ACID/BASE I and II

Q6. Describe the common buffer systems present in industrial, environmental, or biological systems.

Industrial: Controlling pH levels to optimize yields and cut-off undesirable side reactions. Also important in protecting food from spoilage.

Buffered aspirin, Bufferin, is produced based on the fact that a higher pH leads to a faster tablet breakdown and absorption.

Biological:
$H_2CO_{3(aq)}$ / $HCO_3^-{}_{(aq)}$ buffers human blood plasma.

$H_2PO_4^-{}_{(aq)}$ / $HPO_4^{2-}{}_{(aq)}$ buffers human cells.

R1. Write equations representing the formation of acidic solutions or basic solutions from non-metal and metal oxides.

In general: non-metal oxides form acidic solutions; metal oxides form basic solutions (exception: see *N3*.)

Examples:

$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$
$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$

Related Questions: 50, 51

R2. Describe the pH conditions for rain to be called acid rain.

R3. Relate the pH of normal rain water to the presence of dissolved CO_2.

Normal rain has pH = 5.6 with the $CO_{2(aq)}$ from non-human activity. Rain with a lower pH is called acid rain.

R4. Describe sources of NO_x and SO_x.

NO_x: motor vehicle emissions

SO_x: coal-fired power generating stations (sulfur in the coal), non-ferrous ore smelters (sulfide ores)

Related Question: 52

R5. Discuss general environmental problems associated with acid rain.

- leaching of heavy metal ions into lakes on granitic strata
- aquatic life and waterfowl populations threatened
- possible damage to human health, crops, and forests
- many other social problems are likewise related.

Related Question: 53

1. Which of the following is a property of all acidic solutions at 25°C?

   A. They have a pH less than 7.0.

   B. They have a pH greater than 7.0. I

   C. They cause phenolphthalein to turn pink.

   D. They release hydrogen when placed on copper metal.

   Source: August 2003

2. Which of the following is a common property of acid solutions?

   A. They have a pH > 7.

   B. They turn red litmus blue. I

   C. They have a slippery feeling.

   D. They turn pink phenolphthalein colourless.

   Source: January 2004

3. When a small solid sample is added to a solution of H_2SO_4, a precipitate forms and the solution becomes less acidic. Which of the following substances could have caused these results? I

   A. Na_2SO_4 **B.** $Sr(OH)_2$

   C. $Mg(OH)_2$ **D.** $Ca(NO_3)_2$

   Source: August 2003

4. Which net ionic equation best describes the reaction between NaOH and H_2S?

   A. $OH^-_{(aq)} + H^+_{(aq)} \rightleftarrows H_2O_{(l)}$

   B. $2OH^-_{(aq)} + H_2S_{(aq)} \rightleftarrows 2H_2O_{(l)} + S^{2-}_{(aq)}$

   II C. $2NaOH_{(aq)} + H_2S_{(aq)} \rightleftarrows 2H_2O_{(l)} + Na_2S_{(aq)}$

   D. $2Na^+_{(aq)} + 2OH^-_{(aq)} + 2H^+_{(aq)} + S^{2-}_{(aq)} \rightleftarrows 2H_2O_{(l)} + 2Na^+_{(aq)} + S^{2-}_{(aq)}$

Source: April 2004

5. Which of the following is a general characteristic of Arrhenius acids?

   A. They produce H^+ in solution.

   B. They accept an H^+ from water.

   I C. They turn bromthymol blue a blue colour.

   D. They react with H_3O^+ ions to produce H_2.

Source: April 2004

6. Select the equation that best represents the reaction of CH_3NH_2 acting as a base with water.

   A. $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftarrows CH_3NH^+_{3(aq)} + OH^-_{(aq)}$

   I B. $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftarrows CH_3NH^-_{(aq)} + H_3O^+_{(aq)}$

   C. $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftarrows CH_3NH_2OH^-_{(aq)} + H^+_{(aq)}$

   D. $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftarrows CH^+_{3(aq)} + NH_{3(aq)} + OH^-_{(aq)}$

Source: January 2004

7. Consider the following reaction:
   $HCN + CH_3NH_2 \rightleftarrows CN^- + CH_3NH_3^+$

   Which of the following describes a conjugate acid-base pair in the equilibrium above?

	Acid	Base
A.	CN^-	HCN
B.	$CH_3NH_3^+$	CN^-
C.	HCN	$CH_3NH_3^+$
D.	$CH_3NH_3^+$	CH_3NH_2

I

Source: August 2003

8. Identify a conjugate pair from the equilibrium provided:
   $PO_4^{3-} + HCO_3^- \rightleftarrows HPO_4^{2-} + CO_3^{2-}$

   A. CO_3^{2-} and PO_4^{3-}

   B. PO_4^{3-} and HCO_3^- I

   C. PO_4^{3-} and HPO_4^{2-}

   D. HCO_3^- and HPO_4^{2-}

Source: April 2004

9. Which of the following solutions will show the greatest electrical conductivity?

   A. 0.1 M HCl B. 0.5M H_2CO_3 I

   C. 0.5M H_3BO_3 D. 0.1M $H_2C_2O_4$

Source: January 2004

10. Which of the following best describes a weak acid?

    A. Its 0.10 M solution will have pH = 1.00.

    B. It may be very soluble, but only partly ionized. I

    C. It must be very soluble and completely ionized.

    D. It must be of low solubility and completely ionized.

Source: April 2004

11. Which of the following is the weakest base?

A. F^-

B. HS^- I

C. CN^-

D. IO_3^-

Source: August 2003

12. When comparing 0.10 M HPO_4^{2-} and 0.10 M $HC_2O_4^-$ as acids, which of the following is true?

A. $HC_2O_4^-$ is weaker and its pH is larger.

I **B.** HPO_4^{2-} is stronger and its pH is larger.

C. HPO_4^{2-} is weaker and its pH is smaller.

D. $HC_2O_4^-$ is stronger and its pH is smaller.

Source: January 2004

13. Which of the following will have the smallest K_b value?

A. IO_3^-

II **B.** NH_3

C. CN^-

D. HPO_4^{2-}

Source: January 2004

CHALLENGER QUESTION

II **14.** Which of the following solutions will have the lowest $[OH^-]$?

