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,

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

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 $Ca(OH)_2$. $(2 \times 0.02) = 0.04$ 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:

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

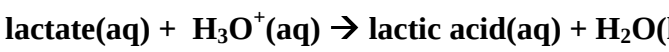
Answer: **3.86**

ANSWERS CONTINUES ON THE NEXT PAGE

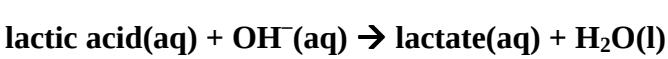
(c) Using equations, comment on how the final solution in (b) will respond to additions of small amounts (e.g. less than 0.01 mol) of acid or base in comparison to additions of the same amounts of acid or base to 1 L of water.

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:



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



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.