Unit D: Chemical Equilibrium Focusing on Acid-Base Systems – Solutions

Question 1

A chemical system is at equilibrium when the rate of the forward reaction is equal to the rate of the reverse reaction!

Option A is correct!

Question 2

$$\underset{\text{colourless}}{N_2O_4(g)} + heat \rightleftharpoons \underset{\text{brown}}{NO_2(g)}$$

Pay attention to what the question is asking specifically. It wants to know which property would NOT be used to determine if this system is at equilibrium.

Three of these options are properties which could be varied to demonstrate reversibility of the chemical system and they are incorrect.

Option B is incorrect. The colour of the system can display which side of the equilibrium predominates. A steady colour shows the system is at equilibrium.

Option C is incorrect. With differing amounts of particles on either side of the equilibrium, the pressure can be varied to push the equilibrium to one side or the other.

Option D is incorrect. The temperature can also be varied to push the equilibrium to one side or the other.

Option A is correct! The mass of the system will never change because of the law of conservation of mass. This could **not** be used to determine if the system is at equilibrium.

Question 5

$$N_2(g) + O_2(g) + 180.4\ kJ \rightleftharpoons 2\ NO(g) \quad K_c = 4.0\times10^{-31}$$

$$K_c < 1$$

K_c for this reaction is considerably lower than 1. This means that reactants are heavily favoured in this equilibrium. The concentration of the product, [NO(g)], is far lower than the concentration of any reactant.

Option D is correct! [NO(g)] < [N_2(g)]

Question 6

This question's solution requires the construction of an ICE table!

Note: If a given temperature or pressure is held constant, there is no need to worry about either.

Remember, it is **molar concentrations that go in an ICE table**, NOT mole amounts. Luckily, the reaction vessel volume is 1.0 L so the mole amounts translate directly to concentrations in mol/L.

	$N_2(g)$	$+$	$3\,H_2(g)$	$\rightleftharpoons$	$2\,NH_3(g)$
I	80.0		90.0		0
C	$-x$		$-3x$		$+2x$
E	$80.0 - x$		$90.0 - 3x$		$2x$ $= 20.0$

The equilibrium amount of $NH_3(g)$ is 20.0 moles, so:

$$2x = 20.0\,mol$$

$$x = 10.0\,mol$$

$$[H_2(g)]_{eq} = 90.0\,mol - 3(10.0\,mol) = \mathbf{60.0\,mol}$$

The correct answer is **60.0**.

Question 9

$$K_c = \frac{[C_2H_5OH(g)]}{[C_2H_4(g)][H_2O(g)]}$$

Given:

$$K_c = 300.0$$

Remember to divide mole amounts by the volume of the reaction vessel, 5000 L.

$$[C_2H_4(g)] = \frac{115\ mol}{5000\ L} = 0.0230\ \frac{mol}{L}$$

$$[H_2O(g)] = \frac{110\ mol}{5000\ L} = 0.0220\ \frac{mol}{L}$$

These values get plugged into the equilibrium law expression and a little algebra is done to solve for the concentration of $C_2H_5OH(g)$.

$$300.0 = \frac{[C_2H_5OH(g)]}{(0.0230)(0.0220)}$$

$$[C_2H_5OH(g)] = (300.0)(0.0230)(0.0220) \doteq \mathbf{0.152\ \frac{mol}{L}}$$

The correct answer is B, 1.52×10^{-1} mol/L.

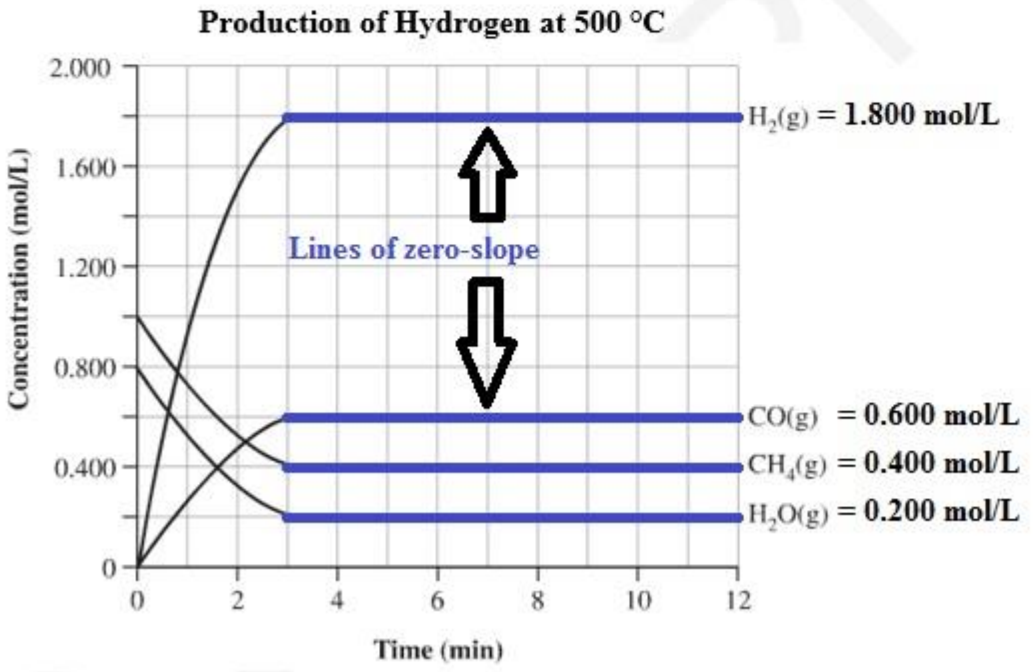
Question 10

$$CH_4(g) + H_2O(g) \rightleftharpoons 3\,H_2(g) + CO(g)$$

The law expression for the above equilibrium is:

$$K_c = \frac{[H_2(g)]^3[CO(g)]}{[CH_4(g)][H_2O(g)]}$$

Looking at the concentration vs. time graph below, the system is at equilibrium when all species' curves reach a line of zero-slope. The equilibrium concentrations of each entity can be read right off the graph.



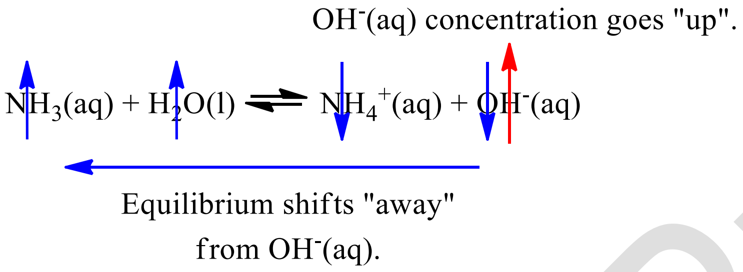
Simply plug the equilibrium concentrations in to the law expression to calculate the equilibrium constant, K_c.

$$K_c = \frac{(1.800)^3(0.600)}{(0.400)(0.200)} \doteq \mathbf{43.7}$$

The correct answer is **43.7**.

Question 13

Adding NaOH(aq) to the equilibrium below is a stress that initially increases the concentration of OH^-(aq).



In the diagram above, the red arrow indicates the stress and the blue arrows indicate what happens because of the stress.

The increase or decrease of the concentration of an entity will NOT change the value of the equilibrium constant, so options A and B are incorrect.

From the diagram above, the concentration of NH_4^+(aq) decreases because of increasing the OH^-(aq) concentration.

Option D is correct!

Question 14

There is only one type of stress on an equilibrium system that will change the equilibrium constant and that is a **change in temperature**.