A. $NaF_{(aq)}$ **B.** $NaCl_{(aq)}$

C. $NaHCO_{3(aq)}$ **D.** $NaHPO_{4(aq)}$

Source: April 2004

CHALLENGER QUESTION

15. Water has the greatest tendency to act as an acid with which of the following?

A. Cl^-

B. NO_2^- I

C. $H_2PO_4^-$

D. CH_3COO^-

Source: April 2004

16. Which of the following relationships is used to calculate K_w at 30°C?

A. $K_w = pH + pOH$

B. $pK_w = -\log[H_3O^+]$ I

C. $K_w = [H_3O^+][OH^-]$

D. $K_w = [H_3O^+] + [OH^-]$

Source: August 2003

17. Which of the following statements is true for an acidic solution at 25°C?

A. $pH > 7.0$

B. $pOH < 7.0$ I

C. $[H_3O^+] < [OH^-]$

D. $[H_3O^+] > [OH^-]$

Source: April 2004

18. What is the $[OH^-]$ in 0.024 M HCl?

A. 2.5×10^{-16} M

B. 4.0×10^{-13} M I

C. 2.5×10^{-2} M

D. 2.5×10^{12} M

Source: April 2004

19. Which of the following equations can be used to calculate pOH?

I

A. $pOH = -\log K_w$

B. $pOH = pK_w + pH$

C. $pOH = pK_w - pH$

D. $pOH = -\log[H_3O^+]$

Source: January 2004

20. What is the pOH of 0.2 M HNO_3?

I

A. 5×10^{-14} **B.** 0.2

C. 0.7 **D.** 13.3

Source: August 2003

21. What is a general characteristic of all Brønsted-Lowry bases?

I

A. They all accept H^+.

B. They all accept OH^-.

C. They will turn litmus a pink colour.

D. They will react with acids to produce H^2 gas.

Source: January 2004

22. What is the equilibrium expression for the predominant equilibrium in $NaHCO_{3(aq)}$?

II

A. $K_a = \dfrac{[HCO_3^-]}{[H_3O^+][CO_3^{2-}]}$

B. $K_b = \dfrac{[HCO_3^-]}{[H_2CO_3][OH^-]}$

C. $K_a = \dfrac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]}$

D. $K_b = \dfrac{[H_2CO_3][OH^-]}{[HCO_3^-]}$

Source: April 2004

23. Which of the following K_a values represents the acid with the strongest conjugate base?

II

A. $K_a = 4.2 \times 10^{-12}$

B. $K_a = 9.5 \times 10^{-9}$

C. $K_a = 2.0 \times 10^{-5}$

D. $K_a = 7.8 \times 10^{-3}$

Source: August 2003

24. What is true about an acid that has a large K_a value?

II

A. The acid is weak.

B. The acid is strong.

C. The acid has a large K_b value.

D. The acid has a large pH value.

Source: April 2004

25. What is the K_b value for $HC_6H_5O_7^{2-}$?

II

A. 1.0×10^{-14} **B.** 5.9×10^{-10}

C. 2.4×10^{-8} **D.** 4.1×10^{-7}

Source: January 2004

[illegible]

26. What is the dissociation equation for Na_2CO_3 in water?

II or I

A. $Na_2CO_{3(s)} \rightarrow Na^{2+}_{(aq)} + CO^{2-}_{3(aq)}$

B. $Na_2CO_{3(s)} \rightarrow 2Na^{+}_{(aq)} + CO^{2-}_{3(aq)}$

C. $CO^{2-}_{3(aq)} + H_2O_{(l)} \rightarrow HCO^{-}_{3(aq)} + OH^{-}_{(aq)}$

D. $Na_2CO_{3(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_2CO_{3(aq)}$

Source: August 2003

27. Which of the following describes the dissociation of calcium chloride?

Solubility

A. $CaCl_{(s)} \rightarrow Ca^{+}_{(aq)} + Cl^{-}_{(aq)}$

B. $Ca_2Cl_{(s)} \rightarrow Ca^{+}_{2\,(aq)} + Cl^{-}_{(aq)}$

C. $CaCl_{2(s)} \rightarrow Ca^{2+}_{(aq)} + Cl^{-}_{2\,(aq)}$

D. $CaCl_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^{-}_{(aq)}$

Source: January 2004

28. Consider the following reaction:
$NO^{-}_{2\,(aq)} + H_2O_{(l)} \rightleftarrows HNO_{2(aq)} + OH^{-}_{(aq)}$

This reaction represents which of the following?

II

A. the titration of NO_2^-

B. the ionization of HNO_2

C. the hydrolysis of $NaNO_2$

D. the dissociation of $NaNO_2$

Source: April 2004

CHALLENGER QUESTION 29.0

29. Which of the following solutions has the highest pH?

A. 0.1 M HCl

II

B. 0.1 M NaF

C. 0.1 M NaHS

D. 0.1 M NH_4Cl

Source: August 2003

30. Which of the following properties is true for a solution of KNO_3?

A. It is neutral.

II

B. It is very basic.

C. It is slightly basic.

D. It is slightly acidic.

Source: January 2004

31. Which of the following salts will be basic?

A. KCl

B. NH_4Cl

II

C. $KHSO_4$

D. K_2HPO_4

Source: April 2004

32. Which term does the following statement best describe? *A mixture of a weak acid and its conjugate base, each with distinguishing colours.*

A. buffer

B. titration

II

C. indicator

D. primary standard

Source: January 2004

33. The indicator phenolphthalein can be described by the following equilibrium equation:

$$\underset{\text{colourless}}{HIn} + H_2O \rightleftarrows H_3O^+ + \underset{\text{pink}}{In^-}$$

HCl is added to a slightly pink sample of this indicator. After equilibrium has been re-established, how do the $[H_3O^+]$ and the colour of the solution compare with the original equilibrium?

II

	$[H_3O^+]$	Colour of Solution
A.	decreases	turns more pink
B.	decreases	turns colourless
C.	increases	turns more pink
D.	increases	turns colourless

Source: August 2003

34. Consider the following indicator equilibrium:

$$\underset{\text{(yellow)}}{HIn} + H_2O \rightleftarrows H_3O^+ + \underset{\text{(blue)}}{In^-}$$

What is the result of adding CH_3COOH to this indicator?

II

	Equilibrium Shift	Colour
A.	left	blue
B.	left	yellow
C.	right	blue
D.	right	yellow

Source: April 2004

35. A weak acid is titrated with a strong base using the indicator phenolphthalein to detect the end point. What is the approximate pH at the transition point?

II

A. 7.0

B. 8.0

C. 9.0

D. 10.0

Source: January 2004

36. An indicator changes colour when 4.0 M HCl is added. If the indicator has a $K_a = 1 \times 10^{-10}$, identify the indicator and the pH at its transition point.

II

	Indicator	pH
A.	phenolphthalein	4.0
B.	phenolphthalein	10.0
C.	thymolphthalein	4.0
D.	thymolphthalein	10.0

Source: April 2004

37. What is the K_a value for the indicator neutral red?

A. 1×10^{-14}

B. 4×10^{-8} II

C. 7.4

D. 14.0

Source: August 2003

CHALLENGE QUESTION

38. Which of the following is not a good use for an acid-base titration curve?

A. to determine the concentration of the base

B. to select a suitable indicator for the titration

C. to determine whether the acid is strong or weak

D. to select a suitable primary standard for the titration

II

Source: August 2003

39. What term describes the chemical that is used to detect the end point of an acid-base titration?

A. buffer

B. standard II

C. indicator

D. primary standard

Source: April 2004

40. What volume of 0.100 M H_2SO_4 is needed to titrate 25.0 mL of 0.200 M NaOH?

A. 12.5 mL

B. 25.0 mL I

C. 50.0 mL

D. 100.0mL

Source: August 2003

41. What volume of 0.500 M NaOH is required to neutralize 25.0 mL of 0.250 M HBr?

A. 5.00 mL

I **B.** 12.5 mL

C. 20.0 mL

D. 25.0 mL

Source: January 2004

CHALLENGER QUESTION

42. The strong acid, $HNO_{3(aq)}$, is titrated with the weak base, $NH_{3(aq)}$. What is the net ionic equation for this reaction?

A. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

II **B.** $H^+_{(aq)} + NH_{3(aq)} \rightarrow NH^+_{4(aq)}$

C. $HNO_{3(aq)} + NH_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

D. $H^+_{(aq)} + NO^-_{3(aq)} + NH_{3(aq)} \rightarrow NH^+_{4(aq)} + NO^-_{3(aq)}$

Source: April 2004

43. Which of the following titrations always results in pH = 7.0 at the equivalence point?

A. A weak acid is titrated with a weak base.

B. A weak acid is titrated with a strong base.

II or I

C. A strong acid is titrated with a weak base.

D. A strong acid is titrated with a strong base.

Source: January 2004

44. What $[H_3O^+]$ results when 25.0 mL of 1.0 M HCl is mixed with 15.0 mL of 0.30 M KOH?

A. 0.020 M

B. 0.50 M I

C. 0.70 M

D. 0.82 M

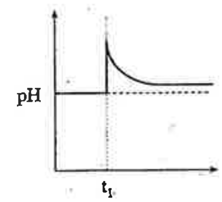
Source: April 2004

CHALLENGER QUESTION

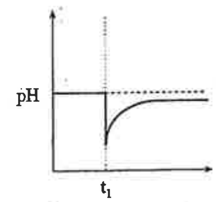
45. Which of the following graphs best describes the effect on the pH of a buffer solution when a small amount of acid is added at t_1?

