Acid bases Practice worksheet

A hydrogen sulfate/sulfate system is represented by the following equation.

$$HSO_4^-(aq) + H_2O(\ell) \leftrightharpoons SO_4^{2-}(aq) + H_3O^+(aq)$$

(a) Predict how

- the forward reaction rate and
- the pH

will differ from their original values after the following changes are imposed on the system and equilibrium has been re-established. Use the terms **increase**, **decrease**, **no change**.

(6 marks)

Change imposed by the addition of	**Effect on forward reaction rate when equilibrium is re-established**	**Effect on pH when equilibrium is re-established**
a few drops of concentrated hydrochloric acid		
a few drops of concentrated lead(II) nitrate solution		
distilled water		

(b) The reaction in part (a) is endothermic in the forward direction as written. Predict what will happen to the pH when the temperature is increased. Justify this prediction. (4 marks)

Question 29 **(10 marks)**

In a beaker 12.00 mL of 0.0334 mol L^{-1} sulfuric acid solution, $H_2SO_4(aq)$, is added to 32.50 mL of 0.0288 mol L^{-1} potassium hydroxide solution, KOH(aq).

(a) Identify the limiting reagent in this reaction. Show **all** workings. (5 marks)

(b) Calculate the final concentration of the excess reagent. Show **all** workings. (3 marks)

(c) Calculate the pH of the final solution. Show **all** workings. (2 marks)

Question 31 **(8 marks)**

Water is capable of self-ionisation.

(a) Write an equation for the self-ionisation of water. (2 marks)

(b) Write the equilibrium constant expression for the self-ionisation of water. (1 mark)

(c) The equilibrium constant for the self-ionisation of water K_w is 1.00×10^{-14} at 25 °C. What does this value indicate about this reaction? (1 mark)

The K values for the self-ionisation of water at 100.0 kPa are given here for a number of different temperatures.

Temperature (°C)	K value
0	0.114×10^{-14}
25	1.00×10^{-14}
50	5.48×10^{-14}
75	19.9×10^{-14}
100	51.3×10^{-14}

(d) Calculate the pH of water at 50 °C. (2 marks)

(e) Is water acidic, basic or neutral at 50 °C? State a reason for your answer. (2 marks)

Question 26 **(10 marks)**

Solid copper(II) hydroxide is added to excess 0.100 mol L^{-1} carbonic acid solution.

(a) Write the balanced equation, with appropriate state symbols, for the reaction that takes place between the copper(II) hydroxide and carbonic acid. (3 marks)

(b) Predict **all** visible changes that would be observed, if any, while the reactants are mixed together and afterwards. (3 marks)

(c) Predict **two** observations that would be different if excess 0.100 mol L^{-1} hydrochloric acid was used instead of the 0.100 mol L^{-1} carbonic acid. (2 marks)

One: _____

Two: _____

(d) State **two** personal safety measures the experimenter should take when conducting these experiments. (2 marks)

One: _____

Two: _____

Question 27 **(12 marks)**

Phosphoric acid, $H_3PO_4(aq)$, is a weak, triprotic acid.

(a) Write the ionisation equation for phosphoric acid in water which shows the **second** proton of the acid being released into solution. (2 marks)

Magnesium carbonate, $MgCO_3(s)$, is an ingredient of a commonly-used antacid.

(b) Other than water, list **three** species (elements, compounds, ions) that would be found in the reacting vessel open to the atmosphere at the completion of the reaction between excess solid magnesium carbonate and an aqueous solution of phosphoric acid. (3 marks)

One: _____

Two: _____

Three: _____

Sodium hydroxide solution, NaOH(aq), was used in a titration to determine the concentration of phosphoric acid.

(c) Other than it having too low a molar mass, state **two** reasons why the concentration of the sodium hydroxide solution cannot be reliably determined by weighing out an amount of solid sodium hydroxide and dissolving it in a known volume of distilled water. (2 marks)

One: _____

_____

_____

Two: _____

_____

_____

Question 29 **(15 marks)**

Wines and other alcoholic drinks can spoil when the alcohol (ethanol) they contain oxidises to acetic acid (ethanoic acid). An acidity regulator, monosodium citrate, is often added to drinks to prevent the formation of acetic acid. The monosodium citrate does this by acting as a buffer.

A citric acid/dihydrogen citrate ion buffer can be prepared from citric acid, $H_3C_6H_5O_7$ and monosodium citrate, $NaH_2C_6H_5O_7$.

(a) Write an equation for the buffer system ($H_3C_6H_5O_7$ / $H_2C_6H_5O_7^-$) containing citric acid, $H_3C_6H_5O_7$ and monosodium citrate, $NaH_2C_6H_5O_7$. (2 marks)

Buffers that contain equal concentrations of both components are most effective. This buffer solution is prepared by mixing 100.0 mL of citric acid solution with 100.0 mL of monosodium citrate solution. The citric acid solution, $H_3C_6H_5O_7(aq)$, has a concentration of 0.200 mol L^{-1}.

(b) Calculate the mass of sodium citrate, $NaH_2C_6H_5O_7$, that would need to be dissolved in 100.0 mL of distilled water to make the most effective buffer solution. (3 marks)

(c) If a citric acid buffer was prepared to a pH of 3.5, what would be the concentration of the hydroxide ion at 25.0 °C? (3 marks)

(d) Explain why only a small change in pH is observed in this buffer solution when a small amount of sodium hydroxide solution is added, compared to adding a similar amount of sodium hydroxide solution to a system that is not a buffer solution. Your answer should refer to the buffer equilibrium in part (a). (4 marks)

1. Consider the following data:

Barbituric acid $HC_4H_3N_2O_3$ Ka = 9.8 x 10^{-4} at 25°C

If you were to put 2 .0 M of $HC_4H_3N_2O_3$ into pure distilled water what would:

a) The pH of the solution be to 2 significant figures [**3 marks**, -1 for incorrect sig figs]

b) Calculate the percentage ionisation of Barbituric acid, to 2 significant figures [**2 marks**, -1 for incorrect sig figs]