

# CHEMISTRY

# CONTENT CHECKLIST (FROM TEXTBOOK)

## DYNAMIC EQUILIBRIUM SYSTEMS

- o rate of chemical reactions
- o equilibrium
- o equilibrium systems in the environment

## ACIDS & BASES 'how do chemicals swap protons'

- o acids & bases
- o buffers
- o indicators
- o volumetric analysis

## OXIDATION & REDUCTION

- o redox reactions
- o galvanic cells
- o electrolysis

UNIT 3

## INDUSTRIAL CHEMISTRY

- o key products from the chemical industry
- o resources & the environment

## ORGANIC CHEMISTRY

- o structure of organic molecules
- o reactions of organic molecules

UNIT 4

## ORGANIC SYNTHESIS

- o from monomers to polymers
- o from fats and oils to soaps and biodiesel

## BIOCHEMISTRY

- o from amino acids to proteins
- o uses of proteins

# equilibrium

CAN ONLY REACH EQUILIBRIUM IF IT IS  
- a closed system      - a reversible reaction → chemical or physical

## AT EQUILIBRIUM THE FORWARD + REVERSE REACTIONS ARE EQUAL

the ratio of the [ ] of products & reactants is reflected by the  
**EQUILIBRIUM CONSTANT (K or Keq)**

this is constant for a given reaction @ a given temp.  
and is calculated by:

$$K = \frac{[\text{Product 1}]^{\text{coefficient 1}} * [\text{Product 2}]^{\text{coefficient 2}}}{[\text{Reactant 1}]^{\text{coefficient 1}} * [\text{Reactant 2}]^{\text{coefficient 2}}}$$

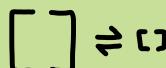
only for gases

MOL L<sup>-1</sup> can be used with solutions  
for gases, volume is often used as it is proportional to concentration

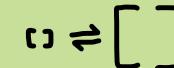
equilibrium can be described as left or right



$$K_{\text{eq}} > 1$$



$$K_{\text{eq}} < 1$$



$$K_{\text{eq}} = 1$$

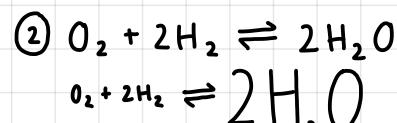
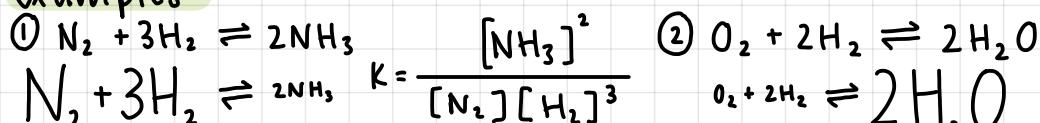


note:

equilibrium constant  
= number

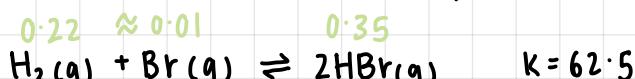
equilibrium expression  
= formula

examples:



equilibrium = ——

③ what is the equilibrium concentration of Br<sub>2</sub> if [HBr] = 0.35M and [H<sub>2</sub>] = 0.22M



$$\frac{(0.35)^2}{[\text{Br}_2](0.22)} = 62.5$$

$$[\text{Br}_2] = \frac{(0.35)^2}{(62.5)(0.22)} = 8.91 \times 10^{-3} \text{ M}$$

note: can't do this without K but may have to work backwards i.e. given 3 [ ]'s

**RATE OF REACTION:** how quickly the reaction occurs

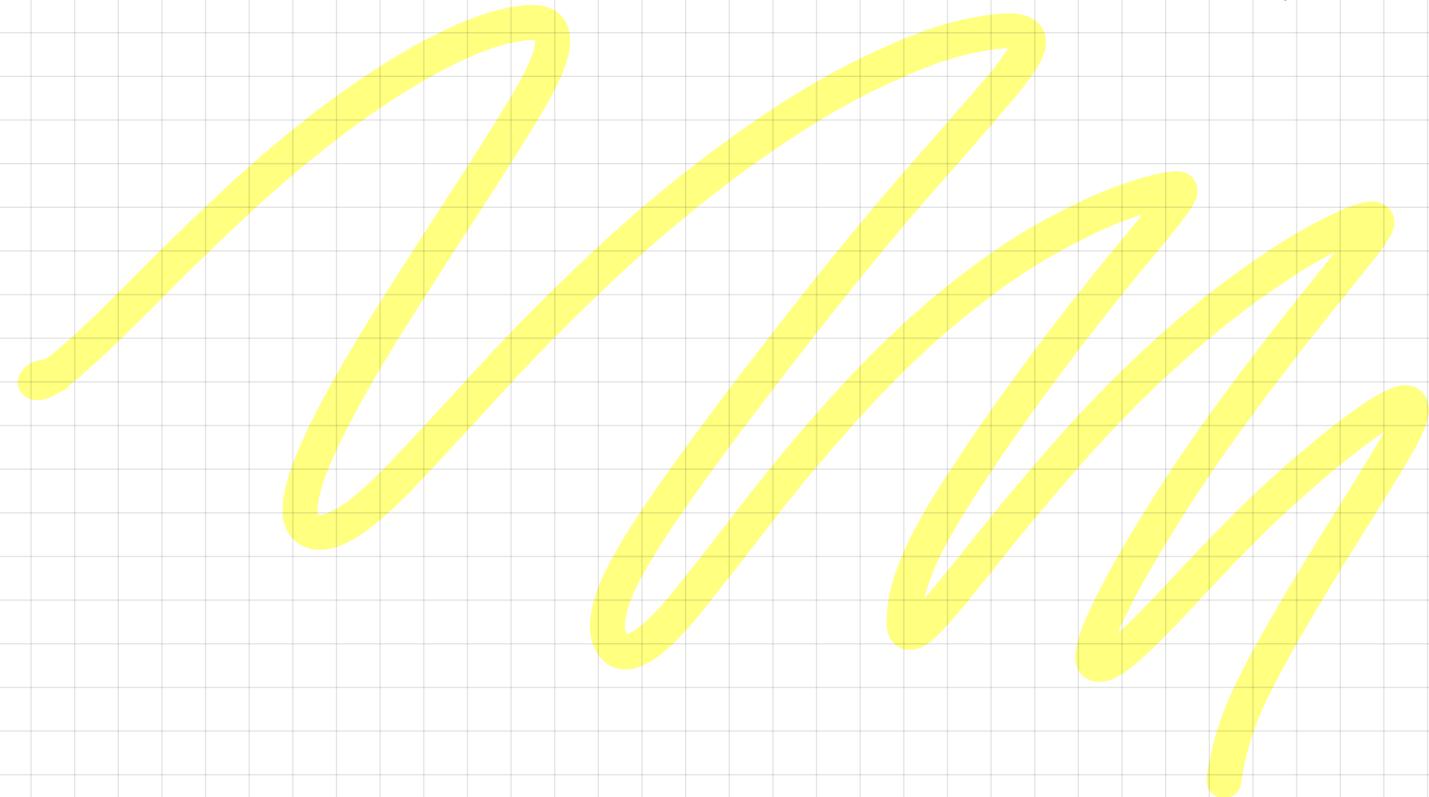
**YIELD:** the amount of product produced

the position of equilibrium affects the yield , reaction rate affects how fast the yield is produced

when changes are made to a system @ equilibrium  
the forward, reverse or both reactions may be affected

if the rate of either or both reaction is altered then the system will no longer be @ equilibrium & the system will adjust to reestablish equilibrium

