

# Chemical Equilibrium

→ Define Le Chatelier's Principle

↪ It states that whenever there is a change in the condition of the equilibrium, the equilibrium will shift in the opposite direction to oppose the change.

→ Factors affecting the position of equilibrium

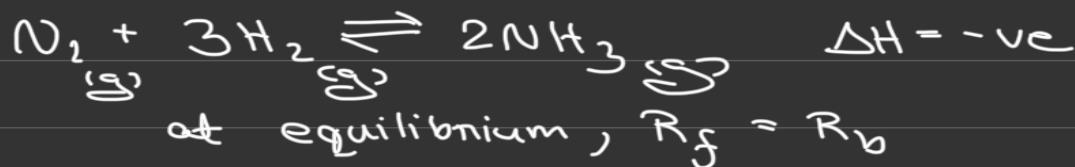
↪ Temperature, Pressure and Concentration

# A catalyst will not affect the position of the equilibrium, it will only affect the rate of the reaction.

→ Define Chemical Equilibrium

↪ It is a stage in a reversible reaction, where the rate of the forward reaction is equal to the rate of the backward reaction.

↪ At equilibrium, the amount of reactants and products remain constant.



→ When temperature is increased:-

↪ Equilibrium shifts to the left

$$R_f \uparrow \quad R_B \uparrow$$

$R_B$  increases more significantly than  $R_f$

$$R_B > R_f$$

→ When temperature is decreased

↪ Equilibrium shifts to the right

$$R_f \downarrow \quad R_B \downarrow$$

$R_f$  decreases more significantly than  $R_B$

$$R_f > R_B$$

→ When pressure is increased

↪ Equilibrium shifts to the right

$$R_f \uparrow \quad R_B \uparrow$$

# when  $A + B_{(g)} \rightleftharpoons C_{(s)}$

only  $R_B$  is affected  
when pressure is changed

$R_f$  increases more significantly than  $R_B$

$$R_f > R_B$$

→ When pressure is decreased

↪ Equilibrium shifts to the left

$$R_f \downarrow \quad R_B \downarrow$$

$R_f$  decreases more significantly than  $R_B$

$$R_B > R_f$$

# Only conc<sup>n</sup> of the reactants will affect the rate of the reaction in that direction

→ When the conc<sup>n</sup> of  $N_2$  is increased

↪ Position of the equilibrium will shift to the right

$$R_f \uparrow \quad R_B =$$

$$R_f > R_B$$

→ When the conc<sup>n</sup> of  $NH_3$  is decreased

↪ Position of the equilibrium will shift to the left

$$R_B \downarrow \quad R_f =$$

$$R_f > R_B$$

Catalyst

↪ At equilibrium  $R_f = R_B$

↪ When catalyst is added,  $R_f \uparrow \quad R_B \uparrow$

$$\therefore R_f = R_B$$

# Equilibrium Constant

→ Equilibrium Constant in terms of conc<sup>n</sup>:  $K_c$

$K_c$  = Product of the conc<sup>n</sup> of products raised to the power given to their stoichiometric equation after reaching equilibrium

Product of the conc<sup>n</sup> of reactants raised to the power given to their stoichiometric equation



$$K_c = \frac{[C]^o \times [D]^p}{[A]^m \times [B]^n}$$



$$K_c = \frac{[NH_3]^2}{[N_2] \times [H_2]^3} = \frac{(mol dm^{-3})^2}{(mol dm^{-3})^4} = mol^{-2} dm^6$$

#  $K_c$  does not exist when there is a solid in the reaction

# Follow the order of the equation given in the question when considering products and reactants in  $K_c$ .

Q:  $H_2 + I_2 \rightleftharpoons 2HI$ , Initially 2 mol of  $H_2$  and 2 mol of  $I_2$  are allowed to react. At equilibrium 1.6 mol of  $H_2$  is present. Find  $K_c$ .