II

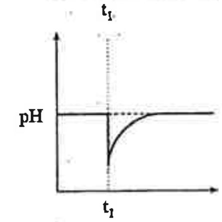
A.



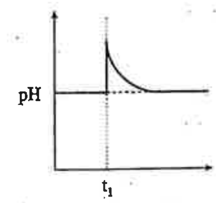
B.



C.



D.



Source: January 2004

46. What typically happens to the pH of a buffer solution when a small amount of acid is added?

A. The pH increases slightly.

B. The pH decreases slightly. II

C. The pH always remains the same.

D. The pH first increases then decreases to its original value.

Source: April 2004

47. Which of the following pairs of chemicals could be used to make a buffer solution?

A. NH_3 and H_2O

II **B.** HCl and NaCl

C. NH_3 and NH_4Cl

D. CH_3COOH and HCl

Source: August 2003

II **48.** A buffer solution is prepared using sufficient amounts of H_2S and NaHS. What limits this buffer's effectiveness when NaOH is added?

A. $[H_2S]$ **B.** $[HS^-]$

C. $[OH^-]$ **D.** $[H_3O^+]$

Source: January 2004

49. Consider the following buffer equilibrium:

$$HF_{(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + F^-_{(aq)}$$

II What would limit the buffering action if acid were added?

A. $[F^-]$ **B.** [HF]

C. $[H_2O]$ **D.** $[H_3O^+]$

Source: April 2004

50. What reaction occurs when sodium oxide dissolves in water?

A. $NaO_{(s)} \rightarrow Na^{2+}_{(aq)} + O^{2-}_{(aq)}$

II **B.** $Na_2O_{(s)} \rightarrow Na^{+}_{2(aq)} + O^{2-}_{(aq)}$

C. $NaO_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)}$

D. $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

Source: August 2003

51. What is produced when MgO is added to water?

A. the metal Mg

B. the acid HMgO II

C. the base $Mg(OH)_2$

D. the amphiprotic species H_2MgO

Source: January 2004

52. Which of the following is a major source of $NO_{2(g)}$, which contributes to the problem of acid rain?

A. a fuel cell

B. an air conditioner II

C. a nuclear power plant

D. the automobile engine

Source: January 2004

53. Identify an environmental problem associated with acid rain.

A. increasing the pH of lakes

B. depletion of the ozone layer II

C. chemical decomposition of rainwater

D. chemical erosion of limestone structures

Source: April 2004

Written Response

II 1. Write the net ionic equation for the acid-base reaction that occurs between $NaCN_{(aq)}$ and $NH_4Cl_{(aq)}$ **(2 marks)**

Source: August 2003

I 2. Define the term *amphiprotic* and give an example of an amphiprotic anion. **(2 marks)**

Source: August 2003

I 3. At 20°C, the ionization constant of water (K_w) is 6.76×10^{-15}. Calculate the $\left[H_3O^+\right]$ of water at 20°C. **(2 marks)**

Source: August 2003

II 4. Calculate the pH of 0.50 M NaF. **(5 marks)**

Source: August 2003

II 5. Outline a procedure to prepare a buffer solution. **(3 marks)**

Source: August 2003

I 6. a) Write the equation to represent the reaction that results when NH_4^+ ions are mixed with HCO_3^- ions. **(2 marks)**

b) Identify the two bases in the reaction in part a). **(1 mark)**

c) Predict whether the reaction will favour the reactants or products. Justify your answer. **(1 mark)**

Source: January 2004

7. Calculate the pH of 0.60 M NH_4I. Start by writing the equation for the predominant equilibrium reaction. **(5 marks)** II

Source: January 2004

8. A solution of $NaOH_{(aq)}$ was standardized by titration using oxalic acid $\left(H_2C_2O_{4(s)}\right)$ as the primary standard. The following data was collected:

Mass of $H_2C_2O_{4(s)}$ used $= 1.02\,g$ I

Volume of $NaOH_{(aq)}$ used $= 40.6\,mL$

Calculate the concentration of the $NaOH_{(aq)}$. **(3 marks)**

Source: January 2004

9. a) Write the formula equation to represent the complete neutralization reaction between household vinegar (acetic acid) and drain cleaner (sodium hydroxide). **(2 marks)** I

b) Write the formula for the conjugate base of the reactant acid. **(1 mark)** I

Source: April 2004

10. A sample of pure $NaOH_{(s)}$ is dissolved in water to make 10.0 L of solution and a pH = 10.75 results. Calculate the mass of pure NaOH that was dissolved. **(3 marks)** I

Source: April 2004

11. Calculate the pH of 0.70 M NH_3. Start by writing the equation for the predominant equilibrium reaction. **(5 marks)** II

Source: April 2004

UNIT TEST 4 – ACIDS, BASES, AND SALTS

1. A substance which produces hydroxide ions in solution is a definition of which of the following?

   A. an Arrhenius acid

   I **B.** an Arrhenius base

   C. a Brønsted-Lowry acid

   D. a Brønsted-Lowry base

Source: June 2003

2. Which of the following is generally true of acids, but not for bases?

   A. pH > 7

   I **B.** release H+ in solution

   C. conduct current when in solution

   D. cause indicators to change colour

Source: June 2003

3. Which of the following 1.0 M solutions will have the highest electrical conductivity?

   A. HI

   I **B.** HF

   C. HCN

   D. HNO_2

Source: June 2003

4. Consider the following equilibrium:

   $HOI + F^- \rightleftarrows OI^- + HF$

   Reactants are favoured in this equilibrium. Which of the following describes the relative strengths of the acids and the bases?

   I

	Stronger Acid	Stronger Base
A.	HF	F^-
B.	HF	OI^-
C.	HOI	F^-
D.	HOI	OI^-

Source: June 2003

5. Which of the following is true for a neutral aqueous solution?

   A. $\left[H_3O^+\right] = 0.0\,M$

   B. $\left[H_3O^+\right] = \left[OH^-\right]$ I

   C. $\left[H_3O^+\right] > \left[OH^-\right]$

   D. $\left[H_3O^+\right] < \left[OH^-\right]$

Source: June 2003

[illegible]

6. Which of the following is a definition of pK_w?

   A. $pK_w = -\log K_w$ I

   B. $pK_w = \text{pH} - \text{pOH}$

   C. $pK_w = 7.0$ at 25°C

   D. $pK_w = \left[H_3O^+\right]\left[OH^-\right]$

Source: June 2003

7. Consider the following equilibrium:

   $H_2CO_{3(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + HCO^-_{3(aq)}$

   What is the equilibrium expression?

