by an element or compound.

Example: $Mg(s) - e^- \rightarrow Mg^{+2}$- (aq)

$Zn (s) - e^- (g) \rightarrow Zn^{+2}$ (aq)

- During **oxidation** in the above examples, Mg and Zn have lost 2 electrons each and have changed to Mg^{+2} and Zn^{+2} respectively. Losing electrons means that Magnesium atom and Zinc atoms ha ve been **oxidized.** The figure has increased.

Reduction is the gain of electrons

Example: $Cu^{+2} (aq) + e^- \rightarrow Cu^0 (s)$

$Ag^{+1} (aq) + e^- \rightarrow Ag^0$ (s)

- We started with copper (Cu^{+2}) ions with a charge of +2 in Cu^{+2} and ended with copper(Cu^0) atoms. This shows that copper ions have gained electrons, this means copper ions have been **reduced.** Ag at first had a charge of +2 but after gaining an electron, it changed to 0. The figure has reduced.

Oxidation Number

- The **charge** on an atom or ion is called its oxidation number

Mg^0 (s) + Cu^{+2} (s) ------------- Mg^{+2} (aq) + Cu (s)

- On the reactant side, Mg has an **oxidation number** of **0** while on the product side Mg has an **oxidation number** of **+2.**, **magnesium is oxidised**

- A substance is **oxidized** when its **oxidation number** is **increased.**

- On the reactant side, Cu has an **oxidation number** of **+2** while on the product side Cu has an **oxidation number of 0.** However, **copper is reduced**

- A substance is **reduced** when its **oxidation number is decreased.**
- An **element** in its **free state** has an **oxidation number of zero (0)**
- An element is considered to be **oxidized** if its **oxidation number** is **positive.**
- An element is considered to be **reduced** if its **oxidation number** is **negative.**

Calculation of Oxidation Number

Rules:

1. Oxygen has oxidation number of +2
2. Chlorine has oxidation number of -1
3. Hydrogen has oxidation number of +1
4. For neutral molecules, the **sum of separate charges is equal to 0.**
5. For **charged ion, the sum of separate charges = final charge on ion.**

Example

Work out the oxidation number of each element in the compounds below.

a. S in SO_2
b. Mg in $MgCl_2$
c. SO_4^{-2}

Working

a. S in SO_2

$$S + 2(O) = 0$$
$$S + 2(-2) = 0$$
$$S - 4 = 0$$
$$\underline{\mathbf{S = +4}}$$

b. Mg in $MgCl_2$

$$Mg + 2(Cl) = 0$$
$$Mg + 2(-1) = 0$$
$$Mg - 2 = 0$$
$$\underline{\mathbf{Mg = +2}}$$

c. S in SO_4^{-2}

$$S + 4(O) = -2$$
$$S + 4(-2) = -2$$
$$S - 8 = -2$$
$$\underline{\mathbf{S = +6}}$$

Redox Reactions

If both **oxidation** and **reduction reactions** happen at the same time, they are called **redox reactions**

Half Reactions Equations

A half reaction is a separate equation showing which substance loses or gains an electron in a redox reaction. If the two equations are added together, they give **overall equation.**

$Mg(s)$ + Cu^{2+} (aq) ------------- Mg^{2+} (aq) + $Cu(s)$

Example:

Oxidation half- equation: Mg(s) -------------- Mg^{2+}(aq) + $2e^-$

Reduction half-equation: Cu^{2+} (aq) + $2e^-$ ------------ Cu^0 (aq

Overall equation : $Mg(s) + Cu^{2+}(aq) \text{-------------} Mg^{2+}(aq) + Cu(s)$

Reducing and Oxidizing Agents.

- An **oxidizing agent** is a substance which **accepts electron**.
- A **reducing** agent is a substance which **donates electron**.
- The **oxidation number** of oxidizing agent is **reduced.**
- The **oxidation number** of reducing agent is **increased.**

Referring to the above examples of equations; **Mg is a reducing agent** because

(1) its oxidation number has increased.
(2) It has donated electron to Cu^{2+}

Cu^{2+} is an oxidizing agent because

(1) its oxidation number has reduced
(2) it has accepted electrons

DISPLACEMENT REACTION

Displacement reaction is a reaction in which an element displaces another element in redox reaction

Example:
A piece of magnesium (Mg) metal is placed in an aqueous solution of silver nitrate as shown in the diagram below.

