Redox Reactions

* in terms of oxygen :-

Redox

Reduction	Oxidation
lose oxygen	gain oxygen

Reduction : Fe in Fe_2O_3

* $Al + Fe_2O_3 \longrightarrow Al_2O_3 + Fe$

oxidation : Al

reduction : Cu in CuO

* $H_2 + CuO \longrightarrow Cu + H_2O$

oxidation : H_2

reduction : Zn in ZnO

* $CO + ZnO \longrightarrow Zn + CO_2$

oxidation : O_2

* in terms of hydrogen :-

Redox

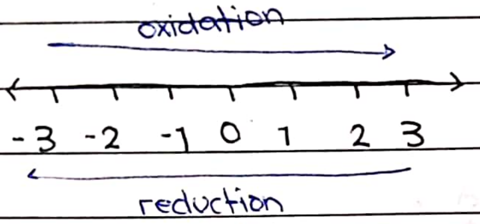
Reduction	Oxidation
gain hydrogen	lose hydrogen

reduction : Cl_2

$H_2S + Cl_2 \longrightarrow HCl + S$

oxidation : S in H_2S

* in terms of oxidation state :-



(for charges)

* Rules of oxidation state :

1) the oxidation state for any free element is zero.

monoatomic — Na, K, Ca
diatomic — H_2, O_2
polyatomic — P_4, S_8

2) the oxidation number of any atom in a compound form.

- group 1 → +1 Li, Na, K, Rb
- group 2 → +2 Mg, Ca, Sr, Ba
- group 3 → +3 only Al
- group 7 → -1 only F

3) the oxidation number of hydrogen (+1)
except with metals in metal hydroxides (-1)

non-metals :- $C\overset{+1}{H_4}$, $N\overset{+1}{H_3}$

metals :- $\overset{+1}{Na}\overset{-1}{H}$, $\overset{+1}{Ca}\overset{-1}{H_2}$

4) the oxidation state for oxygen is (-2)
except in peroxide (-1) except in OF_2 (+2)
↳ O_2^{-2}

* Sodium oxide : $\overset{+1}{Na} + O^{-2} \rightarrow Na_2O^{-2}$

* Sodium peroxide : $\overset{+1}{Na} + O_2^{-2} \rightarrow Na_2O_2^{-1}$

* Water : H_2O^{-2}

* hydrogen peroxide : $H_2O_2^{-1}$

* Calcium oxide : $\overset{+\cancel{2}}{Ca} + O^{-\cancel{2}} \rightarrow CaO$

* Calcium peroxide : $\overset{+\cancel{2}}{Ca} + O_2^{-\cancel{2}} \rightarrow CaO_2$

NOTE: in ions you don't multiply with the charge of the ion but with the oxidation state of the element

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

- **Compounds :-**

* $\overset{+1}{Na}\underset{?}{\underline{Cl}}$ $+1 + Cl = 0$ $\boxed{Cl = -1}$

* $\overset{+1}{H}\underset{?}{\underline{Cl}}\overset{-2}{O}$ $+1 - 2 + Cl = 0$ $-1 + Cl = 0$ $\boxed{Cl = +1}$

* $\overset{+1}{H}\underset{?}{\underline{Cl}}O_2^{-2}$ $Cl + 1 + 2(-2) = 0$ $Cl + 1 - 4 = 0$ $\boxed{Cl = +3}$

* $\underset{?}{\underline{N}}H_3^{+1}$ $N + 3(+1) = 0$ $N + 3 = 0$ $\boxed{N = -3}$

* $\underset{?}{\underline{N_2}}O_5^{-2}$ $2N + 5(-2) = 0$ $2N - 10 = 0$ $2N = 10$ $\boxed{N = +5}$

* $\overset{+1}{H_2}\underset{?}{\underline{S}}$ $S + 2(+1) = 0$ $S + 2 = 0$ $\boxed{S = -2}$

* $\overset{+1}{H_2}\underset{?}{\underline{S}}O_4^{-2}$ $S + 2(+1) + 4(-2) = 0$ $S + 2 - 8 = 0$ $S - 6 = 0$ $\boxed{S = +6}$

- **ions**

* $\underset{?}{\underline{N}}O_3^{-1}$ $N + 3(-2) = -1$ $N - 6 = -1$ $\boxed{N = +5}$

* $\underset{?}{\underline{Mn}}O_4^{-1}$ $Mn - 8 = -1$ $\boxed{Mn = +7}$

* $\underset{?}{\underline{S}}O_3^{-2}$ $S - 6 = -2$ $\boxed{S = +4}$

* $\underset{?}{\underline{Cr_2}}O_7^{-2}$ $2Cr - 14 = -2$ $2Cr = 12$ $\boxed{Cr = +6}$

* **Exercise :-**

Find the oxidation state of each underlined species.

$\underline{Fe_2}O_3^{-2}$
$2Fe + 3(-2) = 0$
$2Fe - 6 = 0 + 6$
$\boxed{Fe = +3}$

$\underline{O}F_2^{-1}$
$O + 2(-1) = 0$
$O - 2 = 0$
$\boxed{O = +2}$

$\underline{S}O_4^{-2}$
$S + 4(-2) = -2$
$S - 8 = -2 + 8 + 6$
$\boxed{S = +6}$

$V_2\overset{-2}{O}_5$

$2V+5(-2)=0$

$2V-10=0$

$2V=10$

$\boxed{V=+5}$

$\overset{-2}{PO_4}{}^{-3}$

$P+4(-2)=-3$

$P-8=-3$

$\boxed{P=+5}$

$CrO_4{}^{-2}$

$Cr-8=-2$

$\boxed{Cr=+6}$

CO^{-2}

$C-2=0$

$\boxed{C=+2}$

$CO_2{}^{-2}$

$C+2(-2)=0$

$C-4=0$

$\boxed{C=+4}$

$CO_3{}^{-2}$

$C+3(-2)=0$

$C-6=0$

$\boxed{C=+6}$

* 'important !

- $\overset{0}{Al}+\overset{+3}{Fe_2}\overset{-2}{O_3}\longrightarrow\overset{+3}{Al_2}\overset{-2}{O_3}+\overset{0}{Fe}$ oxidation : $Al\rightarrow Al^{3+}$ reduction , $Fe^{3+}\rightarrow Fe$

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

  oxidation : $Fe^{2+}\rightarrow Fe^{3+}$ reduction : $Mn^{7+}\rightarrow Mn^{2+}$

- $Cr_2O_7^{2-}+Sn^{2+}+H^{1+}\rightarrow Cr^{3+}+Sn^{4+}+H_2O$

  oxidation : $Sn^{2+}\rightarrow Sn^{4+}$ reduction : $Cr^{6+}\rightarrow Cr^{3+}$

Writing balanced half ionic equation.

* Balance the :-

1) Atoms

2) the charge by adding é's to the side with greater charge by the difference.

H.O.O → half of oxidation.

H.O.R → half of reduction.

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

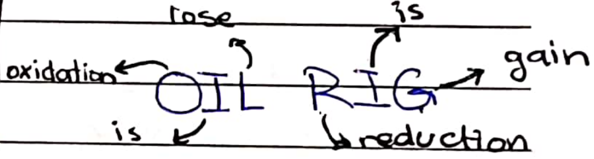
H.O.O : $Ca \rightarrow Ca^{2+} + \cancel{2e}$

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

overall reaction $Ca + Zn^{2+} \rightarrow Ca^{2+} + Zn$

oxidation → é "lose ē"

reduction ē → "gain ē



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

H.O.O : $(2Na \longrightarrow 2Na^{1+} + \cancel{2\bar{e}}) * 2$

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

overall reaction $2Na + 2H \rightarrow 2Na^{1+} + H_2$

* $\overset{0}{Ca} + \overset{3+}{Al}\overset{1-}{Cl_3} \longrightarrow \overset{2+}{Ca}\overset{1-}{Cl_2} + \overset{0}{Al}$

H.O.O : $(3Ca \longrightarrow 3Ca^{2+} + \overset{6}{\cancel{2e}}) * 3$

H.O.R : $(2Al^{3+} + \overset{6}{\cancel{3e}} \rightarrow 2Al) * 2$

overall reaction $3Ca + 2Al^{3+} \rightarrow 3Ca^{2+} + 2Al$

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

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

H.O.R : $Cl_2 + \cancel{2e} \rightarrow 2Cl^{-1}$

overall reaction $S^{2+} + Cl_2 \rightarrow S + 2Cl^{1-}$

$H^{+1} + H^{+1} = \underline{\underline{+2}}$

* $\overset{0}{Al} + \overset{+1\ -1}{HCl} \longrightarrow \overset{+3\ -1}{AlCl_3} + \overset{0}{H_2}$

H.O.O : $(2Al \longrightarrow 2Al^{3+} + \overset{6}{\cancel{3}}\bar{e}) * 2$

H.O.R : $(\overset{6}{\cancel{2}}H^{+} \longrightarrow \overset{6}{\cancel{2}}\bar{e} + 3H_2) * 3$

overall reaction $2Al + 6H^{1+} \longrightarrow 2Al^{3+} + 3H_2$

* $\overset{+7\ -2}{MnO_4}^{-1} + Fe^{2+} + H^{1+} \longrightarrow Mn^{2+} + Fe^{3+} + H_2O$

H.O.O : $(5Fe^{2+} \longrightarrow 5Fe^{3+} + \cancel{5\bar{e}}) * 5$

H.O.R : $Mn^{7+} + \cancel{5\bar{e}} \longrightarrow Mn^{2+}$

overall reaction $5Fe^{2+} + Mn^{7+} \rightarrow 5Fe^{3+} + Mn^{2+}$

* $\overset{6+}{Cr_2}\overset{2-}{O_7} + Sn^{2+} + H^{1+} \longrightarrow Cr^{3+} + Sn^{4+} + H_2O$

H.O.O : $(3Sn^{2+} \longrightarrow 3Sn^{4+} + \overset{6}{\cancel{2}}\bar{e}) * 3$

H.O.R : $(2Cr^{6+} + \overset{6}{\cancel{3}}\bar{e} \longrightarrow 2Cr^{3+}) * 2$

overall reaction $3Sn^{2+} + 2Cr^{6+} \rightarrow 3Sn^{4+} + 2Cr^{3+}$

Oxidising agent + Reducing agent

*** Oxidising agent (oxidant) :-**
the substance that itself reduced and causes the other substance to be oxidised.

*** Reducing agent (reductant):-**
the substance that itself oxidised and causes the other substance to be reduced.

	reducing agent	oxidising agent
	oxidation	reduction
O	gain O	lose O
H	lose H	gain H
oxi. state	↑	↓
ē - transfer	lose ē	gain ē

if the substance is an ion, in a compound the agent is the compound itself
~~oxidation~~

- $\overset{0}{Al} + \overset{3+}{Fe_2}\overset{-2}{O_3} \longrightarrow \overset{3+}{Al_2}\overset{2-}{O_3} + \overset{0}{Fe}$

oxidation : Al
reduction: Fe^{3+}

oxidising agent : Fe_2O_3
reducing agent : Al

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

oxidation: C^{2+}
reduction: Zn^{2+}

oxidising agent : ZnO
reducing agent: CO

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

oxidant : Cl_2
reductant : H_2S

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

oxidant : MnO_4^{-1}
reductant : Fe^{2+}

* most common oxidising agent :

1) oxygen

$\overset{0}{Fe} + \overset{0}{O_2} \longrightarrow \overset{3+}{Fe_2}\overset{2-}{O_3}$ $\overset{0}{S} + \overset{0}{O_2} \longrightarrow \overset{4+}{S}\overset{2-}{O_2}$

2) Acidify pottasium manganate $(KMnO_4) / H^+$

$MnO_4^{-1} \longrightarrow Mn^{2+}$

(+7) (+2)

Purple colourless

3) Acidify pottasium dicromate $K_2Cr_2O_7 / H^+$

$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

(+6) (+3)

orange green

4) Halogens

yellow $Fe_2 + 2\bar{e} \rightarrow 2Fe^-$

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

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

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

black solid red brown solution (aq)

} colourless solution (ions)

* most common reducing agent :-

1) Hydrogen (H_2) $\overset{0}{H_2} + \overset{2+}{Cu}O \longrightarrow \overset{1+}{H_2}O + \overset{0}{Cu}$

2) Carbon and Carbon monoxide

$$\overset{0}{C} + \overset{3+}{Fe_2}O_3 \longrightarrow \overset{0}{Fe} + \overset{2+}{C}O$$

$$CO + \overset{2+}{Zn}O \longrightarrow CO_2 + \overset{0}{Zn}$$

3) Iodide

$$2I^-_{(aq)} \longrightarrow I_{2(aq)} + 2\bar{e}$$

colourless red brown

4) Metals

metals that are more reactive are more likely to lose ē, so more likely to be ~~oxidising~~ a reducing agent

- K → strongest reductant K^+ weakest oxidant
- Na
- Li
- Ca
- Mg
- Al
- CuCO ~~CaCO~~
- Zn
- Fe
- Pb
- H
- Cu
- Ag weakest reductant Ag^+ → strongest oxidant

I^- is a halide
not halogen

- halides are ions
- halogens are atoms

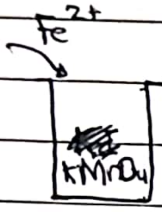
Question :-

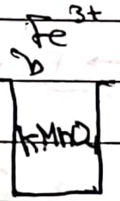
Fe^{2+} is a reducing agent.

