

Particles and Purification

• 3 states of matter:

- **Solid** - Fixed shape + volume, cannot flow, not compressible
- **Liquid** - Fixed volume - can flow - compressible
- **Gas** - Spreads around everywhere - compressible

Particles	
Atom	- Smallest, can't be broken down
molecule	- 2+ atoms joined
ion	- 1 or group of atoms - electrical charge.

- Compressing gases: Pressure increases volume decreases.
- Heating gases: Higher the temp Higher the pressure. Particles move/vibrate faster.
Greater kinetic energy.

Energy put in
Melting + Boiling

Energy put out
Condensing + Freezing.

Sublimation
Gas → Solid

- The idea that particles are constantly in motion is called **Kinetic Particle Theory**.

- **Brownian motion** - Zig-Zag - heavy particles do not show.

Diffusion

- Movement of something (particles) from a high concentrated area to a low concentrated area.
- Diffusion in gases is faster.

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- Diffusion also provides evidence for kinetic particle theory.
- Can occur in liquids which are miscible or solids dissolving in liquids.

Apparatus for measuring

- Standard unit for mass is kilogram sometimes gram. We use a scale/top pan balance.

- We use a stop-watch/clock to measure time.

- We use a thermometer in degrees celsius to measure temperature.

- We measure volume in cm^3 or dm^3 . $1 \text{ dm}^3 = 1000 \text{ cm}^3$, we use four things:

1. measuring cylinder

2. Volumetric pipette - 10 cm^3 - 25 cm^3

3. Burette - 50 cm^3

4. Volumetric Flask

50 cm^3 - 1 dm^3

- We measure gas in a gas syringe or an upturned measuring cylinder.

Chromatography

- Chromatography is used to separate pigments with filter paper.

• Can identify substances in mixture.

• Purifying the separated substances.

- Can be used to identify colourless substances.

Purity

• Only one substance is present.

• Can tell through chromatography or through boiling and melting points. Boiling point rises when impurity is present, melting point decreases.

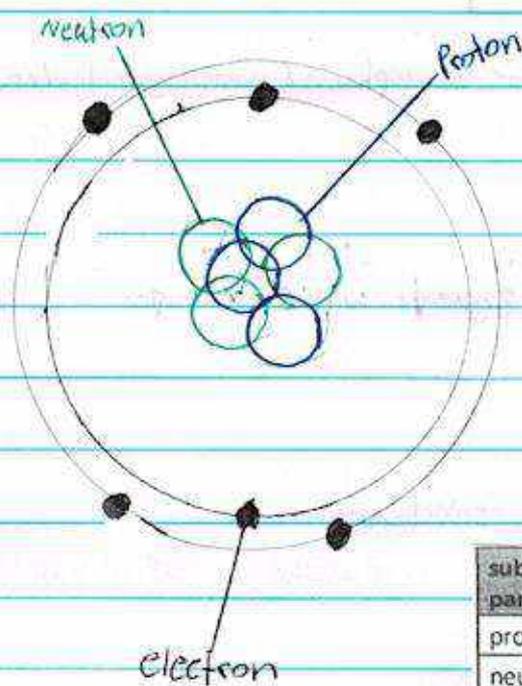
Methods of purification:

- Filtration - Undissolved solid in a solution - Filtrate - residue.
- Decanting - Pouring off solution from top
- Centrifuge - machine spins tube, solid gets pulled to bottom
- Crystallisation -
- Solvent extraction - separates 2 solutes in a solvent, second solvent is used to help.
- Distillation - simple distillation - fractional distillation

Atoms, + Elements, + Compounds.

• Atom is smallest uncharged particle - can take part in chemical change.

Atom: electrically neutral



$$P + N = \text{nucleon no. / mass no.}$$

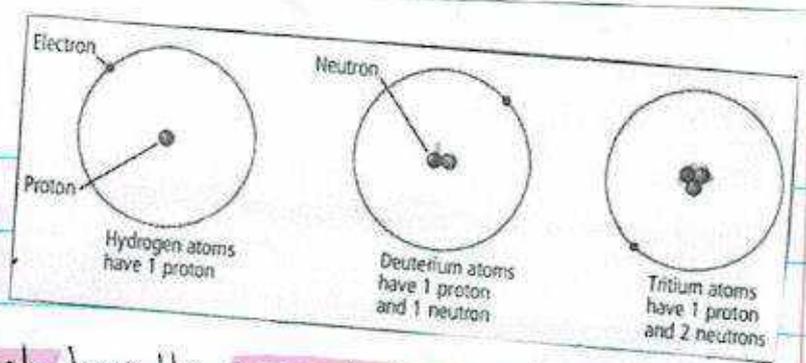
$$P = \text{proton no. / atomic no.}$$



$$\text{No. P} = \text{No. E}$$

subatomic particle	symbol	relative mass	relative charge
proton	p	1	+1
neutron	n	1	no charge
electron	e	0.00054	-1

Isotopes:



- Atoms of the same element have the same proton number always but their neutron number differs.
- $\text{Nucleon no.} - \text{Proton no.} = \text{neutron no.}$
- IF an isotope is radioactive its nucleus is unstable so it decays over time = radioisotopes. (Harmful)
- medical + industry + military + Home uses of radioisotopes.

Electron Shells / Energy levels.

1st shell = 2 Electrons max

2nd shell = 8 Electrons max

3rd shell = 8 Electrons max e.c.t

• Electron configuration / structure of sodium is 2,8,1.

Horizontal Rows in periodic table is periods

- each element has one more electron than the last.

• Vertical column is called Groups

- Same number of electrons in outer shell / valency electrons.

- The number of electrons in outer shell is same as group number

noble gases have a full outer shell of electrons making their electronic structure very stable and unreactive.

Elements : made up of only one type of atom - cannot be broken down.

Compounds : Two or more different types of atoms joined by a chemical bond.

Molecular compounds - ionic compounds

Mixtures : contains two or more elements or compounds that aren't chemically bonded together - can separate with physical methods.

metals + non-metals :

- Density (g/cm^3) = $\frac{\text{mass g}}{\text{volume cm}^3}$

Metals :

- Have high density, high melting + boiling points, conducts, malleable, ductile, lustrous, sonorous.

