RAY-NOTES® 2009

O-Level Sc (Chemistry) v1.5

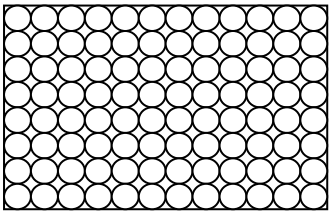
Easy notes summarized for O-Levels

Hong Ray Corporations®

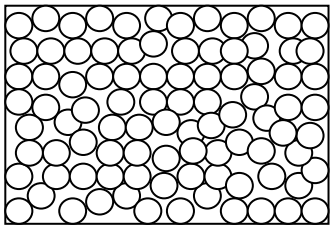
09

Chapter 1 – Summary

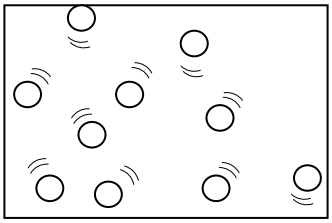
1) Matter is made up of particles.



- Solid (Fixed shape, High Density)
  - Particles close together,
  - Orderly Arranged,
  - Held by strong forces in fixed positions



- Liquid (No fixed shape, High Density)
  - Particles close together,
  - NOT in order,
  - Strong forces, (between particles)
  - Vibrate & Free to move



- Gas (No fixed shape, Low density)
  - Particles Far apart
  - Weak forces (Between particles)
  - Free to move around
  - **Pressure** is due to the particles bouncing off walls of container.

Apparatus To measure:

Mass of chemical, use **Electronic Balance**

Temperature, use **Thermometer**

Vol. of gas, use **Measuring Cylinder** OR **Burette** OR **Pipette**.

Changes of State

1) Melting (Solid → Liquid)
   - Particles gain energy
   - Vibrate more
   - Occurs at melting point
   - Endothermic* Reaction (Because particles gain energy to break bonds)
2) Freezing (Liquid → Solid)
   - Occurs at freezing point (Also the m.p. of pure substance)
   - Exthothermic* Reaction (Because particles give out energy to form bonds)

*__EN__dothermic reaction means heat or energy __EN__TER (absorbed/gained) to break bond!

*__EX__othermic means heat, energy __EX__IT (give out/lose) to form bonds!

Note that for both Freezing & Melting, temp. remains constant during the process!

Chapter 2 – Summary

SEPERATION OF MIXTURES

2.1 – MixTURES & Compounds

1) Compound
   - 2 or more elements chemically joined together
   - Eg. Sodium Chloride
   - Contains Covalent or Ionic Bonds

2) Mixtures
   - 2 or more substances which are not chemically joined together
   - Eg. Iron in Sulphur powder.
   - Can consist of:
     - Elements (Iron & Sulphur powder)
     - Compounds (Salt in water – salt & water are both compounds)
     - Elements + Compounds (Air)

#Pure Substances

1) Fixed composition
2) Fixed Mp/Bp
3) Produces only 1 spot on chromatogram
4) All molecules same (equal)

#Mixture

1) Variable Composition
2) Variable Bp/Mp
3) 2 or more spots on Chromatogram
4) 2 or more diff. molecules

Personal Tips: "Man Becomes what he thinks about" – Mooris Goodman
- Visualize yourself being able to understand this set of notes perfectly. By truly believing, it must be Fact.

Performance improvement Tips: Listen to instrumental musics; (Boroque is NOT recommended) as it can make you tired
- Try ideal musics such as 'Free as a bird' and 'A day with you' by Omar. Also make sure you listen using earpiece, NOT speakers!

2.2 – Purification

(Note that mixtures are Impure substances)

* Soluble means can dissolve in water (eg. Salt)
* Insoluble means cannot dissolve in water (eg. Sand)

1) Insoluble solid & Liquid → Filtration — - SAND & WATER
2) Solid & Liquid solution → Crystallization (Obtaining solid) – Solute — - Copper Sulphate crystals from - Copper Sulphate solution
3) Solid & Liquid solution → Distillation (Obtaining Liquid) – Solvent — - Water from seawater
4) 2 Liquids mixed (Miscible) → Fractional Distillation — - Petroleum to petrol/diesel
5) Mixture of Organic Compounds → Chromatography (Eg. Colour Dyes) — - Separating dyes in inks

Note that "Miscible" means mixable - (Able to be mixed)

 # Filtration – Separating solid from liquid

Possible becoz:

- Liquid particles small enough to go thru filter paper pores
- Solid particles too large.

- Solid obtained – "Residue"
- Liquid obtained – "Filtrate"

#Crystalisation – "Obtaining solid from Solid & Liquid solution" OR
- "Formation of crystals from a cooling liquid/ Saturated solution"

Process:

1. Dissolve solid in solvent to give "solution".
2. Solution heated to evaporate solvent.
3. >>Produces a hot saturated solution.
4. Crystals of pure solid formed on cooling.

General notes:
- A "SATURATED" solution means no more solid can dissolve in it
- Any question that wants the <u>method</u> to obtain (make) ~crystals, the process will be "Crystalisation"!
Eg. "What method is used to make Copper Sulphate Crystals?" Ans: Crystalisation.

#Distillation – Obtaining pure liquid from solution (Eg. To obtain pure water from seawater)

Process:

1. Heat solution in flask.
2. The solution boils
3. Pure liquid turns to vapour, leave the flask.
4. Vapour cooled by condenser, changes back to liquid
5. >>Liquid obtained is called "Distillate" (Collected in conical flask)

Possible because:

- Pure liquid change to gas easily – Low BP
- Solid does not boil (so remains in flask) – High BP

General Note: The constant temperature of process = Boiling Point of pure liquid

#Fractional Distillation – Seperating 2 liquids which are mixed (Miscible)

Process:

1. Mixture heated
2. Liquid with lowest BP comes out 1st, (at top of fractionating colomn)
3. Cooled by condenser, - Changes back to liquid
4. >>Liquid obtained is called "Distillate" (Collected in conical flask)

#Chromatography - For Seperating/ Identifying mixtures of Organic compounds.
- For Seperating mixtures of metal ions.

Personal Tips: Chemistry is actually very easy! I used to get D7 for my Chemistry, but with perseverance for 3 months, I scored A1 in my O-Levels 2008!

- If you can't do a question, use a red-pen to circle it, fold the page! Arrive school earlier, clear doubts with teachers outside the staffroom.
- You CAN ask ANY of the Science Teachers in your school!
- Be Brave to take the 1st step, and everything will be smooth after that.

"Take the first step in faith. You don't have to see the whole staircase, Just take the first step" – Martin Luther King

Chapter 3 – Summary

STRUCTURE OF ATOMS

3.1 – Particles in atoms

Atoms are made up of 3 particles:

1) Protons
   - Mass:1
   - Charge: +1
   - Location: Nucleus

2) Neutron
   - Mass: 1
   - Charge: 0
   - Location: Nucleus

3) Electron
   - Mass: 1
   - Charge: -1
   -Location: on the shells (around nucleus)

General Notes:
- If an atom is "electrically neutral", No. of proton = No. of neutrons
- No. of protons = No. of electron
- Nucleon number → sum of proton & neutron. (Also called "mass no.")

3.2 – Isotopes

Isotopes – Atoms of same element, but with diff no. of neutron.
(Same proton no. , diff neutron no.)

Note:

- All elements form Isotopes
- Isotopes have SAME chemical properties (becoz have same no. of outershell electrons)

*Note that the chemical properties are determined by the number of "Outer-shell" electrons.
- If 2 ~ have same no. of outershell electrons, they will have same chemical properties.

3.3 – Electrons arrangement in atoms

An atom with stable electronic structure will have 2 OR 8 valence electrons*
Eg:

- 2.8.8 → 8 electron in outermost shell (or 2 valence electron)
- 2.8 → 8 electron in outermost shell
- 2 → 2 electron in outermost shell

- "Valence" electrons - electrons on the outermost shell.

3.4 – Ions

Ions – Particles with a positive or negative electric charge.

If an atom does not has 2 or 8 valence electrons, it is **NOT STABLE!**

- **Not stable** means it **will react** by gaining or losing electrons, just to make sure they get either an '2' or '8' in their no. of valence electrons.
- So, when they gain or lose electrons, they will form "Ions"

Eg. An atom with electronic structure 2.8.8.7 will need to gain 1 electron to form 2.8.8.8 (STABLE)
- 2.8.3 will have to lose 3 electrons to form 2.8 (STABLE)

When an atom:

- Gain electron, it forms "NEGATIVE" ions
- Lose electron, it forms "POSITIVE" ions

General notes:

Negative Ion

- Formed when an atom gain electron
- Formed by non-metals

Positive Ion

- Formed when atom lose electron
- Formed by metals
  → Atoms form ions to obtain electronic structure of noble gas (Noble gas are STABLE)

*Important Fact: *"Metals lose electrons, non-metals gain electrons"*

Chapter 4 – Summary

CHEMICAL BONDS

4.1 – Covalent Bonding

Covalent Bonds – Formed between atoms of non-metals & non-metals

A "Double bond" means each atom provides 2 electron
- Total = 4 electrons being shared

4.2 – Ionic Bonding

Ionic Bonding – Formed by "transfer of electrons between metals & non-metals, forming positive & negative ions."

