

PUR ENDERS



The Periodic Table of Elements

Lu Iutetium 175 103 Lr awrencium ا من المراجع الم مراجع المراجع الم مراجع المراجع الم مراجع المراجع مراجع المراجع مل مراجع المراجع المراجع المراجع المراجع المراجع المراجع المراجع المراجع المراح المراجع المراجع المراجع المراجع المراحع المر Yb 173 102 No nobelium ⋝ thulium 169 101 Md Er 167 100 Fm fermium > C 12 12 14 3ilicon 28 28 28 32 66 67 73 73 73 73 119 119 114 114 114 114 114 114 Ho holmium 165 99 ES einsteinium \geq Boron 11 11 11 27 27 27 27 20 49 11 115 81 81 204 204 Dy dysprosium 163 98 Cf Cf Tb 159 97 Bk erkelium Gd 157 96 Cm eurium Eu 152 95 Am americium Group Sm samarium 150 94 Pu Plutonium Co 59 athrodium 103 103 103 103 103 103 103 103 109 Mt aeitnerium hydrogen Pm romethium 93 Pp neptunium iron 56 56 44 44 44 101 101 101 100 8 100 100 100 100 Hs Hs Hs Mn manganese 55 43 43 43 7 7 7 7 7 7 8 8 107 107 107 107 107 Nd 144 92 92 0 238 238 Pr 141 91 Pa protactinium 231 proton (atomic) number atomic symbol name relative atomic mass anadium 51 105 105 105 105 105 Db bubnium Ce 140 90 90 Th Thorium 232 Key La anthanum 139 89 89 actinium Sc 45 39 7 yttrium 89 81 anthanoids 89 – 103 actinoids anthanoids

The volume of one mole of any gas is 24 dm³ at room temperature and pressure (r.t.p.).

actinoids



Pure Chemistry Notes



3EXP CHEMISTRY TOPICS

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Particles	Location	Charge	Mass	Remarks
Proton	Nucleus	+1	1 unit	Protons and neutrons in
Neutron	Nucleus	No charge	1 unit	the nucleus of an atom
Electron	Electron shell	-1	1/1840	are collectively known as
				NUCLEONS

ISOTOPES

Definition: Isotopes are <u>atoms</u> of the <u>same element</u> with the same number of protons and electrons but <u>different number of neutrons</u>.

Hence isotopes have the same:

- (a) Number of protons and electrons
- (b) Chemical properties due to similar electronic configuration

But they have different:

- (a) Number of neutrons
- (b) Physical properties like mass, density

Common Isotopes are:

 $^{17}_{37}Cl$ $^{17}_{35}Cl$

25% of isotopes of Chlorine are 37-Cl and 75% of isotopes of Chlorine are 35-Cl

Hence the average mass of Chlorine is calculated as shown:

Relative atomic mass of Chlorine = 25% x 37 + 75% x 35 = 35.5

Hence Chlorine in the Periodic Table is

¹⁷_{35.5}Cl

Other Isotopes are:



 ${}^{1}_{2}H$

deuterium



tritium



Atom	lon	Ion Formula	Remarks		
Chlorine	Chloride	Cl-	Negative ions formed		
Fluorine	Flouride	F-	ends with ~ide		
Bromine	Bromide	Br⁻	Positive ions formed do		
lodine	lodide	l-	not change their name		
Oxygen	Oxide	O ²⁻	-		
Sulfur	Sulfide	\$ ²⁻			
Nitrogen	Nitride	N ³⁻	-		
Hydrogen	Hydride	H-			
Phosphorus	Phosphide	P ³⁻			

IONS FORMULA LIST

COMPLEX IONS FORMULA

Complex Ion	Formula
Hydroxide	OH-
Sulfate	SO4 ²⁻
Carbonate	CO32-
Nitrate	NO₃ ⁻
Ammonium	NH4 ⁺

COMPLEX IONS FORMULA (ADDITIONAL LIST)

Complex Ion	Formula
Sulfite	SO ₃ 2-
Hydrogen carbonate	HCO3⁻
Hydrogen sulfate	HSO₄⁻
Hydrogen sulphite	HSO₃-
Phosphate	PO4 ³⁻
Thiosulfate	\$2O3 ²⁻

Sodium hydroxide - NaOH	(1:1)
Copper (II) hydroxide - Cu(OH)2	(1:2)
Aluminium nitrate - Al(NO3)3	(1:3)
lron (III) carbonate – Fe2(CO3)3	(2:3)
Copper (II) sulfate – CuSO4	(1:1)
Ammonium sulfate – (NH4)2SO4	(2:1)

Sodium phosphide - Na₃PO₄ ^(3:1) Lithium hydrocarbonate - NaHCO₃ ^(1:1) Sodium hydrocarbonate - NaHCO₃ ^(1:1) Ammonium sulfate - (NH₄)₂SO₄ ^(2:1)



POSITIVE IONS FORMULA (ADDITIONAL LIST)

lon	Formula
Silver	Ag+
Zinc	Zn ²⁺
Copper (I)	C∪+
Copper (II)	Cu ²⁺
Iron (II)	Fe ²⁺
Iron (III)	Fe ³⁺
Lead (II)	Pb ²⁺
Lead (IV)	Pb ⁴⁺

For transitiion metals with 2 oxidation state eg Copper which can be Cu⁺ and Cu²⁺

The charge of the metal must be specified.

CuSO₄ – Copper (II) sulfate

Fe(OH)₃ – Iron (III) hydroxide

IMPORTANT FORMULA

Acid	Formula	
Hydrochloric acid	HCI (aq)	
Nitric acid	HNO ₃	
Sulfuric acid	H ₂ SO ₄	
Sulfurous acid	H ₂ SO ₃	
Phosphoric acid	H ₃ PO ₄	
Carbonic acid	H ₂ CO ₃	
Gas	Formula	
Hydrogen chloride gas	HCI (g)	
Ammonia	NH ₃	
Methane	CH4	
Other chemical	Formula	
Hydrogen peroxide	H ₂ O ₂	
Limestone	CaCO3	
Limewater	Ca(OH) ₂	
Quicklime	CaO	



CHEMICAL FORMULA

Simple lons – There is no bracket used.

Name of	Positive	Negative	Formula	Ratio
compound	ion	ion		Cation : Anion
				in compound
Sodium oxide	Na+	O ²⁻	Na ₂ O	2:1
Lithium oxide	Li+	O ²⁻	Li ₂ O	2:1
Calcium oxide	Ca ²⁺	O ²⁻	CaO	1:1
Aluminium oxide	Al ³⁺	O ²⁻	Al ₂ O ₃	2:3
Sodium nitride	Na+	N ³⁻	Na ₃ N	3:1
Aluminium nitrlde	Al ³⁺	N ³⁻	AIN	1:1

Transition metals with variable oxidation states – The roman numeral in the bracket will give the cation its charge.

Name of	Positive	Negative	Formula	Ratio
compound	ion	ion		Cation : Anion in compound
Copper (I) oxide	Cu+	O ²⁻	Cu ₂ O	2:1
Copper (II) oxide	CU ²⁺	O ²⁻	CuO	1:1
Iron (II) chloride	Fe ²⁺	Cŀ	FeCl ₂	1:2
Iron (III) chloride	Fe ³⁺	Cl-	FeCl ₃	1:3
Lead (II) bromide	Pb ²⁺	Br⁻	PbBr ₂	1:2
Lead (IV) bromide	Pb4+	Br⁻	PbBr₄	1:4



Complex lons

Name of compound Positive Formula Negative Ratio **Cation : Anion** ion ion in compound Sodium hydroxide Na^+ OH-NaOH 1:1 Zinc hydroxide Zn(OH)₂ Zn²⁺ OH-1:2 Aluminium hydroxide Al³⁺ Al(OH)₃ OH-1:3

For Complex ions, a bracket is used when there is more than one set.

Name of compound	Positive ion	Negative ion	Formula	Ratio Cation : Anion in compound
Sodium sulfate	Na ⁺	SO4 ²⁻	Na ₂ SO ₄	2:1
Zinc sulfate	Zn ²⁺	SO4 ²⁻	ZnSO ₄	1:1
Aluminium sulfate	Al ³⁺	SO4 ²⁻	$AI_2(SO_4)_3$	2:3

Name of compound	Positive ion	Negative ion	Formula	Ratio Cation : Anion in compound
Sodium carbonate	Na ⁺	CO32-	Na ₂ CO ₃	2:1
Zinc carbonate	Zn ²⁺	CO32-	ZnCO₃	1:1
Aluminium carbonate	Al ³⁺	CO ₃ ²⁻	$AI_2(CO_3)_3$	2:3

Name of compound	Positive ion	Negative ion	Formula	Ratio Cation : Anion in compound
Sodium nitrate	Na ⁺	NO₃ ⁻	NaNO ₃	1:1
Zinc nitrate	Zn ²⁺	NO ₃ -	Zn(NO ₃) ₂	1:2
Aluminium nitrate	Al ³⁺	NO₃ ⁻	Al(NO ₃) ₃	1:3

Name of compound	Positive	Negative	Formula	Ratio
	ion	ion		Cation : Anion
				in compound
Ammonium nitrate	NH_4^+	NO ₃ -	NH ₄ NO ₃	1:1
Ammonium carbonate	NH_4^+	CO32-	(NH ₄) ₂ CO ₃	2:1
Ammonium sulfate	NH_4^+	SO4 ²⁻	(NH ₄) ₂ SO ₄	2:1
Ammonium hydroxide	NH_4^+	OH	NH ₄ OH	1:1



NAMING OF COMPOUNDS WITH TRANSITION METALS

Metal / Oxidation State	Formula of ion
Copper (I)	Cu+
Copper (II)	CU ²⁺
Iron (II)	Fe ²⁺
Iron (III)	Fe ³⁺
Lead (II)	Pb ²⁺
Lead (IV)	Pb4+

The metals listed below can have more than one oxidation state (charge). Hence it is important to state which oxidation state is present in the compound during naming

Formula	Name of compound
Cu_2S	Copper (I) sulfide
CuCl ₂	Copper (II) chloride
FeCl ₂	Iron (II) chloride
Fe ₂ O ₃	Iron (III) oxide
Pbl ₂	Lead (II) iodide
PbO ₂	Lead (IV) oxide

NAMING OF COVALENT COMPOUNDS (NON METALLIC COMPOUNDS)

The following prefixes are used:

1 ~ Mono 2 ~ Di 3 ~ Tri

Formula	Name of compound
СО	Carbon <u>mono</u> xide
CO ₂	Carbon <u>dio</u> xide
SO3	Sulfur <u>trio</u> xide
SO ₂	Sulfur <u>dio</u> xide
F ₂ O	Oxygen <u>di</u> fluoride



BALANCING CHEMICAL EQUATIONS

All chemical equations consists of two portion naming reactants reacting to form products.

Reactant A + Reactant B \rightarrow Product C + Product D

The number of elements on both sides of the equation must be the same for an equation to be balance.







STATE SYMBOLS

General classification	State symbol (at room temperature)
Metal	(s) except for Mercury (I)
Metal oxide	(s)
Metal carbonate	(S)
Note: All SPA carbonates can be in (aq) form if stated.	
Acid (Eg: HCI, HNO ₃ , H ₂ SO ₄)	(aq) except for conc acid (I)
Alkali (soluble bases)	
Group I metal hydroxides	(aq)
(Eg: NaOH, LiOH, KOH)	
Group II metal hydroxides (last 4 elements)	(aq)
(Eg: Ca(OH)2, Ba(OH)2)	
 Ammonium hydroxide (NH₄OH) 	(aq)
Gas (eg. H ₂ , CO ₂ , NH ₃ , O ₂)	(g)
Water	(I)
Mercury	(1)

Note: All SPA (Sodium, Potassium & Ammonium) salts are (aq)

Salts Solubility Table:

General Classification	Soluble Salt (aq)	Insoluble Salt (s)
Nitrates	All nitrates are (aq)	None
Chlorides	All chlorides are (aq) except for	Lead(II) chloride, PbCl ₂ (s)
		Silver chloride, AgCl (s)
	All sulfates are (aq) except for	Calcium sulfate, CaSO4 (s)
Sulfates		Lead(II) sulfate, PbSO4 (s)
		Barium sulfate, BaSO4 (s)



EXAMPLES

Compound	State symbol	Compound	State symbol
Ca(NO ₃) ₂	(aq)	LiNO3	(aq)
Fe(NO ₃) ₂	(aq)	KNO3	(aq)
FeCl₃	(aq)	BaCl ₂	(aq)
AICI ₃	(aq)	CaCl ₂	(aq)
PbCl ₂	(s)	MgCl ₂	(aq)
Fe ₃ (SO ₄) ₂	(aq)	ZnSO₄	(aq)
CaSO4	(S)	BeSO4	(aq)
Li ₂ SO ₄	(aq)	PbSO ₄	(s)
Fe(OH) ₃	(S)	NaOH	(aq)
Zn(OH)₃	(S)	Ca(OH)2	(aq)
Ba(OH) ₂	(aq)	Cu(OH) ₂	(s)
КОН	(aq)	LiOH	(aq)

Substance	State symbol	Substance	State symbol
Mg	(s)	К	(S)
Fe	(S)	Са	(S)
Zn	(s)	Li	(S)
ZnO	(s)	Fe ₂ O ₃	(S)
PbO	(s)	K ₂ O	(S)
CuO	(s)	MgO	(S)
BeCO ₃	(s)	CaCO3	(S)
C∪CO₃	(s)	FeCO ₃	(S)
HCI	(aq) if hydrochloric acid		
HCI	(g) if hydrogen ch	(g) if hydrogen chloride gas	
HCI	(I) if concentrated hydrochloric acid		



ELEMENT, MIXTURES AND COMPOUNDS

DEFINITIONS

Element	Mixture	Compounds
A substance which	Is formed when 2 or more	A substance that is made
cannot decompose	elements or compounds	up of two or more
further into 2 or more	are added together	elements chemically
simpler substances by	without any chemical	bonded in a fixed ratio
chemical means	bonds formed.	

EXAMPLES

Element	Mixture	Compounds
Metals Eg. Mg, Fe, Cu	Seawater	Covalent compound
Non Metals Monoatomic element (1 atom) – Neon, Argon	 (Salt, Sand + Water) Sugar/Salt solution (Sugar + Water) (Salt + Water) 	 Methane, CH₄ Ammonia, NH₃ Carbon dioxide, CO₂
Diatomic molecule (2 atoms)– H ₂ , O ₂ , Cl ₂ , F ₂ , Br ₂ Triatomic molecule (3 atoms) – O ₃	Alloy (Brass – Zinc + Copper) (Steel – Iron + Carbon) 	 Ionic Compound Sodium chloride, NaCl Iron (III) nitrate. Fe(NO₃)₃



DIFFERENCES

Mixture	Compounds	
Composition		
The components present can be in <u>any</u> <u>composition or ratio.</u>	Elements present must be in a <u>fixed</u> composition or ratio.	
E.g. Air is a mixture of gases. The proportion of the different gases in air is varying at different places.	 E.g. H₂O can only consists of 2H atoms and 1 O atom. NH₃ can only consists of 1N atom and 3 H atoms. 	
Melting and	Boiling point	
Has a <u>range</u> of melting and boiling points Heating curve showing on a gradual change.	Has ONE <u>fixed</u> melting and boiling point Heating curve showing one fixed point $t^{0}C$	
250	200- 150- 100- 50- 2' 4' 6' 8' 10' 12 14 Time of heating (minutes)	
Chemico	Il properties	
Have the same chemical properties with its constituent substances .	Different chemical properties from its constituent elements.	
Separation	n techniques	
Can be easily separated by physical methods such as using	Difficult to break down into constituent elements.	
 Filtration - soluble/insoluble, Magnet - Iron & other subst Distillation - Different bp, Sublimation - Solid to gas directly 	Requires <u>chemical methods</u> such as electrolysis and thermal decomposition.	



Experimental Techniques

1. Measurement of Time

- Measured with a digital stopwatch or an analogue stopwatch
- Has an SI unit of **seconds**
- Accuracy of ±0.01s for digital stopwatch.

2. Measurement of Temperature

- Measured with a mercury-in-glass thermometer or alcohol-in-glass thermometer
- Has the SI unit of Kelvin (K)
- Usually with an accuracy of $\pm 0.5^{\circ}C$.

• This depends on the thermometer used. If the markings have more divisions between adjacent integers then it is more accurate.

Accuracy:

0.05 cm³



3. Measurement of Mass

- Measured with an <u>electronic balance</u> or the <u>beam</u> <u>balance</u>
- Has an accuracy of <u>0.01g</u>
- SI Unit: **kg**

4. Measurement of Volume

• Has an SI unit of **m**³

Conversion of units between dm³ and cm³











For laboratory purposes, measured usually in <u>cm³</u> or for calculation purposes, in <u>dm³</u>.

Apparatus	Accuracy
Beaker	Used to measure an approximate volume of a liquid.
Chinikan	This is NEVER used for accurate measurement. Example: approximately 100 cm ³
Measuring	Measures up to the nearest whole number.
Cylinder	E.g 13 cm ³
Burette	Accurately measures out the volume of a liquid to the nearest 0.05 cm³
	Scaled marked (graduated) in 0.10 cm ³ divisions
■	In the laboratory, it measures up to maximum of 50.00 cm³ of liquid. (Readings to 2 d.p)
U	
Pipette	Accurately measures fixed volumes of liquid.
N	(Readings to 1 d.p)
	Usually come in fixed volumes liked <u>20.0cm³, 25.0cm³</u>



Using a burette

- For a burette, the volume is **measured 'top-down'** so be careful when reading the value.
- For accurate reading, **align your eyes to the liquid level**. Then, read off the scale at the **bottom of the meniscus**.





Collection of Gases

- There are 4 different methods of collecting gases:
 - Gas Syringe
 - Displacement of Water
 - Upward Delivery
 - Downward Delivery
- To determine which method to use, we have to look at:
 - o <u>Density of Gas</u>
 - o <u>Solubility of Gas</u>

Gas	Chemical Formula	Solubility in Water	Density compared to Air	Method of collection
Ammonia	NH3	Extremely soluble	Less dense M _r = 14 +1+1+1 =17	Upward delivery
Sulfur Dioxide	\$O ₂	Very soluble	More dense Mr = 32 +16+16 =64	Downward delivery
Hydrogen Chloride	HCI	Very soluble	More dense Mr = 1+35.5 =36.5	Downward delivery
Chlorine	Cl ₂	Soluble	More dense Mr = 35.5 +35.5 = 71	Downward delivery
Oxygen	O ₂	Sparingly soluble	More dense M _r = 16+16 =32	Downward delivery, Displacement of water
Carbon Dioxide	CO ₂	Slightly soluble	More dense Mr = 12 +16+16 =44	Downward delivery, Displacement of water
Hydrogen	H ₂	Insoluble	Less dense Mr = 1+1 =2	Upward delivery, Displacement of water

Note:

Mr refers to the molecular mass of the gas. This is calculated by addition of the mass number of the atoms in one molecule of gas.

The mass of gases is compared to the mass of N_2 ($M_r = 28$), any gas whose M_r is heavier than 28 is a dense gas.



	T	D1
collection	Type of gases collected	Diagram
Gas Syringe		
 Has units of <u>cm³</u> Can contain a maximum volume of <u>100cm³</u> of gas 	This method of gas collection can be used for any gas . Gases can be collected and measured .	gas syringe
Downward Delivery	Used to collect gases which are denser than air) (tap (unnel
The gas collects by displacing the air out the gas jar. The denser gas flows into the jar and air is pushed up and out of the jar. Upward Delivery The gas flows into	Examples of such gases include: • CO ₂ • HCl • Cl ₂ • SO ₂ Used to collect gases which are less dense than air	Collactor of heavy gas (Downward Balvory)
upward into the gas jar and pushes the air down and out of the jar.	Examples of such gases include: • H ₂ • NH ₃	upward delivery (downward displacement of air)
Used to collect gases which are <u>insoluble</u> , <u>slightly soluble or</u> <u>sparingly soluble in</u> <u>water</u>	Gases collected via this method should not be soluble in water. Examples of such gases include N ₂ CO ₂ (sparingly soluble) O ₂ (sparingly soluble) H ₂ (insoluble)	thistic funnal delivery tube insoluble gas Collection of insoluble gas over water







TEST FOR GASES

Gas	Test	Observation	Diagram		
Hydrogen, H2	Place a lighted splint at the mouth of the test tube	If H ₂ is present, lighted splint will extinguish with a "pop" sound	burning splint		
Oxygen, O2	Inset a glowing splint into the test tube	If O2 is present, the glowing splint will rekindle	Glowing Splint Oxygen Gas		
Carbon dioxide, CO2	Bubble the gas into limewater	If CO2 is present, a white precipitate /ppt will form in limewater	acid marble chips		
Ammonia, NH₃	Place a damp red litmus at the mouth of the test tube	If NH₃ is present the red litmus will turn blue	Red Blue		
Chlorine, Cl ₂	Place a damp red/blue litmus at the mouth of the test tube	If Cl ₂ is present, both red and blue litmus will be bleached	Red Bleaches		



SEPARATION TECHNIQUES

SEPARATING SOLID-SOLID MIXTURES

Dissolve and filter: when only one solid is soluble

Procedure:

- 1. Choose a suitable solvent such that only glass rod one solid is soluble.
- 2. Add the solvent to the mixture to **dissolve** the soluble solid.
- 3. The mixture is then poured through the filter funnel.
- 4. The <u>residue</u>, which is the insoluble solid, remains on the filter paper.
- 5. The *filtrate* is collected in the beaker.
- The <u>residue</u> is then washed with distilled water and pressed dry between the filter paper.