Non-metals :

- opposite of metals.

oxide = oxygen

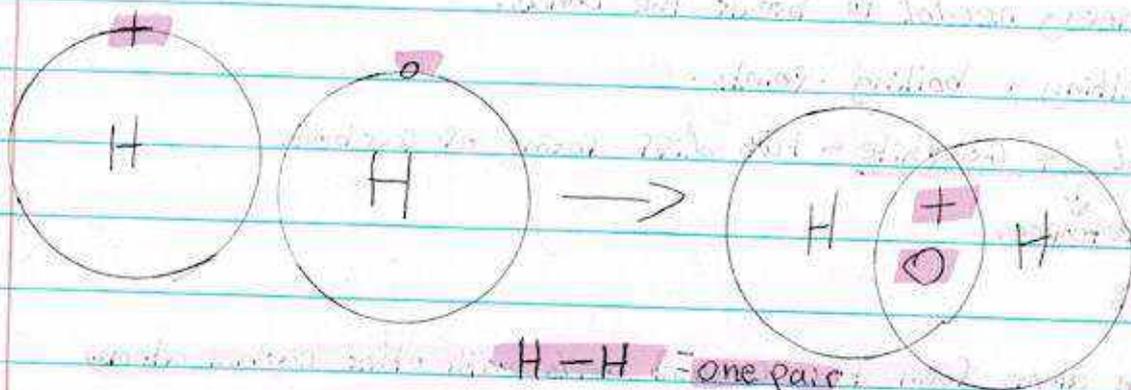
Structure + Bonding

Ionic Bonding : metal - non-metal - transfer

- ions are formed when atoms lose or gain 1+ electrons.
- positive charge = more protons = negative charge = more electrons
- stable octet = full outer shell of 8 electrons

• Attraction between positive + negative ions is an ionic bond.

• Covalent bonding is non-metals - share



• The electrons that aren't used in bonding = lone pairs

• Double Bond = $N=N$ = Triple bond = $N\equiv N$

- Lattice: A Regular repeated arranged particles of a 3 dimensional structure.

• The weak attractive forces between molecules = intermolecular forces.

• The strong forces within the molecules = covalent bonds.

Ionic compounds	Covalent compounds
<ul style="list-style-type: none">• High melting + boiling points. Strong attractive forces.• Soluble in water, insoluble in organic solvents.• Conducts electricity when molten or dissolved in water. mobile.	<ul style="list-style-type: none">• low melting + boiling points. weak intermolecular forces.• insoluble in water• Don't conduct electricity cuz have no ions, only uncharged molecules.

* Giant Covalent Structures:

- Network of covalent bonds throughout a whole structure = macromolecules.
- Lot of energy needed to break the bonds.
- High melting + boiling points.

→ • Diamond + Graphite - two diff forms of carbon.

↓
allotropes

* Diamond:

- The carbon atoms form four covalent bonds with other carbon atoms.
- They link together to form a giant lattice.

* Graphite:

- Carbon atoms arranged in layers. Each atom linked to three others.
- Arranged in hexagons.
- Bond between layers is weak = Graphite can flake off which is why it's used as a lubricant + pencil lead, slippery, easily scratched.
- Conducts electricity: Carbon has four valence electrons but graphite only bonds to three so the fourth is a delocalised electron which can drift along the layers when voltage is applied causing it to be able to conduct electricity.

* Silicon (IV) oxide:

- Silicon dioxide • Similar structure to diamond.
- Each silicon atom is bonded to 4 oxygen atoms, Each oxygen atom is bonded to two silicon atoms.

* metallic bonding:

- metal atoms packed closely together = valence electrons move away from atom causing a sea of free delocalised electrons to form, surrounding a positively charged lattice.
- The positively charged metal ions in the lattice are held together by their strong attraction to the mobile electrons that move between the ions. This equals metallic bonding.
- High melting + boiling points
- Good conductors - due to vibration of atoms.
- malleable + ductile - layers can slide over each other which means new bonds can form leaving the metal with a different shape.

Formula + Equations

- Chemical symbol = O (Oxygen), S (Sulfur)
- Some symbols are not as clear because they will come from a different language - Greek, Arabic, Latin - e.g. Fe is iron - K is Potassium

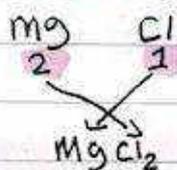
- we can work out the formula of a compound by knowing the valencies or combining powers of the element it contains.

Group	I	II	III	IV	V	VI	VII	VIII
valency	1	2	3	4	3	2	1	0

-> Taking away 8 - 8

magnesium chloride =

magnesium is in group II so its valency is 2
chlorine is in group VII so its valency is 1



<- swap numbers

if numbers are same they cancel out.

Simple rules for naming compound of 2 elements =

2 elements =

• If compound contains metal + non-metal we write metal first + ending of non-metal becomes 'ide'. e.g. Chlorine + magnesium = magnesium chloride.

• If compound contains 2 non-metals we write the one with the lower group number first (except hydrogen, this goes first). e.g. Nitrogen - group V - + oxygen - group VI - = nitrogen dioxide. If they are in same group the one further down goes first.

3 elements =

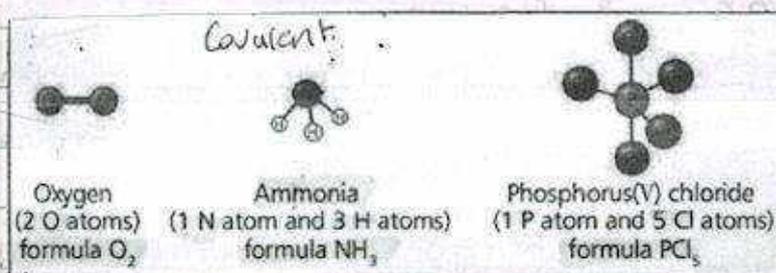
- oxygen + other element

x OH = hydroxide - NaOH = sodium hydroxide

x NO₃ = nitrate - SO₄ = sulfate - CO₃ = carbonate

Working out formulae

• If given a pic of a molecule showing all atoms + bonds it'll be easy to find formula. e.g. Diagrams

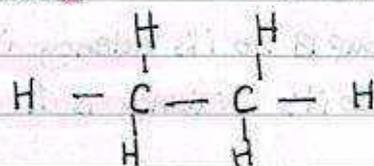


• For some molecules we can write 2 types of formulae

1. Molecular formula = Number of each type of atom present

2. Empirical formula = Simplest whole number ratio of atoms present.

e.g. structural formulae



• molecular formula = C₂H₆

• empirical formula = C₂H₆

1 Group I	2 II	3 H ⁺	4 III	5 IV	6 V	7 VI	8 VII	9 VIII
Li ⁺	Be ²⁺					O ²⁻		none
Na ⁺	Mg ²⁺		Al ³⁺					none
K ⁺	Ca ²⁺	Transition metals	no. metal					none

- Ionic compound formula:

1. Count numbers of positive ions
2. Count numbers of negative ions
3. Cancel the numbers down to simplest whole number ratio.

e.g. Zinc chloride =

$$\text{Zn} = 4 \quad \text{Cl} = 8 \quad \div \text{each by } 4 = \text{Zn} = 1 \quad \text{Cl} = 2$$

Formula = ZnCl_2

- You can work out the formula for an ionic compound knowing the charge of each ion.