∴ anything that changes the rate of a reaction may affect equilibrium



# LECHATELIER'S PRINCIPLE

if a change is made to a system that is at equilibrium, the system will act in such a way so as to partially counteract the change

i.e. for every action, there's a partial reaction

## CHANGES TO CONCENTRATION

when  $[ ]$  of a reactant or product changes, the RR which uses the substance will be altered

### LCP:

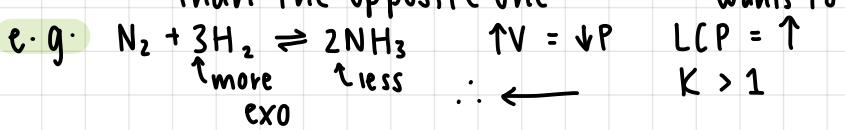
- $[ ] \uparrow$ , reaction using the substance will be favoured
- $[ ] \downarrow$ , reaction producing the substance will be favoured

note: adding more of a solid will not change the concentration of the substance.  
∴ will not affect the position of equilibrium

## CHANGES TO VOLUME OR PRESSURE

when volume or pressure of a gas system is changed, the  $[ ]$  of all gases is altered

- if the reactants & products have different no's. of mols, they will be affected to different extents
- reaction w/  $\uparrow$  mols of gas as reactants will be effected more
  - ↳ ∵ this reaction will  $\uparrow$  or  $\downarrow$  in rate than the opposite one



### LCP:

- volume  $\downarrow$  or pressure  $\uparrow$ , system favours reaction that produces less moles of gas
- volume  $\uparrow$  or pressure  $\downarrow$ , system favours reaction that produces more moles of gas
- diluting or concentrating an (aq) system has the same effect on the (aq) substances

## CHANGES TO TEMPERATURE

- a change in temp is the only thing that can change the equilibrium constant
- $\uparrow \text{temp} = \uparrow FRR + \uparrow RRR$ 
  - ↳ endothermic increases more
- $\downarrow \text{temp} = \downarrow FRR + \downarrow RRR$ 
  - ↳ exothermic will be faster
- if forward is favoured = more products @ new equilibrium  
 $\therefore K_{eq} \uparrow$
- if reverse is favoured = more reactants @ new equilibrium  
 $\therefore K_{eq} \downarrow$

## ADDING A CATALYST

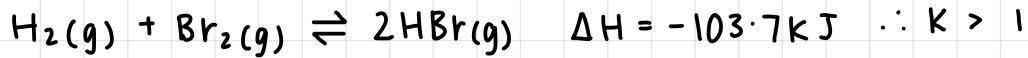
- doesn't affect the position of equilibrium
- EA for both FRR & RRR is lowered  
 $\therefore$  both reaction rates  $\uparrow$ 
  - ↳ equilibrium is reached more quickly

## INCREASING THE SURFACE AREA

- $\uparrow SA = \uparrow RRR$ 's but not the position of equilibrium

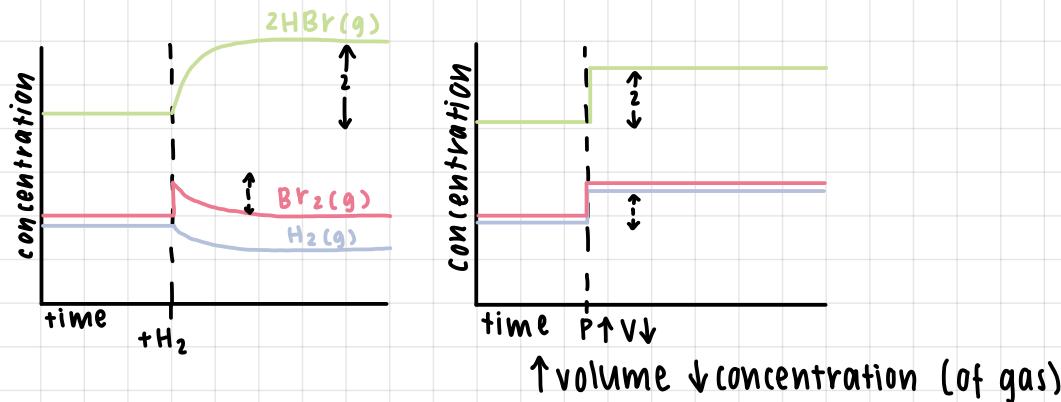
# equilibrium constant changes

example :

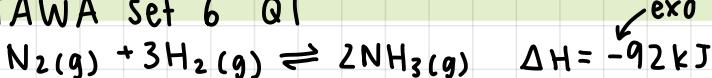


negative  $\Delta H$   
 $\therefore$  exothermic

|                       | [H <sub>2</sub> ] | [Br <sub>2</sub> ] | [HBr]     | K value  |   |
|-----------------------|-------------------|--------------------|-----------|----------|---|
| add H <sub>2</sub>    | ↓                 | ↓                  | ↑         | same     | wants to 'get rid' of the added H <sub>2</sub><br>so reaction using it = favoured |
| add HBr               | ↑                 | ↑                  | ↓         | same     | wants to 'get rid' of the added HBr<br>so reaction using it = favoured            |
| remove H <sub>2</sub> | ↑                 | ↑                  | ↓         | same     | wants to make more H <sub>2</sub> so reaction<br>that is producing is favoured    |
| remove HBr            | ↓                 | ↓                  | ↑         | same     | wants to make more HBr so reaction<br>that is producing is favoured               |
| temp ↑                | ↑                 | ↑                  | ↓         | decrease | favours endothermic $\therefore$ the<br>reverse is favoured                       |
| temp ↓                | ↓                 | ↓                  | ↑         | increase | favours exothermic $\therefore$ the<br>forward is favoured                        |
| P ↑ + V ↓             | no change         | no change          | no change | same     |   |



