Redox

reduction (lose O) — oxidation (gain O)

① oxygen

reduction for Fe in Fe_2O_3

- $Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$

oxidation for Al

red for Zn in ZnO

- $CO + ZnO \rightarrow Zn + CO_2$

oxd carbon in CO

red Cu in CuO

- $H_2 + CuO \rightarrow Cu + H_2O$

oxd for H_2

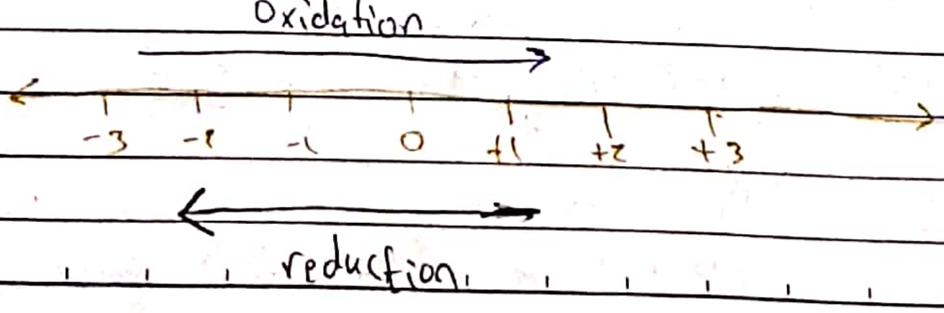
② Hydrogen — gain H — lose H

red Cl in HCl

- $H_2S + Cl_2 \rightarrow HCl + S$

oxd S in H_2S

③ oxidation state — decrease — increase



Rduction

- $\overset{0}{Zn} + \overset{2+}{Cu}\overset{1-}{Cl_2} \longrightarrow \overset{2+}{Zn}\overset{1-}{Cl_2} + \overset{0}{Cu}$

Oxid.

* Rules for oxidation state.

1) The oxidation state for any free element = zero.

Polyatomic (P_4, S_8)

monoatomic (Na, Li, Cu, Ca...)

diatomic (H_2, N_2, I_2, Cl_2)

2) The oxid. num. of any atom in a compound from

g1 = +1 Li, Na, K, Rb, Cs, Fr

g II = 2+ Mg, Ca, Sr, Ba

g III = 3+ Al

g VII = -1 F

3) The oxid. num. of H (+1), except with metal is metal hydroxide (-1).

$\overset{+1}{CH_4}$, $\overset{+1}{NH_3}$, $\overset{\text{metal}}{Na}\overset{-1}{H}$, $\overset{\text{metal}}{Ca}\overset{+1}{H_2}$

4) oxidation state of oxygen (-2), except in

OF_2 : 2+

Peroxide : -1 O^{2-} : oxide , O_2^{2-} : peroxide

- sodium oxide : $\overset{+1}{Na} + O^{2-} \longrightarrow NaO_2$
- sodium peroxide: $\overset{+1}{Na} + \overset{2-}{O_2} \longrightarrow \overset{+1}{Na_2}\overset{-1}{O_2}$

5) The sum of all oxid. state in a compound = 0, an ion = charge of the ion.

$\overset{+1}{Na}Cl$ $1+x=0$ $x=-1$

$\overset{+1}{H}ClO^{2-}$ $1+2+x=0$ $\cancel{-1}+x=0$ $x=+1$

$N_2O_5^{2-}$ $5(-2)+2x=0$

$\cancel{-10}, +2x = \cancel{0}$ $2x=10$ $x=5$

$\underline{N}\overset{2-}{O}_3^{-1}$ $x + \overset{-6}{3(-2)} = -1$ $x - 6 = -1$ $\overset{+6}{}$ $\underline{x = \cancel{4}5}$

$\underline{S}\overset{-2}{O}_3^{2-}$ $x + \overset{-6}{3(-2)} = -2$ $x - 6 = -2$ $\underline{x = 4}$.

$\underline{Cr}_2\overset{2-}{O}_7^{2-}$ $2x + \overset{-14}{7(-2)} = -2$ $2x - 14 = -2$ $2x = 12$ $\underline{x = 6}$

* Find the oxidation state of each of the underlined species.

$\overset{3+}{\underline{Fe}_{(2)}}O_3^{2-}$ $2x + \overset{-6}{3(-2)} = 0$

$2x - 6 = 0$ $\cancel{2}x = \cancel{6}$ $\underline{x = 3}$.

$\overset{2+}{\underline{O}}\overset{-1}{F_2}$ $x + \overset{-2}{2(-1)} = 0$ $x - 2 = 0$ $\underline{x = 2}$.

$\overset{6+}{\underline{S}}\overset{-2}{O}_4^{2-}$ $x + \overset{-8}{4(-2)} = -2$ $x - 8 = -2$

$\underline{x = 6}$.

$\overset{5+}{\underline{P}}\overset{2-}{O}_4^{3-}$ $x + \overset{-8}{4(-2)} = -3$

$= x - 8 = -3$

$= \underline{x = 5}$

Oxidation + Reduction

* $\overset{0}{Al} + \overset{3+}{Fe}_2\overset{2-}{O}_3 \longrightarrow \overset{3+}{Al}_2\overset{2-}{O}_3 + \overset{0}{Fe}$

- oxid : $Al^{0} \longrightarrow Al^{3+}$
- redu: $Fe^{3+} \longrightarrow Fe^{-2}$

* $\underset{\downarrow}{\underline{\overset{+7}{\underset{x}{Mn}}\overset{2-}{O}_4}}{}^{-1} + Fe^{2+} + H^{+} \longrightarrow Mn^{2+} + Fe^{3+} + \overset{+1}{H}_2O^{-2}$

$x + \overset{-8}{\widehat{4(-2)}} = -1 \qquad x - \cancel{8} = -\cancel{1}$

$\underline{x = 7}.$

- Oxid : $Fe^{2+} \longrightarrow Fe^{3+}$
- redu : $Mn^{7+} \longrightarrow Mn^{2+}$

* $\underset{\downarrow}{\underline{\overset{+6}{\underset{x}{Cr}}_2\overset{2-}{O}_7}}{}^{2-} + Sn^{2+} + H^{+} \longrightarrow Cr^{3+} + Sn^{4+} + \overset{+1}{H}_2\overset{2-}{O}$

$2x + \overset{-14}{\widehat{7(-2)}} = -2 \qquad 2x - \cancel{14} = -\cancel{2}$

$2x = 12 \qquad \underline{x = 6}.$

- oxid: $Sn^{2+} \longrightarrow Sn^{4+}$
- redu: $Cr^{+6} \longrightarrow Cr^{3+}$

- Writing balanced Half ionic equation

1) Atoms
2) the charge: by adding $\bar{e}$s to the side with greater charge by the difference.

* we add $\bar{e}$ to the bigger charge
($- < 0 < +$)

- $\overset{0}{Al} + \overset{3+}{Fe}_2\overset{2-}{O}_3 \rightarrow \overset{3+}{Al}_2\overset{2-}{O}_3 + \overset{0}{Fe}$

Half of oxidation: $Al \rightarrow Al^{3+} + 3\bar{e}$
Half of reduction: $Fe^{3+} + 3\bar{e} \rightarrow Fe$
overall reaction: $Al + Fe^{3+} \rightarrow Al^{3+} + Fe$

- $\overset{0}{Ca} + \overset{2+}{Zn}\overset{1-}{Cl}_2 \rightarrow \overset{2+}{Ca}\overset{1-}{Cl}_2 + \overset{0}{Zn}$

H.O.O.: $Ca \rightarrow Ca^{2+} + 2\bar{e}$
H.O.R.: $Zn^{2+} + 2\bar{e} \rightarrow Zn$

Reactants → products
oxid → $\bar{e}$ "lose $\bar{e}$"
red $\bar{e}$ → "gain $\bar{e}$"

Overall reaction: $Ca + Zn^{2+} \rightarrow$ ~~Zn~~ $Ca^{2+} + Zn$

- $\overset{0}{Na} + \overset{1+}{H}\overset{1-}{Cl} \rightarrow \overset{1+}{Na}\overset{1-}{Cl} + H_2$

Oxsn Na

H.O.O: $2\,(2Na \rightarrow 2Na^{1+} + 1\bar{e})$
H.O.R: $\overset{+1}{2H} + 2\bar{e} \rightarrow \overset{0}{H_2}$
Overall rxn: $2Na + 2H^{+1} \rightarrow 2Na^{+1} + H_2$

- $\overset{0}{Ca} + \overset{3+}{Al}\overset{1+}{Cl}_3 \rightarrow \overset{2+}{Ca}\overset{1-}{Cl}_2 + \overset{0}{Al}$

H.O.O: $3\,(3Ca \rightarrow 3Ca^{2+} + 2e^{-})$
H.O.R: $2\,(2Al^{3+} + 3\bar{e} \rightarrow 2Al)$
O.rxn: $3Ca + 2Al^{3+} \rightarrow 3Ca^{2+} + 2Al$

- $H_2\overset{2-}{S} + \overset{0}{Cl_2} \rightarrow H\overset{-1}{Cl} + S^{0}$

H.O.O $S^{2-} \rightarrow S + 2\bar{e}$

H.O.R $Cl_2 + 2\bar{e} \rightarrow \textcircled{2}Cl^-$

O-rxn: $S^{2-} + Cl_2 \rightarrow S + 2Cl^-$

- $\overset{0}{Al} + \overset{+1}{H}\overset{-1}{Cl} \rightarrow \overset{3+}{Al}\overset{-1}{Cl_3} + \overset{0}{H_2}$

H.O.O: $2\,Al^{0} \rightarrow 2Al^{3+} + 3\bar{e}$

H.O.R: $2H^{+1} + 2\bar{e} \rightarrow 3H_2^{0}$

-3

O.rxn: $2Al + 6H^{+1} \rightarrow 2Al^{3+} + 3H_2$

- $\overset{7+}{Mn}\overset{2-}{O_4}^{-1} + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + \overset{+1}{H_2}\overset{-2}{O}$

H.O.O: $5Fe^{2+} \rightarrow 5Fe^{3+} + 1\bar{e}$ $3-2=1$ $3>2$

H.O.R: $Mn^{7+} + 5\bar{e} \rightarrow Mn^{2+}$ $7-2=5$ $7>2$

O.rxn: $5Fe^{2+} + Mn^{7+} \rightarrow 5Fe^{3+} + Mn^{2+}$

- $\overset{+6}{Cr_2}\overset{-2}{O_7} + Sn^{2+} + H^+ \rightarrow Cr^{3+} + Sn^{4+} + \overset{+1}{H_2}\overset{2-}{O}$

H.O.O: 3 ($3Sn^{2+} \rightarrow 3Sn^{+4} + 2\bar{e}$ $4-2=2$ $4>2$

H.O.R: 2 ($2Cr^{+6} + 3\bar{e} \rightarrow 2Cr^{3+}$ $6-3=3$ $3<6$

O.rxn: $3Sn^{2+} + 2Cr^{6+} \rightarrow 3Sn^{+4} + 2Cr^{3+}$

Oxidising / reducing agent

- **Oxidising agent / oxidant:**
  the substance that itself reduced & causes the other substance to be oxidise.

- **reducing agent "reductant"**
  the substance that itself oxidised & causes the other substance to be reduced.

	reducing oxidation	oxidising reduction
O	gain O	lose O
H	lose H	gain H
oxid state	↑	↓
ē transfer	lose ē	gain ē

if the substance is an ion in a compound the agent is the compound itself

- $\overset{0}{Al} + \overset{3+}{Fe_2}\overset{2-}{O_3} \longrightarrow \overset{3+}{Al_2}\overset{2-}{O_3} + \overset{0}{Fe}$
  - oxidation: Al
  - reduction: Fe^{3+}
  - reducing agent: Al
  - oxidising agent: Fe_2O_3

- $\overset{2+}{C}\overset{2-}{O} + \overset{2+}{Zn}\overset{2-}{O} \longrightarrow \overset{0}{Zn} + \overset{4+}{C}\overset{2-}{O_2}$

  $C + 2(-2) = 0$ $+4 = C$
  - oxidation: C^{2+}
  - reduction: Zn^{2+}
  - reducing agent: CO
  - oxidising agent: ZnO

- $\overset{+1}{H_2}S^{2-} + \overset{0}{Cl_2} \longrightarrow \overset{+1}{H}Cl^{-1} + S^{0}$

  S $2- \to 0$
  - oxidant: ~~S~~ Cl_2
  - reductant: H_2S

  $Cl^{0} \to Cl^{-1}$ $0 > -1$

- $\overset{7+}{Mn}\overset{2-}{O_4}^{-1} + Fe^{2+} + H^{+} \longrightarrow Mn^{2+} + Fe^{3+} + \overset{+1}{H_2}O^{-2}$
  - oxidant: MnO_4^{-1}
  - reductant: Fe^{2+}

* Most common oxidising agents

① oxygen

- $\overset{0}{Fe} + O_2^{0} \rightarrow \overset{3+}{Fe_2}O_3^{2-}$

- $\overset{0}{S} + O_2^{0} \rightarrow \overset{4+}{S}\overset{2-}{O_2}$

② Acidify pottasium manganate $KMnO_4 / H^+$

- $\overset{7+}{Mn}\overset{2-}{O_4}^{-1} \longrightarrow Mn^{2+}$

Purple colourless

③ Acidify potassium dichromate $K_2Cr_2O_7 / H^+$

- $\overset{6+}{Cr_2}\overset{2-}{O_7}^{2-} \longrightarrow Cr^{3+}$

orange green

④ Halogens

- $2\bar{e} + F_2 \rightarrow 2F^-$

yellow

- $2\bar{e} + Cl_2 \rightarrow 2Cl^-$

green-yellow gas

- $2\bar{e} + Br_2 \rightarrow 2Br^-$

red-brown

- $2\bar{e} + I_2 \rightarrow 2I^-$

purple gas

black solid

red-brown solid

colourless solution

* Most common reducing agent

① Hydrogen

- $\overset{0}{H_2} + \underbrace{\overset{2+}{Cu}O^{2-} \longrightarrow \overset{+1}{H_2}\overset{2-}{O} + Cu^{0}}$ H_2 reductant,

② carbon & carbon monoxide

- $\overset{0}{C} + \underbrace{\overset{3+}{Fe_2}O_3^{2-} \longrightarrow \overset{0}{Fe}} + \overset{2+}{C}O^{2-}$ C reductant,

- $\overset{2+}{C}O^{2-} + \underbrace{\overset{2+}{Zn}O \longrightarrow \overset{4+}{C}O_2 + Zn^{0}}$ CO reductant.

③ Iodine / halide

- $2I^{-} \longrightarrow I_2 + 2\bar{e}$

④ Metals

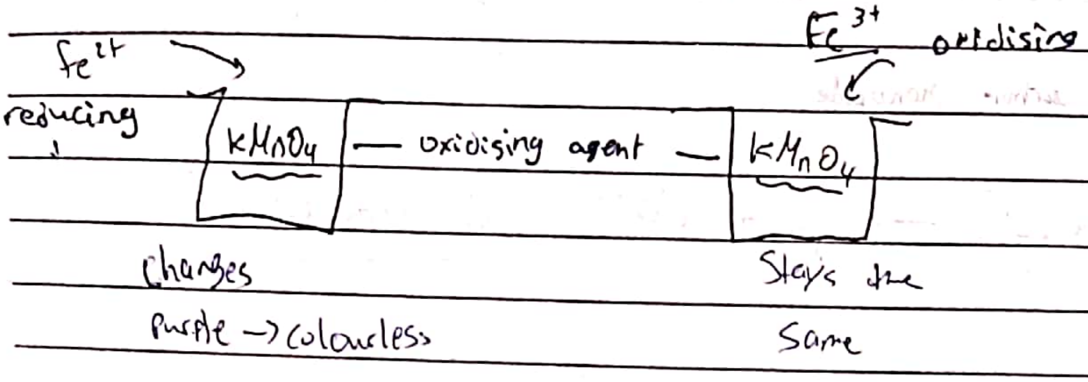
metal more reactive more likely to lose $\bar{e}$ more likely to oxidise. more likely to be a reductant.

- K → strongest reductant
  - K^{+} → weakest oxidant
- Na
- Li
- Ca
- Mg
- Al
- C, CO
- Zn
- Fe
- Pb
- H
- Cu
- Ag → weakest reductant
  - Ag^{+} → strongest oxidant

Q:- Fe^{2+} is a reducing agent

Fe^{3+} is an oxidising agent

(1) record the observation in each of the following reactions.



