I am able to

1 Define a Brønsted-Lowry acid — A proton (H^+) donor

2 Define a Brønsted-Lowry base — A proton (H^+) acceptor

3 Classify each of the species in the following reactions as Brønsted-Lowry acids and bases and identify conjugate acid-base pairs:

acids bases	conjugate pairs	
$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$	$CH_3COOH(aq)$ $CH_3COO^-(aq)$	$H_2O(l)$ $H_3O^+(aq)$
$CH_3NH_2(aq) + H_2O(aq) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$	$CH_3NH_2(aq)$ $CH_3NH_3^+(aq)$	$H_2O(aq)$ $OH^-(aq)$
$NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(aq) + H_2O(l)$	$NH_4^+(aq)$ $NH_3(aq)$	$OH^-(aq)$ $H_2O(l)$

4 Identify the conjugate base of each of the following:

H_2SO_4	HCO_3^-	HPO_4^{2-}
HSO_4^-	CO_3^{2-}	PO_4^{3-}

5 Identify the conjugate acid of each of the following:

$HCOO^-$	$H_2PO_4^-$	SO_4^{2-}
$HCOOH$	H_3PO_4	HSO_4^-

6 Explain the difference between the terms *amphoteric* and *amphiprotic*

Amphoteric refers to any species that can act as an acid and a base
Amphiprotic refers specifically to a species that can donate and receive a proton (H^+) [Brønsted-Lowry definition]

7 Identify which of the following is/are amphiprotic?

HCO_3^- NO_3^- H_2O $H_2PO_4^-$ NH_4^+ OCl^-

8 Complete the following equations:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O$

$NH_3 + HCl \rightarrow NH_4Cl$

$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$

$NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$

$2CH_3COOH + Ca \rightarrow (CH_3COO)_2Ca + H_2$

9 State whether the neutralisation reaction between an acid and an alkali is exothermic or endothermic.

EXOTHERMIC

10 State the names of the acid and alkali needed to make each of the following salts using titration:

Potassium chloride Sodium bromide ammonium nitrate potassium sulfate

Potassium chloride	hydrochloric acid + potassium hydroxide
Sodium bromide	hydrobromic acid + sodium hydroxide
ammonium nitrate	nitric acid + ammonia
potassium sulfate	sulfuric acid + potassium hydroxide

11 State the names of an acid and another substance that could be reacted to make each of the following salts:

Copper(II) nitrate calcium chloride barium chloride magnesium sulfate

	acid	**other substance**
Copper(II) nitrate	nitric acid	copper(II) oxide/hydroxide/carbonate
calcium chloride	hydrochloric acid	Calcium Calcium oxide/hydroxide/carbonate/hydrogencarbonate
barium chloride	hydrochloric acid	Barium Barium oxide/hydroxide/carbonate/hydrogencarbonate
magnesium sulfate	sulfuric acid	Magnesium Magnesium oxide/hydroxide/carbonate/hydrogencarbonate

12 Explain the difference between a strong acid and a weak acid.

Strong acid: dissociates completely in aqueous solution
Weak acid: dissociates partially in aqueous solution

13 Classify the following acids and bases as strong or weak:

ACIDS	BASES
Hydrochloric acid	Sodium hydroxide
Ethanoic acid	Ammonia
Sulfuric acid *(for first dissociation)*	CH_3NH_2
Carbonic acid	Potassium hydroxide
Nitric acid	Barium hydroxide

14 Write equations for the dissociation of the following acids:

CH_3COOH, HCl, H_2SO_4

$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$ ***or*** $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ ***or*** $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$ ***or*** $H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$

HSO_4^- can then dissociate further:
$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$ ***or*** $HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$

15 Arrange the following in order of decreasing conductivity of solutions of equal concentration:

CH_3COOH, HCl, H_2SO_4, H_2O

$H_2SO_4 > HCl > CH_3COOH > H_2O$

16 Write equations for the ionisation of the following bases:

$NaOH$, NH_3, $Ba(OH)_2$

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$ $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$

17 Arrange the following in order of increasing conductivity of solutions of equal concentration:

$NaOH$, NH_3, $Ba(OH)_2$

$NH_3 < NaOH < Ba(OH)_2$

18 Explain which of the following solutions of equal concentration conducts electricity better:

CH_3COOH or CH_3COONa
NH_3 or NH_4Cl

CH_3COONa conducts electricity better – fully ionised in solution as it is an ionic salt
CH_3COOH is a weak acid – only partially ionised
CH_3COONa – higher concentration of ions, therefore better conductor.