Initial : 2 2 0

Final : 1.6 1.6 0.8

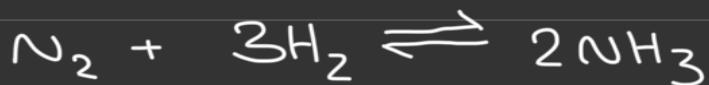
$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$= \frac{\left(\frac{0.8}{V}\right)^2}{\left(\frac{1.6}{V}\right) \left(\frac{1.6}{V}\right)} \frac{(mol dm^{-3})^2}{(mol dm^{-3})^2}$$

0.25

# When  $K_c$  has no unit, no need to work with  $V$ ,  
 the  $V$  will automatically cancel out

Q: Initially 3 mol of  $N_2$  is allowed to react with 8 mol of  $H_2$ . At equilibrium, 0.8 mol of  $NH_3$  is present. Given that all the gases are in a closed vessel of volume  $5\text{dm}^3$ . Find the value of  $K_c$ .



$$I: \quad 3 \quad 8 \quad 0$$

$$F: 2.6 \quad 6.8 \quad 0.8$$

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

$$= \frac{0.8^2}{6.8^3 \times 2.6 \times 0.2^2}$$

$$= 0.0196 \text{ mol}^{-2} \text{ dm}^6$$

Q: A mixture of 2 mol of  $SO_2$ , 1 mol of  $O_2$  and 0.5 mol of  $SO_3$  are left in a closed vessel of volume  $200\text{cm}^3$  to reach equilibrium. At equilibrium, 1.2 mol of  $O_2$  is present. Find the value of  $K_c$  and its unit



$$I: \quad 2 \quad 1 \quad 0.5$$

$$F: \quad 2.4 \quad 1.2 \quad 0.1$$

$$K_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}$$

$$= \frac{0.1^2}{}$$

$$2 \cdot 4^2 \times 1 \cdot 2 \times 5 \\ = 2 \cdot 89 \times 10^{-4} \text{ mol}^{-1} \text{ dm}^3$$

- Assumptions  
(need confirmation)
- Raj bhai imp notes
- Stuff I added  
(confirmed)

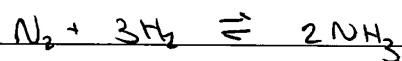
## Equilibrium Constant in terms of Pressure ( $K_p$ )

→ Define Partial Pressure # Only applicable for a mixture of gases

↳ Pressure exerted by a single gas in a gaseous mixture ↗ after reaching equilibrium

↳ Partial Pressure of a gas = molar fraction  $\times$  total pressure

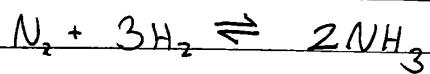
$K_p = \frac{\text{product of the partial pressure of the products raised to the power of their mols in their stoichiometric equation}}{\text{reactants}}$



$$K_p = \frac{(P_{NH_3})^2}{(P_{N_2})(P_{H_2})^3} \quad \frac{atm^2}{atm^4} : \Rightarrow atm^{-2}$$

\*# Other than temperature, no other factor will affect the values of  $K_p$  and  $K_c$  of a reaction

A. Initially 3 mole of  $N_2$  and 8 mole of  $H_2$  are allowed to reach equilibrium. At equilibrium 2.5 mole of  $NH_3$  is present. Given that the total pressure of the system is 5 atm, find  $K_p$ .



	$N_2$	$H_2$	$NH_3$
Initial :	3	8	0
Final :	1.75	4.25	2.5

$$P_{N_2} = \frac{1.75}{8.5} \times 5$$

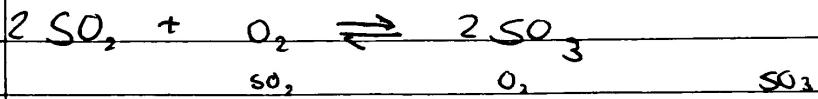
$$= 1.03 \text{ atm}$$

$$P_{H_2} = \frac{4.25}{8.5} \times 5 = 2.5 \text{ atm}$$

$$P_{NH_3} = \frac{2.5}{8.5} \times 5 = 1.47 \text{ atm}$$

$$K_p = \frac{(1.47)^2}{(35/34)(2.5)^3} = 0.134 \text{ atm}^{-2}$$

Q. 1 mol of  $\text{SO}_2$ , 1 mol of  $\text{O}_2$  and 0.2 mols of  $\text{SO}_3$  are placed in a closed vessel to reach equilibrium. At equilibrium 0.8 mol of  $\text{SO}_3$  gas is present. The total pressure of the system is  $6.0 \times 10^5 \text{ Pa}$