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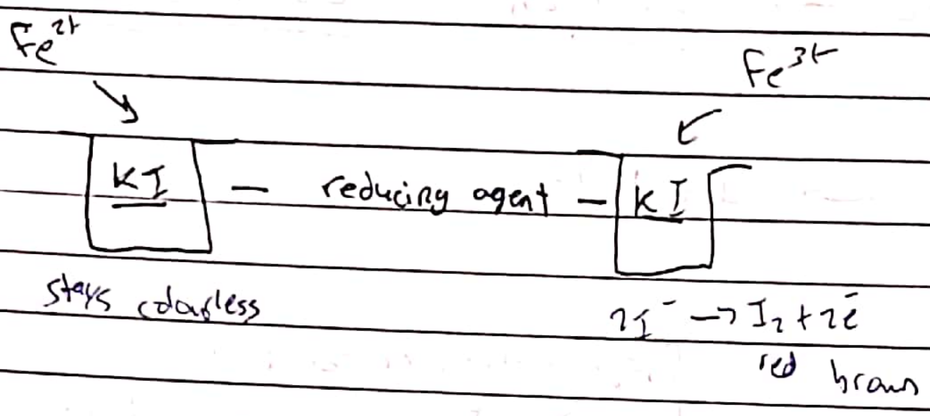

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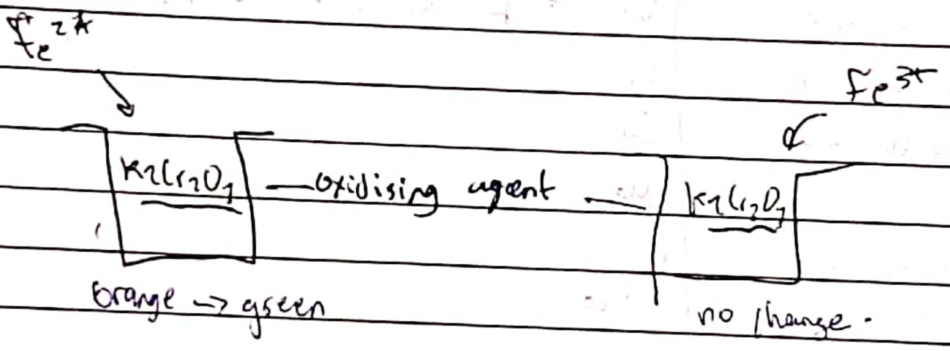


* - $KMnO_4$ Purple → colourless

- $K_2Cr_2O_7$ orange → green
- Iodine black solid or red-brown → red-brow liquid → purple gas

- F_2 (flurine) → yellow

Cl_2 → green yellow

Br_2 red-brown

Electrolysis

- electricity
Analysis
"Breaking down"

- **Electrolysis**: Breaking down chemical compounds (ionic) when molten or aqueous by passing electricity.

- why the ionic compounds don't conduct elec when solid?
no mobile ions.

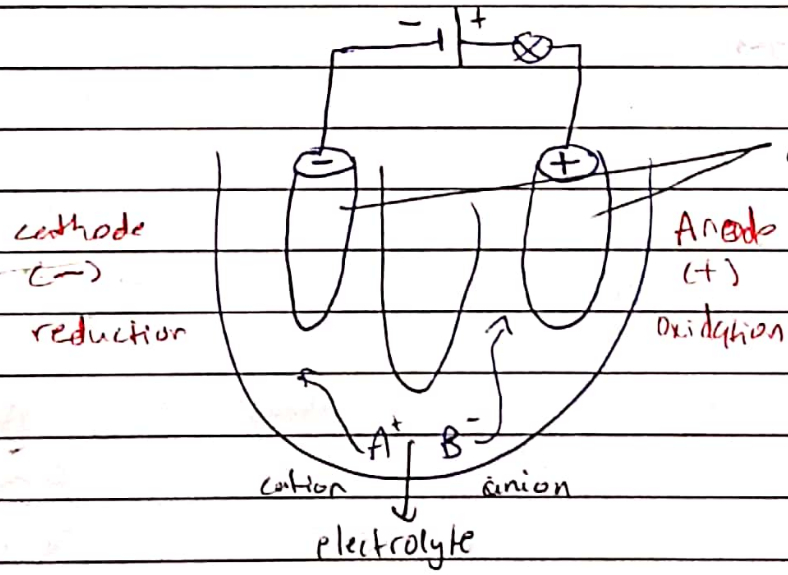
- ——— conduct elec when molten?
mobile ions (free moving).

$NaCl_{(s)} \xrightarrow{heat} Na^+_{(L)} + Cl^-_{(L)}$

$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$

- **Electrolyte**: a chemical compound that conducts elec. when molten or aq.

Electrolysis cell



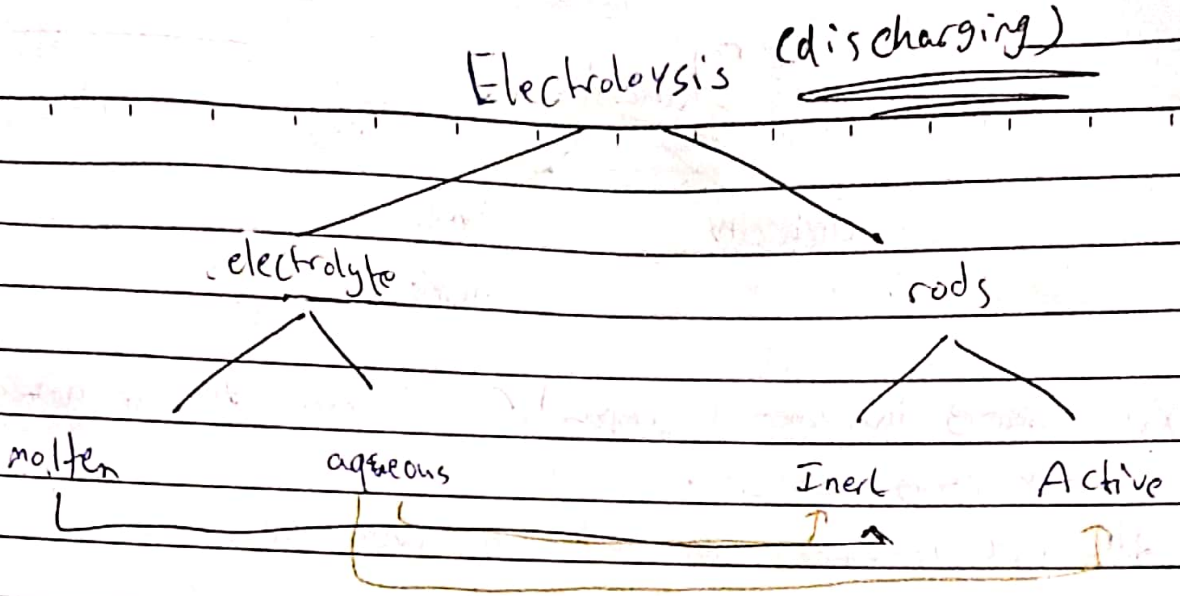
electrods (rods)

Inert
unreactive
- graphite
- Pt

Reactive
- Cu
- Ag

- cathode (−): the -ve rod that attracts +ve ions where reduction occurs.
- anode (+): the +ve rod that attracts -ve ions where oxidation occurs.

Electroloysis (discharging)

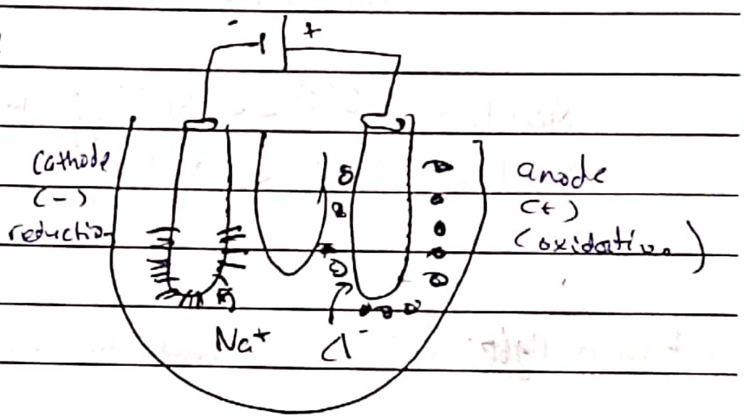


* electrolysis for Molten electrolyte using Intert rods (graphite)

NaCl • electrolyte is used w/

Total ions: Na^+ Cl^-

Cathode anode



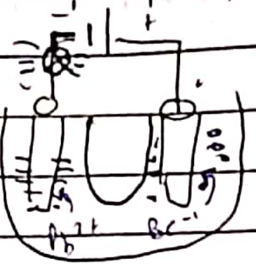
$1e^- + Na^+ \rightarrow Na$ $2Cl^- \rightarrow Cl_2 + 2e^-$

deposite of metal green-yellow bubbles

F_2	Yellow gas
Cl_2	green-yellow gas
Br_2	red-brown
I_2	purple gas black solid red brown solution

Cu red-brown solid.

* molten lead (II) Bromide

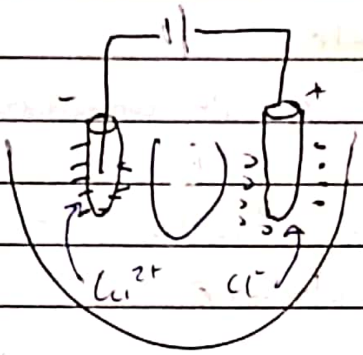


$PbBr_2$

Pb^{2+} Br^{-1} ← Total ions

cathode	anode	
$2\bar{e} + Pb^{2+} \rightarrow Pb$	$2Br^{-1} \rightarrow Br_2 + 2\bar{e}$	
deposit of metal	bubbles of red-brown gas	- electrolyte : used up. - bulb : lights on.

* CuCl2/graphite

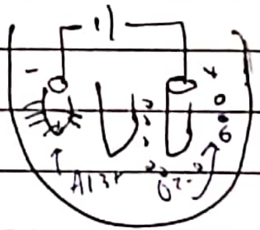


$CuCl_2$

Cu^{2+} Cl^-

Cathode	anode	
$2\bar{e} + Cu^{2+} \rightarrow Cu$	$2Cl^- \rightarrow Cl_2 + 2\bar{e}$	
deposit of red-brown solid	bubbles of green yellow gas	electrolyte: Consumed

* Al_2O_3 /graphite



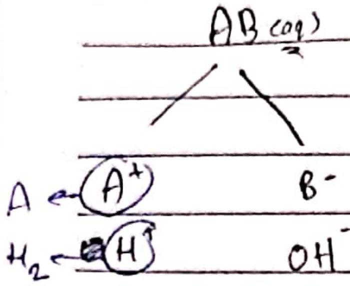
Al_2O_3

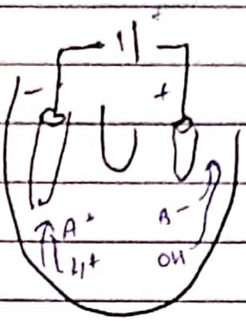
Al^{3+} O^{2-}

cathode	anode	
$3\bar{e} + Al^{3+} \rightarrow Al$	$2O^{2-} \rightarrow O_2 + 4\bar{e}$	
deposite of metal	bubbles of colourless gas	electrolyte: consumed.

graphite/pt

* electrolysis of aqeous electrolyte using Inert.





K^+
Na^+
Li^+
Ca^{2+}
Mg^{2+}
Al^{3+}
Zn^{2+}
$Fe^{2+/+3}$
Pb^{+2}
H^+
$Cu^{+2/+1}$
Ag^{+1}
Au^{+3}

At the cathode

The less reactive Ion is more likely to reduce & the more reactive stays in the electrolyte.

At the anode

Always OH^- except concentrated Halides
(Cl^-, Br^-, I^-)

- when the halide oxidise.

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

- when OH oxidise.

$$4OH \rightarrow 2H_2O + O_2 + 4e^-$$

bubbles
of colourless gas

* concentrated NaCl/graphite
(aq)

NaCl

Na^+ Cl^-

H^+ OH^-

cathode	anode	
$2H^+ + 2e^- \rightarrow H_2$	$2Cl^- \rightarrow 2Cl_2 + 2e^-$	-electrolyte:
bubbles of colourless gas	bubbles of green-yellow gas	NaOH

* $CuSO_4$ (aq) / graphite

$CuSO_4$

Cu^{2+} SO_4^{2-}

H^+ OH^-

- electrolyte: H_2SO_4.

cathode	anode
$2e^- + Cu^{2+} \rightarrow Cu$	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$
deposite of red-brown solid	bubbles of colourless gas

* dilute NaCl(aq)/graphite

NaCl

Na^+ Cl^-

H^+ OH^-

cathode	anode	
$2H + 2e^- \rightarrow H_2$	$4OH \rightarrow 2H_2O + O_2 + 4e^-$	electrolyte: NaCl
bubbles of colourless gas	bubbles of colourless gas	

* concentrated $KI_{(aq)}$ / graphite

KI

K^+ I^-

H^+ OH^-

cathode	anode	electrolyte: KOH
$2\bar{e} + H^+ \rightarrow H_2$	$2I^- \rightarrow I_2 + 2\bar{e}$	
bubbles of	bubbles of	
colourless gas		

* Concentrated $CuCl_2$ (aq) / graphite

$CuCl_2$

Cu^{2+} Cl^-

H^+ OH^-

cathode	anode	forget the electrolyte ~~[illegible]~~ for this compound
$2\bar{e} + Cu^{2+} \rightarrow Cu$	$2Cl^- \rightarrow 2\bar{e} + Cl_2$	
deposit of	bubbles of	
red-brown solid	green-yellow gas	electrolyte: less conc. $CuCl_2$ / not used up

* Concentrated sodium Chloride called Brine solution

$NaCl_{(aq)}$

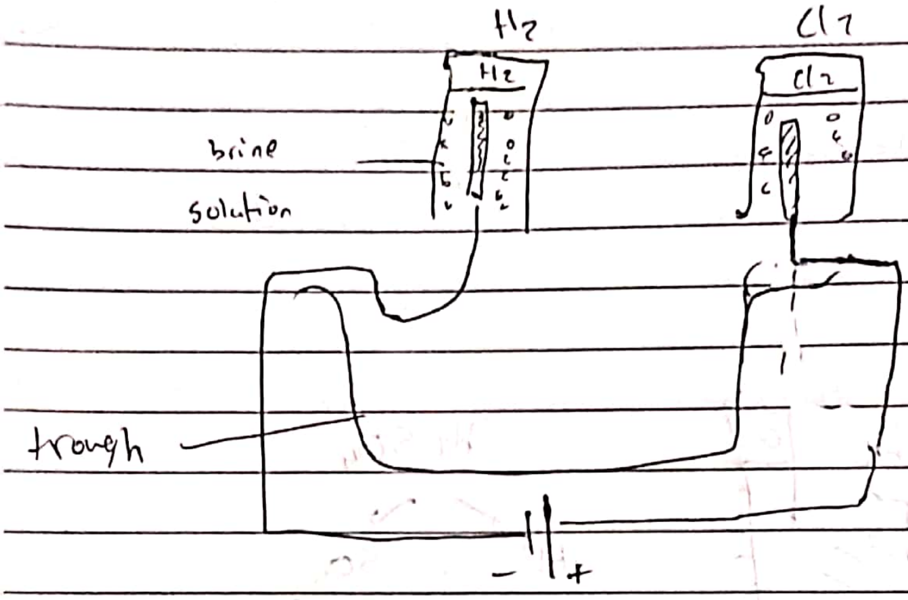
Na^+ Cl^-

H^+ OH^-

cathode	anode	
$2H^+ + 2\bar{e} \rightarrow H_2$	$2Cl^- \rightarrow Cl_2 + 2\bar{e}$	- electrolyte: NaOH
bubbles	bubbles	
colourless gas	of green- Yellow gas	

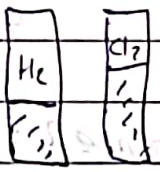
$NaCl$ = neutral = 7 pH

Q1: Plan an experiment to collect & measure the volume of H_2 & Cl_2 produced

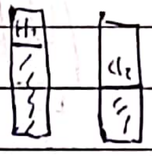


Q2:- the final appearance of the two measuring cylinders are:-

(a)



(b)



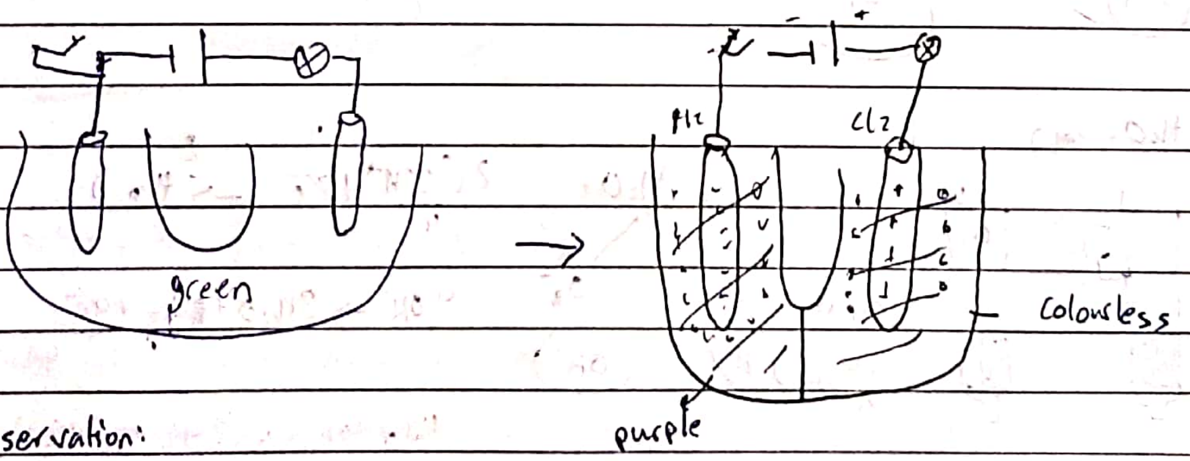
(c)

Half of oxid. : $2Cl^- \rightarrow Cl_2 + 2e^-$

Half of redu. : $2H^+ + 2e^- \rightarrow H_2$

overall: $2H^+ + 2Cl^- \rightarrow H_2 + Cl_2$

Q3:- Brine with universal indicator

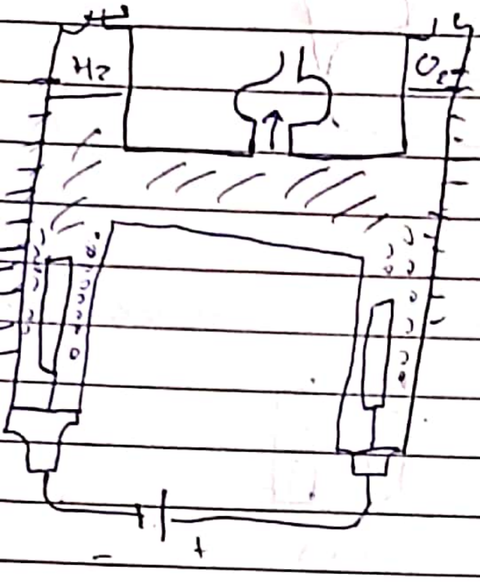


- observation:
  - bulb lights up
  - green-yellow gas (bubbles)
  - bubbles of colourless bubbles
  - around cathode the solution → purple (NaOH is alkali)
  - the ~~then~~ ~~solution~~ anode the solution becomes colourless since Cl_2 bleaches the colour

* the Cl_2 not immediately appear as H_2 produced
some Cl_2 dissolves in solution.

