REDOX (topic 9 & 19)
REVISION NOTES

9.1 Oxidation & reduction

- oxidation is the **loss** of one (or more) electrons from a substance ⟶ **OIL**
- reduction is the **gain** of one (or more) electrons from a substance ⟶ **RIG**

- oxidation states can be used to determine what has been REDUCED / OXIDIZED

  ⚠ +2 NOT 2+   e.g. $NaCl$ is Na^+ and Cl^- ∴ Na has +1, Cl has -1

— RULES
  - sum of ox no = charge on molecule / atom   e.g. SO_4^{2-} has 2- as sum
  - compound ox no = 0
  - group 1 = +1 , group 2 = +2 , group 3 = +3

— EXCEPTIONS
  - O_2 always has -2 ox no. unless H_2O_2 (where it has 1-)
  - H when w/ metals (hydrides) has -1 oxidation number

- AGENTS...
  - reduced = oxidizing agent   (can accept electrons)
  - oxidized = reducing agent   (can gain electrons)

  e.g. $2FeCl_3 + 3Zn \longrightarrow 3ZnCl_2 + 2Fe \ (s)$

  OX numb.   3+ $\underset{-3}{-1}$ 0   loss $\underset{-2}{-1}$ 2+   0   ∴ Cl is a spectator ion
                                NO loss   gain

                                                  electrons must balance

  $Fe^{3+} + 3e^- \longrightarrow Fe$ ∴ $2Fe^{3+} + 6e^- \longrightarrow 2Fe$
  $Zn \longrightarrow Zn^{2+} + 2e^-$ ∴ $3Zn \longrightarrow 3Zn^{2+} + 6e^-$

  final ionic equation : $2Fe^{3+}_{(aq)} + 3Zn_{(s)} \longrightarrow 3Zn^{2+}_{(aq)} + 2Fe_{(s)}$

  e.g. $Cr^{3+} \longrightarrow Cr_2O_7^{2-}$ ∴ $2Cr^{3+} \longrightarrow 2Cr^{1+} + 6e^-$
          (3+)   (+6)(-2)
                   (+12)(-14)

  use H_2O to balance Os / Crs $2Cr^{3+} + 7H_2O \longrightarrow Cr_2O_7$
  balance H^+ $2Cr^{3+} + 7H_2O \longrightarrow Cr_2O_7 + 14H^+$
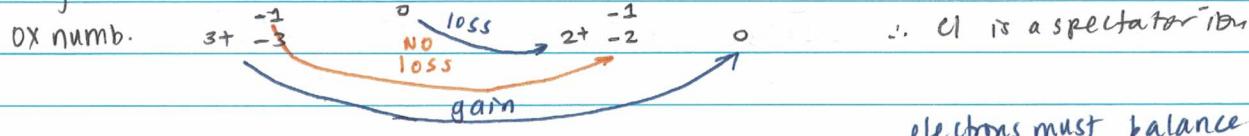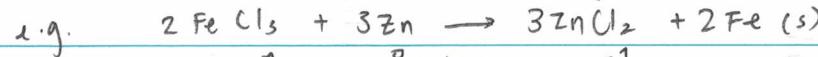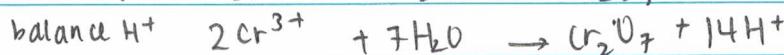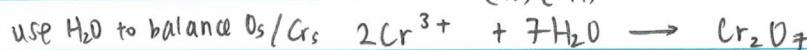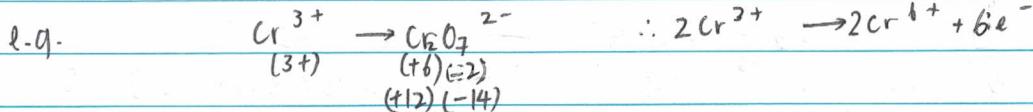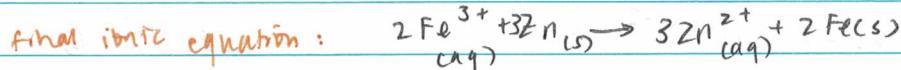
  when combining half equations...
  - e^- SAME on both half eqn (x common factor)
  - cancel out what is left by combining.
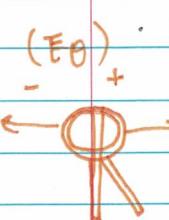
- (Eθ) activity series is ranked by how EASILY a metal gets oxidized (loses electrons)
  - the more reactive, the more likely it will be oxidized. (acts as reducing agent!)
  - e.g. zinc will react with copper ions to form zinc ions
    [best reducing agent, gets oxidized = caesium = ↓ electroneg. ↓ attraction]
    [best oxidizing agent, gets reduced = fluorine = ↑ electron- ↑ attraction]

(from reactivity series
  more +ve = reduced ⟶ . more -ve oxidized ⟸

$$\left[\begin{array}{l}\text{additional BOD notes:} \\ \uparrow °C \downarrow \text{solubility of } O_2 \therefore \uparrow BOD \\ \text{bacteria respire aerobically} \\ O_2 \text{ can dissolve from air}\end{array}\right]$$ $C_6H_{10}O_5 = $ organic matter

the winkler method (redox titrations)

can be used to measure **biochemical oxygen demand (BOD)** $\nearrow$ anaerobic bacteria

$\hookrightarrow$ which is amt of O_2 required to decompose organic matter over time!