* Compound ions:

ions containing more than one type of atom:

x NH_4^+ = ammonium ion

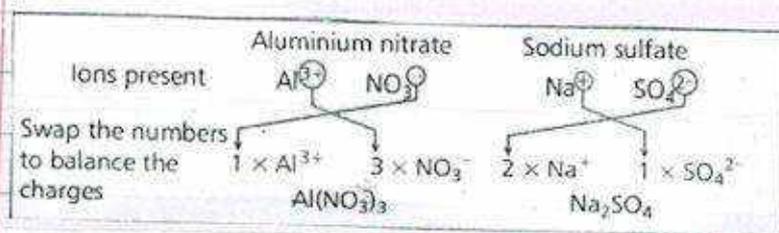
x OH^- = hydroxide ion

x NO_3^- = Nitrate ion

x SO_4^{2-} = Sulfate ion

x CO_3^{2-} = Carbonate ion

x HCO_3^- = hydrogen carbonate ion



* Chemical equations:

1. Word equations:

- Simplest type of equation

Magnesium + oxygen \rightarrow magnesium oxide

Nitrogen + oxygen $\xrightarrow[\text{pressure}]{\text{heat + catalyst}}$ ammonia

not very useful:

• Don't show numbers of molecules

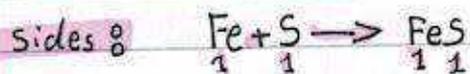
reactants = left - products = right

Chemical name

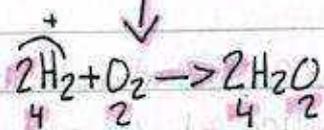
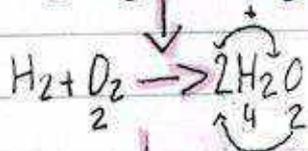
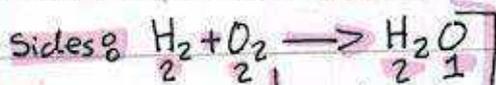
- Name can be very long
- Names for diff chemicals are not the same in diff langs.

2. Symbol equations

- Short hand way of describing a chemical.
- Balanced equation = must be same number of each type of atom on both



- Not balanced equation = Diff amount of numbers of each atom on both



How to make balanced

→ now a balanced equation

More symbol equations

When brackets are used, the small number at the bottom right of the brackets multiplies through what is in the brackets. So $\text{Mg}(\text{NO}_3)_2$ has 1 'atom' of magnesium, (1×2) 2 atoms of nitrogen and (2×3) 6 atoms of oxygen. If we write $2\text{Mg}(\text{NO}_3)_2$, we would have twice as many of each of these atoms: 2 magnesium, 4 nitrogen and 12 oxygen.

Some examples are given here:

Example 1: aluminium + water \longrightarrow aluminium + hydrogen
chloride hydroxide chloride



Balance the chlorine atoms $\text{AlCl}_3 + \text{H}_2\text{O} \longrightarrow \text{Al}(\text{OH})_3 + 3\text{HCl}$

Balance the oxygen atoms $\text{AlCl}_3 + 3\text{H}_2\text{O} \longrightarrow \text{Al}(\text{OH})_3 + 3\text{HCl}$

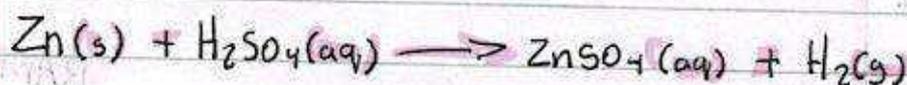
Diatomic molecules

- Hydrogen = H_2
- Nitrogen = N_2
- Oxygen = O_2
- Fluorine = F_2
- Chlorine = Cl_2
- Bromine = Br_2

- State symbols

shows if substance is liquid, solid, gas or dissolved in water.
 (l) (s) (g) (aq) = aqueous solution

e.g



3. Ionic equations: Shows only those ions that react

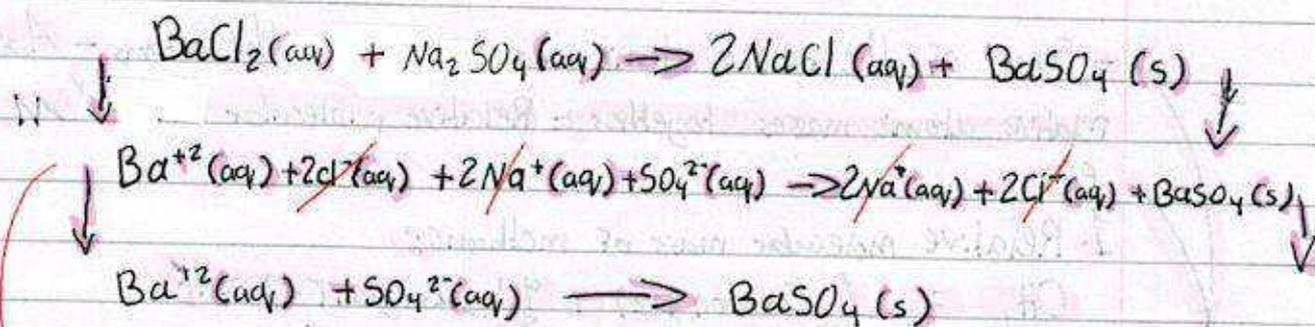
• ionic compounds dissolve in water = ions separate



1. Balanced ionic equation + state symbols

First = +
 Second = -
 separate

S X
 L X
 g X
 aq ✓



- Spectator ions: ions we cancel out

Chemical Calculations

- Atoms of diff elements have diff masses.
- Relative mass - how heavy one atom is compared to another
- Single atom mass is too small you can't weigh it. Scientists weigh a lot of atoms then compare them with the mass of the same number of 'standard' atoms
- Mass found = Relative atomic mass
- Average mass of naturally occurring atoms of an element on a scale

Naturally occurring isotopes

e.g
 Carbon-12
 = mass of 12 units

- It's easy to find the Relative atomic mass of an atom because it's usually found on most periodic tables:
 e.g.

17 → atomic no.
Cl
chlorine
35.45

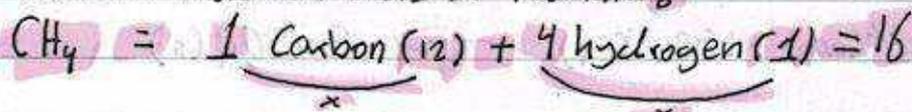
Relative atomic mass = $35.45 = \underline{35.5}$ ^{RAM}

- Can use relative atomic mass to compare how heavy an atom is with another

- Sum of the relative atomic masses of all the atoms - Add all the relative atomic masses together = Relative molecular mass / M_r

e.g.

