- Ascorbic acid (Vitamin C) is a monoprotic acid of formula $C_6H_8O_6$. Calculate the pH of a 0.10 M solution of ascorbic acid, given the K_a of ascorbic acid is 8.0×10^{-5} M. 3

As ascorbic acid is a weak acid, $[H_3O^+]$ must be calculated:

	$C_6H_8O_6$	H_2O	$\rightleftharpoons$	H_3O^+	$C_6H_7O_6^-$
initial	**0.1**	**large**		**0**	**0**
change	$-x$	**negligible**		$+x$	$+x$
final	$0.10 - x$	**large**		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+(aq)][C_6H_7O_6^-(aq)]}{[C_6H_78O_6(aq)]} = \frac{x^2}{(0.10 - x)}$$

As $K_a = 8.0 \times 10^{-5}$ is very small, $0.10 - x \sim 0.10$ and hence:

$$x^2 = 0.1 \times (8.0 \times 10^{-5}) \text{ or } x = 2.8 \times 10^{-3} \text{ M} = [H_3O^+(aq)]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H_3O^+(aq)] = -\log_{10}[0.0028] = 2.5$$

Answer: **pH = 2.5**

- Write equations to show what happens to a buffer solution containing equimolar amounts of $C_6H_5CH_2COOH$ and $C_6H_5CH_2COOK$ when: 2
(a) H_3O^+ is added, (b) OH^- is added.

(a)
$C_6H_5CH_2COO^-(aq) + H_3O^+(aq) \rightarrow C_6H_5CH_2COOH(aq) + H_2O(l)$

(b)
$C_6H_5CH_2COOH(aq) + OH^-(aq) \rightarrow C_6H_5CH_2COO^-(aq) + H_2O(l)$

- Calculate the pH of a solution that is 0.010 M in benzoic acid, C_6H_5COOH, and 0.010 M in $C_6H_5CO_2Na$. The K_a of benzoic acid is 6.4×10^{-5} M. **4**

This solution contains an acid and its conjugate base so the Henderson-Hasselbalch equation can be used:

$$\mathbf{pH = p}K_\mathbf{a} + \mathbf{log_{10}}\left(\frac{\mathbf{[base]}}{\mathbf{[acid]}}\right)$$

As [acetic acid] = [sodium acetate], $\mathbf{log_{10}}\left(\frac{\mathbf{0.010}}{\mathbf{0.010}}\right) = \mathbf{log_{10}(1) = 0}$ **and so**

$$\mathbf{pH = p}K_\mathbf{a} \mathbf{= -log(6.4 \times 10^{-5}) = 4.19}$$

Answer: **4.19**

Would this solution make a good buffer system? Give reasons for your answer?

Because the concentrations of weak acid and conjugate base are equal, this solution is a good buffer system. (Good buffers require this ratio to be between 0.1 and 10.)

As the concentrations are only 0.010 M, the buffer does not have a very great capacity. It will buffer effectively for small amounts of added H^+ or OH^-, but large amounts will quickly cause the weak acid/conjugate base ratio to move outside the 0.1-10 range.

- The gases NO_2 and N_2O_4 are in equilibrium according to the following equation. **2**

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = +57 \text{ kJ mol}^{-1}$$

In which direction will the reaction move when the following changes are made?

The pressure is increased by decreasing the volume.

The reaction involves 1 mol → 2 mol so the system will react to an increase in pressure by shifting to lower to the left to reduce it: shift to reactants.

The temperature is increased.

The reaction is endothermic so the system will react to an increase in temperature by shifting to the right to reduce it: shift to products.