$\hookrightarrow$ HIGH BOD = low O_2 content $\therefore$ high pollution $\nwarrow$ important

STEP 1. manganese hydroxide + oxygen $(Mn(OH)_2$ is made frm $MnSO_4 +$ alkali)

$$2Mn(OH)_2 {}_{(s)} + O_2 {}_{(aq)} \rightarrow 2MnO_2 {}_{(s)} + 2H_2O {}_{(l)}$$

STEP 2. Add H_2SO_4 to form $Mn(SO_4)_2$ and then react this with KI

$$Mn(SO_4)_2 + 2I^-{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + I_2 + 2SO_4{}^{2-}{}_{(aq)}$$

STEP 3. Determine I_2 concentration using sodium thiosulphate (titration)

$$2S_2O_3{}^{2-} + I_2 \rightarrow S_4O_6{}^{2-} + 2I^- \qquad \text{(starch, blue} \rightarrow \text{colourless)}$$

half equations

$$2Mn^{2+} + O_2 + 4OH^- \rightarrow 2MnO_2 + 2H_2O$$
$$\tfrac{1}{2}x \qquad\qquad\qquad x$$

$$MnO_2 + 2I^- + 4H^+ \rightarrow I_2 + 2H_2O + Mn^{2+}$$
$$\tfrac{x}{2} \qquad\qquad\qquad \tfrac{x}{2}$$

$$2S_2O_3{}^{2-} + I_2 \rightarrow 2I^- + S_4O_6{}^{2-}$$
$$x$$
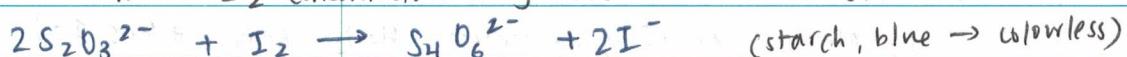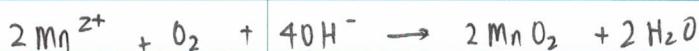
❗ $S_2O_3 : O_2$ is $2:1$

❗ $I_2 : S_2O_3$ $2 : 4$

other titrations

$$2Cu^{2+} + 4I^- \rightarrow 2CuI + I_2$$

③ moles $1.3\times10^{-2} \longleftarrow - - - - - - 6.6\times10^{-3}$ ②

working back!

$$I_2 + 2S_2O_3{}^- \rightarrow S_4O_6{}^{2-} + 2I^-$$

① moles 6.1×10^{-3} $\quad 0.013$ [be sure to ADD volumes!]

❗ titration conc $\frac{old}{new} \times$ moles

harder half equations

○ $2H_2O_2 \rightarrow 2H_2O + O_2$ HOW TO FORM ½ EQUATION?

$\downarrow\ \downarrow \qquad \downarrow\ \downarrow\ \downarrow$
$+1\ -1 \qquad +1\ -2\ \ 0$

$\hookrightarrow$ must both have H_2O_2 on RHS so they add up to $2H_2O_2$

$\begin{cases}\text{2 oxygen} \rightarrow -2 \\ \text{you must have } 2e^-\end{cases}$

① $H_2O_2 + 2e^- + 2H^+ \rightarrow 2H_2O$
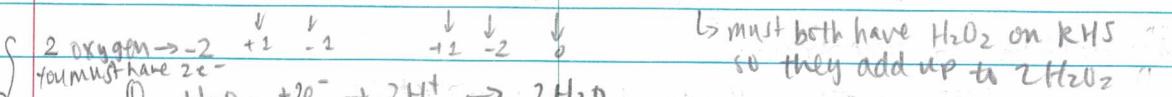
must have H^+ to balance out H_2S

② $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$

$\nearrow SO_4{}^{2-} :$ spectator $\therefore$ ignore

○ $Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$

individually: $Pb^0 \rightarrow Pb^{2+}$ and $PbO_2 \rightarrow Pb^{2+}$

$\hookrightarrow Pb \rightarrow Pb^{2+} + 2e^-$ $\hookrightarrow PbO_2 + 2e^- \rightarrow Pb$

balance H_2O, H^+: $PbO_2 + 2e^- * \rightarrow Pb + 2H_2O$

$PbO_2 + 2e^- + 4H^+ \rightarrow Pb + 2H_2O$
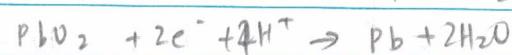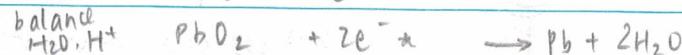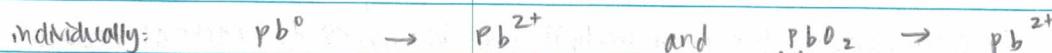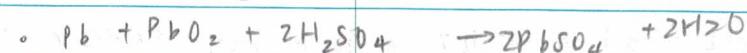
9.2 Electrochemical cells

voltaic cells

- chemical energy $\Rightarrow$ electrical energy (spontaneous)