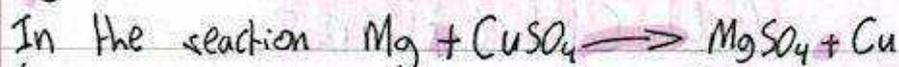
1. Relative molecular mass of methane:



- Same method for ionic compounds = Relative formula mass

- Work out how much product we get from given amount of reactant by simple proportions

e.g.



6.4g of copper are formed from 2.4g of magnesium - what mass of magnesium is needed to get 32g of copper?

6.4g copper is formed from 2.4g magnesium
 to get 32g = $\frac{32}{6.4} \times 2.4 = 12\text{g}$ magnesium

No. given
 Product × reactant

* Avogadro constant and the mole

- If we have 6×10^{23} atoms, ions or molecules of a substance, we have an amount which is easily weighed - no. of atoms, ions or molecules is called the Avogadro number/constant.

- Amount of substance with Avogadro number of particles is called the mole.

* Pair: 2

* dozens: 12

* Mole: 6.022×10^{23} \rightarrow Avogadro number = N_A

&

- 1 mole is simply the relative atomic mass/relative formula in grams.

\downarrow
2 moles = double

\downarrow
or Molar mass

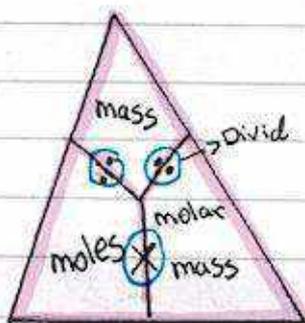
- To find number of moles of particles in something, use formula:

$$\boxed{\frac{\text{g}}{\text{g/mol}}} \leftarrow \text{Number of moles} = \frac{\text{mass of sub taken} \rightarrow \text{grams}}{\text{mass of 1 mole of sub} \rightarrow \text{molar mass}} \rightarrow \boxed{n = \frac{m}{M}}$$

- Rearrange formula to find mass of substance in given number of moles:

(a) mass of substance = number of moles \times mass of 1 mole of sub
 \downarrow
molar mass

or

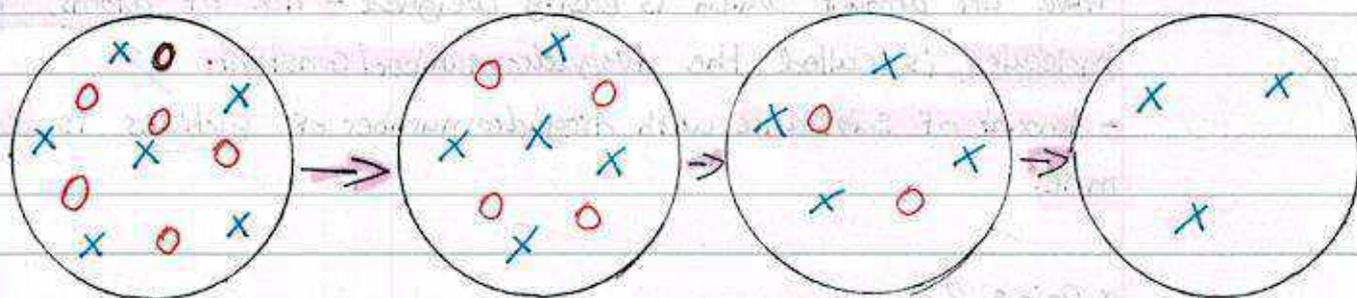


What mass of one reactant needed to add to another so they react and there is no waste. We need to know the ratio (stoichiometry). See bottom of page 61.

$$\boxed{\frac{?}{?} \times ? = ? \text{ g}}$$

- When carrying out a reaction we sometimes use/are left with an excess of one of the reactants while in the other reactant there is no excess.

- The reactant that gets used up is called the limiting reactant/reagent.



X Calcium carbonate

O Hydrochloric acid

- Each 1 mole of calcium carbonate 2 moles of hydrochloric acid is converted to product.

Gets used up faster. limiting reactant.

- How to find which reactant is limiting = which reactant has lower no. of moles?

1. Find mol

2. From the equation: 1 mol of magnesium reacts with 2 mol of HCl so to react completely, 0.05 mol magnesium will need to react with $2 \times 0.05 = 0.1$ mol of HCl.

But we have only 0.075 mol of HCl, so the HCl is the limiting reactant.

From this type of calculation, you can also find out by how much one reactant is in excess. In the example all the hydrochloric acid was used up.

Hydrochloric acid used up = 0.075 mol

From the equation: 2 mol hydrochloric acid reacts with 1 mol of magnesium.

So 0.075 mol hydrochloric acid will react with $\frac{0.075}{2} = 0.0375$ mol magnesium

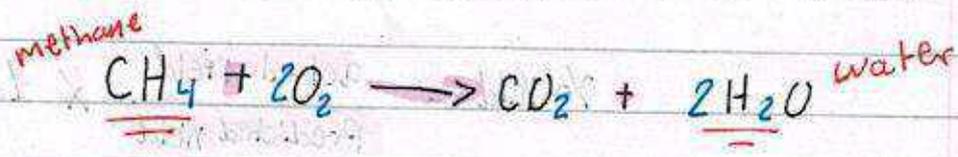
Therefore

Excess magnesium = moles of magnesium at start - moles of magnesium reacted
 $= 0.05 - 0.0375 = 0.0125$ mol

- How to find amount of product from given amount of reactant?

Calculate mass of water formed when 4g of methane is completely burned in oxygen?

C=12
H=1
O=16



$12 + (4 \times 1) = 2 \times (2 \times 1) + 16 =$
 $16\text{g methane} + 64\text{g oxygen} = 20\text{g}$

16g methane will give = 36g water

4g methane will give = $\frac{4}{16} \times 36 = 9\text{g water}$

* Percentages + Volumes

- Use formulae + Relative molecular masses to work out percentage by mass of particular element in a compound. Formula:

page 64 if need

$\% \text{ by mass} = \frac{\text{Relative atomic masses of Particular element}}{\text{Relative formula mass of Compound}} \times 100$

- Volume Calculations (Gas)

DM^3 Decimeter cubed

- o Same volume of gas = Same number of moles
- o molar gas volume = At Room temperature and pressure the volume of one mole of any gas is $24 \text{ dm}^3 / 24000 \text{ cm}^3$.
- o Reactions where gas is produced formula: $1 \text{ mol} = 24 \text{ dm}^3$

Volume of gas (dm^3) = number of moles of gas $\times 24$

Yield + Purity:

Yield = how much of a particular product you can get from the reactants in a chemical reaction. Percentage yield formula:

$$\% \text{ Yield} = \frac{\text{actual yield}}{\text{Predicted yield}} \times 100$$

- actual yield = Amount of product we get in a reaction. (given)
- Predicted/theoretical yield = Calculate maximum amount of product we can get from given amount of reactants using relative formula masses with equation.