(Diagram)

Results

- Magnesium ribbon (metal) dissolved
- The ribbon became covered with a shiny silvery-black coating of silver metal.

Explanation:

- When a metal dissolves, it **ionizes** as shown in the **oxidising half equation** below;

$$Mg(s) \text{ -------------- } Mg^{2+}(aq) + 2e^-$$

- The silver metal coating on the magnesium metal is made up of **silver atoms** as shown by the **reducing half equation** below

$$Ag^+(aq) + e^- \text{ ------------ } Ag^0(s)$$

From the example above, magnesium which was a solid metal at first dissolved and turned into solution, that is magnesium ($Mg^{2+}(aq)$) ions . Silver ($Ag^+(aq)$) which were in solution at first, have been displaced by magnesium.

- We can say that magnesium has **donated electrons** to silver. In the **displacement table (series),** the metal which **displaces (donates electrons**) the other **is placed on top** and the one which **is displaced (accepts electrons)** is placed **at the bottom**

- Using the explanation above, we can develop a **displacement table (series)** as shown below

Potassium K
Sodium Na
Lithium Li
Calcium K
Magnesium Mg
Aluminum Al
Zinc Zn

Iron	Fe
Tin	Sn
Lead	Pb
Hydrogen	H
Copper	Cu
Silver	Ag

ELECTROCHEMICAL SERIES

- Another **table** of the same nature is called **electrochemical series.**
- This table compares the **position of the metal in the table with its reducing power (ability to donate electrons)**
- The metal **on top donates its electron easily,** so it **displaces** other metals **below it** easily. The metal **on top has higher reducing power**

$K^+ (aq) \quad + \quad e- \quad \rightarrow \quad K$

$Na^{+}(aq) + e^{-} \rightarrow Na$

$Li^{+}(aq) + e^{-} \rightarrow Li$

$Ca^{2+}(aq) + 2e^{-} \rightarrow Ca$

$Mg^{2+}(aq) + 2e^{-} \rightarrow Mg$

$Al^{3+}(aq) + 3e^{-} \rightarrow Al$

$Zn^{2+}(aq) + 2e^{-} \rightarrow Zn$

etc

Predicting Displacement reaction

- From the table above, we can predict the metal that would displace the other from a solution of its salt. For example, if sodium metal is placed in a solution of zinc salt, reaction will take place as follows

$Na(s) + Zn^{2+}(aq) \rightarrow Na^{+}(aq) + Zn(s)$

- Reaction will take place because sodium metal is on top of zinc in the table and will donate electrons easily.

If magnesium metal is placed in a solution of potassium chloride. Reaction will not take place as follows;

Mg (s) + K^+ (aq) $\rightarrow$ no reaction (this is because magnesium is at the bottom and can not displace potassium which is above it.)

RUSTING / CORROSION

Rust is an orange –red powder consisting of hydrated iron (III) oxide ($Fe_2O_3.XH_2O$)

This reaction is a **redox reaction** as shown in the half equations below.

Both iron and water are essential for iron to rust.

How rust is formed

1. Iron dissolves first and **ionise** as shown in the half equation

Oxidation : $Fe(s) - 2e \rightarrow Fe^{2+}(aq)$

2. The electrons are used by water and air as shown below;

Reduction: $2H_2O\ (l) + O2\ (g) + 4\ e- \rightarrow 4\ OH^-\ (aq)$

3. the hydroxyl react with $Fe^{2+}(aq)$ ions more water and air to form rust.

$4Fe\ (OH)_2\ (s) + 2H_2O\ (l) + O_2\ (g) \rightarrow 4Fe\ (OH)_3\ (s)$ (rust)

Conditions for Rusting

- Conditions can be known by carrying out an experiment as follows.

Aim: To find conditions for rusting

Materials:
3 test tubes, nails, tap water, boiled water, salt water, rubber bung, oil.

Procedure
Set up 3 boiling tubes as shown below.

(Diagrams)

Test tube A (Air and tap water present)
The iron nail is immersed in tap water in the tube. After two weeks the nails show rusting.