$\Delta G = -ve$
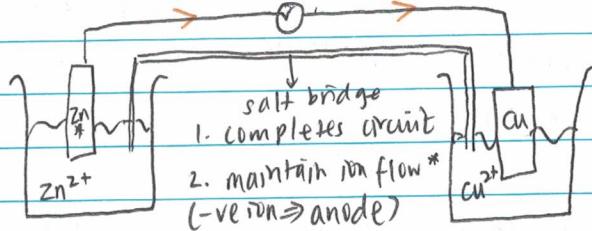
gets oxidized = LOSS

e.g. zinc has $E^\theta = -0.76$ $Cu\, E^\theta = +0.15$

voltage is the difference of 2 metals

zinc transfers electrons to copper $\rightarrow\rightarrow\rightarrow$

$$E^\theta = E^\theta_{\underline{RED}} - E^\theta_{\underline{OXT}}$$

salt bridge
1. completes circuit
2. maintain ion flow *
   (-ve ion $\Rightarrow$ anode)

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-_{(aq)}$$

$$Cu^{2+}_{(aq)} + 2e^-_{(aq)} \rightarrow Cu_{(s)}$$
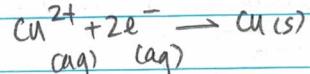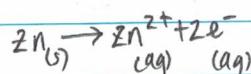
oxidation @ Anode (ANOX)   reduction @ cathode (REDCAT)

- anode is -ve   cathode is +ve

anode will go wonk

* maintain ion flow as Zn^{2+} is charged, needs NO_3^- to balance charges

* if reaction contains solid metal = must have electrode made out of solid.
  $\rightarrow$ if NOT = platinum! (inert) (conducts)

because electrons are "produced"

not theoretical value...
- 1 $mol\,dm^{-3}$?
- stp? 1 atm, 293 k.
- not completion $\rightleftharpoons$

$\dashrightarrow$ charge from n-ve to p+ve

electrolytic cells (MOLTEN SALT)

- electrical energy $\Rightarrow$ chemical energy (non-spont, ΔG +ve)

$$K^+ + e^- \rightarrow K_{(s)}$$
reduction @ CATHODE [REDCAT]

cathode
-ve!

anode
+ve!

$$2Cl^- \rightarrow Cl_2 + 2e^-$$
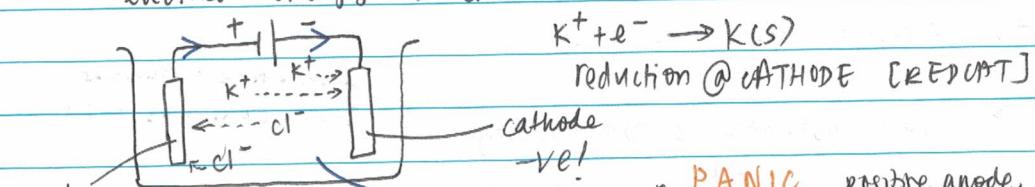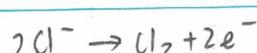
oxidation @ ANODE [ANOX]

electrolyte

* PANIC positive anode, negative cathode
* charge from +ve to -ve

application

- purify substances (reactive metals)
- plating (forms a layer of K on cathode)

factors affecting electrolysis: TIME, CURRENT, e^- (oxidation state)

$Q = I \times t \rightarrow$ (seconds)

charge   current (amps)

importance = produce $Na, Cl_2, NaOH$!

$$n(e^-) = \frac{Q\,(charge)}{96500}$$

moles of e^-

19.1 Standard electrode potential

- a voltaic cell has an EMF (electromotive force) ∴ electrons move -ve ⇒ positive.
  ↳ this is called the cell potential (E^θ) ⎰ reactivity

  dependant on... ⎱ concentration   at 298K and $1 mol dm^3$
                     temperature   it is a STANDARD.

  WILLING TO LOSE = -ve value (E^θ)

  WILLING TO GAIN = +ve value (E^θ)

- standard hydrogen potential   (hydrogen = standard reference electrode)

H_2 at 1atm

filter paper in KNO_3

copper

Cu^{2+}

$1 mol [H^+]$ └ platinum electrode serves as a catalyst
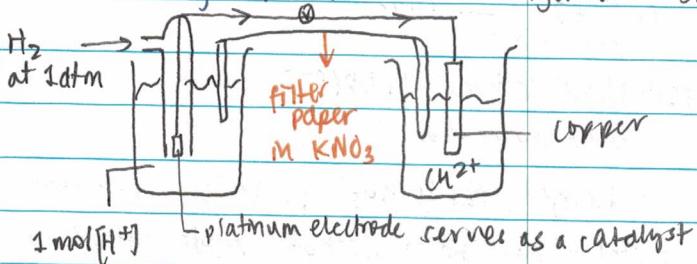
if it is more reactive than hydrogen ⇒ -ve value

less reactive ⇒ +ve value

- calculating cell potential
  - Zn and Cu. since Zn is more -ve ⇒ oxidized ⇒ loss

  $$Zn(s) / Zn^{2+} \;\|\; Cu^{2+} / Cu(s)$$

  $$Zn^{2+} + 2e^- \rightleftharpoons Zn \quad -0.76V \quad ox.$$

  $$Cu^{2+} + 2e^- \rightleftharpoons Cu \quad +.34V \quad red$$
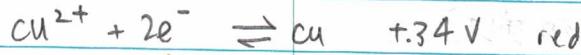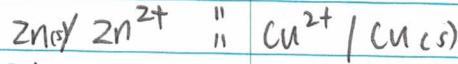
  $$E_{cell} = E^\theta_{red} - E^\theta_{ox} \quad (0.34 - (-0.76) = 1.10V)$$