- electrolysis for H_2SO_4 (aq)

Hoff - Man
device



H_2SO_4

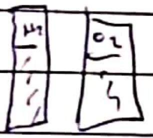
H^+ SO_4^{2-}

H^+ OH^-

H_2 O_2

appearance? ; more H in ratio
so more H will be used up.

Q2 :- the final appearance of two ...



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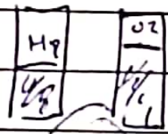
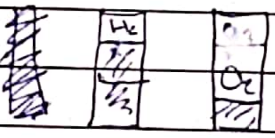
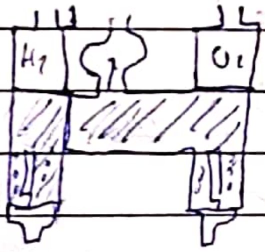


(b)



(c)

* H_2O_2 (aq)



H_2O_2

any of them can reduce [H^+ H^+

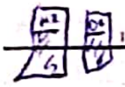
O_2^- OH^-

$2(\underset{4}{\cancel{2}}H^+ + \overset{4}{\cancel{2}}e^- \rightarrow \overset{2}{H_2})$

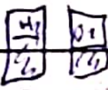
$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

$4H + 4OH \rightarrow 2H_2 + 2H_2O + O_2$

- final appearance



(a)



(b)

(c)

$2H \rightarrow O_2$

2 : 1

COLLEGE

* electrolysis for aqeous electrolyte using active rod.

- the rod must be the same metal ion in the electrolyte

$CuSO_{4(aq)} \xrightarrow{using} Cu$ rod

$AgNO_{3(aq)} \xrightarrow{using} Ag$ rod

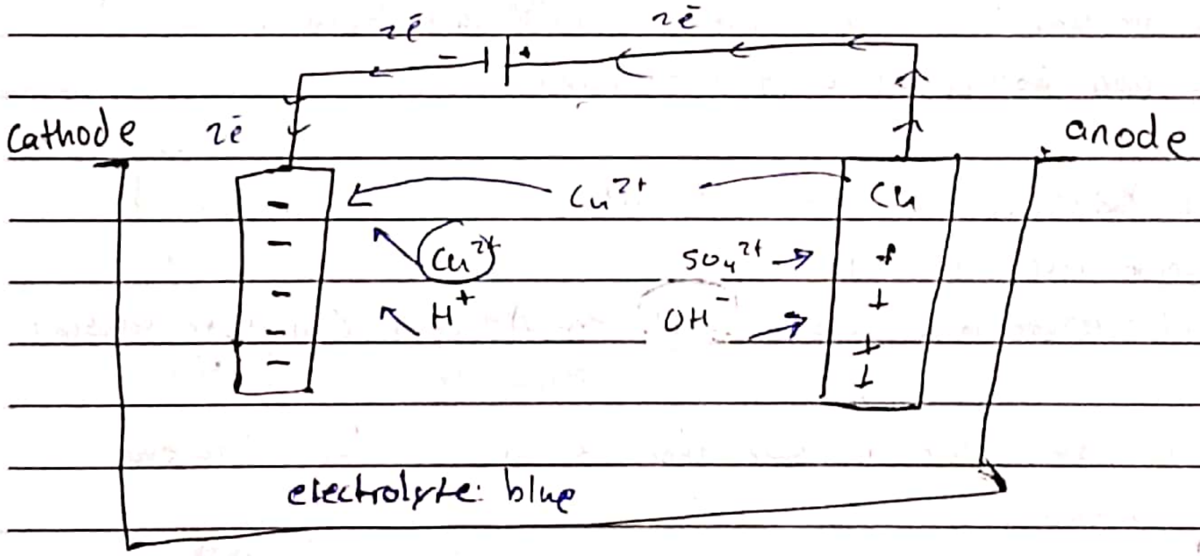
* electrolysis for aqeous $CuSO_4$ using Cu rod.

$CuSO_{4(aq)}$

Cu^{2+} H^+ | SO_4^{2-} OH^-] won't oxidise → the rod itself will

cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

anode: $Cu \rightarrow Cu^{2+} + 2e^-$



* electrolyte just to allow ions to move through it.

* why electrolyte doesn't change its colour [?]

because the anode oxidised + replaced the Cu^{2+} in the electrolyte with the same rate.

APplication on electrolysis

Molten/Inert
⇓
extraction of metals

ageous/Inert
⇓
prepare solution

ageous/active
- electroplating
- purfiying metals

① Electroplating coating metal with another metal using electrolysis.

Why? 1) to prevent rusting.
2) more attractive.

the spoon must be fully emerged

How to electroplate a metal spoon with silver? [8]

1- clean the spoon froom any impurities as oxide layer using sand paper to ensure a good sticking.

2- Make the spoon the cathode (-ve rod)

3- the anode must be Ag.

4- the electrolyte must contain Ag^+ eg. $AgNO_3$ (must be soluble)

5- turn on the circuit " be sure that the spoon is fully imerged in the solution.

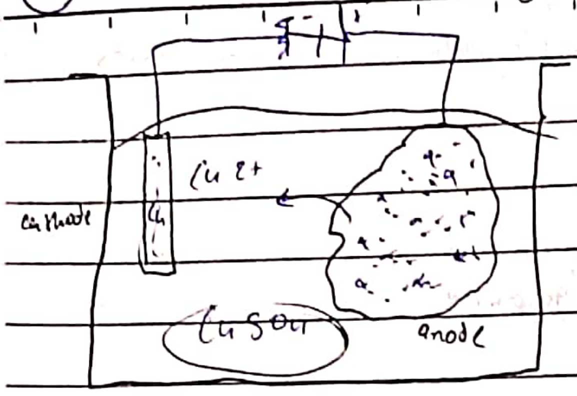
6- rotate the spoon to ensure equal distribution.

7- rinse with distilled water -

8- dry in oven

adeposition happens
On cathode

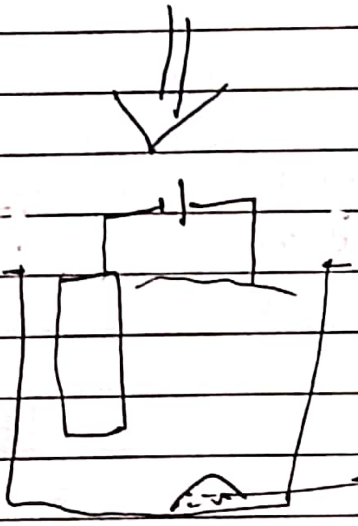
② Purifying metals / Refining copper



- Cu
- impurities Ag, Au, Zn

- Elements that are less (anode) reactive than the active (cathode) rod settle down, while more reactive elements displace the element in the electrolyte.

$Cu^{2+} + 2e^- \rightarrow Cu$ $Cu \rightarrow Cu^{2+} \rightarrow 2e^-$



$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

NO_2^{-1}	nitrite	SO_4^{-2}	sulfate
NO_3^{-1}	nitrate	SO_3^{-2}	sulfite
MnO_4^{-1}	manganate	$Cr_2O_7^{-2}$	dichromate
CH_3COO^{-1} $(C_2H_3O_2^{-1})$	ethanoate	CO_3^{-2}	carbonate
PO_4^{-3}	phosphate		

$Cu < ^{+1}_{+2}$

$Fe < ^{2+}_{3+}$

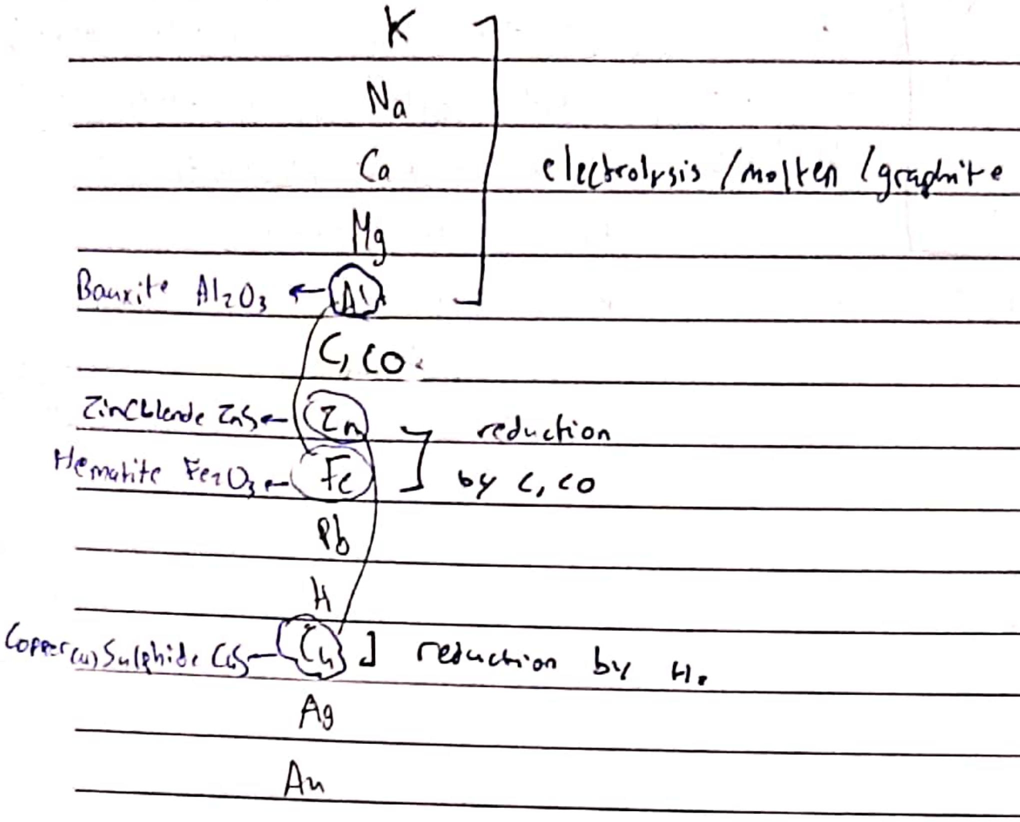
$Pb < ^{+2}_{+4}$

Ag +1

Zn +2

Extraction of metals

- The method of extraction depends on the position of metal in reactivity series.



extraction of Aluminum

ore :- Al_2O_3 - Bauxite

method :- Electrolysis for molten / graphite

\+ m.p. of Al_2O_3 is abt 2000°C

↑ how to solve this problem?

we dissolve Al_2O_3 in a molten cryolite Na_3AlF_6

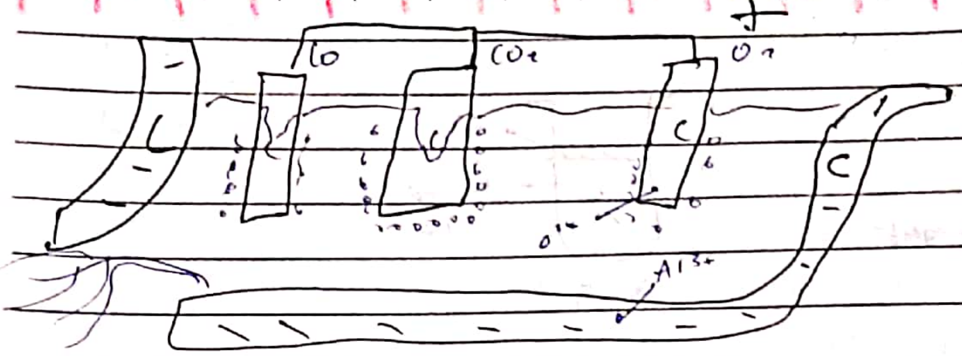
- to lower m.p. to 900°C so less cost
- to increase the electrical conductivity.

Al_2O_3

Al^{3+} O^{2-}

cathode anode

$Al^{3+} + 3e^- \rightarrow Al$ $2O^{2-} \rightarrow O_2 + 4e^-$



Al(l)

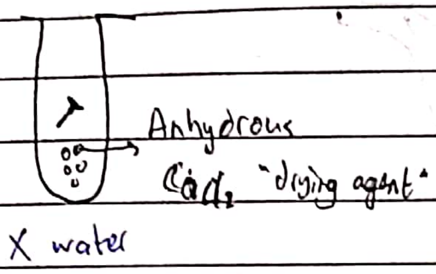
+ gases produced at anode.
  1- O_2
  2- CO_2] reaction of rods
  3- CO] with O_2 so we must replace them periodically.

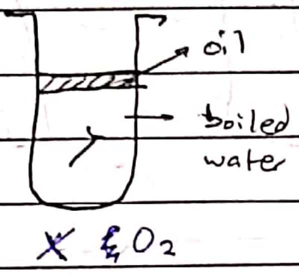
+ Al uses & why...

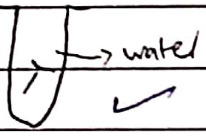
property	use
- low density	- aircraft bodies
- ductile	- electrical wires
- malleable	- window frame
	cooking utensils
- conduct electricity	- wires
- form an oxide layer which non-toxic	- food cans

slow reaction
6-7 days

Rust (iron hydroxide)







Rust: reaction of iron with both O_2 + H_2O

- Plan an experiment to show which rust prevension solution is better.

take a known mass of iron nail, apply a known volume of the first solution. Put them in a known volume of water for 1 week. Dry them + measure the mass again. Repeat the exp with 2nd solution.

conclusion: the exp which cause more increase in mass, is the worse solution.

* equation: $Fe + O_2 \xrightarrow{H_2O} Fe_2O_3 \cdot nH_2O$

- How to prevent rusting

1- painting
3- greasing
2- oiling
4- cover with plastic

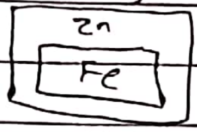
to prevent O_2 + H_2O from reaching iron.

better methods: 1- Galvanizing
2- Sacrificial protection
3- electroplating
4- Cathodic protection

better than electroplating bc if in elec the object is
↳ scartched it will rust, while in gal it will form an oxide layer bc Zn will react.

Galvanizing

Coating

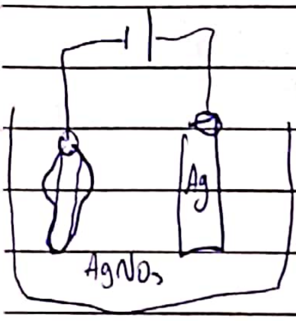


sacrifical

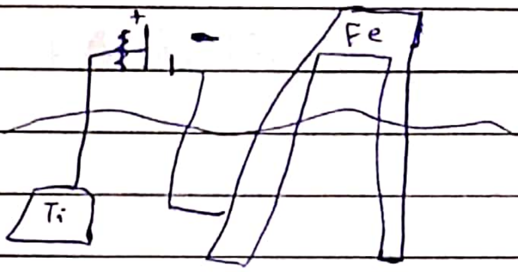
connecting

Zn + Mg are
more reactive than Fe so
they are more likely to oxidise, more
likely to lose $\bar{e}$'s so Fe is less likely to rust.

electroplating



cathodic

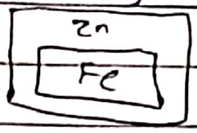


- in this case Fe will be the cathode meaning ~~it will~~ reduction will occur preventing it from oxidising

better than electroplating bc if in elec the object is
↳ scartched it will rust, while in gal it will form an oxide layer bc Zn will react.