Option B is correct, an increase in temperature.

Question 15

$$\begin{array}{ccccc} H_2(g) & + & I_2(g) & + 53.0\ kJ \rightleftharpoons & 2\ HI(g) \\ colourless & & purple & & colourless \end{array}$$

Luckily, the equilibrium takes place in a 1.0 L reaction vessel, so mole quantities can be translated directly to molar (mol/L) concentrations. However, the amounts of each entity are given in millimoles (mmol), not moles. The millimole quantities can be converted to mole quantities by simply dividing each by 1000 mmol/mol.

$$[H_2(g)] = \left(0.057\ \frac{mmol}{L}\right)\left(\frac{1\ mol}{1000\ mmol}\right) = 0.000057\ \frac{mol}{L}$$

$$[I_2(g)] = \left(1.07\ \frac{mmol}{L}\right)\left(\frac{1\ mol}{1000\ mmol}\right) = 0.00107\ \frac{mol}{L}$$

$$[HI(g)] = \left(1.87\ \frac{mmol}{L}\right)\left(\frac{1\ mol}{1000\ mmol}\right) = 0.00187\ \frac{mol}{L}$$

The molar concentrations can then be plugged into the equilibrium law expression to calculate the equilibrium constant (K_c) for the system.

$$K_c = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]} = \frac{(0.00187)^2}{(0.000057)(0.00107)} \doteq \mathbf{57}$$

The correct answer is A, 57.

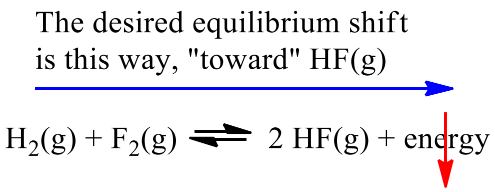
Question 16

Options B and C are incorrect. Because of the law of conservation of mass the density and mass of the system at equilibrium will be the same as those for the system not at equilibrium.

Option D is incorrect. There are two particles on the left and there are two particles on the right. The total pressure of the system at equilibrium will be the same as that for the system not at equilibrium.

Option A is correct! The system will maintain a consistent hue of purple when it is at equilibrium. If the system lies to the left before equilibrium, it will start heavily purple and the colour will become progressively less intense as the system moves towards equilibrium. Conversely, if the system lies to the right before equilibrium, it will start as almost colourless but become progressively more purple as the system moves towards equilibrium.

Question 19

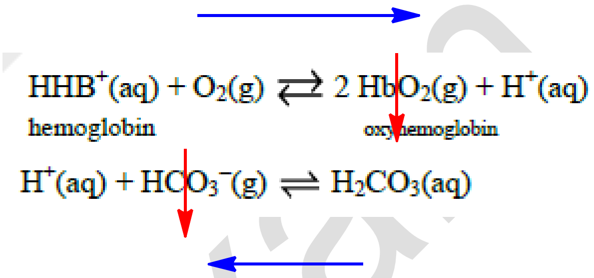


Adding an energy term to the products side of the equilibrium is necessary because of the negative Δ*H* value shown.

Decreasing the temperature will make the energy go "down" and will cause the equilibrium to shift "toward" the side of the energy term and HF(g).

Option D is correct!

Question 20



Of the three equilibria depicted, there are two that involve $H^+(aq)$.

To increase the $[H^+(aq)]$, the first equilibrium could be shifted to the right, or the second equilibrium could be shifted to the left.

Based on the options, this can only be done by decreasing the concentrations of other entities. To shift the equilibria in the desired directions, "toward" $H^+(aq)$, the $[HbO_2(aq)]$ can go "down" or the $[HCO_3^-(aq)]$ can go "down".

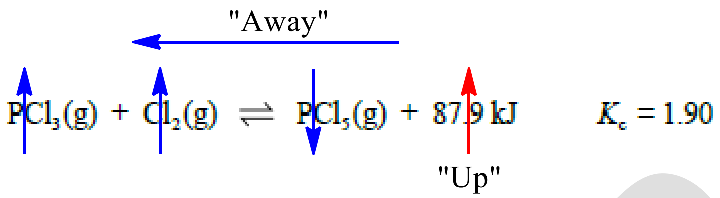
There is no option for $[HbO_2(aq)]$.

Option C is correct! The $[H^+(aq)]$ can be increased by decreasing the $[HCO_3^-(aq)]$.

Question 24

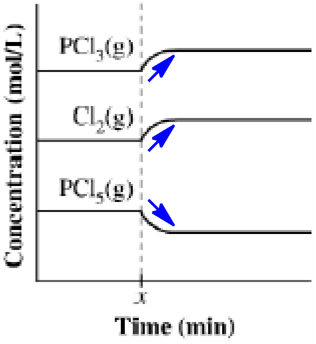
Heating the system, or increasing the system's temperature, will generate a concentration vs. time graph that has NO SPIKES at the time of stress.

Option B is incorrect. The concentration of $PCl_5(g)$ spikes down at the time of stress.



The correct option should show both the concentrations of $PCl_3(g)$ and $Cl_2(g)$ increasing and the concentration of $PCl_5(g)$ decreasing after the time of stress.

D.



Option D is correct!

Question 25

The law expression for the equilibrium above is:

$$K_c = \frac{[PCl_5(g)]}{[PCl_3(g)][Cl_2(g)]}$$

The K_c and equilibrium concentrations get plugged into the equilibrium law expression and a little algebra is done to solve for the equilibrium concentration of Cl_2(g).

$$1.90 = \frac{(0.255)}{(0.165)[Cl_2(g)]}$$

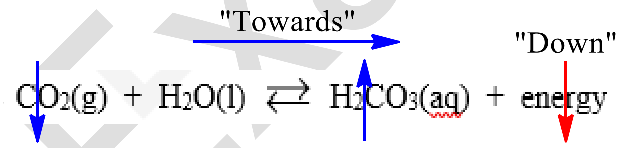
$$[Cl_2(g)] = \frac{0.255}{(1.90)(0.165)} \doteq \mathbf{8.13\times10^{-1}} \, \frac{\boldsymbol{mol}}{\boldsymbol{L}}$$

$$\boldsymbol{a} = \mathbf{8} \;\; \boldsymbol{b} = \mathbf{1} \;\; \boldsymbol{c} = \mathbf{3} \;\; \boldsymbol{d} = \mathbf{1}$$

The correct answer is 8131.

Question 27

The refrigerator will lower the temperature of the soft drink and shift the equilibrium toward the right and produce **more H_2CO_3(aq)** (**i**).



The equilibrium law expression involves products/reactants and in the above diagram, with a decrease in temperature, the product concentration increases and the reactant concentration decreases.

$$\frac{[product]\ increase}{[reactant]\ decrease} = K_c\ increase!$$

Increasing the numerator while decreasing the denominator of the K_c expression will result in an **increase of K_c** (**ii**).

Option C is correct! $i = H_2CO_3(aq), ii = increase$

Question 28

$$\begin{array}{ccc} N_2(g) + 3\,H_2(g) & \rightleftharpoons & 2\,NH_3(g) + energy \\ 4\,particles & & 2\,particles \end{array}$$

Increasing the pressure (**i**) would shift the equilibrium towards the products side which contains fewer particles (2 particles) than the reactants side (4 particles). With the energy term on the products side, a **decrease in temperature** (**ii**) would shift the equilibrium to the same side, increasing the concentration of the products.