Percentage Purity - similar to Percentage Yield. Formula:

$$\% \text{ Purity} = \frac{\text{mass of pure product}}{\text{mass of impure product}} \times 100$$

ionic compounds

Empirical formula = Finding formula of compound using masses of each element that combine to form the compound.

Shows simplest whole number ratio.

Analysis of a compound of tin (Sn) and chlorine (Cl) showed that the tin chloride contained 29.75g of tin and 35.5g of chlorine. Calculate the empirical formula of tin chloride. $A_r[\text{Sn}] = 119$; $A_r[\text{Cl}] = 35.5$

	Sn	Cl
Step 1: note the mass of each element	29.75g	35.5g
Step 2: divide by the relative atomic masses	$\frac{29.75}{119} = 0.25 \text{ mol}$	$\frac{35.5}{35.5} = 1.0 \text{ mol}$
Step 3: divide each by the lowest number of moles	$\frac{0.25}{0.25} = 1$	$\frac{1.0}{0.25} = 4$
Step 4: write the formula	SnCl_4	

works same way with percentages.

- Molecular formula = The actual number of atoms in a molecule.
e.g

methane :

x Empirical formula : CH_3

x molecular formula : C_2H_6

To find the molecular formula we need the relative formula mass of compound and empirical formula.

Worked example:

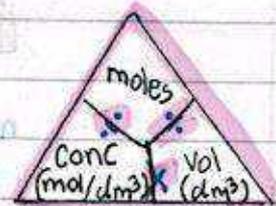
A compound has the empirical formula CH_2 . Its relative formula mass is 84. Calculate its molecular formula. $A[\text{C}] = 12$; $A[\text{H}] = 1$

Step 1: find the empirical formula mass $12 + (2 \times 1) = 14$

Step 2: divide relative formula mass by empirical formula mass $\frac{84}{14} = 6$

Step 3: multiply the empirical formula by the number calculated in Step 2 $6 \times \text{CH}_2 = \text{C}_6\text{H}_{12}$

~~Titration~~ :



Concentration = amount of solute dissolved in 1 dm^3 of solution.
Formula :

$$\text{dm}^3 = \text{cm}^3 / 1000$$
$$\text{Concentration (mol/dm}^3\text{)} = \frac{\text{number of moles of solute}}{\text{Volume of solution (dm}^3\text{)}}$$

Rearrange to find moles in given concentration + volume :

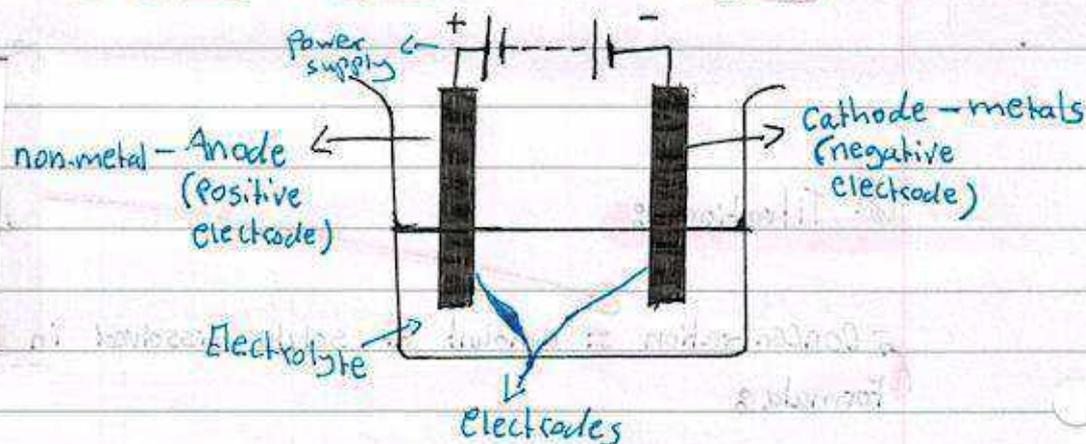
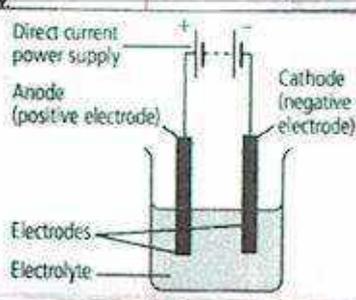
$$\text{Number of moles of solute} = \text{Concentration (mol/dm}^3\text{)} \times \text{Volume of solution (dm}^3\text{)}$$

The procedure needed

- Titration = Find the concentration of alkali needed to completely react with an acid using a procedure.

Electricity + Chemistry

- **Electrolysis**: Breaking down of ionic compound in molten or aqueous solution by passage of electricity.
- The electric current causes a chemical reaction that breaks down the ionic compound. The compound decomposes.
- **Electrolyte** is the compound that conducts electric and breaks down
- **Electrodes** are the rods that carry electric current. Normally inert.
- **Positive electrode = anode** - non-metal forms here
- **negative electrode = cathode** - metal forms here - hydrogen formed here



- Concentrated aqueous solutions of ionic compounds usually get hydrogen at the cathode, this is because it is low in the discharge series.

The lower down the series the more likely to get discharged

The series:

Positive ions: Na^+ Mg^{2+} Al^{3+} H^+ Cu^{2+}
 more likely to get discharged \rightarrow

Negative ions: SO_4^{2-} NO_3^- OH^- Cl^- Br^- I^-
 more likely to be discharged \rightarrow

- The electrolysis of brine is used to produce:

- o Chlorine = used to make solvents for treating drinking water and making bleaches
- o Hydrogen = used for making ammonia, margarine, as a fuel
- o sodium hydroxide = used for making soap + in extraction of aluminium

- When water is electrolysed a bit of sulfuric acid is added to improve its conductivity. hydrogen is formed at cathode and oxygen at anode (or halogens)

ions to atoms:

when an electric current is applied to a molten ionic compound the positive ions (cations) move to cathode, negative ions (anions) move to anode. ions reach electrode they gain (reduction) or lose (oxidation) an ion.

- equation that shows what's happening at one electrode = half equation

- metals gain electrons - non-metals lose electrons.

- when purifying copper we remove any impurities, we do this with electrolysis - (copper refining) / (metal refining)

- An impure strip of copper gets connected to anode (positive) and a thin pure strip of copper is connected to cathode (negative).

- Pure copper from the anode is transferred to cathode. impurities fall to bottom.

x inert electrodes g. isn't reactive electrode. not active electrode
product may differ when using

- Electroplating g. putting a thin layer of one metal on top of another.

• Connect object to be electroplated to cathode (negative). plating metal is connected to anode (positive).

• The process happens same way as metal refining.

x reasons for electroplating g

• Protection of metals from corrosion + improving appearance.