Galvanizing

Couting

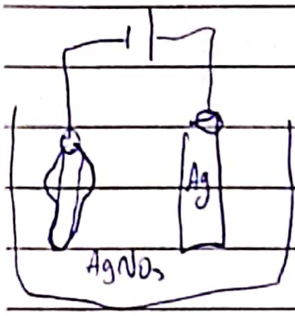


sacrifical

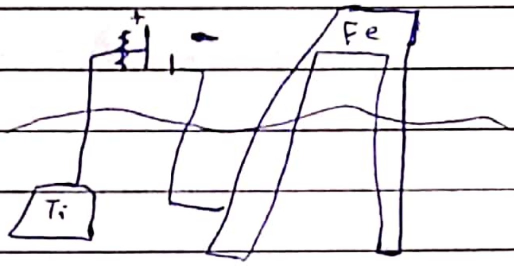
connecting

Zn + Mg are
more reactive than Fe so
they are more likely to oxidise, more
likely to lose $\bar{e}$'s so Fe is less likely to rust.

electroplating



cathodic



- in this case Fe will be the cathode meaning ~~it will~~ reduction will occur preventing it from oxidising

1 mol = 24 dm^3

Rate of reaction

rate = $\frac{\text{change in a quantity}}{\text{change in time}}$

Q : quantity

$R = \frac{\Delta Q}{\Delta t}$ / $\frac{\Delta \text{Mass}}{\Delta \text{time}}$ / $\frac{\Delta \text{volume of gas}}{\Delta \text{time}}$ / $\frac{\Delta pH}{\Delta \text{time}}$ / $\frac{\Delta \text{conc}}{\Delta \text{time}}$

unitless ↓ (pH)

g/s cm³/s /s mol/dm³·s

+ temperature (°C) ila in gases (kelvin).

* Concentration is measured through colour change / light intensity

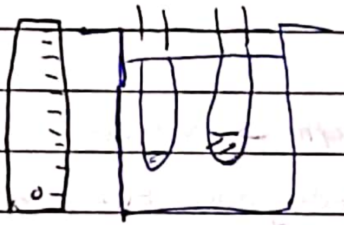
* $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + \cancel{A}\ NaNO_{3(aq)}$

↗ precipitation & replacement reaction

2 (aq) → 1 (s)

* How do measure rate of this reaction? (mass of precipitate)

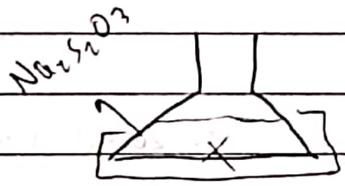
- using ruler, change in distance



change in distance

yellow solid

* $Na_2S_2O_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + SO_{2(g)} + \boxed{S_{2(g)}} + H_2O_{(aq)}$

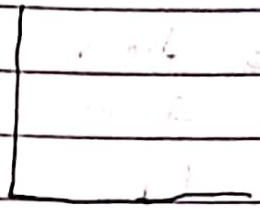


- when you put $Na_2S_2O_3$ start stopwatch.
- when cross dissappears stop the time.

Rate of reaction

$$\text{rate} = \frac{\text{change in quantity}}{\text{change in time}} = \frac{\Delta Q}{\Delta t}$$

ALWAYS y-axis → Q

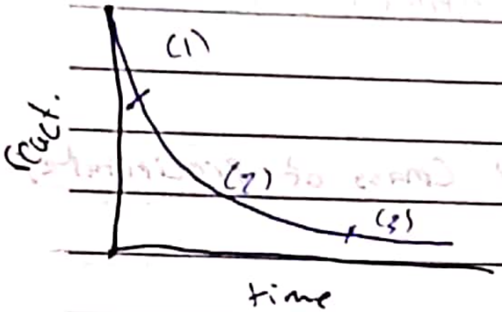


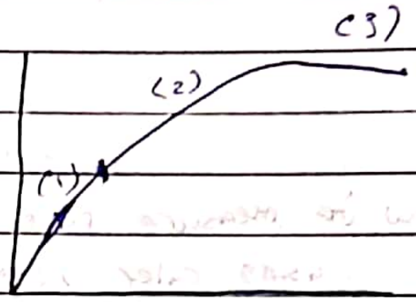
t ← ALWAYS x axis.

Measure rate of reaction

how fast the Reactants **consumed** per unit time

how fast the products **produced** per unit time





- Region (1): fastest rate => from graph → steepest

[3] At the beginning more amount of reactants, more particles, more effective collisions per unit time.

- Region (2): Slower rate => from the graph => less steep

less number of particles, so less number of effective collisions per unit time.

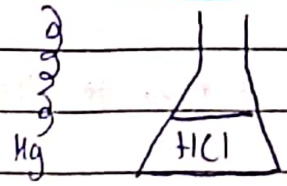
- Region (3): Reaction is over → gradient = 0 / horizontal line

no more limiting factor, so no more effective collisions

- For any chemical reaction there are 3 main conditions.

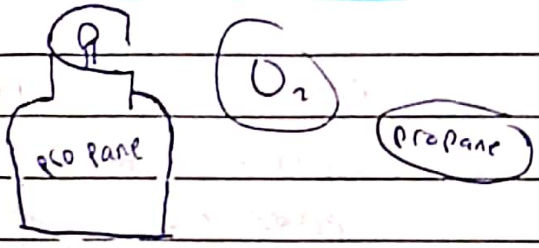
① $Cu + HCl \longrightarrow$ no reaction   the reactants must be suitable

② the reactants must collide



③ The collisions must be effective / min amount of energy to start the energy (activation energy)

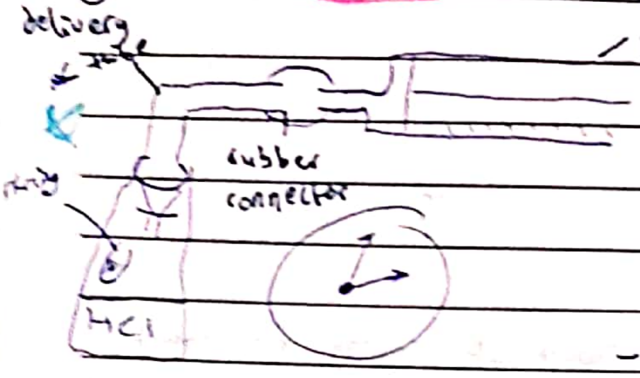
E_a



Measuring rate of reaction

$CaCO_{3(s)} + 2HCl \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

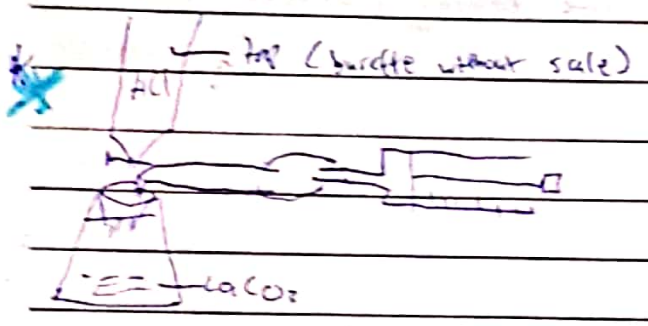
① Measure volume of gas per unit time.



suggest a practical method to prevent gas to escape.

- pull the string & start the stopwatch

-

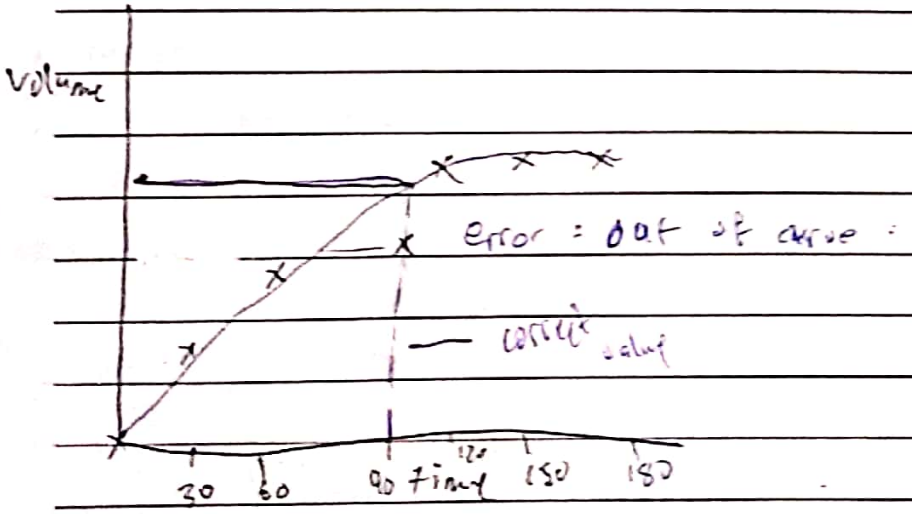


rubber connector has a small hole to release pressure.

time (s)	0	30	60	90	120	150	180
volume of CO_2 (cm³)	0	10	15	17	17.5	17.5	17.5

+10 +15 +7 0.5

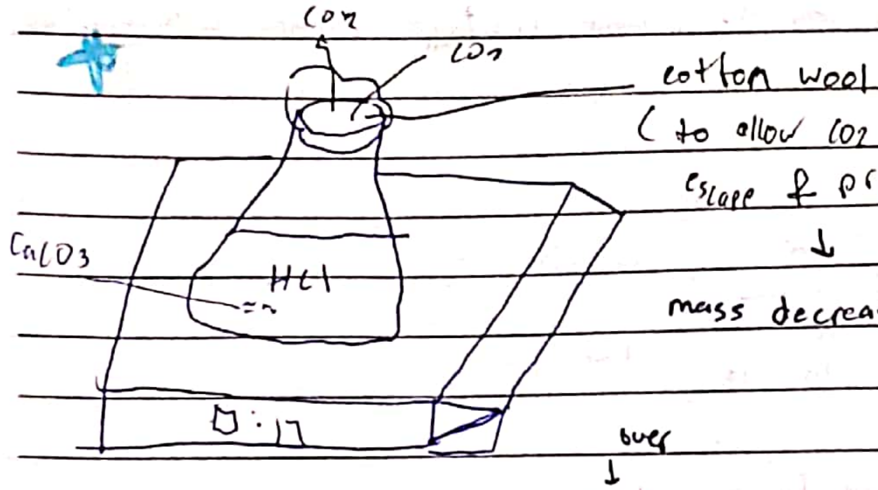
reaction is over



error : out of curve :

— correct value

② Measure the mass of conical flask + contents per unit time.

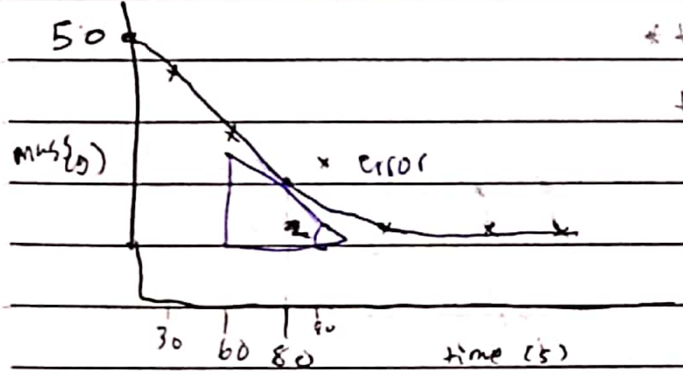


cotton wool
(to allow CO_2 to escape & prevent splashing)
↓
mass decrease bc CO_2 escapes

over
↓

time (s)	0	30	60	90	120	150	180
mass (g)	50	40	35	33	32.5	32.5	32.5

-10 -5 -2 -0.5

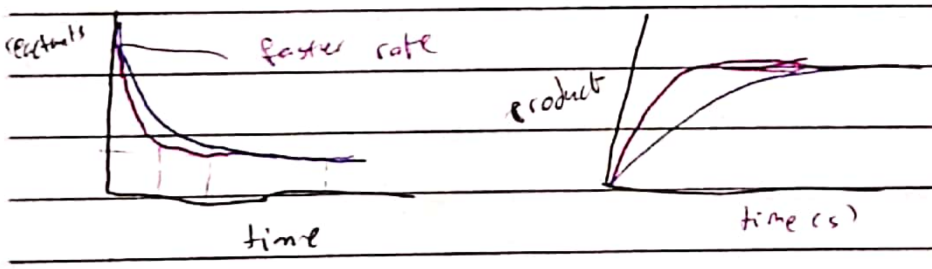


* to find the rate at specific time at = 20s, draw tangent

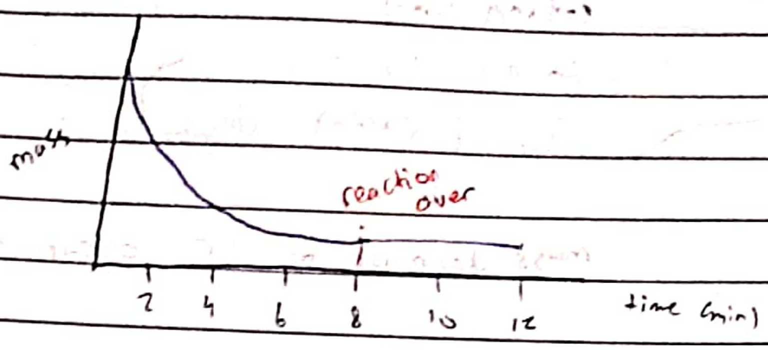
Increase rate of reaction

more product per same period of time
OR
same products per less time.

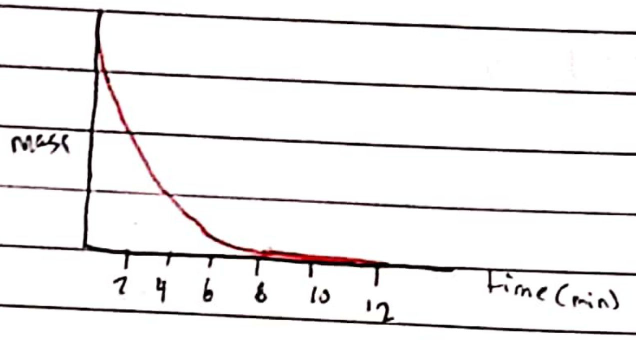
] steeper curve



* The graph below shows how the amount reactant changes with time



— Draw rate vs time graph for this reaction



* Factors affect the rate of reaction

① Temperature
② surface area
③ concentration / amount
④ pressure
⑤ light intensity
⑥ catalyst

① Temperature

★ State how the temp. affect the rate of reaction

- As temperature increases rate of reaction increases.

- Explain how the temperature affects the rate of reaction. [5]

As the temp. increases the particles gain kinetic energy so they move faster. The particles will have energy greater or equal to the activation energy, so more effective collisions per unit. resulting in faster rate of reaction.

- Plan an exp. to show how the temp. affects rate of reaction.

exp1: mass = 2.0g lumps

V HCl = 0.1 dm³

M HCl = 1 mol/dm³

Temp. = 25°C

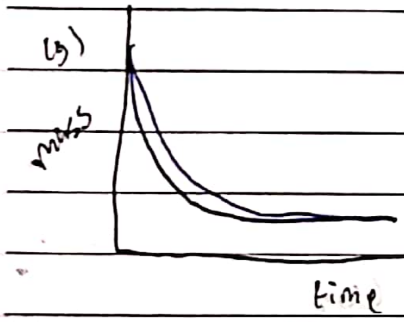
$CaCO_{3(s)} + 2HCl_{(l)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

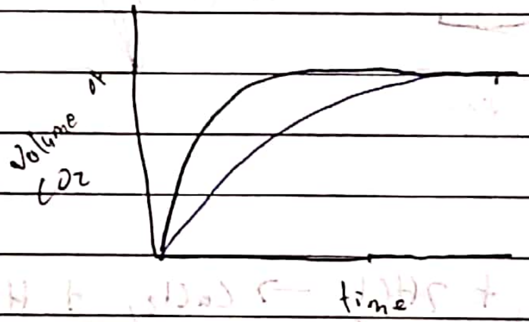
exp2: mass = 2.0g lumps

VHCl = 0.1dm³

MHCl = 1mol/dm³

Temp. = 50°C





• take a known mass of lumps $CaCO_3$ add them to a known volume of known conc. of HCl at 25°C. Measure the volume of CO_2 produced per unit time. Repeat the exp. at 50°C. The exp. at 50°C will produce CO_2 with less time.

using gas syringe

② Surface area

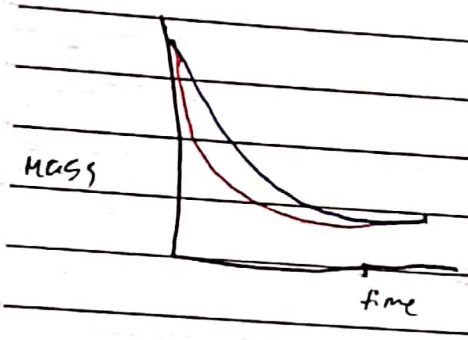
State how surface area affect the rate of reaction.
As surface area increases rate of reaction increases.