STAWA Set 6 Q1



a) write equilibrium constant expression

$$\frac{[NH_3]^2}{[N_2][H_2]^3}$$

b) predict the effect on [NH<sub>3</sub>]

i - ↑T, equilibrium shifts left (reverse reaction favoured)

$$\therefore [NH_3] \downarrow$$

ii - ↑[N<sub>2</sub>(g)], forward favoured [NH<sub>3</sub>] ↑

iii - ↓V LCP: ↓V = ↑P  
favours side w/ fewest particles  
= forward  $\therefore [NH_3] \uparrow$

iv - remove catalyst, doesn't change []

random note: endo

$\downarrow T$   
FRR ↓  
RRR ↓

= more particles  
more impact  
 $\therefore$  reverse favoured

explain

i - FRR ↑ RRR ↑ more products than reactants  
more particles = more successful collisions  
new equilibrium  $\xrightarrow{\text{slow}} \xleftarrow{\text{fast}}$  . : left

ii) Increasing T means particles are moving faster (have more energy) resulting in an increase of both the forward and reverse reaction rates (more successful collisions). The reaction is exothermic meaning equilibrium is on the right. More products means reverse reaction rate will increase more than the forward reaction rate resulting in reverse reaction being favoured.

iii) Increasing the amount of N<sub>2</sub> molecules will result in more successful collisions between reactants favouring the forward reaction.

iv) Reducing V will increase P meaning particles are closer together and an increase of both the forward and reverse reactions (more successful collisions). There is a higher ratio of particles on the left resulting in forward reaction rate increasing more than the reverse reaction rate resulting in reverse reaction being favoured.

v) Catalysts affect reaction rate but have no effect on equilibrium.

# dynamic equilibrium

when a reversible reaction in a closed system is able to continue until the rate of reactants forming products is the same as the rate of products forming reactants

when forward reaction rate (FRR) = reverse reaction rate (RRR) :

- amounts of products & reactants is constant

↳ \* this doesn't mean that the amount of products & reactants are equal \*

**AT DYNAMIC EQUILIBRIUM** the FRR = RRR and both are constant

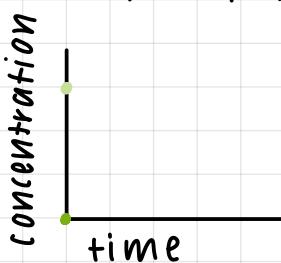
## reaching equilibrium

more reactants @ start  $\therefore [ ] \uparrow$

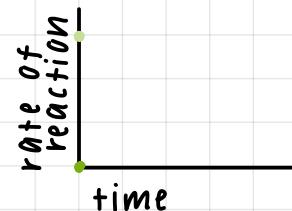
$\therefore$  initial FRR = fast

↑ @ this time there are no products

$\therefore RRR = 0$

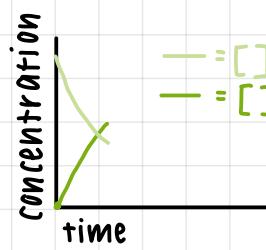


— =  $[ ]$  of products  
— =  $[ ]$  of reactants

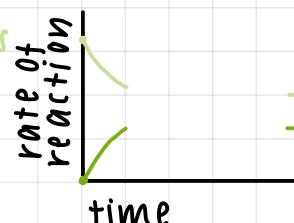


— = RRR  
— = FRR

- as reactants form products :
  - ↳  $[ ]$  reactants  $\downarrow$
  - ↳  $[ ]$  products  $\uparrow$
- = FRR  $\downarrow$ , RRR  $\uparrow$

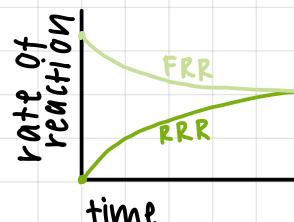
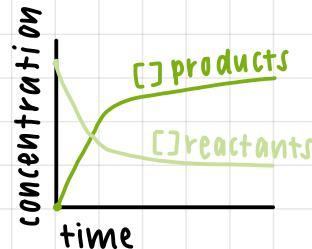


— =  $[ ]$  reactants  
— =  $[ ]$  products



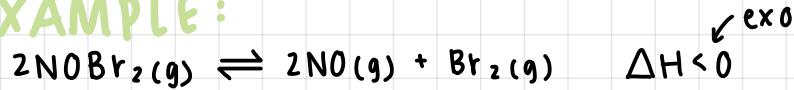
— = FRR  
— = RRR

- continues until:
  - ↳ FRR = RRR
  - ↳  $[ ]$  reactants +  $[ ]$  products remain constant

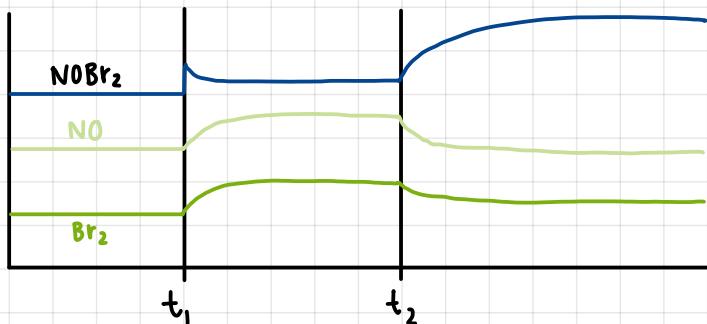


# concentration graphs

## EXAMPLE:



a) @  $t_1$ ,  $\text{NOBr}_2$  was added at a constant temperature

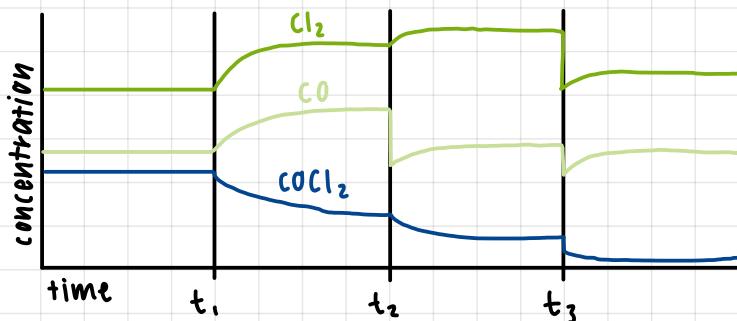
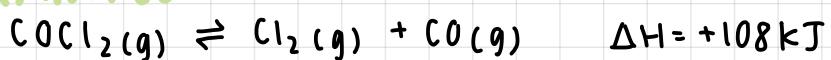


b) @  $t_2$ , temp ↑

$2:2$  ratio  $\therefore$  fall = rise  
 $1:2 \therefore \frac{1}{2}$   
 $\text{Br}_2(g) \rightarrow \text{NOBr}_2$

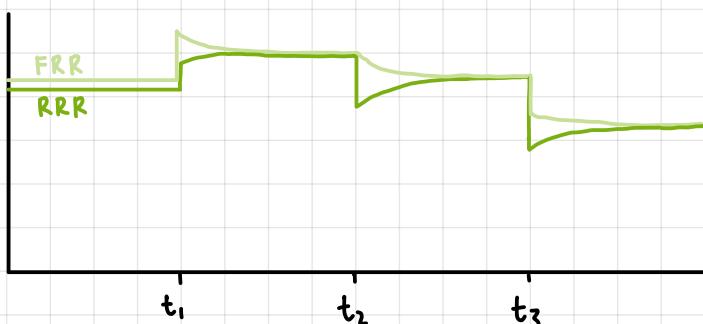
# reaction rate graphs

## EXAMPLE:



$t_1$  = temp increased  
 $t_2$  =  $\text{CO}$  removed  
 $t_3$  =  $\uparrow V \therefore \downarrow P$