Initial :    2              1              0.2

Final :    1.4              0.7              0.8

$$K_p = \frac{P_{\text{SO}_3}^2}{P_{\text{SO}_2} P_{\text{O}_2}} = \frac{0.8^2}{\frac{14}{29} \times 600}$$

$$P_{\text{SO}_2} = \frac{14}{29} \times 600$$

$$P_{\text{SO}_3} = \frac{8}{29} \times 600$$

~~$(3 \times 600)$~~

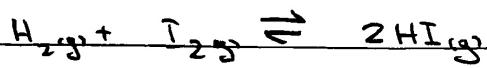
$$(14 \times 600)^2 \times \frac{7 \times 600}{29}$$

$$\underline{0.136 \text{ Pa}^{-1}}$$

$$\underline{0.108 \text{ Pa}^{-1}}$$

$$\underline{\frac{3.75}{0.225} \times 10^{-5} \text{ Pa}^{-1}}$$

Q. 5 mole of  $H_2$  and 4 mole of  $I_2$  are allowed to reach equilibrium. At equilibrium 7 mole of  $HI$  is present. Find  $K_p$ .



T	$H_2$	$I_2$	$HI$
Initial :	5	4	
Final :	1.5	0.5	7

$$K_p = \frac{(7)^2}{(1.5 \times 0.5)} = 65.3 \text{ no unit} \rightarrow \# \text{ Write } \underline{\text{no unit}}$$

if space for  
unit is given

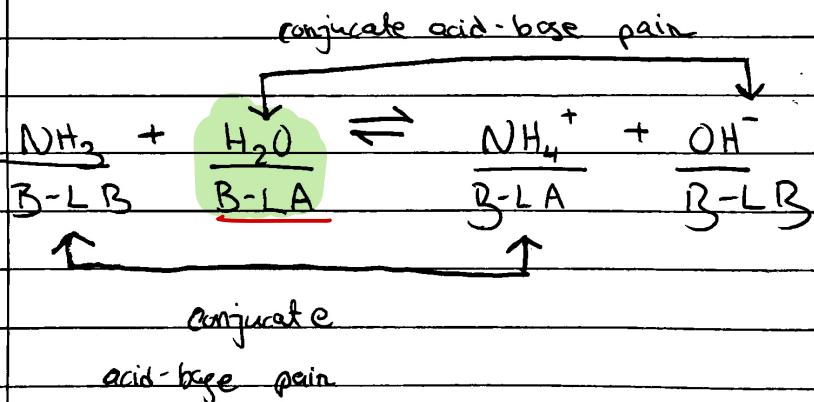
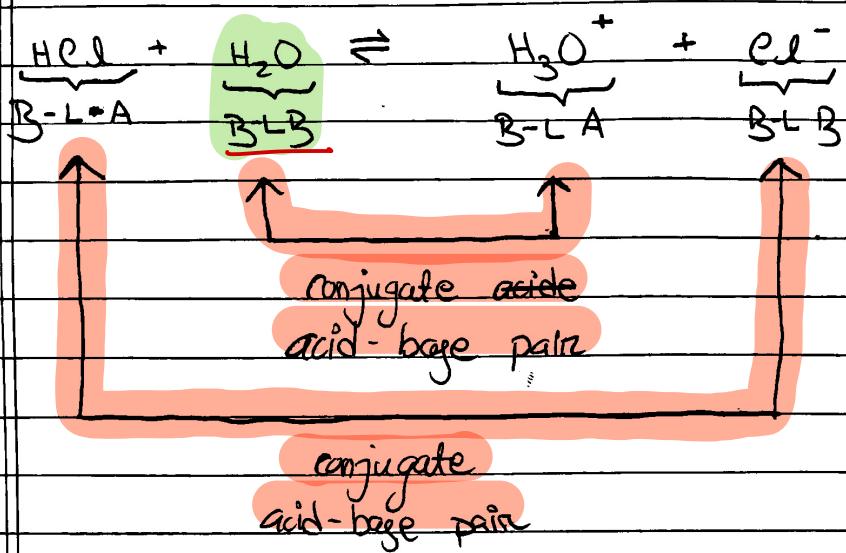
# When Unit gets cancelled out, no need to work with  $P_x$  (Partial Pressure) you can directly use mole.