Fe^{3+} is an ~~oxidin~~ oxidising agent.

record the observation in each of the following :-

①

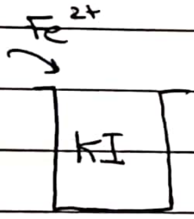


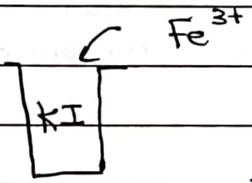


colour change from purple to colourless

stays purple

②





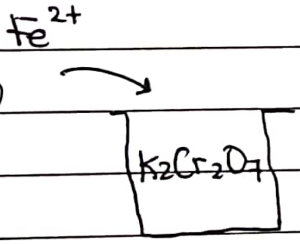
stays colourless

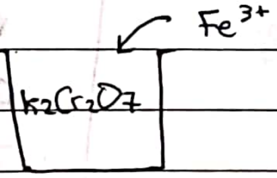
red brown

$2I^- \rightarrow \underbrace{I_2}_{(aq)} + 2\bar{e}$

$Fe^{3+} + I^- \rightarrow Fe^{2+} + I_2$

③





change from orange to green.

no change

Electrolysis

electrol/ysis

↙ ↘

electricity — "flow of electrons"

analysis — "breaking down"

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

* Why the ionic compounds don't conduct electricity when solid?

the ions are not free to move.

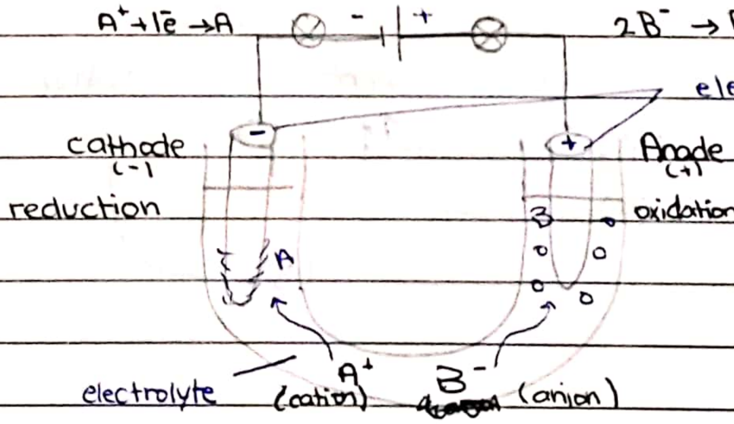
* Why the ionic compounds conduct electricity when dissolved in water or being molten?

they are free to move (mobile electrons).

- **Electrolyte**: Chemical compounds that conducts electricity when molten or ageous.

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

* Electrolysis Cell :-



Inert (unreactive)
- graphite
- pt

Reactive
- Cu
- Ag

*** Cathode :-** the -ve rod that attracts the +ve Ions (cations) where the reduction occurs.

*** Anode :-** the +ve rod that attracts the -ve ions (anions) where the oxidation occurs.

Electrolysis = Discharging

Ion ⟶ element

Cations	Anions
$1\bar{e} + Na^{+} \rightarrow Na$	$2F^{-1} \rightarrow F_2 + 2\bar{e}$
$2\bar{e} + Mg^{2+} \rightarrow Mg$	$2Cl^{-1} \rightarrow Cl_2 + 2\bar{e}$
$2\bar{e} + Ca^{2+} \rightarrow Ca$	$2Br^{-1} \rightarrow Br_2 + 2\bar{e}$
$3\bar{e} + Al^{3+} \rightarrow Al$	$2I^{-1} \rightarrow I_2 + 2\bar{e}$
$2\bar{e} + 2H^{+} \rightarrow H_2$	$2O^{-2} \rightarrow O_2 + 4\bar{e}$

~~* electrolysis depends on 2 things~~

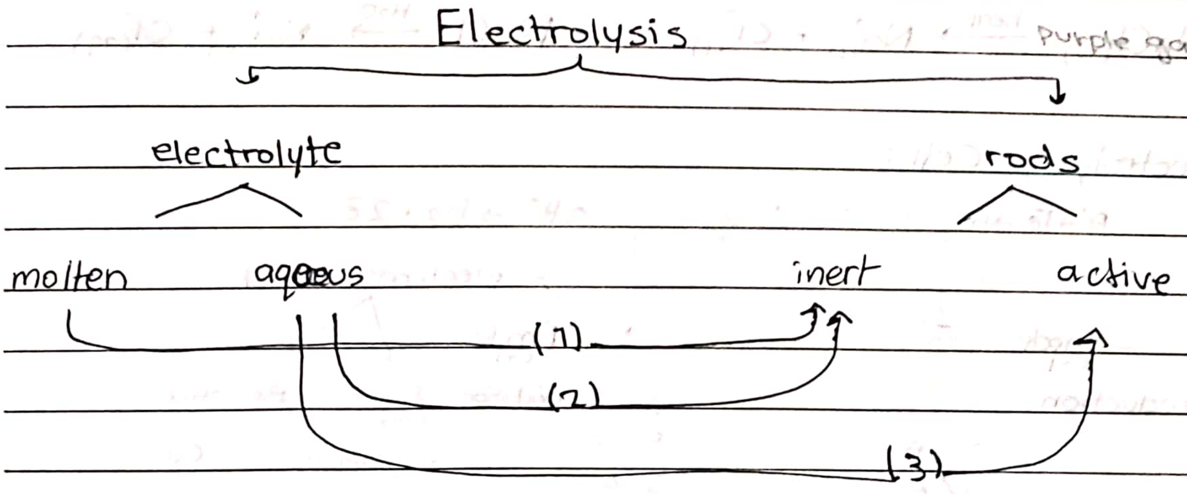
~~elec~~

F_2: yellow gas

Br_2 : red-brown gas

Cl_2 : green yellow gas

I_2 : Black solid / Purple gas / red-brown solution



electrolyte always used up / consumed

* electrolysis for molten electrolyte using inert rods (graphite) :-

* $NaCl_{(l)}$ - electrolyte : used up

Total ions Na^+ Cl^-

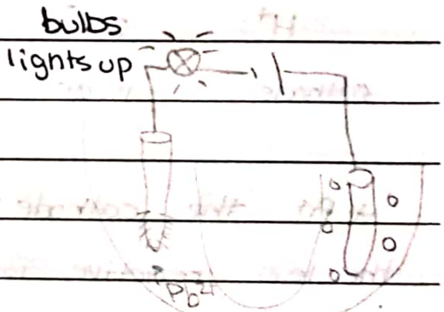
cathode	anode
$Na^+ + 1\bar{e} \rightarrow Na$	$2Cl^- \rightarrow Cl_2 + 2\bar{e}$
- deposite of metal	- bubbles of green yellow gas.

* molten lead (II) Bromide :-

$PbBr_{2\,(l)}$ - electrolyte : used up

Total ions Pb^{2+} Br^-

cathode	anode
$2\bar{e} + Pb^{2+} \rightarrow Pb$	$2Br^{-1} \rightarrow Br + 2\bar{e}$
- deposite of metal	- bubbles of red-brown gas.



* $CuCl_2$ / graphite

$CuCl_{2\,(l)}$ - electrolyte : consumed

Cu^{2+} Cl^{1-}

cathode	anode
$2\bar{e} + Cu^{2+} \rightarrow Cu$	$2Cl^- \rightarrow Cl_2 + 2\bar{e}$
- deposite of red-brown solid.	- bubbles of green yellow gas.

* Al_2O_3 / graphite

Al_2O_3 (l) — electrolyte : used up

$Al^{3+} \leftarrow \quad \rightarrow O^{2-}$

cathode — anode

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

- deposite of metal — - bubbles of colourless gas

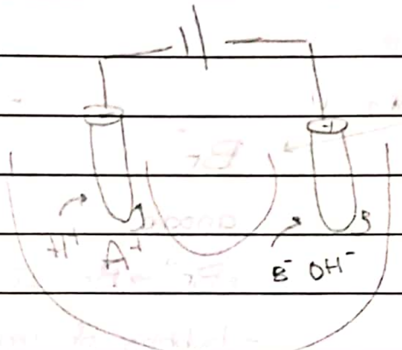
* Electrolysis of ageous Electrolyte using inert :-

AB (aq)

$A \leftarrow A^{+}$ — B^{-}

$H_2 \leftarrow H^{+}$ — OH^{-}

cathode — anode



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 and the more reactive stays in the electrolyte.

* At the anode :-

the OH^{-} always discharges (oxidises) except with concentrated halides . Cl^{-}, Br^{-}, I^{-}

- when halide oxidises :-

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

very important

- when OH^{-} oxidises :-

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

- bubbles of colourless gas

* Concentrated $NaCl_{(aq)}$ / graphite

$NaCl$

$Na^+ \leftarrow \quad \rightarrow Cl^-$

$H^+ \qquad OH^-$

cathode: $2\bar{e} + {}_2H^+ \rightarrow H_2$
- bubbles of colourless gas

anode: ${}_2Cl^- \rightarrow Cl_2 + 2\bar{e}$
- bubbles of green yellow gas

electrolyte : $NaOH$

* $CuSO_4$ / graphite

$CuSO_4$

$Cu^{2+} \leftarrow \quad \rightarrow SO_4^{2-}$

$H^+ \qquad OH^-$

cathode: $2\bar{e} + Cu^{2+} \rightarrow Cu$
- deposite of red-brown metal

anode: $4OH^- \rightarrow 2H_2O + O_2 + 4\bar{e}$
- bubbles of colourless gas.

electrolyte : H_2SO_4

* dilute $NaCl_{(aq)}$ / graphite

$NaCl$

$Na^+ \leftarrow \quad \rightarrow Cl^-$

$H^+ \qquad OH^-$

cathode: $2\bar{e} + {}_2H^+ \rightarrow H_2$
- bubbles of colourless gas

anode: $4OH^- \rightarrow 2H_2O + O_2 + 4\bar{e}$
- bubbles of colourless gas.

electrolyte : $NaCl$

* conc. $KI_{(aq)}$ / graphite

KI

$K^+ \leftarrow \quad \rightarrow I^-$

$H^+ \qquad OH^-$

cathode: $2\bar{e} + {}_2H^+ \rightarrow H_2$
- bubbles of colourless gas

anode: $2\bar{e} + {}_2I^- \rightarrow I_2$
- red-brown solution

electrolyte : KOH

* conc. $CuCl_{2(aq)}$ / graphite

$CuCl_2$

$Cu^{2+} \leftarrow \quad \rightarrow Cl^-$

$H^+ \qquad OH^-$

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

anode: $2\bar{e} + {}_2Cl^- \rightarrow Cl_2$

electrolyte : no need.

* Concentrated Sodium Chloride called Brine Solution.

$NaCl(aq)$

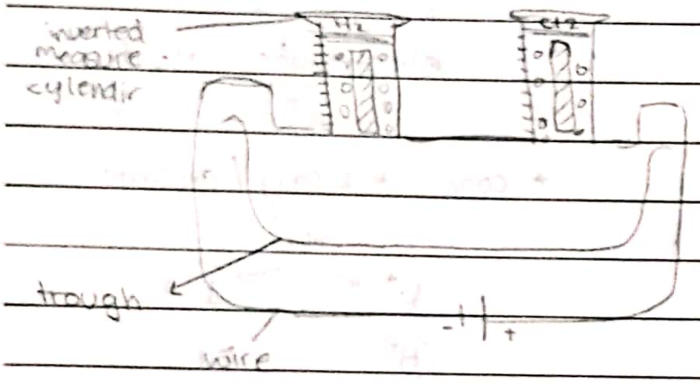
Na^+ Cl^-

H^+ OH^-

callode anode electrolyte : NaOH

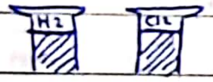
$2\bar{e} + H^+ \rightarrow H_2$ $2\bar{e} + Cl^- \rightarrow Cl_2$

Q1 :- plan an experiment to collect & measure the volume of the H_2 & Cl_2 produced.



inverted measuring cylender method :

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



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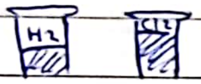

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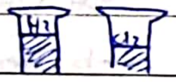


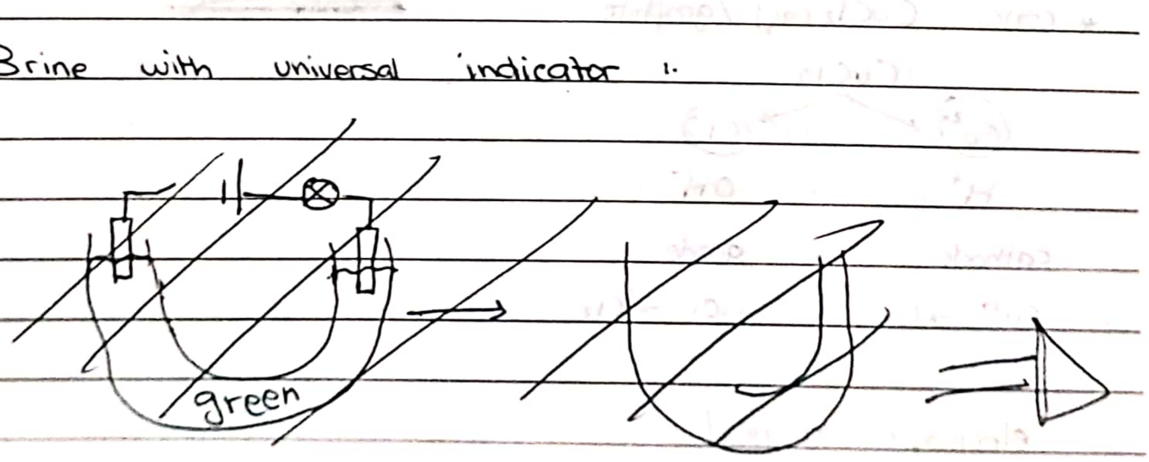
(c)