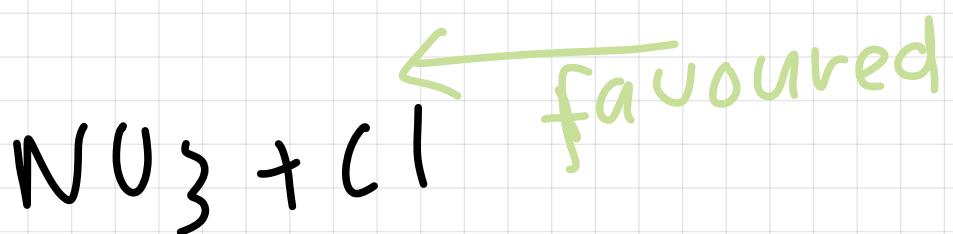
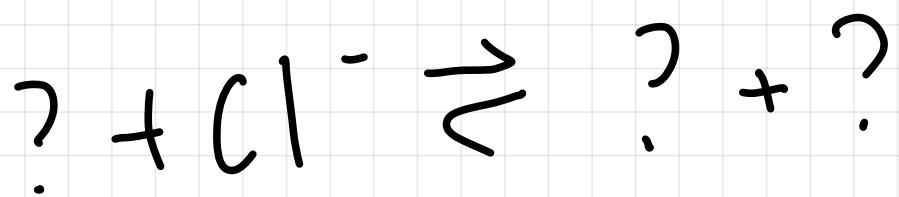
reaction rate graph from ↗



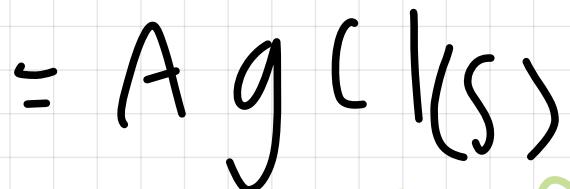
All  $\text{NO}_3^-$  =  
soluble



Adding  $\text{AgNO}_3(\text{aq})$



= precipitation



= disappear

∴ Shift to make  
more

# PROPERTIES OF ACIDS + BASES

## ACIDS

- sour
- conduct electricity 
- ↳ can be strong electrolytes in (aq) solution
  - still strong when diluted
- reacts w/ metals (forms H<sub>2</sub> gas)
- reacts w/ bases (metal hydroxides) to form water and a salt
- pH = < 7
- reacts w/ carbonates + bicarbonates to produce a salt, water + CO<sub>2</sub> gas

covalent molecular

they start w/ Hydrogen usually

(e.g. HCl, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub> etc.)

strong acids      not water though

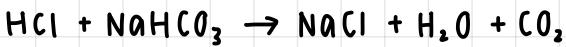
## BASES

- bitter
- react w/ acids (form water + a salt)
- feel slippery
- can be strong or weak electrolytes in (aq) solution

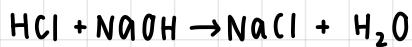
acid + active metal = salt + H<sub>2</sub> gas



acid + carbonates / bicarbonates



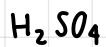
acids neutralise bases



↓  
ALWAYS produces  
a salt (ionic  
compound) + H<sub>2</sub>O

# acids + bases examples

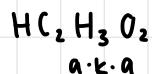
SULFURIC ACID



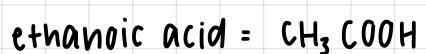
HYDROCHLORIC ACID



ACETIC ACID



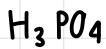
a.k.a



NITRIC ACID



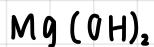
PHOSPHORIC ACID



SODIUM HYDROXIDE



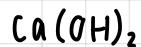
MAGNESIUM HYDROXIDE



POTASSIUM HYDROXIDE



CALCIUM HYDROXIDE

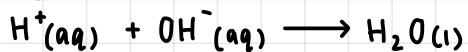


# acids + bases example questions

## EXAMPLE 1

dilute ethanoic acid is added to sodium hydroxide solution

equation:



observation:

no visible reaction (2 colourless  $\rightarrow$  colourless)

## EXAMPLE 2

which of the following is true for all bases?

- (a) they contain hydroxide ( $\text{OH}^-$ ) ions — also correct (if 'basic solutions'  
then = wrong)
- (b) they are able to accept hydrogen ions ←
- (c) they are soluble in water
- (d) they react with metals to produce hydrogen gas

## EXAMPLE 2

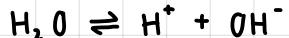
which of the following solutions will have a pH value closest to 7?

- (a)  $1.0 \text{ mol L}^{-1}$   $\text{H}_2\text{SO}_4(\text{aq})$  strong  $= \uparrow$  ionising  $\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$
- (b)  $2.0 \text{ mol L}^{-1}$   $\text{NaOH}(\text{aq})$   $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$   $\text{H SO}_4 \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$
- (c)  $1.0 \text{ mol L}^{-1}$   $\text{HNO}_3(\text{aq})$  strong
- (d)  $2.0 \text{ mol L}^{-1}$   $\text{NH}_3(\text{aq})$  weak base  
 $= \downarrow$  ionising  
least effect on pH

all bases accept H to produce  
 $\text{OH}^-$  ions in water

# HYDROGEN IONS, ACIDITY AND PH

## H<sup>+</sup> IONS FROM WATER



turns into ions - "self ionisation" occurs to a very small extent:

$$[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7} \text{ M}$$

↑ since they're equal, water = a neutral solution

$$K_w = [\text{H}^+] \times [\text{OH}^-] = 1 \times 10^{-14} \text{ M}^2$$

↑ "ion product constant" for water

## THE pH CONCEPT

from 0 → 14

$$\text{pH} = -\log[\text{H}^+]$$

↑ in neutral,  $-\log(1 \times 10^{-7}) = 7$   
acidic,  $[\text{H}^+] > 10^{-7}$