Sublimation: when only one solid can sublime

Procedure:

- 1. The mixture is placed in an evaporating dish
- 2. An inverted filter funnel is placed over the mixture.
- Upon gentle heating, the solid, e.g. <u>ammonium</u> <u>chloride or iodine</u> will sublime and it will <u>deposit</u> on the cold filter funnel.

Substances that will sublime a) Ammonium chloride, NH4CI b) Iodine, I2

- c) Naphthalene (Moth balls)
- d) Dry ice





Using Magnet: when only one solid is a magnetic material

Example: Separating a mixture of iron filings and sulfur

Note: As the above mentioned is a mixture, no chemical bonding has occurred between Iron and sulfur. The magnetic properties of Iron is still present and iron is thus still attracted to a magnet. Hence in a mixture the constituents still retain their chemical properties.



SEPARATING SOLID-LIQUID MIXTURES

These methods are used to separate a solid (solute) that has dissolve in a liquid (solvent). The objective is to remove the liquid and obtain the solid.

Method	Procedure	Diagram
Evaporation to dryness This method is only used for solutes that are stable to heat and <u>do not</u> <u>decompose on heating</u> . E.g. separating sodium chloride crystal from salt solution	 Place the sodium chloride solution in an evaporating dish. During heating, evaporation of <u>WATER/SOLVENT/LIQUID</u> occurs until only sodium chloride is left behind. 	
Crystallisation Crystallisation is used to obtain hydrated crystals which are unstable to heat. These hydrated crystals will dehydrate or lose their water of crystallization by heating Examples: Hydrated Copper (II) Sulfate CuSO4.5H2O	 After filtration, the filtrate is heated gently until saturated. Crystals will form readily on cooling. Allow the saturated solution to cool over time Crystals will form on cooling Filter/decant to collect the crystals Press dry the crystals between filter paper 	evaporating basin the hot solution is allowed to cool. The solid appears as pure crystals. The solution is heated to evaporate most of the solvent. The cold solution is poured off to obtain the crystals. The crystals may be dried by pressing them between sheets of filter paper.











Separating Funnel (Immiscible liquids)

Immiscible liquids are liquids that do not dissolve in each other completely to form a solution.

E.g. Oil and water.

- 1. Pour the mixture of oil and water into the separating funnel.
- 2. Wait for the 2 layers to separate out. Since water is denser than oil, it will form the bottom layer.
- 3. Open the tap to allow the water to run into the beaker.
- 4. Close the tap before the oil layer runs out of the funnel
- 5. The remaining oil layer can be poured out from the separating funnel





PAPER CHROMATOGRAPHY

Paper chromatography is used to separate a mixture of solutes in a solution/<u>separate inks into its dyes</u>.

The dyes are separated due to differences in the same solvent used.

The dye that is more soluble will travel further up the filter paper.

E.g. inks, food colourings, amino acids, etc.







Retention Factor (R_f)





sous and the same R_f value will appear on the same position on the chromatogram.

```
R_{f} = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}} = \frac{Y}{X}
```





Criteria of Purity

The purity of a substance cannot be determined by chemical tests, it is usually determined by physical means such as measuring the **melting or boiling point** of a substance or by **chromatography**

If a substance is pure, it will melt/boil at **one fixed temperature.**

If impurities are present, impurities will **increase the boiling point** (boiling will only start at a higher temperature) and **decrease the melting point** (melting will occur at an even lower temperature) of a substance. Impure substances also melts and boils over a range of temperature.



Temperature – Time graph of Pure Substance



Melting of Impure Solids

An impure solid will melt at a lower than the pure solid

Boiling of Impure Liquids

Presence of impurities in liquids:

 (a) If the impurities are solids and are insoluble in liquid, there is no effect on boiling point Eg.
 Chalk in water, water will still boil at 100°C



(b) If the impurities dissolves in liquid, the liquid will boil over a range of temperature at a high boiling point

Eg Salt solution boils at 103°C over a range of temperature

Case Studies

	A pure solid X melts at 94°C. At which -		
	temperature will impure X melt? 88 °C, 96 °C, 94 °C 98 °C Ans: 88 °C (The presence of impurities lower the mp. Hence impure X must melt at temperature below the pure mp.	N te o(A no	laphthalene (impure) melts at 78°C. At which emperature will pure Naphthalene melt? 73°C, 72 C 75°C, 80°C ns: 80°C (The presence of impurities lower the mp. Hence pure aphthalene must melt at a higher temperature compare to that f the impure
	An impure liquid boils at 42 °C. Which		
	43 °C, 49 °C Ans: 38 °C (The presence of impurities rises the bp Hence pure Y must boil at a lower temperature compare to that of the impure	э.	Pure solid Z has the same mp as Benzoic acid which is at 80°C. Describe an experiment to determine whether solid Z is pure benzoic acid. Ans: Mix Z with a little benzoic acid and measure the melting
		-	point of the mixture.
			If Z is actually benzoic acid, the melting point will be constant and fixed at 80°C.
Explain why salt is sprinkled on surface of roads during winter. Ans: Ice is slippery and we do not want it to form. Hence when a floor is wet, we sprinkle salt to lower the melting point of ice, so that a much lower temperature is required before water can freeze. Thus ice can only form at -4°C instead of 0°C			If Z is not benzoic acid, then pure benzoic acid will be acting like an impurity causing Z to melt at a lower temperature 78°C and it will melt over a range of temperature
		JS _	
			Which has the highest boiling point? Tap water, river water or sea water?
			Ans: Sea water (Has the highest amount of impurities) Since the rise in bp is affected by the amount of dissolved impurities in the liquid, thus sea water will have the highest bp.



KINETIC PARTICLE THEORY

PART I: ARRANGEMENT, MOVEMENT AND ENERGY OF SOLIDS, LIQUIDS AND GASES

	Solid	Liquid	Gas
Diagram			
Drawing			
Arrangement	Particles are tightly packed in an orderly arrangement	Particles are tightly packed but in a disorderly arrangement	Particles are spaced far apart from one another and randomly arranged
Movement	Particles vibrate in their fixed positions	Particles slide past each other	Particles move at high speeds randomly in all directions
Energy	Low	Moderate	High



PART II: CHANGES IN STATE



PART II A: THE BOILING & MELTING POINTS OF A PURE SUBSTANCE

A **pure substance** has a <u>fixed</u> melting and boiling point.

When *impurity is present:* Mixture has a range of melting and boiling point

- **Boiling point** of the original pure substance **increases**
- Melting point of the original pure substance decreases.

PART III: STATE OF MATTERS APPLICATION



Substance with temperature:

- (a) Below the melting point exists as **Solid**
- (b) Above the boiling point exists as **Gas**
- (c) Between the melting and boiling point exists as Liquid

Practice question:

Substance	Melting Point / °C	Boiling Point / °C	State at room temperature 30°C
A	13	79	Liquid
В	36	185	
С	-22	15	
D	-190	-12	



PART IV: EXPLAINING BOILING AND MELTING CURVE IN TERMS OF KPT



When a liquid changes to a gas (heats up until boiling point),

- As the liquid is heated, the particles **gain Kinetic Energy** and start to move faster when **temperature rises**.
- As it reaches the **boiling point**, the temperature remains **CONSTANT**!

During boiling, the <u>heat energy supplied is absorbed</u> by the particles to <u>overcome the attractive forces</u> between the particles to change from liquid to gas.

• When all the liquid has completely boiled, the temperature rises again as the kinetic energy of the particles increases due to heating.



B) MELTING CURVE





When a solid changes to a liquid (heats up until melting point),

- As the solid is heated, the solid particles gain <u>Kinetic Energy</u> and start to vibrate faster as the <u>temperature rises</u>.
- As it reaches the **melting point**, the temperature remains **CONSTANT!**
- During melting, the <u>heat energy supplied is absorbed</u> by the particles to overcome the attractive forces between the particles to change from solid to liquid.
- When the solid has completely melted, the temperature rises again as the kinetic energy of the particles increases due to heating.

HEATING CURVE - THE COMBINED GRAPH INVOLVING MELTING THEN BOILING




PART V: EXPLAINING FREEZING CURVE IN TERMS OF KPT

A) THE FREEZING CURVE

Freezing curve



When a liquid changes to a solid (cools down until melting point),

- As the liquid is cooled, the particles <u>lose kinetic energy</u> and move slower when <u>temperature decreases</u>.
- As it reaches the **melting point**, the temperature remains **CONSTANT**!
- During freezing, the <u>heat energy is released by the particles due to the</u> <u>strengthening of attractive forces</u> between the particles to change from liquid to solid.
- The heat released by the particles is equal to the heat lost to the surroundings during cooling. (E.g. The freezing of water in the freezer)
- When all the liquid has completely frozen into solid, the temperature decreases again as the kinetic energy of the particles decreases due to cooling.



B) COOLING CURVE – THE COMBINED GRAPH INVOLVING CONDENSATION THEN FREEZING



PART VI: EVAPORATION <u>WITHOUT</u> HEATING – [THE <u>COOLING EFFECT</u> OF EVAPORATION]

Question: It is observed that the beaker of alcohol experiences a drop in its temperature with time. Explain.

- The particles at the surface of the alcohol have higher kinetic energy as compared to the rest of the alcohol.
- The surface particles of alcohol have enough energy to break away from the liquid and leave as gas. This is the evaporation process and takes place at the surface of the liquid only.
- This evaporation process thus leaves behind particles with lower kinetic energy, leading to lower temperature.



During evaporation, the more energetic particles escape from the surface leaving the less energetic ones behind.



CHEMICAL BONDING

An atom / ion is stable when it has a noble gas configuration. This is achieve it has an octet structure (8 valence electrons) or a duplet structure (2 valence electrons in the 1st shell)

Atoms will try to		XX	Gaining electrons \rightarrow forming cations (+)
stable noble	(He)	Ne ¥	Loosing electrons \rightarrow forming anions (-)
gas structure of:			Sharing electrons
	×	~~	

To achieve a noble gas structure

(a) Metals which have 1, 2, 3 valence electrons will tend to lose electrons to form positive ions (cations)

Eg. Al \rightarrow Al³⁺ 2.8.3 2.8 K \rightarrow K⁺ 2.8.8.1 2.8.8

(b) Non Metals which have 5, 6, 7 valence electrons will tend to gain electrons to form negative ions (anions)

Eg.	O 2.6	→	O ²⁻ 2.8
	N 2.5	\rightarrow	N³- 2.8

When different / similar elements combined to form a compound / element, the following bonding may occur:

(a) Ionic bonding	- occurs between metallic and non metallic elements
(b) Covalent bonding	- occurs between non metallic elements
(c) Metallic bonding	- occurs only for metals



IONIC BONDING

An ionic compound is formed between a metal and non metal.

To achieve a stable noble gas structure, a metal atom will loose electrons to become a positive ion. The electron(s) lost will be taken by the non metal atom to form a negative ion. This will results in an electrostatic attraction (ionic bonding) between the positive and negative ions, resulting in the formation of an ionic compound.

Example to describe how an ionic compound is formed.

Sodium chloride is formed when a sodium atom (2.8.1) loses 1 electron to for a sodium ion, Na⁺ (2.8) and achieve an octet structure. This electron is transferred to a chlorine atom (2.8.7) which becomes a chloride ion, Cl⁻ (2.8.8). An ionic bonding is thus formed between the Na⁺ and Cl⁻ ion.

Examples of "dot and cross" diagrams of ionic compounds

Sodium Chloride, NaCl

Calcium Fluoride, CaF2

Lithium Oxide, Li₂O

Aluminium Oxide, Al₂O₃



Properties of Ionic Compounds

No.	Properties	Explanation
1	Are always solids at room temperature	Oppositely charged ions attract each other very strongly (electrostatic forces) and therefore pulled each other close together.
2	Have high melting point	A lot of energy is required to break the strong electrostatic attraction between millions of appositively charged ions in the giant lattice structure.
		The forces of attraction increases as the charges between the ions increases. Eg MgO has a higher mp than NaCl
3	Cannot conduct electricity in the solid state but,	In the solid state, the ions are held very strongly by electrostatic attraction in a fixed lattice structure and are not free to move. Hence the solid cannot conduct electricity
	when molten or dissolved in water can conduct electricity	In the molten or aqueous state, the ionic lattice structure is no longer present. The ions are farther apart and not so strongly attracted to each other. They are mobile now and can conduct electricity
4	Most are soluble in water and all are not soluble in organic solvents.	Insoluble salts and hydroxides are insoluble in water eg. AgCl, PbCl ₂ , CaSO ₄ , BaSO ₄ , PbSO ₄ , Cu(OH) ₂



The diagram shows the giant lattice structure of sodium chloride. The ions are kept in a fixed position and held closely by the strong electrostatic attraction.

In the solid state, the ions are not mobile. No electricity is conducted. The light bulb does not light up.





COVALENT BONDING

A Covalent **compound** is formed when two or more non-metals are bonded.

Eg. Hydrogen and Oxygen forms Water; Nitrogen and Hydrogen forms Ammonia

A Covalent **element** can be formed when two atoms of the elements are bonded.

Eg. Nitrogen gas, N₂, Oxygen gas O₂

Types of Covalent Bonding

A **single** covalent bond is formed each non metal atom share an electron with the other. Thus a single covalent bond has a pair of electrons (2 electrons shared).

A double covalent bond has 2 pairs of electrons (4 electrons shared)

A triple covalent bond has 3 pairs of electrons (6 electrons shared)

Each of this can be shown below.

Single Covalent bonding	Double Covalent bonding	Triple Covalent bonding
Methane, CH₄	Oxygen gas, O ₂	Nitrogen gas, N ₂

Common Covalent Molecules are

Covalent elements – O₂, Cl₂, F₂, I₂, H₂ (Di-atomic molecule); O₃ (Tri-atomic molecule)

Covalent compounds – H₂O, SO₂, CH₄, CCl₄, NH₃

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No.	Properties	Explanation
1	Majority are gases, but a few are liquids at room temperature. • O ₂ ,CO ₂ , CO, H ₂ (gases) • H ₂ O, CCl ₄ (liquids)	In covalent molecules, the intermolecular forces (Van Der Waals forces) between the molecules/particles is weak. The particles are not strongly attracted and are further apart. Hence they are gas/liquid.
		The intermolecular force increases with the size of the molecule Hence the small covalent molecules are always gases while the larger ones are liquids.
2	Low melting and boiling point	In covalent compounds the molecules are held together held together by the much weaker intermolecular forces. Much lesser energy is needed to overcome this force. Hence it has a low melting point.
3	Does not conduct electricity in any state.	There is no mobile ions or electrons for the conduction of electricity.
4	Most are insoluble in water but are soluble in organic solvents.	





GIANT STRUCTURES



GIANT COVALENT STRUCTURES

		CDADUITE
DIAMOND	SILICON DIOXIDE	GRAPHIE
	Also known as Quartz/Sand	
Each carbon atoms are	Each silicon atom is	Carbon atoms are covalently
covalently bonded to 4	covalently bonded to 4	bonded to 3 other carbon
other carbon atoms in a	oxygen atom in a	atoms in hexagonal rings in
tetrahedral arrangement	tetrahedral arrangement	lavers
The bonds of the	However each O atom is	Between each hexagonal
tetrahedrons form an	bonded to 2 Silicon atom.	laver, there are weak
extensively interlocked	Hence formula SiO ₂	intermolecular forces that hold
network resulting in a gight		each laver loosely. This results
structure which is very bard	In SiO ₂ , there is less	in graphite layer been able to
and rigid	intorlocking of bonds	he denosited on paper when
ana ngia.		be deposited on puper when
	compared to diamond as	The infermolecular forces are
Has a high melting point as	each O atom has only 2	broken.
a lot of heat is needed to	bonds (compared to 4	
break the large number of	bonds for C atoms in	As each carbon atom is
strong covalent bonding	diamond)	covalently bonded to 3 other
between C atoms.		carbon atoms, this results in
	Hence it is not as rigid and	each carbon atom having a
	has much lesser bonds to	free electron not bonded. As
	break. This results in less heat	a result these free electrons
	required to break the bonds	can move round the
	and SiO_2 has a lower mp	hexagonal ring forming a pool
	than diamond	of mobile electrons



 Properties: High mp and bp Hard and rigid Insoluble in water and almost all solvents 	 Properties Similar to diamond but with a lower mp due to less interlocking bonds 	 Properties High mp and bp Soft and slippery Conduct electricity in solid state
 Uses: As drill bits for drilling rocks due to its hardness To cut other diamond As a precious gem due to its high durability and high refractive index 		 Uses As lubricants to reduce friction in engines due to its being soft and slippery. (Due to weak intermolecular forces between layers, allowing each layer to slip over each other. Pencil lead due to its ability to be deposited on paper. Pressure breaks the intermolecular forces between layer As electrodes for electrical conduction (Due to the presence of free electrons as each Carbon is only bonded to 3 other within a layer)
		Graphite strong covalent bonds between the carbon atoms in each layer weaker forces hold the layers together

ALLOTROPES

Allotropes are different physical forms of the same element that results due to different arrangement in atoms in the structure.

Graphite and Diamond are allotropes of each other.

Allotropes are physically different but they are chemically identical (same chemical properties) eg. 12g of diamond and graphite will both burn at high temperature to produce 1 mole of CO₂



SUMMARY

	Structure	Мр/Вр	Electrical Conductivity	Solubility
Metals	Metallic structure	High	Conduct in all state	Not soluble in both water and organic solvent
lonic Compound	Ionic Lattice	High	No conduction in solid state but conduction in molten and when aqueous solution	Most are soluble in water except insoluble salts Not soluble in organic solvent
Covalent Compound	Simple Molecular	Low	No conduction	Soluble in organic solvent but not in water
Diamond	Giant Covalent Structure	High	No conduction	Not soluble in both water and organic solvent
Graphite	Giant Covalent Structure	High	Conduct in all state	Not soluble in both water and organic solvent
Silicon dioxide	Giant Covalent Structure	High	No conduction	Not soluble in both water and organic solvent





Definition: A compound that **produces H+ ions as the only positive ions** when dissolve in water.

Acids only have properties when dissolved in water.

For Example:

A compound such as the following:

Citric acid crystals	Solid	Have no effect on dry litmus paper
Hydrogen chloride gas	Gas	H ₂ O is required for the formation of H ⁺ or H₃O⁺ (hydroxonium ion), is the hydrated hydrogen ion.

Physical properties of Acid

- Colourless
- Sour in taste

- pH less than 7
- Turns blue litmus red
- Turns universal orange (weak) / red
 (strong)

Basicity of Acids

Refer to the number of replaceable hydrogen atoms in one molecule of the acid.

Acid	Basicity	Equation
H ₂ SO ₄	2 (Dibasic)	H ₂ SO ₄ → 2H ⁺ + SO ₄ ²⁻
HCI	1 (Monobasic)	HCI → H + + CI ⁻
HNO3	1 (Monobasic)	HNO3 → H + + NO3 ⁻
H ₃ PO ₄	3 (Tribasic)	H ₃ PO₄ → 3H⁺ + PO₄ ³ -



Chemical Properties & Reactions of Acids

• Pure metal + Acid \rightarrow Salt + Hydrogen gas Mg (s) + 2HCl (aq) \rightarrow MgCl₂ (aq) + H₂ (g) Zn (s) + 2HNO₃ (aq) \rightarrow Zn(NO₃)₂ (aq) + H₂ (g)

(Unreactive pure metal like copper, silver, gold and platinum have no reaction with dilute acid)

Cu (s) + HCl (aq) \rightarrow No reaction

Metal oxide + Acid → Salt + Water
 Li₂O (s) + H₂SO₄ (aq) -> Li₂SO₄ (aq) + H₂O (l)

(Neutralization)

Metal hydroxide + Acid → Salt + Water (Neutralization)
 Ca(OH)₂ (aq) + 2HNO₃ (aq) -> Ca(NO₃)₂ (aq) + H₂O (I)

(Metal oxide and hydroxides are Bases)*

Metal carbonate + Acid → Salt + Water + Carbon dioxide gas
 CuCO₃ (s) + H₂SO₄ (aq) -> CuSO₄ (aq) + H₂O (l) + CO₂ (g)

Test for Hydrogen gas	Test for carbon dioxide gas
"POP"! burning splint (hydrogen burns with a 'pop' sound) magnesium ribbon dilute hydrochloric acid	Delivery tube Hydrochloric acid limewater Sodium carbonate
In the presence of hydrogen gas, the lighted splint extingishes with a 'pop' sound	With some CO2: CO2 (g) + Ca(OH)2 (aq) → CaCO3 (s) + H2O (l) white precipitate forms With excess CO2: CaCO3 (s) + H2O (l) + CO2 (g) → Ca(HCO3)2 (aq) calcium bicarbonate White precipitate disappears with an excess amount of CO2!