  +ve E cell = spontaneous

  cell potential & spontaneity

  amps × time = coulombs
  96500 coulombs = 1 faraday

  $$\Delta G = -nFE^\theta$$

  moles of e^- faraday → cell potential

  ⇨ $\Delta G = -ve$ = spontaneous

electrolysis of water

+ | -

$H_2S O_4^{2-} OH^-$ CAT $H^+ H_2O$
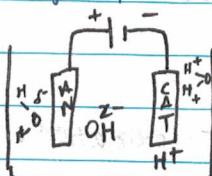
at cathode: reduction (RIG)

★ $2H_2O + 2e^- \rightarrow 2OH^- + H_2 \quad -0.83$

★ $2H^+ + 2e^- \rightarrow H_2 (g) \quad 0.00$ theoretically +ve = RED
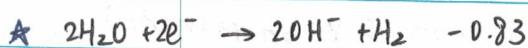
but! $[H^+] = 1\times10^{-7}$ ∴ v small
∴ doesn't usually happen

at anode:
★ $2H_2O \rightarrow O_2 + 4H^+ + 4e^- \quad +1.23$ | theoretical (but NOT)
★ $4OH^- \rightarrow O_2 + 2H_2O + 4e^- \quad +0.4$ | ☑ water is polar = occurs.

OVERALL : $2H_2O \rightarrow 2H_2 + O_2$ [hydrogen + oxygen produced]
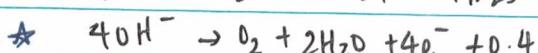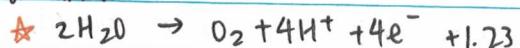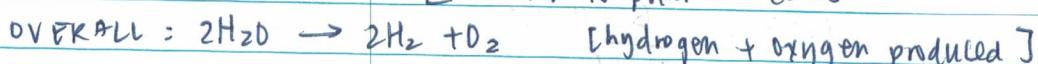
not molten! (molten = will def produce the solid/gas)

electrolysis of aqueous solutions

inert platinum

at the cathode ...

★ $2H_2O + 2e^- \rightarrow O_2 + 2OH^-$ $\quad$ -0.83 $\quad$ (more +ve)

$Na^+ + e^- \rightarrow Na(s)$ $\quad$ ~2.71
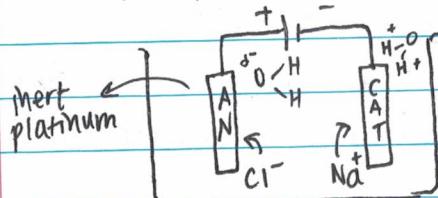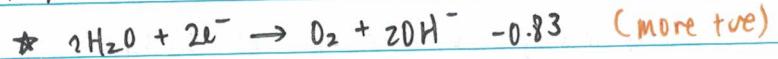
O_2 produced in both cases $\quad$ UNLESS $[Cl^-] \uparrow$ 25%

then... solution will have $Cl_2(g)$

and Na^+ ions and OH^- $\therefore$ NaOH!

at the anode

★ +1.23 $\quad H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $\quad$ (more -ve)

+1.36 $\quad 2Cl^- \rightarrow Cl_2 + 2e^-$ $\quad \uparrow [Cl_2/Br_2/I_2] \uparrow \rightarrow O_2$
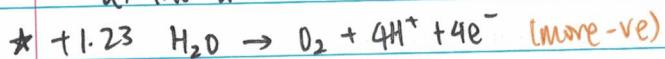
factors that determine products @ anode + cathode $\begin{cases} \text{trans metals} \\ \text{hydrogen} \end{cases}$

① position in electrochemical series compared to hydrogen

MORE +ve = REDUCED $\quad$ MORE ~ve = OXIDIZED

② concentration (e.g. $[Cl^-]$ in example above)

③ nature of electrode $\quad$ e.g. copper in $Cu^{2+} SO_4^{2-}$, anode itself will be OXIDIZED

$\therefore$ Cu forms in cathode, Cu lost in anode, CONC = SAME (in solution)

factors affecting quantity discharged (products)

① no. of electrons : _current_ doubles, product doubles ∵

② charge (Ox. No) : e.g. need $1e^-$ or $2e^-$? if you have Na^+ and Pb^{2+}

$\cdot$ Na^+ produced 2x rate of Pb

$\therefore$ cells are connected in series and SAME current = SAME moles of e^-

additional example — $CuSO_4$

→ platinum

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

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

$CuSO_4 (aq)$

solution loses blue
forms H_2SO_4

copper approach

at the cathode:
$Cu^{2+} + 2e^- \rightarrow Cu$
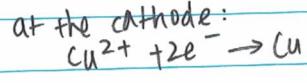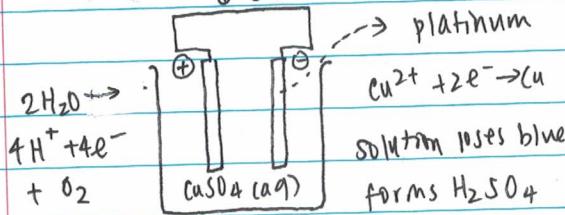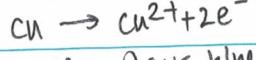
at the anode
$Cu \rightarrow Cu^{2+} + 2e^-$
solution stays blue!