Ionic Bond is the force of attraction between +ve & -ve ions.

Eg. 2 atoms are ionically bonded; Sodium & Chloride.

Sodium: 2.8.1 (Needs to lose 1) → after losing 1 to Chloride, it becomes POSITIVE ion with 2.8
Chloride: 2.8.7 (Needs to gain 1) → after gaining 1 from Sodium, become NEGATIVE ion with 2.8.8

>>Then, they'll be called "Sodium Chloride" (A compound) - Becoz it's chemically joined together..

General Examples:

"Sodium Chloride" – Ionic Bond; becoz Sodium is a metal, Chlorine is non-metal.
"Magnesium Sulphate – Ionic Bond; becoz Magnesium is a metal, Sulphur is a non-metal

*Remember to name the METAL 1st then the NON-METAL!
Eg. We do not name "Chloride Sodium", We only name "Sodium Chloride"!

So why isn't it called "Sodium Chlorine"? → All elements in Group VII changes their tail name with "ide"

i.e. - Fluorine becomes "Fluoride" when it forms a compound (eg. Lithium Fluoride)
- Iodine becomes "Iodide" when it forms "Potassium Iodide" → Compound

"Carbon Dioxide" (CO_2) – Covalent Bond; becoz Carbon is a non-metal, Oxygen is also non-metal
"Methane"(CH_4)– Covalent Bond; becoz Methane consist of "Carbon" & "Hydrogen"
→ Both is non-metal

*Must remember that "Hydrogen" is considered a NON-METAL!

4.3 – Formular of Ionic Compounds

The formular can be explained in a story mode..

Finding formular of **"Barium Chloride"**, -- It is NOT as simple as *"BaCl"!!*

*Remember that Ionic Compounds → Transfer of electrons!
*"Metals lose electrons, non-metals gain electrons"
*Gain electron→ +ve ions / Lose electrons→ -ve ions

Read story mode:
(For weaker students)

- Barium is from group 2, hence it has 2 valence (outershell) electrons
- It is a metal → So it has to LOSE 2 electrons to become stable!
- So when it loses 2 electrons, it forms "positive" ions with the charge "+2"
- So Barium is Ba^{2+}

- Chlorine is from group 7, it has 7 valence electrons
- It is a non-metal → so It has to GAIN 1 electron to become stable!
- When it gains 1 electron, it forms "negative" ions with the charge "-1"

- Thus when you put 2 together, it will be "$Ba^{2+}Cl^{1-}$"
- Because the rule is that the +ve charge must balance the –ve charge,
- So to make is balance, just write is as "$BaCl_2$" //Answer

Personal Tips: Feel grateful for receiving this set of notes, and you will experience more great things coming..

The Shortcut method below will work everytime!
It's the best method to determine the formular of ionic compound!

Question: Find the formular of "Magnesium Sulphate" (MgS)

1. Write symbols of elements side by side (Metal den non-metal)

   Mg S

2. Write the ion formed when they become stable

   Mg^{2+} S^{1-}

   Mg becomes Mg^{2+} because it is from group 2, has 2 valence so it lose 2 to form "Mg^{2+}" ion.
   S becomes S^{1-} because it is from Group 7, has 1 valence, so has to gain 1 to form "S^{1-}" ion

3. Cross the charges (Ignoring the + and – signs)

   Mg^{2+} S^{1-}

   → Mg_1 S_2 → MgS_2 (Ans)

Practice makes perfect! Do not give up on your chemistry → Clear doubts as often as possible
If you are not interested in this subject, continue to clear doubt, and when you get good results, you'll love chemistry! - Hong Ray (Former Student)

4.4 – Properties of Covalent & Ionic Compounds

Note that is this topic has always been tested every year, and will continue to be tested in GCE O-Level exams!!

Just memorise the properties:

 # Covalent

1. Does not conduct electricity in any state (Becoz does not contain ions)
2. Low melting point "less than 200^{o}C" (Becoz weak forces of attraction between molecules)
3. Insoluble in water, Soluble in organic solvents!

 # Ionic Compound

1. Conduct electricity is molten, Aqueous state
   - Because it contain moving ions
   - Does not conduct on solid as ions cannot move

2. High Melting point "More than 1000^{o}C" → hence ionic compounds are generally solids at room temp.
   -Because of strong forces of attraction between ions

3. Soluble in water

*If a **question** (>5m) asks you to explain why Ionic compounds have high BP, Covalent has Low BP, answer in this format:*

1) Ionic compounds consist entirely of ions.
2) The opposite charged ions are held close to one another by very strong electrostatic attraction, known as ionic bonds.
3) Hence large amount of energy needed to break ionic bonds. → therefore high B.P

4) Covalent compounds consist entirely of molecules as they are formed by sharing of electrons.
5) Forces between molecules are very weak.
6) Only small amount of energy needed to break binds. → therefore low B.P

General notes:

- Electricity conductivity – linked to presence of moving ions
- Boiling point – linked to forces of attraction between molecules/ions.

Chapter 5 – Summary

METALS

5.1 – Physical Properties of Metals

#Metals

1. Good conductors of heat/ electricity
   - Contain electrons that are free to move through the metal.
   - For making cooking utensils/ wires

2. High Mp,Bp
   - Have strong bonds between atoms, → a lot of energy is needed to weaken & break bonds.

3. Malleable (Can be pressed into diff. shapes)
   - Layers of metals can slide over each other easily

#Alloys - Mixture of metal + other element

- Brass = Zinc + Copper
- Steel = Iron + Carbon

Note: Alloys are used because they are stronger & harder than pure metals.
- Atoms have different sizes, which prevents them from sliding over each other

Personal Tips: Always remember that if you experience doubts, NEVER HESITATE TO ASK!
Asking the teachers is far better than discussing with your classmates!

- If one teacher can't make you understand, try another teacher. You'll surely get one that suits you!

#5.2 – Reactivity Series

More Reactive

Potassium

Sodium

Calcium

Magnesium

Aluminium

Zinc

Iron

Lead

Hydrogen

Copper

Silver

Gold

Least Reactive

1) Potassium
   - Explosion with cold water
   - Explosion with Dilute hydrochloric acid

2) Sodium
   - Very fast reaction with cold water
   - Explosion with HCl

3) Calcium
   - Fast reaction with water
   - Fast reaction with HCl

4) Magnesium
   - Very slow reaction in cold water
   - Fast reaction with HCl

5) Zinc
   - No reaction with cold water, Burns in steam
   - Fast reaction in HCL (heated)

6) Lead
   - No reaction in water
   - Very slow reaction in HCl

7) Copper
   - N-R in water
   - N-R in HCl

8) Silver
   - N-R in water
   - N-R in HCl

#5.3 – Extraction of Metals from Ores

Ore – a rock which useful metals can be obtained.

- Metal is obtained from its ore by chemical reaction: "Extraction"
  - Extraction is done inside a furnace.

***Methods of extraction**

-The method used depends on how reactive is the metals!

Very Reactive Metals

Electrolysis
(Decomposing metal compound with electricity)

For:
- Potassium
- Sodium
- Calcium
- Magnesium
- Aluminium

Moderately Reactive Metals

Heating metal oxide with coke

For:
- Zinc
- Iron
- Lead

Low Reactivity Metals

Heating metal compounds in air

For:
- Copper
- Silver

Note: *Electrolysis is expensive, hence aluminum is expensive.*

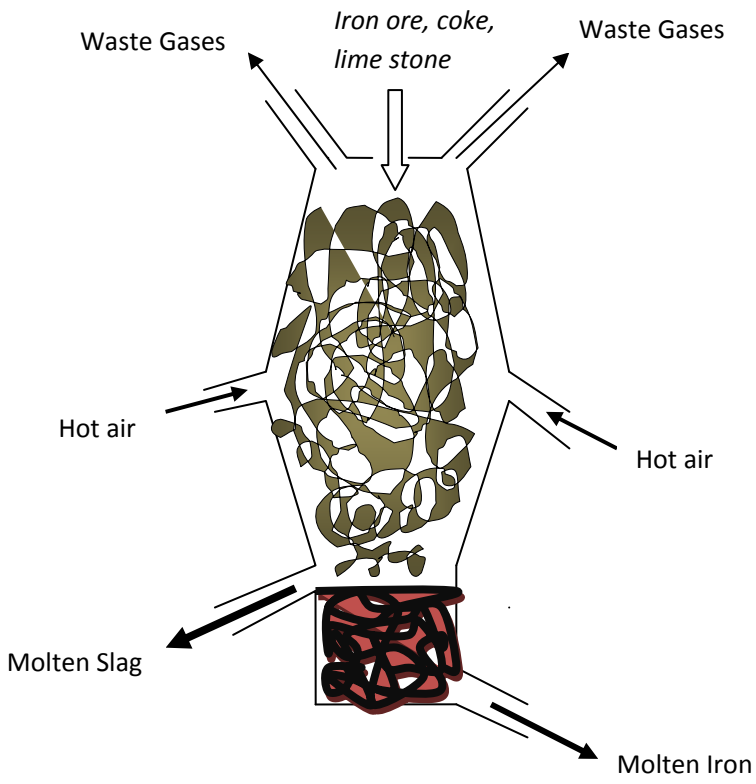
#Extraction of Iron

Iron - Extracted from ore "Haematite" (Impure iron (III) oxide) in the blast furnace

Blast furnace contains:

1. Iron ore
2. Coke
3. Lime stone

- Blasts of hot air (containing oxygen) are blown into the furnace near the bottom.