Option B is correct! $i = increased, ii = decreased$

Question 29

Addition of a catalyst does NOT change the position of the equilibrium (**i**) and it does NOT change the value of $\Delta H°$ (**ii**).

Option D is correct! $i = does\ not\ change, ii = does\ not\ change$

Question 32

Option A is incorrect. This is not a gaseous equilibrium so pressure is not considered.

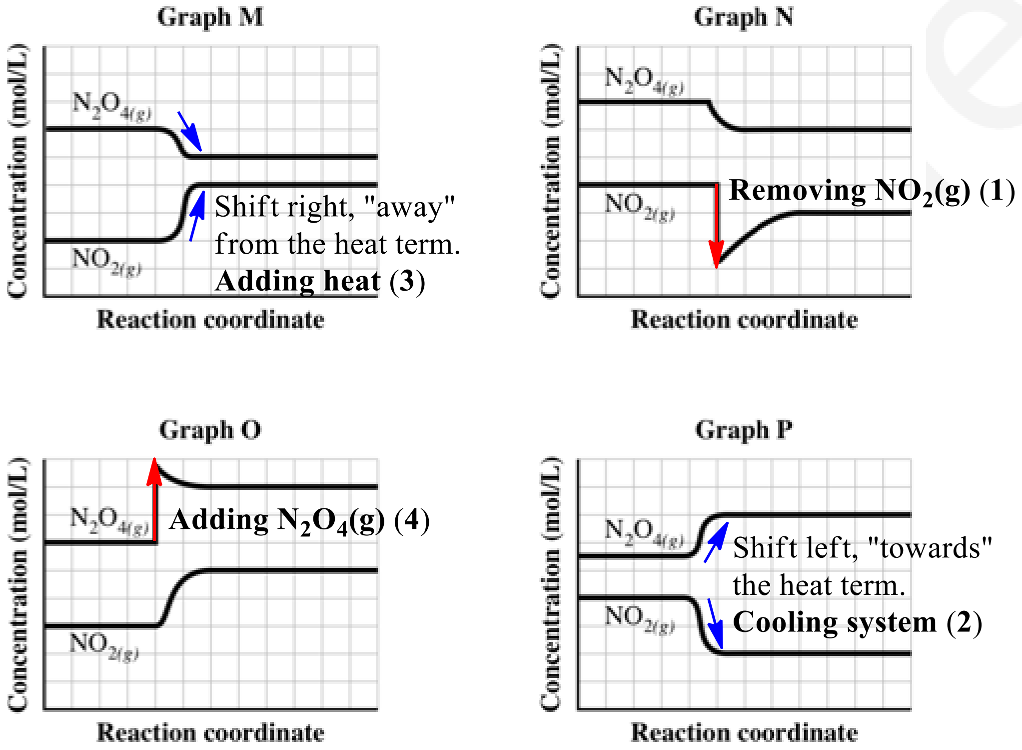
Option B is incorrect. An equilibrium is a dynamic one. The forward and reverse processes have NOT stopped. They are still occurring, but at the same rate.

Option C is incorrect. If the solid citric acid is completely dissolved, there is no equilibrium established as all the solid is simply dissolved in the water.

Option D is correct! A constant mass of solid citric acid remaining is a consistent observable property indicating equilibrium. This dynamic equilibrium involves the dissolution of solid citric acid with the reprecipitation of the dissolved particles occurring at the same time.

Question 35

$$N_2O_4(g) + \text{heat} \rightleftarrows 2\,NO_2(g)$$



Shift right, "away" from the heat term. **Adding heat** (3)

Removing NO_2(g) (1)

Adding N_2O_4(g) (4)

Shift left, "towards" the heat term. **Cooling system** (2)

The correct answer is **3142**.

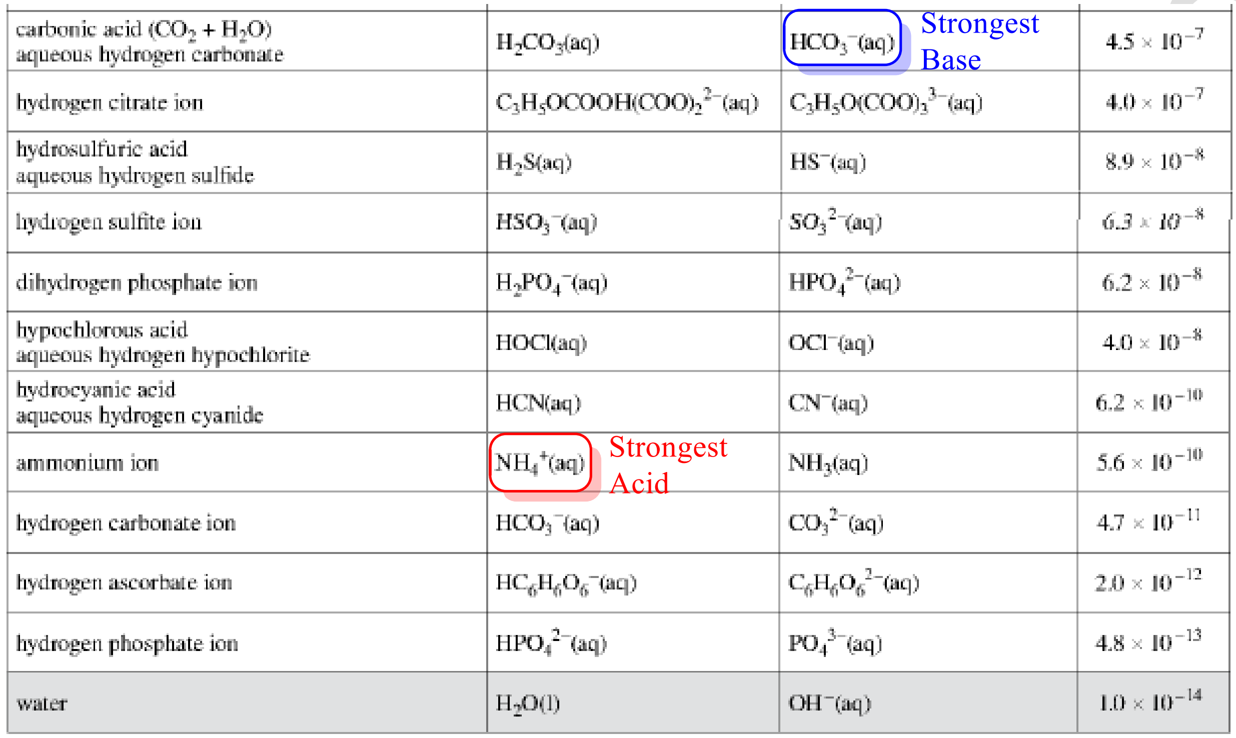
Question 38

One of the options contains a strong acid and the other three are weak acids. The solution that contains the strong acid will have the lowest pH. Strong acids are the top six entries in the data booklet's relative strengths of acids and bases table.