Strong and Weak acids

Strong Acids	Weak acids
Completely ionised in aqueous	Partially ionised in aqueous solution/H2O
(H)	CH,COOH CH,COOH CH,COOH CH,COOH CH,COOH Partially dissociated
After ionisation, all the molecules have ionized, there are no un-ionised molecules.	After ionisation, a few molecules are ionised, there are a lot of un-ionised molecules in solution.
High H ⁺ concentration	Low H ⁺ concentration
Low pH	High pH
pH 0 – 2	pH 3 - 5 (Weak acid [—] Ethanoic acid and all organic acids, Sulfurous acid) pH 5 - 6+ (Very weak acids — Carbonic acid
Examples Strong Acids:	Examples Weak Acids:
HCI – Hydrochloric acid	H2CO3-Carbonic acid
HNO3-Nitric acid	H ₂ SO ₃ – Sulfurous acid
H ₂ SO ₄ –Sulfuric acid	All organic acids eg. Ethanoic acid
Brightly lit SWITCH FLEECTRODE ELECTRODE Strong acid Higher electrical conductivity due to presence of ions 100% ionisation of molecules $HCI \rightarrow H^+ + Cl^-$ $HNO_3 \rightarrow H^+ + NO_3^-$ $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$	Dimly lit Switch ELECTRODE Weak ION acid Some, not high, as there are very few ions, most remained as neutral molecules. 0.4 % will ionised, 96.6 % will remain as neutral CH ₃ COOH molecules CH ₃ COOH → CH ₃ COO ⁻ + H ⁺





Measurement of pH	Compare rate of bubbling / evolution of gas	Comparison of electrical conductivity
Use pH paper / Universal indicator / pH meter	Add the acid to metal or metal carbonate.	Conductivity can be measure with a ammeter / volt meter / bulb in a circuit
Dip pH paper or add universal indicator. Match the colour obtained with the colour reference chart to obtain the pH value	Observe the effervescence (bubbling) or measure the volume of gas collected in unit time/ fixed time.	Conductivity depends on the presence of ions in the solution. Only charged ions can conduct electricity. Molecules cannot conduct.
	Example Acid A – 50cm ³ of H ₂ (g) produced in 5 min. Rate = 50/5 = 10cm ³ /min	Example Acid A – 1 mol/dm ³ HCl Acid B – 1 mol/dm ³ CH ₃ COOH
	Acid B – 20cm ³ of H ₂ (g) produced in 5 min. Rate = 20/5 = 4cm ³ /min	Image: second system Image: second system Image: second





		_
	Three Methods to determine the Basicity of an Acid	R
$\langle $,	

From the Chemical Formula	By Titration / Balancing an equation	From the number of Sodium salts formed by the acid
HCI is monobasic	Titration	HCI forms 1 Na Salt
(Basicity = 1)	Determine the basicity of an acid by	→ NaCl
	calculating the number of moles of	HCI has a basicity 1
H ₂ SO ₄ , H ₂ CO ₃ are dibasic	alkali required to neutralize one mole	(monobasic)
(Basicity = 2)	of the acid.	
		H ₂ SO ₄ can form 2 Na Salts
H ₃ PO ₄ is tribasic	Applying formula,	\rightarrow Na ₂ SO ₄ and NaHSO ₄
(Basicity = 3)	(MU) acid m (acid)	H ₂ SO ₄ has a basicity 2
	$\frac{(MV)alkali}{(MV)alkali} = \frac{n(alkali)}{n(alkali)}$	
		$H_{2}PO_{4}$ can form 3 Na Salts
	Where M (Molarity), V (Volume) and	\rightarrow Ng ₃ PO ₄ NgH ₂ PO ₄ and
	n (Mole)	NgHPO4
	Balancing equation	H₃PO₄ has a basicity 3
	Acid A + NaOH \rightarrow NaA + H ₂ O	(dibasic)
	From equation, 1 mole of NaOH is neutralized by 1 mole Acid A. Hence Acid A must have a basicity of 1 (monobasic)	
	From equation, 2 mole of NaOH is neutralized by 1 mole Acid B. Hence Acid B must have a basicity of 2 (dibasic)	



Strength of Acid	Concentration of Acid
Refer to the number of molecules that	Refers to the number of particles / mass
are ionised.	dissolved per unit volume of solution
pH 0 – 14	mol/dm ³ or g/dm ³
HCI 1	Acid A is 1.5 mol/dm ³
HNO ₃ Strong	Acid B is 1.0 mol/dm ³
HCI decid	Acid A is more concentrated than B
H_2SO_3 Weak CH $_3COOH$ acid	

Strength Vs Concentration of Acid





Definition: A substance that neutralizes an acid to form salt and water only.

Bases are thus metal oxide and metal hydroxides.

Most bases are insoluble in water.

Bases that can dissolve in water form soluble hydroxides known as **alkali**.



pH more than 7

Physical Properties of Alkali

- Colourless
- Bitter in taste
- Soapy to touch
- Pungent smell

Turns red litmus blue Turns universal indicator blue

• Turns universal indicator blue (weak) / purple (strong)

Alkalis

The ion responsible for its alkali property is the OH- (hydroxide ion)

•

Alkali	No. of OH [.]	Equation
NaOH	1	NaOH → Na+ + OH -
Ca(OH) ₂	2	Ca(OH) ₂ → Ca ²⁺ + 2OH -

Three Categories of Alkali

1.	All Group I hydroxides are alkalis	Eg KOH, NaOH
2.	Aqueous Ammonia or ammonium	NH₄OH
	hydroxide is a weak alkali	
3.	Group II hydroxide (Last four in the Group	Eg. Ca(OH) ₂ , Ba(OH) ₂
	are alkalis)	



Formation of Alkali

 (a) Alkali Metal + Water → Alkali + hydrogen gas (Group I Metals) (Violent Reaction as sodium, potassium are very reactive metals) 2Na (s) + 2H₂O (I) → 2NaOH (aq) + H₂ (g) 	(b) Ammonia gas dissolving in water to form a weak alkali known as ammonia hydroxide or <u>aqueous</u> <u>ammonia</u>
(c) Metal oxide (Soluble ones) dissolve in water	$NH_3 (g) + H_2O \iff NH_4OH (aq)$
$Li_2O(s) + H_2O(aq) \rightarrow 2LiOH(aq)$	Aqueous ammonia is a weak alkali and the reaction is a reversible reaction
	(In a reversible reaction, the products can react to produce the original reactants. At equilibrium, the concentrations of reactants and products do not change)

Solubility of Bases

	Solubility of bases		
	(aq)	(\$)	
Bases Are metal oxides and hydroxides	Soluble bases are known as alkalis. These are the three category shown in Page 53	Metal oxides / hydroxides that do not fall into the three categories of alkalis above are insoluble.	
Examples	Na ₂ O (s) + H ₂ O (l) → 2NaOH (aq) K ₂ O (s) + H ₂ O (l) → 2KOH (aq)	Fe(OH)2 (s)Green pptFe(OH)3 (s)Brown pptCu((OH)2(s)Blue pptCuO (s)Black solid	

Chemical Properties & Reactions of Alkalis





Strong alkali	Weak alkali
Completely ionised in aqueous solution/H2O.	Partially ionised in aqueous solution/H2O
After ionisation, all the molecules have ionized, there are no un-ionised molecules.	After ionisation, a few molecules are ionised, there are a lot of un-ionised molecules in solution.
High concentration of OH ⁻ (hydroxide ions)	Low concentration of OH ⁻ (hydroxide ions)
pH 12 – 14	pH 8 – 11
Strong alkalis are KOH, NaOH	Weak Alkali is NH4OH

Colours with pH Indicators for Acids and Alkalis

Indicators	Acids	Alkalis	
Litmus	Turns blue litmus red	Turns red litmus blue	
Phenolphthalein	Remain colourless	Turns pink	
Methyl orange	Remain red	Turns yellow	
Universal Colour ranges from Orange → Red indicator		Colour ranges from Blue-green → Blue → Violet	



• The smaller the pH value, the more acidic the solution and the more H⁺ ions it contains. A solution of pH 1 has more H⁺ ions than a solution of pH 5.



APPLICATION FOR ACIDS AND ALKALIS

ACIDS

Ethanoic acid:

Making detergents

Making fertilisers

In car batteries

In vinegar to

preserve food

ALKALI / BASES

Ammonia solution:

- In window cleaning solutions
- In fertilisers

Magnesium hydroxide:

- In toothpaste to neutralise acid on teeth
- In antacids to relieve indigestion

Sodium hydroxide:

- In making soaps and detergents
- In industrial-cleaning detergents

Calcium oxide (Quicklime)

• In neutralising acidic soil.

pH control of the soil is important because this affects the growth and development of plants. Most plants grow best when the soil is neutral or slightly acidic (pH 6-7).

Soil can become too acidic from **acid rain** and **excessive use of fertilisers** eg. ammonium sulfate.

• To make iron, concrete and cement

Calcium hydroxide (Slaked lime)

In neutralizing acidic soil

Calcium hydroxide must <u>not</u> be added with ammonium fertilisers at the same time. This is because calcium hydroxide is an alkali, and will react with ammonium salt to liberate <u>ammonia gas</u>. The ammonia gas will <u>escape into</u> <u>the atmosphere</u> resulting in the <u>loss of nitrogen</u> needed for plant growth. Calcium hydroxide which is required to reduce the soil acidity will also be lost.

 Desulfurization – Removal of acidic gases from flue gas (waste gas) eg Sulfur dioxide from factories and power station

 $Ca(OH)_2 + SO_2 \rightarrow CaSO_3 + H2O$



NEUTRALIZATION IN DAILY LIFE

- Indigestion Tablets (For Gastric) These often contain oxides, hydroxides or carbonates of metals like Mg or Ca. Indigestion is often caused by excess acid in the stomach This is neutralized by reaction with these bases found in indigestion tablets (milk of magnesia containing MgO and MgCO₃)
- Toothpaste Decaying food articles in your teeth produces acid. In such acid conditions, bacteria thrive and can cause tooth decay. Toothpaste is alkaline and neutralizes the acid, preventing tooth decay.



- 3. Cleaning of Metals Brass and bronze ornaments can be cleaned by rubbing them with a piece of lemon. The citric acid in the lemon juice can react with the copper (II) oxide layer on the surface to form copper (II) citrate and water which can be washed off. Hence the ornament will become shiny.
- 4. Hair Conditioners Shampoos contain mild alkalis. These cause small scales on each hair to open out which can make our hair unmanageable. Hair conditioner is a mild acid which neutralizes the shampoos and causes the scales to close up.

5. Insect Stings –



- **Bee Stings are acidic (formic acid)**. To treat the sting, we can place a weak alkali, such as baking soda on the affected area. The weak alkali neutralises the acid and relieve the pain.
- **Wasp stings are alkaline**. To treat the sting, we can place a weak acid, such as vinegar on the affected area. The weak acid neutralises the alkali and relieve the pain.

Yellow Jackets have straight stingers and can

sting repeatedly whereas Honey bees have barbed stingers and can only sting their victim once.

Before treating a Honey bee sting make sure to remove the embedded stinger. **Do not pull out the stinger** either with your fingers or tweezers as this will cause **more venom** to be squeezed from the stinger sac into the sting. Rather, **knock out the stinger** by scraping the surface of the skin at a 45 degree angle with a card.



QUESTIONS

Question 1:

Dry ammonia gas has no effect on pH paper/litmus paper but will turn damp litmus paper blue.

Question 2:

Why does aqueous ammonia smell strongly of ammonia gas?

Question 3:

Hydrogen chloride gas dissolved in water is acidic and turn red litmus paper blue, but when hydrogen chloride gas is dissolved in organic solvent such as methylbenzene, there is no acidic properties.

Question 4:

Hydrogen chloride gas dissolved in water conduct an electric current but when it is dissolved in an organic solvent it does not conduct electricity

ANSWER 1:

In the gaseous state, ammonia gas is a covalent compound. When it is dissolved in water, hydroxyl ions are produced. The presence of hydroxyl ions reacts with the pH indicators causing a change of colour to blue.

 NH_3 (g) + H_2O (l) $\rightarrow NH_4^+$ (aq) + OH^- (aq)

ANSWER 2:

When ammonia molecules dissolve in water only a few molecules react with water to form hydroxyl ions. Most of the ammonia molecules remain as neutral molecules. These molecules are responsible for the smell of the gas. The ions NH₄⁺ and OH⁻ ions have no smell. Ammonia is a weak alkali hence there is a reversible reaction between the reactants (ammonia gas and water) and ammonium hydroxide.

ANSWER 3 & 4 (Similar)

Hydrogen chloride gas dissolved in water to produce H_3O^+ , this ion is responsible for all properties of acids.

HCl (g) + H₂O (l)
$$\rightarrow$$
 H₃O⁺ (aq) + Cl⁻ (aq)

In methylbenzene, Hydrogen chloride gas dissolved physically, remaining as neutral covalent molecules. This resultant mixture has does not contain H_3O^+ ions for acidic properties/electrical conductivity.





NEUTRALIZATION REACTION – TITRATION

Titration can be measured in the following 3 ways.

pH Titration	Thermometric Titration	Conductimetric Titration
Measuring the changes in pH during titration	Measuring the changes in temperature during titration	Measuring the change in electrical conductivity of the solution during titration
To obtain the point of neutralization, volume of alkali/acid where there is a sudden change in pH is taken.	To obtain the point of neutralization, the volume of alkali/acid where the maximum temperature is reached.	To obtain the point of neutralization, the volume of alkali/acid where the electrical conductivity is the lowest.

Part I

pH Titration



Strong Acid - Strong Base Titration

At A: No alkali is present. There is only acid (H⁺) present. pH is 2 indicating it is a strong acid (HCl/H₂SO₄) – H⁺ concentration is high

From A to B: OH- ions from alkali are introduced and react with H+

 H^+ (aq) + OH^- (aq) \rightarrow H_2O (I)

At B: All H⁺ ions are neutralized. (Point of Neutralization) 1 drop of strong alkali will rise the pH instanteously.

At C: pH increases slowly as OH- added. The pH continues to increase till the final pH of the alkali.

Weak Acid - Strong Base Titration

At A1: No alkali is present. There is only acid (H⁺) present. pH is 4 indicating it is a **weak acid** (HCI/H₂SO₄) – H⁺ concentration is low

Strong Base - Strong Acid Titration

At X: No acid is present. There is only alkali (OH-) present. pH is 14 indicating it is a strong alkali – OH- concentration is high

From X to Y: $\rm H^{+}$ ions from acid are introduced and react with OH-

 H^+ (aq) + OH^- (aq) \rightarrow H_2O (I)

At Y: All OH-ions are neutralized. (Point of Neutralization) 1 drop of strong acid will drop the pH instanteously.

At **Z**: pH decreases slowly as H⁺ added. The pH continues to decrease till the final pH of the acid.

At Z1: As the acid is a weak acid. The decrease will be till pH 3 – 4







Overall Reaction: $H_3PO_4 + 2NaOH \rightarrow 3H_2O + Na_3PO_4$

Conductimetric Titration



Volume of Acid / cm³

At A: No acid is present. There is only alkali (OH-) present. High conductivity as the Na⁺ and OH⁻ concentration is high

From A to B: H+ ions from acid are introduced and react with OH-.

 H^+ (aq) + OH^- (aq) \rightarrow H_2O (I)

The number of ions are reduced. Electrical conductivity decreases.

At B: All OH- ions are neutralized. (Point of Neutralization) No H+ or OHions are present. Conductivity is not zero as Na+ and CI-ions are present.

From B - C: Conductivity increase as more H⁺ and Cl⁻ are added.

At A1: If a weak alkali or acid was used, the number of ions present will be fewer. Hence the initial conductivity will be fewer.

From A1 – B1: If a strong acid was titrated, this will introduce more H+ ions and thus the overall conductivity will increase till C1.





 $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + H_2O(I)$

At A: High conductivity as the $H^{\scriptscriptstyle +}$ and $SO_{4^{2-}}$ concentration is high

From A to B: Conductivity decreases as the ions are loss due to:

(a) H⁺ neutralised by OH- ions to form covalent water molecules

 H^+ (aq) + OH^- (aq) \rightarrow H_2O (I)

(b) $SO_{4^{2-}}$ and Ba^{2+} are removed to for insoluble BaSO4 (s)

 $Ba^{2+}(aq) + SO_{4^{2-}}(aq) \rightarrow BaSO_{4}(s)$

At B: All H⁺ ions are neutralized. (Point of Neutralization) No H⁺ or OH⁻ ions are present. Conductivity is zero as Ba²⁺ and SO₄²⁻ ions are present.

From B - C: Conductivity increase as more Ba²⁺ and OH⁻ are added.



Part III

Thermometric Titration



Acid – Alkali reactions are exothermic reaction

An exothermic reaction is one where **energy is released in the form of heat** causing the temperature of the reaction to increased.

At A: Temperature of the Alkali is at room temperature.

From A to B: Temperature is increasing as H⁺ react with OH⁻ ions

At B: Point of Neutralization is reached, no more OH- is present for reaction. All OH- ions has reacted. No further neutralization can occur. Hence no more heat is evolved. The temperature gradually decreased to room temperature.

Comparing Graph I and II

Half the volume of acid is needed to neutralise the alkali for Graph I than Graph II, indicating that the acid used in Graph I is a dibasic acid, containing twice the number of H⁺ ions.

Graph I – Dibasic Acid (1 mol/dm³) contain twice the H⁺ Graph II – Monobasic Acid (1 mol/dm³)



DXIDES

There are 4 types of oxides. They can be classified as follows

	Ionic compounds	Basic oxides (Bases)
	High melting/boiling pt	
Metal Oxides	Generally solid (rtp)	Amphataria avidas
	Mostly insoluble, if	Amphotene oxides
	soluble will form alkalis	
	Covalent compounds	Neutral oxides
Non Motol Ovideo	 low melting/boiling pt 	
Non Merai Oxides	 Generally gases (rtp) 	Acidic oxides
	Mostly soluble form acids	

Metal Oxides – Basic Oxides

Are oxides of metal	Basic oxides react with acids to form salt
• Eg. Na ₂ O, CaO, Fe ₂ O ₃ , CuO	and water
Are ionic compounds	CuO (s) + 2HCl (aq) → CuCl ₂ (aq) + H ₂ O (l)
High melting/boiling point	CaO (s) + 2HCl (aq) \rightarrow CaCl ₂ (aq) + H ₂ O (l)
Solid at room temp	
• MgO and Al ₂ O ₃ are have high	• Metal oxide for alkali metals can dissolve in
melting point and are used as	water to form soluble metal hydroxide
refractories in furnaces.	(alkali). The alkali will then have a reaction
	with pH indicators / litmus paper.
	□ Na ₂ O (s) + H ₂ O (aq) → 2NaOH (aq)
	□ $Li_2O(s) + H_2O(aq) \rightarrow 2LiOH(aq)$
	□ BaO (s) + H ₂ O (aq) \rightarrow Ba(OH) ₂ (aq)

Uses of Basic Oxides

- CaO (Lime) is used to neutralize acidic gases (SO₂, SO₃) released from power stations
- CaO is also used to neutralize / reduce the acidity of soil, caused by acid rain or excessive fertilizers. If the soil is too acidic, the plants are not able to absorb the nutrients well.