Chemical Reactions in Blast Furnace

1. Coke burns in air to produce carbon dioxide and lots of heat.
   $C + O_2 \rightarrow CO_2$
2. Carbon Dioxide reacts with more coke to produce Carbon Monoxide.
   $C + CO_2 \rightarrow 2CO$
3. Carbon Monoxide react with iron(III)Oxide to produce molten iron & carbon dioxide.
   $3CO + Fe_2O_3 \rightarrow 3CO_2 + 2Fe$
4. Impurities in molten iron are removed by limestone.
   $CaCO_3 \rightarrow CaO + CO_2$
5. Calcium Oxide formed combines with Silica present in the ore to form slag, which is tapped out.
   $CaO + SiO_2 \rightarrow CaSiO_3$

#5.4 – Recycling

- Recycling metals – "Collecting, melting down strap metals to make blocks of fresh metals" (For making new metal objects)

Advantages of Recycling:

1. Metal ores in the ground can last longer.
2. Money saved in energy needed to extract new metals from ore.
3. Scrap metal is removed from the environment, prevents land & water pollution due to corrosion.

Disadvantages of Recycling:

1. Expensive to collect strap metals from many sources
2. Metal fumes produced in melting of scrap metals can cause pollution.

#5.5 – Aluminium

1) Quite a reactive metal (Higher in reactivity series)

- Corrosion Resistant
- Thin layer of Aluminium Oxide prevents corrosion.
  (Prevents air, water from reaching the metal underneath)
- Uses:
  - Food/drink container → *Corrosion* Resistant.
  - Aircraft bodies → Low *density* so that aircraft body is light.
  - Overhead power cables → Good *electrical* conductor, less *dense* compared to copper (Cable can be lighter)

#5.6 – Iron & Steel

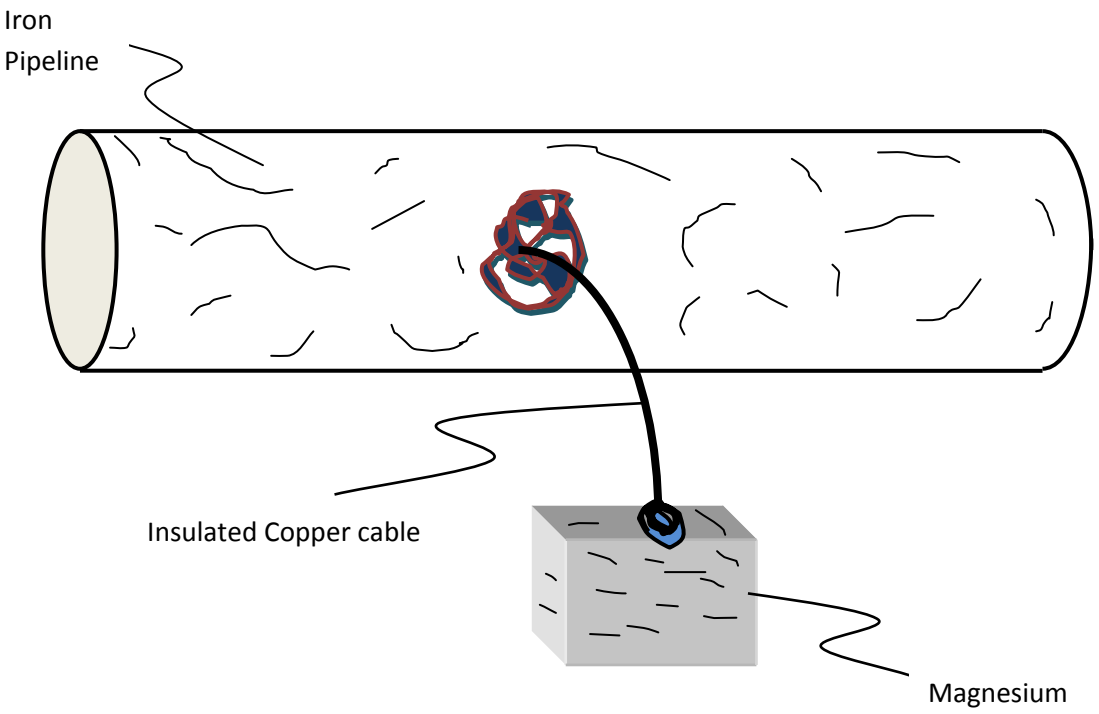
Rusting → Due to air **and** water reacting with iron.
- After reaction, iron becomes Iron(III) Oxide, FeO3

#Rusting Prevention

I. Barrier Method
   - Use of paint, grease, oil or another metal. eg. tin
II. Sacrificial Protection
   - Coating iron or steel with a more reactive metal.

Eg.

- Galvanising - Coating iron/steel with zinc. (Iron corrodes instead of zinc)
- Magnesium attached to iron pipelines to protect from rusting.



Personal Tips: To remember Chemistry Facts, write out in a piece of plain paper (NOT FOOLSCAP).. Just repeat writing them out..

Chapter 6 – Summary

THE PERIODIC TABLE

6.1 – The Structure of the Periodic Table

1) Elements arranged in order of proton numbers.

2) Verticles Colomns – **"Groups"**
   - All elements in Group 2 have 2 valence electrons
   - All elements in Group 3 have 3 valence electrons.. etc

3) Horizontal row – **"Period"**
   - Left → Right = Metals → Non-metals

   - All elements in Period 1 have 1 outershell of electrons
   (Eg. Hydrogen: 1. Helium: 2.)

   - All elements in Period 2 have 2 outershells of electrons eg
   (Eg. Lithium: 2.1 Oxygen: 2.6)

   - All elements in Period 3 have 3 outershell of ... ~
   (Eg. Magnesium: 2.8.2 Chlorine: 2.8.7)

*Note that Hydrogen is considered a ***non-metal*** (For answering O-Level examinations)

Elements in same group:

- Same chemical properties
- Same no. of valence electrons
- Diff physical properties
- Different empirical formula

Note that the chemical properties of an element is linked to the no. of valence electrons.
→ So becoz elements in same group have same no. of valence electrons, they will have same chemical properties.

#6.2 – Group 1 Elements

- Group 1 elements are called Alkali metals
- Because it's from group 1, it has 1 valence electron
- And because it's a metal, it can **lose** 1 electron and form a positive charged ion: "+1" (which is more stable)
  Eg. Potassium: 2.8.8.**1** Lithium: 2.**1** Sodium: 2.8.**1**
- Group 1 elements are very reactive!!! (**Bottom** *MORE REACTIVE* than **TOP**)
- They react with <u>cold water</u> to form <u>Alkaline solution</u> of <u>Metal hydroxide</u> & <u>Hydrogen gas.</u>
- Soft, Silvery metals with low density and melting point (**Bottom** *LOWER m.p* than **TOP**)!!

Example reactions:

1) Lithium + Water → Lithium Hydroxide + Hydrogen Gas

   $2Li + 2H_2O \rightarrow 2LiOH + H_2$

2) Sodium + Water → Sodium Hydroxide + Hydrogen Gas

   $2Na + 2H_2O \rightarrow 2NaOH + H_2$

3) Potassium + Water → Potassium Hydroxide + Hydrogen Gas

   $2K + H_2O \rightarrow KOH + H$

#6.3 – Group 7 Elements

- Group 7 elements are called Halogens
- Because it's from group 7, it has 7 valence electrons

- And because it's a **non-metal**, it can **gain** 1 electron and form a negative charged ion: "-1" (which is more stable)
  Eg. Fluorine: 2.**7** Chlorine: 2.8.**7**

- Group 1 elements are very reactive!!! (**Bottom** *LESS REACTIVE* than **TOP**)

- They form acidic solutions

- Coloured substances with low melting point (**Bottom** *Higher m.p* than **TOP**)!!
- They are diatomic molecules (Each molecule contains 2 atoms)
  Eg. Cl_2, Br_2, I_2, etc

- Colours of elements become darker down the group, Liquid → Solid down the group
  - Fluorine: Pale Yellow (Gas)
  - Chlorine: Yellowish Green (Gas)
  - Bromine: Reddish Brown (Liquid)
  - Iodine: Black (Solid)
  - Astatine: Black (Solid)
- More reactive elements displace less reactive elements
  ("Displace" means replace other less reactive elements)

Eg. When Chlorine reacts with Potassium Iodide,: (Displacement Reaction)

- Chlorine is more reactive than Iodide
- Which means Chlorine have the power to kick off Iodide, and replace its position!
- So after the reaction, it'll be left with Potassium Chlorine AND Iodide.