Option B is correct! 0.50 mol/L HNO_3(aq) (strong acid)

Question 40

The mixture contains NH_4^+(aq), HCO_3^-(aq), and H_2O(l). From these components, the strongest acid and strongest base must be identified from the table of relative strengths of acids and bases. The strongest acid is highest on the left. The strongest base is lowest on the right.



carbonic acid ($CO_2 + H_2O$) aqueous hydrogen carbonate	H_2CO_3(aq)	HCO_3^-(aq)	4.5×10^{-7}
hydrogen citrate ion	$C_3H_5OCOOH(COO)_2^{2-}$(aq)	$C_3H_5O(COO)_3^{3-}$(aq)	4.0×10^{-7}
hydrosulfuric acid aqueous hydrogen sulfide	H_2S(aq)	HS^-(aq)	8.9×10^{-8}
hydrogen sulfite ion	HSO_3^-(aq)	SO_3^{2-}(aq)	6.3×10^{-8}
dihydrogen phosphate ion	$H_2PO_4^-$(aq)	HPO_4^{2-}(aq)	6.2×10^{-8}
hypochlorous acid aqueous hydrogen hypochlorite	HOCl(aq)	OCl^-(aq)	4.0×10^{-8}
hydrocyanic acid aqueous hydrogen cyanide	HCN(aq)	CN^-(aq)	6.2×10^{-10}
ammonium ion	NH_4^+(aq)	NH_3(aq)	5.6×10^{-10}
hydrogen carbonate ion	HCO_3^-(aq)	CO_3^{2-}(aq)	4.7×10^{-11}
hydrogen ascorbate ion	$HC_6H_6O_6^-$(aq)	$C_6H_6O_6^{2-}$(aq)	2.0×10^{-12}
hydrogen phosphate ion	HPO_4^{2-}(aq)	PO_4^{3-}(aq)	4.8×10^{-13}
water	H_2O(l)	OH^-(aq)	1.0×10^{-14}

Strongest Base

Strongest Acid

The equilibrium that is formed in this mixture is:

$$\underset{acid}{NH_4^+(aq)} \quad + \quad \underset{base\ (\boldsymbol{i})}{HCO_3^-(aq)} \quad \rightleftharpoons \quad \underset{conjugate\ base}{NH_3(aq)} \quad + \quad \underset{conjugate\ acid\ (\boldsymbol{ii})}{H_2CO_3(aq)}$$

Option A is correct! $i = HCO_3^-(aq), ii = H_2CO_3(aq)$

Question 43

$$HI(g) \xrightarrow{H_2O(l)} HI(aq)$$

$$HI(aq) + H_2O(l) \rightarrow I^-(aq) + H_3O^+(aq)$$

HI(aq) is a strong acid! It is one of the top six acids listed in the relative strengths of acids and bases table.

In any acidic solution, the pH is less than the pOH (**1**).

The solution conducts electricity (**3**) because of almost complete ionization of the HI(aq) in solution (**5**).

In any acidic solution, the concentration of H_3O^+(aq) is greater than the concentration of OH^-(aq) (**8**).

The correct answer is **1358**.

Question 47

There is a typo in this question. The equation should read:

$$H_2SO_3(aq) + F^-(aq) \rightleftharpoons HSO_3^-(aq) + HF(aq)$$

An amphiprotic entity is one which can act as either an acid (donate H^+) or a base (accept H^+).

There is only one entity among those in the equation above that is amphiprotic.

$$SO_3^{2-}(aq) \xleftarrow{-H^+} HSO_3^-(aq) \xrightarrow{+H^+} H_2SO_3(aq)$$

The correct answer is B, HSO_3^-(aq).

Question 50

A saturated solution of HP(aq) has a concentration of 1.20 × 10^{-2} mol/L and a pH of 2.420.

The concentration of HP(aq) is an initial concentration, before it ionizes to form P^-(aq) and H_3O^+(aq).

$$[HP(aq)]_i = 1.20\times10^{-2}\ \frac{mol}{L}$$

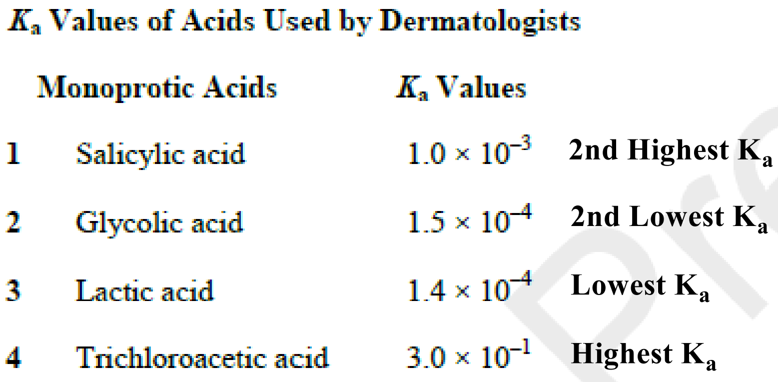
$$-log[HP(aq)]_i = -log(1.20\times10^{-2}) \doteq 1.921 < 2.420$$

Here, the negative log of the $[HP(aq)]_i$ is lower than the pH. Therefore, HP(aq) is a **weak acid** whose K_a is lower than 1 ($\boldsymbol{K_a < 1}$).

Option D is correct!

Question 51

The strongest acid will have the highest K_a. The weakest acid will have the lowest K_a.



K_a Values of Acids Used by Dermatologists

	Monoprotic Acids	K_a Values	
1	Salicylic acid	1.0×10^{-3}	**2nd Highest K_a**
2	Glycolic acid	1.5×10^{-4}	**2nd Lowest K_a**
3	Lactic acid	1.4×10^{-4}	**Lowest K_a**
4	Trichloroacetic acid	3.0×10^{-1}	**Highest K_a**

The correct answer is **4123**.

Question 55

Ammonia, $NH_3(aq)$, acts primarily as a **base** in aqueous solution. In the relative strengths of acids and bases table, $NH_3(aq)$, is only found in the bases column.

In aqueous solution, it forms an equilibrium where it accepts a proton from water to form the ammonium ion and hydroxide ion.

$$\underset{base}{NH_3(aq)} + \underset{acid}{H_2O(l)} \rightleftharpoons \underset{conjugate\ acid}{NH_4^+(aq)} + \underset{conjugate\ base}{OH^-(aq)}$$

A single H^+ transfer! Note how $NH_4^+(aq)$ differs from $NH_3(aq)$ by a single H^+ and the overall +1 charge of $NH_4^+(aq)$ where $NH_3(aq)$ is neutral.

Option C is correct!

Question 57

This is a great question because it introduces a whole new compound, pyridine, whose reactivity can follow the principles learned here in Chemistry 30.

The question states that pyridine, $C_5H_5N(l)$ acts as a **base**. It can **accept a single proton** to form its conjugate acid.

The correct option differs from $C_5H_5N(l)$ by a single H^+!

Option C is incorrect. It differs from $C_5H_5N(l)$ by an N where it has gained one N ($C_5H_5N/C_5H_5\mathbf{N_2}$).

Option D is incorrect. It differs from $C_5H_5N(l)$ by having lost an N and replacing it with a CH ($C_5H_5\mathbf{N}/C_6H_6$).

Option B is incorrect. The overall charge is incorrect. Adding an H^+ to C_5H_5N should give a product with an overall charge of +1, NOT give it a charge of -1.

Option A is correct! ($C_5H_5N/C_5H_5N\mathbf{H^+}$) $C_5H_5NH^+$ differs from C_5H_5N by a single H^+ and has the correct overall charge of +1.

Question 58

Remember that a **Bronsted-Lowry base** is an entity that can **accept an H^+** from a Bronsted-Lowry acid.

Two of the options are ionic compounds and need to be dissociated to better see if either contains an entity that is a Bronsted-Lowry base.

$$A)\ Na^+(aq) + \boldsymbol{Cl^-(aq)}$$

$$D)\ Na^+(aq) + \boldsymbol{HCO_3^-(aq)}$$

Option A is incorrect. $Cl^-(aq)$ is the conjugate base of a top-six acid. It has absolutely no base-strength.