# Other than Temperature, NO other factors will affect the value of  $K_c$  and  $K_p$ .

## Bronsted Lowry

Bronsted Lowry Theory of Acids and Bases

- It states that an acid is a proton donor ( $H^+$ ) and base is a ~~base~~ proton acceptor.



# In a soln of a weak and strong acid, the weak acid will act as a base IF it has the ability to accept the  $H^+$  ion.

the conjugate base  
of the <sup>weak</sup> acid  
will act  
as the base

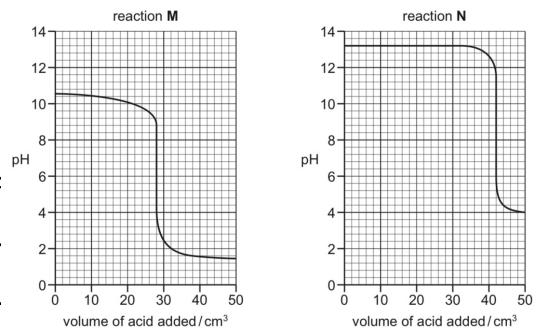
## Titration Curves and Indicators

↪ Indicators are weak acids



↪ Different indicators change colour at different pH. This change occurs over a range of pH, which is known as pH range

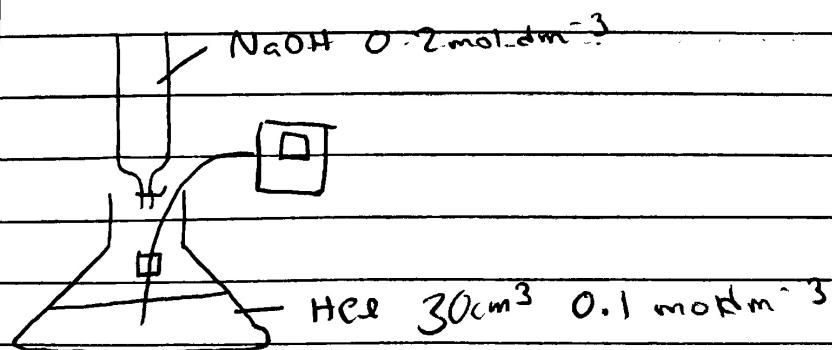
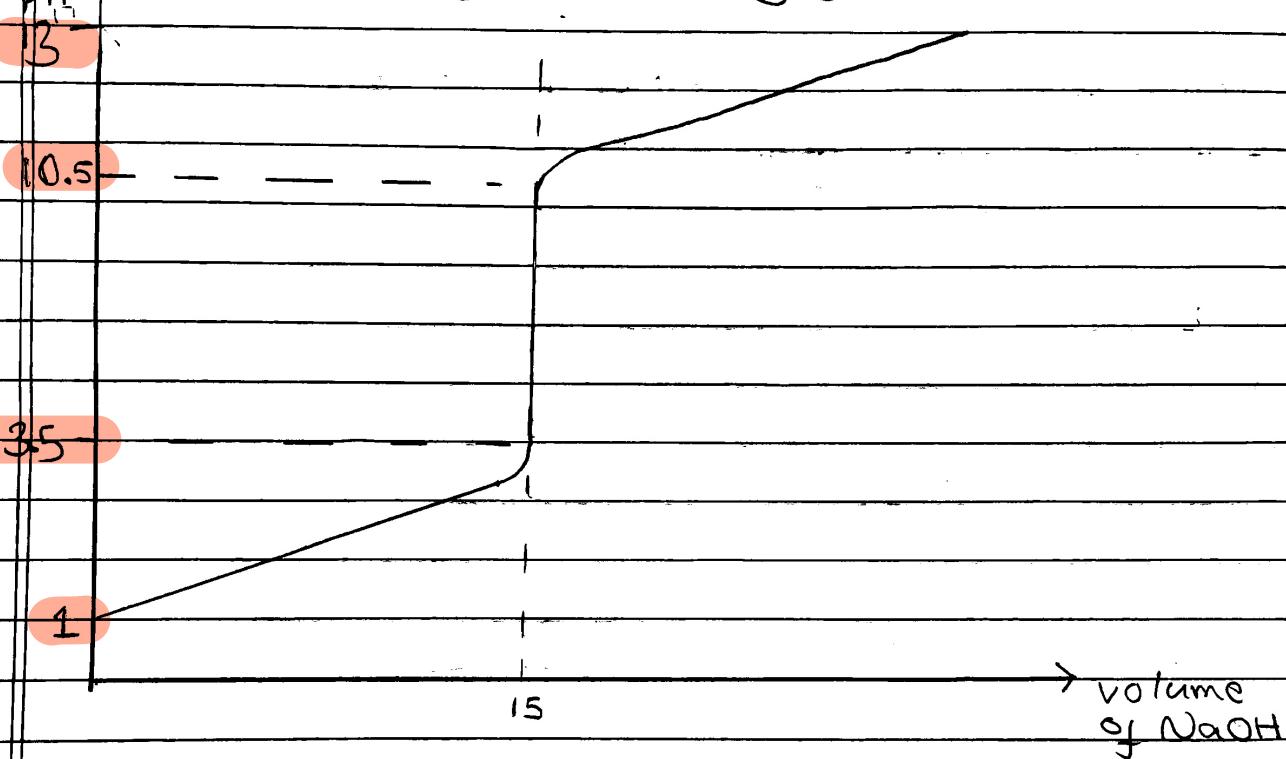
# During the titration of a strong base and a weak acid, we will use an indicator whose pH range is in the alkaline range because if it were in the acidic range, way too much acid would have to be added to the soln to bring about the colour change ~~long ago~~  
far exceeding ~~neutralization~~ volume # Refer to page 15  
and titration graph