---copper

· electroplating

— electrolytic coating of a metal $\quad$ e.g. copper $\quad\quad\quad\quad\rightarrow$ RHS : spoon!

THE CATHODE (-ve) is made from the METAL TO-BE PLATED

THE ANODE (+ve) is made from the IMPURE ONE

$\rightarrow$ as oxidation happens, producing metal _ION_ $\quad (Cu \rightarrow Cu^{2+} + 2e^-)$

$\rightarrow$ which then deposits itself into the cathode $(Cu^{2+} + 2e^- \rightarrow Cu$

— prevent from rusting = more easily/readily oxidized in comparison to another metal!

be careful w/ agents!

past paper questions

- "what happens to $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$"? ... or oxidation number $+6 \rightarrow +3$
- "describe standard conditions" · 1 moldm^{-3} of SOLUTION & 298K
- explain why the electrolysis of $ZnSO_4$ is not used for the electroplating of zinc
   - more +ve ions in solution (Zn^{2+} and H^+) ∴ [H_2] will be produced, reduced readily.
- define "standard electrode potential".
   · potential of a <u>reduced</u> <u>half-equation</u> under standard conditions
   · measured compared to a S. Hydrogen Electrode.
- displacement reactions . bromine when displaced = ORANGE & · iodine = BROWN
- how is electricity conducted in an electrolytic cell? ① electron movement in metal wires
② ion movement in electrolyte (molten ionic liquid)

explain the mechanism of reactivity/displacement.
a strong reducing agent has the greatest ability to donate electrons and displace a
weaker reducing agent.

oxidation number of N in $NH_4Cl \rightarrow NH_4^+$ and Cl^- ∴ $H_4 = 4+$, $N = -3$
   if Mg > reactive cu, what is strongest ox. agent? o 😊   oxid agent = gets reduced
      ∴ gets reduced = GAIN electrons + won't be Mg. Cu^{2+} (NOT cu, cu → gain e$^-$)
3 things that impact electrolysis
   ① magnitude of current ② time of electrolysis ③ oxidation/charge on ion
define "biochemical oxygen demand"
      amt of dissolved oxygen needed by AEROBIC organisms to BREAK DOWN
      organic material in water over time.   ↗ CU! RMB   ↗ Cl_2!
factors affecting aqeous electrolysis ① electrode potential ② electrode ③ concentration

voltaic cells

$E^\theta = E_{cath} - E_{anode}$ e.g. $Cu^{2+} +0.34$ $Mg -2.37$

$E^\theta = RED - OX$
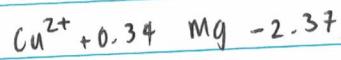
Cu gets reduced / Mg is oxidized

$(+ve - -ve)$

$\therefore$ $0.34 - (-2.37) = +2.71$ is voltage

experimental voltages vs literature values
- reaction does not go to completion
- standard (θ) temperature

ANODE $Mg^{2+}_{(aq)} + 2e^- \rightleftharpoons Mg_{(s)} -2.37$

cathode $Cu^{2+} + 2e^- \rightleftharpoons Cu_{(s)}$ $+0.34$

when you ADD water to anode...
the conc of Mg^{2+} (aq) decreases
$\therefore$ equilibrium moves to left, E^θ becomes more $-ve$ e.g. 2.40

THEREFORE   e.cell value $= 0.34 - (-2.40)$

$= 2.74$ (e.g. no water $+2.71$)

when you ADD water to cathode

$\downarrow [Cu^{2+}]$ $\therefore$ $\uparrow [e^-]$ (equilibrium moves to left)