NH_4Cl conducts electricity better – fully ionised in solution as it is an ionic salt
NH_3 is a weak base – only partially ionised
NH_4Cl – higher concentration of ions, therefore better conductor.

19 Explain two experimental procedures to distinguish between strong and weak acids and bases

Test solutions of equal concentrations

Measure electrical conductivity (use a conductivity meter) – strong acids/bases fully ionised (strong electrolytes) but weak acids/bases only partially ionised (weak electrolytes). Strong acids/bases have higher concentration of ions, therefore conduct electricity better.

Measure pH (use a pH meter) – strong acids fully ionised but weak acids only partially ionised. Strong acids have higher concentration of H^+ ions, therefore lower pH.
Strong bases fully ionised but weak bases only partially ionised. Strong bases have higher concentration of OH^- ions (lower concentration of H^+ ions), therefore higher pH.

React acids with magnesium/calcium carbonate – strong acids fully ionised but weak acids only partially ionised. Strong acids have higher concentration of H^+ ions, therefore react more vigorously with $Mg/CaCO_3$.

20 Complete the following sentences:

A strong acid has a WEAK conjugate BASE

A strong base has a WEAK conjugate ACID

21 If methanoic acid is a stronger acid than ethanoic acid, explain whether the methanoate ion or the ethanoate ion is the stronger base.

Ethanoate ion is the stronger base – the weaker the acid the stronger the conjugate base.

$HA(aq) \rightleftharpoons A^-(aq) + H^+(aq)$

A weaker acid has less tendency to dissociate (position of above equilibrium lies more to left) therefore the conjugate base has greater tendency to pick up proton, i.e. it is stronger.

22 Define pH

$pH=-\log_{10}[H^+(aq)]$

23 Complete the following table without using a calculator:

pH	$[H^+(aq)]$/mol dm^{-3}	acidic/alkaline/neutral?
4	1×10^{-4}	acidic
6.0	1.0×10^{-6}	acidic
7	1×10^{-7}	neutral
11	1.0×10^{-11}	alkaline
13	1.0×10^{-13}	alkaline

24 Calculate the pH for each of the following solutions: 0.1 M HCl(aq) 0.020 M HNO_3(aq)

Both strong acids 0.1 M HCl(aq) pH=1 0.020 M HNO_3(aq) pH=1.7

25 Deduce by how much does the pH changes when a solution of a strong acid with pH=3 is diluted by a factor of 10.

Increases by 1

26 Calculate the relationship between the $[H^+(aq)]$ in a solution of pH=2 and one of pH=6?

6-2=4

Solution of pH=2 has 10^4 times the $[H^+(aq)]$ as the pH=6 solution

27 State two methods for measuring the pH of a solution.

Using a pH meter
Using universal indicator and comparing colour to a colour chart.

28 Write an expression for the equilibrium that exists in any aqueous solution and for the ionic product constant of water.

$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ $K_w = [H^+(aq)][OH^-(aq)]$

29 State the value of K_w at 25 °C? 1.0×10^{-14}

30 Complete the following table for aqueous solutions at 25 °C:

$[H^+(aq)]$/mol dm^{-3}	$[OH^-(aq)]$/mol dm^{-3}	acidic/alkaline/neutral?
1.0×10^{-6}	1.0×10^{-8}	acidic
1.0×10^{-11}	1.0×10^{-3}	alkaline
2.5×10^{-5}	4.0×10^{-10}	acidic
8.3×10^{-4}	1.2×10^{-11}	acidic
3.6×10^{-12}	2.8×10^{-3}	alkaline