### Titration Curve (pH curve)

# What is being titrated will be present in the conical flask.

SA - SB (Strong Acid - Strong Base)

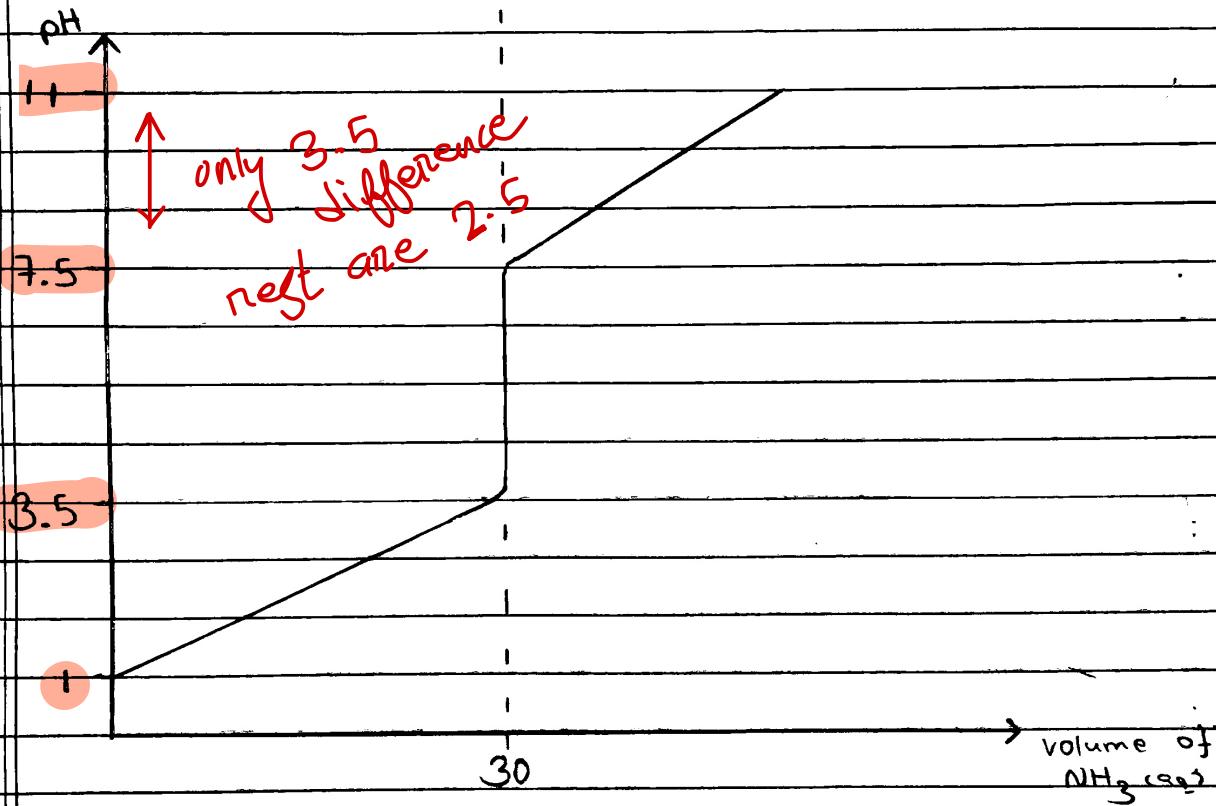


# Unless mentioned otherwise, take the pH of a strong acid as 1.