- = +
+ = -
oil rig

any thing made of metal



found in the mineral ore bauxite

most abundant metal in earth crust

- In 1886 first small drops of liquid aluminium were extracted from aluminium oxide.

- Extracting aluminium oxide from bauxite ore we need to purify the ore

• Crush + mix with sodium hydroxide. Aluminium oxide reacts to sodium hydroxide and dissolves.

• Sodium hydroxide impurities are insoluble so filtered out.

• Then gets heated to make pure aluminium oxide.

- Extracting aluminium from aluminium oxide

• Aluminium needs to be molten. (gets dissolved in molten cryolite + calcium fluoride

• Then we carry out electrolysis using graphite electrodes (to lower melting point).

• At cathode aluminium ions gain electrons \rightarrow reduced to aluminium metal.

• At anode the oxide ions lose electrons + get oxidised.

- Conductors of metals

• Allow electricity to flow through them easily.

• metals and graphite conduct electricity. Cus they have mobile outer-shell of electrons

- insulators e.g. plastics, glass, wood, clay, ceramic stuff

• does not conduct electricity \rightarrow does not have mobile electrons

• Thermosetting plastics used in areas with high heat (avoid melting)

Chemical Changes

- Physical Changes (Topic 1.2)

- no new substance is formed
- Easily reversible

- Chemical changes

- Formation of new substances
- Heat is always taken in or given out
- Cannot be easily reversed (some can be by changing conditions)

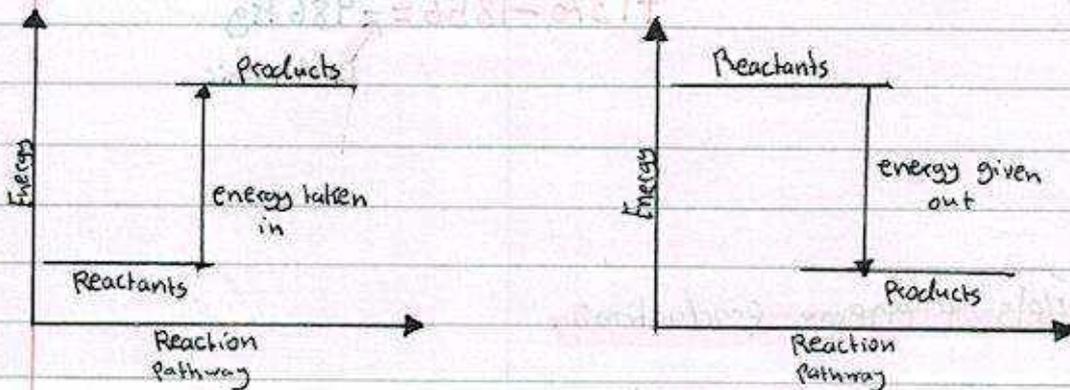
- Exothermic reactions

- Processes that release heat energy to surroundings

- Endothermic reactions

- Processes that take in heat energy from surroundings (absorb heat)

- Energy level diagrams



Endothermic
Reaction

Exothermic reaction

- Bond breaking is endothermic - Energy taken in greater than given out
- Bond making is exothermic - Energy given out is greater than taking in

- Diff between energy of reactants and products shown by ΔH (delta H)

- Energy given out ΔH given negative sign $\Delta H = -212 \text{ kJ/mol}$
- Energy is absorbed/taken in ΔH given positive sign $\Delta H = +212 \text{ kJ/mol}$

exo
endo

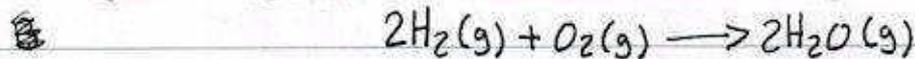
- Bond energies

• Amount of energy needed to break a bond. Symbol = E

• E.g $E(\text{O}=\text{O}) = +498 \text{ kJ/mol}$

• Calculate how much energy is released or absorbed

calculate change in



Energy values

H-H 436 , O=O 498 , O-H 464

Bonds Broken : endothermic

• $2\text{H}-\text{H} = 2 \times 436 = 872 \text{ kJ}$

• $1\text{O}=\text{O} = 498$

• $872 + 498 = 1370 \text{ kJ}$

Bonds Formed : exothermic

$4\text{O}-\text{H} = 4 \times 464$

$= 1856 \text{ kJ}$

$$+1370 - 1856 = -486 \text{ kJ}$$

Exothermic

Fuels + energy production

- Fuels : Substance that can be burned to release energy. exothermic reaction

* Coal : Decay of plants in swampy areas in absence of oxygen. plant remains \rightarrow coal

x Petroleum: Complex mixture of compounds containing Carbon + oxygen. Formed from bodies of tiny animals + plants that sank to sea bed. Pressure from rocks above changed them into petroleum.

x Natural gas: largely methane. Found underground trapped in layers of ice or near areas rich with petroleum.

- The radioisotope uranium-235 (^{235}U) used as a nuclear fuel. Uranium fuel rods lowered into reactor, then bombarded with high speed neutrons, collisions cause ~~react~~ nucleus to split, then large amount of energy released.

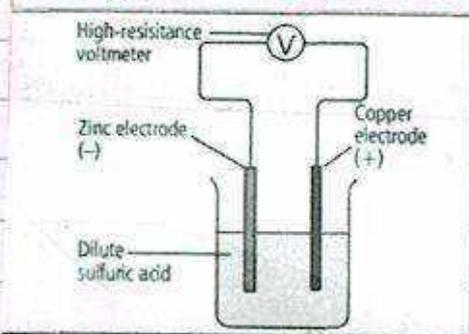
- Electrochemical cells: 2 metals of different reactivity dipping into an electrolyte.

o The more reactive the metal is the better at releasing electrons.

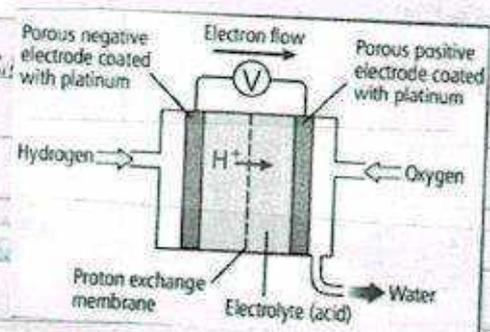
o Difference in ability of atoms to release electrons causes a voltage. Voltage we get depends on metals used.

o loses power after a time - no longer produce voltage. Cus one of the reactants is used up.

o Many substances found are harmful + difficult to dispose of.



Electrochemical



Fuel cell

x Hydrogen burns in oxygen = water. can be used to supply electrical energy continuously. we use a fuel cell.

- Fuel cells: tub platinum electrodes ^{coated onto fibrous material - allows gases to pass}

o Hydrogen is bubbled through the negative electrode. oxygen bubbled through positive electrode.