   A. $K_a = \dfrac{\left[H_3O^+\right]\left[HCO_3^-\right]}{[H_2CO_3]}$ II

   B. $K_a = \dfrac{[H_2CO_3]}{\left[H_3O^+\right]\left[HCO_3^-\right]}$

   C. $K_a = \dfrac{[H_2CO_3][H_2O]}{\left[H_3O^+\right]\left[HCO_3^-\right]}$

   D. $K_a = \dfrac{\left[H_3O^+\right]\left[HCO_3^-\right]}{[H_2CO_3][H_2O]}$

Source: June 2003

8. Which of the following describes the net ionic equation for the hydrolysis of a $NaNO_2$ solution?

A. $NaNO_{2(s)} \rightleftarrows Na^+_{(aq)} + NO^-_{2\,(aq)}$

II

B. $NO^-_{2\,(aq)} + H_2O_{(l)} \rightleftarrows HNO_{2(aq)} + OH^-_{(aq)}$

C. $Na^+_{(aq)} + 2H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + NaOH_{(aq)}$

D. $NaNO_{2(s)} + H_2O_{(l)} \rightleftarrows NaOH_{(aq)} + HNO_{2(aq)}$

Source: June 2003

9. The $HC_2O_4{}^-_{(aq)}$ ion will act as

A. a base since $K_a < K_b$

II

B. a base since $K_a > K_b$

C. an acid since $K_a < K_b$

D. an acid since $K_a > K_b$

Source: June 2003

10. What do a chemical indicator and a buffer solution typically both contain?

A. a strong acid and its conjugate acid

II

B. a strong acid and its conjugate base

C. a weak acid and its conjugate acid

D. a weak acid and its conjugate base

Source: June 2003

11. What is the approximate pH and K_a at the transition point for phenol red?

II

A. $pH = 6.6,\ K_a = 3 \times 10^{-7}$

B. $pH = 7.3,\ K_a = 1 \times 10^{-14}$

C. $pH = 7.3,\ K_a = 5 \times 10^{-8}$

D. $pH = 8.0,\ K_a = 1 \times 10^{-8}$

Source: June 2003

12. When performing a titration experiment, the indicator must always have

A. a distinct colour change at pH = 7.0.

B. the ability to change from colourless to pink.

II

C. a transition point that is close to the equivalence point.

D. an equivalence point that is close to the stoichiometric point.

Source: June 2003

13. A 25.0 mL sample of H_2SO_4 is titrated with 30.0 mL of 0.150 M NaOH. Calculate the concentration of the H_2SO_4.

A. 0.0409 M

I

B. 0.0900 M

C. 0.125 M

D. 0.180 M

Source: June 2003

14. Consider the following buffer equilibrium:

$$H_2CO_{3(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + HCO^-_{3\,(aq)}$$

What happens when a small amount of $NaOH_{(aq)}$ is added?

II

A. $[H_3O^+]$ increases, then the equilibrium shifts to the left.

B. $[H_3O^+]$ decreases, then the equilibrium shifts to the left.

C. $[H_3O^+]$ increases, then the equilibrium shifts to the right.

D. $[H_3O^+]$ decreases, then the equilibrium shifts to the right.

Source: June 2003

CHALLENGING QUESTION

15. What is a common source of $SO_{2(g)}$?

II

A. a fuel cell
B. a car battery
C. a lead smelter
D. corrosion of iron

Source: June 2003

16. Identify the common acid found in the stomach.

I or II

A. nitric acid
B. sulphuric acid
C. perchloric acid
D. hydrochloric acid

Source: June 2004

17. Consider the following equilibrium:

$$HCO_3^- + H_2PO_4^- \rightleftarrows HPO_4^{2-} + H_2CO_3$$

What are the Brønsted-Lowry acids in this equilibrium?

I

A. HCO_3^- and H_2CO_3
B. HCO_3^- and HPO_4^{2-}
C. $H_2PO_4^-$ and H_2CO_3
D. $H_2PO_4^-$ and HPO_4^{2-}

Source: June 2004

18. Which of the following solutions would typically show the greatest electrical conductivity?

I

A. 1.0 M weak acid
B. 0.8 M weak base
C. 0.5 M strong acid
D. 0.1 M strong base

Source: June 2004

19. Which of the following are amphiprotic in aqueous solutions?

I.	H_3BO_3
II.	$H_2BO_3^-$
III.	HBO_3^{2-}
IV.	BO_3^{3-}

I

A. I only
B. IV only
C. I and II only
D. II and III only

Source: June 2004

20. What happens to the ion concentrations in water when a small amount of $HCl_{(aq)}$ is added?

A. $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M I
B. $[H_3O^+]$ and $[OH^-]$ both increase
C. $[H_3O^+]$ increases and $[OH^-]$ decreases
D. $[H_3O^+]$ increases and $[OH^-]$ is unchanged

Source: June 2004

21. Which of the following is a typical pH value for dishwashing solutions?

A. 2.0
B. 4.0 I
C. 10.0
D. 14.0

Source: June 2004

22. What is the pOH of 0.05 M $Sr(OH)_2$?

A. 1.0
B. 1.3 I
C. 12.7
D. 13.0

Source: June 2004

23. Consider the following equilibrium:

$$CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftarrows CH_3NH^+_{3(aq)} + OH^-_{(aq)}$$

Which of the following is true?

A. $K_{eq} = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2][H_2O]}$

II **B.** $K_a = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

C. $K_b = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

D. $K_{sp} = [CH_3NH_3^+][OH^-]$

Source: June 2004

24. What is the K_b value for $H_2PO_4^-$?

A. 1.3×10^{-12}

II **B.** 6.2×10^{-8}

C. 1.6×10^{-7}

D. 7.5×10^{-3}

Source: June 2004

25. Which of the following is the net ionic equation that describes the hydrolysis that occurs in a K_2CO_3 solution?

A. $2K^+_{(aq)} + CO^{2-}_{3(aq)} \rightleftarrows K_2CO_{3(s)}$

II **B.** $K_2CO_{3(aq)} \rightleftarrows 2K^+_{(aq)} + CO^{2-}_{3(aq)}$

C. $CO^{2-}_{3(aq)} + H_2O_{(l)} \rightleftarrows HCO^-_{3(aq)} + OH^-_{(aq)}$

D. $K_2CO_{3(aq)} + 2H_2O_{(l)} \rightleftarrows H_2CO_{3(aq)} + 2K^+_{(aq)} + 2OH^-_{(aq)}$

Source: June 2004

26. Which of the following amphiprotic ions will act predominantly as a base in solution?

A. HSO_3^-

B. HSO_4^- II

C. HPO_4^{2-}

D. $H_2PO_4^-$

Source: June 2004

27. Consider the indicator equilibrium:

$$\underset{\text{(yellow)}}{HIn} + H_2O \rightleftarrows H_3O^+ + \underset{\text{(red)}}{In^-}$$

Which of the following is true about the transition point of this indicator?

A. pH = 7.0

B. [HIn] = [In$^-$] II

C. [HIn] > [In$^-$]

D. moles of H_3O^+ = moles of In$^-$

Source: June 2004

28. What is one of the K_a values for thymol blue?

A. 2×10^{-9}

B. 2×10^{-7} II

C. 1×10^{-7}

D. 6×10^{-2}

Source: June 2004

29. A 25.0 mL sample of a diprotic weak acid is titrated with 20.2 mL of a strong, monoprotic base with conc. = 0.10 What is the concentration of the acid?

A. 0.040 M I

B. 0.080 M

C. 0.16 M

D. 0.12 M

Source: June 2004

30. Which of the following is the complete ionic equation for the titration of $HCl_{(aq)}$ with $KOH_{(aq)}$?

I or II

A. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

B. $HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

C. $H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

D. $H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$

Source: June 2004

31. What is always true about the pH at the equivalence point when a weak acid is titrated with a strong base?

II

A. $pH < 6.8$

B. $pH > 7.0$

C. $pH = 7.0$

D. $pH = 8.8$

Source: June 2004

32. What happens to the pH of a buffer solution if a small amount of base is added?

II

A. The pH remains constant.

B. The pH increases slightly.

C. The pH decreases slightly.

D. The pH decreases significantly.

Source: June 2004

33. What would be a reasonable $[H_3O^+]$ value for a sample of rainwater to be classified as *acid rain*?

II

A. 1.58×10^{-8} M

B. 3.16×10^{-7} M

C. 6.31×10^{-5} M

D. 1.00×10^{-1} M

Source: June 2004

34. Consider the following redox equation:

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Which of the following statements is false?

A. Iron is oxidized.

B. Hydrogen is reduced.

C. Manganese is reduced.

D. The equation is balanced.

Source: June 2004

35. Consider the following unbalanced redox equation:

$Pt + NO_3^- + Cl^- \rightarrow PtCl_6^{2-} + NO + H_2O$

Which chemical species is oxidized?