# Unless mentioned otherwise, take the pH of strong Alkali as 13.

## SA-WB (Strong Acid - Weak Base)

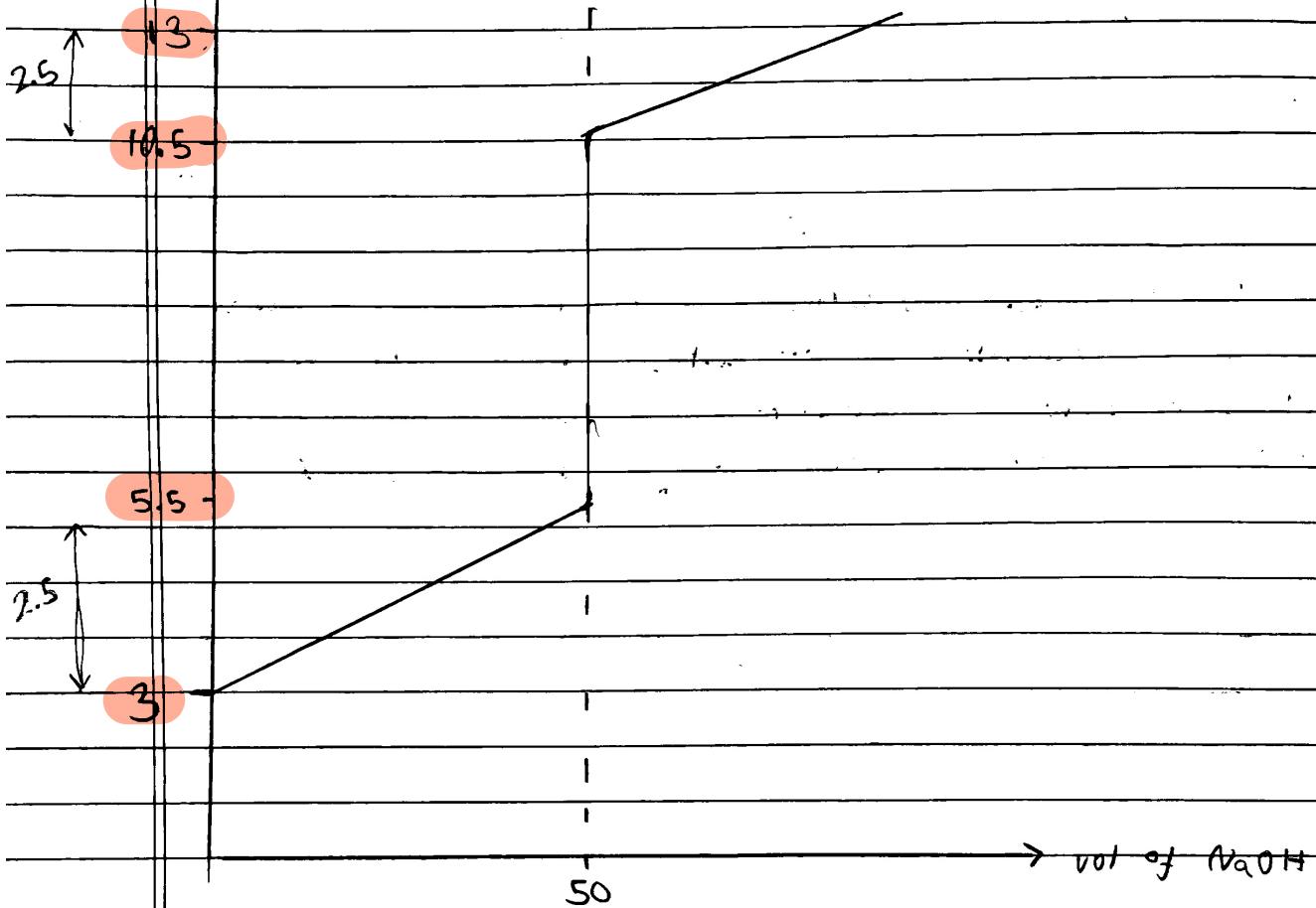
- $30 \text{ cm}^3$  of  $0.1 \text{ mol dm}^{-3}$   $\text{HCl}$  (conical flask)
- $0.1 \text{ mol dm}^{-3}$   $\text{NH}_3$  (burette)