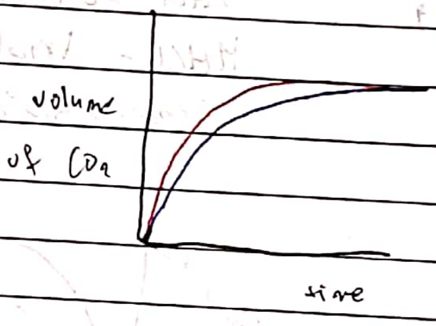
Explain how surface area affect the rate of reaction. [4]
As the surface area increases (decrease the particle size ~~using~~ by crushing using mortar + pestile) more particles exposed to the reaction. more effective collisions per unit time. So faster rate.

Plan an exp. to show how the surface area affect the rate of reaction

exp 1 mass = 2.0 g
lumps
V_{HCl} = 0.1 dm^3
M_{HCl} = 1 mol/dm^3
Temp. = 25°C

exp 2 mass = 2.0 g
Powder
V_{HCl} = 0.1 dm^3
M_{HCl} = 1 mol/dm^3
Temp. = 25°C





$CaCO_{3(s)} + 2HCl_{(l)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

6) Catalyst (they are solid)

- A chemical substance that speeds up the reaction without being used up.

HOW?

- it provides an alternative way with lower activation energy ~~Ea~~ so more particles will have energy equal to or more than the activation energy, so more effective collisions per unit time - So faster rate of reaction.

The reaction:

$$2H_2O_{2(aq)} \xrightarrow[② MnO_2]{① CuO} 2H_2O_{(l)} + O_{2(g)}$$

- Plan an experiment to show that CuO is a catalyst for this

Q1 reaction. [6]

→ take a know volume

→ with a know conc. of H_2O_2 → at know temperature.

→ Measure the volume of O_2 produced per unit time.

→ Repeat the experiment using CuO.

→ the exp. using CuO will produce more O_2 per the same unit time.

- Plan an exp. to show which of the two catalyst is better CuO or MnO_2.

→ Same as Q1 + same mass of catalyst. The exp. which produce more O_2 per same unit time used better catalyst.

- Plan an exp. to show that CuO isn't used up during the reaction. Measure the mass of CuO. Add H_2O_2 until no more fizzes. Filter the mixture, dry the solid in the oven. Remeassure the mass. Same initial + final mass.

COLLEGE

HCl if liquid → acid {
if gas → hydrogen chloride

Reversible reaction

Types of chemical rxn.

one way

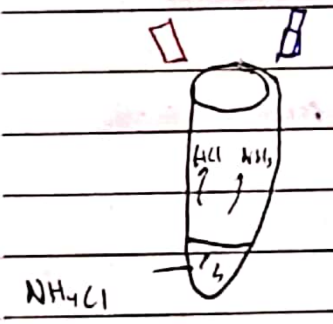
$$\text{Reactants} \xrightarrow{\text{forward}} \text{products}$$

both way

$$\text{Reactants} \underset{\text{backward}}{\overset{\text{forward}}{\rightleftharpoons}} \text{product}$$

$$* \quad NH_4Cl_{(s)} \xrightarrow{\Delta} NH_{3(s)} + HCl_{(g)}$$

ammonia chloride — ammonia — hydrogen chloride



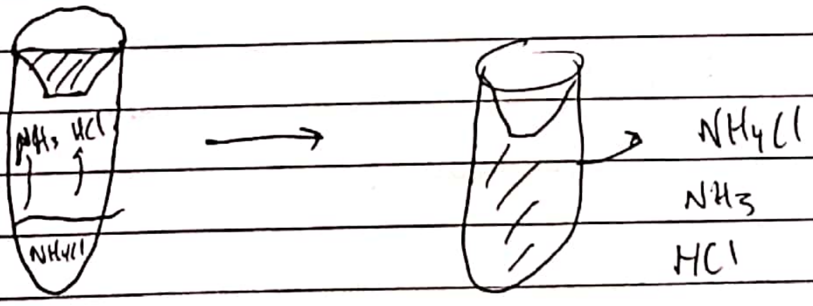
Q :- which damp litmus paper will change its colour first, why? [3]

- the damp red litmus paper changes to blue first bc $NH_{3(aq)}$ is an alkali and lighter than $HCl_{(aq)}$ which is acidic.

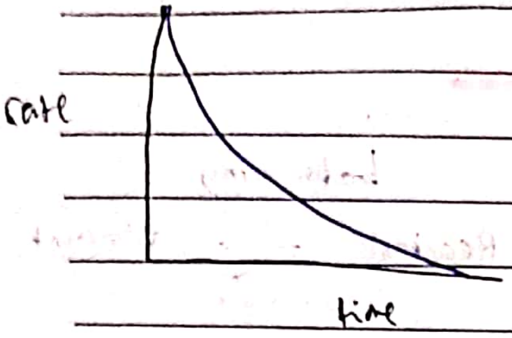
$$HCl_{(g)} + NH_{3(g)} \longrightarrow NH_4Cl_{(s)}$$

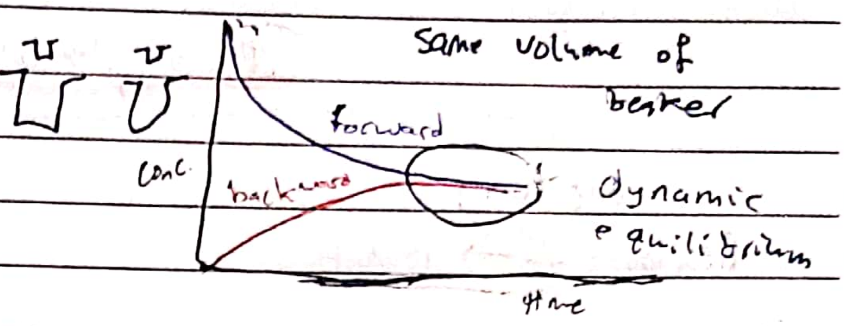
products are suitable

* so how do we keep both reactions happening



$R \rightarrow P$ $R \rightleftharpoons P$



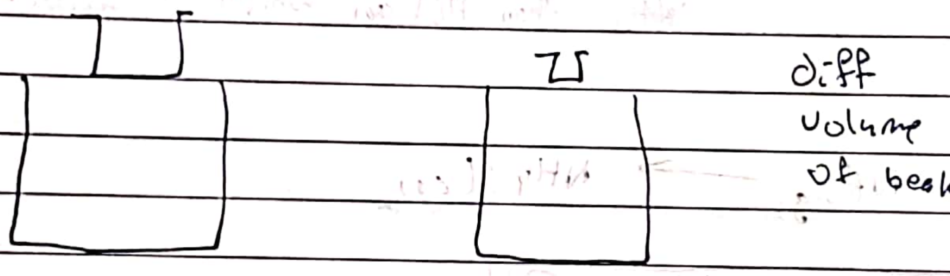


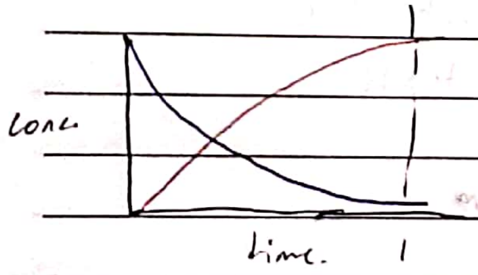
Conc. R ↓ until constant

Conc. P ↑ until constant

dynamic equilibrium:

- In terms of rate when the rate of forward reaction is equal to the rate of backward reaction.
- In terms of conc. when the conc. of reactant & product are constant.

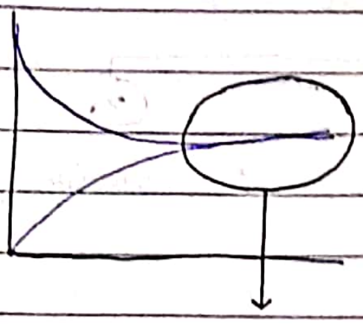




when the line is flat = dynamic equilibrium

- beakers represt moles

In terms of rate



when reactions are EQUAL

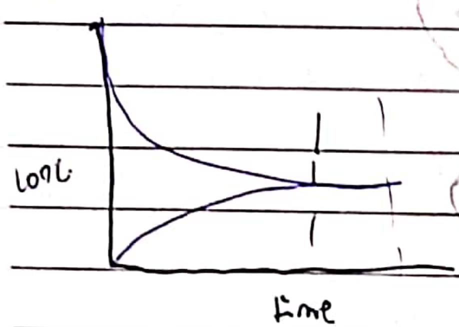
- rate of forward ↓
- less reactants so less particles so less effective collisions per unit time

- rate of backward ↑
  more products
  so more particles
  so more effective collisions

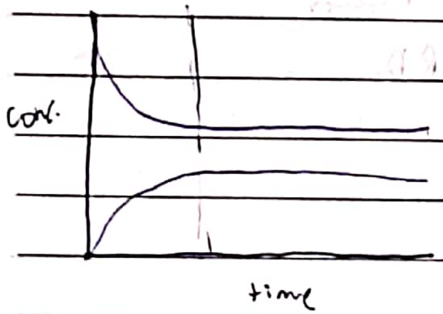
In terms of conc.

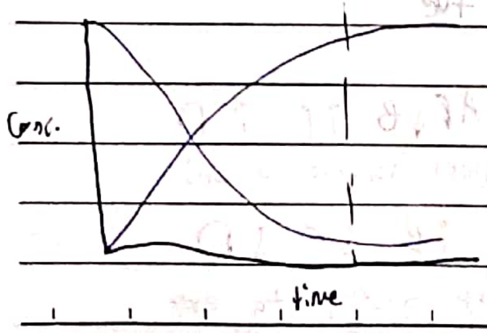
Le chatelier principle

- if system is at equil. ⇌ + any external factor disrubt the equil.



- the equil. can shift itself eithe to the forward ⟶ or to the backward ⟵ to return back to equil.





exo → releases heat

endo → absorbs heat

Factors that affect the position of equilibrium

① temp.	② Pressure	③ concentration
↑temp. shift to endo	↑Pressure shift to less gas mole	
↓temp. shift to exo	↓pressure shift to more gas mole	

↑temp. rate ↑

↓temp. rate ↓

↑Pressure ↑↑ rate of less gas mole

↑ rate of more gas mole

↓pressure ↓↓ rate of less gas mole

↓ rate of more gas mole

↑R, ↓P] shift forward

↓R, ↑P] shift backward

① Temperature

↑temp. ↑↑ rate of endothermic
↑ rate of exothermic
shift to endo

↓temp. ↓↓ rate of endo
↓ rate of exo
shift to exo

ΔH = enthalpy change

+ve — gain — endo

-ve — lose — exo

The sign of ΔH always represents FORWARD reaction

$$A + B \underset{exo}{\overset{endo}{\rightleftharpoons}} C + D \qquad \Delta H = +ve$$

↑temp. ↑↑ rate of forward
↑ rate of backward
↓A ↓B ↑C ↑D
shift forward to endo

↓temp. ↓↓ rate of forward
↓ rate of backward
↑A ↑B ↓C ↓D
shift backward to exo

$x + w \underset{endo}{\overset{exo}{\rightleftharpoons}} Y + Z$ $\Delta H = -ve$

↑ temp. ↑ rate of forward $\left.\begin{matrix} \\ \end{matrix}\right]$ shift to endo ↑ x ↑ w ↓ Y ↓ Z
↑↑ rate of backward

↓ temp. ↓ rate of forward $\left.\begin{matrix} \\ \end{matrix}\right]$ shift to exo ↓ x ↓ w ↑ Y ↑ Z
↓↓ rate of backward

Q. $3H_{2(g)} + N_{2(g)} \underset{endo}{\overset{exo}{\rightleftharpoons}} 2NH_{3(g)}$ $\Delta H = -ve$ exo

to produce more yield of NH_3 we must lower the temp. to favour the forward reaction which is the exo.

- Mixture of NO_2 + N_2O_4 at equil. in a sealed tube.

$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$

dark brown pale brown

+ If we put this ~~mixture~~ sealed tube in a cold waterbath the mixture becomes paler?, why?

~~Becomes~~ Because the forward reaction is exo enhanced by cooling.

② Pressure

- As the pressure increases the equil. will shift to the side with less pressure.

- As the pressure decreases the equil. will shift to the side with more pressure.

more ↓ moles.

GAS

more pressure ←

$$\overbrace{3H_{2(g)} + 1N_2}^{4 \text{ moles}} \rightleftharpoons \overbrace{2NH_{3(g)}}^{2 \text{ mole}}$$

(4 → 2)

↑ pressure equil. will shift forward to the side with less gas moles. ↑ % NH_3

$$\overbrace{CaCO_{3(s)}}^{0} \rightleftharpoons CaO_{(s)} + \overbrace{CO_{2(g)}}^{1}$$

(0 → 1) ↓ pressure equil. will shift forward to the side with more gas moles.

$$\overbrace{H_2CO_{3(aq)}}^{0} \rightleftharpoons \overbrace{H_2O_{(l)} + CO_{2(g)}}^{1}$$

↓ pressure equil. will shift forward to the side with more gas moles.

$$\overbrace{H_{2(g)} + I_{2(g)}}^{2} \rightleftharpoons \overbrace{2HI_{(g)}}^{2}$$

Changing the pressure has no effect on the position of equilibrium, since both side has the same num. of gas moles.

$$\overbrace{H_2 + I_2}^{2} \rightleftharpoons \overbrace{2HI_{(g)}}^{2}$$

purple   colourless

* the equil. isn't affected by increasing pressure : why
by increasing the pressure the mixture becomes more purple.
The I_2 molecules become closer to each other
so the colour seems to be darker.

* Sealed container contains mixture of $NO_{2(g)}$ & $N_2O_{4(g)}$ at equilibrium.

$$\overbrace{2NO_{2(g)}}^{2} \rightleftharpoons \overbrace{N_2O_{4(g)}}^{1}$$

$\boxed{\text{dark brown}}$ $\boxed{\text{pale brown}}$

AS

by ↑ pressure the colour of mixture :

* a) becomes paler , then goes darker
(b) " darker — paler
* c) " paler & stay paler
d) — darker — darke

*Q : $2SO_{2(g)} + O_{2(g)} \underset{\text{endo}}{\overset{\text{exo}}{\rightleftharpoons}} 2SO_{3(g)}$ $\Delta H \overset{\text{exo}}{=} -ve$

	rate of forward	rate of backward	% SO_3
↑ Temp.	increase	increase	decrease
↓ Temp.	decrease	decrease	decrease

③ Concentration

- $A + B \rightleftharpoons$

- ↑ [A] shift forward ↓ B ↑ C ↑ D
- ↓ [B] shift backward ↑ A ↓ C ↓ D
- ↑ [C] shift backward ↑ A ↑ B ↓ D

Indicator:

$HIn_{(aq)} \rightleftharpoons H^+_{(aq)} + In^-_{(aq)}$

colour 1 colour 2

red yellow

add HCl: proton donor ↑ H^+ shift backward

More HIn colour 1

less In colour 2

add NaOH: proton acceptor ↓ H^+ shift forward

More ~~to~~ In: colour 2

Less HIn colour 1

+ The reversible reaction below is at equilibrium

$Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightleftharpoons Cr^{3+} + Fe^{3+} + H_2O$

orange green

- explain why by adding HCl to the reaction mixture, the colour of the mixture becomes green?

HCl is an acid (proton donor)

more H^+

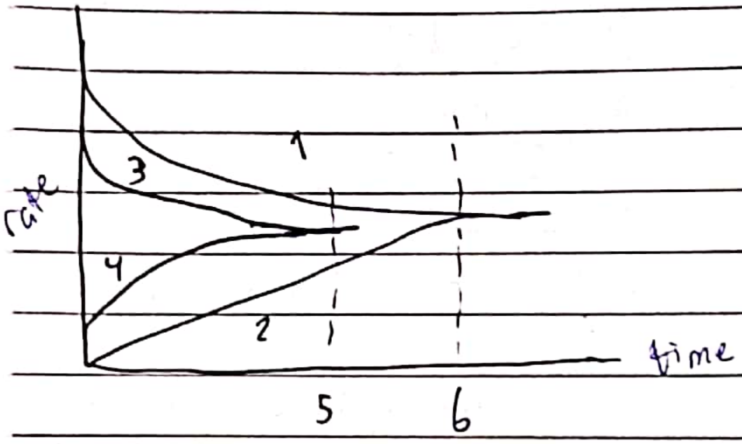
shift forward

more Cr^{3+} → more green

less $Cr_2O_7^{2-}$ → less orange

* Catalyst

has no effect on the position of the equilibrium since it speeds up the rate of forward and backward reaction.