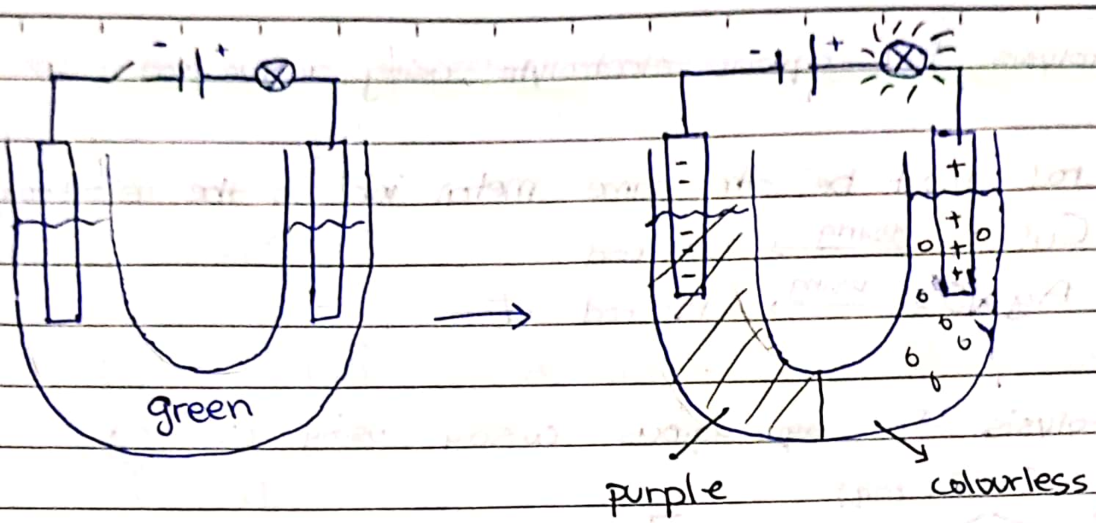
H.O.O : $2Cl^- \rightarrow Cl_2 + \cancel{2\acute{e}}$

H.O.R : $2H^+ + \cancel{2\acute{e}} \rightarrow H_2$

overall reaction: $2H^+ + 2Cl^- \rightarrow \boxed{H_2 + Cl_2} \rightarrow 1 : 1$

Q3 :- Brine with universal indicator :





* explain why the Cl_2 didn't immediatly appear as H_2 produce?

Some Cl_2 dissolves in solution.

- Observation :-

1) Bulb lights up.
2) Bubbles of green yellow gas on the anode (oxidation of Cl^-)
3) Bubbles of colourless gas on the cathode (reduction of H^+)
4) Around the cathode, the solution becomes purple. because NaOH is an alkali.
5) Around the anode, the solution becomes colourless since Cl_2 bleaches the colour.
~~6) Rotate the spoon to ensure an equal distribution.~~
~~7) Rinse with distilled water.~~
~~8) Dry in oven.~~

~~* Purifying Metals / Refining Copper~~ :

* Electrolysis for aqeous electrolyte using active rod :-

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

$CuSO_4 \xrightarrow{\text{using}}$ Cu rod

$AgNO_3 \xrightarrow{\text{using}}$ Ag rod

ex:

* elecrolysis for ~~aqu~~ aqous $CuSO_4$ using Cu rod.

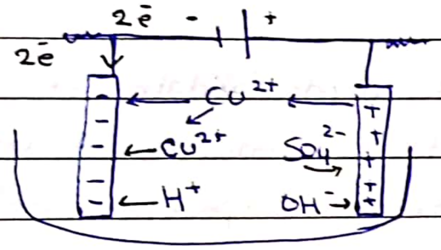
$CuSO_4$ (aq)

Cu^{2+} ← → SO_4^{2-} X

H^+ OH^- X } the anode itself oxidises.

cathode anode

$2\bar{e} + Cu^{2+} \rightarrow Cu$ $Cu \rightarrow Cu^{2+} + 2\bar{e}$ (anode)

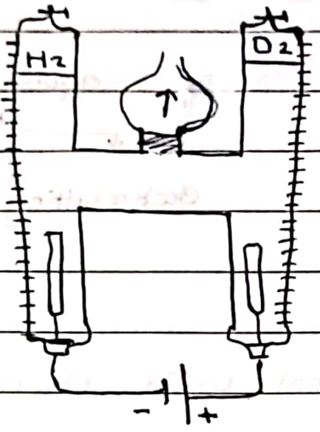


* electrolyte just to allow ions to move through it

* why electrolyte doesn't change its conc. [2]

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

* electrolysis for H_2SO_4 (Hoffman) :-



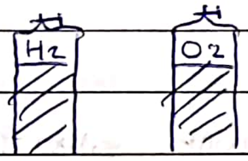
H_2SO_4

$H^+ \leftarrow \quad \rightarrow SO_4^{2-}$

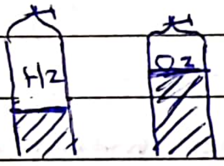
$H^+ \rightarrow H_2$ cathode

$OH^- \rightarrow O_2$ anode

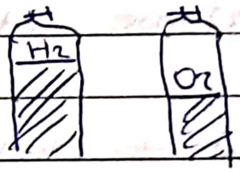
Q :- the final appearance :-



(a)



[✓ (b)]



(c)

$$4H^+ + \cancel{2e} \rightarrow 2H_2$$

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

$$4H^+ + 4OH^- \rightarrow 2H_2O + \underbrace{2H_2 + O_2}_{2:1}$$

Applications on Electrolysis

molten/ inert → extraction of metal

aqueous/inert → prepare solution

aqueous active → electroplatting, purifiying metal

* Electroplatting :-

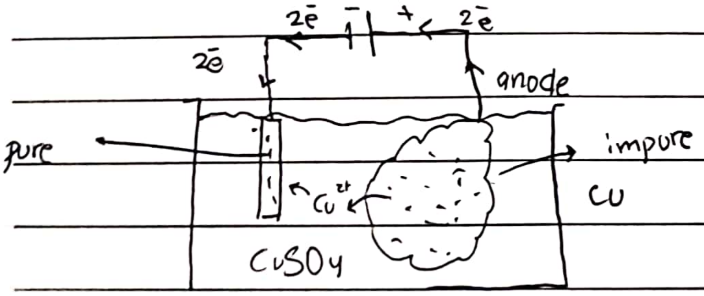
- coating a metal with another metal using electrolysis.

why? 1) to prevent rusting
2) more attractive

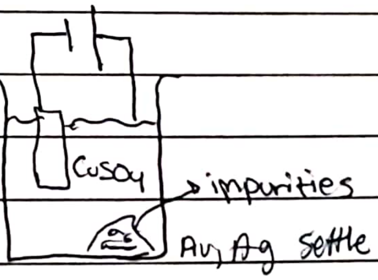
* How to electroplate a metal spoon with silver?

1) Clean the spoon from any imparities of oxide layer using sand paper, to ensure a good sticking.
2) make the spoon the cathode (-ve)
3) the anode must be Ag.
4) the electrolyte must contain Ag^+ eg. $AgNO_3$
5) turn on the circuit "be sure that the spoon is fully immersed in the electrolyte."
6) Rotate the spoon to ensure an equal distribution.
7) Rinse with distiled water
8) dry in oven.

* Purifying Metals / Refining Copper :



Cu
* impurities
Ag, Au, Zn



Au, Ag settle down (less reactive)

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

* Extraction of Metals from their ions :-

The method of extraction depends on the position of metals in the reactivity series :

K
Na
Li — Electrolysis
Ca — Molten / Graphite
Mg
Bauxite Al_2O_3 ← Al

C, CO

Zinc Blend ZnS ← Zn] reduction by
Hematite Fe_2O_3 ← Fe] C, CO
Pb

H

Copper Sulfide CuS ← Cu] reduction by H_2
Ag
Au

- Extraction of Aluminium :

Ore → Al_2O_3 Bauxite

- method : electrolysis for molten Al_2O_3 / graphite.

* m.p of Al_2O_3 is about 2000°C

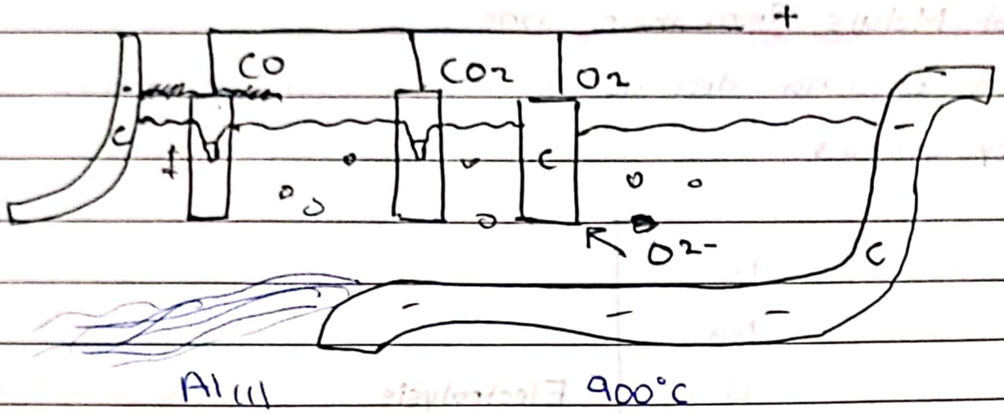
So we dissolve Al_2O_3 in a molten Cryolite Na_3AlF_6

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

$Al_2O_3(l)$

cathode Al^{3+} ↶↷ O^{2-} Anode

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



gases produced at Anode :-

1- O_2

2- CO_2] reaction of rods - with O_2

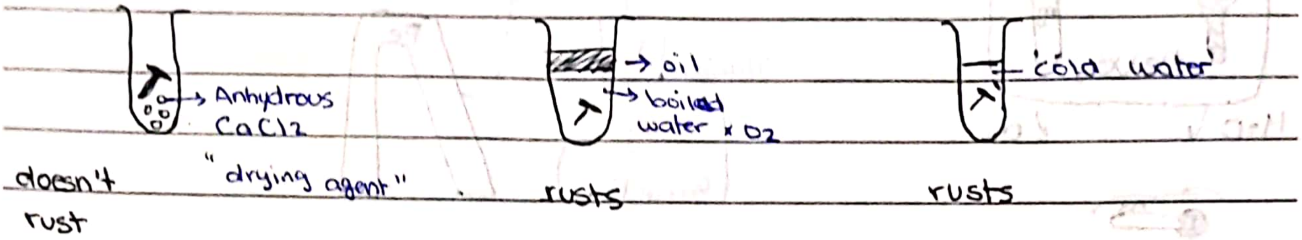
3- CO] so we must replace them periodically

property	use
- low density	- Aircraft bodies
- ductile	- electrical wires
- malleable	- window frame / cooking utensils.
- conduct electricity	- wires
- form an oxide layer which is non-toxic.	- Food cans.

$$Fe + O_2 \xrightarrow{H_2O} \overbrace{Fe_2O_3 \cdot nH_2O}^{\text{brown colour}}$$

*** Rusting :-** → slow reaction (6-7 days)

- Reaction of Iron with both O_2 and H_2O. ~~STP~~



- Plan an expirement to show which rust prevention 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 and measure the mass again. Repeat the experement with the 2nd solution.

Conclusion :- the exp. which causes more increase in mass, is the worst solution.

- How to prevent rusting :-
  - painting
  - oiling
  - greasing
  - cover with plastic

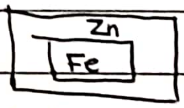
⟹ to prevent O_2 and H_2O from reaching the iron.

Rusting Prevention

- Galvanizing
- electroplating
- cathodic protection
- Sacrificial Protection

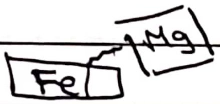
*** Galvanizing :-**

Coating Fe with Zn



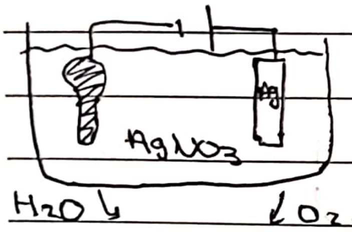
*** Sacrificial Protection :-**

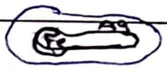
connecting Fe with Mg



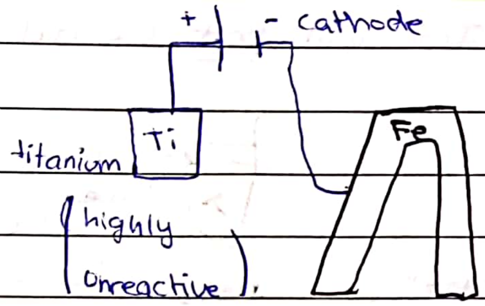
* Zn & Mg are more reactive than Fe, so they are more likely to oxidise, more likely to lose ē, so Fe is less likely to rust.