Marks

- Quinine is a natural product that has anti-malarial properties. It was originally extracted for therapeutic use from the bark of the cinchona tree, but is now synthesised by the pharmaceutical industry. Quinine is not very soluble in water and is generally administered as the more soluble hydrochloride salt ($C_{20}H_{24}N_2O_2{\cdot}HCl$). The p$K_a$ of this salt is 4.32. What is the pH of a 0.053 M solution of quinine hydrochloride? **3**

As ascorbic acid is a weak acid, $[H_3O^+]$ must be calculated:

	acid	H_2O	$\rightleftharpoons$	H_3O^+	base
initial	**0.053**	**large**		**0**	**0**
change	**$-x$**	**negligible**		**$+x$**	**$+x$**
final	**$0.053 - x$**	**large**		**x**	**x**

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+(aq)][base]}{[acid]} = \frac{x^2}{(0.053 - x)}$$

As pK_a = 4.32, $K_a = 10^{-4.32}$ and is very small, $0.053 - x \sim 0.053$ and hence:

$$x^2 = 0.053 \times (10^{-4.32}) \text{ or } x = 1.59 \times 10^{-3}\text{ M} = [H_3O^+(aq)]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H_3O^+(aq)] = -\log_{10}[0.00159] = 2.80$$

Answer: **pH = 2.80**

- Use chemical equations to illustrate how $HPO_4^{2-}/H_2PO_4^-$ can act as a buffer. **2**

The HPO_4^{2-}(aq) acts as a base and can take up added H^+(aq):

$$HPO_4^{2-}(aq) + H_3O^+(aq) \rightarrow H_2PO_4^-(aq) + H_2O(l)$$

The $H_2PO_4^-$(aq) acts as an acid and can take up added OH^-(aq):

$$H_2PO_4^-(aq) + OH^-(aq) \rightarrow HPO_4^{2-}(aq) + H_2O(l)$$

The system thus has the capacity to maintain the pH.

Marks 6

- Butyric acid, $CH_3CH_2CH_2COOH$, is found in rancid butter and parmesan cheese. The pK_a of butyric acid is 4.83.
  (a) What is the pH of a 0.10 M water solution of butyric acid?

As $pK_a = -\log K_a = 4.83$, $K_a = 10^{-4.83}$. Denoting butyric acid as HA, the initial concentration of [HA(aq)] = 0.10 M. The reaction table is then:

	[HA(aq)]	⇆	$[H^+(aq)]$	$[A^-(aq)]$
t = 0	**0.10.**		**0**	**0**
change	**-x**		**+x**	**+x**
equilibrium	**0.10 – x**		**x**	**x**

Hence, $$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]} = \frac{(x)(x)}{0.10-x} = \frac{x^2}{0.10-x}$$

As K_z is small, the amount of dissociation, x, is also small so 0.10 – x ~ 0.10. Using this approximation, $K_z = \dfrac{x^2}{0.10} = 10^{-4.83}$ **hence $x = 1.22 \times 10^{-3}$ M.**

As $x = [H^+(aq)]$, $pH = -\log[H^+(aq)] = -\log(1.22 \times 10^{-3}) = 2.92$

Answer: **2.92**

(b) Calculate the pH of the solution formed when 0.050 mol of NaOH(s) is added to 1.0 L of 0.10 M butyric acid.

As NaOH is a strong base, it will dissociate completely and each mole of OH^- will react with butyric acid to form one mole of $A^-(aq)$.

1.0 L of 0.10 M HA contains 0.10 mol. After addition of 0.050 mol of OH^-, the number of moles of HA = (0.10 – 0.050) = 0.05 mol and the number of moles of A^- = 0.050 mol.

As 1.0 L of solution is present, [HA(aq)] = 0.05 M and $[A^-(aq)]$ = 0.05 M. Substituting into the expression for K_a gives:

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]} = \frac{[H^+(aq)]\times(0.05)}{(0.05)} = 10^{-4.83} \text{ so } [H^+(aq)] = 1.5 \text{ M}$$

Hence, $pH = -\log[H^+(aq)] = 4.83$

Answer: **4.83**

(c) Using equations, comment on how the final solution in (b) will respond to additions of small amounts of acid or base in comparison to 1 L of water.