$+0.34$ $\Rightarrow +0.3$ $\therefore$ $0.30 - (-2.37)$

$\therefore$ $2.67 = E^\theta$ E^θ has $\downarrow$

ANODE	CATHODE	EFFECT E^θ
increase[ion]	—	more $-ve$
—	increase [ion]	more $+ve$

ANODE $Mg^{2+} + 2e^- \rightleftharpoons Mg_{(s)}$ $-2.$

CATHO $Cu^{2+} + 2e^- \rightleftharpoons Cu_{(s)}$ $+$

what will reduce $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$ $+0.77$?   OILRIG
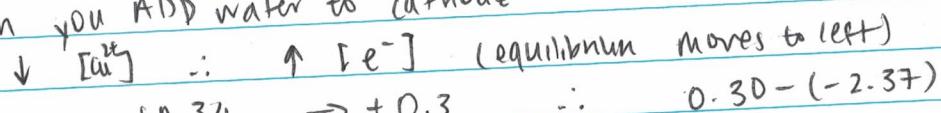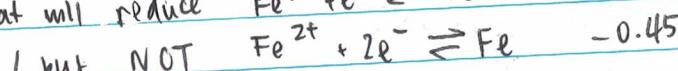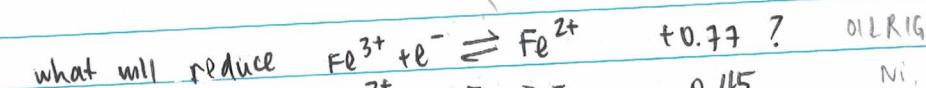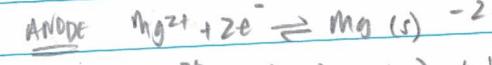
but NOT $Fe^{2+} + 2e^- \rightleftharpoons Fe$ -0.45 Ni, Sn, Pb, H

$\rightarrow$ must be oxidized itself $\therefore$ must be more $-ve$ than $+0.77$
but cannot reduce Fe^{2+} $\therefore$ must be less $-ve$ than -0.45
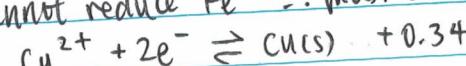
$Cu^{2+} + 2e^- \rightleftharpoons Cu_{(s)}$ $+0.34$

(if Ox anything with oxid.)   reduce $Cu^{2+} + e^- \rightleftharpoons Cu^+$ $+0.15$ $\therefore$ tve
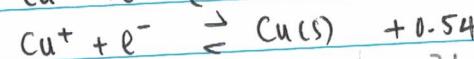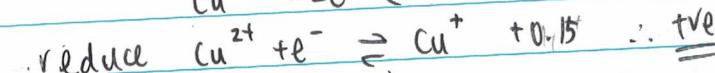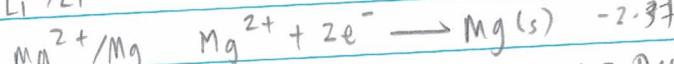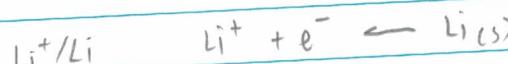
$Cu^+ + e^- \rightleftharpoons Cu_{(s)}$ $+0.54$

Cu^{2+}/Cu $Cu^{2+} + 2e^- \leftarrow Cu + 0$

ion e⁻ cell

Li^+/Li $Li^+ + e^- \leftarrow Li_{(s)}$ -3.04

Mg^{2+}/Mg $Mg^{2+} + 2e^- \longrightarrow Mg_{(s)}$ -2.37

F_2/F^- $\frac{1}{2}F_2 + e^- \rightarrow F^- +$

Ecell $+ve - (-ve) = -2.71 + 3.04$ $E^\theta = -2.53$ (can power itself once there is additi... powe...)

$= 0.67$ more likely

$2Li + Mg^{2+} \longrightarrow 2Li^+ + Mg$

$Cu + F_2 \longrightarrow Cu^{2+} + 2F$
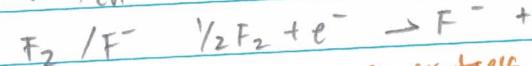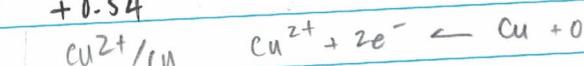
spontaneity $= \Delta G = \Delta H - T\Delta S$

· entropy / disorder

· exo/endothermic = enthalpy change

Cell POTENTIAL & SPONTANEITY

$$\Delta G = -n \; F \; E_{cell}$$

moles → faraday's constant (96500C) C/mol⁻ → E⁰

$Cu + F_2 \longrightarrow Cu^{2+} + 2F^-$

+2.53V , 2 moles of e⁻ transf

ΔG in JOULES $= -2 \times 2.53 \times 96500$

$= -498290$ J

$= -488.29$ KJ

∴ if $E\theta$ is +ve $= \Delta G = -ve$ ∴ spontaneous.

if $\Delta G = -350$, reacton = 1e⁻

ΔG is always kJ

$-350,000 = -FE_{cell}$

$$\frac{-350000}{-350000} \quad E_{cell} = +3.63$$

[Cu] + [F₂] ⟶ [Cu²⁺] + [2F⁻]

copper   platinum

Cu being oxidized , RIG, ∴ anode.

OIL   ANox

HALF CELL DIAGRAM

Fe^{3+}/Fe^{2+} +0.77 oxidzed

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$

Cl_2/Cl^- +1.36 reduced

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

1 mol Fe^{3+} & Fe^{2+} $Fe^{2+} Fe^{3+}$ Pt Cl_2 @ 1 atm   Pt Cl^- -ve

anode NO_3^- K⁺ → cathode

(-ve)   (+ve)