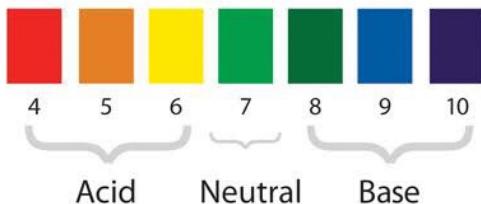
$$\text{pH} < -\log(10^{-7})$$

↳ pH < 7

0-7 pH = acid range

7-14 pH = base range

### Universal Indicator pH Color Chart



## ION PRODUCT CONSTANT

K<sub>w</sub> is constant in every (aq) solution

if  $[\text{H}^+] > 10^{-7}$  then  $[\text{OH}^-] < 10^{-7}$  acidic

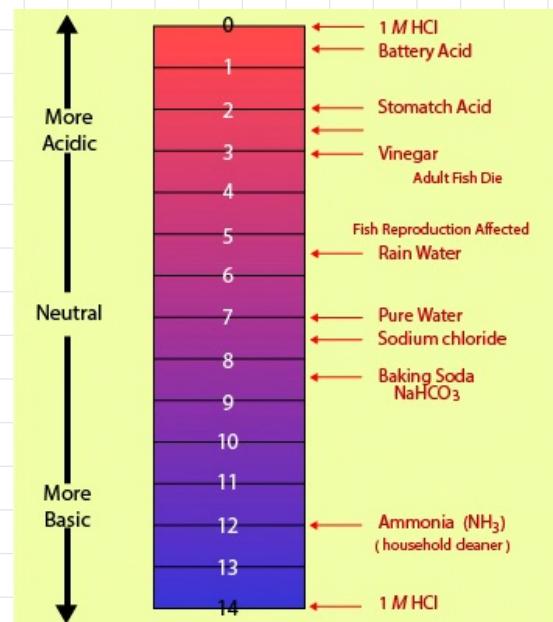
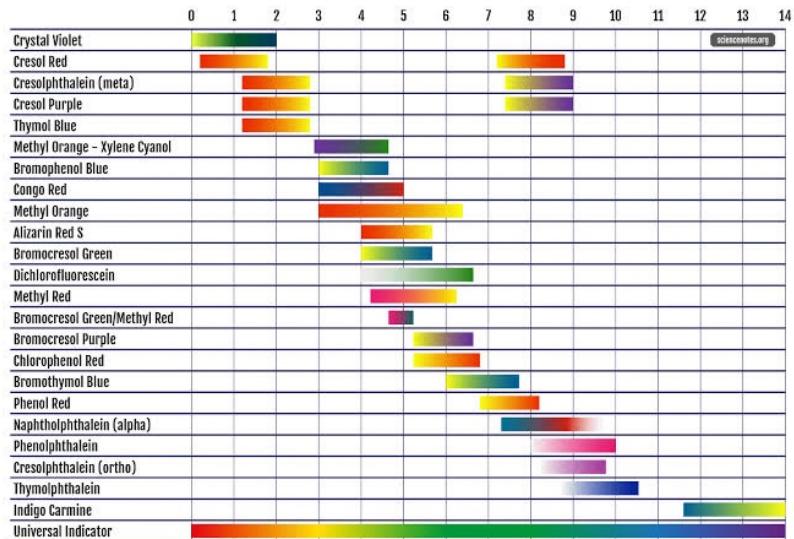
if  $[\text{H}^+] < 10^{-7}$  then  $[\text{OH}^-] > 10^{-7}$  basic

basic = can form a base

alkaline = dissolved in water

## ACID-BASE INDICATORS

given @ 25°C (may change at different temps.)



# pH/ acidity examples

## EXAMPLE 1

calculate the hydroxide ion concentration in the following situations:

(a) a standard  $0\cdot109 \text{ mol L}^{-1}$  hydrochloric acid used to analyse floor cleaning solutions

$$C(\text{HCl}) = 0\cdot109 \quad [OH^-] = \frac{10^{-14}}{0\cdot109} = 9\cdot62 \times 10^{-14} \text{ M}$$

when  $[H^+] \times [OH^-] = 10^{-14}$

(b) A  $0\cdot125 \text{ mol L}^{-1}$  sulfuric acid solution used to dissolve samples of malachite (copper ore containing mostly copper carbonate). Assume the first hydrogen is fully ionised + the second is 10% ionised

$$[H^+] = 0\cdot125 + 0\cdot0125 = 0\cdot1375 \text{ M} \quad [H^+] [OH^-] = 10^{-14} \quad [OH^-] = \frac{10^{-14}}{0\cdot1375} = 7\cdot27 \times 10^{-14} \text{ M}$$

## EXAMPLE 2

a laboratory chemist wants a low concentration hydrochloric acid solution with a pH of 5.00, for calibrating conductivity meter. What volume of water must be added to a 25mL sample of HCl solution

$$\begin{array}{lll} \text{pH} = 3.60 & \text{pH} = 5 & = 0.628 \text{ L} \\ [H^+] = 10^{-3.6} = 2.51 \times 10^{-4} \text{ M} & [H^+] = 10^{-5} \text{ M} & = 628 \text{ mL} \\ n(H^+) = (2.51 \times 10^{-4})(0.025) & V = \frac{n}{c} & \\ = 6.28 \times 10^{-6} \text{ mol.} & = \frac{6.28 \times 10^{-6}}{10^{-5}} & \end{array}$$

# ACID + BASE THEORIES

## ARRHENIUS DEFINITION

1887

ACIDS produce  $H^+$  ions in (aq) solution  
BASES produce  $OH^-$  ions when dissolved in water



limited to (aq) solutions + only for hydroxides

)

$NH_3$  (ammonia) couldn't be an Arrhenius base: no  $OH^-$  produced

### POLYPROTIC ACIDS

some compounds have more than one ionisable hydrogen to release

$HNO_3$  - monoprotic  
 $H_2SO_4$  - diprotic ( $2H^+$ )  
 $H_3PO_4$  - triprotic ( $3H^+$ )

↑ however,  
doesn't mean  
it's stronger

### ACIDS

not all the hydrogen in an acid may be released as ions

↑ only those that have very polar bonds are ionisable  
(when the hydrogen is joined to a very electronegative element)

## BRØSTED-LOWRY

1923

- broader definition than arrhenius

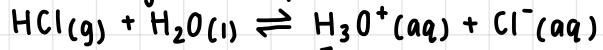
ACID is hydrogen-ion donor ( $H^+$  or proton)

BASE is hydrogen-ion acceptor

acids and bases always come in PAIRS

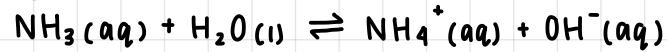
HCl is an acid

↑ when it dissolves in water, it gives its protons to water



↑ water is a base, makes hydronium ion

### WHY AMMONIA IS A BASE



ammonia = hydrogen-ion acceptor (base)

+ water = hydrogen-ion donor (acid)

] causes the  $[OH^-]$  to be

greater than in pure water

∴ the ammonia solution is basic

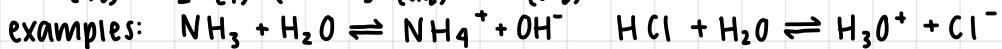
### ACIDS + BASES COME IN PAIRS

conjugate base: remainder of the original acid (after it donates its hydrogen ion)

conjugate acid: the particle formed when the original base gains a hydrogen ion

∴ a conjugate acid-base pair is related by a LOSS OR GAIN of a single hydrogen ion

acid + base  $\rightleftharpoons$  conjugate base + conjugate acid



base acid

c.a.