A. Pt

B. Cl^-

C. H_2O

D. NO_3^-

Source: June 2004

Written Response

1. An acid-base reaction occurs between HSO_3^- and IO_3^-.

I

a) Write the equation for the equilibrium that results. **(1 mark)**

b) Identify one conjugate acid-base pair in the reaction. **(1 mark)**

c) State whether reactants or products are favoured, and explain how you arrived at your answer. **(2 marks)**

Source: April 2003

2. At 10°C, $K_w = 2.95 \times 10^{-15}$.

II

a) Determine the pH of water at 10°C. **(3 marks)**

b) State whether water at this temperature is acidic, basic or neutral, and explain. **(1 mark)**

Source: April 2003

II **3.** Calculate the pH of 0.50 M H_2S. **(4 marks)**

Source: April 2003

4. I a) Write an equation to represent the predominant reaction when $HC_2O_4^-$ is mixed with HSO_4^-. **(1 mark)**

II b) Justify your statement by comparing K_a values. **(1 mark)**

I c) Identify a conjugate acid-base pair. **(1 mark)**

I d) Predict whether the equilibrium will favour the formation of reactants or products. Explain. **(2 marks)**

Source: June 2003

5. Write an equation representing the ionization of water and state both ion concentrations that exist for pure water to have a pH = 7.20. **(3 marks)** I

Source: June 2003

6. Calculate the pH of 0.25 M $NaHCO_3$, a basic salt. **(5 marks)** II

Source: June 2003

7. Explain why the action of a buffer solution is limited. **(2 marks)** II

Source: June 2003

8. Using calculations, show why the electrical conductivity of 1.0 M H_2CO_3 will be less than that for 0.10 M HCl. **(4 marks)** II

Source: June 2004

9. Water, at 60°C, has a $K_w = 9.55 \times 10^{-14}$.

a) Write an equation representing the ionization of water. Include the heat of reaction (57.1 kJ) in the equation. **(2 marks)** I

b) If a small amount of NaOH is added to water, what happens to the value of K_w? **(1 mark)**

Source: June 2004

10. Calculate the pH of 3.0 M Na_2CO_3. Start by writing the equation for the predominant equilibrium reaction. **(5 marks)** II

Source: June 2004

ANSWERS AND SOLUTIONS
UNIT REVIEW – ACIDS, BASES, AND SALTS

1. A	**12. D**	**23. A**	**34. B**	**45. B**
2. D	**13. A**	**24. B**	**35. C**	**46. B**
3. B	**14. B**	**25. B**	**36. D**	**47. C**
4. B	**15. D**	**26 B**	**37. B**	**48. A**
5. A	**16. C**	**27. D**	**38. D**	**49. A**
6. A	**17. D**	**28. C**	**39. C**	**50. D**
7. D	**18. B**	**29. C**	**40. B**	**51. C**
8. C	**19. C**	**30. A**	**41. B**	**52. D**
9. A	**20. D**	**31. D**	**42. B**	**53. D**
10. B	**21. A**	**32. C**	**43. D**	**WR1–11. See Solution**
11. D	**22. D**	**33. D**	**44. B**	

1. A

All acids have a pH of less than 7.0. No acids will release hydrogen gas on reaction with copper metal since $H^+_{(aq)}$ is below $Cu_{(s)}$ on the Standard Reduction Potential table of the *Data Booklet* (page A8).

2. D

According to your Acid-*Base Indicators* chart on page 7 of your *Data Booklet*, phenolphthalein will be colourless at pH less than 8.2, and pink at pH greater than 10.0.

3. B

$$Sr(OH)_{2(aq)} + H_2SO_{4(aq)} \rightarrow 2H_2O_{(l)} + SrSO_{4(s)}$$

As the equation shows the acid is being neutralized, and $SrSO_4$ precipitates out.

4. B

$H_2S_{(aq)}$ is a weak acid and will be only slightly ionized. In a *net*-ionic equation it will be written in the un-ionized form.

formula equation:
$$2NaOH_{(aq)} + H_2S_{(aq)} \rightleftharpoons Na_2S_{(aq)} + 2H_2O_{(l)}$$

complete ionic *equation:*
$$2Na^+_{(aq)} + 2OH^-_{(aq)} + H_2S_{(aq)} \rightleftharpoons 2Na^+_{(aq)} + S^{2-}_{(aq)} + 2H_2O_{(l)}$$

net-ionic equation:
$$2OH^-_{(aq)} + H_2S_{(aq)} \rightleftharpoons S^{2-}_{(aq)} + 2H_2O_{(l)}$$

5. A

They produce H^+ in solution.

6. A

Referring back to the definition of Brønsted-Lowry bases, $CH_3NH_{2(aq)}$ must react to accept a proton. This occurs in the first reaction.

7. D

A conjugate acid base pair will always differ by one proton (H^+).

8. C

Members of a conjugate acid/base pair will always be on opposite sides of a Brønsted-Lowry acid-base reaction equation. They will differ in composition by one proton (H^+).

9. A

The solution with the greatest ion concentration will have the greatest conductivity. Since $HCl_{(aq)}$ is a strong acid, its dissociation (ionization) will be virtually 100% and its ion concentration will be many times higher than that of weaker acids of equal or slightly greater concentrations.

10. B

By definition, weak acids are only partially ionized into $H^+_{(aq)}$ and another ion (their conjugate base). Solubility is not a criterion; though strong acids all have high solubilities; weak acids vary in solubility.

11. D

On the Relative Strengths of Brønsted-Lowry Acids and Bases table of the *Data Booklet* (page A6), the strongest base is present in the bottom right, the weakest base in the top right. The weakest base of those listed is $IO_3^-{}_{(aq)}$.

12. D

$HC_2O_4^-$ has a K_a value of 6.4×10^{-5}, while HPO_4^{2-} has a K_a of 2.2×10^{-13}. The larger the K_a the stronger the acid and the lower the pH.

13. A

The weaker the base the smaller the K_b. The weakest base will be the one highest up on the base side of your acid-base chart on page A6 of your *Data Booklet*.

14. B

NaCl is a neutral ionic compound while the others are weak bases. Even though Cl^- is present on the base side of the table on page A6 of the *Data Booklet*, it is a weaker base than water and is not really a base. Note that the formula in choice D is incorrectly balanced. It should be either NaH_2PO_4 or Na_2HPO_4.

15. D

Water is a very weak Brønsted-Lowry acid (and base). It will have the greatest tendency to act as an acid with the strongest base listed which is CH_3COO^-.

16. C

$K_w = \left[H_3O^+\right]\left[OH^-\right]$. This is true at 30°C or any other temperature where water is liquid. The *value* of K_w does change, however.

17. D

An acidic solution will have $[H_3O^+] > [OH^-]$, pH < 7.0, and pOH > 7.0.

18. B

In 0.025 M HCl, $[H_3O^+] = 0.025$ M since it is a strong acid. On the basis of this,

$$\left[OH-\right] = \frac{1.0\times10^{-14}}{0.025\text{ M}} = 4.0\times10^{-13}\text{ M}$$

19. C

$pK_w = pH + pOH$
$pOH = pK_w - pH$

20. D

Since HNO_3 is a strong acid,

$$\left[HNO_3\right]_{initial} = \left[H_3O^+\right]_{equilibrium} = 0.2\text{M}$$

$$pH = -\log\left[H_3O^+_{(aq)}\right] = -\log 0.2\text{M} = 0.7$$

$$pOH = 14.00 - pH = 14.00 - 0.7 = 13.3$$

21. A

The Brønsted-Lowry theory defines acids as proton donors, and bases as proton acceptors. So, the general characteristics of all Brønsted-Lowry bases are that they all accept H+.

For example, in the reaction:

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

HCl is a Brønsted acid and H_2O is a Brønsted base

22. D

Since HCO_3^- is amphiprotic, it is necessary first to determine whether it will have acidic or basic properties in aqueous solution. K_a of HCO_3^- is 5.6×10^{-11} read directly from the chart on page A6 of the *Data Booklet.*

$$K_b = \frac{1.0\times10^{-14}}{K_a \text{ for } H_2CO_3} = \frac{1.0\times10^{-14}}{4.3\times10^{-7}} = 2.3\times10^{-8}$$

Since K_b is greater than K_a, HCO_3^- will act as a base leading to the equation:
$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$, and

$$K_b = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]}$$

23. A

The strongest conjugate base would be in a pair with the weakest acid. The weakest acid will have the smallest K_a.