Solution (b) consists of a mixture of a weak acid and its conjugate base: it is a buffer system and will resist changes in pH. If acid is added, the system can respond by removing it using A^-:

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

If base is added, the system can respond by removing it using HA:

$$OH^-(aq) + HA(aq) \rightarrow H_2O(l) + A^-(aq)$$

Marks 6

- Lactic acid, $CH_3CHOHCOOH$, is produced in the body during normal exercise. It is a monoprotic acid with a pK_a of 3.86.

(a) What is the pH of a 0.10 M water solution of lactic acid?

The reaction table is:

	lactic acid(aq)	H_2O(l)	$\rightleftharpoons$	lactate(aq)	H_3O^+(aq)
start	**0.10**	**large**		**0**	**0**
change	**-x**	**-x**		**+x**	**+x**
equilibrium	**0.10-x**	**large**		**x**	**x**

As $pK_a = -\log_{10}(K_a) = 3.86$, $K_a = 10^{-3.86}$ and:

$$K_a = \frac{[\text{lactate(aq)}][H_3O^+(aq)]}{[\text{lactic acid(aq)}]} = \frac{(x)(x)}{(0.10-x)} = \frac{x^2}{(0.10-x)} = 10^{-3.86}$$

As K_a is very small, x is tiny and 0.10 – x ~ x. Hence,

$$K_a \sim \frac{x^2}{(0.10)} = 10^{-3.86} \text{ or } x^2 = (0.10) \times (10^{-3.86}) \text{ so } x = [H_3O^+(aq)] = 3.72 \times 10^{-3}\ M$$

As $pH = -\log_{10}([H_3O^+(aq)]$:

$$pH = -\log_{10}(3.72 \times 10^{-3}) = 2.43$$

Answer: **2.43**

(b) Calculate the pH of the solution formed when 0.02 mol of $Ca(OH)_2$(s) is added to 1.0 L of 0.10 M lactic acid.

1.0 L of 0.10M lactic acid contains 0.10 mol of acid.

$Ca(OH)_2$ is a strong base. It will dissociate completely to give $2OH^-$(aq) for every 1 mole of $Ca(OH)_2$. (2 × 0.02) = 0.04 mol of OH^-(aq) will be produced. This will neutralize 0.04 mol of the acid leaving (0.10 – 0.04) = 0.06 mol of acid. Assuming that the volume does not change from the addition of the solid, [lactic acid] = 0.060 M. The neutralization produces lactate anion with [lactate] = 0.040 M.

The solution now contains acid and its conjugate base. It is a buffer and the pH can be calculated using the Henderson-Hasselbalch equation can be used:

$$pH = pK_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = 3.86 + \log_{10}\left(\frac{0.040}{0.060}\right) = 3.86$$

Answer: **3.86**

ANSWERS CONTINUES ON THE NEXT PAGE

(c) Using equations, comment on how the final solution in (b) will respond to additions of small amounts (e.g. less than 0.01 mol) of acid or base in comparison to additions of the same amounts of acid or base to 1 L of water.

The solution in (b) will act as a buffer. As it contains both an acid (lactic acid) and a base (lactate), it can react with both added base and acid to maintain a near constant pH.

Added H_3O^+ will be consumed by the reaction of the lactate anion:

lactate(aq) + H_3O^+(aq) → lactic acid(aq) + H_2O(l)

Added OH^- will be consumed by the reaction of the lactic acid:

lactic acid(aq) + OH^-(aq) → lactate(aq) + H_2O(l)

As long as the amounts of lactic acid and lactate are large in comparison to the added acid or base, the pH is approximately constant and is described by the Henderson-Hasselbalch equation.

If acid or base is added to water, the [H_3O^+(aq)] or [OH^-(aq)] will change according to the amount added and the pH will change rapidly.