Chemical Equation:

$$Cl_2 + 2KI \rightarrow I_2 + 2KCl$$

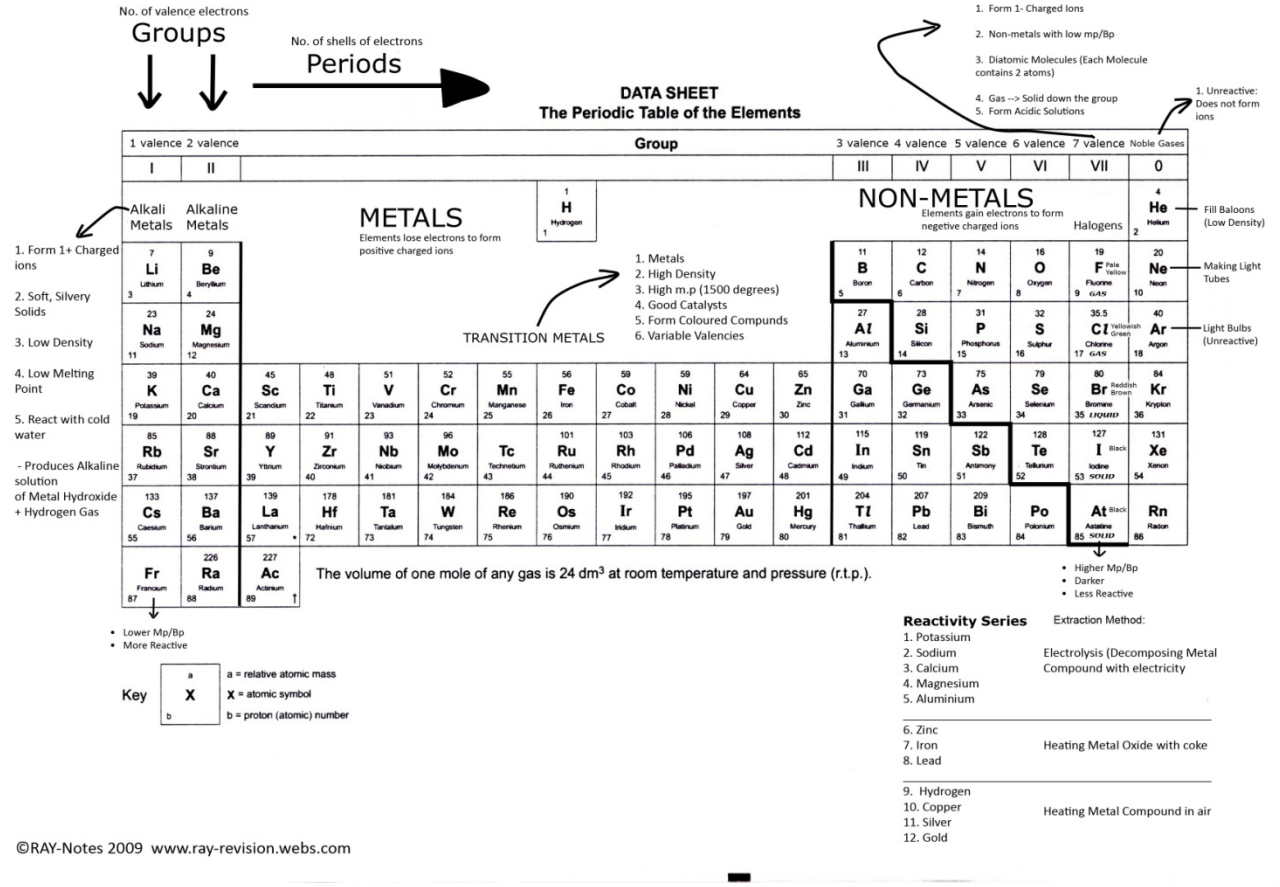
6.4 - Transition Metals (For students taking Pure Chem)

- Transition Elements are all metals, they do not belong to any group
- They have high density and melting point (1500^{o}C)
- Acts as good catalysts
- Form coloured compounds
- They have <u>variable</u> valencies (do not have fixed no. of valence electrons)

*Also note that Group 2 elements are called "Alkaline Metals"

*RAY-Resource 2009: Refer to the attached 'modified' periodic table to get a more detailed view of the entire chapter. This will help you remember the facts better!
Alternatively, this resource is available for free download at ray-revision.webs.com

Below: Preview of Periodic Table



Chapter 7 – Summary

THE ENVIRONMENT

7.1 – Composition of Air

- Air is a mixture of gases.
- 79% Nitrogen
- 20% Oxygen
- 1% Other gases (mainly argon)
- Also contain small amount of Carbon Dioxide & water vapour.

7.2 – Uses of Oxygen

1) Oxygen is used for combustion.
   "Combustion" means 'burning'
   Eg. Combustion of Carbon in blast furnace to produce CO_2

   Other info:
2) Oxygen tents helps people to breathe
3) Making steel by burning impurities to oxides then removed.
4) Oxyacetylene – Acetylene gas burn in O_2, high temp. melt steel

5) Limestone rock – a form of Calcium Carbonate

7.3 – Air Pollution

- Air contains large number of harmful substances (Pollutants):

1) Carbon Monoxide
   - Due to incomplete combustion of Carbon Fuels
   - Cause breathing difficulties/ headaches!!
   - Solution: Supply excess air to ensure complete combustion
     Eg. Fit vehicles with catalytic converters.

More FACTS:

- "Incomplete combustion" means lack of oxygen to burn.
- Carbon Monoxide is absorbed by haemoglobin in our blood,
- Prevents our blood from absorbing oxygen
- → Breathing difficulties, etc.

2) Methane
   - Due to Bacteria Decay of vegetation (Farm animals dung)
   - Cause Global warming
   - 'No easy solutions'

3) Oxides of Nitrogen
   - Lightning & Vehicle engines
   - Breathing difficulty/ Acid rain/ Produce ozone!!
   - Fit vehicles with Catalytic converters

4) Ozone
   - Sunlight acting on unburned hydrocarbons & Nitrogen dioxide
   - Irritates eyes & lungs!!
   - Reduce vehicle emissions of pollutants.

5) Sulphur dioxide
   - Combustion of fossil fuels like coal
   - Breathing difficulty/ Acid rain
   - Burn less sulphur-containing fuels

6) Unburned hydrocarbons
   - Vehicle engines
   - Produces Ozone
   - Fit vehicles with catalytic converters.

Acid Rain

- Sulphur Dioxide and Nitrogen Dioxide are main causes of acid rain.
- Kills fish in fresh water lakes
- Prevented by - Fitting motor vehicles with catalytic converters to reduce emissions of nitrogen oxides.

General Note: "Metals form basic oxides, Non-metals form acidic oxides."
Thus Both Sulphur & Nitrogen form Acidic oxides becoz they are both non-metals.

Chapter 8 – Summary

ACID & BASES

8.1 – Properties of Acids

Common Acids:

1. Hydrochloric Acid HCl
2. Nitric Acid HNO_3
3. Sulphuric Acid H_2SO_4
4. Ethanoic Acid CH_3COOH

- When an acid dissolve in water, it produces Hydrogen Ions, H^+.
- Acid turns blue litmus paper red
- Acid turns universal indicator red
- pH value less than 7

Rules: (MUST REMEMBER!)

1) ACID + METAL → SALT + HYDROGEN GAS

(Please note that acid do not react with any metals below hydrogen in the reactivity series!!!)

SO what I mean by 'salt'?
You just need to know that *Nitrate*, *Sulfate*, *Chloride, Carbonate* ARE Salts!
The name of salt produced simply depends on what acid is used.

So if you use Hydrochloric acid to react with a metal, it will form 'metal'-Chloride & Hydrogen Gas.

Eg. If you put "Hydrochloric Acid" and Sodium, It forms Sodium Chloride & Hydrogen Gas
Eg. If you put "Sulphuric Acid" and "Zinc", It forms Zinc Sulfate & Hydrogen Gas

2) ACID + BASE → SALT + WATER

*Must remember: A base is Metal **Oxide** or **Hydroxide**
(Oxide contains "O", Hydroxide contains "OH")

-Which means any metal that contains an "oxide" or "hydroxide" in it's name is a BASE!!

Eg. H_2SO_4 + CuO → $CuSO_4$ + H_2O
ACID **BASE** **SALT** **WATER**

Eg. HCl + NaOH → NaCl + H_2O
ACID **BASE** **SALT** **WATER**

3) ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE GAS

Note: Any thing that contains a "CO_3" in it's name is a carbonate.
i.e. $MgCO_3$, $CuCO_3$, etc are carbonates

Eg. $2HCl + CuCO_3 \rightarrow CuCl_2 + H_2O + CO_2$
ACID CARBONATE SALT WATER CARBON DIOXIDE

Eg. $H_2SO_4 + MgCO_3 \rightarrow MgSO_4 + H_2O + CO_2$
ACID CARBONATE SALT WATER CARBON DIOXIDE

*Extra notes: Before you balance the equation, always check if you have written the correct formulas.
Eg. $CuCl_2$ not CuCl. If you are unable to balance, it must be due to incorrect formulas.

- Approach teachers for help on balancing equations as this type of topic is better taught verbally.
Again, BE BRAVE! "Take the 1st step in faith.."

8.2 – Properties of Bases

Remember? Bases are Metal Oxides or Hydroxides, so anything that contains this 2 words are Bases!