* Electroplating :-





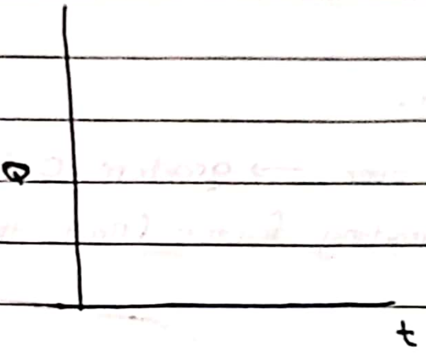
* Cathodic Protection :-



Rate of Reaction

$$\text{Rate} = \frac{\text{change in Quantity}}{\text{change in time}}$$

$$\text{Rate} = \frac{\Delta Q}{\Delta t}$$

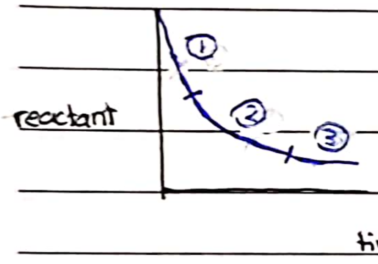


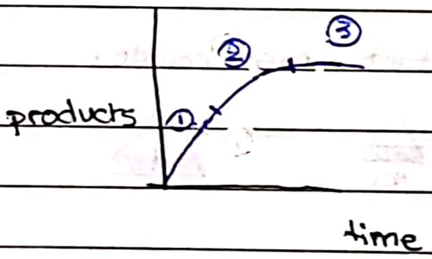
* Quantity can be :·

- $\frac{\Delta \text{mass}}{\Delta \text{time}}$
- $\frac{\Delta \text{PH}}{\Delta \text{time}}$
- $\frac{\Delta \text{temp.}}{\Delta \text{time}}$
- $\frac{\Delta \text{light intensity}}{\Delta \text{time}}$
- $\frac{\Delta \text{conc.}}{\Delta \text{time}}$
- $\frac{\Delta \text{volume}}{\Delta \text{time}}$
- $\frac{\Delta \text{heigh of ppt}}{\Delta \text{time}}$

measure the rate of reaction

how fast the reactants consumed per unit time ← → How fast the products producted per unit time





* region ① :-

fastest rate → from the graph, steepest

more amount of reactants/products , more particles , more effective collisions per unit time.

* region ② :-

Slower rate → from the graph less steep.

→

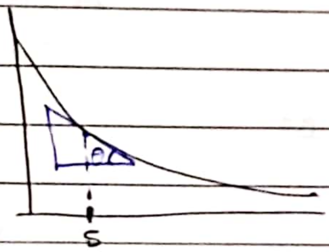
less number of particles, so less number of effective collisions per unit time.

* region ③ :-

reaction is over → gradient = 0 (horizontal)

no more limiting factor (limits the product), no more effective collisions

* To measure the gradient :-



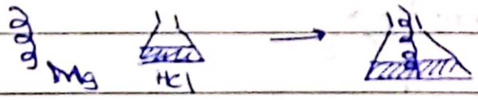
$\theta = \tan(\frac{O}{A})$

* For any chemical reaction there are three main conditions :-

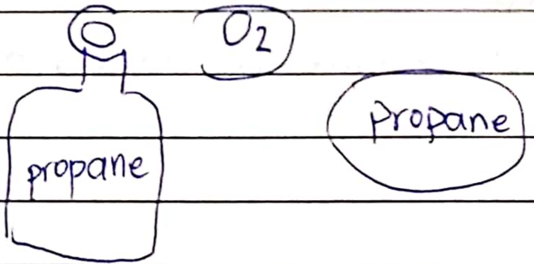
1) the reactant must be suitable

$Cu + HCl \rightarrow$ no reaction

2) the reactants must collide.



3) The collisions must be effective / min. amount of energy to start the reaction. Activation Energy Ea.

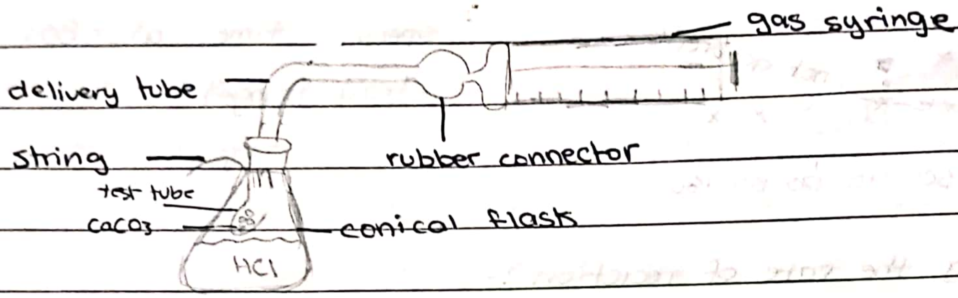


→ ALWAYS not constant

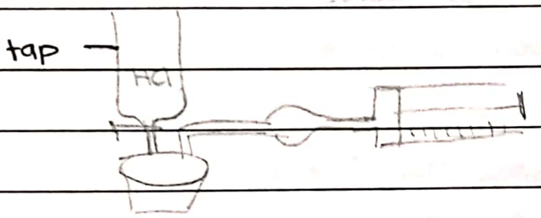
*** measuring rate of reaction :**

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

1) measure the volume of gas per unit time.



OR

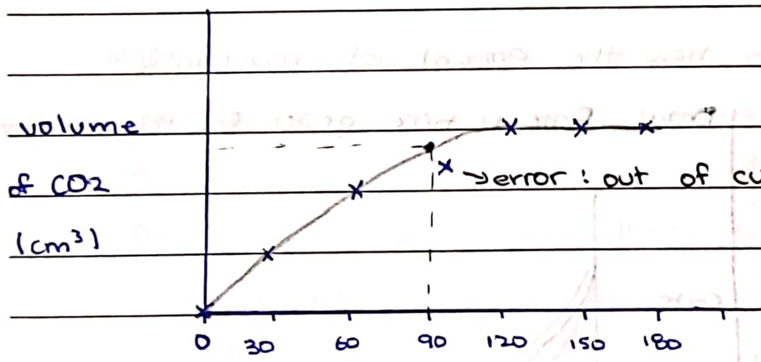


difference between tap & beurette is tap has no measurements while the beurette does.

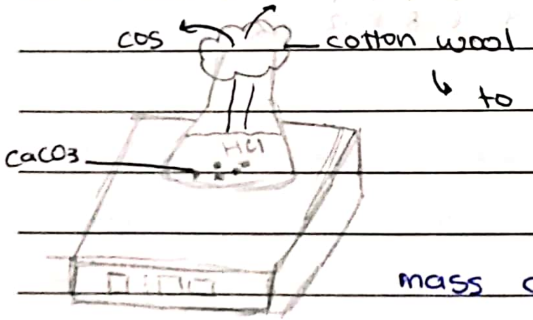
↓ over

time (s)	0	30	60	90	120	150	180
vaume of CO_2 (cm^3)	0	10	15	17	17.5	17.5	17.5

+10 +5 +2 +0.2



2) measure the mass of the conical flask + contents per unit time

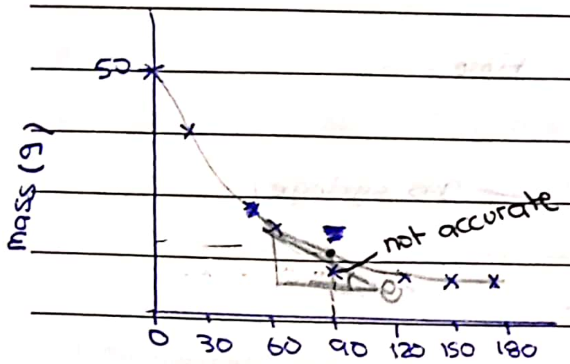


↳ to allow CO_2 to escape & prevent splashing

mass decreases because 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

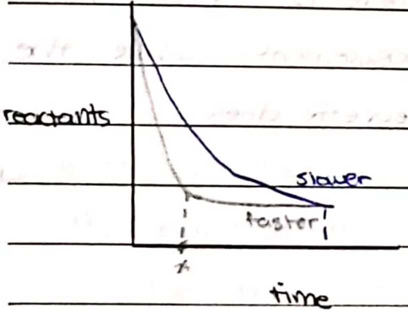


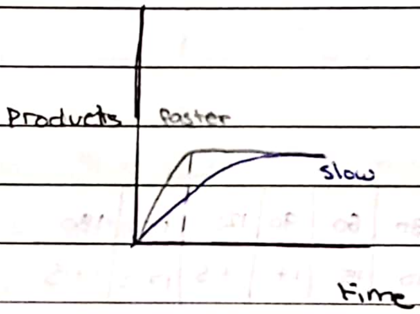
to find the rate at specific time at = 80s draw tanget

* Increasing the rate of reaction :-

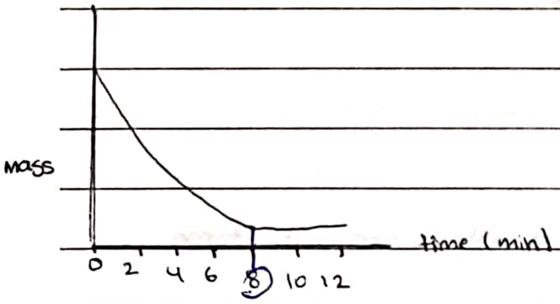
- more product per same period of time.
OR
- same products per less time

} steeper curve

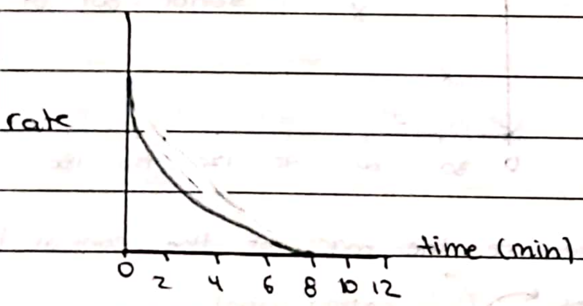




Q :- the graph below shows how the amount of reactants changes with time.



Draw Rate vs time graph for this reaction



* Factors that affect the rate of reaction :-

1) Temperature :-

- State how the temp. affects the rate of reaction.

as the temp. increases, the rate of reaction increases.

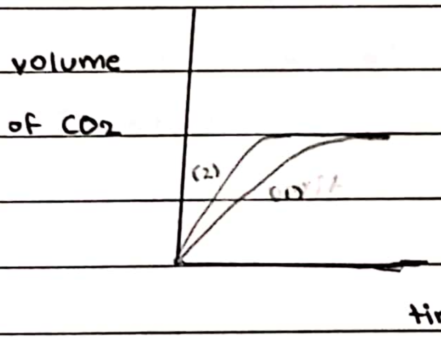
- Explain how the temperature affects the rate of reaction. [5]

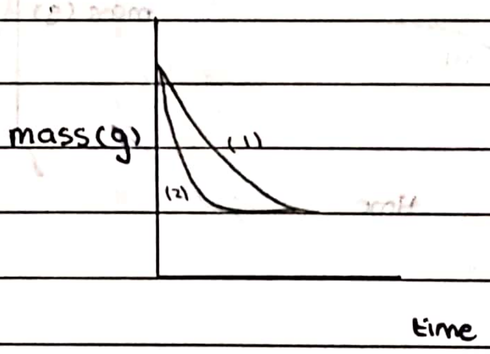
as the temp. increases, the particles will gain kinetic energy so they move faster. The particles will have more kinetic energy so they move faster. The particles will have energy equal to or greater than the Ea. So more effective collisions per unit time. So faster rate of reaction

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

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

- conical flask 1 :-	- conical flask 2 :-
mass = 2g (lumps)	mass = 2g (lumps)
V_{HCl} = 0.1 dm^3	V_{HCl} = 0.1 dm^3
M_{HCl} = 1 mol/dm^3	M_{HCl} = 1 mol/dm^3
Temp = 25°C	Temp = 50°C





- take a known mass of lumps $CaCO_3$, add them to a known volume of known conc. of 25°C, measure the volume of CO_2 produced. Repeat the exp. at 50°C. The exp. at 50°C produces CO_2 with less time.

2) Surface Area :-

- State how the surface area affects the rate of reaction.

as the surface area increases, the rate of reaction increases.

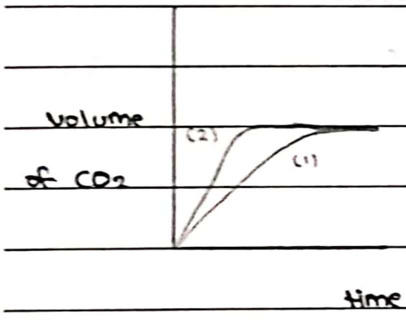
- Explain how the surface area affects the rate of reaction.

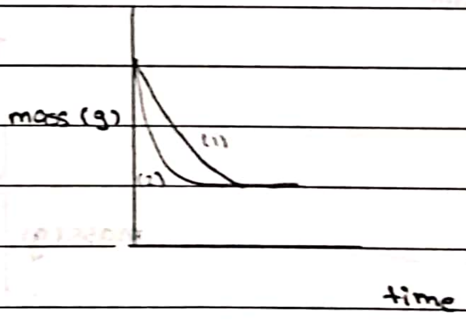
as the surface area increases, (decrease the particle size by crushing using mortar & pestile) more particles exposed to the reaction, more effective collisions per unit time so faster rate.