# Unless mentioned otherwise, take the pH of weak alkali to be 11.

## WA - SB (Weak Acid - Strong Base)

- 25 cm<sup>3</sup> of ~~0.2~~ 0.2 mol dm<sup>-3</sup> ethanoic acid
- pH 0.1 mol dm<sup>-3</sup> of NaOH



# Unless mentioned otherwise, take the pH of the weak acid as 3

# In a WA-WB curve, it will not reach the end-point at a specific volume, It reaches the endpoint over a range of volumes

# A WA-WB titration has no sharp Endpoint Volume.

# They We will never be asked to draw a WA-WB titration curve.

# If the pH Range of an Indicator falls within the pH range at the end point volume of the titration then that indicator can be used for that titration

# Does not have to completely overlap, it can only partially overlap and still be suitable

# Misc Notes

## Equilibria

8

Ethyl ethanoate undergoes the following reaction.



Equal amounts of ethanoic acid and ethanol were mixed together and allowed to reach equilibrium.

At equilibrium, the concentrations of both ethanoic acid and ethanol were  $0.42 \text{ mol dm}^{-3}$ .

What is the concentration of ethyl ethanoate at equilibrium?

- A  $0.22 \text{ mol dm}^{-3}$
- B  $0.65 \text{ mol dm}^{-3}$
- C  $0.81 \text{ mol dm}^{-3}$
- D  $1.54 \text{ mol dm}^{-3}$

Paper 1 Variant 2 Summer 2019 I Q10

Answer:

C

→ Here mol of  $\text{H}_2\text{O}$  is equal to mol of  $\text{CH}_3\text{CO}_2\text{C}_2\text{H}_5$ . Thus their conc' also the same.

↳ Since the units cancel out anyways, you can directly work with the mol

# When calculating  $K_c$ , you must account for the conc' of the Water ( $\text{H}_2\text{O}$ ) participating in the reac'.

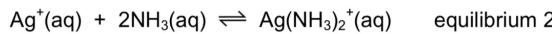
↳ This makes sense of you are finding the conc' of the REACTING Water within the total volume of water in the sol'

#  $\text{HBr(aq)}$  is a strong acid.

14  $\text{AgNO}_3\text{(aq)}$  is added to a solution of a halide ion,  $X^-\text{(aq)}$ , and aqueous ammonia is then added.



The ionic equations for the two reactions that occur are shown.

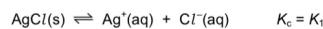


Which statement is correct?

- A The position of equilibrium 1 lies to the left when  $X^- = \text{I}^-$ .
- B Increasing the concentration of ammonia causes the position of equilibrium 1 to move to the left.
- C  $K_c$  for equilibrium 2 is larger when  $X^- = \text{Cl}^-$  than when  $X^- = \text{I}^-$ .
- D Equilibrium 2 is a redox reaction



\* 16 Silver chloride and silver iodide form equilibria when added to water.