24. B

Acids with large K_a values are at the top of the chart on page A6 of the *Data Booklet.* The strong acids are the ones at the top of the chart.

25. B

$$K_b\ HC_6H_5O_7^{2-} = \frac{K_w}{K_a\ H_2C_6H_5O_7^-} = \frac{1.00\times10^{-14}}{1.7\times10^{-5}} = 5.9\times10^{-10}$$

26. B

In a dissociation equation in water, the electrolyte, in this case Na_2CO_3, dissociates into its aqueous ions, in this $Na^+_{(aq)}$ and $CO_3^{2-}{}_{(aq)}$.

27. D

Only C and D start out with the correct formula of calcium chloride, and only D does the dissociation correctly.

28. C

Hydrolysis is defined as the reaction of a chemical entity (ion or molecule) with water to produce H_3O^+ or OH^-.

29. C

The only base of among the solutions listed is NaHS, ($HS^-_{(aq)}$). Bases have higher pH than acids. The other solutions are all acids, though NH_4Cl, ($NH_4^+{}_{(aq)}$) is a very weak acid.

30. A

KNO_3 is a neutral electrolyte. NO_3^- *could* accept a proton, but has less tendency to do so than even water.

31. D

HPO_4^{2-} is amphiprotic but forms a basic solution in water. HSO_4^- is also amphiprotic but forms an acidic solution in water. This can be determined by comparing the relative sizes of K_a and K_b as shown in the solution to question 22.

32. C

A mixture of a weak acid and its conjugate base could describe a buffer or an indicator, but the phrase *each with* distinguishing *colours* makes indicator the only correct answer.

33. D

If HCl ($H_3O^+_{(aq)}$ and $Cl^-_{(aq)}$) is added to the equilibrium system shown, [$H_3O^+_{(aq)}$] will be increased (though it will decrease *slightly* as the equilibrium shifts left). As the equilibrium shifts left [HIn] will increase, and [In^-] will decrease causing the solution to become colourless.

34. B

If CH_3COOH is added to the equilibrium mixture, it will cause an increase in [H_3O^+] which will shift the equilibrium left according to Le Châtelier's Principle and cause the indicator to become more yellow.

35. C

A titration of weak acid with a strong base will always have an endpoint of greater than 7. This is because a new base will be produced along with water. For example:

$$HNO_{2(aq)} + OH^-_{(aq)} \longrightarrow \underset{\text{new base}}{\underset{\uparrow}{NO^-_{2\,(aq)}}} + H_2O_{(l)}$$

36. D

$pK_a = -\log\left(1\times10^{-10}\right) = 10.0 =$ pH of the transition point. According to the chart on page A7 of the *Data Booklet*, thymolphthalein changes over a pH range of 9.4 to 10.6. The middle of this range, 10.0, is its transition point.

37. B

Neutral red changes over the pH range 6.8 – 8.0. The transition point is in the middle of this range, pH = 7.4. Therefore pK_a for the indicator is 7.4 and $K_a = 10^{-pK_a} = 10^{-7.4} = 4\times10^{-8}$.

38. D

A primary standard is selected on the basis of its stability in aqueous solution. A titration curve does not give any information about this.

39. C

Acid-base indicators on page A7 of the Data Booklet detect the endpoint of an acid-base titration.

40. B

$$2\,NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$$

n_1	n_2
25.0 mL	$v = ?$
0.200 M	0.100 M

$$n_1 = 0.200\,M \times 0.0250\,L = 5.00\times10^{-3}\,mol$$

$$n_2 = \frac{1}{2}\times 5.00\times10^{-3}\,mol = 2.50\times10^{-3}\,mol$$

$$v = \frac{2.50\times10^{-3}\,mol}{0.100\,M} = 0.0250\,L = 25.0\,mL$$

41. B

Since HBr is a strong acid, and NaOH is a strong base the reaction equation will be:
$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)}$

n_1	n_2
0.250 M	0.500 M
25.0 mL	$v = ?$

$$n_1 = 0.250\,M \times 0.0250\,L = 6.25\times10^{-3}\,mol$$

$$n_2 = 6.25\times10^{-3}\,mol \times \frac{1}{1} = 6.25\times10^{-3}\,mol$$

$$v_{NaOH} = \frac{6.25\times10^{-3}\,mol}{0.500\,M} = 0.0125\,L = 12.5\,mL$$

42. B

The strong acid $HNO_{3(aq)}$ ionizes completely to $H^+_{(aq)}$ and $NO_3^-{}_{(aq)}$. The wcak base, $NH_{3(aq)}$ undergoes hydrolysis to a very small degree. Its predominant form is $NH_{3(aq)}$. The correct equation for the reaction of the strongest acid with the strongest base present is

$H^+_{(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)}$

43. D

A strong acid titrated with a strong base will produce the following net ionic equation: $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)}$.

pH at the endpoint will therefore be 7.0

44. B

$$\left[H_3O^+\right]_{initial} = 1.0\,M \times 0.0250\,L = 0.025\,mol$$

$$\left[OH^-\right]_{initial} = 0.30\,M \times 0.0150\,L = 0.0045\,mol$$

$$\text{excess } H_3O^+ = 0.025\,mol - 0.0045\,mol = 0.021\,mol$$

$$\left[H_3O^+\right]_{final} = \frac{0.021\,mol}{0.0400\,L} = 0.51\,M$$

45. B

When a small amount of an acid is added to a buffer solution, the pH will end up only slightly lower than its original value. It will take a moment for it to establish this new equilibrium as shown in graph B.

46. B

A buffer is composed of a mixture of weak acid, HA, and its conjugate base A^-. If acid is added to the buffer, the added H_3O^+ reacts with the A^-, converting it completely, or at least partially, to HA and H_2O.

$H_3O^+ + A^- \rightleftharpoons$ or $\rightarrow HA + H_2O$.

The only thing that changes is the ratio of [HA] to [A^-]. [HA] will be slightly more, [A^-] slightly less. K_a calculations demonstrate that the pH will drop only slightly with this change.

47. C

Buffer solutions are composed of a weak acid and its conjugate base. An acid and its conjugate base differ from each other in that the acid has one more proton (H^+) than its conjugate base. NH_4Cl (NH_4^+) differs from NH_3 by one proton.

48. A

When a base like NaOH is added to a buffer, the acid component of the buffer reacts with the base. The acid component of the buffer described is H_2S. Its concentration will determine the effectiveness of the buffer in reacting with the added base.

49. A

If acid is added it will react with the F^-. The system will continue to buffer against acid addition until all of the F^- is used up. Once it is used up, it will no longer be able to buffer against acid addition.

50. D

$O^{2-}{}_{(aq)}$ from the Na_2O is an extremely strong base. It will accept a proton from water to become $OH^-_{(aq)}$ which even after accepting a proton is still a strong base.

51. C

$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(s)}$.

52. D

Automobile exhaust emissions are the main source of $NO_{2(g)}$, sometimes seen as a brown haze in large cities.

53. D

Limestone (major component $CaCO_3$) is a base. It will react with the acid to produce soluble HCO_3^- compounds which will wash off the structure, causing it to be eroded.

Written Response

1. *Write the net ionic equation for the acid-base reaction that occurs between $NaCN_{(aq)}$ and $NH_4Cl_{(aq)}$.* **(2 marks)**

$$CN^-_{(aq)} + NH_4{}^+_{(aq)} \rightleftarrows HCN_{(aq)} + NH_{3(aq)}$$

2. *Define the term amphiprotic and give an example of an amphiprotic anion.* **(2 marks)**

Definition: Amphiprotic describes a substance that can act as either an acid or a base.