Marks

- Codeine, a cough suppressant extracted from crude opium, is a weak base with a pK_b = 5.79. What is the pH of a 0.020 M solution of codeine? **2**

As codeine, is a weak base and so [OH$^-$] must be calculated. for example using a reaction table:

	codeine	H_2O	$\rightleftharpoons$	codeineH^+	OH^-
initial	**0.020**	**large**		**0**	**0**
change	**-x**	**negligible**		**+x**	**+x**
final	**0.020 – x**	**large**		**x**	**x**

The equilibrium constant K_b is given by:

$$K_b = \frac{[\text{codeineH}^+][\text{OH}^-]}{[\text{codeine}]} = \frac{x^2}{(0.020 - x)}$$

As pK_b = -log$_{10}K_b$ = 5.79, K_b = $10^{-5.79}$. Hence,

$$\frac{x^2}{(0.020 - x)} = 10^{-5.79}$$

As K_b is very small, 0.020 – x ~ 0.020 and hence:

$x^2 = 0.020 \times 10^{-5.79}$ **or** $x = 1.8 \times 10^{-4}$ **M = [OH$^-$(aq)]**

Hence, the pOH is given by:

pOH = –log$_{10}$[OH$^-$] = –log$_{10}$[1.8×10^{-4}] = 3.74

Finally, pH + pOH = 14.00 so

pH = 14.0 – 3.74 = 10.26

Answer: **10.26**

- A buffer solution is formed with 0.250 M CH_3COOH and 0.350 M CH_3COONa. What is the pH of this buffer solution? (K_a of acetic acid = 1.8×10^{-5} M.) **3**

The pH of a buffer solution is given by the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log\left(\frac{\text{base}}{\text{acid}}\right)$$

As pK_a = -log$_{10}K_a$ = -log$_{10}$(1.8×10^{-5}) = 4.74. With [base] = [CH_3COONa] = 0.350 M and [acid] = [CH_3COOH] = 0.250 M,

ANSWER CONTINUES ON THE NEXT PAGE

$$\mathrm{pH} = 4.74 + \log\left(\frac{0.350}{0.250}\right) = 4.89$$

As the buffer contains a higher concentration of base than acid, the pH > pK_a.

Answer: **4.89**

Calculate the pH of the solution formed when 6.3×10^{-2} mol of NaOH is added to 1.0 L of the buffer solution.

In a 1.0 L solution of the buffer, there is 0.250 mol of CH_3COOH and 0.350 mol of CH_3COO^-.

The OH^- will react with the CH_3COOH to produce CH_3COO^-. The concentration of the former will therefore decrease whilst the concentration of the latter will increase. After the OH^- is added:

number of moles of CH_3COOH = $(0.250 - 6.3 \times 10^{-2})$ M = 0.187 mol
number of moles of CH_3COO^- = $(0.350 + 6.3 \times 10^{-2})$ M = 0.413 mol

As the volume of solution does not change, these are also the new acid and base concentrations. Hence, the buffer now has:

$$\mathrm{pH} = 4.74 + \log\left(\frac{0.413}{0.187}\right) = 5.09$$

As base has been added, there is an increase in the pH. As it is being added to a buffer system, this change is small. Addition of this quantity of base to water would increase the pH by 1.20 units.

Answer: **5.09**

Marks 6

- Write the balanced chemical equation for the dissolution of solid $Ca(CH_3CO_2)_2$ in water.

$Ca(CH_3CO_2)_2(s) \rightarrow Ca^{2+}(aq) + 2CH_3CO_2^-(aq)$

What is the pH of a solution that has 158.2 g of $Ca(CH_3CO_2)_2$ dissolved in 1.000 L of water? The pK_a of acetic acid, CH_3COOH, is 4.76.