Metal Oxides – Amphoteric Oxides

	Are oxides of metal	•	Amphoteric oxides react with both acids and alkali to
	• Only ZnO, Al ₂ O ₃ , PbO		form salt and water
	Are ionic compounds		
	• High melting/boiling point		ZnO (s) + 2HCl (aq) → ZnCl2 (aq) + H2O (l)
	• Solid at room temp		Al_2O_3 (s) + 2HCl (aq) \rightarrow AlCl ₃ (aq) + H ₂ O (l)
	 Al₂O₃ has have high 		ZnO and Al_2O_3 are reacting as bases.
	melting point and is used		
	as refractories in		ZnO (s) + 2NaOH (aq) \rightarrow Na ₂ ZnO ₂ (aq) + H ₂ O (I)
	furnaces.		AI_2O_3 (s) + NaOH (aq) \rightarrow NaAlO ₂ (aq) + H ₂ O (l)
			ZnO and Al_2O_3 are reacting like acids.
		•	Amphoteric oxides are insoluble in water.
		•	Amphoteric oxides have both basic and acidic
			properties.
		•	They are soluble in both acids and alkalis to form salts.
_			



Non Metal Oxides – Acidic Oxides

Are oxides of non-metal	Acidic oxides dissolves in water to form acids.		
• Eg. SO ₂ , SO ₃ , CO ₂ , NO ₂ , P ₄ O ₁₀ (s)	These acids will then have an effect on pH		
Are covalent compounds	indicators and litmus paper.		
Low melting/boiling point	CO ₂ (g) + H ₂ O (I) → H ₂ CO ₃ (aq)		
Generally gases at room	SO2 (g) + H2O (I) → H2SO3 (aq)		
temp	SO_3 (g) + H ₂ O (I) \rightarrow H ₂ SO ₄ (aq)		
	React with alkalis to form salt and water		
	CO ₂ (g) + 2NaOH (aq) → Na ₂ CO ₃ (aq) + H ₂ O (I)		
	CO_2 (g) + Ca(OH) ₂ (aq) \rightarrow CaCO ₃ (s) + H ₂ O (I)		
	$SO_{2}(a) + 2NaOH(aa) \rightarrow Na_{2}SO_{2}(aa) + H_{2}O(l)$		
	$SO_2(g) + Ca(OH)_2(aa) \rightarrow CaSO_2(a) + H_2O(I)$		
	302(g) + Ca(CH)2(aq) + Ca(CS(3) + H2C(1))		
	O_2 and O_2 is acting as acids.		
Acid rain & its Effect			
Acid rain in industrial areas contai	ns H2SO3 (sulfurous acid) – Weak acid		
H ₂ SO ₃ –	$H_2SO_3 \xrightarrow{\text{Oxidized by } O_2 \text{ in air}} H_2SO_4$		
Sulfurous aci	d Sulfuric acid		
• This acid can be oxidised by oxygen in the air to form H_2SO_4 (sulfuric acid) – Strong			
acid. Sulfuric acid may corrode building and statues as it is a strong acid			

Non Metal Oxides – Neutral Oxides

•	Are oxides of non-metal	Are insoluble in water
•	Eg. CO, NO, H ₂ O	Are insoluble in acids and alkalis
•	Are covalent compounds	 Have no effect on pH indicators
	Low melting/boiling point	
	Generally gases at room temp	





SALT Solubility

To understand salt preparation, the solubility of salts must first be understood.

Type of Salts	Soluble	Insoluble
Nitrates (NO3-)	All nitrates are soluble	None
Chlorides (Cl ⁻)	Most chlorides are soluble	Only AgCl and PbCl ₂ are
Similar application for Br- and I-		insoluble
Sulfates (SO ₄ ²⁻)	Most sulfates are soluble	Only CaSO4, BaSO4, PbSO4
		are insoluble
Hydroxides (OH-)	Only Alkali are (aq)	All other bases
	Group I OH	
	 NH₄OH 	
	Group II OH (Last 4)	

Special consideration

Type of Salts	Soluble	Insoluble
General rules	All types of Na ⁺ , K ⁺ and NH ₄ ⁺	
SPA Salts	salts are soluble	
Carbonates (CO ₃ ²⁻)	Only Na+, K+ and NH4+ salts	All other carbonates are (s)
Hydrogen	All hydrogen carbonates are	
carbonate (HCO3-)	soluble	



SALTS PREPARAT^{ION}

A salt is defined as a compound formed when the replaceable hydrogen of an acid has been replaced by a metal / metallic radical (NH_4^+)

Eg.

Acid	Salt formed	Salt formed
HCI	NaCl	NH4CI
H_2SO_4	Na ₂ SO ₄	NaHSO _{4 (Acid Salt)}
H_2CO_3	Na ₂ CO ₃	NaHCO _{3 (Acid Salt)}
H ₃ PO ₄	Na ₃ PO ₄	NaH2PO4 (Acid Salt)

Acid salts

- Acid salts are compounds in which the replaceable hydrogen of an acid have been partially replaced by a metal.
- Acid salts still contain H⁺ ion and are acidic.
- Acid salts ionise to form H⁺ ions in aqueous solution
 Eg KHSO₄ → K⁺ + H⁺ + SO₄²⁻
- They are not acids as H⁺ ions are not the only positive ions released in the solution.

Preparation

Salts can be prepared by the following ways:

Mtd	Equation	Remarks
1	Acid + Metal → Salt + H2	Reactive metals from Group I should not be placed in acid as the reaction can be explosive/violent Unreactive metals like Cu, Ag, Au and Pt do not react with acid as they are below H in the reactivity series and cannot displace H.
2	Acid + Base → Salt + H ₂ O	A base is a metal oxide or hydroxide If a base is soluble in water it is known as an alkali
3	Acid + Carbonate \rightarrow Salt + H ₂ O + CO ₂	



The three methods of salt preparation are:

No.	Method of Preparation	When to apply?
1	Titration	Used to prepare Group I and Ammonium salts
		Eg.
		(a) KNO3 (b) CaCl2 (c) NH4Cl
2	Ionic Precipitation	Used to prepared insoluble salts like: (d) AgCl (e) PbCl ₂ (f) CaSO ₄ (g) BaSO ₄ (h) PbSO ₄
3	Adding excess solid to acid (Crystalization)	Used to prepare all other salts not in 1 and 2. Eg. (a) CuSO4 (b) FeCl3 (c) ZnSO4 (d) Mg(NO3)2



TITRAT^{ION}

Method involves the following reaction:

- (a) Alkali + Acid \rightarrow Salt + Water (Neutralization)
- (b) Aqueous Carbonate solution + Acid \rightarrow Salt + Water + Carbon dioxide

An indicator such as methyl orange, phenothalein is used

Objective: To find out the volume of acid/alkali required to neutralize a fixed volume of alkali/acid in a conical flask. End point is reached when acid/alkali is completely neutralized.



No.	Procedure	Remark
1	Pipette 20.0cm ³ of alkali (NaOH) into a conical flask	
2	Add 1-2 drops of methyl orange indicator	Methyl orange will turn yellow in alkaline solution
3	Titrate acid (H ₂ SO ₄) from the burette into the conical flask until the solution turns pale orange .	End point is reached when the indicator changes colour permanently.
4	Record the volume of acid used.	At this point the solution of salt obtained in impure due to the presence of an indicator



Describe either one method

No.	Procedure	Remark
5a	Now that the end point volume is known, the titration is repeated using the same volume of acid and alkali but without any indicator. A pure salt solution is obtained.	The salt solution obtained is then crystalized or evaporated to dryness (for NaCl only) The theoretical description of salt preparation via titration does not require the average reading to be calculated. The average readings is required for actual experimentation purpose for scientific accuracy.

<u>Or</u>

No.	Procedure	Remark
5b	Add activated charcoal to the mixture obtained in 3 and boil.	Activated charcoal will absorbed the indicator/dye (methyl orange)
6	Once the indicator is absorbed, filter the charcoal to obtain a filtrate containing pure salt solution.	

Once the pure salt solution is obtained, the following is applied.

Crystalization

No.	Procedure	Remark
1	Heat the solution until the solution becomes saturated.	To remove excess water so that crystals will form readily on cooling.
2	As the solution reaches saturation point, gentle heating is done.	Strong heating will dehydrate the crystals producing anhydrous salts. CUSO ₄ .5H ₂ O $\xrightarrow{\Delta}$ CUSO ₄ + 5H ₂ O (White) (Crystals lose water of crystallization on heating to form anhydrous crystals)
3	The solution is then left to cool	Crystals form because solubility decreases with decrease in temperature.
4	Filter off or decant crystals from solution	



5	Wash the crystals with a small amount of distilled water to remove any impurities or contaminants	Water is usually squirt on the crystals and allow to drain off.
6	Press dry the crystals between filter papers to obtain dry crystals.	

Or

Evaporation to Dryness (For NaCl)

No.	Procedure	Remark
1	For NaCl, the solution at endpoint is evaporate to dryness to obtain crystal of sodium chloride.	NaCl cannot crystalize as it solubility does not increase with a rise in temperature. Refer to the graph below for explanation.

Graph (Solubility Curves)



Explanation of Graph

Temperature	Remarks
At 30°C	$KNO_3 - 5g/100g H_2O$ of salt is dissolved.
	NaCl – 15g/100g H_2O of salt is dissolved.
At 80°C	As the temperature increases,
	$\rm KNO_3$ – 55g/100g $\rm H_2O$ of salt is dissolved. The increased in the mass of dissolved salt has increased largely compared to NaCl
	NaCl – only 20g/100g H ₂ O of salt is dissolved.



Cooling from 80°C to 30°C	If the both crystals are cooled from 80°C to 30°C, the mass of crystals that can be dissolved in for:
	 (a) KNO₃ - can be 5g/100g of water. This means that (55-5g) 50g/100g of water of KNO₃ will form crystals. (b) NaCl - few crystals will be formed as the mass of crystal forming will be only (20-15g) 5g/100g of water of NaCl will form crystals.
	Hence crystals of NaCl cannot be obtained by crystallization, as the solubility of NaCl does not increase much with rise in temperature.

Titration to form Acid Salt (Eg. NaHSO4, KHSO4)

No.	Procedure	Remark
1	Pipette 20.0 cm ³ of alkali, NaOH into a conical flask	
2	Add 1-2 drops of methyl orange.	Solution turns yellow
3	Add acid, H ₂ SO ₄ until the colour changes to pale orange (End point is reached).	
4	Record the volume of acid used. Eg 22.3 cm ³	
5	Repeat titration without using indicator using 20.0 cm ³ of NaOH (same volume of alkali) but add twice the volume of acid from Step 4. Eg 22.3 x 2 = 44.6 cm ³ Solution obtained contain pure acid salt eg NaHSO ₄ , KHSO ₄	

To form Ammonium salts (NH₄Cl, (NH₄)₂SO₄)

No.	Procedure	Remark
1	Titration of NH₄OH with acid	

Or

No.	Procedure	Remark
1	Bubble ammonia gas into acid until solution smells strongly of ammonia. Solution of the salt is obtained. Excess ammonia is boiled off during evaporation.	


IONIC PRECIPITAT

Method involves the following reaction:

- (a) Aqueous salt + Acid \rightarrow Insoluble Salt + Acid
- (b) Aqueous salt + Aqueous salt \rightarrow Insoluble Salt + Aqueous salt

When? - To obtain insoluble salts like

AgCl	Ca\$O4	PbSO₄
PbCl ₂	Ba\$O₄	Insoluble Carbonates

Theory

Insoluble salts above cannot be prepared by using a metal / metal oxide/ metal carbonates with acid. This is because when a solid is added to an acid, the insoluble salt formed will coat itself around the solid and prevent further reaction.

 $BaO + H_2SO_4 \rightarrow BaSO_4 + H_2O$

 $Ba + H_2SO_4 \rightarrow BaSO_4 + H_2$





H₂SO₄









Concept to Form Insoluble Salts

- Both chemical used must be aqueous solution (free ions)
- One must contain the required cation (positive/metal ion)
- The other must contain the required anion (negative ion)

Examples

To obtain PbSO4			
Pb ²⁺	\$O4 ²⁻		
From: Pb(NO ₃) ₂	From any of the following: CUSO4 K ₂ SO4 Na ₂ SO4 H ₂ SO4		



 $\mathsf{Pb}(\mathsf{NO}_3)_{2\,(\mathsf{aq})} + \mathsf{Na}_2\mathsf{SO}_{4\,(\mathsf{aq})} \boldsymbol{\rightarrow} \mathbf{PbSO}_{4\,(\mathsf{s})} + 2\mathsf{Na}\mathsf{NO}_{3\,(\mathsf{aq})}$

To obtain CaSO4			
Ca ²⁺ SO ₄ ²⁻			
From any of the following:	From any of the following:		
 CaCl₂ Ca(NO₃)₂ 	 MgSO4 Li2SO4 Na2SO4 H2SO4 		

To obtain AgCl			
Ag+ Cl-			
From: • AgNO3	From any of the following: • MgCl ₂ • HCl • NaCl		





 $\text{CaCl}_{2 \text{ (aq)}} + \text{H}_2\text{SO}_{4 \text{ (aq)}} \rightarrow \textbf{CaSO}_{4 \text{ (s)}} + 2\text{HCl}_{2 \text{ (aq)}}$



 $\mathsf{AgNO}_{3\,(\mathsf{aq})} + \mathsf{NaCl}_{\,(\mathsf{aq})} \rightarrow \mathsf{AgCl}_{(\mathsf{s})} + \mathsf{NaNO}_{3\,(\mathsf{aq})}$



To obtain CaCO3			
Ca ²⁺	CO32-		
From any of the following:	From any of the following:		
 CaCl₂ Ca(NO₃)₂ 	 MgCO₃ Li₂CO₃ Na₂CO₃ 		
	H ₂ CO ₃ cannot be used as a source of CO ₃ ²⁻ as this acid is too weak there is insufficient ions for reaction		



 $CaCl_{2 (aq)} + Na_{2}CO_{3 (aq)} \rightarrow CaCO_{3 (s)} + 2NaCl_{(aq)}$

No.	Procedure	Remark
1	Add together the two aqueous solution eg. Aqueous barium nitrate + aqueous sodium sulfate	Ba{NO ₃ } _{2 (aq)} + Na ₃ SO _{3 (aq)} → BaSO (s) + NaNO _{3 (aq)}
2	Stir and mix. Eg. Insoluble barium sulfate precipitate	
3	Filter off precipitate to obtain the insoluble salt eg. Barium sulfate.	
4	Wash the residue with distilled water.	To remove contaminants which may be from the excess unreacted barium nitrate or sodium sulfate.
5	Leave to dry between filter papers.	



Adding Excess SOLID to ACID

Method involves the following reaction:

- (a) Metal + Acid \rightarrow Salt + H₂
- (b) Metal oxide + Acid \rightarrow Salt + H₂O
- (c) Metal hydroxide + Acid \rightarrow Salt + H₂O
- (d) Metal carbonate + Acid \rightarrow Salt + H₂O + CO₂

This method is not used when the following salts are prepared:

- Group I salts
- Ammonium salts
- Insoluble salts (AgCl, PbCl₂, CaSO₄, BaSO₄, PbSO₄ or any insoluble carbonates)

(NEUTRALIZATION)

(NEUTRALIZATION)

No.	Procedure	Remark
1	To a fixed volume of dilute acid (HCI) in a beaker, add solid (CuO)	CuO + HCI →CuCl ₂ + H ₂ O
2	Warm the mixture	Warming increases the solubility. More crystals can be obtained giving better yield.
3	Stir the mixture	Increase the rate of dissolving
4	Add excess solid (CuO) until the solid is unable to dissolve.	This ensures that all the acid has reacted and will not remain to contaminate the salt
5	Filter to remove the excess solid (CuO)	
6	Heat the filtrate to obtain a saturated solution. As the solution reaches saturation point, gentle heating is done.	Strong heating will dehydrate the crystals producing anhydrous salt



			$\langle \rangle$
7	Leave aside to cool	Crystals will form as solubility decreases	
8	Filter off or decant the crystals from solution		
9	Wash with some distilled water to remove impurities		
10	Press dry between filter paper to obtain dry salt crystals		

SUMMARY

No.	Method of Preparation	When to apply?
1	Titration	Used to prepare Group I and Ammonium salts
	Acid + Alkali → Salt + Water	
2	Ionic Precipitation	Used to prepared insoluble salts like:
	2 Ways:	(a) AgCl
	(a) Aq Salt + Aq Salt	(b) $PbCl_2$ (c) $CqSO_4$
	(b) Aq Salt + Acid	(d) BaSO4
		(e) PbSO4
3	Adding excess solid to acid (Crystalization)	Used to prepare all other salts not in 1
	4 Ways:	and 2.
	(a) Metal +Acid	Eg.
	(b) Metal oxide + Acid	(a) CuCl ₂
	(c) Metal hydroxide + Acid	(b) FeSO4
	(d) Metal carbonate + Acid	



METALS

Structure of Metals

Metal atoms are packed very closely together in neat, regular rows in a definite pattern forming the metallic lattice. This lattice consists of positive ions in a "sea of delocalised" electrons. Metal atoms tend to expel its valence electrons to achieve a stable noble gas configuration (Octet) and form a positive cation.



A metallic bond is thus formed between these positive ions and the delocalised electrons between them.

Physical Properties	Property of metals	Explanations	Exceptions
Hardness	Very hard		Group I Alkali metals are soft and can be cut
Density	Dense		Group I Alkali metals such as Li, K and Na floats on water
Melting points	High mp	Strong metallic bonding between cations and delocalized electrons holds the metallic lattice together. A lot of energy is required to break this force. Alloys have lower mp than pure metal due to the presence of impurities.	Group I metals have low mp. Group I metals have 1 valence electron. They will tend to lose 1 electrons to form a +1 charged ion. Since every atom loses just 1 electron, this will contribute to a smaller electron pool compared to other metals. Thus the metallic bonding is weaker than other atoms.

Malleability	All metals are malleable	For pure metal, the atoms have the same size and are regularly arranged in neat rows. When a force is applied, the atoms in the rows can easily slip over each other and fit into new positions. Hence pure metals are very malleable.	
		Pure metal Regular layer slips easily	
		Alloy (Different sizes) Regular layers are disrupted	
		However for alloys, the sizes (bigger/smaller) of the atoms are different. The irregular sizes disrupt the regular layers results in a resistance created. When pressure /force is applied, making it difficult for atoms to slip. Hence alloys are less malleable and tougher / stronger compared to pure metals.	
Ductile	All metals are ductile	Rows of metallic atoms slide over each other when pulled, thus lengthening the metal.	
Conductors of Heat and Electricity	All metals are good conductors of heat and electricity	In a metal lattice, each atom loses its valence electrons to achieve a stable noble gas configuration. This give rise to pool of mobile delocalized electrons which can conduct electricity and heat. When a metal is connected to an electrical circuit, electrons are 'pushed into' the metal (from the negative terminal of the battery), and are 'pulled out" at the positive terminal. Thus a "sea of electrons" flow through the metal. As equal number of electrons are pushed in and pulled out, no chemical changes occur. When a metal is heated, the delocalized electrons gain energy and move faster. They will collide with each other and energy is transferred from on electron to another. Heat energy is thus passed along from one end of the metal to the other end.	



CHEMICAL PROPERTIES OF METALS

(a)	Metals have 1, 2 or 3 valence electrons and always lose electrons to form positive ions such as Na ⁺ , Ca ²⁺ , Al ³⁺ , depending on the number of electrons lost.			
	Down a group, the distance between the positive nucleus and valence shell increases as shown for Group II metals.			
	Beryllium	, Be Magnesium, Mg Ca	lcium, Ca Strontium, Sr	Barium, Ba
		G alka	roup 2, line earth metals	
	As such, metals ten	d to loose electrons	easier and form po	sifive ions easier down
	increases as the at	omic size increases	se by which merdin	orn positive ionj
(b)	Reaction with air:			
	(i) Corrosion			
	Most metals	may react very slow	ly with components	s in air (O2, CO2, H2O
	etc) at room temperature.			
	However, for reactive metals like K, Na and Ca, corrosion occurs readily.			
	Hence these must be stored in oil to prevent contact with air and water.			
	When Na reacts with air and water, Na2O and NaOH are formed			
	respectively. These alkaline compounds are able to absorb the CO_2			
	(acidic oxide) in the air to form sodium carbonate.			
	INC (react with O2	2/H2O) 7 INU2O/INUON	(react with CO2) 7 INU2C	\bigcup_3
	(ii) Burning (Hec	ating is required)		
	Metal such a	as K, Na (Group I) ar	nd Ca, Ba and Mg (Group II) burns to form
	oxide. When	they are burn, a fla	me with the followin	ng characteristic colour
	is seen.			
	Metal	Colour of flame	Chemical Fauatio	n
	K	Lilac	$2K + O_2 \rightarrow K_2O$	
	Na	Yellow	$2Na + O_2 \rightarrow Na_2O$	
	Ca	Red	$2Ca + O_2 \rightarrow 2CaC$)
	Mg	Dazzling white	$2Mg + O_2 \rightarrow 2MgC$)
	The evidenter	mod are basic ovideo	which react with act	tc) The evider that are
	soluble in water will form alkalis (soluble metal hydroxides)			



(c)	Reaction with acids
	Metal + Acid \rightarrow Salt + H ₂
	(Metal below H in the reactivity series do not react with acid)
	Refer to the reactivity series for their reaction.
	Note: Certain metals like Ca, Pb may formed insoluble layer of salt when reacting
	to acids like HCI or H ₂ SO ₄ (refer to notes on insoluble salts Pg57)
(d)	Reaction with water/steam
	Reactive metals can react violently with both cold water and steam to form
	alkali and hydrogen gas. Metal + H ₂ O (I) \rightarrow Metal hydroxide + H ₂
	Less reactive metals can react only with steam and not with cold water, a metal
	oxide and hydrogen gas is formed. Metal + H_2O (g) \rightarrow Metal oxide + H_2
	Refer to the reactivity series for their reaction.
(e)	Displacement Reaction
	A more reactive metal can displace a less reactive metal from its salt or from its oxide.