Test tube B (air but no water)

- The anhydrous calcium chloride is put in to absorb water vapour. The tube is tightly sealed to prevent more air entering.

- After 2 weeks, no rusting occurred.

Test tube C (water but no air)

- The water is boiled for five minutes to remove all dissolved gases.
- The boiled water is poured into the test tube while still hot.
- The nail is inserted and the tube is sealed with a layer of oil to prevent any air entering.
- After 2 weeks, no rusting occurred.

Conclusion

- Rust did not form in test tube B where there was no water and again did not form in test tube C where there was no air. However, rust formed in test tube A where there was air and water present. This shows that

both air and water **must** be present for rust to occur.

PREVENTION OF CORROSION / RUST

- To prevent iron from rusting, it is necessary to stop air (oxygen) or water coming into contact with iron.
- Some of the ways of preventing rusting are:

1. Painting

- Ships, lorries. Cars bridges are painted to prevent water and air from coming into contact with iron

2. Oiling / greasing

- The iron or steel in the moving parts of the machinery are smeared with oil or grease to prevent air and water coming into contact with iron.

3. Galvanising

- Galvanising means coating a metal with zinc metal.
- The zinc corrodes instead of iron. As it corrodes, it donates electrons to iron thereby preventing iron form rusting as shown in the equation;

Zinc *oxidizes* as: $Zn\ (s)\ -\ 2e$ ---------- $Zn^{2+}(aq)$

Iron is *reduced* as: $Fe^{2+}\ (aq)\ +\ 2e$ --------- $Fe\ (s)$

4. Sacrificial protection

- Bars of zinc are attached to the hulls of ships and oil rigs. Zinc is the one which corrodes instead of iron since it is on top in the reactivity series. Iron is reduced while zinc oxidizes as shown in the equations below;

Zinc *oxidizes* as: $Zn\ (s)\ -\ 2e$ ---------- $Zn^{2+}(aq)$

Iron is *reduced* as: $Fe^{2+}\ (aq)\ +\ 2e$ --------- $Fe\ (s)$

5. Coating with plastic

- Air and water is prevented from reaching the metal because of plastic which covers the metal

ELECTROPLATING

Is the covering of a metal with another metal by means of electric current (electrolysis)

- Electrolysis is the process of using electric current to decompose a compound into ions.

Process of electroplating

Example: Electroplating iron with zinc

(Diagram)

- Using the electrolytic apparatus shown above, a piece of pure zinc metal is placed at the anode and iron metal is placed at the cathode.
- Both zinc and iron are immersed in a solution of zinc (Zn^{2+}) ions.
- The direct current (dc) is switched on for several hours.
- The zinc (Zn^{2+}) ions from solution are attracted to the cathode. When they reach the cathode, they gain electrons (**are reduced)** and turn into zinc (Zn) atoms as shown by the chemical equation below:

  At the cathode: $Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$
- Zinc metal at the anode loses electrons (**oxidizes),** and turn into

zinc (Zn^{2+}) ions as shown in the equation below:

At the anode: $\mathbf{Zn\ (s)\ -\ 2e\ \rightarrow\ Zn^{2+}(aq)}$

The electrons passes through the wire to cathode but the ions get into the solution to maintain the concentration of zinc ions.

Note :

The following facts should be known in electroplating

1. The **metal to be electroplated** should be the **cathode**
2. The **metal to be used for electroplating** should be the **anode**
3. The **solution** should contain the **ions of the anode**

PROTON TRANSFER REACTIONS

- A proton is a hydrogen (H^+) ion.
- An acid is a proton (H^+)donor (giver)
- It releases H^+ ions in the solution.
- A base is a proton acceptor.

Examples of Acid –Base reactions.