c.b.

acid base

c.a.

c.b.

water = amphoteric

(can act as both acid and base)

# LEWIS ACIDS + BASES

lewis acid : electron pair acceptor

lewis base : electron pair donor

acids often don't even need hydrogen

# strong or weak? - acids

## STRENGTH

acids + bases are classified based on the degree to which they ionise in water

**STRONG** = completely ionised in (aq) solution (ionise 100%)

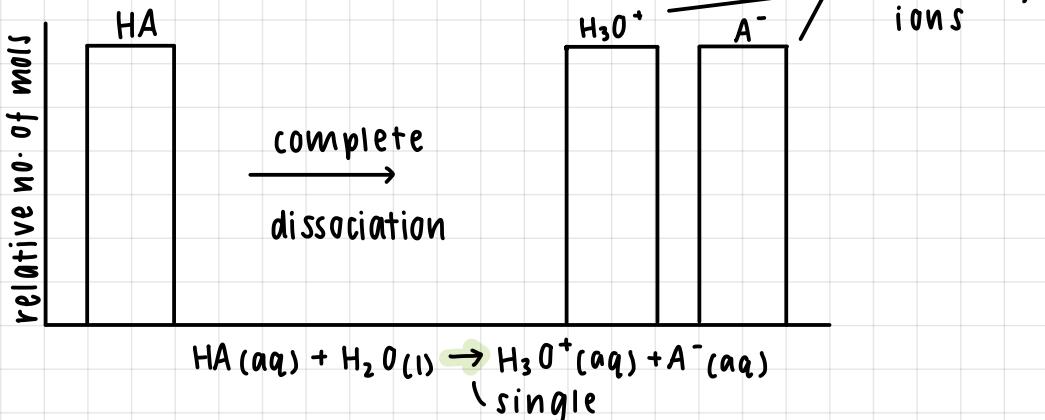
**WEAK** = slightly ionise in (aq) solution

\* strength is very different from concentration \*

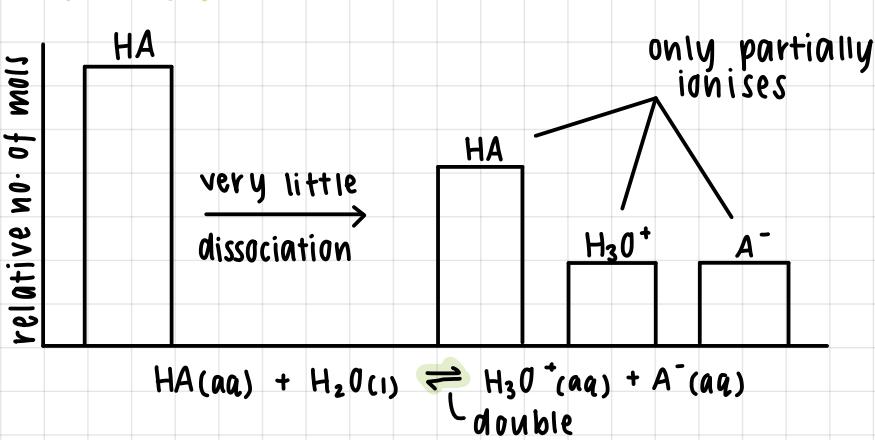
**strong**: it forms many ions when dissolved

example:  $\text{Mg}(\text{OH})_2$  is a strong base & falls completely apart when dissolved but not much dissolves so it is not concentrated

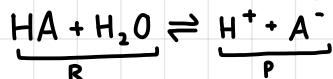
## STRONG ACID



## WEAK ACID



## MEASURING STRENGTH

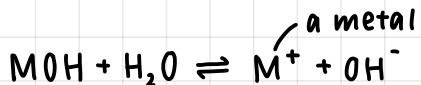


acid dissociation constant =  $K_a$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

i.e.  $\frac{\text{P}}{\text{R}}$

bigger = stronger  
(more products (ions))



base dissociation constant =  $K_b$

$$K_b = \frac{[\text{M}^+][\text{OH}^-]}{[\text{MOH}]}$$

bigger = stronger  
(more dissociated ions produced)

# strong or weak? acids questions

write the  $K_a$  expression for  $\text{HNO}_2$

$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

assume it's weak

write the  $K_b$  expression for  $\text{NH}_3$  (as  $\text{NH}_4\text{OH}$ )

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

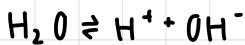
when it dissociates



( $\text{NH}_3$  dissolves in water)

equilibrium constant for pure water =  $5.13 \times 10^{-13} \text{ mol}^2 \text{ L}^{-2}$  at  $100^\circ\text{C}$ , which is correct?