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

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

- conical flask 1 :-	- conical flask 2 :-
mass = 2g (lumps)	mass = 2g (powder)
V_{HCl} = 0.1 dm³	V_{HCl} = 0.1 dm³
M_{HCl} = 1 mol/dm³	M_{HCl} = 1 mol/dm³
Temp = 25°C	Temp = 25°C





3) Concentration (Amount) :-

- State how the conc. affects the rate of reaction.

as the conc. increases, the rate of reaction increases.

- Explain how the conc. affects the rate of reaction.

as the conc. increases, more particles, → more effective collisions per unit time, so faster rate of reaction.

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

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

- conical flask 1 :-

mass = 2g (lumps)

$V_{HCl} = 0.1 dm^3$

$M_{HCl} = 0.1 mol/dm^3$

Temp = 25°C

- conical flask 2 :-

mass = 2g (lumps)

$V_{HCl} = 0.1 dm^3$

$M_{HCl} = 0.2 mol/dm^3$

Temp = 25°C

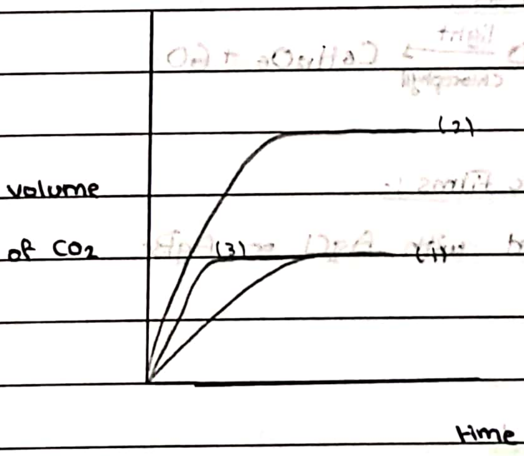
- conical flask 3 :-

mass = 4g (lumps)

$V_{HCl} = 0.1 dm^3$

$M_{HCl} = 0.1 mol/dm^3$

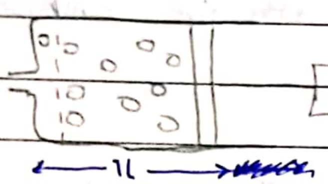
Temp = 25°C

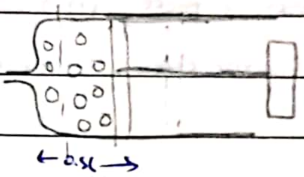


4) Pressure :- (affects only gas)

- Explain how the pressure affect the rate of reaction.

as the pressure increases (by reducing the volume) more particles particles per unit volume, more effective collisions per unit time so faster rate of reaction.





$\frac{10 \text{ particles}}{1L} = 10$ $\frac{10 \text{ particles}}{0.5L} = 20$

5) Light (only for photochemical reactions) :-

* reaction that need light to occur.

- Photosynthesis :-

$6CO_2 + 6H_2O \xrightarrow[\text{Chlorophyll}]{\text{light}} C_6H_{12}O_6 + 6O_2$

- Photographic films :-

* films coated with AgCl or AgBr

$AgCl \xrightarrow{\text{light}} Ag_{(s)} + Cl_2$

$AgBr \xrightarrow{\text{light}} Ag_{(s)} + Br_2$

6) Catalysts :-

Define :- Chemical substance that speeds up the reaction without being used up.

How?

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

- The reaction : $2H_2O_{2(aq)} \longrightarrow 2H_2O_{(l)} + O_{(g)}$ has 2 catalysts CuO / MnO_2

1) Plan an exp. to show that CuO is a catalyst for this reaction.

take a known volume with a known conc. of H_2O_2 at known temp, measure the volume of O_2 produced per unit time. Repeat the exp. using CuO. The exp. using CuO will produce more O_2 per the same unit time.

2) Plan an exp. to show which of the 2 catalysts is better CuO or MnO_2.

Same as 1 + same mass of catalysts. The exp. which produces more O_2 per unit time. used the better catalyst.

3) Plan an exp. to show that CuO is not used up during the reaction.

measure the mass of CuO, add H_2O_2 until no more fizzes. Filter the mixture, dry solid in oven. Remeasure the mass. Same initial and final mass.

Reversible Reaction

Types of chemical reactions

one way

reactants $\xrightarrow{\text{forward reaction}}$ products

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

irreversible reaction

both ways

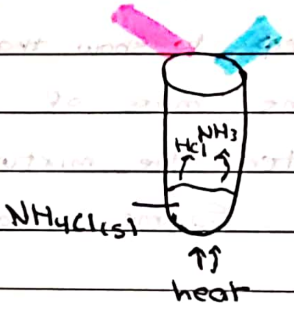
reactants $\underset{\text{backwards}}{\overset{\text{forward}}{\rightleftharpoons}}$ products

$CaCO_3 \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

reversible reaction

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

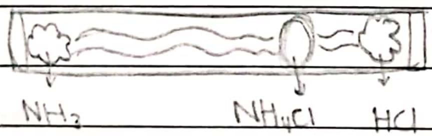
Ammonium Chloride — Ammonia — Hydrogen Chloride



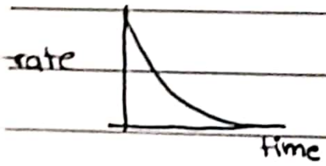
* Which damp litmus paper will change its colour first, why?

the damp red litmus paper changes to blue, ~~First~~ because NH_3 is an alkali and lighter than HCl which is acidic.

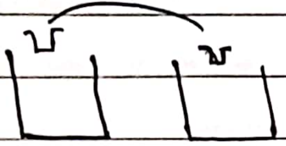
- $HCl_{(g)} + NH_{3(g)} \longrightarrow NH_4Cl_{(s)}$

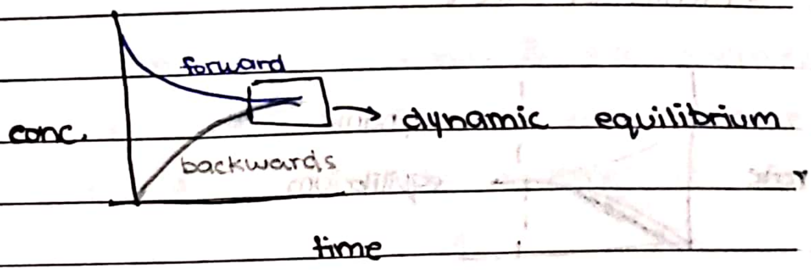


* R → P



* Same volume of beaker

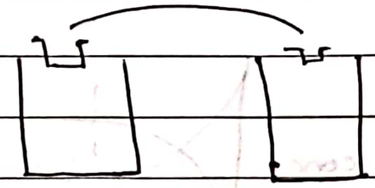


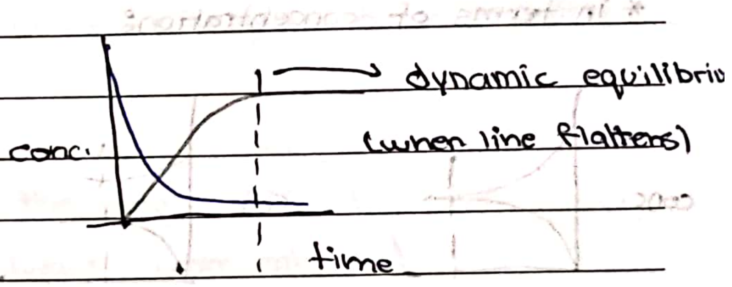


conc. R ↓ until constant

conc. P ↑ until constant

* different volume of beaker





*Dynamic Equilibrium :-

- in terms of rate :-

when rate of forward reaction is EQUAL to the rate of backward reaction.

- in terms of conc. :-

when the conc. of reactant and product are CONSTANT.

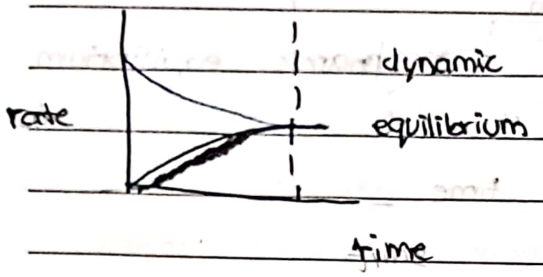
* Why the rate of forward reaction decrease with time ?

less reactants so less particles, so less effective collisions per unit time.

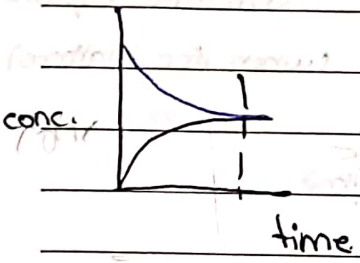
* why the rate of backward reaction increase with time ?

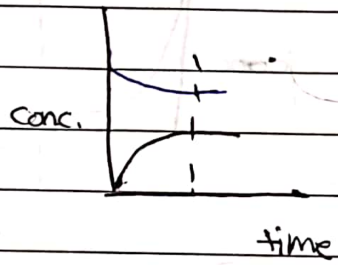
more product so more particles, so more effective collisions per unit time.

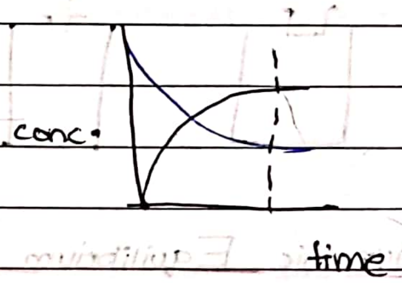
* in terms of rate:



* in terms of concentration:







* Le Chatelier principle :-

If the system is at equilibrium $\rightleftharpoons$ and any external factor disrubts the equilibrium so the equilibrium can shift itself either to

forward $\rightleftharpoons$

backward $\rightleftharpoons$

to return back to the equilibrium.

Factors that affect the position of equil.

Temperature	Pressure	Concentration
endo. & exo.	- ↑ press. shifts to less gas moles.	↑ R ↓ P } shifts Forward
- ↑ Temp. Shifts to endo	- ↓ press. shifts to more gas moles.	
- ↓ Temp. Shifts to exo.	- ↑ press. ↑ rate of less gas moles ↑ rate of more gas moles.	↓ R ↑ P } shifts backward
- ↑ temp. rate ↑ - ↓ temp. rate ↓	- ↓ press. ↑ rate of less/more	

1) Temperature :-

* temp. increases :-

↑↑ rate of endothermic

↑ rate of exothermic

both rate of endo. & exo. increase but the rate of endo. increases more, so it shifts to endothermic.

* temp. decreases :

↓↓ rate of endothermic

↓ rate of exothermic

both rate of endo. & exo. decrease but the rate of endo. decreases more, so it shifts to exothermic.

* ΔH =

enthalpycharge

+ve ↙ ↘ -ve

gain lose

endo. exo.

the sign ΔH always represents the forward reaction

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

→ forward

← backward

↑ Temp. ↑↑ rate of forward

↑ rate of backward

↓A ↓B ↑C ↑D

shifts forward to ~~the~~ endo.

↓ Temp. ↓↓ rate of forward

↓ rate of backward

↑A ↑B ↓C ↓D

shifts backward to exo.

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

↑ Temp. ↑ rate of forward

↑↑ rate of backward

] → ↑X ↑W ↓Y ↓Z

shifts to endo.

Question 1 :-

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

- to produce more yeild of NH_3 you must :

use low temp.

to favor the forward reaction which , in this case, is the exothermic reaction.

Question 2 :-

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

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

Question 3 :-

mixture of NO_2 and N_2O_4 at equil. in a sealed tube.

$$\underset{\text{dark brown}}{2NO_{2(g)}} \rightleftharpoons \underset{\text{pale brown}}{N_2O_{4(g)}}$$

- if we put the sealed tube in a cold water bath , the mixture becomes paler . why ?

because the forward reaction is exothermic , enhanced by cooling.

2) Pressure :-

- As the pressure increases, equilibrium shifts to side with less pressure and has less gas mole.

- As the pressure decreases, equilibrium shifts to side with more pressure and has more gas mole.

* $\underbrace{3H_2(g) + N_2(g)}_{4\text{ mole}} \rightleftharpoons \underbrace{2NH_3(g)}_{2\text{ mole}}$

↑ pressure : shifts forward to side with less gas mole, ↑% NH_3
↓ pressure : shifts backward to side with more gas mole.

* $\underbrace{CaCO_3(s)}_{0\text{ mole}} \rightleftharpoons \underbrace{CaO(s) + CO_2(g)}_{1\text{ mole}}$

why? CaO & $CaCO_3$ are solid

↑ pressure : shifts backward to the side with less gas mole.
↓ pressure : shifts forward to side with more gas mole.

* $\underbrace{H_2(g) + I_2(g)}_{2\text{ mole}} \rightleftharpoons \underbrace{2HI(g)}_{2\text{ mole}}$

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

* $\underset{\text{purple}}{H_2(g) + I_2(g)} \rightleftharpoons \underset{\text{colourless}}{2HI(g)}$

- the equilibrium isn't affected by increasing the pressure. Why by increasing the pressure the mixture becomes more purple?

The I_2 moles become closer to each other so it seems more purple.