The molar mass of $Ca(CH_3CO_2)_2$ is:

molar mass = (40.08 (Ca) + 4×12.01 (C) + 6×1.008 (H) + 4×16.00 (O)) g mol^{-1}
= 158.168 g mol-1

Thus, 158.2 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{158.2\text{ g}}{158.168\text{ g mol}^{-1}} = 1.000\text{ mol}$$

When $Ca(CH_3CO_2)_2$ is dissolved, it produces $Ca^{2+}(aq) + 2CH_3CO_2^-$. Hence, if 1.000 mol of $Ca(CH_3CO_2)_2$ is dissolved in 1.0 L, $[CH_3CO_2^-]_{initial}$ = 2.000 M.

As $CH_3CO_2^-$ is a weak base, $[CH_3CO_2^-]$ must be calculated by considering the equilibrium:

	$CH_3CO_2^-$	H_2O	$\rightleftharpoons$	CH_3CO_2H	OH^-
initial	**2.000**	**large**		**0**	**0**
change	$-x$	**negligible**		$+x$	$+x$
final	$2.000 - x$	**large**		x	x

The equilibrium constant K_b is given by:

$$K_b = \frac{[CH_3CO_2H][OH^-]}{[CH_3CO_2^-]} = \frac{x^2}{(2.000-x)}$$

For an acid and its conjugate base:

pK_a + pK_b = 14.00 so pK_b = 14.00 – 4.76 = 9.24

As pK_b = 9.24, $K_b = 10^{-9.24}$. K_b is very small so 2.000 – x ~ 2.000 and hence:

$x^2 = 2.000 \times 10^{-9.24}$ or y = 0.0000393 M = $[OH^-]$

Hence, the pOH is given by:

pOH = $-\log_{10}[OH^-] = \log_{10}[0.0000393]$ = 4.47

Finally, pH + pOH = 14.00 so

pH = 14.000 – 4.47 = 9.53

pH = **9.53**

ANSWER CONTINUES ON THE NEXT PAGE

Calculate the pH of this solution after the addition of 0.250 mol of HCl gas?

The solution contain 2.000 mol of $CH_3CO_2^-$. This will react with 0.250 mol of HCl(g) to produce 0.250 mol of CH_3CO_2H, leaving (2.000 – 0.250) mol = 1.750 mol of $CH_3CO_2^-$ in unreacted.

As the solution has a volume of 1.000 L, $[CH_3CO_2H]$ = 0.250 M and $[CH_3CO_2^-]$ = 1.750 M.

The solution contains a weak acid and its conjugate base. The pH of this buffer solution can be calculated using the Henderson-Hasselbalch equation;

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]} = 4.76 + \log\frac{1.750}{0.250} = 5.61$$

pH = **5.61**

- Both HCO_3^-(aq) and CO_2(aq) are present in human blood. How does their presence ensure that the pH of blood is maintained at ~7.2, even if H^+(aq) or OH^-(aq) are produced by processes in the body? **4**

CO_2(aq) and HCO_3^-(aq) constitute a buffer system:

$$\mathbf{CO_2(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H^+(aq)}$$

Excess H^+ is removed by:

$$\mathbf{HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(aq) + H_2O(l)}$$

Excess OH^- is removed by:

$$\mathbf{CO_2(aq) + OH^-(aq) \rightarrow HCO_3^-(aq)}$$

How does hyperventilation (very rapid breathing) interfere with this balance? What is the effect?

Hyperventilation removes CO_2 from the blood. This shifts the above buffer equilibrium to the left (Le Chatelier's principle), [H^+(aq)] decreases and the blood pH increases.

- The pH value of pure water at 25 °C is 7.00. How, if at all, does that value change when the temperature is changed to 37 °C (a person's body temperature)? Explain. **3**

The auto-ionisation of water is an endothermic reaction

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \qquad \Delta H > 0.$$

Increasing the temperature will push this reaction to the right (Le Chatelier's principle), so the $[H_3O^+(aq)]$ will increase and the pH will therefore *decrease*.