There are 2 types of Bases; Soluble Base & Insoluble Base

- **Alkali are Soluble Bases**

- You only need to remember 4 alkalis:
  1. Potassium Hydroxide
  2. Sodium Hydroxide NaOH
  3. Calcium Hydroxide $Ca(OH)_2$
  4. Ammonia NH_3

Just remember the 4 soluble Bases by "PO, CA, SO, NH_4"

- When an Alkali (Soluble Base) dissolves in water, it produces Hydroxide Ions, OH^-.
- Alkali turns Red litmus paper Blue
- Alkali turns universal indicator blue
- pH value more than 7

*Extra Info: Farmers put alkali, Calcium Hydroxide, onto the fields to neutralize excess acids.

8.3 – The pH Scale

- The smaller the pH, the more acidic is the solution.
  (Means higher concentration of Hydrogen Ions, H^+)

- The Larger the pH, the more alkaline is the solution.
  (Means higher concentration of hydroxide ions, OH^-)

Eg. Something with pH value of 3 is more acidic than something with pH value 9.

pH values:

- 1 → 6 = Acidic
- 7 = Neutral
- 8 → 14 = Alkaline

The pH of a solution can be measure using:

1. Universal Indicator
2. pH meter

8.4 – Oxides

Oxides – Are "Compounds of oxygen with other elements"

3 Types of Oxides:

1. Acidic
2. Basic
3. Amphoteric

Remember? "Metals form basic oxides, Non-metals form acidic oxides."

Hence Acidic Oxides - Oxides of Non-metals
Basic Oxides – Oxides of metals

Eg. SO_3, CO_2, SO_2, NO_2 are Acidic Oxides becoz S, C, N are non-metals.

Eg. MgO, CuO, CaO are all Basic Oxides becoz Mg, Cu, Ca are metals.

Acidic Oxide + Water → Acid

Eg. $SO_3 + H_2O \rightarrow H_2SO_4$

Acid + Base → Salt + Water
- Remember that a base refers to metal oxides or hydroxides.

Thus Basic Oxide (Base) + Acid → Salt + Water

Eg. $CaO + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O$

Amphoteric oxides – "Metal Oxides that react with both acid & alkalis" to form salts

For O-Levels, you only need to
memorise the 3 Amphoteric Oxides:

1. Aluminium Oxide A_2O_3
2. Zinc Oxide ZnO
3. Lead Oxide PbO

Personal Tips: You must memorize the "Acid+Metal→Salt+Hydrogen Gas" etc.. A full list of equations summarized from all chapters is attached with this .zip package. To remember better, photocopy that list, paste it on the wall of your room or restroom - It works!

8.5 – Salt Preparation

Personal Tips: This might seem the hardest chapter to you now. But in truth, it interesting and simple to do!
- You MUST Clear Doubts with your teachers regularly, and you'll see miracles in your next test results.

First, You must memorise the 'secret' method to see if a salt is soluble.
(Solubility tables from your textbooks are harder to memorise)

	Soluble	**Insoluble**
Nitrate	All	-
Sulfate	Rest	Ba, Ca, Pb
Chloride	Rest	Ag, Pb
Carbonate	Group1	Rest

Remember via:
"Baked Chicken Pasta At Pastamania"

'Secret' Table contributed by Ms Aida (BtVSS, MOE Singapore)

This table helps you know whether a salt is soluble or not

How to use?

Facts of table:

- **All Nitrates are soluble**
- All Carbonates that are from group 1 in the periodic table are soluble; the rest are insoluble. *Eg. Potassium Carbonate is soluble, Barium Carbonate is insoluble*

- *All sulfates are soluble except Barium Sulfate, Cacium~, Lead~.*
  Eg. Zinc sulfate is soluble while Calcium sulfate is insoluble

- *All Chlorides are soluble except Silver Chloride, Lead Chloride..*
  Eg. Calcium Chloride is soluble, Lead Chloride is insoluble

So is Magnesium Sulfate soluble? → YES

Is Sodium Carbonate soluble? → YES

You still need to know 1 fact: All group 1 salts are SOLUBLE!

- Which means Lithium Carbonate, Potassium Chloride, etc are confirmed soluble since they're from group 1 of periodic table.

You can prepare (make) salts in many ways, depending on what type of salt it is.

**Important Notes:*

Titration Method (Soluble Base + Acid)

- **To prepare any salt that is from Group 1.** (Eg. Preparing Sodium Sulfate)
  1) Add 25cm of acid to a conical flask
  2) Add a few drops of indicator. (eg. Phenol-Phthalein)
  3) Add Sodium Hydroxide using a burette until indicator changes colour.
  4) Repeat experiment with $25cm^3$ of Sulphuric Acid but NO indicator!
  5) Add same volume of Sodium Hydroxide.
  6) Sodium Sulfate is obtained by evaporating & crystalising the salt.

Precipitation Method (Soluble + Soluble)

- **To prepare any Insoluble salt.** (Eg. Preparing Silver Chloride)
  1) Add Silver Nitrate with Sodium Chloride.
  2) Filter out the precipitate.
  3) Wash the residue with distilled water.
  4) Leave the residue to dry.

To use precipitation, just make sure the salt you want to make is insoluble.
The method requires you to use "Soluble salt + Soluble salt" so using the table, just find 2 salts that are soluble and contain part of the name.

For example, I want to make an insoluble salt called "Calcium Sulfate".

1st, I need to find "Calcium-~ " that is soluble. And I also know that "All nitrates are soluble"
→ So the 1st salt I use is **Calcium Nitrate**.

Next, I need to find "something-sulfate" that is soluble. I also know that all Group 1 salts are soluble..
→ So I can use **Lithium Sulfate** as my 2nd salt.

By reacting 2 salts, I'll get what I want, which is Calcium Sulphate!!

$$Ca(NO_3)_2 + Li_2SO_4 \rightarrow CaSO_4 + 2LiNO_3$$

When they react, they simply just change partners..

To prepare Soluble salt, but NOT from Group 1:

- There are 2 ways:
  1) Soluble salt + soluble salt
  2) Soluble Salt + Insoluble Salt

Eg. To prepare Magnesium Sulfate:

I use Magnesium Nitrate & Sodium Sulfate (Both are soluble)

$Mg(NO_3)_2 + Na_2SO_4 \rightarrow MgSO_4 + 2NaNO_3$

Alternatively, I can also use "Soluble + Insoluble" way.
I use Magnesium Nitrate & Barium Sulfate (1 is soluble, other is not)

$Mg(NO_3)_2 + BaSO_4 \rightarrow MgSO_4 + Ba(NO_3)_2$

	Soluble	Insoluble
Nitrate	All	-
Sulfate	Rest	Ba, Ca, Pb
Chloride	Rest	Ag, Pb
Carbonate	Group1	Rest

Remembering the table

3rd row: *"Baked Chicken Pasta At Pastamania"*

1st row: *"No Super Childish Children"*

Chapter 9 – Summary

Chemical Analysis

Chemical Analysis is about finding the name of an unknown salt by carrying out several experiments/tests..

All salts contains 2 parts: Cation (+ve charges ions) & Anion (-ve charged ions)

The question usually shows the observations of tests, and you will have to find out what is the salt according to the observation.

Testing for CATIONS:

With this table, you don't need to memorise the colours of different salts!

Just draw out this table during your O-Levels

CATION	NaOH		NH_3	
	Colour	Soluble or Insoluble	Colour	Soluble or Insoluble
Cu^{2+}	B	I	B	S
Fe^{2+}	G	I	G	I
Fe^{3+}	R	I	R	I
Ca^{2+}	W	I	-	-
NH^{4+}	N	-	-	-
Zn^{2+}	W	S	W	S
Al^{3+}	W	S	W	I
Pb^{2+}	W	S	W	I

Colour of Observation

Soluble or Insoluble in excess Alkali

The CATIONS table shows the colour changes observed when the salt is added to Sodium Hydroxide, NaOH, or Aqueous Ammonia, NH_3.

It also shows whether the salt is soluble in excess alkali (NaOH, NH_3)

- For example, I carry out an experiment to find out what an unknown salt contain.
- When I add NaOH to the salt, colour changes to white, and when I add excess of NaOH, it dissolved (Soluble).
- Next, I tried adding NH_3 to the salt. The colour also changes to white, and it is also soluble in excess NH_3.
- Then, I match the test results with the table. The unknown salt contains Zinc, Zn^{2+}

Testing for ANIONS:

ANION	Tested with	Observations	
Chlorine	$AgNO_3$	AgCl	White
Carbonate	Acid	CO_2	Carbon Dioxide (Turn Limewater chalky)
Sulphate	$Ba(NO_3)$	$BaSO_4$	White
Nitrate	Al, NaOH	NH_3	Ammonia (Turn red litmus blue)
Iodide	$Pb(NO_3)_2$	PbI_2	Yellow

Eg. I tested salt x with Silver Nitrate $AgNO_3$, a white precipitate is formed. This means that salt x contains Chlorine.. (The white precipitate is "AgCl")

Test for GASES:

- **Ammonia** → Use damp litmus Paper → Turns from red to blue
- **Carbon Dioxide** → Use Limewater → White ppt formed (Turns Chalky)
- **Chlorine** → Use Damp blue litmus paper → Bleaches Litmus paper (Litmus Turns colourless)
- **Hydrogen** → Use Burning Splint → 'Pop' Sound Heard
- **Oxygen** → Use Glowing Splint → Glows Brighter or Burst into flames
- **Sulphur Dioxide** → Place a drop of Potassium Dichromate(VI) on Filter paper
  - Orange Dichromate turns green.