Each equilibrium position lies well to the left.

Silver iodide will not dissolve in aqueous ammonia. Silver chloride will dissolve in aqueous ammonia. Another equilibrium is formed.



The position of this equilibrium lies to the right.

What is the order of magnitude for these three equilibrium constants?

	smallest	→	largest
A	$K_3$	$K_2$	$K_1$
B	$K_3$	$K_1$	$K_2$
c	$K_2$	$K_1$	$K_3$
D	$K_1$	$K_2$	$K_3$



32 Four solutions, each of concentration  $0.1 \text{ mol dm}^{-3}$ , were tested with a pH meter. The results are shown.

solution	formula of acid or base	pH
acid 1	$\text{CH}_3\text{CO}_2\text{H}$	4
acid 2	$\text{HNO}_3$	1
base 1	$\text{CH}_3\text{NH}_2$	11
base 2	$\text{NaOH}$	14

C

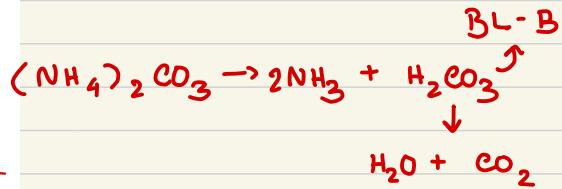
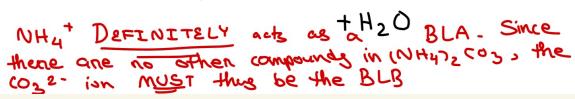
Which statements explain these results?

- 1 Acid 2 has a lower pH than acid 1 because it is more soluble.
- 2 Base 2 has a higher concentration of hydroxide ions in solution than base 1. ✓
- 3 Acid 1 dissociates less than acid 2. ✓

- 15 Ammonium carbonate is a crystalline solid. On gentle warming a reaction occurs, forming ammonia as one product.

How are the carbonate ions behaving during this reaction?

- A Brønsted-Lowry acid  
✓ B Brønsted-Lowry base  
✓ C oxidising agent  
D reducing agent



# The anions of acids always act as BLB.

# To obtain an acid from a sol<sup>n</sup> of its salt, add (react it with) a strong acid. The anion of the acid present in the salt accepts the  $\text{H}^+$  ion from the strong acid and forms  $\text{H}_2\text{O}$  reverting back to its acid.

→ not sure



↳ Dibasic acids lose 1  $\text{H}^+$  ion first, losing the next  $\text{H}^+$  ion taking a lot of energy as the  $\text{H}^+$  ion has to be removed from a Positively charged ion.



Indicator	Colour in acid solution	Colour in alkaline solution
Blue litmus	Red	Blue
Methyl orange	Pink	Yellow
Thymolphthalein	Colourless	Blue
Red litmus	Red	Blue



19 Sulfur dioxide,  $\text{SO}_2$ , reacts with calcium hydroxide in aqueous solution.

What is the main product that is first formed?

- A  $\text{Ca}(\text{HSO}_4)_2$     B  $\text{CaS}$     C  $\text{CaSO}_3$     D  $\text{CaSO}_4$



## # 2 features of a reaction in DYNAMIC EQUILIBRIUM

M1 equal rates of forward and backward reactions

M2 closed system OR macroscopic properties unchanging

$$K_p = \frac{P_{\text{O}_2}}{P_{\text{Cl}_2}^2}$$

At  $1.00 \times 10^5 \text{ Pa}$  and  $500 \text{ K}$ , 70% of the initial amount of  $\text{Cl}_2(\text{g})$  has reacted.

Calculate  $K_p$  and state its units.

$\text{Cl}_2$	$\text{O}_2$
$0.3x$	$0.35x$
46000	54000

$$\frac{0.3}{0.65} \times 10^{10000}$$

$$K_p = \frac{54000}{46000^2}$$
$$= 2.53 \times 10^{-5}$$

$$K_p = 2.53 \times 10^{-5}$$

units =  $\text{Pa}^{-1}$

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## # Amphoteric means reacting with both acids and BASES.

When we increase the pressure, the reaction goes in the direction that results in fewer molecules of gas being formed.