COLOURS OF METAL/SALTS

No.	Substance	Colour
1	Cu	Reddish Brown
2	CuO	Black
3	CuCO ₃	Green
4	Cu ²⁺ solution (aq)	Blue
5	Fe	Grey
6	Fe ²⁺ solution (aq)	Pale Green
7	Fe(OH) ₂ (s)	Green
8	Fe ³⁺ solution (aq)	Orange/Yellow
9	Fe(OH)₃ (s)	Brown

Note:

Aqeuous salt (Not in the table above) tends to be **colourless**. Insoluble salt (Not in the table above) tend to be **white**.



REACTIVITY SERIES OF METALS

Metals can be arranged in order of increasing or decreasing reactivity / electropositivity. A reactivity list is used to predict how metals react with water and acids.

Metals	Reaction with water and steam	Reaction with acid
К	Very reactive metal. Violent reaction with both cold water and steam to form soluble metal	Very reactive metal. Violent reaction with dilute
Na	hydroxides (alkali)	acid.
Ca	Metal + H ₂ O (I) \rightarrow Metal hydroxide + H ₂ (Violent Rxn)	Metal + Acid \rightarrow Salt + H ₂ (Violent Rxn)
	Metal + H ₂ O (g) \rightarrow Metal hydroxide + H ₂ (Violent Rxn)	
Mg	Very slow reaction with cold water Metal + H2O (I) → Metal oxide + H2 (very slow)	Reaction becomes less vigorous down the series.
	Reaction with steam to form metal oxide Metal + H₂O (g) →Metal oxide + H₂	Metal + Acia -> Sali + H2
AI	No reaction with cold water.	
Zn	Reaction with steam to form metal oxide. Metal + H ₂ O (g) \rightarrow Metal oxide + H ₂	
re	Note: Aluminium does not react with both steam and water as a thin insoluble layer of Al_2O_3 will always form for AI thus preventing pure AI from reacting.	
Pb	Very slow reaction with steam Metal + H₂O (g) →Metal oxide + H₂ (very slow)	
Н		
Cu Ag/Au	Unreactive metal, no reaction with both cold	No reaction with dilute acid
Pt	water and steam	



APPLICATIONS

No.	Reaction	Equation	Observation
1	Magnesium is reacted with dilute hydrochloric acid	Mg(s) + 2HCl (aq) → MgCl ₂ (aq) + H ₂ (g)	 Vigorous bubbling Silvery metal of Mg dissolves in acid to form a colourless solution MgCl₂ Hydrogen gas is produced (Pop sound with lighted splint) Test tube feels warm (Exothermic rxn)
2	Zinc is reacted with dilute hydrochloric acid	Zn(s) + H2SO4 (aq) → ZnSO4 (aq) + H2 (g)	 Slight bubbling Slow reaction Grey metal of Zn dissolves in acid to form a colourless solution ZnSO4 Hydrogen gas is produced (Pop sound with lighted splint)
3	Iron with dilute sulfuric acid	Fe(s) + H ₂ SO ₄ (aq) → FeSO ₄ (aq) + H ₂ (g)	 Reaction is slow Very slow bubbling is observed Grey metal dissolves to form a pale green solution FeSO₄ (Similar for FeCl₂) Hydrogen gas is produced.
4	Calcium with dilute sulfuric acid	Ca(s) + H2SO4 (aq) → CaSO4 (s) + H2 (g)	 Reaction is fast in the beginning Vigorous bubbling at the start, slows down and stop. Calcium remain undissolved and turns white. Explanation: Ca is a reactive metal, hence there is a vigorous rxn at the start. However when insoluble layer of CaSO₄ coats around Ca, this slows down and eventually stops the rxn
5	Potassium with water	2K(s) + 2H ₂ O (I) → 2KOH (aq) + H ₂ (g)	 Vigorous reaction Potassium darts about on surface of water (Less dense). It melts into a silvery bead and then disappears. Temperature of reaction rises



6	Calcium with	$2Ca(s) + 2H_2O(I) \rightarrow Ca(OH)_2(aq) + H_2(g)$	٠	Less vigorous compared to Na
	water	Ca(OH) ₂ (s) Calcium hydroxide is slightly soluble in water	•	Vigorous bubbling of H ₂ observed. White ppt is formed around Calcium due to the formation of Calcium hydroxide which is only slightly soluble in water. Reaction is stop when a coating of insoluble calcium hydroxide (slaked lime forms around calcium)
1	1		1	

REACTION OF STEAM & WATER WITH METALS





DISPLACEMENT REACTIONS

Displacement reaction for metals can occur in 2 ways via:

No.	Reaction	Equation
1	Displacement	$Zn(s) + CuO(s) \rightarrow ZnO(s) + Cu(s)$
	from oxides	ZnO is an amphoteric oxide that is yellow when hot and white when cold
	The oxides (reactants) must	2AI (s) + 3ZnO (s) \rightarrow 2AI ₂ O ₃ (s) + 3Zn (I)
	heated so that a reaction will	2AI (s) + Fe ₂ O ₃ (s) \rightarrow Al ₂ O ₃ (s) + 2Fe (I)
	occur. This reactions are	The temperature for the displacement of Zn/Fe is so high that Zn/Fe melts when formed.
	exothermic and releases a lot of	Thermite Reaction (Application) – Repair crack railway tracks.
	near	2AI (s) + Fe ₂ O ₃ (s) \rightarrow AI ₂ O ₃ (s) + 2Fe (I)
		A mixture of aluminium and iron (III) oxide powder can be pour into cracks in railway tracks, A piece of Magnesium is used as a fuse and the magnesium is lighted. The heated Magnesium will provide a high temperature needed to start the reaction.
		This results in the more reactive AI displacing Fe from Fe_2O_3 . The reaction is strongly exothermic resulting in a lot of heat produced that can cause the iron displaced to melt. Molten iron then deposit and solidify in the crack.
2	Displacement	$Zn(s) + CuSO_4 (aq) \rightarrow ZnSO_4 (aq) + Cu (s)$
	from its salt solution (aq)	Observation: The blue solution of CuSO₄ fades and a colourless solution of ZnSO₄ is formed. A reddish brown solid, Cu is deposited.
		Explanation: Zinc is more electropositive than copper. Zinc atoms lose electrons more readily to Cu^{2+} . The Cu^{2+} ions take up the 2 electrons and is reduced to Cu atom. Half Equation: $Cu^{2+} + 2e \rightarrow Cu$ As the concentration of Cu^{2+} decreases, the blue colour of the solution fades . A reddish brown copper deposits.
		$Fe(s) + CuSO_4 (aq) \rightarrow FeSO_4 (aq) + Cu (s)$
		Observation: The blue solution of CuSO4 fades and a pale green solution of FeSO4 is formed. A reddish brown solid, Cu is deposited.
		Explanation: Iron is more electropositive than copper. Iron atoms lose electrons more readily to Cu^{2+} . As the concentration of Fe ²⁺ increases in the solution, the solution gradually turns green . The Cu^{2+} ions take up the 2 electrons and is reduced to Cu atom. Half Equation: $Cu^{2+} + 2e \rightarrow Cu$ As the concentration of Cu^{2+} decreases, the blue colour of the solution fades. A reddish brown copper deposits.
		$Mg(s) + Pb(NO_3)_2 (aq) \rightarrow Mg(NO_3)_2 (aq) + Pb (s)$
		Observation: The solution remains colourless as Pb(NO ₃) ₂ and Mg(NO ₃) ₂) are colourless. Shiny black crystals of Pb is formed.



STABILITY OF METAL OXIDES

The more reactive the metal, the stronger the bonds formed in its compound. Thus the more stable its compound is.

Metals		Oxides	
К	(a) Action of heat	(b) Reduction of oxides (Removal of oxygen atom)	(b) Reduction of oxides (Removal of oxygen atom)
Να	All metal oxides are stable except oxides of metal	By a displacement reaction via	Heating with reducing gases such as:
Ca	below Copper (which is less electropositive)	heating a more reactive metal with any less reactive metal oxide.	 Coke, C CO
Mg	Example: K2O, CuO are stable oxide.	These reaction are exothermic and a lot of heat is released. Metals	 H₂ NH₃
AI		displaced are usually molten. Example: Thermite Reaction	Oxides of reactive metals such from K – Al cannot be reduced by this method.
		2AI (s) + Fe ₂ O ₃ (s) → AI ₂ O ₃ (s) + 2Fe (I)	Al ₂ O ₃ cannot be reduced by C, CO or H ₂ but can be reduced by heating with a displacement with a more reactive metal. $3Mg(s) + Al_2O_3(s) \rightarrow 3MgO(s) + 2AI(I)$
Zn			Less reactive metal can be reduced by
Fe			these gases as shown.
Ph			• Heating with carbon/coke $2nO + C \rightarrow 2n + COc$
			$CUO + C \rightarrow CU + CO_2$
Cu			• Heating with H ₂ (In has no rxn with H ₂) PbO + H ₂ \rightarrow Pb + H ₂ O CuO + H ₂ \rightarrow Cu + H ₂ O • Heating with carbon monoxide Fe ₂ O ₃ + 3CO \rightarrow 2Fe + 3CO ₂ CuO + CO \rightarrow Cu + CO ₂
Hg	Metal oxides in this section		
Ag	can decompose to their pure metal on heating		
Pt	Example:		
	$2Ag_2O + heat \rightarrow 4Ag + O_2$ $2HgO + heat \rightarrow 2Hg + O_2$ (Red solid) (Silvery liquid)		



REACTION OF HYDROGEN GAS WITH METAL OXIDES



SOLUBILITY OF METAL OXIDES

Insoluble in water	Soluble in water
Most oxides are insoluble in water	All Group I Oxides are soluble in water. They dissolves in water to form alkalis.
	Na2O (s) + H2O (I) → 2NaOH (aq)
	$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$
	However for Group II oxides, CaO is slightly soluble, MgO is very slightly soluble.
	 (a) CaO (s) + a few drop of H₂O (I) → Ca(OH)₂ (s) – Rxn is strongly exothermic, solid feel hot and steam maybe evolved. (b) CaO (s) + a large amount of H₂O (I) → Ca(OH)₂ (aq) Limewater is formed.



STABILITY OF METAL HYDROXIDES

Upon heating, most metal hydroxides decompose to a metal oxide and steam as shown below (except NaOH, KOH and **Group I hydroxides which do not decompose as they are very stable**)

Metal Hydroxides + Heat \rightarrow Metal oxide + Steam

 $Ca(OH)_2$ (s) + Heat \rightarrow CaO (s) + H₂O (g)

 $\begin{array}{c} \mathsf{Cu}(\mathsf{OH})_2 \text{ (s) } + \mathsf{Heat} \rightarrow \mathsf{CuO} \text{ (s) } + \mathsf{H}_2\mathsf{O} \text{ (g)} \\ \\ \mathsf{Blue} \text{ ppt} & \mathsf{Black} \text{ solid} \end{array}$

For less reactive metal (below Copper), their hydroxides decomposes to pure metal, oxygen and steam as their oxides are unstable to heat.

AgOH + Heat \rightarrow Ag₂O (Unstable) + H₂O

Further decomposition: Ag₂O (Unstable) + Heat \rightarrow Ag + O₂

 $Hg(OH)_2 + Heat \rightarrow HgO (Unstable) + H_2O$

Further decomposition: HgO (Unstable) + Heat \rightarrow Hg + O₂

SOLUBILITY OF METAL HYDRDXIDES

Metal Hydroxides that are soluble in water are known as alkalis. Refer to notes on Alkalis for solubility.



STABILITY OF METAL CARBONATES

Upon heating, most metal carbonates decompose to form metal oxides and carbon dioxide gas. (Group I metal are very reactive and their metal carbonates are stable. Group I metal carbonates do not decompose on heating)

Metal carbonate + Heat \rightarrow Metal oxide + Carbon dioxide

 $ZnCO_3 + Heat \rightarrow ZnO + CO_2$ White (Hot - Yellow; Cool - White)

 $CuCO_3 + Heat \rightarrow CuO + CO_2$ Green Black

Metal that are below Copper in the reactivity series, will decompose further due to its unstability. It will decompose to the pure metal and oxygen.

Metal carbonate + Heat → Metal Oxide (Unstable) + Carbon dioxide **Further Decomposition:** Metal Oxide (Unstable) + Heat → Metal + Oxygen

 $Ag_2CO_3 + Heat \rightarrow Ag_2O$ (Unstable) + CO_2

Further decomposition: Ag₂O (Unstable) + Heat \rightarrow Ag + O₂

Ammonium carbonate does not leave a solid residue during thermal decomposition as all the products formed are gaseous.

 $(NH_4)_2CO_3$ (s) + Heat \rightarrow NH₃ (g) + CO₂ (g) + H₂O (g) White crystals

SOLUBILITY OF METAL CARBONATES

Only Sodium, Potassium and Ammonium carbonates are soluble in water (SPA salts), the rest of the metal carbonates are insoluble.



SUMMARY OF STABILITY & SOLUBILITY

Metals	Stability of oxides	Solubility of oxides
K	Stable – Cannot decompose further on	Soluble in water to form soluble metal
Να	neating	nydroxides (aikalis)
Ca		Slightly soluble in water
Mg		Very slightly soluble in water
Al		Insoluble in water
Zn		
Fe		
Pb		
Cu		
Hg	Unstable – Decomposes to Pure metal and	
Ag	oxygen gas on nealing.	
Pt	1	

Metals	Stability of Carbonates	Solubility of Carbonates
K	Stable – Do not decompose on heating.	Only Sodium, Potassium and
Na		Ammonium carbonates are soluble in water
Ca	Decomposes to metal oxide and carbon	Most of metal carbonates are
Mg		
AI		
Zn		
Fe		
Pb		
Cu		
Ag	Decomposes to metal oxide (unstable) and	
Au	metal oxide formed, decomposes further to metal and oxygen gas.	



Metals	Nitrates	Carbonates
К	Metal Nitrate $ ightarrow$ Metal nitrite + Oxygen	Stable and unaffected by heat
Na	$2NaNO_3$ (s) $\rightarrow 2NaNO_2$ (s) + O_2 (g)	No decomposition
Ca	Metal nitrate \rightarrow Metal oxide + O ₂ (g) + NO ₂ (g)	Metal carbonate \rightarrow Metal oxide + CO ₂ (g)
Mg	$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + O_2(g) + 4NO_2(g)$	$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$
AI		
Zn		
Fe		
Pb		
Cu		
Ag	Metal nitrate \rightarrow Metal + O ₂ (g) + NO ₂ (g)	Metal carbonate \rightarrow Metal oxide + O ₂ (g) + CO ₂
Au	$2AgNO_3(s) \rightarrow 2Ag(s) + O_2(g) + 2NO_2(g)$	Ag ₂ CO ₃ (s) \rightarrow 2Ag (s) + O ₂ (g) + CO ₂ (g) Decomposes to metal oxide (unstable) and
		carbon dioxide on heating. The unstable metal oxide formed, decomposes further to metal and oxygen gas.

Metals	Hydroxides
К	No decomposition
Na	
Ca	Metal hydroxide \rightarrow Metal oxide + H ₂ O (g)
Mg	$Mg(OH)_2(s) \rightarrow MgO(s) + H_2O(g)$
Al	
Zn	
Fe	
Pb	
Cu	
Ag	Metal hydroxides do not exist
Au	



Metals oxide	Reduction with Carbon	Reduction with hydrogen
K ₂ O	Oxides are not reduced by Carbon	Oxides are not reduced by Hydrogen
Na ₂ O	K_2O (s) + C (s) \rightarrow No Reaction	K_2O (s) + H_2 (g) → No Reaction
CaO		
MgO		
Al ₂ O ₃		
ZnO	Oxides reduced by heating with	
FeO	$7nO(s) + C(s) \rightarrow 7n(s) + CO(a)$	Oxides are reduced by heating with
PbO		$PbO_{(s)} + H_{2}(a) \rightarrow Pb_{(s)} + H_{2}O_{(a)}$
CuO		
Ag ₂ O	Oxides decomposed to metal by heating only	
	$2Ag_2O(s) \rightarrow 4Ag(s) + 2O_2(g)$	



METAL EXTRACTION

- A pure metal can be extracted from it ore/oxide via two methods:
- (a) Chemical Reduction of the metal oxide using reducing agents like Coke (C) or carbon monoxide gas
- (b) Electrolytic reduction/Electrolysis via its molten chloride or oxide. This method is used for metals whose oxides cannot be reduced by C or CO. Stable oxides which cannot be reduced by chemical reduction are Group I oxides like K, Na; Group II oxides like Ca, Mg and Group III oxides like AI. These are reactive metals.

Metals	Method of extraction	
К	Extracted via electrolytic reduction of Chlorides or Oxides	
Na		
Ca		
Mg		
AI		
Zn	Extraction by chemical reduction using C/CO	
Fe	Stage 1: Roasting of Ore	
Pb	The metal ore is strongly heated in a stream of air. This converts any metal sulfides and carbonate to a metal oxide which can then be reduced by C/CO	
Cu	2ZnS (s) + 3O ₂ (g) → 2ZnO (s) + 2SO ₂ (g) [Oxidation]	
	$CuCO_3$ (s) + Heat \rightarrow CuO (s) + CO ₂ (g) [Decomposition]	
	Stage 2: Reduction of the Oxide	
	$ZnO + C \rightarrow Zn + CO$	
	$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$	
	Stage 3: Purification of extracted metal	
	Impurities such as Sand and Silica can be removed by heating with limestone, CaCO3 (Refer to Blast Furnace)	
Hg	Unreactive metal, tend to be found uncombined often as "free metal"	
Ag		
Pt		

BLAST FURNACE EXTRACTION OF IRON

Raw materials:

Haematite, Fe₂O₃ Coke, C Limestone, CaCO₃

Main impurities:Silicon dioxide (Sand), Basic in natureWaste gases:CO2, N2 (Unreacted component of air blown into the furnace)

No.	Stages of Extraction	Chemical Equation	
1	Coke is burn in hot air (blown from the base of the furnace) Reaction is strongly exothermic and a high temperature of 1900°C is produced at the base of the furnace	$C + O_2 \rightarrow CO_2$	NO and NO2 are also produced due to the high temperature in the furnace Coke is used to produce the reducing gas CO
2	Carbon dioxide formed is reduced by hot coke to form the reducing gas CO	$CO_2 + C \rightarrow 2CO$	Coke is also a fuel to burn and maintain the high temperature of the furnace.
3	Haematite, Fe ₂ O ₃ is reduced by carbon monoxide to form molten iron and carbon dioxide	$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$	
4	Main impurities is Silicon dioxide (Sand). This is removed in a 2 stage process		
4a	Limestone is decomposed by heating to produce lime, CaO and carbon dioxide	CaCO ₃ → CaO + CO ₂	Limestone is used to produce lime, CaO which removes the impurity.
4b	The basic lime, CaO reacts with acidic silicon dioxide (sand) to form a layer of calcium silicate which is removed as slag. This is a neutralization reaction.	CaO + SiO₂ → CaSiO₃	The molten slag floats on the molten iron and is skimmed off. A thin layer is usually left behind to prevent the contact with air and re-oxidation of the hot molten iron. This iron is known as pig iron.



BLAST FURNACE



Diagram taken from <<u>https://www.bestchoice.net.nz/chemistry/1305/p21243.htm</u>>

This iron is known as pig iron.

oxidation of the hot molten

contact with air and re-

iron.

Oxygen



STEE MANUFACTURE PROCESS

Steel is manufactured from pig iron via the following process:

(a)	Oxygen is blown into the molten pig iron to oxidise the non- metallic impurities such as C and Si to their oxides. C (s) + O ₂ (g) \rightarrow CO ₂ (g) Si (s) + O ₂ (g) \rightarrow SiO ₂ (s)	verter pure Iron
(b)	Limestone, CaCO ₃ is added. CaCO ₃ decomposes to produce lime, CaO, which is basic and reacts with the SiO ₂ (acidic) to form slag, CaSiO ₃ . Slag can be removed. $CaO (s) + SiO_2 (s) \rightarrow CaSiO_3 (I)$ When the slag is tapped off or removed, wrought iron (pure iron) is left. Wrought Iron is soft and malleable and has little tensile strength. It is used to make gates and chains.	Limestone added
(c)	 The mass of Carbon is then calculated and added to form steel. Steel can be classified into: (a) Mild steel (low carbon content – up to 0.25% C) is strong but malleable. It is used to make car bodies and machinery. (b) High carbon steel (from 0.5 – 1.5% C) is strong but brittle (not malleable). It is used to make knives, hammer and drills 	Pure Iron is poured out
(d)	Other elements such as Cr, Ni may be added to form different types of steel (Stainless steel)	Carbon & other elements



ALLOYS

PROPERTIES & USAGE

No.	Question	Explanation	
(a)	Is the entire car body made of mild steel? Why / Why not?	No, iron has a high density. The car will be heavy and consumes a lot of fuel due to it mass.	
		Thus alloys of aluminium are used as they are stronger (but not as strong as steel) and are less dense. Fuel consumption will be more efficient.	
(b)	Why are kitchen cutlery such as knives made of stainless steel?	Stainless steel is strong and does not corrode.	
(C)	Why is pure iron/cast or pig iron not used to make an iron bridge?	Both have no tensile strength. Pure iron is too soft. Cast iron is too brittle.	
(d)	What is brass made up of? What is it used for?	Brass is an alloy consisting of zinc and copper. It is used as a decorative ornaments, musical instruments and coins.	
(e)	Why metals are often used in the form of alloys?	 Alloys are harder and stronger eg. Steel is stronger than pure iron Alloys are more resistant to corrosion eg. Copper corrodes faster than brass. Alloys have a lower melting point eg. Solder has a lower mp than both pure lead and tin. Alloys can be more attractive eg. Brass is used as a decorative metal. 	