Important Tips: You MUST memorise the 3 powerful tables above! They're going to help you memorise things so easily. You won't be able to find any easy stuffs like that in your textbook! So it's your choice to use it or leave it. Whatever you choose is right.

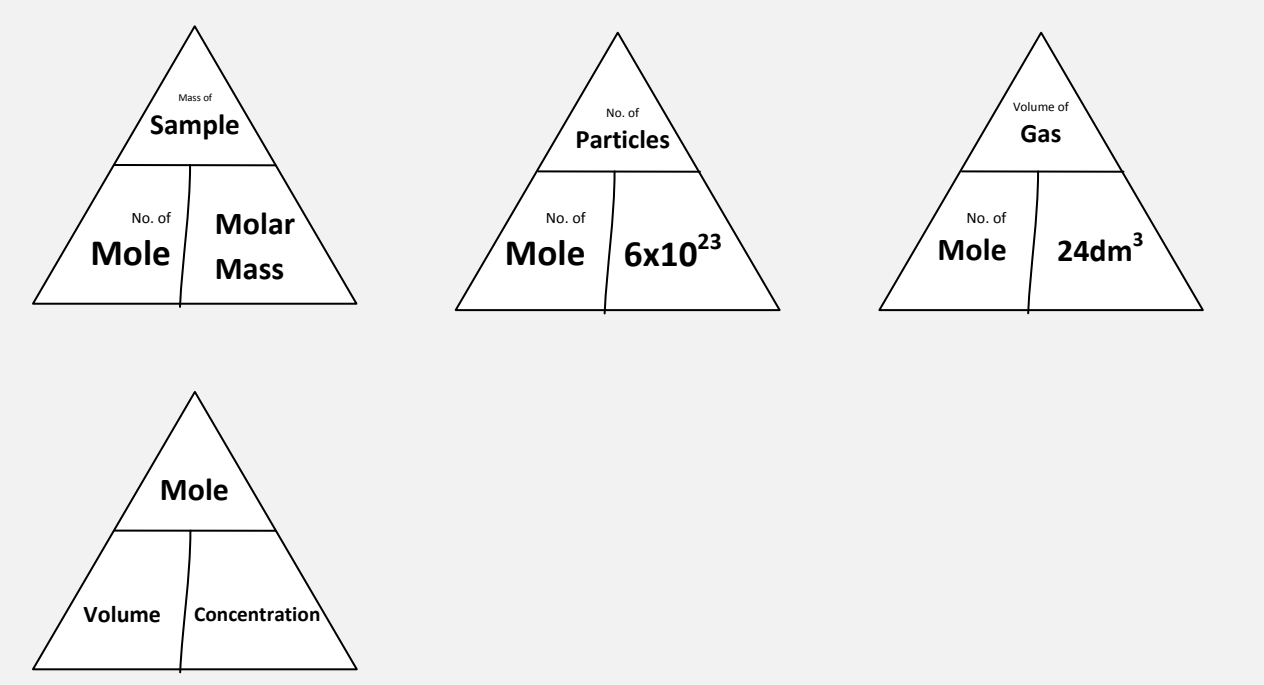
Chapter 10 – Summary

MOLE CONCEPT

10.1 – Counting Atoms: The Mole

*This chapter is better to be taught verbally by your teachers as it's not easy to explain in words. Thus, approach teachers when in any doubt!

Memorise this 4 IMPORTANT Formulas:



Some facts:

1. 1 Mole = 6×10^{23} particles
2. Molar Mass – Mass of 1 mole of any substance
3. Relative Atomic mass = Molar mass, the diff. is that Molar mass has a "grams" on it.

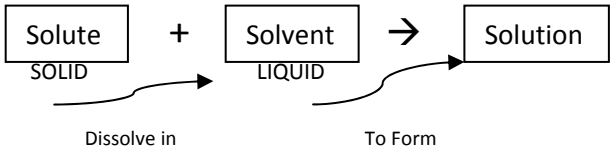
$$^{32}S_{16}$$

4. Eg. Sulfur: S_{16} has a molar mass of 32g. Which also means it contain 32g per mole.

#10.2 – Molar Volume of Gases

- 1 Mole of any gas, has the same volume of $24dm^3$ at room conditions (25°C, 1 Atmosphere)
- Eg. Ammonia NH_3 → Molar Mass = 17g
  Carbon Dioxide CO_2 → Molar Mass = 44g
  } 24dm
- $1dm^3 = 1000cm^3$

10.3 – Concentration of solutions



- Concentration is measured in mol/dm^3

#10.4 – Molecular Formula

- The Molecular formula shows the actual number of each type of atoms present in a compound

To find the molecular formula of a compound, you need:

1. Empirical Formula of compound
2. Relative molecular mass

For example:

The relative molecular mass of a compound is 62
The empirical formula of the compound is (COH_3)

1) **Calculate: (COH_3) = 12+16+3 = 31**
2) **Write the statement: 62 = n(31)**
3) **Find n: $n = {}^{62}/_{31} = 2$**
4) **n=2**
5) **Add 2 into the formula: $(COH_3)_2 = C_2O_2H_6$ (ans.)**

#10.5 – Empirical Formula

- **Empirical Formula** – the simplest formula of a compound which shows the ratio between the atoms of each element.

Finding the empirical formula is about finding the no. of moles of both items, den devide their values with the smaller value.

Eg. A substance contain 80% Copper, 20% Sulfur
Find the empirical formula.

1) **Find the no. of mole of both items!**

no. of mole of Copper: $^{80}/_{64} = 1.25mol$
no. of mole of Sulfur: $^{20}/_{32} = 0.625mol$

2) ***You compare 2 values; 0.625 is a smaller value right?***

3) ***So, devide both values with 0.625:***
$^{1.25}/_{0.625} = 2$
$^{0.625}/_{0.625} = 1$

4) ***Den, you put the numbers in this form:*** Cu_2S (ans)

Chapter 12 – Summary

OXIDATION & REDUCTION

Must memorise:

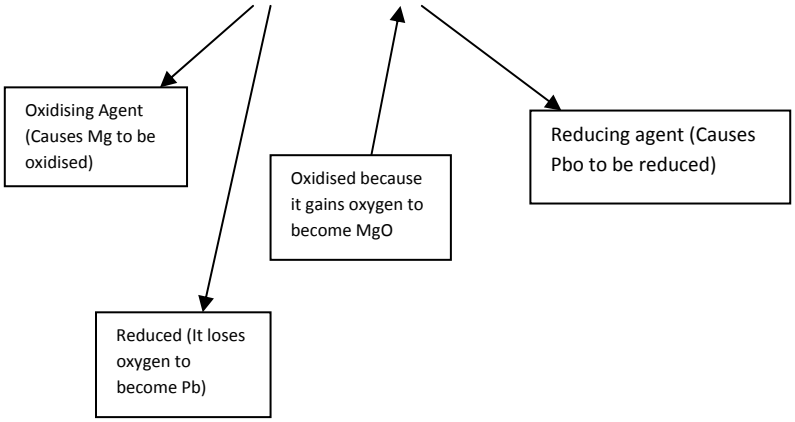
Oxidation

1. Gain Oxygen
2. Loss Hydrogen
3. Loss electrons
4. Increase in oxidation state

Reduction

1. Gain hydrogen
2. Gain electron
3. Loss Oxygen
4. Decrease in oxidation state

Eg. $PbO + Mg \rightarrow Pb + MgO$



General notes:

- A substance that causes something to be oxidised is an oxidising agent.
- A substance that causes something to be reduced is an reducing agent.

- An oxidising agent itself is being reduced in the reaction
- An reducing agent itself is being oxidised in the reaction

*Redox reaction – A reaction in which Oxidation and reduction takes place.

Loss / Gain of electrons

Eg. $2Na^0 + Cl_2^0 \rightarrow 2Na^+Cl^-$

Sodium loses electron → it is oxidised

Chlorine Gains electron → Reduced

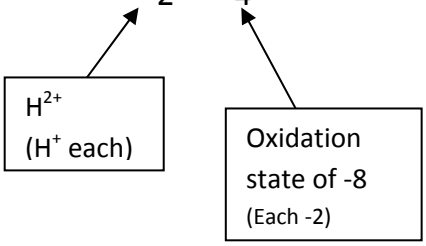
Oxidation State

*All elements only have charges when they form compounds.

Eg. Mg → NO CHARGE
Mg^+O^- → Got charge

Oxidation state is the <u>charge on an ion</u> → depends on which group of the periodic table the elements are in.

Eg. What is the oxidation state of Sulfur in H_2SO_4?



Note: The number of positive charges must balance no. of negative charges.

Hence Oxidation state of Sulfur is 8 – 2 = +6

Personal Tips: You must ASK YOUR TEACHERS if you don't understand any question or topic! Do Not Delay or you might regret soon… If you still don't understand after his/her explanations, <u>DO NOT pretend</u> to have understood! Just ask him/her to explain again.

Oxidation & Reduction as changes in Oxidation state

Remember:

- Increase in Oxidation state – Oxidation
  Eg. $Mg \rightarrow Mg^{2+}$ (Oxidation state gains from 0 to +2) - Mg is oxidised.