1. $NH^+(aq) + OH^-(aq) \leftrightarrow H_2O(l) + NH3$
2. $CH3COOH(aq) + OH^- (aq) \leftrightarrow CH_3COO^- + H2O(l)$
3, $HBr(aq) + CN^-(aq) \leftrightarrow HCN(aq) + Br^-$
4. $HCOOH + NH3 \leftrightarrow HCOO^- + NH_4^+$
5. $HSO_4^- + H_2O \leftrightarrow H_3O^+ + SO_4^{-2}$

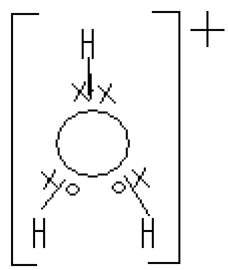
HYDRONIUM

- Hydronium ion is formed when water molecule accepts a proton as shown below

$$H_2O \quad + \quad H^+ \quad \rightarrow \quad H_3O^+$$

(Hydronium ion)

- ➢ The ion is positively charged because the proton (H^+) carries a positive charge
- ➢ Structure of Hydronium is shown below:



CONJUGATE BASES AND CONJUGATE ACIDS

- Conjugate means having common thing.

Conjugate acids to conjugate base

Conjugate acid – H^+ $\rightarrow$ **Conjugate base**

Stronger acid $HCl - H^+ \rightarrow Cl^-$

$H_2SO_4 - H^+ \rightarrow HSO_4^-$

$HNO_3 - H^+ \rightarrow NO_3^-$

$H_3O^+ - H^+ \rightarrow H_2O$

$CH_3COOH - H^+ \rightarrow CH_3COO^-$ *Stronger base*

- From the table above, the **acid on top (strong acid) will donate a proton to the acid below** it.
- Bases **below water** in the table **are bases of strong basic strength** and will accept proton from water to form hydroxyl OH^- ions that will give alkaline solution.
- Bases **above water** will not accept proton from water because they are

bases of weak basic strength. The can not make water lose its hydrogen.

Conjugate base to conjugate acid

Conjugate + H^+ → **Conjugate acid**

$OH^- + \mathbf{H^+} \rightarrow H_2O$

$H_2O + \mathbf{H^+} \rightarrow H_3O^+$

$CH_3COO^- + \mathbf{H^+} \rightarrow CH_3COOH$

$Cl^- + \mathbf{H^+} \rightarrow HCl$

$SO_3^{-2} + \mathbf{H^+} \rightarrow HSO_3^-$

STRENGTH OF ACIDS AND BASES

A. **Strong and weak acids**

- Strength of acid is the ability to donate protons to a base.
- **Strong acid** is the acid which donates protons well or easily. It is

the acid which produces high concentration of hydrogen **H^+ ions** in water solution.

- The strong acid is **completely ionised** in water and produces high concentration of hydrogen ions as shown in the diagram below:

$$HCl + H_2O \rightarrow H_3O^+ (aq) + Cl^- (aq)$$

- For hydrochloric acid, all the molecules break to form **H^+ and Cl^-**.
- A weak acid produces few ions when dissolved in water because it ionises (dissolves) partially as shown below;

$$CH_3COOH + H_2O \leftrightarrow CH_3COO^- + H_3O^+$$

- The (↔) sign means the reaction is **reversible.** This means that when products form, the combine again to reform the reactants.

B. **Strong and weak bases**

- Strong alkali is the base which produces a high concentration of OH^- ions in water eg sodium hydroxide
- When strong base dissolves in water, it breaks up (ionises) completely to form ions.

  Eg: NaOH +

STRONG AND WEAK BASES

- Strong alkali is a base which produces a high concentration of

OH^- ions in water e.g. Sodium hydroxide.

- When strong base dissolves in water, it breaks up completely to form ions e.g.

NaOH(s) ---------→ Sodium ions + Hydroxide ions

- A weak base produces fewer hydroxide ions when it dissolves in water e.g.

NH_3 (g) + H_2O ⇔ NH_4^+ + OH^- (aq)

- Fewer ammonia molecules break so only a low concentration of hydroxide ions is produced

Activity

Complete the following equations

a) CH_3COOH + NaOH ⇔
b) H_2SO_4 + KOH →
c) NH_3 + HBr →

INVESTIGATING STRENGTH OF ACIDS AND BASES USING CONDUCTIVITY APPARATUS

(DIAGRAM)

- ✓ The conductivity apparatus is set as shown above.
- ✓ Connect the circuit with battery and ammeter to the electrodes
- ✓ Immerse the electrodes in the electrolytic bath under test.
- ✓ Check the size of the current on the ammeter
- ✓ The greater the current, the stronger the acid or the base and the weaker the current the weaker the acid or base.