$2Fe^{2+} + Cl_2 \longleftrightarrow Fe^{3+} + 2Cl^-$
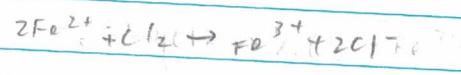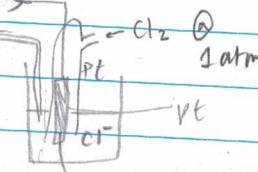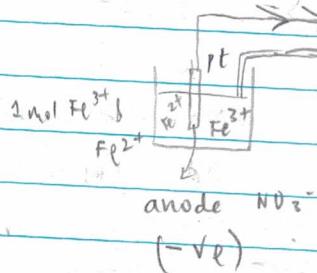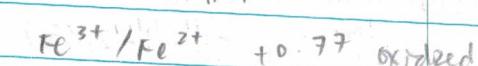
$E^\theta = 1.36 - 0.77$

$= 0.6$ V

e flow = -ve → +ve

$$\Delta G = \frac{-2 \times 96500 \times 0.6}{1000} = -115.8 \; kJ$$

ΔG diff - reaction X completion

- conditions X standard

- not 1 mol dm⁻³

· more reactive species displaces less reactive

↳ MORE REACTIVE METAL REDUCES less reactive metal

(other way is NOT spontaneous)

$$Mg^{2+} + 2e^- \rightleftharpoons Mg(s) \qquad [-2.37]$$

$E^\theta (V)$

(magnesium)
(displaced Pb)
↓

$$Pb^{2+} + 2e^- \rightleftharpoons Pb(s) \qquad [-0.13]$$
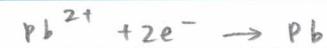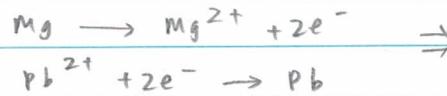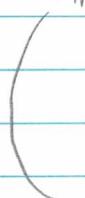
Mg will reduce led.
Pb oxidising magnesium

$$Mg \longrightarrow Mg^{2+} + 2e^- \qquad \Rightarrow \qquad Pb^{2+} + Mg \longrightarrow Mg^{2+} + Pb$$
$$Pb^{2+} + 2e^- \longrightarrow Pb$$

red. agent

oxidising agent

→ SOLID (s)!

best reducing agents e.g. Caesium → low electronegativity (↑shielding ↑shells) ∴ ↓attract

oxidising agents e.g. Fluorine → high electron-v. (↓radius ↓shielding) ↑attract

↳ $F_2(g)$

$\boxed{E^\theta}$

OXIDISED ←——————→ REDUCED

reducing agent   oxidising agent

e.g.

$$Al^{3+} + 3e^- \xleftarrow{} Al(s) \qquad E^\theta = -1.66$$
$$Mn/Mn^{2+} \qquad E^\theta = -1.18$$

Al is a better reducing agent ←

$$Al^{3+} + 3e^- \rightleftharpoons Al(s) \qquad \Rightarrow \quad 2Al^{3+}_{(aq)} + 6e^- \xleftarrow{} 2Al(s)$$
$$Mn^{2+} + 2e^- \rightleftharpoons Mn(s) \qquad \rightarrow \quad 3Mn^{2+}_{(aq)} + 6e^- \longrightarrow 3Mn(s)$$

$$2Al + 3Mn^{2+} \longrightarrow 2Al^{3+}_{(aq)} + 3Mn($$
(s)   better (aq)   (aq)
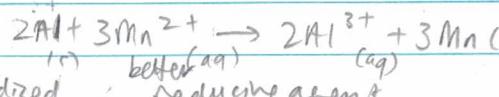
Al has lost electrons - oxidized ∴ reducing agent,
reduces less reactive
metal.

$$2Al + 3MnO \longrightarrow Al_2O_3 + 3Mn(s)$$

→ ∴ negative goes to left, gets reduced
positive goes to right & get oxidising agent

$$Al + 1.5\ I_2 \longrightarrow Al^{3+} + 3I^- \qquad bc \quad \tfrac{1}{2}I_2 + e^- \rightarrow I^-$$
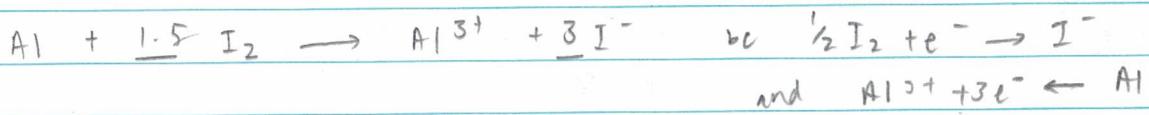
and $Al^{3+} + 3e^- \leftarrow Al$

(purple)   (colourless)   electrons MUST be balanced.
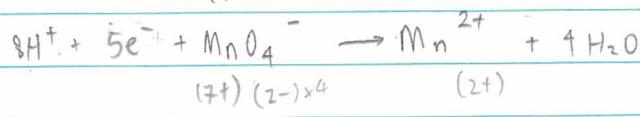
$$8H^+ + 5e^- + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$$
(7+) (2-)×4   (2+)
$^{-8}$

- ? glowstick
- x colourimetry

- activation energy ... sulphurous?
- titration of thiosulphate?
- MnO_2 (vid c conc?)
-

REDOX TITRATION

$$5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^- \quad [reducing agent]$$

$$8H^+ + 5e^- + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O \quad (oxidizing\ agent)$$

$$\underset{purple}{} \qquad \underset{colourless}{}$$

$$\therefore \quad 8H^+ + 5Fe^{2+} + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