RUSTING **CORROSION OF IRON**

Definition:

- (a) Corrosion: gradual destruction of metal due to the reaction with air, water or chemicals in the air.
- (b) Rusting is the corrosion of iron in the presence of air and water.



(Hydrated Iron (III) oxide ~ Rust)

Experiment to show the conditions required for rusting



(a) Only air: No rusting

(b) Only water: No rusting

(c) Air and water together: rusting takes place

Test Tube	Condition	Explanation
Α	Anhydrous calcium chloride (Drying	No water is present, rusting will not take
	agent) is present.	place
В	Water is boiled to remove oxygen in it,	No oxygen/air is present, rusting will not
	and a layer of oil is placed over water	take place
	to keep air out.	
С	No drying agent or layer of oil. It	Presence of both water and oxygen,
	functions as a control experiment.	rusting occurs.
D	Sodium chloride or seawater is used.	The iron rusts fastest as the ions, Na+ and
	Contains Na ⁺ and Cl ⁻ ions.	Cl- in the water catalyse (speed up) the
		rusting process.

Note: Fe must be oxidized to Fe²⁺, before rusting can take place. When this occur, the reaction cannot be stopped as long as Fe is in contact with air. The faster the Fe²⁺ forms, the faster the rate of rusting.

2 electrons lost

 $\xrightarrow{\text{to }O_2 \text{ in air}} Fe^{2*} \xrightarrow{H_2O} Fe (OH)_2 \xrightarrow{\text{Air}} Fe(OH)_3$ Fe Fe₂O₃.xH₂O

Rust Hydrated Iron (III) oxide



APPLICATION

In which experiment will Fe rust faster? Predict and explain the results in the following setups.



Fe will in Setup 2 will be rusting faster than Setup 1. This is because Fe (being more electropositive) will lose electrons readily to copper than to oxygen in the air. This results in Fe forming Fe²⁺ faster, resulting in a faster rate of rusting.

PREVENTION OF RUSTING

No.	Method of rust prevention	Explanation
(a)	Coating the Fe with a layer of:	
	 Paint, oil or grease Another less reactive metal (below Fe in the reactivity series) 	To prevent iron from coming into contact with water and air. When either factor is not present, rusting is prevented. When the coating is broken, the iron will rust.
(b)	Galvanization – Coating with Zinc (more reactive than Fe) Without Zn,	The zinc protects the iron even when the layer is scratched. As long as there is some zinc in contact with iron, electrons will be transferred to iron, causing zinc to become Zn^{2+} ($Zn - 2e \rightarrow Zn^{2+}$).
	$Fe - 2e \rightarrow Fe^{2+}$ (lost to oxygen in air)	This prevents the 1 st stage of rusting (Fe – 2e \rightarrow Fe ²⁺), thus rusting cannot take place.
	With Zn, $\mathbf{r}^{Fe} - 2e \rightarrow Fe^{2+}$ Zn - 2e \rightarrow Zn ²⁺ Sacrificial protection - Rxn in which a	The electrons passed from zinc to iron prevents the formation of Fe ²⁺ . Fe loses electrons to the air are thus replaced by electrons from zinc.
	more reactive metal corrodes in the place of the less reactive metal it is in contact with. (Zn and Mg are used)	Zinc metal corrodes instead of iron. This is knowns as sacrificial protection.



APPLICATIONS



State the advantages and disadvantages of coating food cans with (a) tin and (b) zinc

(a) Using tin,

Advantage: Tin (Sn) is less reactive than Zinc and does not react readily with acidic and alkaline substances in food. Zinc dissolves more readily forming Zn²⁺ (which in large quantity can be poisonous)

Disadvantage: Tin is less electropositive than Iron. Electrons flow from Iron to Tin very readily. Therefore, Fe²⁺ forms more readily than when there is no tin present. Fe rust readily if tin is scratched off.

(b) Using zinc,

Advantage: Zinc offers sacrificial protection and will corrode in place of iron even when scratched off.

UNDERGROUND IRON PIPES

To prevent such pipes from rusting, bags of magnesium are placed in holes in the ground (at ground level) and connected to the pipe with a wire. The magnesium which is more reactive ionises.

Mg – 2e → Mg²⁺

The electrons are transferred via the wire to the iron thus replacing the electrons lost by iron to the air. This reduces the formation of Fe²⁺ and prevents rusting.

The magnesium which is slowly corroded can be replaced easily which a new block/bag easily (since it is at ground level).





Block of zinc (sacrificial metal)

CORROSION OF SHIPS

Ships are made of iron. To slow down rusting, blocks of zinc are attached to the sides of ships. These will act as sacrificial protection and corrode in place of iron.



RECYCLING

Metals are finite resources – there in only a limited amount of metal that can be mined from the earth. The only way to make metals last longer is to recycle them.

RECYCLING

Recycling is a process of collecting and processing materials that would be otherwise be thrown away as trash and turning them into new products.

Metal objects that unwanted and discarded, are crushed and then melted to produce blocks of clean metal. Lead, iron and aluminium are mainly recovered by scrap metal recycling.



Lead is recovered from car batteries.

Iron and steel are largely produced from recycled scrap metals.

Aluminium is recycled from drink cans and food containers.

ADVANTAGES OF RECYCLING

- Better conservation of natural resources (reserves last longer)
- Higher savings of energy and financial resources. Recycling saves the cost of extracting new metals from ore.

Eg. Recycling aluminium saves 95% the cost of extracting pure aluminium from its oxide, bauxite as a lot of electricity is needed to electrolyse aluminium oxide (bauxite). Producing Al from scrap Al requires 70-80% less energy compared to obtaining Al from bauxite via electrolysis

• More effective waste disposal. Fewer landfills are required to dispose waste metals. Land can be save for other uses (housing, commercial), cost of building landfill sites is also reduced.



Bauxite (Mineral of Aluminium oxide)





•Litter accumulation and pollution from abandoned metal objects (slowly corroded by air and water) is solved. Metals leach into the soil and eventually into rivers water causing contamination.

•Mining of metal destroys the natural environment. It produces a large amount of waste material and causes landslides due to soil erosion. Water pollution due to sediments washed into waterways may also occur.

DISADVANTAGES OF RECYCLING

- The smelting process used to melt old batteries (to extract lead) can produce a lot of metal fumes into the air.
- The valuable metal from abandoned objects are usually extracted leaving less valuable metal unwanted and thrown away, leading to accumulation of unwanted metal.







Landfills of waste.



MOLE CONCEPT

Relative Molecular Mass (M_r) is defined as the number of times one molecules of an element or compound is heavier than $\frac{1}{12}$ of an atom of $_{12}$ C.

To calculate Mr:

Step 1: Write the chemical formula of the compound Step 2: Find the sum total of the mass (Ar) of all the individual atoms in the molecule.

Name	Chemical Formula	Molecular Mass (Mr)
Sodium sulfate	Na ₂ SO ₄	$M_r = 23 + 23 + 32 + (16x4) = 142$ (no unit)
Copper (II) nitrate	Cu(NO3) 2	$M_r = 64 + 2 \times (14 + 18 \times 3) = 188$
Hydrated Iron (II) sulfate	FeSO ₄ .7H ₂ O	$M_r = 56 + 32 + 4 \times 16 + 7 \times (1 + 1 + 16) = 278$
(Iron (II) sulfate with water of crystallization)		

PRACTICE

Question 1:

Metal M forms a sulfate of formula MSO₄ which has M_r of 136

(a) Calculate the relative atomic mass, Ar of M.	[40]
(b) Suggest the identity of M (from the periodic table)	[Ca]

Question 2:

Element E forms a compound of formula E_2CO_3 which has M_r of 74

- (a) Calculate the relative atomic mass, Ar of E. [7]
- (b) Suggest the identity of E (from the periodic table) [Li]



PERCENTAGE & MASS COMPOSITION

Percentage (of an element in a compound)

% of an element = $\frac{mass of element}{Mr of compound} x 100\%$

Examples

% of O in Na₂SO₄ =
$$\frac{4 \times 16}{23 \times 2 + 32 + 16 \times 4} \times 100\%$$

= <u>45.1 %</u> (All answers to 3 sig fig)

% of N in (NH₄)₂CO₃ =
$$\frac{14 x 2}{2 x (14+4)+12+16 x 3} x 100\%$$

= 29.2% (All answers to 3 sig fig)

% of H₂O in CuSO₄.5H₂O =
$$\frac{5 x (16+2)}{64+32+(16 x 4)+5(16+2)} x 100\%$$

= 36.0% (All answers to 3 sig fig)

Mass (of an element)

Calculation of the mass of an element, given the mass of a compound

Examples

Calculate the mass of iron in 50g of iron (III) sulfate

 M_r of $Fe_2(SO_4)_3 = 56 \times 2 + 3 \times (32 + 16 \times 4) = 400$

Since Mr has no unit, any unit of mass can be assigned to it

400g of Fe₂(SO₄)₃: 112 g (2 x 56g) of Fe

Hence, 50g of Fe₂(SO₄)₃ = $\frac{112}{400}x$ 50g = 14.0 g of Fe (All answers to 3 sig fig)

Calculate the mass of copper in 1000 tonnes of copper (II) chloride

 M_r of $CuCl_2 = 64 + 2 \times 35.5 = 135$

Since Mr has no unit, any unit of mass can be assigned to it

135 tonnes of CuCl₂: 64 tonnes of Cu

Hence, 1000 tonnes of $CuCl_2 = \frac{64}{135}x \ 1000 \ tonnes = \frac{474.1 \ tonnes \ of \ Cu}{(All \ answers \ to \ 3 \ sig \ fig)}$



AVOGADRO'S CONSTANT – MOLE – MASS

One mole of any substance contain 6.022×10^{23} of particles.

$6.022{\times}10^{23}\,\text{is}$ known as the Avogadro's constant, N_{A}

The mass of one mole of a substance is referred to as the **molar mass**. The molar mass of any element or compound is the sum total mass of the atom(s) in the element or compound in grams.

Substance	Mole	Number of particles	Mass
Mg	1	6.022×10 ²³ Mg atoms	Mass of 1 mole Mg = 24 g
H ₂ O	1	2 x 6.022×10 ²³ H atoms	Mass of 1 mole H ₂ O
		6.022×10 ²³ O atoms	= 1 + 1 + 16 = 18 g
		Or	
		6.022×10 ²³ H ₂ O molecules	
Zn(NO ₃) ₂	1	2 x 6.022×10 ²³ NO ₃ ⁻ ions	Mass of 1 mole Zn(NO ₃) ₂
		6.022×10 ²³ Zn ²⁺ ions	$= 65 + 2(14 + 3 \times 16) = 189 \text{ g}$
CO ₂	2	<u>2</u> x 2 x 6.022×10 ²³ O atoms	Mass of 1 mole CO ₂
		<u>2</u> x 6.022×10 ²³ C atoms	$= 12 + 2 \times 16 = 44 \text{ g}$
		Or	Mass of 2 mole CO ₂
			= 2 x 44 = 88 g
		2 x 6.022×10 ²³ CO ₂ molecules	
Fe ₂ O ₃	1.5	<u>1.5</u> x 2 x 6.022×10 ²³ Fe ³⁺ ions	Mass of 1 mole Fe ₂ O ₃
		<u>1.5</u> x 3 x 6.022×10 ²³ O ²⁻ ions	= 56 x 2 + 3 x 16 = 160 g
			Mass of 1.5 mole Fe ₂ O ₃
			= 1.5 x 160 = 240 g

From the table above, two relationships can be established, namely





Total mass of E = 160 - (3 x 16) = 112 g (Subtract mass of 3xO atoms)

Mass of E = $\frac{112}{2}$ = 56 g (A_r of E = 56)

Hence, E is <u>Iron</u>



Calculate the mass of nitrogen gas, N₂ that contains the same number of molecules as 35.5g of chlorine gas, Cl₂. (Assume Avogadro's constant is 6.0×10^{23})

Step 1: Find the no. of moles of Cl₂

No. of moles of $Cl_2 = \frac{mass}{Mr \text{ of } Fe0} = \frac{35.5}{35.5+35.5} = 0.5 \text{ moles } Cl_2$

Step 2: 0.5 moles of Cl_2 will contain the same no. of molecules as 0.5 mol of N_2 Find the no. of mole N_2 present.

No. of moles of $N_2 = 0.5$ moles

Step 3: Mass of N_2 = no. of moles x $M_r N_2$ = 0.5 x (14 + 14) = 14g

Calculate the mass of sulfur that contains the same number of atoms as 4.8g of oxygen gas, O₂. (Assume Avogadro's constant is 6.0×10^{23})

Step 1: Find the no. of moles of O_2

No. of moles of $Cl_2 = \frac{mass}{Mr \text{ of } 02} = \frac{4.8}{16+16} = 0.15 \text{ moles } O_2$

Step 2:

1 mole of O_2 contains = 6 x 10²³ O_2 molecules = 2 x 6 x 10²³ O_2 atoms

Step 3: Find the number of moles of sulfur

0.15 moles of Cl_2 will contain the same no. of molecules as 0.5 mol of N_2 Find the no. of mole N_2 present.

No. of moles of $N_2 = 0.5$ moles

Step 3: Mass of N₂ = no. of moles x M_r N₂ = 0.5 x (14 + 14) = 14g



VOLUME OF GAS

One mole of all gases occupy the same volume (molar volume) of 24dm³ or 24 000cm³ at room temperature and pressure (r.t.p).

Mass of gas	Mole of gas	Volume of gas
44g CO ₂	1 mole CO ₂	24 dm ³
71g Cl ₂	1 mole Cl ₂	24 dm ³
17g NH₃	1 mole NH ₃	24 dm ³

Mass of gas	Mole of gas	Volume of gas
88g CO ₂	2 mole CO ₂	2 x 24 dm ³
8g H ₂	4 mole H ₂	4 x 24 dm ³
56g N ₂	2 mole N ₂	2 x 24 dm ³

Thus the relationship between mole and volume of a gas can be summarized as follows:





APPLICATIONS




Calculate the mass of oxygen in 6 dm³ of oxygen gas

Step 1: Find the no. of moles of O2

No. of moles of $O_2 = 6 \text{ dm}^3 / 24 \text{ dm}^3 = 0.25 \text{ mol}$

Step 2: Use the Mr of O_2 to find the mass

Mass of O₂ = No. of mol x Mr = 0.25 x (16+16) = <u>8 g</u> O₂

Calculate the Mr of an unknown gas if 600cm3 of the gas has a mass of 1.1 g

Step 1: Find the no. of moles of gas

No. of moles of gas = $\frac{0.6}{24}$ = 0.025 moles

Step 2: Find the Mr of the gas

Mr of gas = mass / no. of mole

= 1.1 / 0.025 = <u>44 (</u>Mr has no units)





STOICHIOMETRY (MOLE RATIO)

In a chemical equation, a specific amount of reactants will react with each other to form a specific amount of product(s). This value is "fixed" for every balanced chemical equation.

From the "fixed" value of the chemical equation, we can obtain the mole ratio of the chemical reaction or stoichiometry.

Example 1:	Example 2:
The chemical equation is	The chemical equation is
Fe + 2HCI → FeCl ₂ + H ₂	$Ca(OH)_2 + 2HCI \rightarrow CaCl_2 + 2H_2O$
The corresponding mole ratio is	The corresponding mole ratio is
1:2:1:1	1 : 2 : 1 : 2
Fe + 2HCI → FeCl ₂ + H ₂	Ca(OH)₂ + 2HCI → CaCl₂ + 2H₂O
Example 3:	Example 4:
Example 3: The chemical equation is	Example 4: The chemical equation is
Example 3: The chemical equation is N₂ + 2H₂ → 2NH₃	Example 4: The chemical equation is CH₄ + 2O₂ → CO₂ + 2H₂O
Example 3: The chemical equation is N₂ + 2H₂ → 2NH₃ The corresponding mole ratio is	Example 4: The chemical equation is CH ₄ + 2O ₂ → CO ₂ + 2H ₂ O The corresponding mole ratio is
Example 3: The chemical equation is $N_2 + 2H_2 \rightarrow 2NH_3$ The corresponding mole ratio is 1 : 2 : 3	Example 4: The chemical equation is CH ₄ + 2O ₂ → CO ₂ + 2H ₂ O The corresponding mole ratio is 1 : 2 : 1 : 2
Example 3: The chemical equation is $N_2 + 2H_2 \rightarrow 2NH_3$ The corresponding mole ratio is 1 : 2 : 3 $N_2 + 2H_2 \rightarrow 2NH_3$	Example 4: The chemical equation is $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ The corresponding mole ratio is 1 : 2 : 1 : 2 $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Hence from the chemical equation, we can use the mole ratio (stoichiometry) to determine:

- (a) No. of moles of products / reactants
- (b) Mass of products / reactants
- (c) Volume of gas formed / used



APPLICATIONS

Example 1:

In an experiment 1.4 g of iron was reacted with excess HCl.

 $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$

Calculate:

- (a) the no. of moles of HCl reacted
- (b) vol of gas produced
- (c) mass of salt formed

From the chemical equation, the mole ratio is

1 : 2 : 1 : 1Fe + 2HCl \rightarrow FeCl₂ + H₂

SOL⊵

(a) No. of moles of Fe = mass / Mr Fe = 1.4 / 56 = 0.025 mol Fe

Using mole ratio, 1 mole Fe : 2 mol of HCI Hence,

0.025 mol Fe : 2 x 0.025 mol HCl

No. of mol HCl reacted = <u>0.05 mol</u>

(b) No. of moles of Fe = 0.025 mol – From (a)

Using mole ratio, 1 mole Fe : 1 mol of H₂ Hence,

0.025 mol Fe : 0.025 mol H₂

Vol H₂ = mol x 24 dm³ = 0.025 x 24 = <u>0.6 dm³</u>

(c) No. of moles of Fe = 0.025 mol – From (a) Using mole ratio, 1 mole Fe : 1 mol of FeCl₂ Hence, 0.025 mol Fe : 0.025 mol FeCl₂

Mass of FeCl₂ = mol x Mr FeCl₂ = 0.025 x (56 + 35.5 + 35.5) = <u>3.175 g</u>

SOLN



Example 2:

0.3 mol of copper (II) sulfate reacts with sodium hydroxide to according to the equation shown

 $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$

Calculate:

(a) the no. of moles of NaOH reacted

(b) mass of ppt formed

From the chemical equation, the mole ratio is

1 : 2 : 1 : 1 $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$

(a) Using mole ratio,

1 mole CuSO₄ : 2 mol of NaOH

Hence,

0.3 mol CuSO4 : 2 x 0.3 mol NaOH

No. of mol NaOH reacted = <u>0.6 mol</u>

(b) No. of moles of $CuSO_4 = 0.3 \text{ mol} - \text{From}$ (a)

Using mole ratio, 1 mole CuSO₄ : 1 mol of Cu(OH)₂ Hence,

0.3 mole CuSO₄: 0.3 mol of Cu(OH)₂

Mass of $Cu(OH)_2 = mol x Mr Cu(OH)_2 = 0.3 x (64 + 2(16+1)) = 29.4 g$

Example 3:

Zn(NO₃)₂ decomposes when heated as shown

 $Zn(NO_3)_2 \rightarrow 2ZnO + 4NO_2 + O_2$

If 600 cm³ of oxygen were produced in the reaction, Calculate:

- (a) the mass of zinc nitrate decomposed
- (b) mass of solid residue formed
- (c) the volume of NO2 released

SUI ₽



From the chemical equation, the mole ratio is 1 : 2 : 4 : 1 $Zn(NO_3)_2 \rightarrow 2ZnO + 4NO_2 + O_2$

(a) No. of mol of O_2 produced = 0.6 / 24 = 0.025 mol O_2

Using mole ratio,

1 mole O₂: 1 mol of Zn(NO₃)₂

Hence,

0.025 mole O_2 : 0.025 mol of Zn(NO₃)₂

Mass of Zn(NO₃)₂ reacted = mol x Mr Zn(NO₃)₂ = 0.025 x (65 + 14 +48) = <u>3.175g</u>

(b) No. of moles of $O_2 = 0.025 \text{ mol} - \text{From}$ (a)

Using mole ratio, 1 mole O₂ : 2 mol of ZnO Hence, 0.025 mole O₂ : 2 x 0.025 mol of ZnO