31 Calculate $[H^+(aq)]$ and $[OH^-_{(aq)}]$ for each of the following solutions:

0.1 M HCl, 0.01 M NaOH, 0.020 M $Ba(OH)_2$.

	$[H^+(aq)]$ / mol dm^{-3}	**$[OH^-_{(aq)}]$ / mol dm^{-3}**
0.1 M HCl	0.1	1×10^{-13}
0.01 M NaOH	1×10^{-12}	0.01
0.020 M $Ba(OH)_2$	2.5×10^{-13}	0.040

32 Write expressions for K_a or K_b (as appropriate) for each of the following reactions:

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$	$K_a = \frac{[H_3O^+(aq)][CH_3COO^-(aq)]}{[CH_3COOH(aq)]}$
$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	$K_b = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(aq)]}$
$HCOOH(aq) \rightleftharpoons H^+(aq) + HCOO^-(aq)$	$K_a = \frac{[H^+(aq)][HCOO^-(aq)]}{[HCOOH(aq)]}$
$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$	$K_b = \frac{[C_2H_5NH_3^+(aq)][OH^-(aq)]}{[C_2H_5NH_2(aq)]}$

33 Calculate the value of K_a of HA given that the pH of a 0.0100 M solution of HA is 5.20.

HA is a weak acid (pH if it were a strong acid would be 2)

$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]}$$

pH=5.20 $[H^+(aq)]=10^{-5.20} = 6.31\times10^{-6}$ mol dm^{-3}

$$K_a = \frac{[6.31\times10^{-6}][6.31\times10^{-6}]}{[0.0100-6.31\times10^{-6}]} = 3.98\times10^{-9}$$

34 Define pK_a and pK_b $pK_a = -\log_{10}K_a$ $pK_b = -\log_{10}K_b$

35 Arrange the following acids in order of decreasing acid strength (strongest first)

ACID	K_a
Propanoic	1.3×10^{-5}
Nitrous (HNO_2)	4.7×10^{-4}
Hydrofluoric (HF)	5.6×10^{-4}
Chloric(I)	3.7×10^{-8}

Hydrofluoric (HF) > Nitrous (HNO_2) > Propanoic > Chloric(I)

36 Arrange the following acids in order of decreasing acid strength (strongest first)

ACID	pK_a
Ethanoic (CH_3COOH)	4.8
Carbonic (H_2CO_3)	6.4
Benzoic (C_6H_5COOH)	4.2
Ammonium Ion (NH_4^+)	9.3
Phosphoric(V) (H_3PO_4)	2.1

Phosphoric(V) > Benzoic > Ethanoic > Carbonic > Ammonium Ion

37 Calculate the pH for each of the following weak acids. State any assumptions you make?

ACID	pK_a	concentration / mol dm^{-3}	pH
HA_1	4.60	0.100	2.80
HA_2	5.24	0.0250	3.42
HA_3	7.40	0.0640	4.30

$HA_1(aq) \rightleftharpoons H^+(aq) + A_1^-(aq)$ $K_a = 10^{-4.60} = 2.51\times10^{-5}$

$$K_a = \frac{[H^+(aq)][A_1^-(aq)]}{[HA_1(aq)]}$$

Assume that dissociation of the acid is negligible compared to its concentration

$$2.51\times10^{-5} = \frac{[H^+(aq)]^2}{[0.100]}$$

$[H^+(aq)] = 1.58\times10^{-3}$ mol dm^{-3}

$pH = -\log_{10}[H^+(aq)] = -\log_{10} 1.58\times10^{-3} = 2.80$

38 Define pOH and pK_w

$pOH = -\log_{10}[OH^-(aq)]$ $pK_w = -\log_{10}K_w$

39 Calculate the pH for each of the following weak bases. State any assumptions you make?

BASE	pK_b	concentration / mol dm^{-3}	pH
B_1	4.20	0.100	11.40
B_2	6.46	0.0200	9.92
B_3	8.80	0.0400	8.90