Option B is incorrect. $CH_3OH(aq)$ is not even found in the bases column of the relative strengths of acids and bases table.

Option C is incorrect. HCOOH (aq) is found in the acids column of the relative strengths of acids and bases table, NOT in the bases column.

Option D is correct! $HCO_3^-(aq)$ is indeed a Bronsted-Lowry base. It can accept an H^+ from an acid and is found in the bases column of the relative strengths of acids and bases table.

Question 60

Sulfurous acid, H_2SO_3(aq), is a weak acid and in aqueous solution forms the following equilibrium:

$$\underset{acid}{H_2SO_3(aq)} + H_2O(l) \rightleftharpoons \underset{\boldsymbol{conjugate\ base}}{\boldsymbol{HSO_3^-(aq)}} + H_3O^+(aq)$$

The K_b of HSO_3^-(aq) can be calculated by dividing K_w (1.0×10^{-14}) by the K_a of its conjugate acid, H_2SO_3(aq), where the K_a can be looked up in the relative strengths of acids and bases table.

$$K_{a_{H_2SO_3(aq)}} = 1.4\times10^{-2}$$

$$K_{b_{HSO_3^-(aq)}} = \frac{K_w}{K_{a_{H_2SO_3(aq)}}} = \frac{1.0\times10^{-14}}{1.4\times10^{-2}} \doteq \mathbf{7.1\times10^{-13}}$$

$$\boldsymbol{a = 7\ \ b = 1\ \ c = 1\ \ d = 3}$$

The correct answer is **7113**.

Question 61

A weak base system at equilibrium in aqueous solution looks like this:

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

The weak base, B(aq), accepts an H^+ from water in equilibrium to form its conjugate acid, BH^+(aq), and the hydroxide ion. Note the **equilibrium arrows** used for the weak base system.

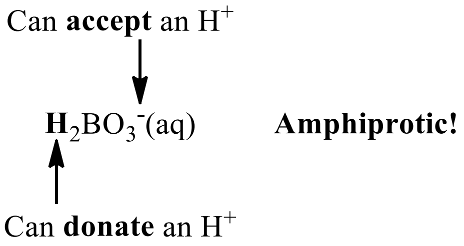
The option that most closely resembles the above equation is:

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

Option A is correct!

Question 63

Remember that an amphiprotic entity can both donate an H^+ AND accept an H^+.



The amphiprotic species is $\mathbf{H_2BO_3^-(aq)}$ (**i**).

A **conjugate acid-base pair** is a pair of entities that **differ by only a single H^+**. Of the options listed, there is only one that displays this:

$$\boldsymbol{ii = HNO_2(aq)\ and\ NO_2^-(aq)}$$

Option C is correct! $i = H_2BO_3^-(aq), ii = HNO_2(aq)\ and\ NO_2^-(aq)$

Question 68

This question can be looked at as a titration situation where oxalic acid, $HOOCCOOH(aq)$, is being titrated by some hydroxide base. All the mixtures and entities would be present at key points on the titration curve.

The following reactions would occur from the beginning of the titration to the end, in sequence.

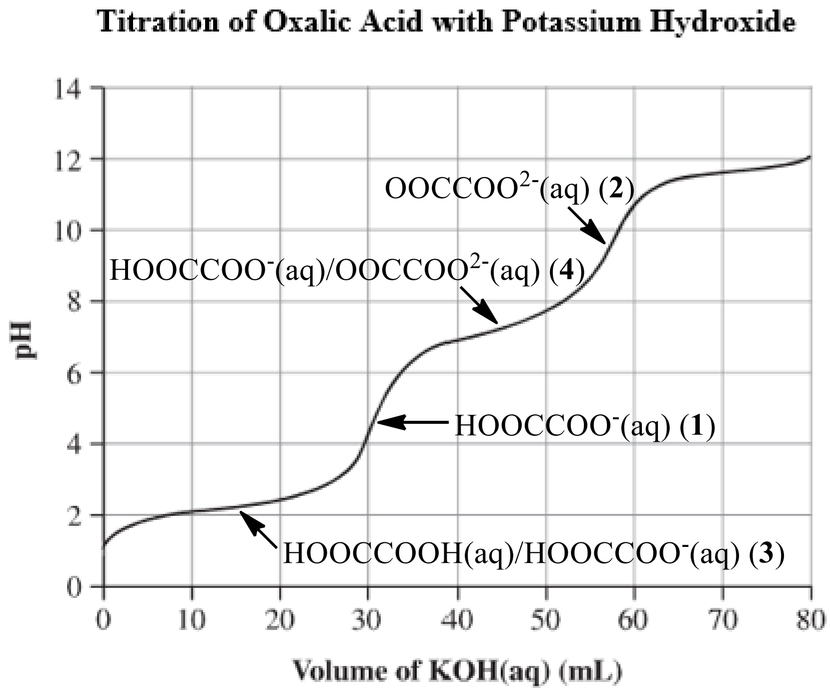
First reaction (removal of first proton):

$$HOOCCOOH(aq) + OH^{-}(aq) \rightarrow HOOCCOO^{-}(aq) + H_2O(l)$$

Second reaction (removal of last proton):

$$HOOCCOO^{-}(aq) + OH^{-}(aq) \rightarrow OOCCOO^{2-}(aq) + H_2O(l)$$

At the start of the titration, the pH would be low, at the pH of oxalic acid. As the titration progresses, the pH would gradually get higher save for the equivalence points where marked increases in pH would be observed.



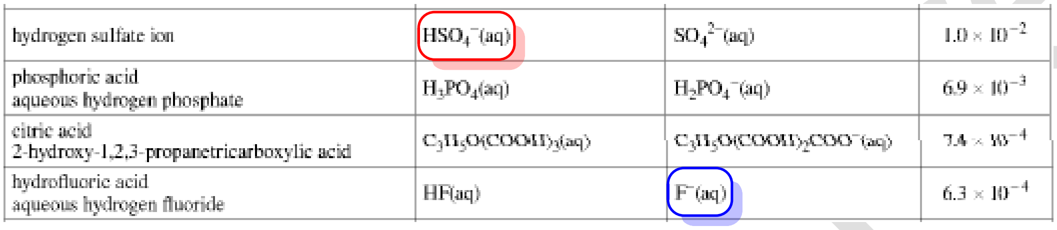
The correct answer is **3142**.