Mass of ZnO = mol x Mr ZnO = 0.05 x (65 + 16) = <u>4.0 g</u>

(c) No. of moles of $O_2 = 0.025 \text{ mol} - \text{From}$ (a)

```
Using mole ratio,

1 mole O<sub>2</sub> : 4 mol of NO<sub>2</sub>

Hence,

0.025 mole O<sub>2</sub> : 4 x 0.025 mol of NO<sub>2</sub>
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Vol of NO₂ = mol x 24 dm³ = 0.1 x 24dm³ = <u>2.4 g</u>



Question 1

CO₂ reacts with NaOH as shown: CO₂ + 2NaOH \rightarrow Na₂CO₃ + H₂O

300 cm³ of the gas r.t.p. were passed into excess NaOH. Calculate:

(a) the mass of NaOH that was required to absorb the reacting aas	[10]	
(d) the mass of record and was required to absolb the reacting gas	[19]	

- (b) mass of anhydrous Na₂CO₃ produced. [1.325g]
- (c) mass of hydrated Na₂CO_{3.}10H₂O produced. [3.575g]

Question 2

Ethanol, C_2H_5OH is mixed with petrol to form gasoline. On complete combustion, ethanol produces CO_2 and H_2O . What volume of O_2 measured at r.t.p. is needed for the complete combustion of 23.0g of ethanol? [36 dm³]

Question 3

0.5 g of an alloy of iron, carbon and silicon (carbon steel) when reacted with an excess of hydrochloric acid gave 192 cm³ of hydrogen at r.t.p. Calculate the percentage of iron in carbon steel. Hint: Only metals react with acids to liberate hydrogen gas. [89.6%]

Question 4:

Potassium superoxide KO_2 can be used in spacecraft to supply oxygen according to the equation

 $4\mathrm{KO}_2 + 2\mathrm{H}_2\mathrm{O} \rightarrow 4\mathrm{KOH} + 3\mathrm{O}_2$

The potassium hydroxide formed removes carbon dioxide.

- (a) Show that 1.9 g of potassium superoxide will supply 0.25 dm³ of oxygen gas at r.t.p. [0.25 dm³]
- (b) Write the equation for the reaction of potassium hydroxide and carbon dioxide. $[2KOH + CO_2 \rightarrow K_2CO_3 + H_2O]$
- (c) Calculate the volume of carbon dioxide absorbed by 1.0g of alkali

[0.3 dm³]



STOICHIOMETRY (VOLUME RATIO OF GASES)

For chemical equations involving gases,

mole ratio = volume ratio

This is because the no. of moles of all gases occupy the same volume ie. 1 mole of any gas occupies 24 dm^3

There is no need to convert the volume of a gas to mole. Use mole ratio instead.

Example 1:	Example 2:
The chemical equation is	The chemical equation is
$C_{3}H_{8} + 5O_{2} \rightarrow 3CO_{2} + 4H_{2}O$	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
The corresponding mole ratio is	The corresponding mole ratio is
1 : 5 : 3 : 4	1 : 2 : 1 : 2
$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
The corresponding volume ratio if 20cm ³	The corresponding volume ratio if 10cm ³
of C3H8 was used is	of CO ₂ was used is
20 : 100 : 60 : 80	10 : 20 : 10 : 20
$C_{3}H_{8} + 5O_{2} \rightarrow 3CO_{2} + 4H_{2}O$	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
Example 3:	
The chemical equation is	
$N_2 + 2H_2 \rightarrow 2NH_3$	
The corresponding mole ratio is	
1:2:3	
$N_2 + 2H_2 \rightarrow 2NH_3$	
The corresponding volume ratio if 40cm ³	
of N ₂ was used is	
40 : 80 : 120	
N ₂ + 2H ₂ → 2NH ₃	

[2500 cm³]

[3400 cm³]



PRACTICE

Question 1

In the motorcar engine, 200 cm³ of octane (C₈H₁₈) is burnt according to the equation

 $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

Calculate:

- (a) The volume of oxygen used
- (b) The total volume of gases produced
- (c) If the engine is cooled to room temperature what is the total volume of gas produced? [1600 cm³]

Question 2:

40 cm³ of a gaseous oxide of nitrogen were decomposed completely into 40 cm³ of nitrogen and 20 cm³ of oxygen. What is the formula of the oxide? [N₂O]

Question 3:

The gaseous element X is diatomic. One volume of the element combines with one volume of hydrogen gas to form one volume of gaseous hydride. What is the formula of the hydride of X? $[X_2H_2]$

Question 4:

What is the ratio of the volume of 2 g of hydrogen to the volume of 16 g of methane, both volumes at r.t.p? [1:1]

Question 5:

124 g phosphorus has the same volume as 71 g of chlorine at the same temperature and pressure. What is the formula of a molecule of phosphorus? $[P_4]$



CONCENTRATION OF ACIDS AND ALKALIS

A solution is a **mixture** which contain a solid (solute) dissolved in a liquid (solvent)

The concentration of a solution (usually acids/alkalis) is simply the amount of solute dissolved in the solvent.

Concentration can be expressed in 2 ways namely:

- (a) **mol/dm³** or mol dm⁻³ (also known as molarity)
- (b) **g/dm³** or g dm⁻³

Note:

- (a) represent the no. of moles of solute dissolve in a volume of solvent.
- (b) represent the mass of solute dissolve in a volume of solvent.

Example 1:	Example 2:
Calculate the no. of mole of solute in	Calculate the no. of mole of solute in
500 cm ³ of a 0.5 mol/dm ³ solution	0.25 dm ³ of a 2.0 mol/dm ³ solution
From the concentration, 0.5 mol/dm ³	From the concentration, 2.0 mol/dm ³
1 dm ³ contains 0.5 mol solute	1 dm ³ contains 2 mol solute
1000 dm ³ contains 0.5 mol solute	Hence
1 cm ³ contains 0.5/1000 mol solute	0.25 dm dm ³ contains 2 x 0.25 mol solute
Hence,	= 0.5 mol solute
500 cm ³ contains 0.5/1000 x 500 mol	
= 0.25 mol solute	
Example 3:	Example 4:
Calculate the vol of solution which	Calculate the vol of solution which
contain 0.2 mol of solute in 0.5 mol/dm ³	contain 1.5 mol of solute in 1.25
solution	mol/dm ³ solution
From the concentration, 0.5 mol/dm ³	From the concentration, 1.25 mol/dm ³
0.5 mol solute dissolved in 1 dm ³	1.25 mol solute dissolved in 1 dm ³
0.5 mol solute dissolved in 1000 cm ³	1.25 mol solute dissolved in 1000 cm ³
1 mol solute dissolved in 1000/0.5 cm ³	1 mol solute dissolved in 1000/1.25 cm ³
Hence,	Hence,
0.2 mol solute dissolved in	1.5 mol solute dissolved in
1000/0.5 × 0.2	1000/1.25 x 1.5
= 400 cm ³	= 1200 cm ³
= 0.4 dm ³	= 1.2 dm ³



Example 5: (Involving g/dr A solution contains 2.0g of Calculate its concentration (a) g/dm ³ (b) mol/dm ³	n³) NaOH in 250cm³. n in
(a) Concentration of NaOl	H = 2g / 0.25 dm ³ = 8 g/dm ³
(b) No. of moles of NaOH Hence, conc NaOH	= 2 / (23+16+1) = 0.05 mol = 0.05 / 0.25 = 0.2 mol/dm ³

Thus the relationship between concentrations of a solution can be expressed as:





Question 1:	
Calculate the mass of solute in 0.5 dm ³ of 0.5 mol/dm ³ of NaOH	[10g]
Question 2:	
Calculate the mass of solute in 150 cm ³ of 5.0 mol/dm ³ of K_2CO_3	[103.5g]
Question 3:	
A sample of 2 dm3 sodium carbonate solution has a concentration of 0.2 Calculate	mol/dm3.
(a) the no. of mol of sodium carbonate in the solution	[0.4 mol]
(b) the mass of hydrated sodium carbonate Na2CO3.10H2O, used to pr solution. [Hint: Moles of hydrated salt = Moles of anhydrous compou	epare the und] [114.4g]
(c) the no. of mol of water contained in the crystal sample.	[72g]



CONCENTRATION (IN TERMS OF IONS)

What is the concentration of hydrogen ions in 0.05 mol/dm³ sulfuric acid, $\rm H_2SO_4$ in terms of:

(a) mol/dm³

(b) g/dm³

2 possible ways to solve:

(a) As the volume of sulfuric acid is	(b)Let the volume of sulfuric acid be 1 dm ³
fixed, the mol of H ₂ SO ₄ present in this	
vol = 0.05 mol	Thus
	Mol of H ₂ SO ₄ = 0.05 x 1dm = 0.05 mol
Thus,	
0.05 mol H ₂ SO ₄ contains	Thus,
0.05 mol x 2 H+ ions (dibasic acid) and	0.05 mol H ₂ SO ₄ contains
0.05 mol SO42- ions	0.05 mol x 2 H+ ions (dibasic acid) and
	0.05 mol SO42- ions
Hence,	
Concentration of H ⁺ = mol/vol	Hence,
= 0.05 mol/dm ³	Concentration of H ⁺ = mol/vol
	= 0.05/1
	= 0.05 mol/dm ³

Since 0.05 mol H ₂ SO ₄ contains (0.05 mol x 2 H+ ions (dibasi	c acid)
The mass of H+ ions present	= mol x Mr H+	
	= 0.1 x 1	[A _r of H ⁺ is 1]
	=0.1g	
The concentration of H+(assum	ing volume of acid is 1dm	³) = 0.1 /1 = 0.1 g/dm ³

Practice

Question 1:

What is the concentration of sodium ions in a solution containing 0.5 mol/dm³ of sodium carbonate? [1.0 mol/dm³]

Question 2:

What is the mass of potassium ions contained in 20.0 cm³ of 0.2 mol/dm³ of K₃PO₄? [Hint: Find mol of K⁺ salt \rightarrow Mol of K⁺ ions \rightarrow Mass of K⁺ ions] [0.468 g]



MOLE RATIO INVOLVING SOLUTIONS

Example:

Excess solid zinc was added to 50.0 cm³ of 2.0 mol/dm³ of HCl

Calculate:

- (a) the no. of mol of acid in 50.0 cm³ of 2.0 mol/dm³ solution
- (b) the no. of mol of hydrogen evolved in the reaction.
- (c) The volume of hydrogen gas evolved at r.t.p

Chemical Equation:	$Zn(s) + 2HCI (aq) \rightarrow ZnCl_2 (aq) + H_2 (g)$
(a) No. of mol of HCl	= conc x vol (dm ³) = 2.0 x 50/1000 = 0.1 mol HCl
(b) No of mol of H ₂	= $\frac{1}{2}$ mol HCl (Using the mole ratio of the equation) = $\frac{1}{2} \times 0.1 = 0.05$ mol H ₂
(c) Vol of H_2	= mol x 24 dm ³ = 0.05 x 24 = 1.2 dm ³ H ₂

Practice

Question 1:

3.0 dm³ of 0.1 mol/dm³ of sulfuric acid is split on the lab floor. The split acid is neutralized by 1.0 mol/dm³ of aqueous sodium hydrogen carbonate as shown below.

 $H_2SO_4 + 2NaHCO_3 \rightarrow Na_2SO_4 + H_2O + CO_2$

Calculate the volume of sodium hydrogen carbonate needed to neutralize the acid. [0.15 dm³]

Question 2

Sodium carbonate was added to 100 cm³ of 1.0 mol/dm³ hydrochloric acid. Calculate the mass of hydrated sodium carbonate Na₂CO₃.10H₂O required to neutralize the acid. [Hint: Mol of hydrated salt = Mol of anhydrous salt] [14.3 g]



Question 3:

20.0 cm³ of 0.1 mol/dm³ of NaOH reacted with iron (II) sulfate according to the following equation.

 $2NaOH + FeSO_4 \rightarrow Fe(OH)_2 + Na_2SO_4$

Calculate:

(a) The no. of mol of iron (II) sulfate which reacted. (b) The mass of iron contained in the sample of the salt	[0.001 mol] [0.056 g]
Question 4: A small piece of Li of mass 0.35g is added to cold water. The resulting solution is titrated with 2.0 mol/dm ³ HCI.	
What volume of HCl is needed to neutralize the solution? [Hint: Calculate mol of alkali \rightarrow mol HCl \rightarrow vol of HCl]	[0.025dm ³]

Question 5:

The ionic equation for a reaction is given below

 $3\mathrm{SO}_2 + \mathrm{Cr}_2\mathrm{O}_7^{2\text{-}} + 2\mathrm{H}^+ \xrightarrow{} 3\mathrm{SO}_4^{2\text{-}} + 2\mathrm{Cr}^{3\text{+}} + \mathrm{H}_2\mathrm{O}$

A solution contains 0.1 mol/dm³ of dichromate (VI) ions. Calculate the minimum volume of this solution required to remove the sulfur dioxide gas of 2 dm³ of polluted air which contain 3.6% by volume of SO₂ measured at r.t.p.

[0.01 dm³]



DILUTING AND CONCENTRATING SOLUTIONS

[No change in amount of solute, water is added or boiled off]

Example

100 cm³ of 1.0 mol/dm³ of NaOH was concentrated to 20 cm³. Calculate the concentration in mol/dm³ of the resulting solution.

The no. of mole in the solution does not change. The volume has been decreased from 100 cm³ to 20cm³

No. of mole of NaOH	= conc x vol
	$= 1.0 \times 100/1000$
	= 0.1 mol NaOH
Conc of NaOH (in 20cm ³)	= mol / vol
	$= 0.1 / 0.02 = 5 \text{ mol/dm}^3$
	(The solution has become more concentrated)

Practice

Question 1:

Calculate the volume of water added to dilute 100 cm³ of 1.0 mol/dm³ of sodium carbonate to form a 0.2 mol/dm³

 $[0.5 - 0.1 = 0.4 \text{ dm}^3]$





TITRATION CALCULATIONS

Question 1:

A sample of 20.0 cm³ of 0.1 mol/dm³ of NaOH was neutralized by 21.6 cm³ of sulfuric acid. Calculate the concentration of the sulfuric acid in mol/dm³ used. [Hint: Mole of NaOH \rightarrow Mol of acid \rightarrow Conc of acid [0.0463 dm³]

Question 2:

20.0 cm³ of 0.2 mol/dm³ of KOH was neutralized by 25.0 cm³ of sulfuric acid.Calculate the concentration of the sulfuric acid in(a) mol/dm³(b) g/dm³[15.68 g/dm³]

Question 3:

26.6 cm³ of 0.035 mol/dm³ of HCl was neutralized by 20.0 cm³ of NaOH.Calculate the concentration of the alkali in(a) mol/dm³(b) g/dm³[1.86 g/dm³]

Question 4:

20.0 cm³ of 0.055 mol/dm³ of HCl was titrated with 0.01 mol/dm³ of NaOH. Calculate the volume of the alkali needed to neutralized the acid. [0.11 dm³]

Question 5:

BA1 contains 0.75 mol/dm³ of NaOH BA2 contains 6.10 g/dm³ of HXO₃

20.0 cm ³ of the alkali neutralized 19.5 cm ³ of the acid	
(a) Calculate the concentration in mol/dm ³ of HXO ₃	[0.0769 mol/dm ³]
(b) Calculate the molar mass (mass of 1 mole) of the acid, HXO ₃	[793 g]

Question 6:

20.0 cm³ of 0.2 mol/dm³ of KOH was titrated with 25.0 cm³ of H2SO4.Calculate the concentration in(a) mol/dm³(b) g/dm³[0.16 mol/dm³][15.68 g/dm³]

Question 7:

20.0 cm³ of 0.015 mol/dm³ of KOH exactly neutralized 18.3 cm³ of acid HX. The Mr of acid HX is 80.

 $\mathsf{HX} + \mathsf{KOH} \twoheadrightarrow \mathsf{KX} + \mathsf{H_{2}O}$

Calculate the mass of the acid in 1 dm³ of the solution [2.55 g]



PERCENTAGE PURITY

% purity = $\frac{\text{mass of pure}}{\text{total mass of sample (impure)}} \times 100 \%$

Note:

- The sample that is given in the question is often the impure sample.
- The pure sample that is present is usually calculated based on the mole/mass/vol of product formed.

Example

This is the original mass consisting of pure C and impurities

Only pure C can react to form CO₂. This is the vol of CO₂ formed.

When 100 g of coal, C were completely burnt in air, it produced 180 dm³ of carbon dioxide gas at r.t.p. Calculate the % purity of the coal sample.

 $C + O_2 \rightarrow CO_2$

Step 1: Find the mole of CO₂ formed.

No. of mol $CO_2 = 180/24 = 7.5$ mol

Step 2: Find the mole and mass of C reacted. (This will be the pure C reacted)

No. of mol C = 7.5 mol Mass C = 7.5 mol x 12 = 90g

(using mol ratio of eqn)

Step 3: Find the percentage purity

% purity =
$$\frac{90}{100}$$
 x 100 % = 90% pure C (From 100g of coal, 90% of it contain pure C)



Strategy:

From the qns, always identify the impure sample and identify the amount of product formed to deduce the % purity.

Question 1:

1.4 g of sulfur dioxide were obtained /produced when 0.8 g of sulfur was burnt. Calculate the % purity of the sample of sulfur $$_{87.5\%}$$

Question 2:

540 cm³ of hydrogen gas was produced when 2.0 g of zinc was added to dilute HCl under room conditions. Calculate the % purity of the zinc sample 73.13%

Question 3:

3.31 g of zinc oxide was produced when a sample of 6.25 g of zinc carbonate was heated. $$_{\rm 81.73\%}$$

Calculate:

(a) the mass of zinc carbonate needed to produce 3.31 g of zinc a	oxide. 5.12g
(b) the % purity of the zinc carbonate sample.	81.73%
(c) the volume of carbon dioxide gas produced.	0.981dm3

Question 4:

A sample of potassium hydroxide containing 8.5 g/dm³ has potassium chloride as an impurity. 28.6 cm³ of 0.052 moldm³ sulfuric acid was required to neutralize 20.0 cm³ of the alkali. Calculate the % by mass of KOH in the sample. 98.12% [Hint: The concentration KOH contains the impurities]

Question 5:

A sample of sodium hydroxide containing 5.0 g/dm³ has calcium chloride as an impurity. 23.6 cm³ of 0.012 moldm³ HCl was required to neutralize 25.0 cm³ of the alkali. Calculate the % purity of the alkali sample.



PERCENTAGE PURITY

Terms	Explanation
Yield	Amount of products formed in a reaction
Theoretical yield	Calculated amount of product obtained (from mole ratio)
Experimental yield	Actual amount of product obtained in a reaction

% Yield = $\frac{\text{experimental yield}}{\text{theoretical yield}} \times 100 \%$

Note:

- The experimental yield is often given. This value cannot be calculated directly from mol ratio
- The theoretical yield is obtained via calculation from the mol ratio

Example

When 3.84 g of magnesium metal was heated in excess oxygen gas. 6.0 g of oxide was obtained. Calculate the % yield of the oxide.

 $2Mg + O_2 \rightarrow 2MgO$

The experimental yield is 6.0g

Step 1: Calculate the theoretical yield of MgO from Mg

No. of mol of Mg = 3.84 / 24 = 0.16 mol Mg No. of mol of MgO = $0.16 \times 1 = 0.16$ mol MgO Mass of MgO (Theoretical yield) = $0.16 \times (24 + 16) = 6.4$ g

Step 2: Calculate % yield

% yield = 6 / 6.8 x 100% = 93.8 %





Question 1:

2.5 g of calcium oxide was obtained by heating 5.0 g of calcium carbonate. Calculate the % yield of the oxide. $$_{89.3\%}$$

Question 2:

0.7 g of hydrated copper (II) sulfate was obtained by reacting 40 cm³ of 0.1 mol/dm³ sulfuric acid with excess copper (II) oxide. Calculate the % yield of the hydrated salt, CuSO₄.5H₂O. [Hint: the mol of hydrated salt is similar to the mole of aqueous salt formed]

Question 3:

1.15 dm³ of carbon dioxide was collected when 4.2 g of magnesium chloride was heated. Calculate the % of the carbonate decomposed. $_{95.8\%}$

Question 4:

Ethanol, C₂H₅OH is oxidised to ethanoic acid, CH₃COOH as shown

 $C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$

If 9.2 g of pure ethanol was used, calculate

(a) mass of ethanoic acid obtained if the yield is 100%	12g
---	-----

(b) the mass of acid obtained if the yield is 75% 9g

Question 5:

In the reaction,

 N_2 + $3H_2 \rightarrow 2NH_3$

The yield of ammonia is 15%. What is the volume of ammonia produced at r.t.p,

- (a) if 8.4 kg of nitrogen is reacted with excess hydrogen? 2160dm3
- (b) if 72dm³ of hydrogen is reacted with excess nitrogen? 7.2dm³



LIMITING REACTANTS

Note:

- For this category of question, data/facts are given for both reactants
- Only one value will be used in the mol ratio for calculation
- The other reactant will be the excess, this data is not used for calculation, unless the remaining excess reactant is required.
- The mol of the limiting reactant will always be used to calculate the mol of products

Example

30.0 g of magnesium was burnt in 24.0 g of oxygen gas. Calculate the mass of oxide produced.

2Mg + O₂ → 2MgO

Step 1: Convert both mass to mol for comparison

Mol of Mg = 30 / 24 = 1.25 mol Mg Mol of O₂ = 24 / (16 + 16) = 0.75 mol O₂

Step 2: Compare and identify the excess/limiting reactant

Actual mol present:	2Mg 1 .25 mol	+ O	2 i mol	→	2MgO
If 1.25 mol of Mg is us	ed,	1.25, [The exce	/2 = 0.62 re is 0.73 ess]	5 mol (5 mol C	D_2 is used D_2 present. Hence O_2 is the
If 0.75 mol of O2 is use	ed	0.75 [Hov for c	x 2 = 1.5 vever, o i reactio	mol M hly 1.25 n, hend	ig is required 5 mol Mg is present (insufficient ce Mg is limiting the reaction.]

Step 3: Since the limiting reactant is identified (Mg), we will use mol of Mg

No. of mol of Mg = 1.25 mol MgNo. of mol of MgO = 1.25 mol MgOMass of MgO = $1.25 \times (24 + 16) = 50 \text{ g MgO}$



Question 1:

0.006 mol of magnesium carbonate is added to 20 cm³ of 0.2 mol/dm³ hydrochloric acid. Calculate

- (a) no. of mol of magnesium carbonate in 0.5 g
- (b) no. of mol of acid in 20 cm³ of 0.2 mol/dm³
- (c) volume of carbon dioxide formed.
- (d) mass of excess reactant remaining after the reaction.

Question 2:

A solution containing 1 mol of NaOH is added to another solution containing 1 mol of ZnSO₄. Calculate the no. of mol of Zn(OH)2 precipitated.

Question 3:

Copper (II) chloride reacts with sodium hydroxide as shown below

 $CuCl_2 + 2NaOH \rightarrow Cu(OH)_2 + 2NaCI$

5.0 cm³ of 1.0 mol/dm³ of CuCl₂ was added to 10 cm³ of 0.15 mol/dm³ NaOH

Calculate

- (a) no.of mol of CuCl₂ present
- (b) no. of mol of NaOH present
- (c) no. of mol of $Cu(OH)_2$ formed.

Question 4:

0.1 mol of Zn was added to 20 cm³ of 0.5 mol/dm³ hydrochloric acid. Calculate the volume of hydrogen gas evolved at r.t.p.