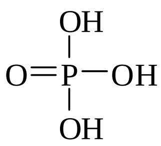
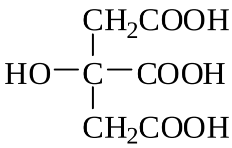
Is pure water at 37 °C acidic, basic or neutral? Circle your choice.

acidic basic **neutral**

(Pure water is neutral as $[H_3O^+(aq)] = [OH^-(aq)]$. This is the criterion for a neutral solution NOT pH = 7.)

Marks
3

- Consider the two triprotic acids, phosphoric acid and citric acid.

Acid	Formula	K_{a1}	K_{a2}	K_{a3}
phosphoric	H_3PO_4	7.1×10^{-3}	6.3×10^{-8}	4.5×10^{-13}
citric	$C_6H_8O_7$	7.1×10^{-4}	1.7×10^{-5}	6.4×10^{-6}

Explain why $K_{a1} > K_{a2} > K_{a3}$ for both acids.

It is more difficult to remove a proton from a negatively charged species, so $K_{a1} > K_{a2} > K_{a3}$ for all acids. *i.e.* K_{a1}, K_{a2} and K_{a3} correspond to removal of H^+ from H_3PO_4, $H_2PO_4^-$ and HPO_4^{2-} respectively. This process gets harder and harder because a positively charged proton is having to be removed from a negatively charged molecule.

For phosphoric acid, the K_a values differ by about 5 orders of magnitude while for citric acid there is a much smaller difference. Explain.

The number of resonance structures for the various conjugate bases are:

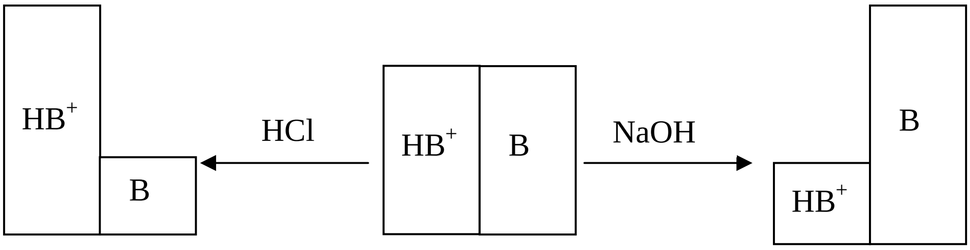
2 for $H_2PO_4^-$; 3 for HPO_4^{2-}; and 4 for PO_4^{3-}
2 for $C_6H_7O_7^-$; 4 for $C_6H_6O_7^{2-}$; and 8 for $C_6H_5O_7^{3-}$

The conjugate bases for citric acid are more stable (because they have greater resonance stabilisation) so the corresponding acids are all stronger.

Alternatively, the increasing negative charges in the conjugate bases are being formed in different parts of the molecule in the case of citric acid, whereas they are all very close to each other in phosphoric acid. Again, the formation of the conjugate base series of citric acid is therefore easier and the acids are therefore stronger.

Marks 7

- A buffer system with a weak base B and its conjugate acid HB^+ is shown in the diagram below with equal concentrations. Complete the diagram by showing the relative concentrations after the addition of some HCl or NaOH.



Write down the balanced net ionic equations for both these reactions.

$HB^+(aq) + OH^-(aq) \rightarrow B + H_2O(l)$
$B(aq) + H^+(aq) \rightarrow HB^+(aq)$

Calculate the pH of a buffer if it contains 0.200 mol of $NaNO_2$ and 0.300 mol of HNO_2 in 1.00 L of water. The pK_a of HNO_2 is 3.15.

The concentrations of acid (HNO_2) and base (NO_2^-) are:

concentration of acid = number of moles / volume
= (0.300 mol / 1.00 L) = 0.300 mol L^{-1}

concentration of base = (0.200 mol / 1.00 L) = 0.200 mol L^{-1}

The pH of this buffer can then be calculated using the Henderson-Hasselbalch equation:

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{base}]}{[\mathrm{acid}]} = 3.15 + \log\frac{0.200}{0.300} = 2.97$$

pH = **2.97**

What is the pH if (a) 0.05 mol of HCl(g) and (b) 0.25 mol of HCl(g) is added?