- Decrease in Oxidation state – Reduction
  Eg. $Cl \rightarrow Cl^{-}$ (Oxidation state decreases from 0 to -1) - Cl is reduced.

Chapter 13 – Summary

RATE OF REACTIONS

- **Slow reactions:**
  - Rotting a piece of wood in the ground
  - Rusting of a steel fence

- **Fast reactions:**
  - Dynamite exploding
  - Burning a piece of magnesium in air

  <PIC OF GRAPH>

Notes:

1. Reactions takes place when particles collide.
2. Most collisions do not produce a reaction because colliding particles need a minimum energy to react when they collide.
   This is called "Activation Energy"

Factors affecting speed of reaction

(Particles Theory)

1) **Concentration of solution**

   - The higher the concentration of solution, the faster the rate of reaction.
   - Because the particles are closer together so they collide more frequently and so there were more frequent reactions.

   Note: Only applies to reactions of solutions.

2) **Pressure**

   - The higher the pressure, th higher is the rate of reaction.
   - Because the particles are squeezed closer together, so they collide more frequently and so there were more frequent reactions.

   Note: Only applies to reactions of gases.

3) **Particle size of solid**

   - The smaller the particle size, the higher the rate of reaction.
   - Because the total surface area of the solid increases so reacting particles of liquid and gases collide more frequently with the surface and so there are more frequent reactions.

   Note: Only applies to reactions of solids.

4) **Temperature**

- The higher the temperature, the faster the rate of reaction.
- Because at higher temperature, the particles have greater kinetic energy, so they react more often when they collide.

Chapter 14 – Heat changes in reactions

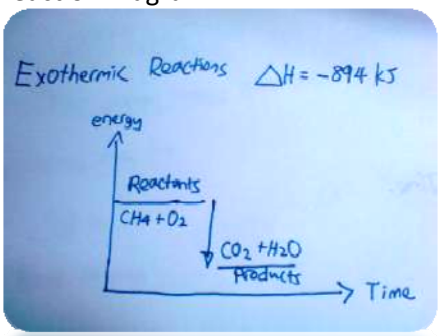
1) Exothermic -ve

- Energy/ Heat Released
- Solution/Testtube Becomes hot
- Bond Forming
- ▲H= -ve (Negative Value)
- Eg. All combustion reaction like burning Magnesium in air/ Reaction of acid with alkalis

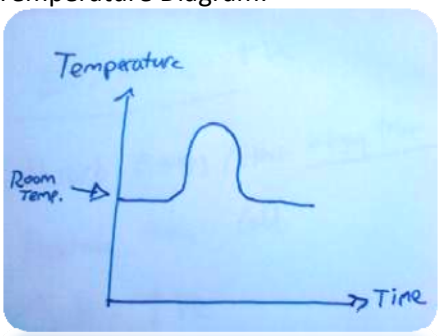
Why solution turns hot when heat is given out?

- Heat is released from the solution to your hands so you feel the test tube is hot.

Reaction Diagram:



Temperature Diagram:



2) Endothermic +ve

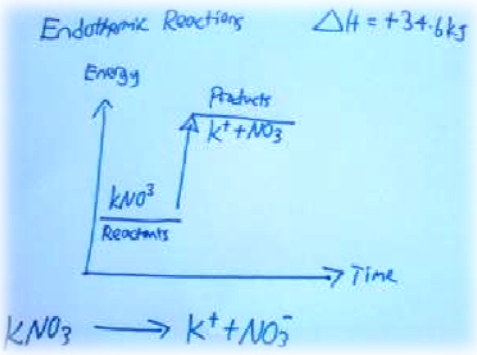
- Energy/ Heat Absorbed
- Solution turns cold
- Bond Breaking
  Eg. Freezing of ice

- ▲H= +ve (Positive Value)

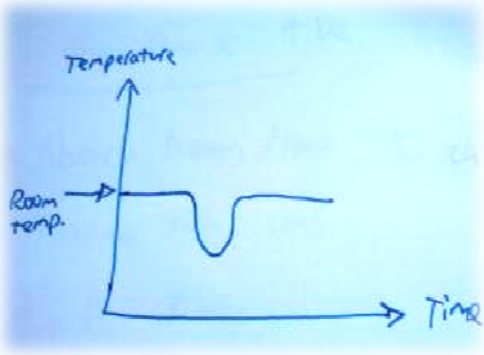
Why solution turns cold when heat is taken in?

- The heat in your hands is absorbed so you'll feel test tube becoming cold.

Reaction Diagram:



Temperature Diagram:



*Most reactions are exothermic as they make the test-tube turn hot.

Info:
Energy is taken in to break bonds.
Energy is given out for bond forming.
▲H refers to Delta H
▲H is the amount of heat energy taken in/ given out during a chemical reaction.

Chapter 15 – Summary

FUELS

15.1 – Fossil Fuels

1) Petroleum & Natural Gas are the 2 main fossil fuels.
   - Fossil fuels were produced from plant and animal remains (Long ago)

2) Petroleum - A sticky black liquid and a mixture of hydrocarbons.
3) Natural Gas consist mainly of methane

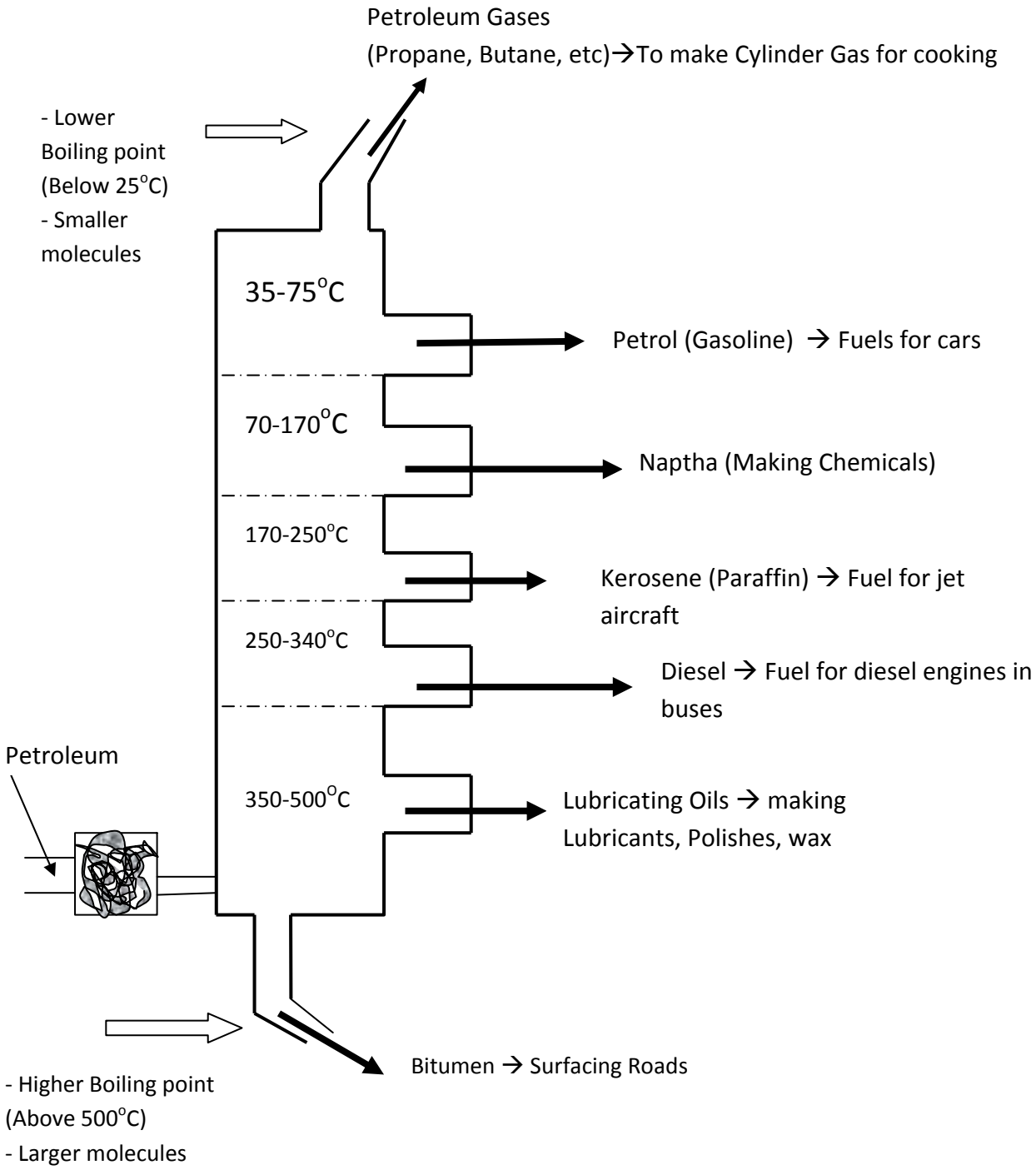
15.2 – Fractional Distillation of Petroleum

- Petroleum is separated to different fractions by fractional distillation.
- Seperation can take place because petroleum is a mixture of substances with **different** boiling points.