Question 5:

4.0 cm³ of 1.0 mol/dm³ of CuSO₄ was mixed with 8.0 cm³ of 1.0 mol/dm³ Na₂CO₃ CuSO₄ + Na₂CO₃ → CuCO₃ + Na₂SO₄ What was the colour of the resulting mixture?

Question 6:

A mixture containing 50 cm³ H₂ and 50 cm³ O₂ is exploded. Find the total volume of the resulting gas mixture at room temperature.

Question 7:

A sample of 30.0 cm³ of CO was burnt in 50.0 cm³ of O₂. What was the composition of the gas remaining after the reaction?

Question 8

12 dm³ of H₂ reacts with 6 dm³ of N₂ to 2 dm³ of NH₃. Which reactant is the excess? What is the % yield of ammonia?



EMPIRICAL & MOLECULAR FORMULA

Molecular formula	Empirical formula
Actual/True formula	Simplest formula
Shows the actual no. of atoms in the	Shows the simplest ratio of atoms in the
compound	compound
Molecular Formula	Empirical Formula
Fe₂Cl₀	FeCl₃
C₄H₄	СН
C ₂ H ₄	CH ₂
C ₃ H ₆	CH ₂
NaCl	NaCI (remain the same)



	CHEMISTRY	DEFINITION LIST	- SECONDARY	3 & 4
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KINETIC PARTICLE THEOR	Y
Volatile liquid	A liquid that evaporates quickly at room temperature
	due to its low boiling point.
Sublimation	The process of changing from the solid state directly to
	the gaseous state without passing through the liquid
	state.
Diffusion	The movement of particles from a region of higher
	concentration to a region of lower concentration due to
	the constant random motion of particles.
PURIFICATION AND SEPA	RATION TECHNIQUES
Pure substance	A single substance which is not mixed with other
	substances. It has a fixed and exact/one melting and
	boiling points.
Solute	The substance that is being dissolved to form a solution.
Solvent	The substance that dissolves another to form a solution
Solution	A mixture of solute and solvent
Saturated solution	A solution that contains the maximum mass of solute that
	can dissolve in a fixed/given volume of solvent at a fixed
	temperature. It is a solution in which no more solute can
	dissolve
Immiscible liquids	Liquids that do not dissolve in each other, and form
	different layers.
Miscible liquids	Liquids that mix together completely to form a solution
Filtration	A method to separate a mixture of a liquid and an
	insoluble solid
Residue	The insoluble solid which remains on the filter paper
	during filtration
Filtrate	The liquid which passes through the filter paper after
	filtration
Evaporation to dryness	A method to obtain a soluble solid from a solution by
	continuous heating until all the water has boiled off.



Crystallization	A method to obtain pure crystals of a solute from a
	saturated solution by cooling.
Simple distillation	A method to separate a solvent from a solution
Fractional distillation	A method to separate a mixture of miscible liquids with
	different boiling points
Separating funnel	An apparatus used to separate immiscible liquids
Chromatography	A method used to separate two or more components
	(Eg. Dyes, pigments) in a mixture using a suitable solvent
	(Eg. Water, ethanol)
Locating agent	A substance sprayed onto a chromatogram to show the
	presence of colourless substances on the paper. The
	locating agent colours the substances.
R _f value	The ratio between the distance travelled by the
	substance and the distance travelled by the solvent in a
	chromatography experiment.
ELEMENTS, MIXTURES AN	
Atom	The smallest particle of an element that has the chemical
	properties of that element.
Element	A substance made from only one type of atom. It cannot
	be separated into simpler substances by chemical
	processes or electricity.
Compound	A substance when two or more elements are chemically
	combined together
Mixture	A substance when two or more elements that are not
	chemically combined together
Molecule	A group of two or more atoms that are chemically
	combined
ATOMIC STRUCTURE	
Isotopes	Atoms of the same element that have the same proton
	numbers but different number of neutrons
Nucleon or Mass	The total number of protons and neutrons in the nucleus
Nucleon or Mass number	The total number of protons and neutrons in the nucleus of an atom



Electronic	The arrangement of electrons in shells around the nucleus
configuration/silociore	of an atom
IONIC BONDING	
lon	A positively or negatively charged particle. It is formed
	when an atom or a group of atoms loses or gains
	electrons
Cation	A positively charged ion
Anion	A negatively charged ion
lonic bond	The strong electrostatic force that holds positive and
	negative ions together in an ionic compound
Giant structure	A three-dimensional network of atoms or ions packed
	together in a regular pattern
Noble gas structure	It refers to a duplet or octet electronic configuration. It
	refers to an electronic structure with fully filled outer shell
Polyatomic ions	lons that are made up of a group of atoms
Lattice	It refers to the orderly arrangement of ions or atoms
	which repeats itself throughout a structure
COVALENT AND METALLI	C BONDING
Covalent bond	It is a bond formed by the sharing of electrons between
	two non-metallic atoms.
Intermolecular force	Weak forces of attraction between molecules with simple
	molecular structure
Allotrope	Allotropes are different forms of the same element with
	different arraignment of atoms and structures (Eg.
	Diamond and Graphite are allotropes of carbon)
Metallic bond	The electrostatic forces of attraction between positive
	metals ions in a sea of delocalised mobile electrons.
WRITING EQUATIONS	
Valency	The combining power of an atom which is determined by
	the number of electrons it will lose, gain or shared when it
	combines with other atoms.
Ionic equation	An equation that shows the ions that take part in a
	reaction



Spectator ions	lons that do not take part in any chemical reactions and are still present as ions at the end of the reaction.
THE MOLE	
Relative atomic mass	The average mass of one atom of an element when
(Ar)	compare to 1/12 of the mass of a Carbon-12 atom
Relative molecular	The average mass of one molecule when compare to
mass (Mr)	1/12 of the mass of a Carbon-12 atom
Relative formula mass	The sum of the relative atomic masses of all the atoms in
	a formula unit of an ionic compound
Avogadro's constant	The number of particles in one mole of a substance. Its
	value is 6 x 10 ²³
Avogadro's law	Avogadro's law states that at constant temperature and
	pressure, the volume of gas is directly proportional to the
	number of moles of the gas present
Mole	The amount of a substance which contain 6 x 10 ²³
	particles
Molar mass	The mass of one mole of a substance
CHEMICAL CALCULATIO	INS CONTRACTOR OF
Stoichiometry	The relationship between the amounts (measured in
Stoichiometry	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical
Stoichiometry	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction
Stoichiometry Limiting reactant	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction
Stoichiometry Limiting reactant	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed
Stoichiometry Limiting reactant Concentration of a	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the
Stoichiometry Limiting reactant Concentration of a solution	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or damage non-living things
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution Carbon cycle	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or damage non-living things The mechanism that maintains the level of carbon
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution Carbon cycle	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or damage non-living things The mechanism that maintains the level of carbon dioxide in the atmosphere
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution Carbon cycle Greenhouse effect	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or damage non-living things The mechanism that maintains the level of carbon dioxide in the atmosphere The warming effect that results when gases like carbon
Stoichiometry Limiting reactant Concentration of a solution THE ATMOSPHERE AND E Air pollution Carbon cycle Greenhouse effect	The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction The reactant that is completely used up in a reaction and it limits the amount of products formed The amount of solute dissolved in a unit volume of the solution NVIRONMENT The condition in which air contain a high concentration of certain chemicals that may harm living things or damage non-living things The mechanism that maintains the level of carbon dioxide in the atmosphere The warming effect that results when gases like carbon dioxide and water vapour trap and retain some of the



Global warming	The phenomenon of an increase in the Earth's average
	temperature due to the build-up of greenhouse gases in
	the atmosphere
Photosynthesis	A process where plants make use of carbon dioxide and
	water in the presence of sunlight and chlorophyll to
	make glucose.
	$(6CO_2 + 6H_2O + Chlorophyll \rightarrow C_6H_{12}O_6 + 6O_2)$
Respiration	A process in which food substances / glucose react with
	oxygen inhaled to produced carbon dioxide, water and
	energy
	(C ₆ H ₁₂ O ₆ + 6H ₂ O → 6CO ₂ + 6H ₂ O)
SALTS	
Salt	A salt is an ionic compound formed when a metallic ion
	or an ammonium ion replaces one or more hydrogen
	ions of an acid
Acid salt	A salt containing one or more replaceable hydrogen ion
	A salt of a polybasic acid which is only partially
	neutralized by a base Eg NaHCO3
Precipitate	An insoluble solid that is produced in a solution as a result
	of a chemical reaction
Water of crystallization	Water molecules that combined with salt to form crystals
Hydrated salts	Salts that contain water of crystallization
Anhydrous salts	Salts that do not contain water of crystallization
OXIDATION AND REDUC	TION
Oxidation	A reaction where a substance:
	Takes in oxygen
	Loses hydrogen
	Loses electron(s)
	Increases in oxidation state after reaction
Reduction	A reaction where a substance:
	Loses in oxygen
	Take in hydrogen
	Take in electron(s)
	Decreases in oxidation state after reaction



Oxidising Agent	It is a substance that causes oxidation in another
	substance.
	It is itself reduced after the reaction
Reducing Agent	It is a substance that causes reduction in another
	substance.
	It is itself oxidised after the reaction
Redox Reaction	A reaction whereby oxidation and reduction take place
	simultaneously
Oxidation State	The charge an atom of an element would have if it
	existed as an ion in a compound.
METALS	
Alloy	A mixture of a metal with non-metal or other metals
Rusting	The slow oxidation of iron in the presence of air and
	water to form hydrated iron (III) oxide (rust)
Corrosion	The wearing away of the surface of a metal by chemical
	reaction
Ore	An ore is a compound ((Usually the oxides, sulphides,
	chlorides or carbonates) mixed with large amounts of
	earth and rock
Recycling	Recycling is a process of collecting and processing
	materials that would otherwise be thrown away as trash
	and turning them into new products
Sacrificial protection	The use of a more reactive metal which is attached to a
	less reactive metal so that the more reactive metal will
	corrode in place of the less reactive metal.
ELECTROLYSIS	
Electrolysis	A process of using electricity to break down or
	decompose a compound (usually an ionic compound in
	molten or aqueous state). The current is carried by
	moving ions.
Electroplating	The process of depositing a layer of metal on another
	substance using electrolysis
Simple / Electric Cell	A device which converts chemical energy into electricity



Cathode	A negatively charged electrode in an electrolytic cell.	
	The electrode that is connected to the negative terminal	
	of the battery.	
Anode	A positively charged electrode in an electrolytic cell.	
	The electrode that is connected to the positive terminal	
	of the battery.	
Electrode	A conductor which carries electricity in or out of an	
	electrolyte during electrolysis	
Electrolyte	A compound which in molten or aqueous state,	
	conducts electricity by the movement of ions and is	
	decomposed in the process.	
THE PERIODIC TABLE		
Displacement reaction	A reaction in which an tom or molecule takes the place	
	of another atom or molecule in a compound	
ENERGY CHANGES		
Endothermic Reaction	A reaction which absorbs heat from the surroundings	
	causing a decrease in temperature of the surrounding	
Exothermic Reaction	A reaction which release heat to the surroundings	
	causing increase in temperature of the surrounding	
Enthalpy change	The amount of energy involved (taken in or given out) in	
	a reaction.	
Activation Energy	The minimum energy that the reactant particles must	
	possess during their collisions in order for a chemical	
	reaction to occur	
Collision Theory	Collision theory states that a chemical reaction can	
	occur only if the reacting particles collide with one	
	another	
Combustion	A reaction in which a substance burns in oxygen with the	
	release of heat energy.	
Fuel	A substance that burns easily in air to produce energy.	
Fossil fuels	Fuels produced many millions of years ago from the	
	decaying remains of animals or plants. The three main	
	fossil fuels are coal, petroleum and natural gas.	



Fuel Cell	A chemical cell in which reactants usually a fuel and	
	oxygen are continuously supplied to produce electricity	
	directly	
SPEED OF REACTION		
Speed of reaction	The amount of reactant used up per unit time	
	The amount of gas produced per unit time	
Effective Collisions	Collisions between reacting particles that result in the	
	formation of product particles	
Catalyst	A substance which increases the speed of a chemical	
	reaction and remains chemically unchanged at the end	
	of the reaction	
Enzyme	A biological catalyst that catalyses the chemical	
	reaction in plants and animals	
AMMONIA		
Reversible Reaction	A reaction that can go both forward and backward at	
	the same time	



INTRODUCTION TO ORG	ANIC CHEMISTRY	
Functional Group	An atom or group of atoms the gives the characteristic	
	properties of the organic compound.	
Homologous Series	A family of organic compounds with members of the	
	family having the same functional group and general	
	formula.	
	Each members differs from the next by -CH2- Group and	
	the members have similar chemical properties.	
Hydrocarbons	Organic compounds made up from the elements	
	hydrogen and carbon only	
Molecular Formula	A formula that shows the actual number of each type of	
	atoms	
Structural Formula	A formula that shows how each of the atoms in the	
	molecule are arranged.	
ALKANES AND ALKENES		
Alkanes	Hydrocarbon having the general formula C_nH_{2n+2}	
Isomers	Compounds that have the same molecular formula but	
	different structural formula.	
Cracking	The breakdown of long chain hydrocarbon molecules	
	with heat and/or catalyst to produce smaller	
	hydrocarbon molecules and/or hydrogen	
Substitution Reaction	A reaction in which the hydrogen atoms of an alkane are	
	substituted (replaced) by other atoms. Ultraviolet light is	
	needed for this reaction	
Saturated	Hydrocarbons that contains only single bonds between	
Hydrocarbons	carbon atoms.	
Unsaturated	Any hydrocarbons that contains one or more carbon-	
Hydrocarbons	carbon double bonds	
Alkenes	Hydrocarbon that contains one or more carbon-carbon	
	double bonds.	
	Alkenes with only one carbon-carbon double bond have	
	the general formula of CnH2n	



Addition Reaction	A reaction in which a molecule (element or compound)	
	adds to an unsaturated compound to form a single new	
	compound	
Hydrogenation	The addition of a hydrogen molecule across a double	
	bond	
ALCOHOLS AND CARBO		
Carboxylic Acid	An organic acid containing the carboxyl group - COOH	
Alcohol	An organic acid containing the hydroxyl group - OH	
Fermentation	The conversion of glucose by micro-organisms such as	
	yeast into ethanol an carbon dioxide	
MACROMOLECULES		
Monomer	A simple molecule that joins to other similar molecules to	
	form a long chain molecule (polymer)	
Polymer	A macromolecule/ giant molecule formed by a large	
	number of identical molecules (monomers) joined	
	together by covalent bonds	
Polymerisation	A chemical reaction in which a large number of	
	identical or similar molecules called monomers react with	
	each other to form a single long chain molecule/	
	macromolecule	
Addition Polymerization	A chemical reaction in which unsaturated monomers join	
	together to form a single giant molecule (polymer)	
	without losing any molecules or atoms	
Condensation	A chemical reaction in which a small molecule such as	
Polymerization	water is produced as a by-product of a reaction when	
	monomers reacts to form a polymer	
Non-biodegradable	It means cannot be decomposed by bacteria in the soil.	



COLOURS OF COMMON SUBSTANCES

Substance	Colour		
Non-Metals			
Sulfur powder Yellow			
Carbon	Black		
Metals			
Group I Metal	Silvery		
Most metals (Fe, Pb, Zn, Al, Ni etc)	Grey		
Copper	Reddish brown/pink		
Halogens			
Fluorine gas	Pale yellow		
Chlorine gas	Yellow green		
Bromine liquid / Aqueous bromine	Reddish brown		
lodine solid, l2 (s)	Purplish black		
lodine solution , l2 (aq)	Brown		
lodine vapour, , l2 (g)	Purple		

Insoluble metal hydroxides		
Aluminium hydroxide (s)	White	
Copper (II) hydroxide (s)	Blue	
Iron (II) hydroxide (s)	Dirty Green	
Iron (III) hydroxide (s)	Brown	
Lead (II) hydroxide (s)	White	

	Salts	
Group I - III salts	White (when solid)	
	Colourless (in aqueous solution)	
Cu ²⁺ salts	Mainly Blue	
	Exceptions:	
	CuCO3 (s) - Green	
	CuO (s) - Black	
	Cu ₂ O (s) - Red	
Fe²+ salts (Eg FeSO₄)	Mainly Green	
Ni ²⁺ salts (Eg NiSO4)		
Cr ³⁺ salts (Eg Cr ₂ (SO ₄) ₃)	Exceptions	
	FeO (s) - Black	
	Fe ₃ O ₄ (s) - Black	
Fe ³⁺ salts (Eg Fe ₂ O ₃)	Brown	
Silver chloride White		
Lead iodide	Yellow ppt	
Silver iodide		
Lead (II) oxide		
Barium sulfate	White ppt	
Mn ²⁺ salts	Colourless	
MnO ₂	Black solid	



TESTS FOR PRESENCE OF WATER

Substance	Colour
Test 1; Anhydrous Copper (II) Sulfate	White
Test 2; Hydrated Copper (II) Sulfate	Blue
Test 1; Anhydrous Cobalt Chloride	Blue
Test 2; Hydrated Cobalt Chloride	Pink

OXIDIZING AGENTS

Substance	Colour	
Acidified Potassium Manganate (VII)	KMnO₄ (aq) + RA → Mn ²⁺	
	Purple Colourless	

REDUCING AGENTS

Substance	Colour	
Potassium lodide solution	$\begin{array}{c} \text{KI (aq)} + \text{OA} \rightarrow I_2\\ \text{Colourless} & \text{Brown} \end{array}$	
Iron (II) Sulfate solution	$\begin{array}{c} FeSO_4 (aq) + OA \rightarrow Fe^{3+} \\ Green & Brown \end{array}$	

METAL OXIDES

Substance	Colour	
Zinc Oxide	Before heating: White While heating: Yellow (hot) After heating: White (cool)	
	$ZnCO_3$ + Heat \rightarrow ZnO + CO_2	
	White Yellow (hot)	
	White (Cool)	

INDICATORS

Substance	Acids	Alkali
Methyl Orange	Red	Orange
Phenolphthalein	Colourless	Red
Screened Methyl Orange	Red	Green
Red litmus paper	Remains Red	Turns Blue
Blue Litmus paper	Turns Red	Remains Blue

GASES

Substance	Colour
Nitrogen dioxide gas, NO2	Brown