* Sealed tube contains mixture of NO_2

3) Concentration :-

↑ Reactant } shifts forward
↓ Product

↓ Reactant } shifts backward
↑ Product

$A+B \rightleftharpoons C+D$

* ↑(A) ↑(D) shifts
↓(B) ↑(C) forward

* ↑(A) ↓(D) shifts
↓(B) ↓(C) backward

* ↑(A) ↓(D) shifts
↑(B) ↑(C) backwards

- Indicator :- example : mythyl orange

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

red yellow

↑H^+ shifts backward

* add HCl : proton donor more HIn , more red
less In^- , less yellow

↓H^+ shift forward

* add NaOH : proton accepter more In^- , more yellow
less HIn , less red

Question :-

the reversible reaction below is at equilibrium:-

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

orange green

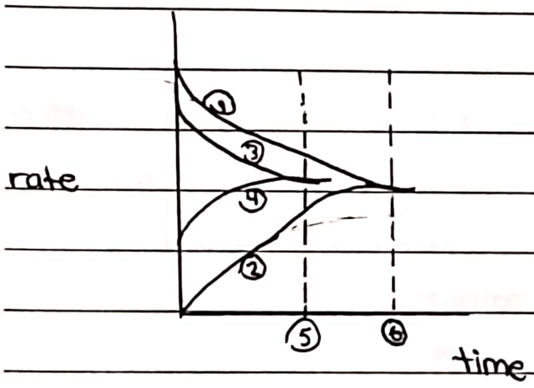
* Explain why adding HCl to the reaction mixture, the colour of the mixture becomes green?

HCl is an acid (proton donor)

more H^+, shifts forward, more Cr^{3+}, more green

less Cr_2O_7, less orange.

* Catalyst : doesn't affect position of equilibrium, since it speeds up the rate of forward & backward reactions.



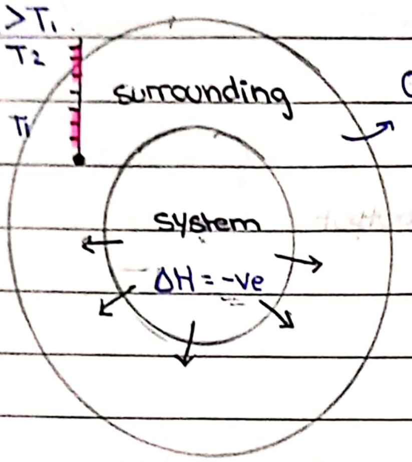
1) rate of forward without catalyst.
2) rate of backward without catalyst.
3) rate of forward with catalyst
4) rate of backward with catalyst
5) time taken to reach equilibrium with catalyst.
6) time taken to reach equilibrium without catalyst.

Energetics

* Exothermic :-

Reaction that gives out (release) energy to surrounding.

$T_2 > T_1$.



$Q = mc\Delta T$

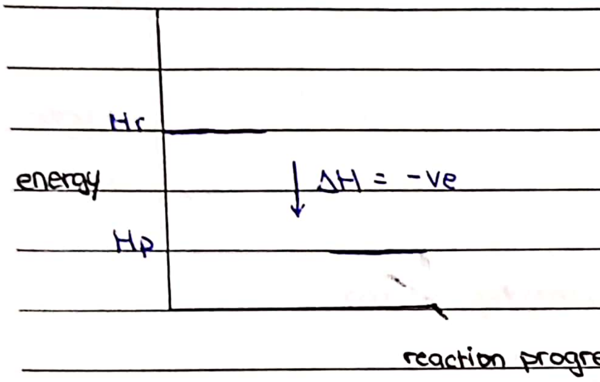
Enthalpy : Heat contents

"stored energy"

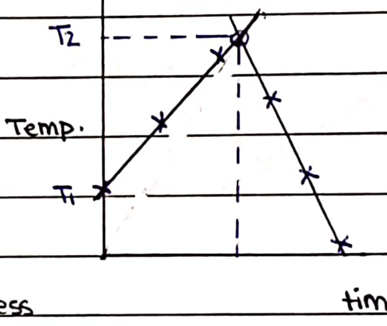
Hr → Hp

enthalpy of reactants → enthalpy of products

for system (energy level diagram)



for surrounding (temp. diagram)



* why did the temp. decrease with time?
reaction is over so it turns back to room temp.

mass (g) → m

change in temp. (°C) → ΔT

energy transfer (J) ← $Q = mc\Delta T$

c → specific heat capacity

* how do we know if the reaction is more exothermic?

more ΔT which cause an increase in Q (energy transfer).

* examples on exothermic reactions :-

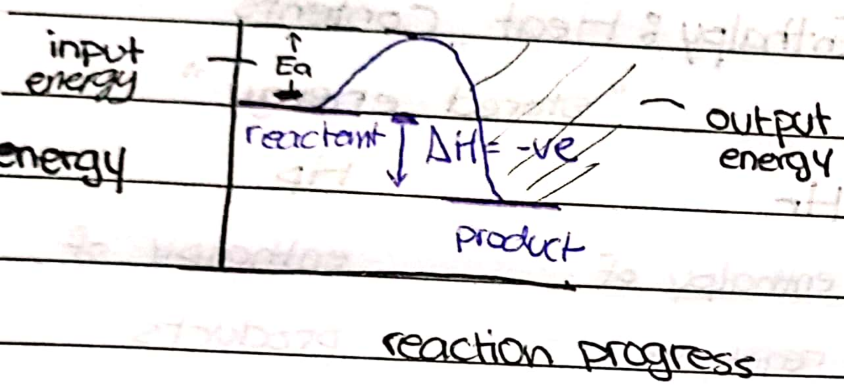
1) Freezing, Condensation.
2) Respiration
3) Combustion
4) Neutralization
5) Voltic cell (battery)
6) Displacement
7) Building up bonds.

* How to express exothermic reaction :-

1) Reactants ⟶ Products $\Delta H = -ve$

2) Reactants ⟶ Products + ENERGY

3) Profile diagram

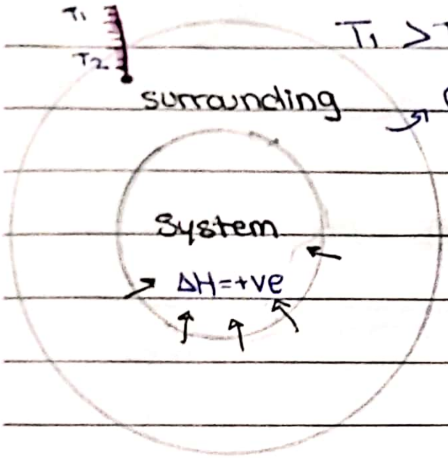


input < output

* Endothermic :-

Reaction that absorbs (takes in) energy from surrounding.

$T_1 > T_2$



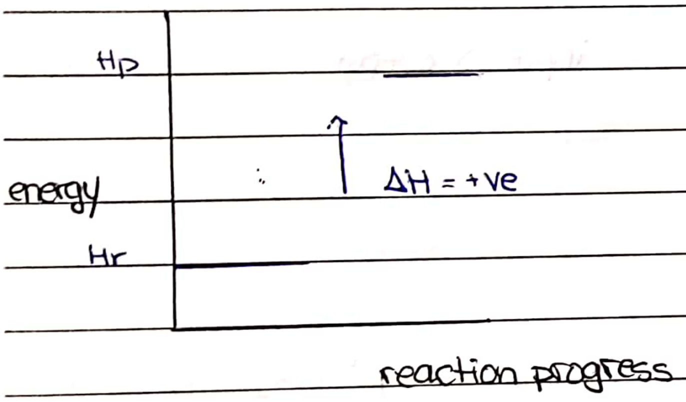
$Q = mc\Delta T$

Enthalpy & Heat contents

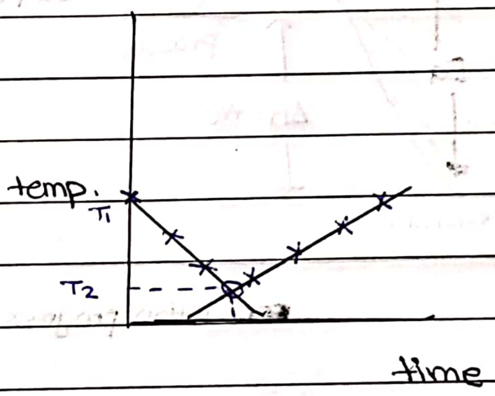
"stored energy"

Hr	Hp
enthalpy of reactants	enthalpy of products.

for system (energy level diagram)



for surrounding (temp. diagram)



$$\underset{\text{energy transfer (J)}}{Q} = \overset{\text{mass (g)}}{m}\underset{\text{specific heat capacity}}{c}\,\Delta T \rightarrow \text{change in temp. (°C)}$$

* how do we know that the reaction is more endothermic?

more ΔT which causes an increase in Q (energy transfer)

* examples on endothermic reactions :-

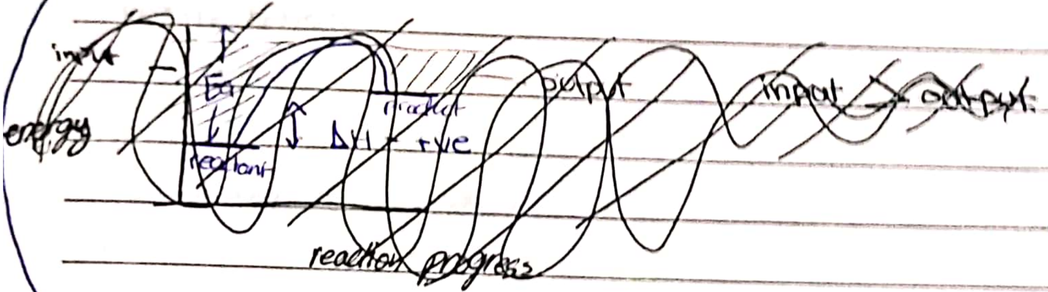
1) boiling, melting
2) photosynthesis
3) thermal decomposition
4) electrolysis
5) Photographic films
6) dissolving ammonium salts.
7) breaking down bonds.

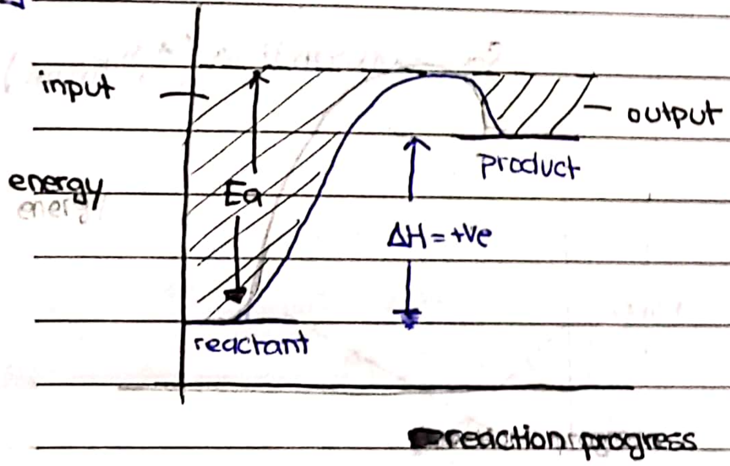
* How to express endothermic reaction :-

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

2) Reactants + ENERGY → Products

3) Profile diagram.





input > output

Measuring ΔH reactions

theoritical
- using bond energy

experimental
combustion displacement neutralization

* ΔH reactions using bond energy :- (theoritical)

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

bond	bond energy kj/mol
H - H	436

- $H-H + 436\,kj/mol \rightarrow H_{(g)} + H_{(g)}$ (endo.)
  $H_{(g)} + H_{(g)} \longrightarrow H-H + 436\,kj/mol$ (exo.)

* why gaseous ?

If it is in the solid state it will need energy to overcome the forces of attraction between particles to become liquid → gas, the bond is then broken using bond energy.

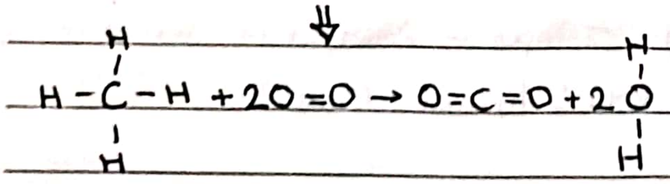
* ΔH reaction = Σ input energy − Σ output energy
to break down bonds in reactant — to break down bonds in product

ΔH → $\Delta H = +ve$ endo. — input > output
ΔH → $\Delta H = -ve$ exo. — input < output

to use equation :
1) balance equation
2) covalent structure
3) bond energy

- examples :

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



$$H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H + 2\,O=O \rightarrow O=C=O + 2\,\overset{\displaystyle H}{\underset{\displaystyle H}{O}}$$

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

bond broken		bond build	
4 * C-H →	4 * 413	2 * C=O	2 * 799
2 * O=O	2 * 495 +	4 * O-H	4 * 463 +
	2642 kj		3450 kj

ΔH = Σ input - Σ output

= 2642 - 3450   ΔH = $\boxed{-}$808 kj/mol

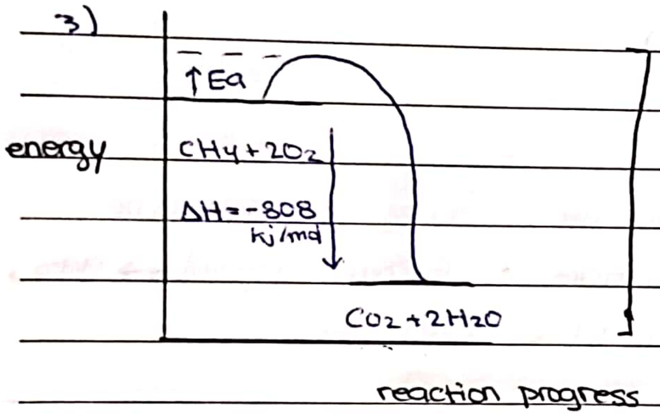
exo.