Question 69

An acid-base reaction which favours the products is one in which the acid and the base are found in this orientation on the relative strengths of acids and bases table.

$$\begin{array}{ccc} Acid & & \\ & + & \\ & & Base \end{array}$$

Of the options listed, there is an acid-base pair that have this orientation on the table.



hydrogen sulfate ion	$HSO_4^-(aq)$	$SO_4^{2-}(aq)$	1.0×10^{-2}
phosphoric acid aqueous hydrogen phosphate	$H_3PO_4(aq)$	$H_2PO_4^-(aq)$	6.9×10^{-3}
citric acid 2-hydroxy-1,2,3-propanetricarboxylic acid	$C_3H_5O(COOH)_3(aq)$	$C_3H_5O(COOH)_2COO^-(aq)$	7.4×10^{-4}
hydrofluoric acid aqueous hydrogen fluoride	$HF(aq)$	$F^-(aq)$	6.3×10^{-4}

Option C is correct! The equilibrium which favours products is:

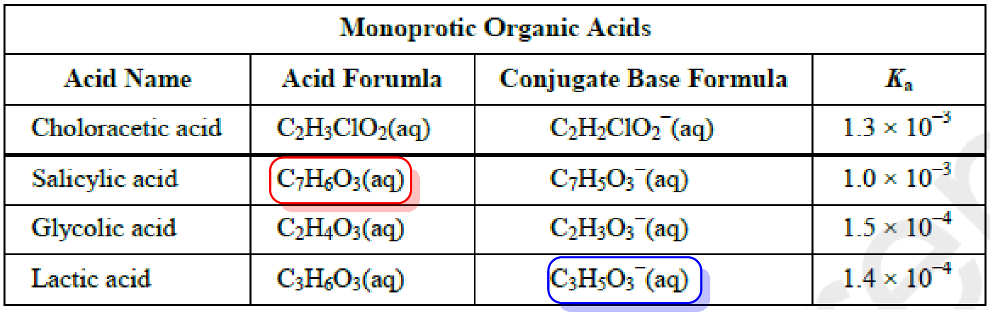
$$HSO_4^-(aq) + F^-(aq) \rightarrow SO_4^{2-}(aq) + HF(aq)$$

Question 72

The given table has acids listed from strongest to weakest as evidenced by the descending K_a values as one reads down the table.

Read each formula carefully and locate the acid-base pair that are found in this orientation to get the equilibrium that favours the products.

$$\begin{array}{ccc} Acid & & \\ & + & \\ & & Base \end{array}$$



Monoprotic Organic Acids

Acid Name	Acid Forumla	Conjugate Base Formula	K_a
Choloracetic acid	$C_2H_3ClO_2(aq)$	$C_2H_2ClO_2^-(aq)$	1.3×10^{-3}
Salicylic acid	$C_7H_6O_3(aq)$	$C_7H_5O_3^-(aq)$	1.0×10^{-3}
Glycolic acid	$C_2H_4O_3(aq)$	$C_2H_3O_3^-(aq)$	1.5×10^{-4}
Lactic acid	$C_3H_6O_3(aq)$	$C_3H_5O_3^-(aq)$	1.4×10^{-4}

Option B is correct!

$$\underset{acid}{C_7H_6O_3(aq)} + \underset{base}{C_3H_5O_3^-(aq)} \rightarrow C_7H_5O_3^-(aq) + C_3H_6O_3(aq)$$

Question 73

Start by labelling each reactant as an acid or a base.

$$(\mathbf{1})\ \underset{\color{red}{acid}}{HNO_2(aq)} + \underset{\color{blue}{base}}{HCOO^-(aq)} \rightleftharpoons NO_2^-(aq) + HCOOH(aq)\ \ (Products\ favoured)$$

$$(\mathbf{2})\ \underset{\color{red}{acid}}{HN_3(aq)} + \underset{\color{blue}{base}}{OBr^-(aq)} \rightleftharpoons N_3^-(aq) + HOBr(aq)\ \ (Products\ favoured)$$

$$(\mathbf{3})\ \underset{\color{red}{acid}}{HN_3(aq)} + \underset{\color{blue}{base}}{HCOO^-(aq)} \rightleftharpoons N_3^-(aq) + HCOOH(aq)\ \ (Reactants\ favoured)$$

Place acids and bases in a table (in **bold**), orienting them based on whether products or reactants are favoured.

	Acids	**Bases**	
Strongest Acid	**HNO_2(aq) (1)**		Weakest Base
		$HCOO^-$(aq) (1)	
	HN_3(aq) (3)		
Weakest Acid		**OBr^-(aq) (2)**	Strongest Base

Finish by filling in the blanks with each acid's or base's conjugate (in *italics*) as seen on the products side of the reactions:

	Acids	**Bases**	
Strongest Acid	**HNO_2(aq) (1)**	*NO_2^-(aq)*	Weakest Base
	HCOOH(aq)	**$HCOO^-$(aq) (1)**	
	HN_3(aq) (3)	*N_3^-(aq)*	
Weakest Acid	*HOBr(aq)*	**OBr^-(aq) (2)**	Strongest Base

The acids ranked from strongest to weakest are: nitrous acid (**4**), formic acid (**1**), hydrazoic acid (**2**), hypobromous acid (**3**).

The correct answer is **4123**.

Question 74

Option A is incorrect. H_3O^+(aq) signifies that a strong acid is present. This mixture does not contain one.

Option B is incorrect. CH_3COO^-(aq) is not one of the reactants specified in the mixture.

Option D is incorrect. A net ionic equation should not contain sodium ions as these ions are spectators in the reaction.

Option C is correct! The CH_3COOH(aq) is the acid and it is the HCO_3^-(aq) ion that acts as a base in this reaction and the H_2CO_3(aq) formed decomposes to CO_2(g) and H_2O(l). The CO_2(g) gives the "fluffy texture" to the baked good.

Question 80

$$XeF_6(s) + 3\,H_2O(l) \rightarrow XeO_3(aq) + 6\,HF(aq)$$

The HF(aq) produced in this reaction is an acid!

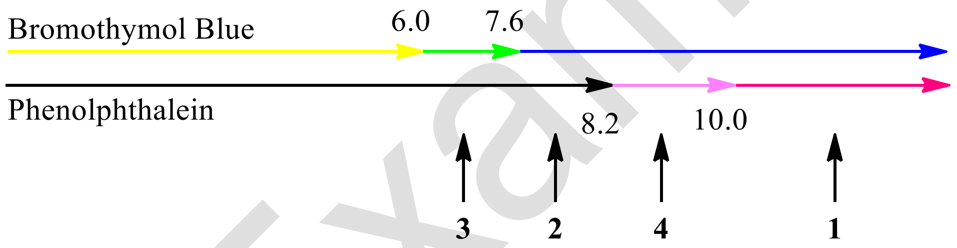
As more acid is produced, the acidity of the solution increases which causes the **pH of the solution to decrease** (**i**).

If the pH of the solution decreases, the **pOH of the solution increases** (**ii**).

Option C is correct! $i = decrease, ii = increase$

Question 82

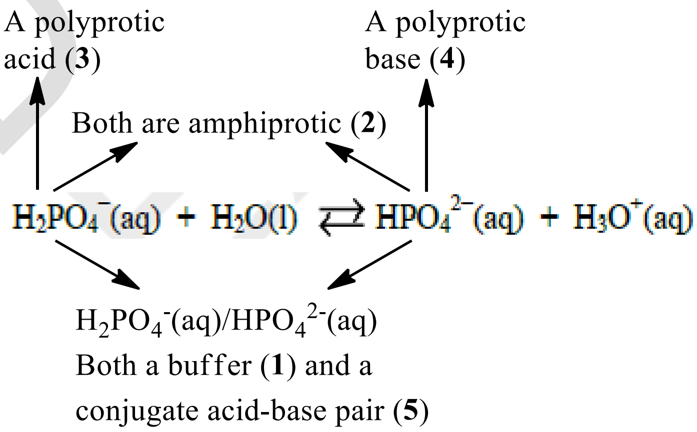
Draw a two overlapping number lines for the indicators and mark the pH values at which the colour changes occur for each indicator and the colour of the indicator over the pH ranges.



Each solution can be placed on the number line based on the colour of each indicator.

The correct answer is **3241**.