Example: $HCO_3{}^-$

3. *At 20°C, the ionization constant of water (K_w) is 6.76×10^{-15}. Calculate the $[H_3O^+]$ of water at 20°C.* **(2 marks)**

$$K_w = [H_3O^+][OH^-] = 6.76\times10^{-15}$$

$$\text{Since } [H_3O^+] = [OH^-], K_w = [H_3O^+]^2 = 6.76\times10^{-15}$$

$$[H_3O^+] = 8.22\times10^{-8}\text{ M}$$

4. *Calculate the pH of 0.50 M NaF.* **(5 marks)**

	F^-	+ H_2O ⇄	HF	+ OH^-
[I]	0.50		0	0
[C]	$-x$		$+x$	$+x$
[E]	$0.50-x$		x	x

(assume x is negligible)

$$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{3.5\times10^{-4}} = 2.86\times10^{-11} = \frac{[HF][OH^-]}{[F^-]} \quad \leftarrow \textbf{1 mark}$$

$$2.86\times10^{-11} = \frac{x^2}{0.50} \quad \leftarrow \textbf{1 mark}$$

$$x = [OH^-] = 3.78\times10^{-6}\text{ M} \quad \leftarrow \textbf{1 mark}$$

$$\left.\begin{array}{l} pOH = 5.42 \\ pH = 8.58 \end{array}\right\} \leftarrow \textbf{1 mark}$$

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

5. *Outline a procedure to prepare a buffer solution.* **(3 marks)**

Prepare an aqueous mixture that contains:

1. a weak acid
2. a salt of its conjugate base
3. the acid and salt in sufficient concentrations

6. **a)** *Write the equation to represent the reaction that results when $NH_4{}^+$ ions are mixed with $HCO_3{}^-$ ions.* **(2 marks)**

$$NH_4{}^+_{(aq)} + HCO_3{}^-_{(aq)} \rightleftarrows NH_{3(aq)} + H_2CO_{3(aq)}$$

b) *Identify the two bases in the reaction in part a).* **(1 mark)**

$HCO_3{}^-$ and NH_3

c) *Predict whether the reaction will favour the reactants or products. Justify your answer.* **(1 mark)**

Prediction: Reactants $\frac{1}{2}$ **mark**

Justification: $K_{a_{H_2CO_3}} > K_{a_{NH_4^+}}$ $\frac{1}{2}$ mark

7. *Calculate the pH of 0.60 M NH_4I. Start by writing the equation for the predominant equilibrium reaction.* **(5 marks)**

	$NH_4^+{}_{(aq)}$ + $H_2O_{(l)}$	$\rightleftarrows$ $H_3O^+{}_{(aq)}$	+ $NH_{3(aq)}$
[I]	0.60	0	0
[C]	$-x$	$+x$	$+x$
[E]	$0.60-x$	x	x

(assume x is negligible)

$$K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]}$$

$$5.6\times10^{-10} = \frac{(x)(x)}{(0.60)}$$

$$x = [H_3O^+] = 1.83\times10^{-5}\,M$$

$$pH = 4.74$$

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

8. *A solution of $NaOH_{(aq)}$ was standardized by titration using oxalic acid $\left(H_2C_2O_{4(s)}\right)$ as the primary standard. The following data was collected:*
Mass of $H_2C_2O_{4(s)}$ used = 1.02 g
Volume of $NaOH_{(aq)}$ used = 40.6 mL
Calculate the concentration of the $NaOH_{(aq)}$. **(3 marks)**

$$H_2C_2O_{4(s)} + 2NaOH_{(aq)} \rightarrow Na_2C_2O_{4(aq)} + 2H_2O_{(l)}$$

$$\text{Moles of } H_2C_2O_{4(s)} = 1.02\,g \times \frac{mol}{90.0\,g}$$

$$= 1.133\times10^{-2}\,mol$$

$$\text{Moles of NaOH} = 2\left(1.13\times10^{-2}\,mol\right)$$

$$= 2.267\times10^{-2}\,mol$$

$$[NaOH] = \frac{2.26\times10^{-2}\,mol}{0.0406\,L} = 0.558\,M$$

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

9. a) *Write the formula equation to represent the complete neutralization reaction between household vinegar (acetic acid) and drain cleaner (sodium hydroxide).* **(2 marks)**

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$$

b) *Write the formula for the conjugate base of the reactant acid.* **(1 mark)**

CH_3COO^-

10. *A sample of pure $NaOH_{(s)}$ is dissolved in water to make 10.0 L of solution and a pH = 10.75 results. Calculate the mass of pure NaOH that was dissolved.* **(3 marks)**

$$pH = 10.75$$
$$pOH = 14.00 - 10.75 = 3.25$$
$$\left[OH^-\right] = 5.6 \times 10^{-4}\ M$$
$$\text{mass NaOH} = 5.6 \times 10^{-4}\ \frac{mol}{L} \times 10.0\ L \times \frac{40.0\ g}{mol} = 0.22\ g$$

↑ $\frac{1}{2}$ **mark** ↑ **1 mark**

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

11. *Calculate the pH of 0.70 M NH_3. Start by writing the equation for the predominant equilibrium reaction.* **(5 marks)**

	$NH_{3(aq)}$ + $H_2O_{(l)}$ $\rightleftarrows$	$NH_4{}^+{}_{(aq)}$ +	$OH^-{}_{(aq)}$
[I]	0.70	0	0
[C]	$-x$	$+x$	$+x$
[E]	$0.70-x$	x	x

(assume x is negligible)

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$
$$= \frac{\left[NH_4{}^+\right]\left[OH^-\right]}{\left[NH_3\right]}$$
$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.70}$$
$$x = \left[OH^-\right] = 3.5 \times 10^{-3}\ M$$
$$pOH = 2.45$$
$$pH = 11.55$$

ANSWERS AND SOLUTIONS
UNIT TEST 4 – ACIDS, BASES, AND SALTS

1. B	7. A	13. B	19. D	25. C	31. B
2. B	8. B	14. D	20. C	26. C	32. B
3 A	9. D	15. C	21. C	27. B	33. C
4. B	10. D	16. D	22. A	28. A	34. B
5. B	11. C	17. C	23 C	29. A	35. A
6. A	12. C	18. C	24. A	30. D	WR1–10. See Solution

1. B

An Arrhenius base is a substance which dissociates in water to produce hydroxide ions.

2. B

According to the Arrhenius theory, an acid is a substance which ionizes in water to produce hydrogen ions. This is not true for bases.

3. A

The greater the concentration of ions in solution, the greater the electrical conductivity. Although these solutions have equal concentrations, only HI is a strong acid. It will have high ion concentrations. The other acid *ion* concentrations will be low.

4. B

Since reactants are favoured, the proton transfer from HF to OI^- is more complete than that of HOI to F^-.

5. B

In a neutral solution at room temperature, $[H_3O^+] = [OH^-] = 1.00 \times 10^{-7}$ M.

6. A

By definition.

7. A

For the reaction, $a\,A + b\,B \rightarrow c\,C + d\,D$,

$$K_{eq} = \frac{[C]^c \bullet [D]^d}{[A]^a \bullet [B]^b}.$$

K_a is just a special case of K_{eq}. $H_2O_{(l)}$ is not included in equilibrium constants of aqueous solutions. Thus for an acid

$$K_a = \frac{[H_3O^+] \cdot [\text{conjugate base}]}{[\text{acid}]}$$

8. B

NO_2^- from $NaNO_2$ is a weak base. It reacts with water or hydrolyzes according to equations B or D. The question asks for a *net ionic equation*, so B is the correct answer.