$B_1(aq) + H_2O(l) \rightleftharpoons B_1H^+(aq) + OH^-(aq)$ $K_b = 10^{-4.20} = 6.31\times10^{-5}$

$$K_b = \frac{[B_1H^+(aq)][OH^-(aq)]}{[B_1(aq)]}$$

Assume that ionisation of the base is negligible compared to its concentration

$$6.31\times10^{-5} = \frac{[OH^-(aq)]^2}{[0.100]}$$

$[OH^-(aq)] = 2.51\times10^{-3}$ mol dm^{-3}

$pOH = -\log_{10}[OH^-(aq)] = -\log_{10} 2.51\times10^{-3} = 2.60$

$pK_w = pH + pOH$

$pH = 14-2.60 = 11.40$ Assume that temperature is 25 °C

40 State the relationship between K_a and K_b for a conjugate acid-base pair.

$K_a \times K_b = K_w$

41 State the relationship between pK_a and pK_b for a conjugate acid-base pair.

$pK_a + pK_b = pK_w$

42 Using the pK_a values above, work out the pK_b values of the following bases at 25°C.

BASE	pK_b
CH_3COO^-	14 - 4.8 = 9.2
HCO_3^-	14 – 6.4 = 7.6
$C_6H_5COO^-$	14 – 4.2 = 9.8
NH_3	14 – 9.3 = 4.7

CH_3COO^- pK_a for CH_3COOH is 4.8 $pK_b = 14-4.8 = 9.2$

43 Calculate the pH of pure water at 50°C given that the value of K_w at 50 °C is 5.48×10^{-14}.

$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ $K_w = [H^+(aq)][OH^-(aq)]$

Pure water: $[H^+(aq)] = [OH^-(aq)]$ $5.48\times10^{-14} = [H^+(aq)]^2$ $[H^+(aq)] = 2.34\times10^{-7}$ mol dm^{-3}

$pH = -\log_{10} 2.34\times10^{-7} = 6.63$

44 Explain whether a solution with pH=6.8 is acidic, alkaline or neutral at 50 °C

Alkaline as neutral pH is 6.63 at this temperature (question 43) and the solution has a higher pH than 6.63.

45 Calculate the pH of 0.100 mol dm^{-3} NaOH at 50°C given that the value of pK_w at 50 °C is 13.3.

Strong base, therefore $[OH^-(aq)] = 0.100$ mol dm^{-3}

$pOH = -\log_{10} 0.100 = 1$

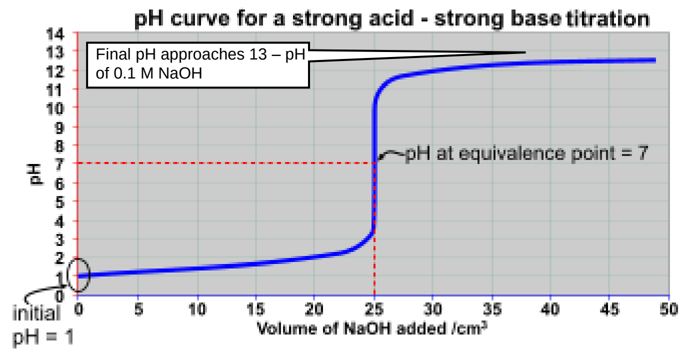
$pK_w = pH + pOH$

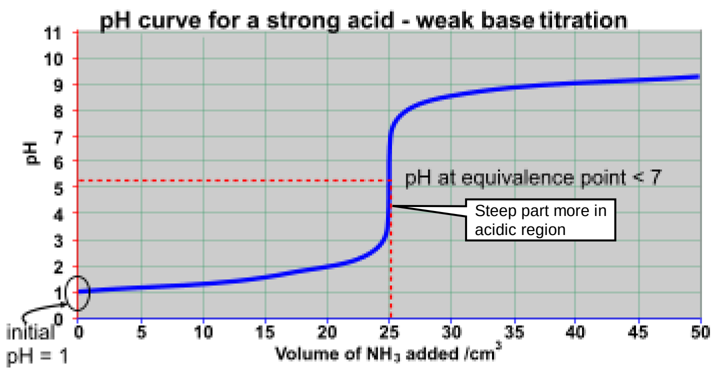
$13.3 = pH + 1$ $pH = 12.3$

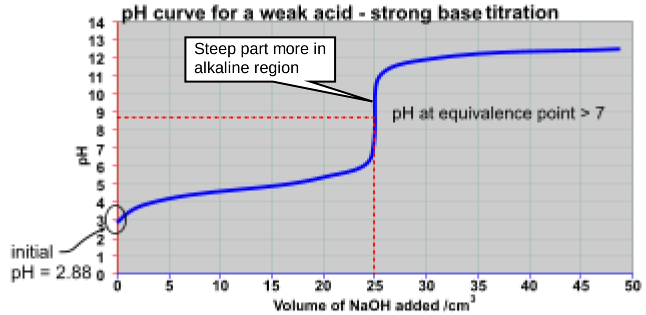
46 Sketch titration curves for the following titration combinations:

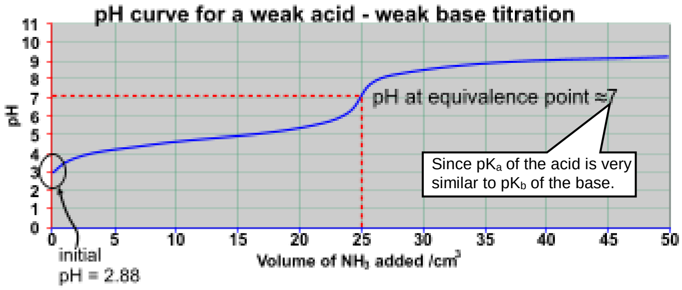
Strong acid - Strong base [25 cm^3 0.100 mol dm^{-3} HCl and 0.100 mol dm^{-3} NaOH]
Strong acid - Weak base [25 cm^3 0.100 mol dm^{-3} HCl and 0.100 mol dm^{-3} NH_3 (pK_b=4.75)]
Weak acid - Strong base [25 cm^3 0.100 mol dm^{-3} CH_3COOH (pK_a=4.76) and 0.100 mol dm^{-3} NH_3]
Weak acid - Weak base [25 cm^3 0.100 mol dm^{-3} CH_3COOH (pK_a=4.76) and 0.100 mol dm^{-3} NH_3 (pK_b=4.75)]

In each case calculate the initial pH, state whether the pH at the equivalence point is less than 7, equal to 7 or greater than 7



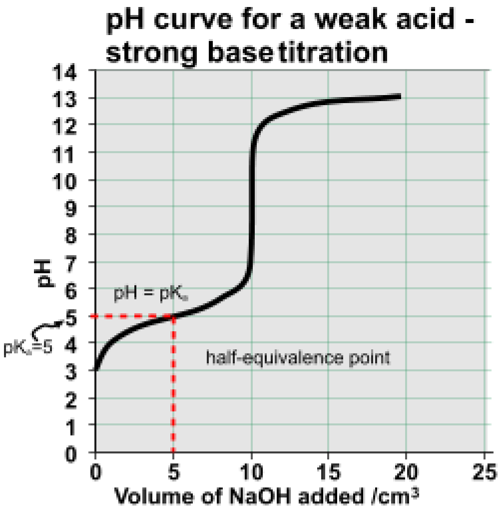
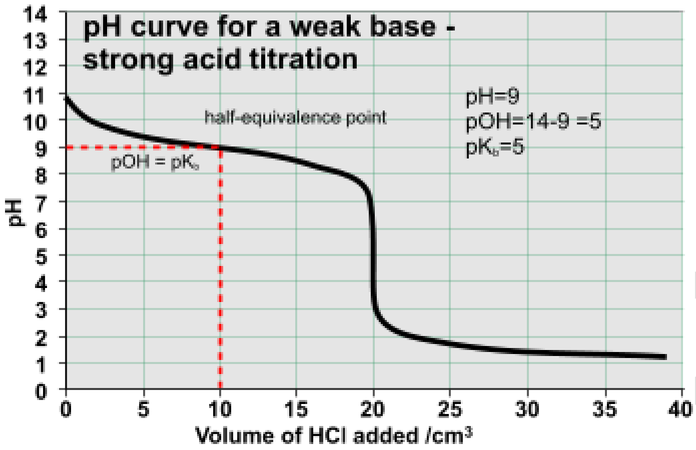






47 Mark on curve I the point at which pH=pK_a of the weak acid and deduce pK_a of the weak acid.

Mark on curve II the point at which pOH=pK_b of the weak base and deduce pK_b of the weak base.