- because it's exo. it can be written as:

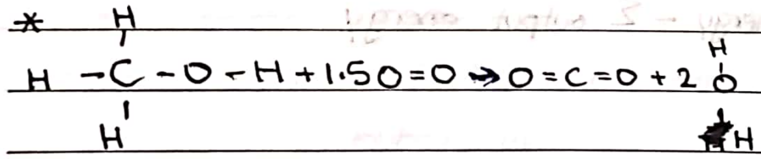
1) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 808$ kj/mol

2) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ ΔH = -808 kj/mol

3)



profile diagram.



$$* \quad H-\overset{H}{\underset{H}{C}}-O-H + 1.5\,O=O \rightarrow O=C=O + 2\,\overset{H}{\underset{H}{O}}$$

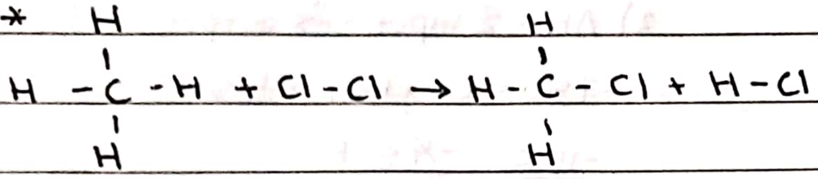
bond	bond energy
C-H	413
O=O	495
C=O	799
O-H	463
+ C-O	358

bond broken		bond build	
3 * C-H	3 * 413	2 * C=O	2 * 799
1 * C-O	1 * 358 +	4 * O-H	4 * 463 +
1 * O-H	1 * 463		3450 kj
1.5 * O=O	1.5 * 495		
	2802.5 kj		

* ΔH = Σ input - Σ output

ΔH = -647.5 kj/mol (exo.)

*



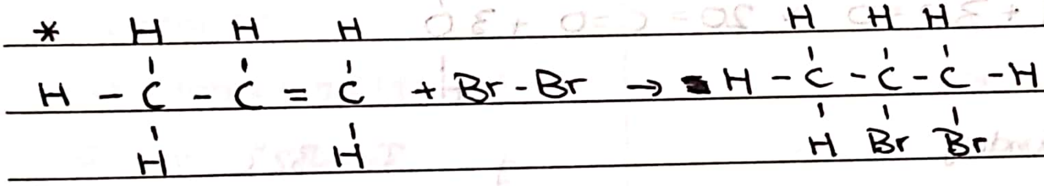
$H-\overset{H}{\underset{H}{C}}-H + Cl-Cl \rightarrow H-\overset{H}{\underset{H}{C}}-Cl + H-Cl$

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

bond broken		bond build	
1 * C-H	1 * 413 +	1 * C-Cl	1 * 328 +
1 * Cl-Cl	1 * 242	1 * H-Cl	1 * 431
	655 kj		759 kj

ΔH = Σ input - Σ output   655 - 759 = -104 kj/mol (exo)

*



$H-\overset{H}{\underset{H}{C}}-\overset{H}{C}=\overset{H}{\underset{H}{C}} + Br-Br \rightarrow H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{Br}{C}}-\overset{H}{\underset{Br}{C}}-H$

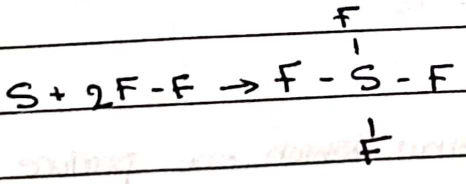
bond	bond energy
C-H	413
C-C	348
C=C	614
Br-Br	193
C-Br	276

bond broken		bond build	
1 * Br-Br	1 * 193 +	2 * C-Br	2 * 276 +
1 * C=C	1 * 614	1 * C-C	1 * 348
	807 kj		900 kj

ΔH = Σ input - Σ output

= 807 - 900   ΔH = -93 kj/mol (exo.

* When sulfur react with florine the reaction gives 780 kj/mol
exo ΔH



$S + 2F-F \rightarrow F-\overset{F}{\underset{F}{S}}-F$

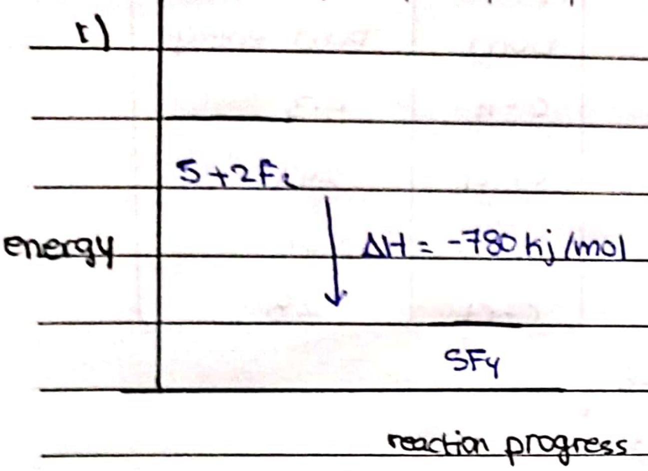
if the bond of F-F is 160 kj/mol,

1) draw an energy level diagram

2) Find bond energy of S-F

(QxΔT)

1)



2) $\Delta H = \Sigma \text{input} - \Sigma \text{output}$

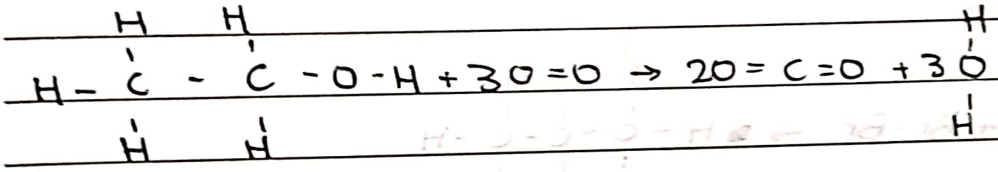
$-780 = (2 * 160) - 4\,S\text{-}F$

$\frac{-1100}{-4} = \frac{-\cancel{4}\,S\text{-}F}{-\cancel{4}}$

$\underline{\underline{S\text{-}F = 275 \text{ kj/mol}}}$

* experimental :-

* combustion :



$H{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}O{-}H + 3\,O{=}O \rightarrow 2\,O{=}C{=}O + 3\,\overset{H}{\underset{H}{O}}$

for surrounding

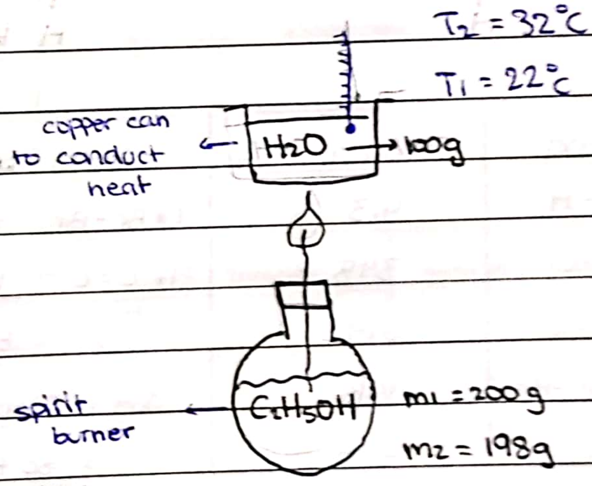
energy transfer ← $Q = mc\Delta T$

$= 100 * 4.2 * 10$

$= 4200\,J = 4.2\,kJ$

↓

burn 2g ethanol



$4200\,J \xrightarrow{\text{burn}} 2g$ ethanol

$\Delta H\ ? \xrightarrow{} 46g$ ethanol (1 mole)] info will be given in question (Mr of ethanol)

$\frac{4200 * 46}{2} = \Delta H$

$= 96.6$

$\boxed{\Delta H = -96.6 \text{ kj/mol}}$

* 2 fuels A&B, plan an exp. to show which one produce 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 & record the final mass and the final temp. of water. repeat the exp. using ~~the~~ fuel B. The fuel which causes more temp rise per gram of fuel produces more energy.

COLLEGE

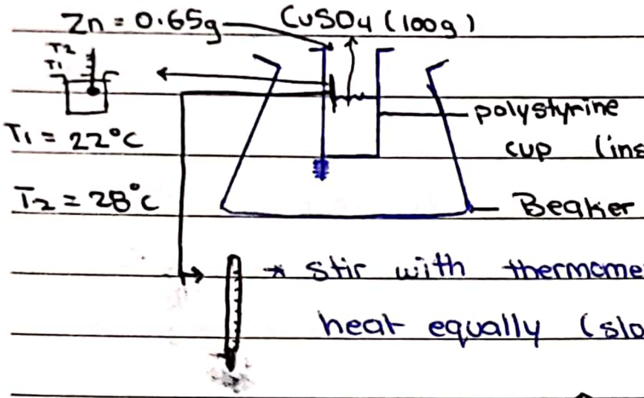
* displacement :

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

* why do we put a beaker below the polystyrine cup?

→ more stable

→ more insulation



* stir with thermometer to distribute heat equally (slowly to avoid overheating)

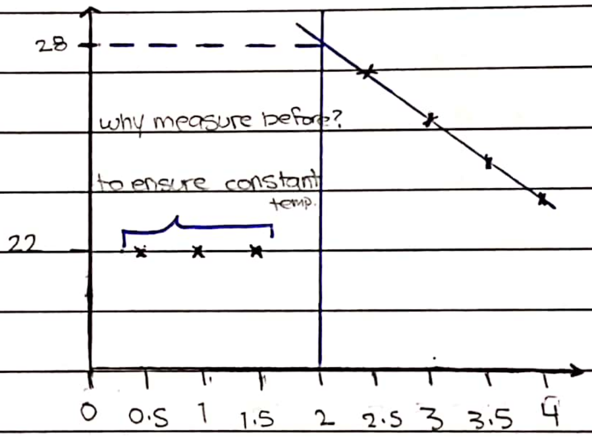
$Q = mc\Delta T$

$Q = 100 * 4.2 * (28-22)$

$Q = 100 * 4.2 * 6$

$Q = 2520\ j$

$Q = 2.52\ kj$

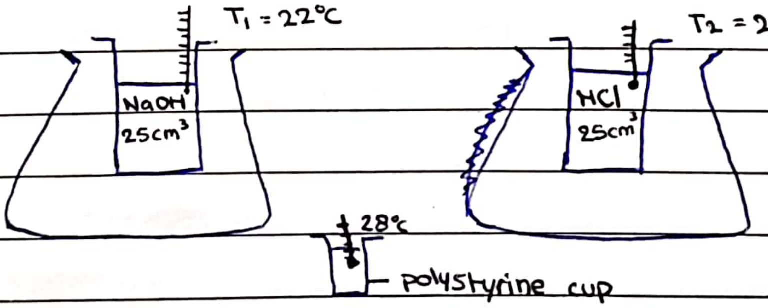


why measure before?

to ensure constant temp.

* neutralization :

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



specific heat capacity = 4.2 J/m

density = $1 g/cm^3$

$Q = mc\Delta T$

$\Delta T = 28 - 22 = 6°C$

$m = ?$

$C = 4.2$

$\rho = m/v$

$1 = m/25$

$m = 25$

$M_{Total} = 25 + 25 = 50$

$50 * 4.2 * 6$

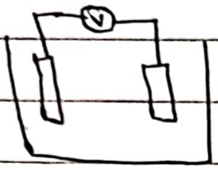
$= \underline{1260 J}$

Alternative Resources of energy

voltic cell

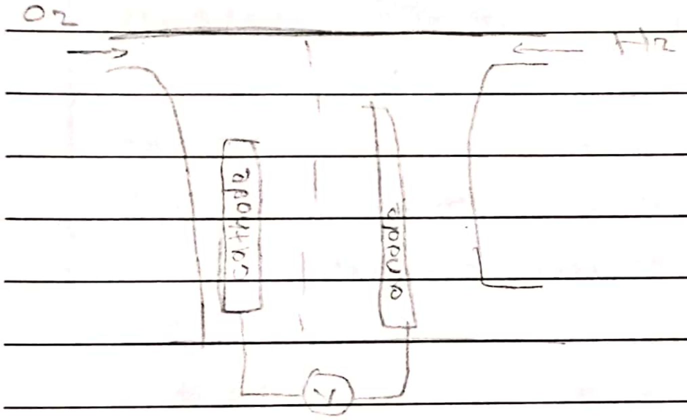
hydrogen fuel cell

Uranium



* Hydrogen fuel cell :

$2H_2^{0} + O_2^{0} \rightarrow 2\overset{1+}{H}_2\overset{2-}{O}$



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

reduction

$H_2 \rightarrow 2H^{+} + 2\bar{e}$

oxidation.

test for water

- physical test : B.p = 100
- Chemical test

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

white blue

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

blue pink

- 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

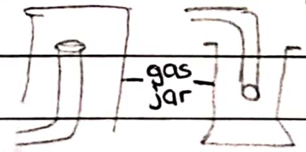
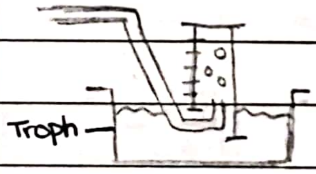
- Dealing with gases
  - Dry
  - Collect
- Industry of
  - NH_3 Haber process
  - H_2SO_4 contact process
  - CO_3^{2-} carbonate cycle
- Extraction of metals
  - Al
  - Fe Zn
  - Cu