1) rate of forward without catalyst.
2) rate of backward without catalyst.
3) rate of forward with catalyst.
4) rate of backward with catalyst.
5) time taken to reach equilibrium with catalyst.
6) " - ——— - - - without catalyst.

$\frac{Q}{\# \text{ of moles}}$

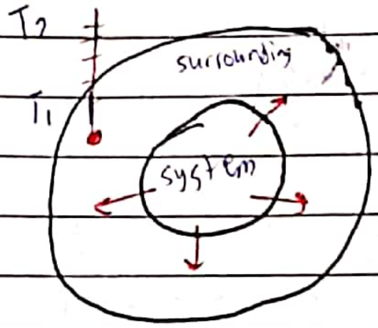
exo

Energetics

- **exothermic** (LOSE)

Reaction that gives out (release) energy to the surrounding.

$T_2 > T_1$

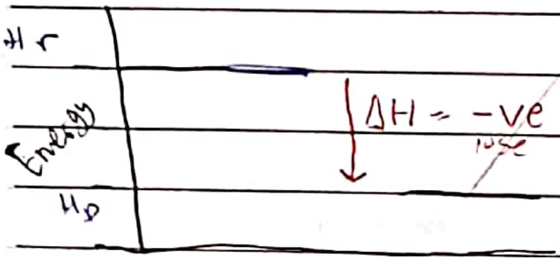


- As energy decreases it becomes more stable.
- Product is more stable (less energy)

- surrounding: $Q = mc\Delta T$

- system: $\Delta H = -ve$

For system (energy level diagram)



rxn progress

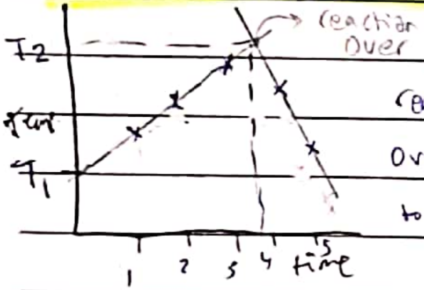
- Enthalpy: Heat contents "stored energy"

Hr enthalpy of reactants

Hp of products

↑Q more exo ΔT ↑

For surrounding



reaction is over so return back to room temp.

$Q = mc\Delta T$

Q: Energy transfer (J)

m: Mass (g)

c: Specific heat capacity (4.2 J/g°C)

ΔT: Change in Temp. (°C)

Examples on exo:-

1- freezing, condensation
2. Respiration.
3. Combustion
4. Neutralization
5. Displacement.
6. Voltaic cell.
7. Building up bonds

COLLEGE

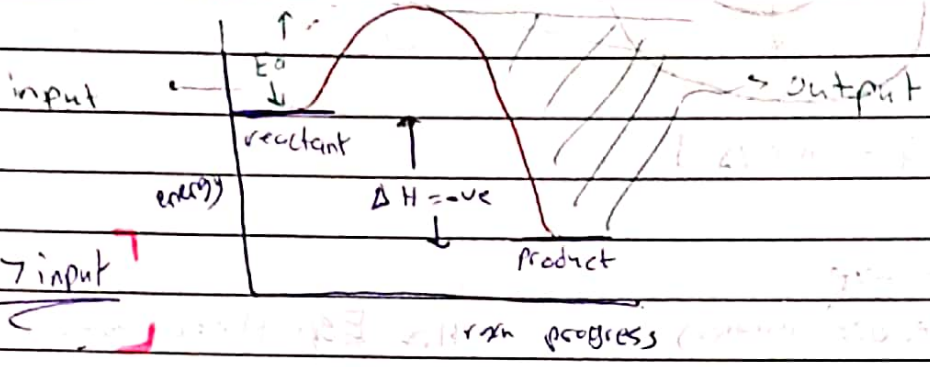
exo

+ How to express exo rxn

① Reactants → Products $\Delta H = -ve$

② Reactants → Product + Energy

③ Profile diagram

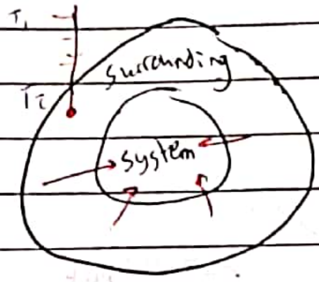


output > input

endothermic

Reactions that absorb (take in) energy from the surrounding.

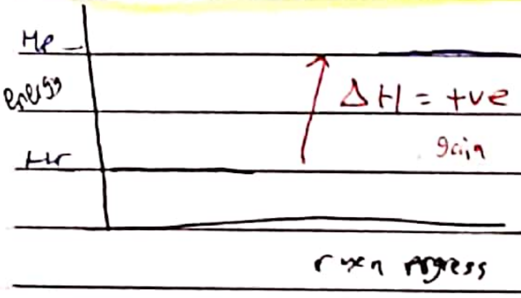
$T_1 > T_2$



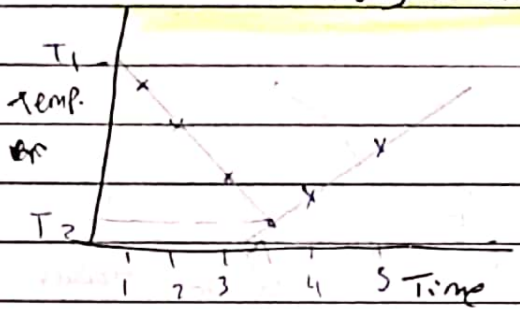
- system : $\Delta H = +ve$
- surrounding : $Q = mc\Delta T$

$\updownarrow Q$ more endo $\Delta T \uparrow$

For System (energy level diagram)



For surrounding (Temp. diagram)



$Q = mc\Delta T$

Examples on endo :

1. Boiling, melting.
2. Photosynthesis.
3. Thermal decomposition.
4. electrolysis.
5. photo-graphic films.
6. dissolve ammonium salts.
7. Breaking down bonds.

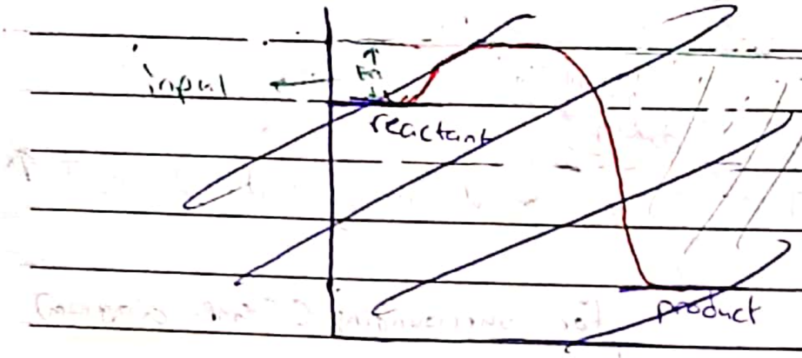
endo

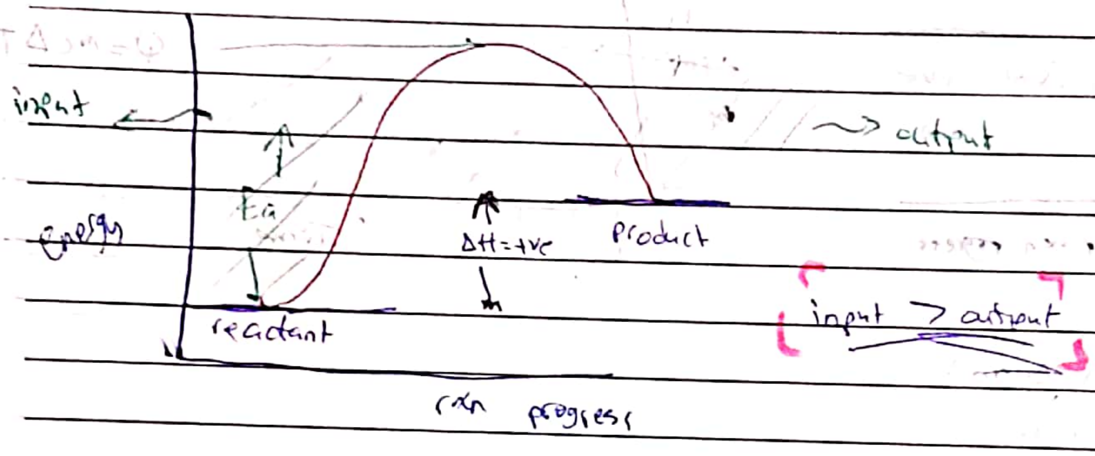
* How to express endo rxn

① Reactants → Products $\Delta H = +ve$

② Reactants + Energy → Products

③ Profile diagram





Measuring ΔH reaction

Theoritical — using bond energy

experimental — Combustion, Displacement, neutralization

ΔH reaction using bond energy :

Bond energy: the amount of energy needed to break 1 mol of a bond in gaseous state.
— — released to build —

bond	bond energy kJ/mol
H - H	436

$H-H + 436 kJ/mol \rightarrow H_{(g)} + H_{(g)}$

$H_{(g)} + H_{(g)} \rightarrow H-H_{(g)} + 436 kJ/mol$

gaseous state?

$H_2O_{(s)} \rightarrow H_2O_{(l)} \rightarrow H_2O_{(g)}$

$\downarrow$ (E) bond energy

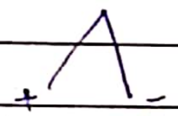
$2H_{(g)}$

$O_{(g)}$

total

ΔH reaction $= \Sigma$input $- \Sigma$output

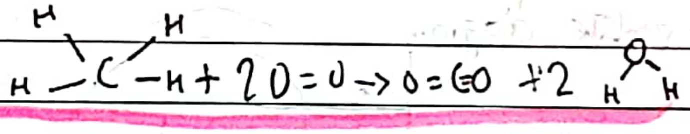
to break down bonds in reactant — to build up bonds in products



endo — exo

input > output   input < output

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$



to use this equation
1. ① balanced equation. ✓
2. ② covalent structure. ✓
3. ③ bond energy.

bond broken

$4 * C-H \rightarrow 4 * 413$
$2 * O=O \rightarrow 2 * 495$ +
2642 kj

bond formed

$2 * C=O \rightarrow 2 * 799$
$4 * O-H \rightarrow 4 * 463$ +
3450 kj

bond	bond energy
C-H	413
O=O	495
C=O	799
O-H	463

C-O	358

$\Delta H = \Sigma in - \Sigma out$
$= 2642 - 3450$
$= -808 kj / mol$

exo

input < output

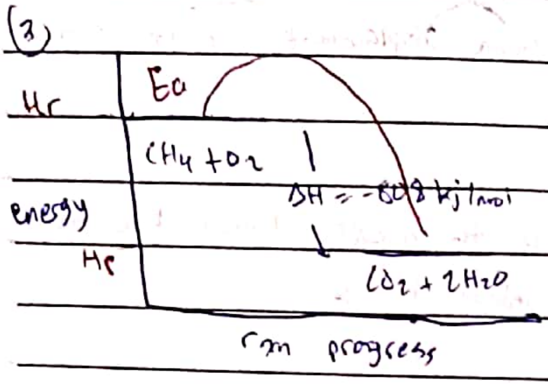
COLLEGE

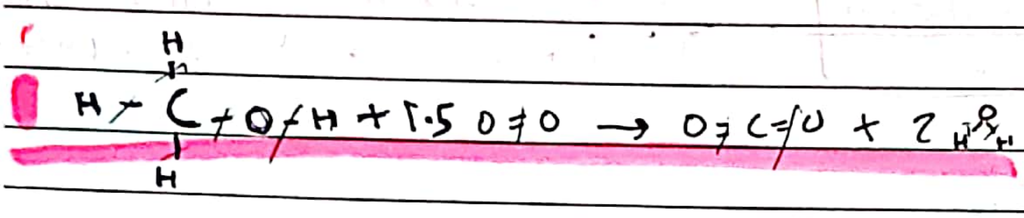
★ bc it's exo it can be written as:

(1) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + 808$ kj/mol

(2) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ $\Delta H = 808$ kj/mol

(3)





$H-\overset{H}{\underset{H}{C}}-O-H + 1.5\ O=O \longrightarrow O=C=O + 2\ H-O-H$

bond broken	
3×C−H	3×413
1×C−O	1×358
1×O−H	1×463
1.5×O=O	1.5×495
	2802.5 kj

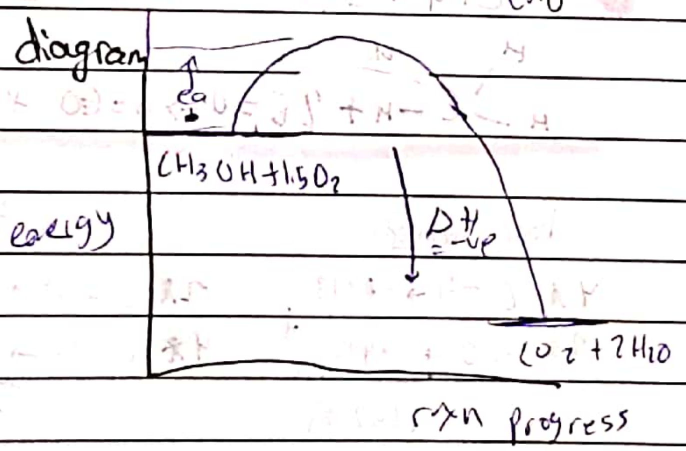
+

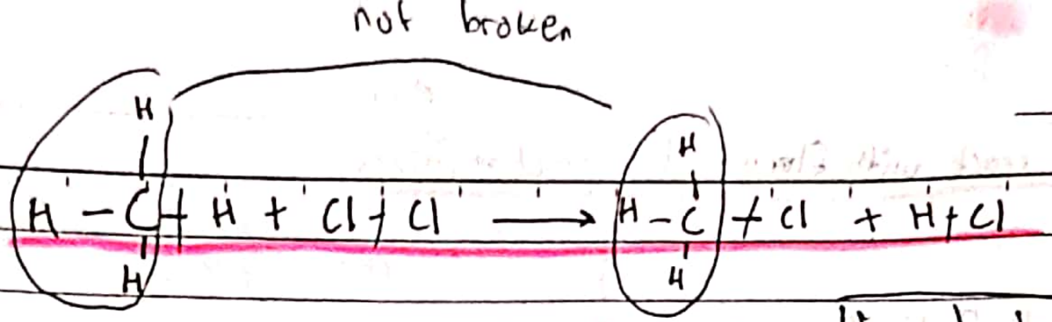
bond formed	
2×C=O	2×799
4×O−H	4×463
	3450 kj

$\Delta H = 2802.5 - 3450 = (-)647.5$ kj/mol

exo

+ profile diagram





bond broken	
1 × C−H	413
1 × Cl−Cl	242 +
	655

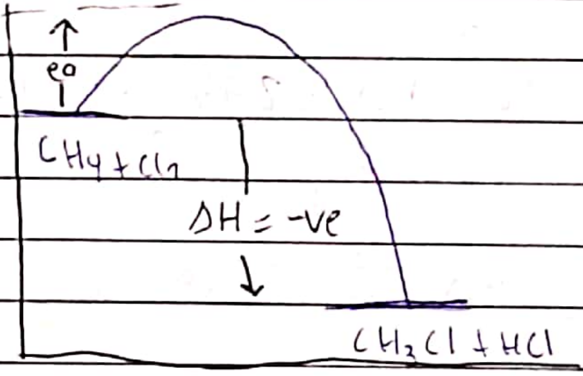
bond formed	
1 × C−Cl	328
1 × H−Cl	431 +
	759

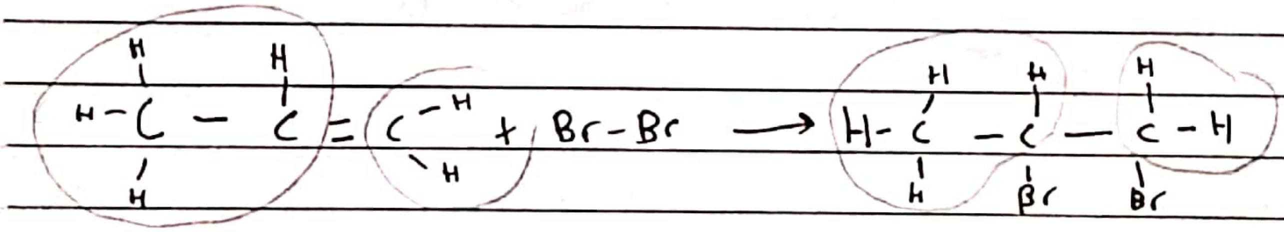
bond	bond energy
C−H	413
Cl−Cl	242
H−Cl	431
C−Cl	328

$\Delta H = 655 - 759 = \boxed{-}104$

exo

* profile diagram





bond broken	
Br−Br	193
C=C	614 +
	807 kj

bond formed	
2 × C−Br	2 × 276
C−C	348 +
	900 kj

bond	bond energy
C−H	413
C−C	348
C=C	614
Br−Br	193
C−Br	276

$\Delta H = 807 - 900 = \boxed{-}93$ kj/mol

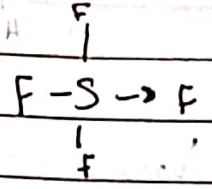
exo

e×o

- When sulfur react with flarine the reaction gives 780 kJ/mol

ΔH = -780

320 $S + 2\,F{-}F \rightarrow$



$$F{-}\overset{F}{\underset{F}{S}}\rightarrow F$$

If bond energy of F-F is 160 kj/mol find bond energy of S-F.