	
curve I **pK_a=5**	**curve II** **pK_b=5**

48 Calculate the pH when 10.0 cm^3 of 0.100 mol dm^{-3} NaOH is added to 20.0 cm^3 of 0.100 mol dm^{-3} weak acid HA (pK_a=6)

Half-equivalence point: pH=pK_a pH=6

Calculate the pH when 10.0 cm^3 of 0.100 mol dm^{-3} HCl is added to 20.0 cm^3 of 0.100 mol dm^{-3} of weak base B (pK_b=4)

Half-equivalence point: pOH=pK_b pOH=4 pH=14-4 = 10

49 Explain how an acid-base indicator works.

Indicator is a weak acid or a weak base

Indicator that is a weak acid (HIn):

$HIn \rightleftharpoons H^+ + In^-$
Colour I Colour II
HIn and In^- are different colours
Add acid (H^+) - position of equilibrium shifts to left (to use up H^+) – indicator appears Colour I
Add alkali (OH^-) – OH^- reacts with H^+ - position of equilibrium shifts to right (to replace H^+) – indicator appears Colour II

Indicator that is a weak base (In^-):

$In^- + H_2O \rightleftharpoons HIn + OH^-$
Colour I Colour II
HIn and In^- are different colours
Add alkali (OH^-) - position of equilibrium shifts to left (to use up OH^-) – indicator appears Colour I
Add acid (H^+) – H^+ reacts with OH^- - position of equilibrium shifts to right (to replace OH^-) – indicator appears Colour II

50 Explain what is meant by the 'pH range of an indicator'.

Range of pH values where intermediate (between the 'acid' and 'alkali') colours of the indicator are seen

51 Explain the connection between the pH range of an indicator that is a weak acid and the value of pK_a for the indicator?

pK_a of the indicator is the mid-point of the pH range ($[HIn]=[In^-]$, $pH=pK_a$)
pH range is approximately $pK_a \pm 1$

52 From the following list of indicators select appropriate ones for each of the titrations in 46.

Indicator	pK_a	pH range
Methyl Yellow	3.5	2.9-4.0
Methyl Red	5.1	4.2-6.3
Phenol Red	7.9	6.8-8.4
Phenolphthalein	9.3	8.2-10.0

Strong acid - Strong base	methyl red/phenol red/phenolphthalein
Strong acid - Weak base	methyl red
Weak acid - Strong base	phenolphthalein
Weak acid - Weak base	no indicator – pH change is gradual

53 Explain what you understand by a *buffer solution*.

A solution that resists change in pH
when **small** amounts of acid/alkali are added.

54 State the composition of an acidic buffer and a basic buffer

Acidic buffer: weak acid and its salt (conjugate base) e.g. ethanoic acid and sodium ethanoate
Basic buffer: weak base and its salt (conjugate acid) e.g. ammonia and ammonium chloride

55 Explain how your chosen buffer will work when small amounts of acid or alkali are added.

Acidic buffer (CH_3COOH/CH_3COO^-)

Equilibrium present in buffer solution: $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ Equilibrium 1

Add acid (H^+): $CH_3COO^- + H^+ \rightarrow CH_3COOH$

The conjugate base reacts with any acid added. If the concentrations of CH_3COO^- and CH_3COOH are large compared to the amount of H^+ added the percentage change in their concentrations is very small and the position of Equilibrium 1 shifts very little, therefore the pH stay approximately the same.

Add alkali (OH^-): $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

The acid reacts with any alkali added. If concentrations of CH_3COO^- and CH_3COOH are large compared to the amount of OH^- added the percentage change in their concentrations is very small and the position of Equilibrium 1 shifts very little, therefore the pH stay approximately the same.

Basic buffer (NH_3/NH_4^+)

Equilibrium present in buffer solution: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ Equilibrium 2

Add acid (H^+): $NH_3 + H^+ \rightarrow NH_4^+$

The base reacts with any acid added. If the concentrations of NH_3 and NH_4^+ are large compared to the amount of H^+ added the percentage change in their concentrations is very small and the position of Equilibrium 2 shifts very little, therefore the pH stay approximately the same ($[H^+]$ and, hence pH, is related to $[OH^-]$).