* Dealing with gases :-

reaction → wet gas → dry → collect

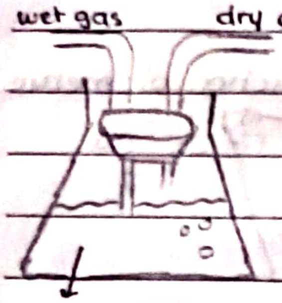
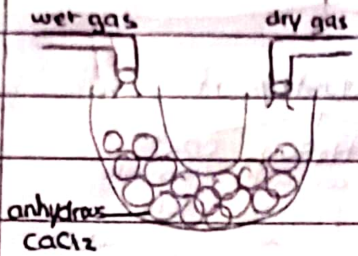
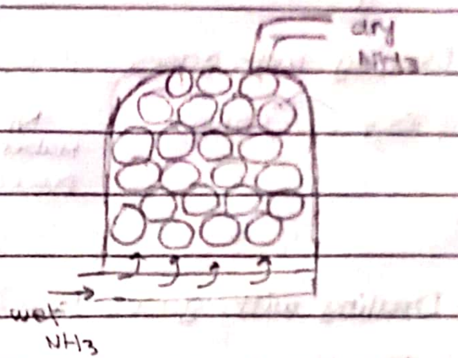
* collect gas :

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

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

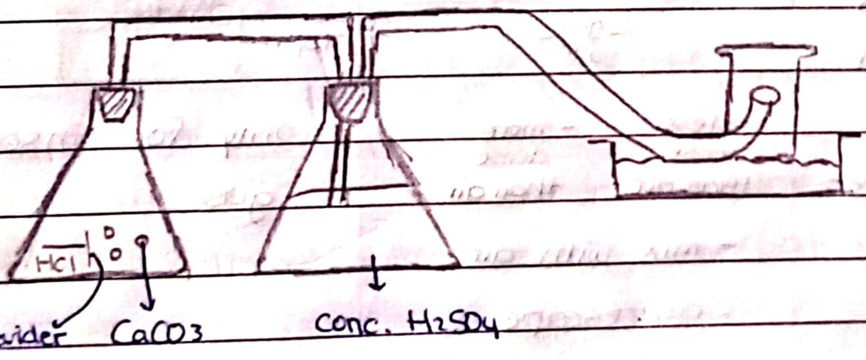
* Drying gas :

- 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 neutralises 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 and measure the volume of CO_2 gas from (O_2 insoluble)

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



uses : smelling salts
cleaning detergents
prepare fertilizers.

* Industry of Ammonia : (Haber process)

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

How to obtain :

1) Nitrogen : fractional distilation of liquid air
↓ different b.p ↓ cooling under high pressure.

2) Hydrogen :- cracking of Alkanes (organic)

- $CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + 2H_{2(g)}$

Essential conditions for Haber process :

1) Temp. 400 °C - 450 °C

less than 400°C

adv :

- higher yeild of NH_3
- shift forward to exo side.

dis :

- Slower rate, particles lose kinetic energy so less effective collisions per unit time.

more than 450°C

adv :

- faster rate

dis :

- less yeild
- shift backward to endo side.

- Add excess H_2 & N_2 return back to converter ~~to enhance forward reaction~~
- Remove NH_3 immediately by cooling

2) Pressure 200 atm

adv :

- more yeild of NH_3
Shift forward to side with fewer gas mole.
- faster rate.

dis :

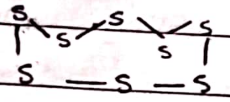
- risk of explosion
- expensive.

* Industry of H_2SO_4 :- (contact process)

Stages $S \underset{+O_2}{\xrightarrow{(1)}} SO_2 \underset{+O_2}{\overset{(2)}{\rightleftharpoons}} SO_3 \xrightarrow{(3)} H_2S_2O_7 \xrightarrow{(4)} H_2SO_4$

Roasting contact process Oleum

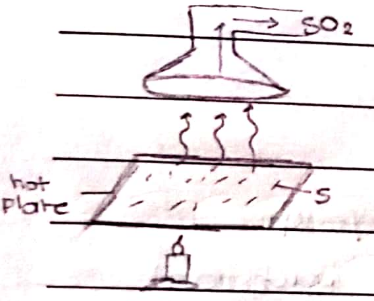
Properties of sulfur :
- group (VI)
- valency (2)
- Yellow solid
- S_8



Uses :
- matches
- medicine
- rubber

Ore :
Zinc blende
ZnS
from fossil fuels

* Roasting :



* How to produce H_2SO_4 :-

1) $SO_3 \xrightarrow{H_2O} 2H_2SO_4$ X
- highly exothermic
- produce low yeild of H_2SO_4

$SO_3 \xrightarrow{conc.\ H_2SO_4} H_2S_2O_7 \xrightarrow{H_2O} H_2SO_4$

2) Electrolysis of copper sulfate

* Contact Process :

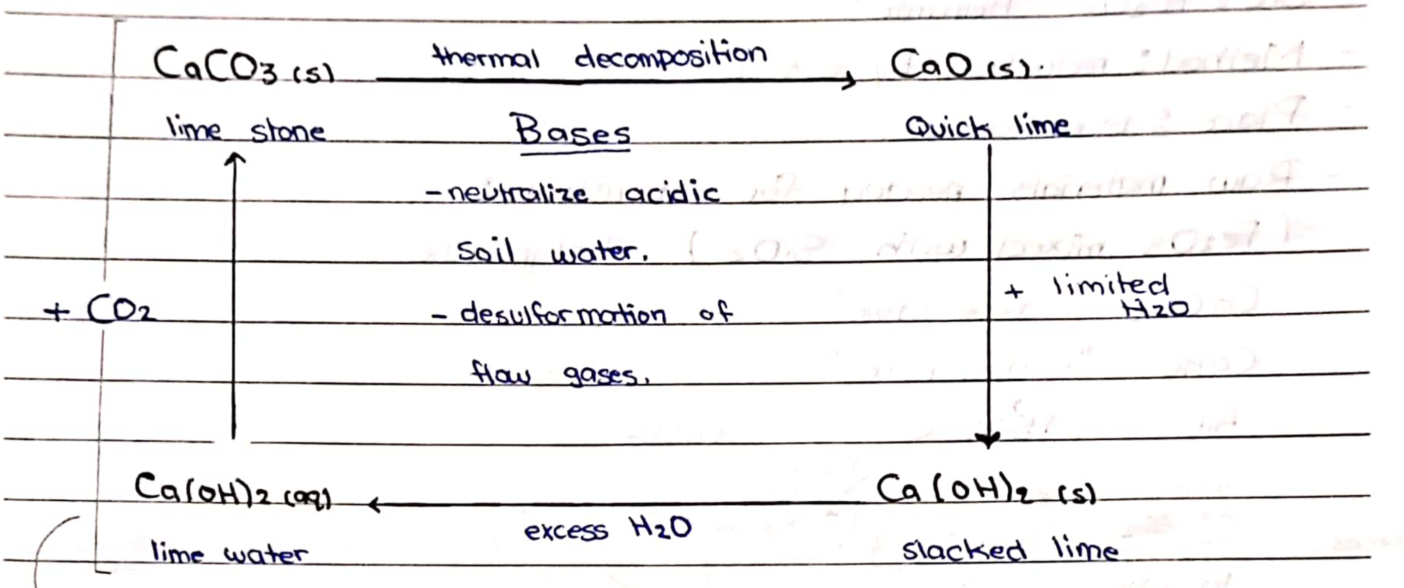
$2SO_{2(g)} + O_{2(g)} \underset{endo}{\overset{V_2O_5\ exo}{\rightleftharpoons}} 2SO_{3(g)}$ $\Delta H = -ve$ exo

Essential Conditions :-
1) Temp. 400–450°c gives max yeild of SO_3
2) Pressure 2 atm → high pressure favour the forward reaction (fewer gas mole)
3) Catalyst V_2O_5 vanadium (V) oxide.

Uses of sulfuric acid :-
- Fertilizer
- As a drying agent

Uses of SO_2 :
* Causes acid rain
* kills bacteria (sterilization)
* paper industry.

* Carbonate Cycle :-



Chemical equation

$Ca(OH)_{2(aq)} + CO_2 \longrightarrow CaCO_{3(s)} + H_2O$

- $CaCO_3$ uses :-

* building

* extraction of iron

- CaO uses :-

* dry NH_3

* Extraction of Metals :-

K
Na
Li
Ca
Mg
Boxite / Al_2O_3 Al] Electrolysis / molten
C, CO
Zinc blend ZnS ← Zn] reduction by C & CO
Hematite Fe_2O_3 ← Fe] "blast furnace"
Pb
H
CuS ← Cu] reduction by H_2
Ag
Au
Pt

*Extraction of Iron :-

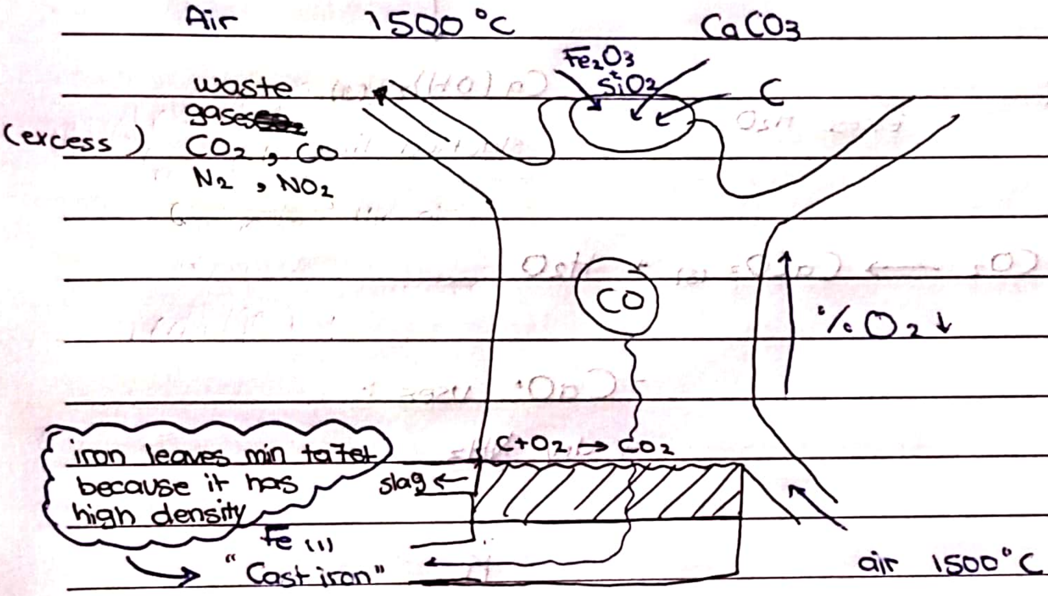
- Ore : Fe_2O_3 "Hematite"
- Method : reduction by C & CO
- Place : blast furnace.
- Raw materials needed for extraction :

(Fe_2O_3 mixed with SiO_2) acid impurities

$CaCO_3$ "lime stone"

Coke "Carbon Pure"

Air 1500 °C



CO is important so reduction would take place.

* complete combustion:

$C + O_2 \rightarrow CO_2$ produce energy

* incomplete combustion:

$2C + O_2 \rightarrow 2CO$

* $CO_2 + C \rightarrow 2CO$

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

$$\left.\begin{array}{l} CaCO_3 \xrightarrow{heat} CaO + CO_2 \\ \underset{base}{CaO} + \underset{acid}{SiO_2} \rightarrow \underset{\text{calcium silicate slag}}{CaSiO_3} \end{array}\right\} \rightarrow$$

CaCO₃ won't have time to neutralize the SiO_2 because it decomposes with heat, so it creates a more stable base to neutralize the SiO_2.

* Slag : forms a protective layer that prevent the reaction of $Fe + O_2$

use : mix with bitumene to make roads.

* Steal Making "oxygen base process"

cast iron

Fe

$CO_2 \uparrow \leftarrow \boxed{C}$ hot O_2 to reduce C.

$SO_2 \uparrow \leftarrow S$

impurities

$SiO_{2(s)} \leftarrow Si$ CaO ⇒ neutralize acids (SiO_2 & P_2O_5) to produce steel

$P_2O_{5(s)} \leftarrow P$

Steel (% of C in steel)

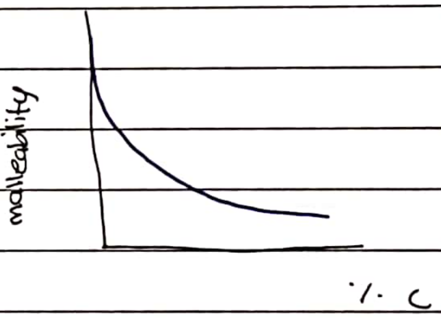
mild steel 0.03 % C (used in car bodies)

medium steel 0.3 % C (used in railways)

stainless steel 3-5 % C (used in cutters/watches)

As the % of C increases, the malliability decreases.

(% C 1/α malliability)



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

Brass Cu, Zn

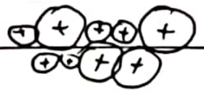
Bronze Cu, Sn

Steel Fe, C, Ni, Cr

Alloy are harder because it has different sizes of particles so harder to slide between each other.

Metal Cu

Brass Cu, Zn



* Extraction of Zinc :

Ore : Zinc blende ZnS

method : reduction ~~of~~ by C & CO

place : blast furnase.

Step 1:- Roasting with hot oxygen.

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