HCl will react with the NO_2^- to form HNO_2.

(a) 0.05 mol of HCl(g) will react with 0.05 mol of the NO_2^- present to form an additional 0.05 mol of HNO_2. The amount of NO_2^- will decrease by 0.05 mol:

concentration of acid = (0.35 mol / 1.00 L) = 0.35 mol L^{-1}

concentration of base = (0.15 mol / 1.00 L) = 0.15 mol L^{-1}

ANSWER CONTINUES ON THE NEXT PAGE

Hence:

$$\mathbf{pH = p}\boldsymbol{K}_{\mathbf{a}} \mathbf{+ log}\frac{[\mathbf{base}]}{[\mathbf{acid}]} \mathbf{= 3.15 + log}\frac{\mathbf{0.15}}{\mathbf{0.35}} \mathbf{= 2.78}$$

(b) 0.25 mol of HCl(g) will react *all* of the NO_2^-. As there is only 0.200 mol of NO_2^- present, 0.05 mol of HCl will remain unreacted. As HCl is a strong acid, this will completely ionize to give $[H^+(aq)]$ = 0.05 mol L^{-1}. Hence:

$$\mathbf{pH = -log_{10}[H^+(aq)] = -log_{10}(0.05) = 1.30}$$

(a) pH = **2.78**	(b) pH = **1.30**

Marks
5

- The concentration of a dissolved gas is related to its partial pressure by c = *k*p. What is the concentration of CO_2 dissolved in blood if the partial pressure of CO_2 in the lungs is 0.053 atm? The *k* for CO_2 is 0.034 mol L^{-1} atm^{-1}.

Using *c* = *k*p,

$$c = (0.034 \text{ mol L}^{-1} \text{ atm}^{-1})(0.053 \text{ atm}) = 0.0018 \text{ mol L}^{-1}$$

Answer: **0.0018 mol L^{-1}**

Calculate the pH of blood if all of this CO_2 reacted to give H_2CO_3.
The K_a of H_2CO_3 is 4.5 × 10^{-7}.

If [H_2CO_3(aq)] = 0.0018 mol L^{-1}, the pH can be calculated using the reaction table:

	H_2CO_3	H_2O	$\rightleftharpoons$	H_3O^+	HCO_3^-
initial	0.0018	large		0	0
change	$-x$	negligible		$+x$	$+x$
final	$0.0018 - x$	large		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{x^2}{0.0018 - x}$$

As K_a = 4.5 × 10^{-7} and is very small, $0.0018 - x \sim 0.0018$ and hence:

$$x^2 = 0.0018 \times (4.5 \times 10^{-7}) \quad \text{or} \quad x = 2.8 \times 10^{-5} \text{ M} = [H_3O^+]$$

Hence:

$$\text{pH} = -\log_{10}[H_3O^+(aq)] = -\log_{10}(2.8 \times 10^{-5}) = 4.54$$

Answer: **4.54**

Hyperventilation results in a decrease in the partial pressure of CO_2 in the lungs. What effect will this have on the pH of the blood? Use a chemical equation to illustrate your answer.

If the CO_2 partial pressure decreases, the equilibrium below will shift to the left. This will decrease [H^+(aq)] and the pH will increase.

$$CO_2(aq) + H_2O \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$$

ANSWER CONTINUES ON THE NEXT PAGE

The pH of blood is maintained around 7.4 by the H_2CO_3 / HCO_3^- buffer system. Explain how a buffer works, illustrating your answer with chemical equations.

A buffer resists changes in pH. It contains substantial quantities of a weak acid and its conjugate base. In the H_2CO_3/HCO_3^- buffer, added acid is removed by the reaction:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq)$$

Added base is removed by the reaction:

$$H_2CO_3(aq) + OH^-(aq) \rightarrow HCO_3^-(aq) + H_2O$$