Add alkali (OH^-): $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

The conjugate acid reacts with any alkali added. If the concentrations of NH_3 and NH_4^+ are large compared to the amount of OH^- added the percentage change in their concentrations is very small and the position of Equilibrium 2 shifts very little, therefore the pH stay approximately the same.

56 Explain which of the following mixtures will result in a buffer solution:

I 25.0 cm^3 0.100 mol dm^{-3} NaOH and 40.0 cm^3 0.100 mol dm^{-3} CH_3COOH
II 25.0 cm^3 0.100 mol dm^{-3} NH_3 and 50.0 cm^3 0.100 mol dm^{-3} HCl
III 25.0 cm^3 0.200 mol dm^{-3} NaOH and 50.0 cm^3 0.100 mol dm^{-3} CH_3COOH
IV 20.0 cm^3 0.200 mol dm^{-3} NH_3 and 30.0 cm^3 0.100 mol dm^{-3} HCl

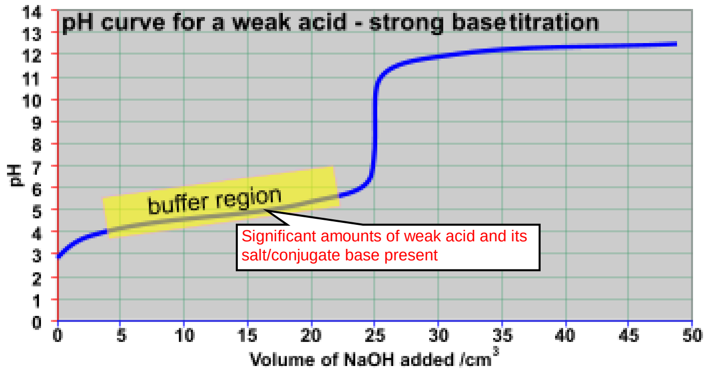
I NaOH reacts with some of the CH_3COOH: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
Not all CH_3COOH reacts as excess CH_3COOH present
Solution contains significant amounts of CH_3COOH (weak acid) and CH_3COO^- (conjugate base/salt), therefore a buffer.

IV HCl reacts with some of the NH_3: $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$
Not all NH_3 reacts as excess NH_3 present
Solution contains significant amounts of NH_3 (weak base) and NH_4^+ (conjugate acid/salt), therefore a buffer.

II Not a buffer because HCl in excess and so no ammonia (weak base) left.

III Not a buffer because equal numbers of moles of NaOH and so all CH_3COOH converted into its salt – no weak acid present.

57 Mark on your pH curve for a weak acid-strong base titration in 46 the *buffer region*.



buffer region

Significant amounts of weak acid and its salt/conjugate base present

58 Predict whether solutions of the following salts will be acidic, alkaline or neutral:
NaCl, NH_4Cl, KNO_3, Na_2CO_3, CH_3COONH_4

	salt of		**acidic, alkaline or neutral?**
NaCl	strong base (NaOH)	strong acid (HCl)	neutral
NH_4Cl	weak base (NH_3)	strong acid (HCl)	acidic
KNO_3	strong base (KOH)	strong acid (HNO_3)	neutral
Na_2CO_3	strong base (NaOH)	weak acid (H_2CO_3)	alkaline
CH_3COONH_4	weak base (NH_3)	weak acid (CH_3COOH)	approximately neutral

The pK_b of NH_3 is approximately the same as the pK_a of ethanoic acid – therefore a solution of ammonium ethanoate will be approximately neutral.

59 Explain using equations why a solution of sodium ethanoate has a pH > 7.

In solution ionises as: $Na^+(aq)$ and $CH_3COO^-(aq)$

CH_3COO^- is the conjugate base of a weak acid and so is a reasonably good base.

$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

Extra OH^- produced, therefore solution is alkaline (pH>7).

[NaOH is a strong base, therefore it is fully ionised, therefore no tendency for Na^+ to re-join with OH^- to reform NaOH and remove the extra OH^- produced]

60 Explain using equations why a solution of $CH_3NH_3NO_3$ is acidic.

In solution ionises as: $NO_3^-(aq)$ and $CH_3NH_3^+(aq)$

$CH_3NH_3^+$ is the conjugate acid of a weak base (CH_3NH_2) and so is a reasonably good acid.