Two main types of fuel cells

- 1. Acidic electrolyte
- 2. Alkaline electrolyte

over All reactions are same and produce water

Fuel cells (advantages)

- x product only water
- x produce more energy per gram of fuel
- x operate with high efficiency

Tells us how rapidly products are formed

Rate of Reaction

Find rate of reactions

- x measure how quickly reactants are used up
- x measure how quickly products are formed

Temperature, concentration, particle size, light can affect rate of reaction

Control variable: keep constant

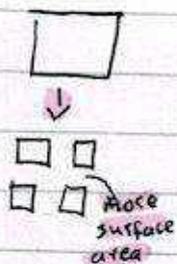
independent variable: changing by researcher

Formula:

$$\text{Rate of reaction} = \frac{\text{change in amount of product/reactant}}{\text{time}}$$

Reactant completely used up - no excess = limiting reactant

Activation energy = amount of energy needed to start a reaction



- Cut something up into smaller pieces, surface + number of particles which can react are increased.

- Catalyst: substance that speeds up chemical reactions

1. Solid Catalysts

o Allows reactants to get closer together on its surface so less energy is needed for reaction to happen.

- Depending on the Gas / particles the higher the concentration the faster or the lower the concentration the slower

→ Similar thing happen when adding pressure. More collision with particles.

- At higher temperatures the rate of reaction is fast

At lower the rate of reaction is slow

x when heated particles gain energy, move faster, collide more.

- Reactions started by light: photochemical reactions

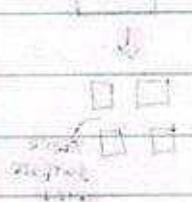
e.g. photosynthesis (catalyst in this is chlorophyll)

plants use sunlight to make glucose

More light = faster reaction

if in volume of container is constant
to start a reaction

- Get something of like smaller pieces
which can react at a faster rate



- Catalyst & activation energy
Catalyst: allows reactants to get closer together on its surface so less energy is needed for reaction to happen

- Depending on the size of particles the higher the concentration the faster
or the lower the concentration the slower

collision theory

2 particles being broken when colliding together. more collisions with particles

- All higher temperatures the rate of reaction is faster

At lower the rate of reaction is slower

2 water molecules. particles gain energy when faster, collide more

- Reactions started by light & photochemical reactions

eg photosynthesis (catalyzed in this is chlorophyll)

light was sunlight to make glucose

more light = faster reaction

Forward reaction
 \rightleftharpoons
 backward/Reverse reaction

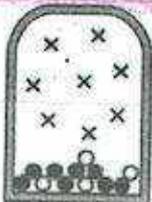
Chemical Reactions

- Reactions which can go in either direction = Reversible reactions
 - The symbol \rightleftharpoons is for a reversible change. Used for reactions we call equilibrium reactions.
 - Hydrated: water is present in structure.
 - Anhydrous: 'without water', water isn't present.
 - Water of crystallisation e.g. water in salt
- We can reverse something that is anhydrous into hydrated simply by adding water.

- Equilibrium: Both the forward and reverse reactions are going on at the same time.

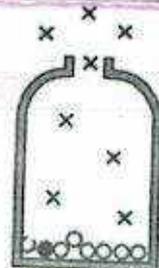
- o Closed system: Reactants or products must not escape from mixture
- o Open system: Reactants or products might/will escape from mixture

● $\text{CaCO}_3(\text{s})$
 ○ $\text{CaO}(\text{s})$
 × $\text{CO}_2(\text{g})$



$\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

This is a closed system. The calcium carbonate is decomposing to calcium oxide but the carbon dioxide is not lost.



$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

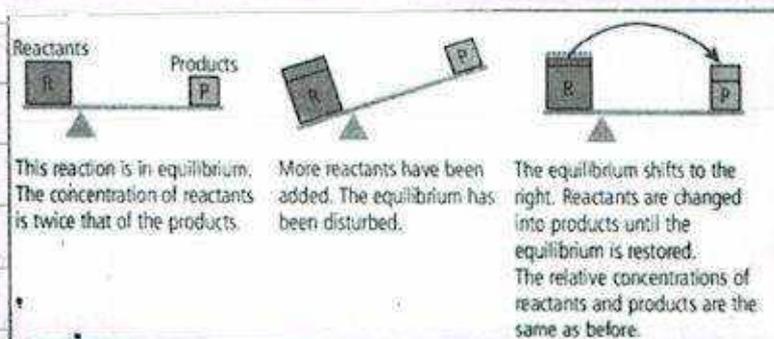
This is an open system. The calcium carbonate is decomposing to calcium oxide and the carbon dioxide is lost.

- Dynamic equilibrium:
 - Reactants $\xrightarrow{\text{changed to}}$ Products
 - Products $\xrightarrow{\text{changed to}}$ reactants
- Rate of forward + Reverse reaction are same at equilibrium. Fixed concentrations.
- Concentration of products greater = equilibrium to the right. Favours products
- Concentration of reactants greater = equilibrium to the left. Favours reactants

* Shifting equilibrium

- Changing concentration, temperature, pressure has an effect on equilibrium.
- * ~~Reaction~~ Reaction tries to oppose changes
- Catalysts speeds up reactions equally. (no effect on position)

- o create equilibrium
- Concentration of Reactants increased = moves to right to form more products
- Concentration of products increased = moves to left to form more reactants



- Changing pressure

o Only affects reactions with gas

o increasing pressure moves to side with smaller gas volume - The right - Endo -

o Decreasing pressure pushes reaction to the left - Exo -

- Changing temperature

x Reaction is exothermic in forward, reverse will be endothermic

o For exothermic reaction & Temperature increases moves in favour of reverse reaction

o for endothermic reaction & Temperature increases moves in favour of forward reaction

* Redox reactions & when Reduction + oxidation take place together

- Oil Rig ^{opposite} → when oxygen is involved = oxidation → gain oxygen Reduction → lose oxygen

- Reducing agents Reduces another substance.

- oxidising agent oxidises another substance.

- Oxidation states

o Roman numeral after name of element in compound e.g iron(II) chloride

Tells us

x Type of ion present in compound

x How oxidised an element in a compound is
o can also be useful for non-metallic elements

- Electron transfer in redox

o Oil Rig

- Set of rules to find oxidation state of element in compound

o Element that is uncombined = oxidation state 0

o Ionic compounds = oxidation state same as charge on ion

o Total oxidation state of all elements in compound is 0

- Potassium manganate(VII) in acidic solution is good oxidising agent. oxidises substance, colour changes from purple to colourless. Test for reducing agents.

- Potassium iodide in acidic solution is good reducing agent. Reduces substance, colour changes from colourless to brown. Test for oxidising agents.