a)  $[\text{H}^+] = 7.16 \times 10^{-7} \text{ mol L}^{-1}$ , water is acidic

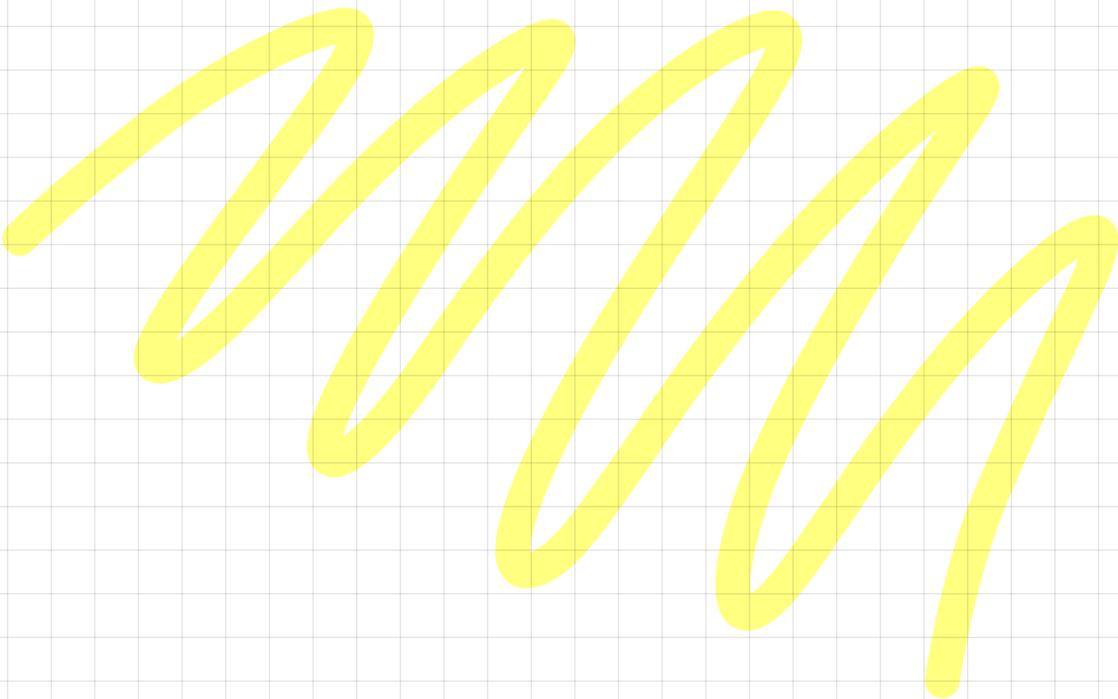


b)  $[\text{H}^+] = 7.16 \times 10^{-7} \text{ mol L}^{-1}$ , water is neutral

$\uparrow \text{temp}$  - more likely to break apart on forward reaction

c)  $[\text{OH}^-] = 7.16 \times 10^{-7} \text{ mol L}^{-1}$ , water is basic

d)  $[\text{OH}^-] = 7.16 \times 10^{-7} \text{ mol L}^{-1}$ , water is acidic



# titrations, salt hydrolysis + buffers

## TITRATIONS

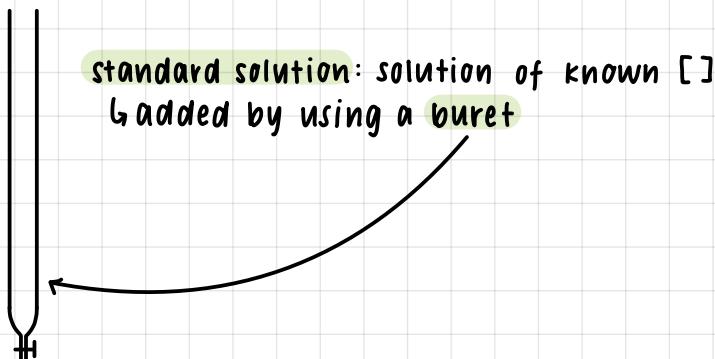
[ ] of base can be determined by performing a **NEUTRALISATION REACTION**

use an indicator to show this occurred

indicator for bases: phenolphthalein

Steps for neutralisation reaction

- measured volume of acid of unknown [ ] is added to a flask
- indicator is added
- base of known [ ] is added slowly until the indicator changes colour, volume is measured



## SALT HYDROLYSIS

salt: ionic compound that comes from
 

- ↳ anion of an acid
- ↳ cation of a base
- ↳ a neutralisation reaction

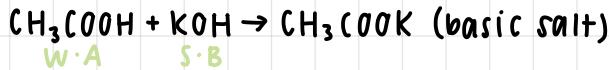
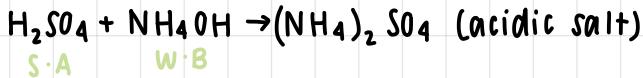
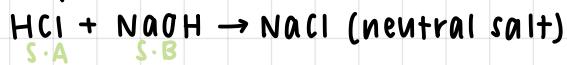
salt hydrolysis: a salt reacts w/ water to produce an acid or a base

**HYDROLYSING SALTS** usually come from:

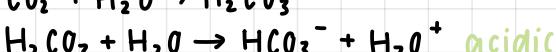
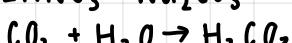
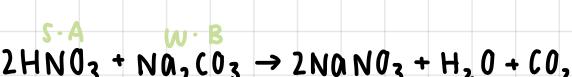
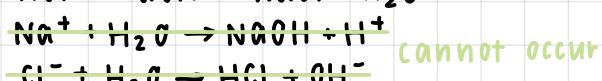
- a strong acid + weak base  
or
- a weak acid + strong base

note: strong acid + strong base  
= neutral salt

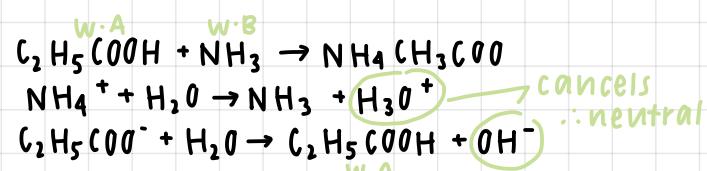
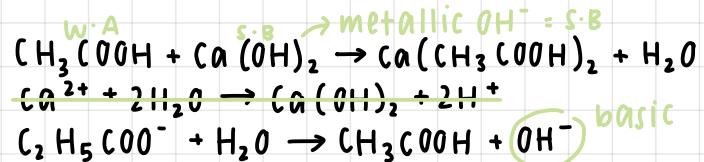
examples:



**WHY NOT S·A + S·B?**



W·B cannot form S·B



## BUFFERS

Solutions where the pH is constant even when small amounts of acid or base are added

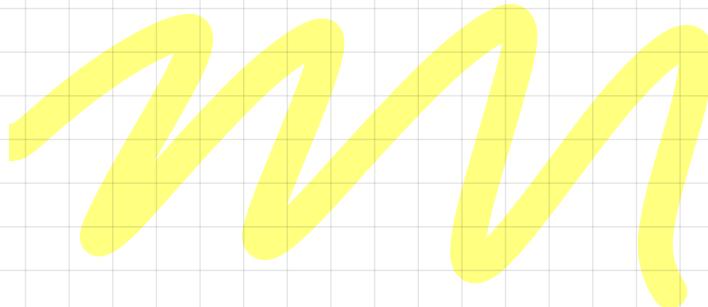
↳ made from a w·a· and one of its salts or w·b· and one of its salts

they RESIST CHANGES IN pH

one chemical neutralises any acid whilst one neutralises any base

↳ they produce each other in the process

**BUFFER CAPACITY** is how much acid or base can be added before it stops working  
(changes pH)



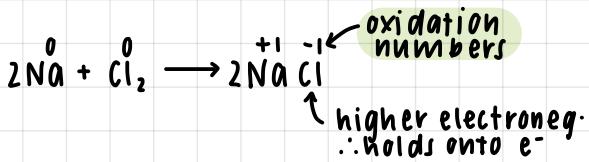
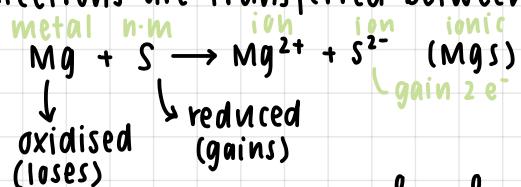
# OXIDATION-REDUCTION REACTIONS

i.e. redox

OXIDATION and REDUCTION always occur simultaneously

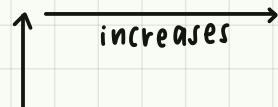
the substance gaining oxygen (losing e<sup>-</sup>s) is oxidised  
the substance losing oxygen (gaining e<sup>-</sup>s) is reduced ] don't need to have oxygen,  
just transfer of e<sup>-</sup>

electrons are transferred between reactants:

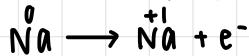


remember:

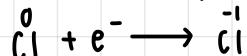
trends for electronegativity



oxidation half equation



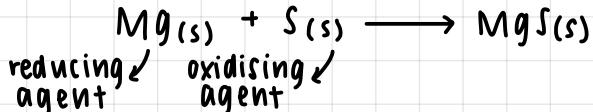
reduction half equation



Lose electrons  
Oxidation  
says  
Gain electrons  
Reduction

Oxidation  
Is  
Losing  
Reduction  
Is  
Gaining

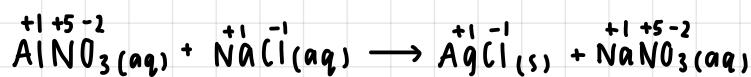
**REDUCING AGENT** = the thing being oxidised (losing e<sup>-</sup>s)  
**OXIDIZING AGENT** = the thing being reduced (gaining e<sup>-</sup>s)



## NOT ALL REACTIONS ARE REDOX

no change in oxidation number = not redox

e.g. precipitation reaction



## CORROSION

damage to metals

even faster w/ presence of salts + acids (because they make electrically conductive solutions ⚡)

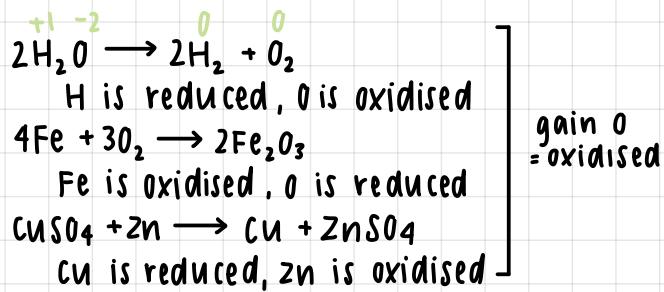
not noble metals (resistant to losing e<sup>-</sup>)

platinum      gold  
by corrosion

aluminium is protected by an oxide coating

prevention: 'sacrificing' a metal to save the second

# redox examples questions/equations



## NOTE

if it is diprotic, you  $\frac{1}{2}$  concentration

# oxidation numbers

- indicates an atom's degree of oxidation or reduction (all based on electronegativity)

## RULES FOR ASSIGNING OXIDATION NUMBERS

① oxidation number of an uncombined element is zero

② monatomic ion = its charge ( $Mg^{2+} = +2$ )

③ oxygen in compounds is -2

in peroxides = -1

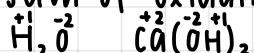
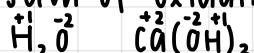
↑ sodium ( $Na_2O_2$ ), hydrogen ( $H_2O_2$ )

④ hydrogen in compounds +1

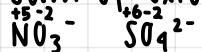
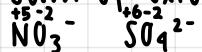
in metal hydrides = -1

↑ metals from group 1

⑤ sum of oxidation no.'s of atoms in compounds must = 0

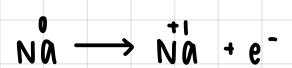


⑥ sum of oxidation no.'s of atoms in polyatomic ions must = ionic charge



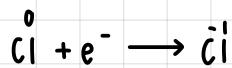
INCREASE IN OXIDATION NUMBER = oxidation

DECREASE IN OXIDATION NUMBER = reduction



sodium is oxidised

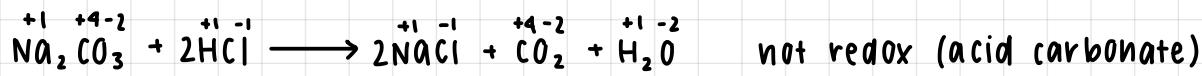
it is the reducing agent



chlorine is reduced

it is the oxidising agent

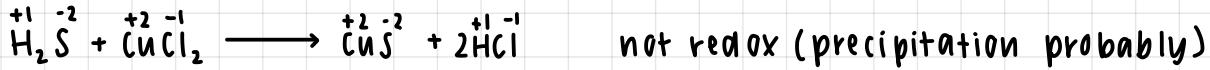
## EXAMPLES



not redox (acid carbonate)



not redox (neutralisation acid-base)



not redox (precipitation probably)

# redox equations

## ACTIVE METALS

- lose e<sup>-</sup> easily
- easily oxidized
- strong reducing agents
- group 16+17

## ACTIVE NONMETALS

- gain e<sup>-</sup> easily
- easily reduced
- strong oxidizing agents
- group 1+2

all chemical reactions can be assigned to one class:

- ① oxidation-reduction (e<sup>-</sup> are transferred)
  - single-replacement, combination, decomposition, combustion
- ② no electron transfer
  - double-replacement, acid-base reactions

always balance equations, 2 methods:

- ① use oxidation-number changes
- ② use 1/2 reactions

## USING OXIDATION-NUMBER CHANGES

start w/ skeleton equation

- ① assign oxidation numbers to all atoms  
(write above their symbols)
- ② identify which are oxidised/reduced
- ③ use bracket lines to connect them
- ④ use coefficients to equalize
- ⑤ make sure they are balanced for both atoms and charge

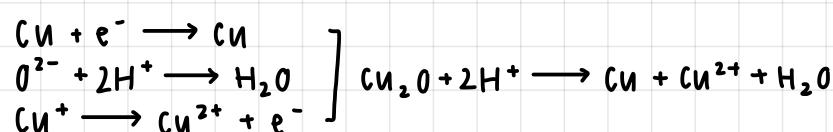
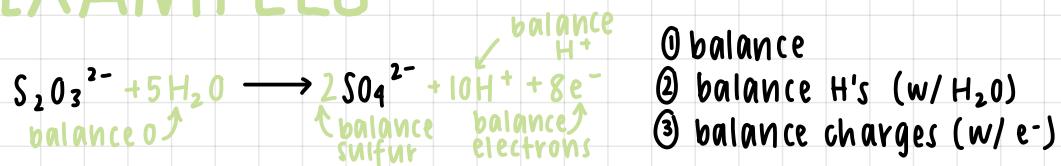
## USING 1/2 REACTIONS

- ① write unbalanced equation in ionic form
- ② write separate 1/2 reactions
- ③ balance 1/2 reactions
- ④ add e<sup>-</sup> to one side of each 1/2 to balance charges
- ⑤ multiply each 1/2 reaction by the e<sup>-</sup> in other 1/2 reaction
- ⑥ combine the 1/2 reactions → overall reaction
- ⑦ add spectator ions to balance (only if asked to include them)

## CHOOSING A BALANCING METHOD

- ① oxidation no. change method
  - works well if the oxidized + reduced species appear only once on each side of the equation (and there are no acids or bases)
- ② the 1/2 reaction method
  - works well for reactions taking place in acidic or alkaline solution

## EXAMPLES



MUST BE ACIDIC - needs H<sup>+</sup> ions

# galvanic cells

## ELECTRON TRANSFER REACTION

oxidation-reduction (redox)

- results in electricity

electrons thru wire