$CH_3NH_3^+(aq) \rightleftharpoons CH_3NH_2(aq) + H^+(aq)$

Extra H^+ produced, therefore solution is acidic (pH<7).

[HNO_3 is a strong acid, therefore it is fully ionised, therefore no tendency for NO_3^- to re-join with the extra H^+ produced to remove it]

61 Explain why $FeCl_3(aq)$ is acidic

Fe^{3+} forms a complex ion in solution: $[Fe(H_2O)_6]^{3+}$

Fe^{3+} highly charged (large charge density) so water polarised very strongly so H^+ fairly readily lost

$[Fe(H_2O)_6]^{3+}(aq) \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+}(aq) + H^+(aq)$

62 State the Lewis definition of an acid and a base.

Acid: electron pair acceptor

Base: electron pair donor

63 State what type of bond is formed when a Lewis acid reacts with a Lewis base.

Coordinate covalent bond / dative covalent bond

64 Classify each of the following as a 🟨 Lewis Acid or a 🟩 Lewis Base:

🟨 H^+ 🟩 NH_3 🟨 NO_2^+ 🟩 CO 🟩 H_2O 🟨 BF_3 🟨 Fe^{2+}

Lewis acids all have space in their outer shell to accept a pair of electrons.

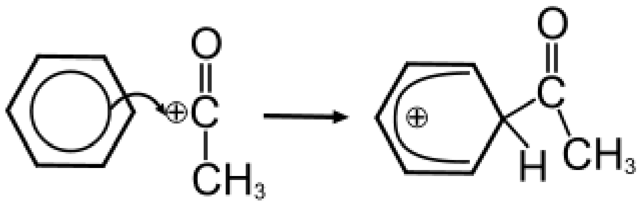
Lewis bases all have a lone pair that they can donate.

65 Classify Lewis acids/bases as electrophiles/nucleophiles

Lewis acid – electrophile – electrophiles accept a pair of electrons

Lewis base – nucleophile – nucleophiles donate a pair of electrons

66 Explain whether H_3CCO^+ acts as a Lewis acid or Lewis base in the following step from an organic mechanism



Curly arrow shows the donation of a pair of electrons from benzene to H_3CCO^+ therefore H_3CCO^+ is a Lewis acid because it accepts a pair of electrons.

67 Explain whether $AlCl_3$ is a Lewis acid or Lewis base in the following reaction:

$AlCl_3 + Cl^- \rightarrow AlCl_4^-$

$AlCl_3$ is a Lewis acid – accepts a pair of electrons from the Cl^-

68 Explain why rain is naturally acidic and state its approximate pH.

Acidic due to dissolved carbon dioxide, which is in equilibrium with carbonic acid

$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

H_2CO_3 is a weak acid: $H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$

Unpolluted rain will have a pH of about 5.6

69 Explain what is meant by *acid deposition*

Any process by which acidic substances leave the atmosphere and are deposited on the surface of the Earth. E.g. acid rain

70 Identify 2 gases that can result in acid deposition and write an equation for the formation of each

SO_2 and NO

$S(s) + O_2(g) \rightarrow SO_2(g)$
$N_2(g) + O_2(g) \rightarrow 2NO(g)$

71 State one source of each of the gases in 70

SO_2 can come from burning fuels, such as coal, that contain sulfur / also formed in volcanic eruptions

NO is formed in internal combustion engines – at the very high temperatures in the engine, nitrogen and oxygen from the air react together / NO is also formed by lightning.

72 Write equations to show how HNO_2, HNO_3, H_2SO_3 and H_2SO_4 can be formed in the atmosphere.

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$
$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

73 Explain the difference between pre- and post-combustion methods for reducing SO_2 emissions.
State which method would be used (a) for producing fuels for cars
(b) in a power station

Pre-combustion method – the sulfur is removed before the fuel is burnt – this is used in producing fuels for cars.

Post-combustion method – the fuel containing sulfur is burnt, which produces SO_2. The SO_2 is then removed from the exhaust gases. This is usually used in power stations.

74 Describe some of the problems associated with acid deposition.

- damage/death of trees
- acidification of lakes/rivers which can cause death of fish
- erosion of buildings/statues made of limestone/marble