Question 83



Option D is correct! "1, 2, 3, 4, and 5"

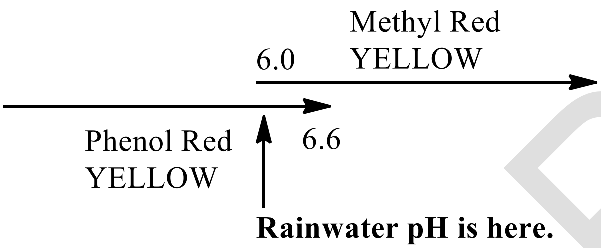
Question 85

Option A is incorrect. Bromothymol blue would NOT be blue at a pH of 6.0.

Option D is incorrect. Bromothymol blue would NOT be yellow at a pH of 7.6.

Methyl red is yellow from a pH of 6.0 and upwards.

Phenol red is yellow up to a pH of 6.6.



Option C is correct! The pH of the rainwater is 6.0 and the bromothymol blue colour is yellow.

Question 86

The pH of the nitrous acid ($HNO_2(aq)$) solution must first be determined.

$$[H^+(aq)] \doteq \sqrt{K_{a_{HNO_2(aq)}} \cdot [HNO_2(aq)]_i} = \sqrt{(5.6\times10^{-4})(0.50)} \doteq 0.0167332\,\frac{mol}{L}$$

$$pH = -log[H^+(aq)] \doteq -log(0.0167332) \doteq 1.78$$

Phenol red is **yellow** at a pH of 1.78.

$HNO_2(aq)$ will react with the basic form of the indicator, **$Pr^-(aq)$**.

$$HNO_2(aq) + Pr^-(aq) \rightleftharpoons NO_2^-(aq) + HPr(aq)$$

Option A is correct!

Question 89

The buffer created by this mixture is:

$$\underset{weak\ acid}{CH_3COOH(aq)} \quad / \quad \underset{conjugate\ base}{CH_3COO^-(aq)}$$

NaOH(aq) will react with the **acid portion of the buffer**.

Note: A net ionic equation should not contain sodium ions as these ions are spectators in the reaction.

$$\boldsymbol{A})\ CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$$

Question 90

$$\underset{\text{weak acid}}{H_2CO_3(aq)} \quad / \quad \underset{\text{conjugate base}}{HCO_3^-(aq)}$$

Hydronium ions, $H_3O^+(aq)$, will react with the **base portion of the buffer**.

$$\boldsymbol{D}) \; H_3O^+(aq) + HCO_3^-(aq) \rightarrow H_2CO_3(aq) + H_2O(l)$$

Question 92

A buffer solution is a mixture of a weak acid and its conjugate base.

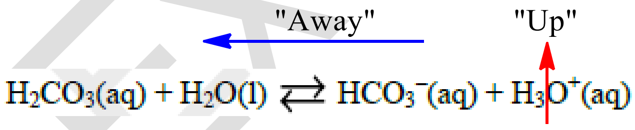
Options A, B, and C are incorrect. None of the pairs are conjugates and all those options involve strong acids.

Option D is correct! $CH_3COOH(aq)$ and $NaCH_3COO(aq)$ form this buffer:

$$\underset{\text{weak acid}}{CH_3COOH(aq)} \quad / \quad \underset{\text{conjugate base}}{CH_3COO^-(aq)}$$

Question 96

Adding HCl(aq) will increase the $H_3O^+(aq)$ concentration.



$$H_2CO_3(aq) + H_2O(l) \rightleftarrows HCO_3^-(aq) + H_3O^+(aq)$$

The equilibrium shifts **towards the reactants**.

A buffer system is meant to resist changes in pH, so there is **no change in the pH** of the system.

Option C is correct!

Question 98

A decrease in pH by 1 unit represents a ten-fold increase in acidity.

$$pH\ 5 \rightarrow pH\ 1 = decrease\ in\ pH\ by\ 4 = increase\ in\ acidity\ by\ a\ factor\ of\ 10^4$$

$$10^4 = 10\,000$$

Option D is correct! 10 000

Question 99

NH_3(aq) is a weak base! It forms this equilibrium in water:

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

The pH cannot be obtained directly. The [OH^-(aq)] must be determined first.

$$[OH^-(aq)] = \sqrt{K_{b_{NH_3(aq)}} \cdot [NH_3(aq)]_i}$$

[NH_3(aq)]$_i$ is given (0.3 mol/L), but the K_b of NH_3(aq) must be determined by dividing K_w by the K_a of NH_4^+(aq).

$$K_{b_{NH_3(aq)}} = \frac{K_w}{K_{a_{NH_4^+(aq)}}} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} \doteq 1.7857 \times 10^{-5}$$

$$[OH^-(aq)] = \sqrt{(1.7857 \times 10^{-5})(0.3)} \doteq 2.31454 \times 10^{-3}\ \frac{mol}{L}$$

From the [OH^-(aq)], the pOH can be determined.

$$pOH = -log[OH^-(aq)] \doteq -log(2.31454 \times 10^{-3}) \doteq 2.6$$

The pH can finally be calculated by subtracting the pOH from 14.0.

$$pH \doteq 14.0 - 2.6 = \mathbf{11.4}$$

The correct answer is **11.4**.

Question 102

One must remember that a K_a is the equilibrium constant for an acid equilibrium with water of this form:

$$HA(aq) + H_2O(l) \rightarrow A^-(aq) + H_3O^+(aq);\ K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

The challenging aspect of this question is that the background information makes no mention of this sort of reaction with respect to hydrazoic acid (HN_3(aq)).

One must realize that HN_3(aq) can donate an H^+ and form an equilibrium with water JUST LIKE the one above. Simply replace “A” with “N_3”.

$$HN_3(aq) + H_2O(l) \rightarrow N_3^-(aq) + H_3O^+(aq);\ \boldsymbol{K_a = \frac{[H_3O^+(aq)][N_3^-(aq)]}{[HN_3(aq)]}}$$

Option B is correct!

Question 103

$$K_{a_{HN_3(aq)}} = 1.9\times10^{-5}$$

$$[HN_3(aq)]_i = 0.28\ \frac{mol}{L}$$

$$[H_3O^+(aq)] = \sqrt{K_{a_{HN_3(aq)}} \cdot [HN_3(aq)]_i} = \sqrt{(1.9\times10^{-5})(0.28)} \doteq 2.3065\times10^{-3}$$

$$pH = -log[H_3O^+(aq)] \doteq -log(2.3065\times10^{-3}) \doteq \mathbf{2.64}$$

The correct answer is **2.64**.

Question 104

The **K_a** for HN_3(aq) is **less than 1**. This means that **reactants are favoured** in the equilibrium.

$$HN_3(aq) + H_2O(l) \rightarrow N_3^-(aq) + H_3O^+(aq)$$

$$[reactants] > [products]$$

$$\underset{\boldsymbol{reactant}}{[\boldsymbol{HN_3(aq)}]} > \underset{\boldsymbol{product}}{[\boldsymbol{H_3O^+(aq)}]}$$

Option B is correct!

Question 110

BE CAREFUL with this question! The likely response for this question will be either option C or D. These K_a values are of the order of 10^{-4} which is appropriate for a weak acid. If one divides the $[A^+(aq)]_i$ (1.62×10^{-3} mol/L) by 10^{-4}:

$$\frac{[A^+(aq)]_i}{K_a} \doteq \frac{1.62\times10^{-3}}{10^{-4}} = 16.2 < 1000$$

The result is a number that is NOT greater than 1000. The shortcut relating $[H_3O^+(aq)]$ and K_a cannot be used. One must build an ICE table!

$$\begin{array}{cccccccc}
 & 2\,H_2O(l) & + & A^+(aq) & \rightleftharpoons & H_3O^+(aq) & + & AOH(aq) \\
I & & & 1.62\times10^{-3} & & 0 & & 0 \\
C & & & -x & & +x & & +x \\
E & & & 1.62\times10^{-3} - x & & \boldsymbol{x} & & x \\
 & & & & & = \mathbf{8.5\times10^{-4}} & &
\end{array}$$

The background information states the equilibrium concentration of $H_3O^+(aq)$ to be 8.5×10^{-4} mol/L. This is the value of x.