- iodometric (from HI)

$$2Cu^{2+}_{(aq)} + 4I^-_{(aq)} \longrightarrow 2CuI + I_2 \quad [measuring\ Cu\ conc\ in\ e.g.\ brass$$
with titration of products]
② x ↑ 6.675×10^{-3} → BACKWARDS

Cu^{2+} 250cm³ ③ 0.01335 MOLES in 50cm ← ×2

$$I_2 + 2S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2I^-$$
(thiosulphate)

④ $\frac{50}{250} \times 0.01335$
$= 0.06675$ MOLES

c= 0.1335 CONC: 0.500
vol: 50cm³ vol: 26.7 cm³

$g = 4.238125 g$
4.24 (3 s.f)

① moles: 6.675×10^{-3} moles: 0.01335
$6.675\times10^{-3} = c\times v$

↓40%
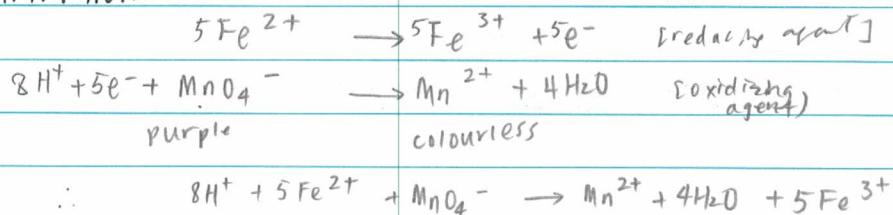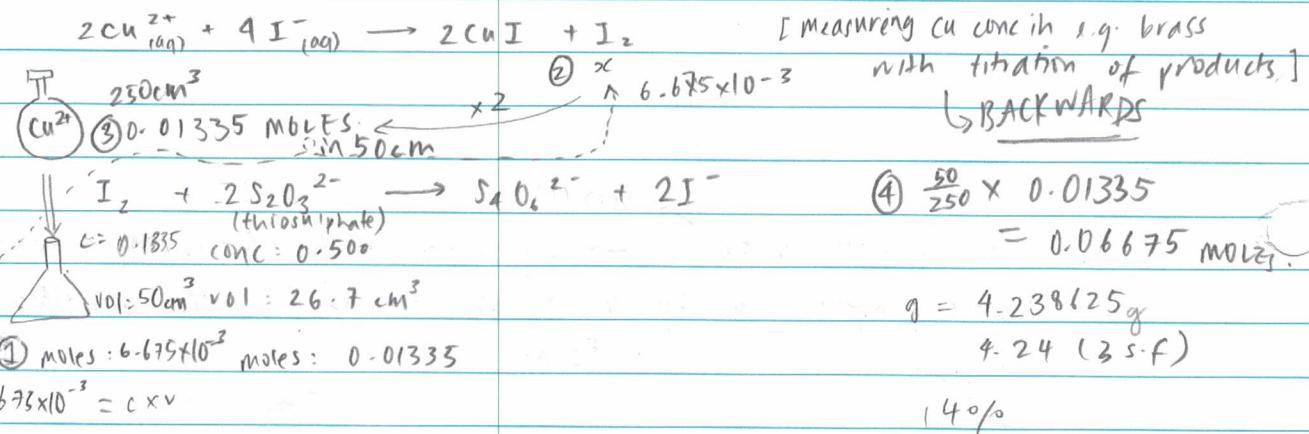
BIOLOGICAL OXYGEN DEMAND

→ anaerobic bacteria

amt of O_2 required to decompose organic matter over a period of time

high BOD : low O_2 content

$$2Mn^{2+} + O_{2(g)} + 4OH^- \longrightarrow 2MnO_2 + 2H_2O$$
$\frac{1}{2}x \quad \frac{x}{4}$ $\cancel{x} \quad \frac{x}{2}$

moles of $2S_2O_3^{2-}$
$\div 4 = O_2$ moles

$$MnO_2 + 2I^- + 4H^+ \longrightarrow I_2 + 2H_2O + Mn^{2+}$$
$x \quad \frac{x}{2}$ $x \quad \frac{x}{2}$

$$2S_2O_3^{2-} + I_2 \longrightarrow 2I^- + S_4O_6^{2-}$$
x $x \quad \frac{x}{2}$
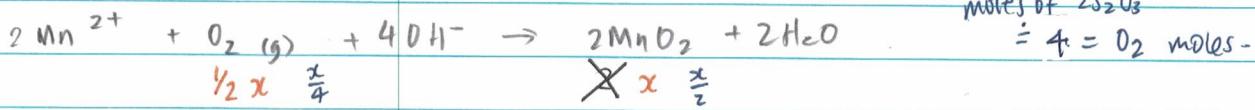
① CONC: 0.0125
vol: 7.71cm³ 25cm³
↓ work
out mass of O_2

moles: 1.6375×10^{-5} $\div 4 = 2.409375\times10^{-5}$ moles of O_2 $\times 32$ $\times 10$
$= 7.71\times10^{-4}$ $= 7.71\times10^{-3}$

of the mass. $\frac{25cm^3}{250cm^3}$