$F{-}F = 160 \times 2 = 320$ $\quad$ $S{-}F \times 4$

$-780 = 320 - 4x$

$-1100 = -4x$ $\quad$ $\boxed{x = 275}$ ← bond energy

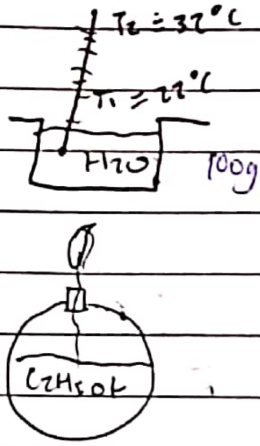
[Q ↑ ΔT ↑]

ΔH (energy change) practically:

combustion   displacement   Neutralization

I - Combustion

$C_2H_5OH_{(l)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(l)}$



$32 - 22 = \underline{\underline{10}}$

$Q = mc\Delta T$ constant

, energy transfer $= 100 \times \underline{4.2} \times 10$

$= 4200\,J \approx \underline{\underline{4.2\,kJ}}$

4.2 → 2g

x → 1 mole = 46 g (Mr → C_2H_5OH)

$x = \boxed{96.6\ kj/mol}$ ΔH

$m_1 = 200g$

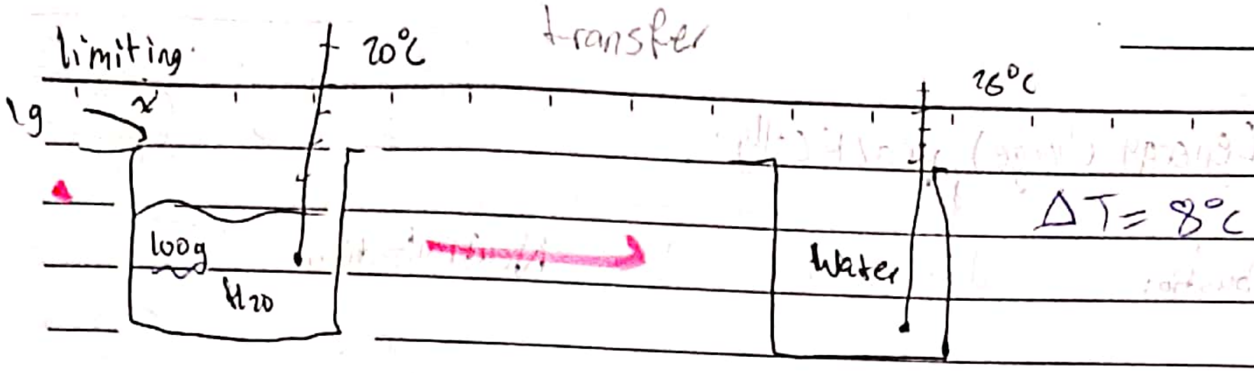
$m_2 = 198g$] $= 2g$, so we need 4.2kg for $\underline{\underline{2g}}$, but what if we want 1 mole?

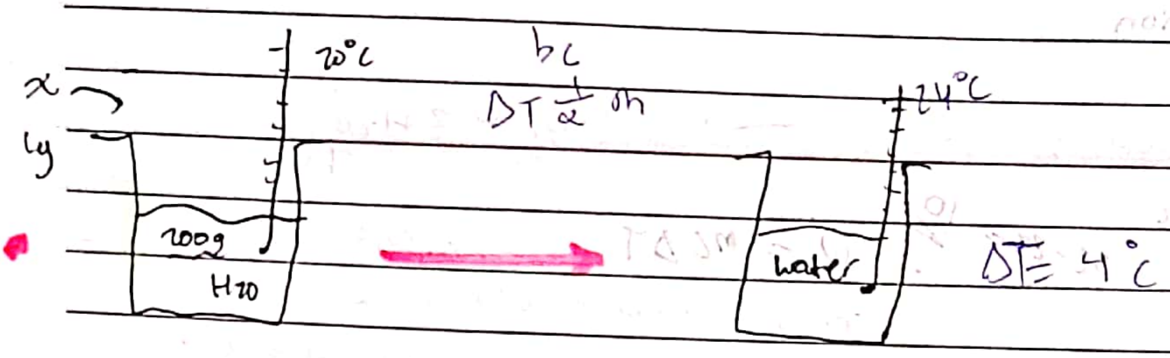
+ Two fuels A & B. Plan an exp. to show which one produces more energy.

→ take a known mass of water with known initial temp. in a copper can.
→ take a known mass of fuel A.
→ ignite the fuel and record the final mass + final temp. of water.
→ Repeat the exp. using fuel B.
→ the fuel which causes more temp. rise per gram of fuel produces more energy.

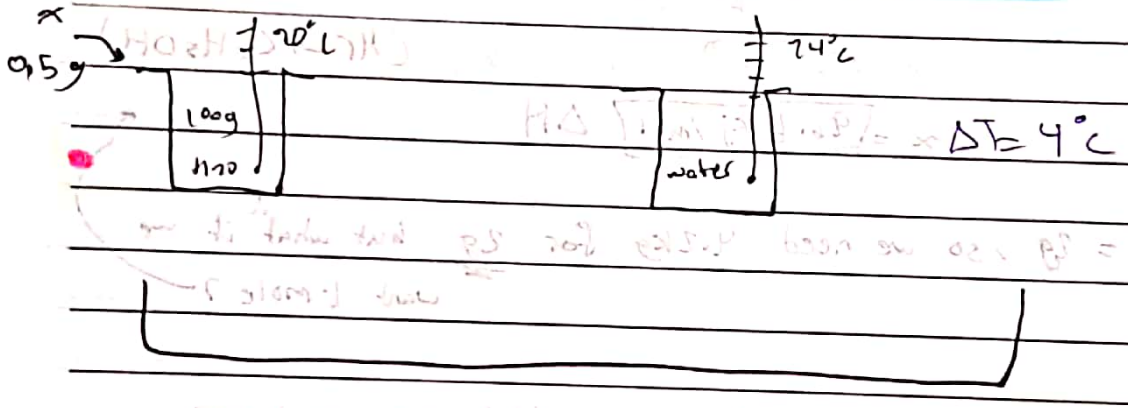
$$\Delta T = \frac{Q}{mc}$$

Same energy transfer





- diff. distribution but same energy transfer.

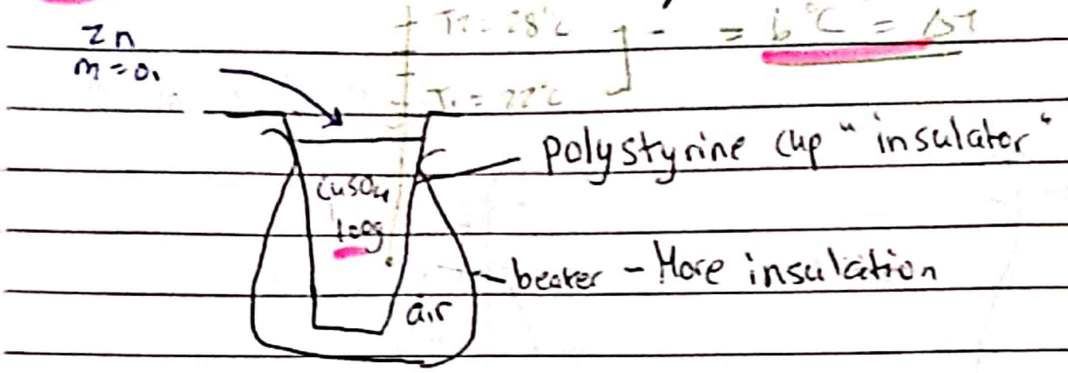


displacement
Introduction

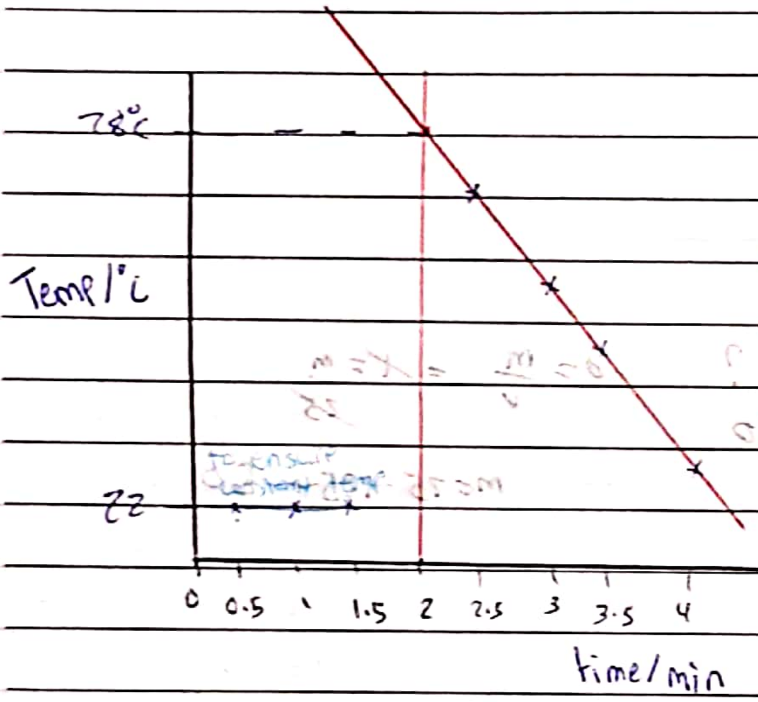
C ALWAYS
4.2

II - displacement

① $Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$



$Q = mc\Delta T = 100 \times 4.2 \times 6 = 2520\ J = \boxed{2.52\ kJ}$

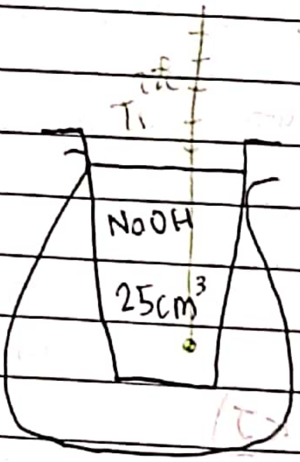


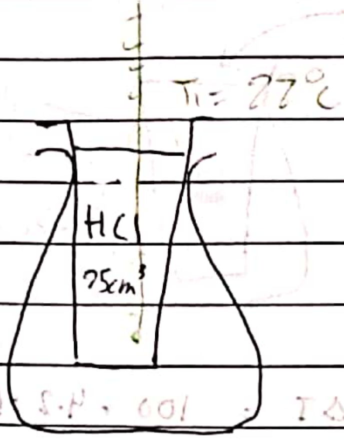
[Stir CuSO4 with thermometer to ~~the~~ distribute heat equally (slowly to avoid overheating)]

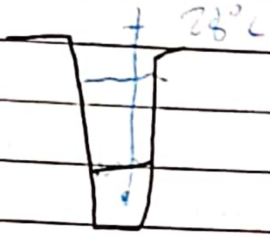
III - Neutralization

$d = 1 g/cm^3$

$$NaOH_{(aq)} + HCl \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$







$\Delta T = 28 - 22 = 6°C$

$c = 4.2$

$m = ?$

$m = 50$

$d = \frac{m}{v} = \cancel{1} = \frac{m}{25}$

$m = 25 + 25 = 50$

$Q = mc\Delta T = 50 \cdot 4.2 \cdot 6$

$= 1260\ J$

extra note:

physical test : Boiling point 100°c

chemical test:

① $CuSO_{4(s)} + 5H_2O \rightleftharpoons CuSO_4 \cdot 5H_2O$

white — blue

(2) $CoCl_2 + 6H_2O \rightleftharpoons CoCl_2 \cdot 6H_2O$

Alternative Resources of Energy

~~Voltaic cell~~ | Hydrogen fuel cell | ~~uranium~~

Hydrogen fuel cell

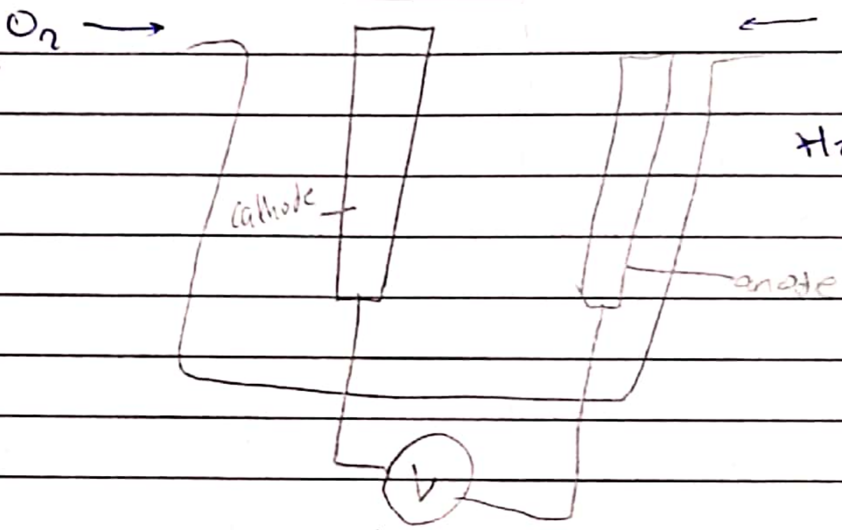
$\overset{0}{2H_2} + \overset{0}{O_2} \rightarrow \overset{+}{H_2}\overset{2-}{O}$

$4e^- + O_2 \rightarrow 2O^{2-}$

reduction

$H_2 \rightarrow 2H + 2e^-$

oxidation



Advantages

- only one waste product (H_2O
- no CO_2
- produce high amount of energy
- generate electricity

disadvantages

- expensive
- hard to store & transport
- risk of explosion

Industrial chemistry

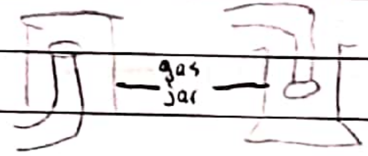
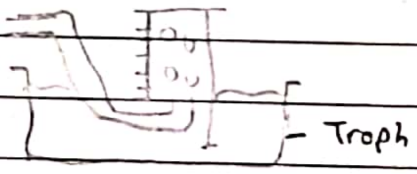
- Dealing with gases
  - dry
  - Collect
- Industry of
  - NH_3 Haber process
  - H_2SO_4 contact process
  - CO_3^{2-} carbonate cycle
- Extraction of metals
  - Al
  - fe
  - Zr
  - Cu

* Dealing with gases:

reaction → wet gas → dry → collect

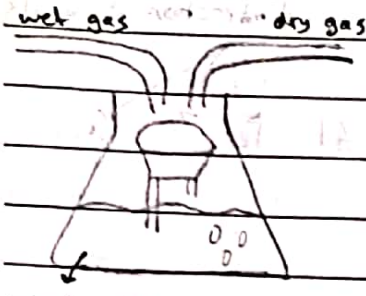
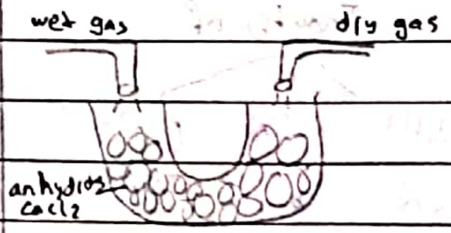
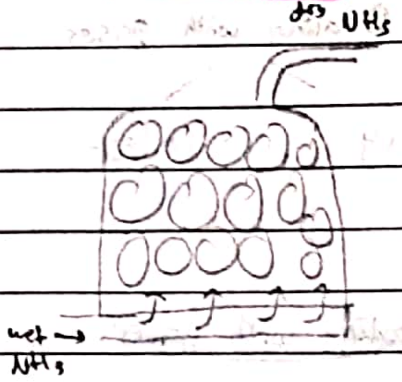
* collect gas:

- O_2: $H_2O_{2(aq)} \xrightarrow{CuO} H_2O_{(l)} + O_{2(g)}$
- H_2: $Mg_{(s)} + HCl_{(aq)} \longrightarrow MgCl_{(aq)} + H_{2(g)}$
- NH_3: $NH_4Cl_{(aq)} + NaOH_{(aq)} \xrightarrow{\Delta} NaCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$
- SO_2: $Na_2SO_3 + HCl \xrightarrow{\Delta} NaCl_{(aq)} + H_2O_{(l)} + SO_{2(g)}$
- CO_2: $CaCO_{3(s)} + HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

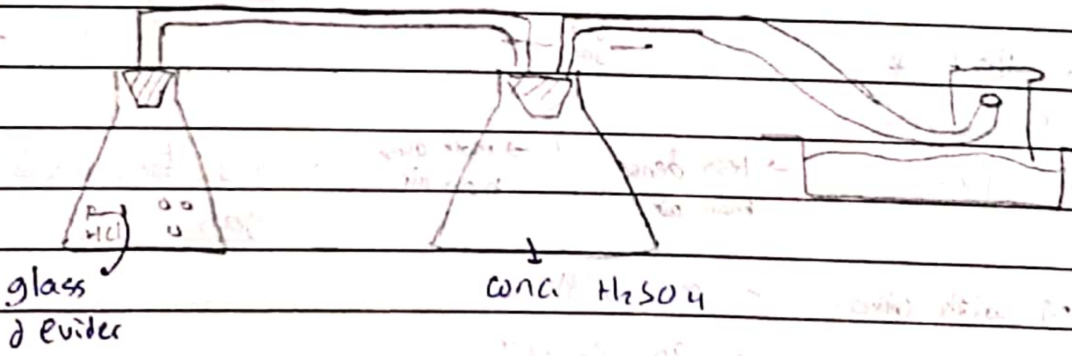
Gas syringe	Delivery tube	Over water
	upward delivery / downward delivery	
		
-used to collect & measure the volume of ~~gas~~ any gas.	→ less dense than air → more dense than air.	- only for insoluble gas
- no mixing with other.	- mix with air - an escape	

* Drying gas:

- concentrated H_2SO_4.
- Anhydrous $CaCl_2$.
- Calcium oxide.