Important Process:

1. The petroleum is heated in a furnace to vapourise it.
2. The vapour condenses to liquid at different heights up the fractionating colomn, where it comes out as different fractions.
3. A fraction is mixture of hydrocarbon with a range of boiling points.
   - The hydrocarbons are alkanes.

15.2 (b) – Fractionating Colomn



You must remember the order of fractions and their uses!

1) *Petroleum*
2) *Petrol*
3) *Naptha*
4) *Kerosene (Paraffin)*
5) *Diesel*
6) *Lubricating Oil*
7) *Bitumen*

Combustion of Fossil Fuels

1) Hydrocarbon burns in air → produce water and carbon dioxide (Complete combustion)

$CH_4 + CO_2 \rightarrow CO_2 + 2H_2O$

Alkane (Hydrocarbon)

Air (Oxygen)

Mainly carbon

2) *Carbon monoxide and soot will also be produced – if not enough air (Incomplete combustion)

#15.2(c)

- Petrol and diesel are needed as fuels for vehicles
- Petroleum does not have enough of them.
- Hence **Cracking** is done

A reaction in which big hydrocarbon molecules are broken down into smaller molecules by heat.

Chapter 16 – Summary

HYDROCARBONS

16.1 – Alkanes

1) Alkanes are **saturated** hydrocarbons (No more atoms can be added to their structure)
2) General formula: C_nH_{2n+2}

Smaller Molecules (Less Carbon Atoms)
Lower B.p

1. Methane CH_4
2. Ethane C_2H_6
3. Propane C_3H_8
4. Butane C_4H_{10}

GAS

5. Pentane C_5H_{12}
6. Hexane C_6H_{14}

Liquid

Larger Molecules
Higher B.p (More Carbon atoms)

- Each carbon atom forms 4 bonds.
- Boiling point increases as molecules get larger

Why higher b.p?
- Intermolecular forces of attraction between molecules increases as molecules become larger.

Homologous series of Alkanes

Homologous series - A set of organic compounds in which the formula of each one differs from the previous one by an extra –CH2- group of atoms.

Members in a Homologous Series have:

- Same chemical reactions
- Same functional group (Eg. –OH, -COOH)
- Same general formula
- Different Physical Properties!

Reactions of Alkanes

Alkanes react with Chlorine in substitution reaction (Slow reaction, requires light)

Note: Alkanes can **only** undergo combustion reactions and Substitution reaction with chlorine!
Otherwise, Alkanes is unreactive.

16.2 – Alkenes

1) Alkenes are unsaturated hydrocarbons (They can react with many substances in addition reactions)
2) General forular: C_nH_{2n}
3) All alkenes contain C=C double bonds

Smaller Molecules (Less Carbon Atoms)
Lower B.p

7. Methene CH_2
8. Ethene C_2H_4
9. Propene C_3H_6
10. Butene C_4H_8
11. Pentene C_5H_{10}
12. Hexene C_6H_{12}

} GAS

Homologous series of Alkenes

Larger Molecules
Higher B.p (More Carbon atoms)

Reactions of Alkenes

1) Alkenes burn in air (Combustion) → Produce Carbon Dioxide & Water

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

- If incomplete combustion, Carbon (Soot) and carbon monoxide will be produced!

*Addition reaction

- C=C double-bonds are broken, extra atoms are added

2) Addition of Bromine

$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

- This reaction is used to test for Alkenes
  (Orange bromine solution decolourises if Alkene is present)

Info: Bromine is from Group 7, so it has colour. Note that the colour can only be present if Bromine remains single(Not reacted).
- If it forms a compound, the colour will disappear. Hence after the reaction, orange bromine turns colourless.

3) **Addition of Water**

Alkene + Steam → Alcohol (Require Phosporic Acid as catalyst, high temperature & pressure)

4) **Addition of Hydrogen** (Nickel Catalyst is needed)

Alkene + Hydrogen → Alkane

- This reaction is used to change vegetable oil to magarine.

16.3 – Cracking of Alkanes

- Cracking – A reaction in which big hydrogen molecules are broken into smaller molecules by heat.

- Done by pressing big alkane molecules over a solid catalyst at high temperature
- Products: - 1 small alkane molecule
  - 1 alkene molecule

*Cracking of alkanes is used to produce:

1) Alkanes
2) More petrol for vehicles
3) Hydrogen

Chapter 17 – Summary

ALCOHOLS AND ORGANIC ACIDS

17.1 – Alcohols

1. Alcohols are organic compounds containing –OH group of atoms.
2. General Formula: $C_nH_{2n+1}OH$

Smaller Molecules
Lower B.p

1. Methanol CH_3OH
2. Ethanol C_2H_5OH
3. Propanol C_3H_7OH
4. Butanol C_4H_9OH

Homologous series of Alcohol

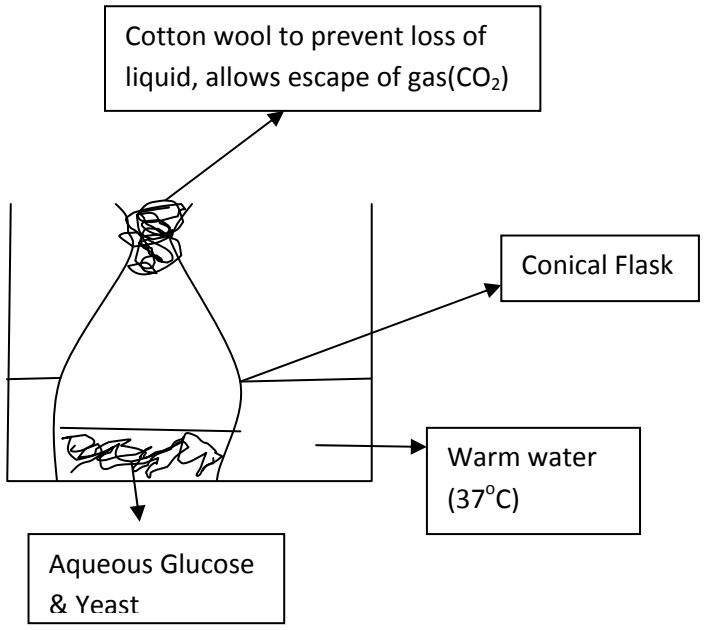
All alcohols are liquids!

Larger Molecules
Higher B.p

Methods of Making Ethanol:

Method 1:

Ethene + Steam → Ethanol
(Alkene) (Alcohol)



Method 2

Fermentation of glucose with yeast.

- The enzymes in yeast change glucose into <u>ethanol</u> and Carbon Dioxide
- Reaction is best at 37^oC
  (If higher temp, enzyme structure will be damaged, no longer acts as catalyst!)
- Products of fermentation:
  Dilute solution of ethanol – Pure ethanol is obtained by fractional distillation.

Conditions for Fermentation:
- 37^oC
- Enzymes in yeast
- No Oxygen present

#Uses of Ethanol:

1. Solvents (Eg. Perfumes)
2. Alcoholic drinks (Beer/wine)
3. Fuel (Petrol for cars)

Reactions of Alcohol

1) Alcohol burn in air → Carbon Dioxide + Water
   to Produce
   Eg. Ethanol + Oxygen → Carbon Dioxide + Water

2) Alcohols are **oxidised** to organic acids
   Eg. Ethanol + Oxygen → Ethanoic Acid + Water
   $C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$

   (Lose 2 hydrogen, Gain 1 Oxygen= Oxidation!)
   Bacteria in air acts as catalyst to the reaction.

17.2 – Carboxylic Acids

1. Carboxylic Acids – Organic compounds containing – CO_2H group of atoms.
2. General Formula: $C_nH_{2n+1}CO_2H$

   *Tip: You should use this alternate working formula: **$C_nH_{2n}O_2$**

   (The original formula $C_nH_{2n+1}CO_2H$ makes it very confusing for most students! The new formula is proven to work on all O-level questions! No Probs!)

Smaller Molecules
Lower B.p

13. Methanoic Acid $C_1H_2O_2$
14. Ethanoic Acid $C_2H_4O_2$
15. Propanoic Acid $C_3H_6O_2$
16. Butanoic Acid $C_4H_8O_2$

Liquid

Larger Molecules
Higher B.p

Formulas generated using this new formula:

$C_nH_{2n}O_2$

Homologous series of Organic Acids

Reactions of Carboxylic Acids

Acid + Alcohol → Ester + Water

- Boil the mixture
- A little concentrated Sulphuric Acid acts as catalyst.

Eg. Ethanoic Acid + Ethanol → "Ethyl Ethanoate". (Ester)

Reaction of making Esters is called "Esterification"

Esters have sweet smell

#**Uses** of Esters:

1. Solvents
2. Flavouring in food

Chapter 18 – Summary

MACROMOLECULES

(Topic is skipped as it is no longer tested in Science (Chemistry) new 2008 syllabus)

If you're taking pure science and this topic is in your syllabus, send me an email to request for summarized notes on this topic. The updated notes will then be made available in the next version of this notes series.

The Periodic Table and Full Chemistry equations is attached with this .zip package.
It is recommended that you photocopy extras and paste them onto walls of your room or even washrooms! It'll help you to remember!

Found mistakes on notes? Send me an email. Corrections will be made within 2 days.

*Hard-copy notes for other subjects are also available at http://www.ray-revision.webs.com