The other equilibrium concentrations can be calculated.

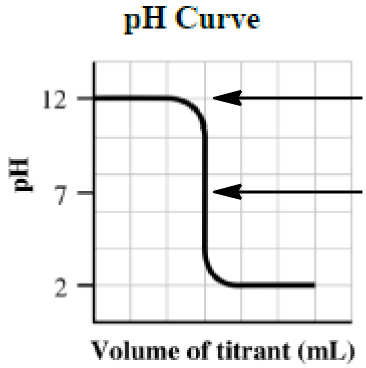
$$[AOH(aq)]_{eq} = 8.5\times10^{-4}\ \frac{mol}{L}$$

$$[A^+(aq)]_{eq} = (1.62\times10^{-3} - 8.5\times10^{-4})\ \frac{mol}{L} = 7.7\times10^{-4}\ \frac{mol}{L}$$

$$K_{a_{A^+(aq)}} = \frac{[H_3O^+(aq)][AOH(aq)]}{[A^+(aq)]} = \frac{(8.5\times10^{-4})(8.5\times10^{-4})}{(7.7\times10^{-4})} \doteq \mathbf{9.38\times10^{-4}}$$

Option C is correct!

Question 120



pH starts high and gets lower as titration progresses

Large vertical portion where the equivalence point is at pH=7

This depicts a titration of a strong base with a strong acid as titrant.

Option A is incorrect. Ammonia is a weak base.

Option B is incorrect. Acetic acid is a weak acid.

Option C is incorrect. Oxalic acid is a weak diprotic acid.

Option D is correct! Sodium hydroxide is a strong base and hydrochloric acid is a strong acid!

Question 121

$$CN^-(aq) + H_2O(l) \rightleftharpoons HCN(aq) + OH^-(aq);\ \ K_b = 1.61\times10^{-5}$$

The pH of the NaCN(aq) solution is given (8.710). As seen in the equilibrium above, $CN^-(aq)$ acts as a weak base in aqueous solution.

The pH can be used to calculate the pOH, which in turn, can be used to calculate the equilibrium concentration of $OH^-(aq)$.

$$pOH = 14.000 - 8.710 = 5.290$$

$$[OH^-(aq)]_{eq} = 10^{-pOH} = 10^{-5.290} \doteq 5.1286\times10^{-6}$$

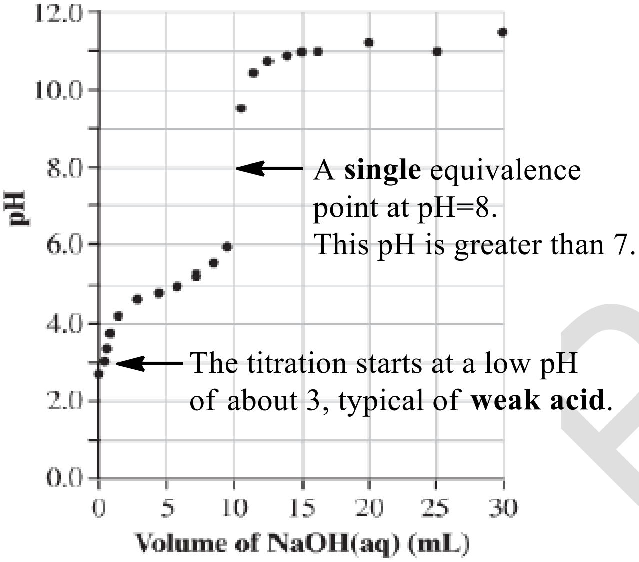
The equilibrium concentration of $CN^-(aq)$ can be easily determined by algebraically isolating for it from the equilibrium law expression for the above equilibrium. Remember that the equilibrium concentration of $OH^-(aq)$ is the same as that of HCN(aq).

$$[CN^-(aq)]_{eq} = \frac{[HCN(aq)]_{eq}[OH^-(aq)]_{eq}}{K_{b_{CN^-}}} \doteq \frac{(5.1286\times10^{-6})^2}{1.61\times10^{-5}} \doteq 1.63\times10^{-6}$$

$$\boldsymbol{a = 1\ \ b = 6\ \ c = 3\ \ d = 6}$$

The correct answer is **1636**.

Question 122



A **single** equivalence point at pH=8.
This pH is greater than 7.

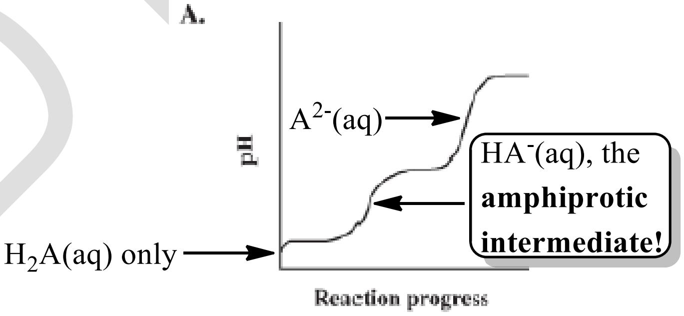
The titration starts at a low pH of about 3, typical of **weak acid**.

With a single equivalence point, the sample is clearly a **monoprotic weak acid**.

Option C is correct!

Question 125

Options C and D are incorrect. In the titration of an acid with a strong base, the acid is the sample and the base is gradually added as titrant. The pH should start low and increase as the titration proceeds.



$A^{2-}(aq)$

$HA^{-}(aq)$, the **amphiprotic intermediate!**

$H_2A(aq)$ only

Option A is correct!

Question 128

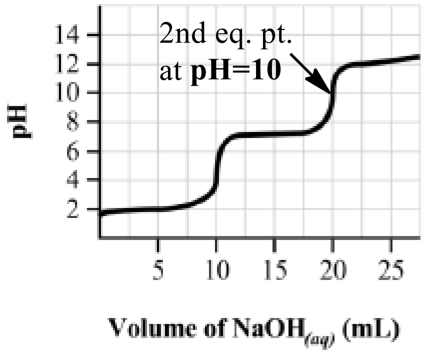
The hydrogen carbonate ion, $HCO_3^-(aq)$, acts as a base and accepts protons from phosphoric acid, $H_3PO_4(aq)$.

$$\underset{}{H_3PO_4(aq)} \xrightarrow{-H^+} \underset{1st\ eq.pt.}{H_2PO_4^-(aq)} \xrightarrow{-H^+} \underset{\boldsymbol{2nd\ eq.pt.}}{\boldsymbol{HPO_4^{2-}(aq)}}$$

The reaction at the 2nd equivalence point should produce $HPO_4^{2-}(aq)$.

$$\boldsymbol{B})\ \underset{acid}{H_2PO_4^-(aq)} + \underset{base}{HCO_3^-(aq)} \rightarrow HPO_4^{2-}(aq) + H_2CO_3(aq)$$

Question 129



The second equivalence point occurs at a pH of about 10.

Of the options listed, the only indicator that changes colour around this pH is **phenolphthalein** which goes from colourless to pink over the pH range of 8.2-10.0.

Option C is correct!