Concentrated H_2SO_4	Anhydrous $CaCl_2$	Calcium oxide
 conc. H_2SO_4 (becomes dilute) - used to dry any gas except NH_3 it neutralizes the H_2SO_4. $NH_3 + H_2SO_4 \rightarrow (OH)_2SO_4$	 - used to dry any gas except ammonia.	

* Draw a suitable apparatus used to collect & measure the volume of CO_2 gas from $CaCO_{3(s)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$ (CO_2 insoluble).



78% N_2 0.03% CO_2
21% O_2 0.01% H_2
0.9% Ar 0.06% others

II - Industry of ammonia "Harber process" → Prepare fertilizers

$$3H_{2(g)} + N_{2(g)} \overset{exo}{\rightleftharpoons} 2NH_{3(g)} \qquad \overset{exo}{\Delta H = -ve}$$

How to obtain:

① **Nitrogen**: fractional distilation of liquid air
$\overset{\downarrow}{\text{diff. B.p.}}$ $\overset{\downarrow}{\text{cooling}}$ under ↑ pressure

② **Hydrogen**: 1- cracking of Alkanes (organic).
2- $CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$

+ essential conditions for ~~harbone~~ haber process

1) temp. 400°C - 450°C adv.

- lower temp. (less than 400°C) → - higher yield of NH_3 → shifts forward exo side
  dis. - slower rate, particles lose KE so less effective collision per unit time
- higher temp. (more than 450°C) → d less yield → shifts backward endo side
  a - faster rate

2) Pressure 200 atm.

adv. - more yield of NH_3 → shift forward to the side with fewer gas mole.
- faster rate.

~~dis.~~

dis. - risk of explosion
- expensive

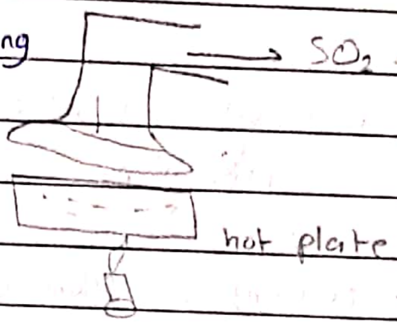
- Uses of ammonia

1- fertelizers.
2- Cleaning detergents
3- Smelling

+ Contact process (industry of H_2SO_4)

stages + $S \xrightarrow{①} SO_2 \underset{\text{contact process}}{\overset{②}{\rightleftharpoons}} SO_3 \xrightarrow{③} H_2S_2O_7 \xrightarrow{④} H_2SO_4$

①: add O_2, process: roasting



S)
- group (VI)
- valency (2)
- yellow solid
- S_8

+ uses: - medicine
- rubber

ore: + zinc blend
+ from fossil fuel

②: $2SO_{2(g)} + O_{2(g)} \underset{}{\overset{V_2O_5}{\rightleftharpoons}} 2SO_{3(g)}$ $\Delta H = -ve$ exo

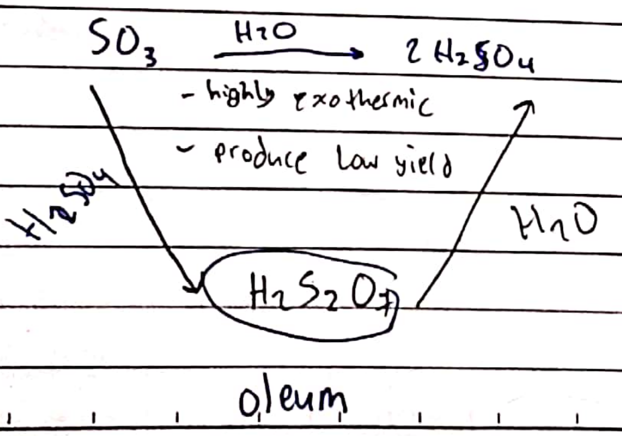
- essential conditions.

Contact process

1) Temp. 400 - 450°C
2) Pressure 2 atm, high pressure favours the forward rxn (fewer gas moles)
2 atm gives max yield of SO_3
3) Catalyst V_2O_5 Vanadium oxide

SO_2 causes acid rain

③ + ④:



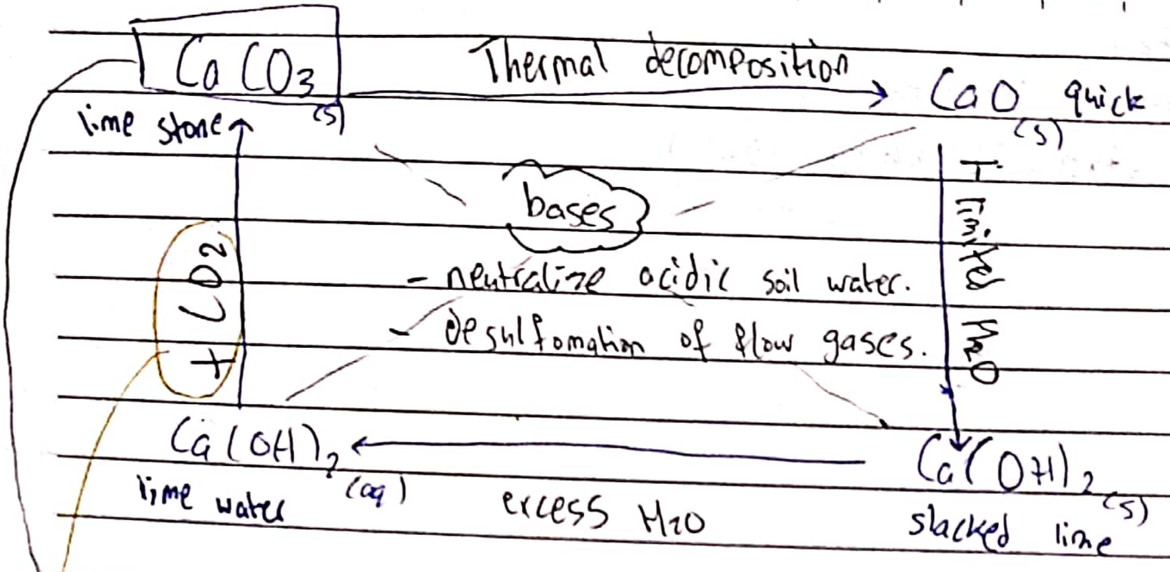
+ we can obtain H_2SO_4 from electrolysis of aqueos copper chloride
↓
electrolyte is H_2SO_4

COLLEGE

*uses of SO_2,

- kills bacteria (sterilization)
- paper industry. (bleaching agent)

Carbonate cycle.



→ uses: - buildings
- extraction of Iron.

→ equation: $Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$

Extraction of Metals

K

Na

Li

Ca

Mg

Bauxite Al_2O_3 ← Al] → electrolysis / molten

C, CO

Zinc blend ZnS ← Zn

Hematite Fe_2O_3 ← Fe] reduction by C + CO

Pb] "blast furnace"

H

CuS ← Cu] reduction by H_2

Ag

Au

Pt

Extraction of Iron

ore: Fe_2O_3 "Hematite".

method: reduction by C + CO.

place: blast furnance.

acidic impurities

raw materials: - Fe_2O_3 mixed with SiO_2 ↙

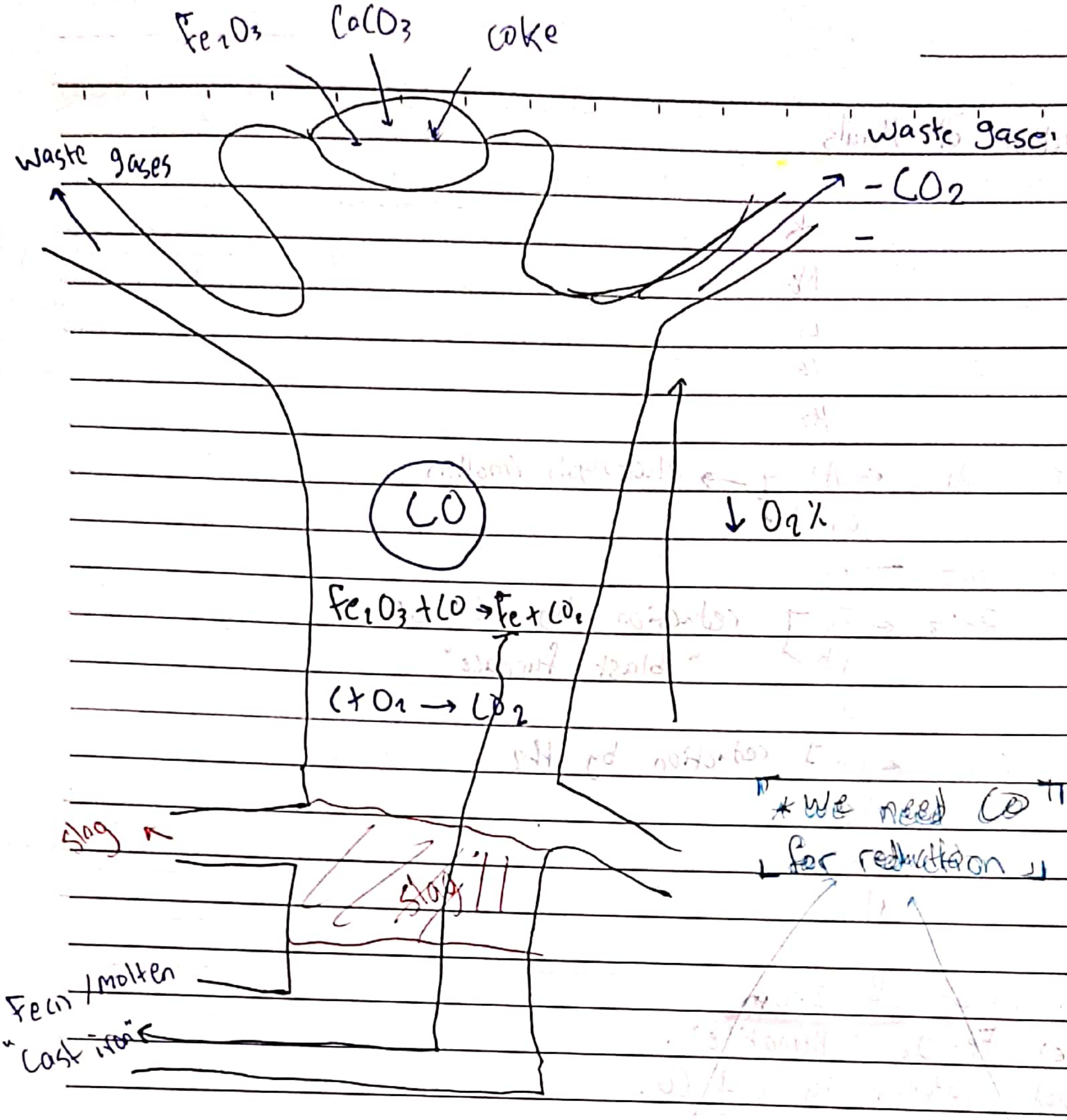
needed for extraction - $CaCO_3$ "lime stone"

- coke "carbon pure" ←

- air 1500°C

- why didn't $CaCO_3$ neutralize SiO_2?

- $CaCO_3$ becomes decomposed from heat. before it gets the chance to neutralize



"*We need CO for reduction"

1st reaction: complete combustion $C + O_2 \rightarrow CO_2$
(produce energy)

2nd reaction: incomplete combustion $2C + O_2 \rightarrow \boxed{2CO}$

redox: - $CO_2 + C \rightarrow \boxed{2CO}$

- $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

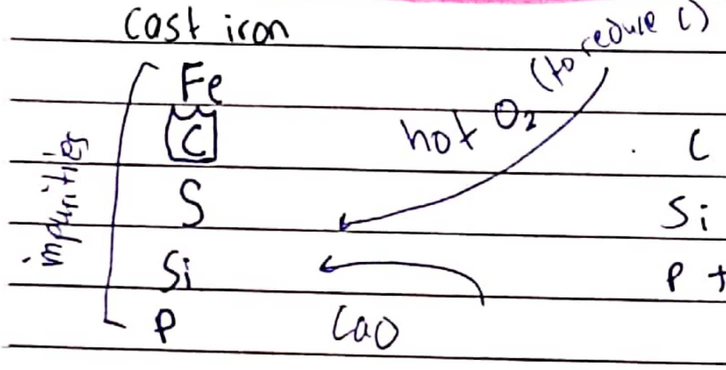
Thermal reaction: $CaCO_3 \xrightarrow{heat} CaO + CO_2$

neutralization: $CaO + \underset{(acid)}{SiO_2} \rightarrow CaSiO_3$ (Calcium silicate slag) importance:

Slag
- forms a protective layer that prevent the reaction of $Fe + O_2$
- mixed with bitumene to make roads. use.

- Nitrogen and nitrogen dioxide are obtained from air (78% N).

- Steel making "oxygen base process"



$C + O_2 \rightarrow CO_{2(g)}$

$Si + O_2 \rightarrow SiO_{2(s)}$

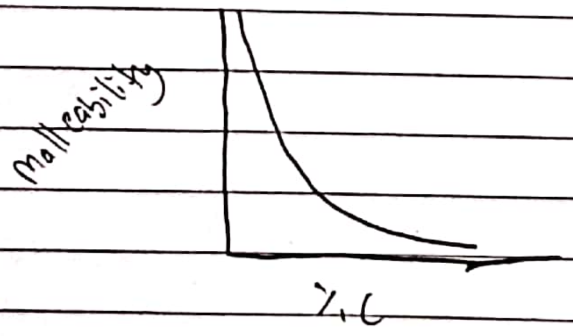
$P + O_2 \rightarrow P_2O_{5(s)}$] acidic

CaO neutralizes them to remove them.

$CaO + SiO_2 / P_2O_5 \rightarrow$ STEEL

Steel			
	mild steel	0.03% C	"car bodies"
	medium steel	0.3% C	"Railways"
	stainless steel	3-5% C	"Cutter, watch"

- Brass is copper + zinc.



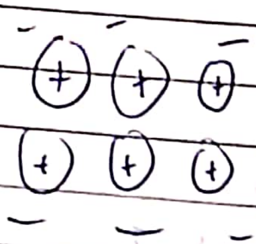
* **Alloy**: mixture of metal with another metal or semi-metals.

Brass: Cu, Zn

bronze: Cu, Sn

Steel: Fe, C, Ni, Cr

metal Cu



* Extraction of zinc

ore: Zinc blend ZnS
method: reduction by C & CO
place: Blast furnance.

C, CO + H_2 can ONLY reduce the less reactive and only from its oxide.

Step 1:- Roasting with hot oxygen.

$$ZnS + O_2 \rightarrow ZnO + SO_2$$

step 2:-

$$C + O_2 \rightarrow CO_2$$
$$C + O_2 \rightarrow 2CO$$
$$CO_2 + C \rightarrow 2CO$$
$$CO + ZnO \rightarrow Zn + CO_2$$

the temp. inside the furnance 1500°C and b.p. of zinc is 907°C. So it produced as pure gas must condense, and other impurities since