Acids +
Bases

- Common lab acids

o Hydrochloric acid HCl o Sulfuric acid H_2SO_4 o Nitric acid HNO_3

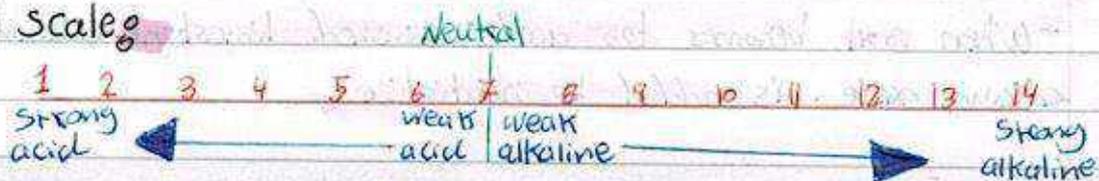
- When dissolved in water acids form hydrogen ions (makes solution acidic). H^+

- Common lab alkalis

o Sodium hydroxide NaOH o Calcium hydroxide Ca(OH)_2 o Ammonia NH_3

- When dissolved in water forms hydroxide ions OH^-

o pH scale



- lower pH = higher concentration hydrogen

- Higher pH = Higher concentration hydroxide

- Universal indicator - Mix of indicators shows range of colours depending on pH

Red → Blue = alkali
Red = Not alkali

- Chlorides = Salts

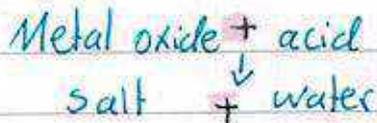
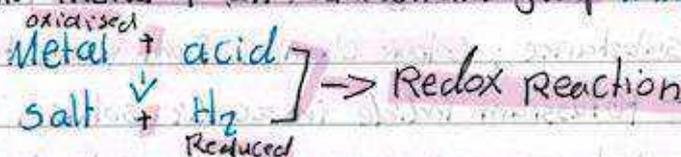
- Litmus is an indicator

Blue → Red = acidic
Blue = Not acidic

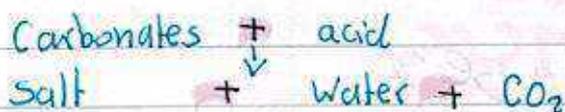
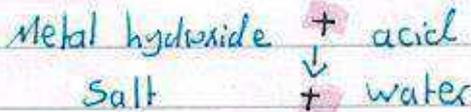
Chemical properties of acids

- Metals react with dilute acids to form salt + hydrogen

- Salt = compound when metal + an ammonium group NH_4 replaces hydrogen in acid.

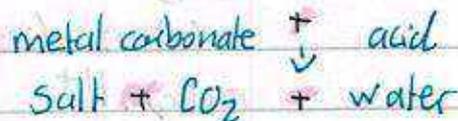
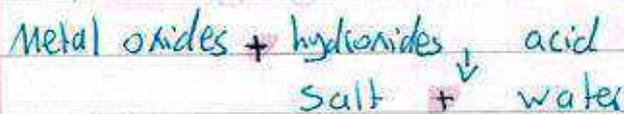


Neutralisation
Reactions



oxides,
carbonates
of metals

- A Base is a substance that can react with an acid soluble in water.

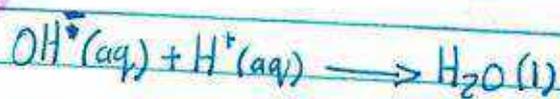


Volatile



- When soil becomes too acidic crushed limestone - calcium carbonate / lime calcium oxide - is added to neutralise.

- Equation for ~~the~~ neutralisation:



- All hydrogen atoms have a single proton

- Hydrogen ions is just a proton

- An acid is a proton donor - gives to base

- A Base is a proton acceptor - Takes from acid

- Strong Acids:

o Lab acids

o Completely ionised when dissolved in water

- Weak Acids:

o Organic acids

o Partly ionised when dissolved in water

- Strong Bases:

o All hydroxides of alkali metals

o Completely ionised in water

- Weak Bases:

o Ammonia

o Partly ionised

- How to tell if strong/weak by electrical conductivity, pH + rate of reaction.

- Oxides are compounds of metals/non-metals with oxygen. Four types.
We tell difference by chemical reactions:

o Basic oxides:

• Metal oxides

• Direct contact/combination with oxygen

• React with acids \rightarrow Salt + water

• Don't react with alkalis.

o Acidic oxides:

- Non-metals
- Direct Reaction with oxygen
- React with alkalis \rightarrow Salt + water
- React with bases (metal oxides) when heated
- React with water \rightarrow acidic solutions

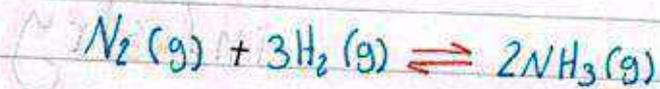
o Neutral oxides:

- Don't react with acids or bases
- lower oxides of non-metals: Nitrogen(I) oxide - Nitrogen(II) oxide
- Carbon monoxide.

o Amphoteric oxides:

- Acidic + Basic properties
- Salt formed when react with acids
- Complex Salts formed when react with alkali.

Temperature 450°C - Pressure 200 atmospheres. They combine.



* About 15% is converted to ammonia.

o The reaction to make ammonia is equilibrium and exothermic. The yield of ammonia increases with an increase of pressure. increasing the pressure shifts the direction to lower volume.

Yield of ammonia decreases with increasing temperature because increase in temperature favours endothermic (Reverse).

- Sulfur + Sulfuric acid

o Sulfur

x Yellow non-metal element

x found uncombined

x manufacture sulfuric acid

o Sulfur dioxide

x Colourless poisonous gas

x Highly acid when touches moist surfaces

x manufacture sulfuric acid

x Bleach, preserve food + drink

o Sulfuric acid

x Dibasic acid (on reaction 2 hydrogen ions can be formed per molecule)

The chemical

industry

- Fertilisers

- o plants need nitrogen, phosphorus and potassium
- o Fertiliser is a substance added to soil to replace the essential plant nutrients that have been lost (NPK fertilisers).
- o Single fertilisers = No potassium - phosphorus.

o Many Fertilisers contain ammonium

made by Haber Process

The sulfuric acid for phosphoric acid is made by Contact Process

o Many crops don't grow in acidic conditions, so lime is added which neutralises it. Stronger alkali displaces ammonia which escapes as a gas.

- Making ammonia

o German chemist Fritz Haber found nitrogen can combine with hydrogen at high pressure + temperature to make ammonia. Haber Process.

o Hydrogen = natural gas or Cracking ethane

o Nitrogen = from air

- Haber Process

x Nitrogen (1 vol) + hydrogen (3 vol) compressed

x Gases pass into large tank (converter). Contains iron catalyst