9. D

$HC_2O_4^-{}_{(aq)}$ is amphiprotic. It may act as a Brønsted-Lowry acid or Brønsted-Lowry base, depending on what it is reacting with. A water solution will display acidic properties since its $K_a > K_b$. Using the chart on page A6 of the *Data Booklet*,

$$K_a = 6.4\times10^{-5},\ K_b = \frac{1.00\times10^{-14}}{5.9\times10^{-2}} = 1.7\times10^{-13}.$$

10. D

Buffers and indicators will contain a mixture of a weak acid with its conjugate base. Both indicators and buffers will react with an added acid or base but for different purposes.

11. C

According to the chart on page A7 of the *Data Booklet*, phenol red changes colour from pH 6.6 – 8.0. The pH in the middle of the transition range for the indicator is 7.3. pKa for the indicator is therefore 7.3, and $K_a = 10^{-pK_a} = 10^{-7.3} = 5 \times 10^{-8}$

12. C

The transition point, or pH at which an indicator changes, should be as close as possible to the pH at the equivalence point, where equivalent numbers of moles of sample and titrant are present.

13. B

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

n_2	n_1
25.0 mL	30.0 mL
$c = ?$	0.150 M

$$n_1 = 0.150\ \text{M} \times 30.0\ \text{mL} = 4.50\ \text{mmol}$$

$$n_2 = 4.50\ \text{mmol} \times \frac{1}{2} = 2.25\ \text{mmol}$$

$$[H_2SO_4] = \frac{2.25\ \text{mmol}}{25.0\ \text{mL}} = 0.0900\ \text{M}$$

14. D

$OH^-_{(aq)}$ from the $NaOH_{(aq)}$ will react with the $H_3O^+_{(aq)}$ converting it to water. As $[H_3O^+_{(aq)}]$ decreases, Le Châtelier's Principle predicts an equilibrium shift to the right.

15. C

Non-ferrous (not iron) ore smelters are major contributors to SO_2 pollution and the resulting acid rain.

16. D

Hydrochloric acid, the aqueous solution of hydrogen chloride, is the main part of gastric acid, which is the one of the main secretions of the stomach. Smaller quantities of potassium chloride and sodium chloride can also be found in the stomach.

17. C

According to the Brønsted-Lowry definition, acids are proton donors. As you read the equation in the forward direction, $H_2PO_4^-$ donates a proton. In the reverse direction, H_2CO_3 donates a proton.

18. C

The solution with the largest ion concentration will have the greatest electrical conductivity. It must be either a *strong* acid or a *strong* base, since weak acids and weak bases are only partially dissociated. Since the *strong* acid concentration in choice C is higher than the *strong* base concentration in choice D, it will have the larger ion concentration and the greater electrical conductivity.

19. D

$H_2BO_3^-$ and HBO_3^{2-} can either or accept or donate a proton depending on what they react with. This is the definition of *amphiprotic*.

20. C

For all aqueous solutions $\left[H_3O^+_{(aq)}\right] \times \left[OH^-_{(aq)}\right]$ = a constant $(1.00 \times 10^{-14}$ at SATP). If $HCl_{(aq)}$ is added to water, $\left[H_3O^+_{(aq)}\right]$ increases and therefore $\left[OH_{(aq)}\right]$ must decrease.

21. C

All soap solutions are basic and will have a somewhat high pH. pH 14.0, choice D, is excessively high.

22. A

$$Sr(OH)_{2(aq)} \rightarrow Sr^{2+}{}_{(aq)} + 2\,OH^{-}{}_{(aq)}$$

0.05 M 2 × 0.05 M 0.1 M

pOH = –log [OH⁻] = –log 0.1 M = 1.0

23. C

$CH_3NH_{2(aq)}$ is acting as a base (accepting a proton) in this reaction equation. This is why the K is a K_b. Following the standard form for equilibrium constants,

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

24. A

$$K_b = \frac{1.00\times10^{-14}}{K_a \text{ of the conjugate acid}} = \frac{1.00\times10^{-14}}{K_a \text{ for } H_3PO_4}$$

$$= \frac{1.00\times10^{-14}}{7.5\times10^{-3}} = 1.3\times10^{-12}$$

25. C

Hydrolysis of a base is a reaction with water to produce $OH^{-}{}_{(aq)}$. C and D both look correct, but only C is a **net-ionic** equation.

26. C

The strongest base from the list will be the answer. This is HPO_4^{2-}.

27. B

When an indicator is at its transition point there will be equal concentrations of HIn and In^-, producing an intermediate colour.

28. A

Thymol blue has 2 transition points, one at about pH 2.0, the other at about pH 8.8, leading to pK_a's of 2.0 and 8.8.

$K_a = 10^{-pK_a} = 10^{-2.0} = 1\times10^{-2}$ or

$K_a = 10^{-pK_a} = 10^{-8.8} = 2\times10^{-9}$

29. A

$$H_2A_{(aq)} + 2NaOH_{(aq)} \rightarrow$$
$$Na_2A_{(aq)} + 2\,H_2O_{(l)}$$

n_2	n_1
25.0 mL	20.2 mL
c = ?	0.10 M

$n_1 = 0.10 \text{ M} \times 20.2 \text{ mL} = 2.0 \text{ mmol}$

$n_2 = 2.0 \text{ mol} \times \frac{1}{2} = 1.0 \text{ mmol}$

$? = c = \frac{1.0 \text{ mmol}}{25.0 \text{ mL}} = 0.040 \text{ M}$

30. D

The question asks for a *complete ionic equation*. Choice A is the correct *net-ionic equation*, choice B is the correct *formula equation*, but choice D is the correct *complete ionic equation*.

31. B

When a weak acid is titrated with a strong base, a weak base plus water will be formed. Therefore pH will be greater than 7.

Example:

$HF_{(aq)}$ *(weak acid)* + $OH^{-}{}_{(aq)}$ *(strong base)*

$\rightarrow H_2O_{(l)} + F^{-}{}_{(aq)}$ *(weak base)*

32. B

A buffer will resist a change in pH despite addition of acid or base. It cannot do this perfectly; pH will change slightly. If too much base is added the acid part of the buffer will be used up and the buffer will no longer be effective.

33. C

"Normal" rain has a pH of approximately 5.6. This is $\left[H_3O^+{}_{(aq)}\right]=10^{-5.6}=3\times10^{-6}\ M$.

Any $\left[H_3O^+{}_{(aq)}\right]$ greater than this would be classified as acid rain. Choices C and D are both greater than $3\times10^{6}M$ but choice D is *unreasonably* high.

34. B

H^+ is neither oxidized nor reduced. Its oxidation number is +1 on both sides of the equation.

35. A

Pt has an oxidation number of 0 in Pt. Its oxidation number in $PtCl_6^{2-}$ is +4. An increase in oxidation number means the element has lost electrons (is oxidized).

Written Response

1. *An acid-base reaction occurs between HSO_3^- and IO_3^-.*

a) *Write the equation for the equilibrium that results.* **(1 mark)**

Solution:

$HSO_3^- + IO_3^- \rightleftarrows SO_3^{2-} + HIO_3$

(equilibrium arrow required) **(1 mark)**

b) *Identify one conjugate acid-base pair in the reaction.* **(1 mark)**

Solution:

HSO_3^- and SO_3^{2-} **OR**

IO_3^- and HIO_3

} ← **1 mark**

c) *State whether reactants or products are favoured, and explain how you arrived at your answer.* **(2 marks)**

Solution:

Reactants are favoured. ← **1 mark**

HSO_3^- is a weaker acid than HIO_3

OR

IO_3^- is a weaker base than SO_3^{2-}

} ← **1mark**

2. *At 10°C,* $K_w = 2.95 \times 10^{-15}$.

a) *Determine the pH of water at 10°C.* **(3 marks)**

Solution:

$$\left.\begin{aligned} K_w &= 2.95 \times 10^{-15} \\ &= [H_3O^+][OH^-] \end{aligned}\right\} \leftarrow \textbf{1 mark}$$

Since $[H_3O^+] = [OH^-]$,

$[H_3O^+]^2 = 2.95 \times 10^{-15}$ ← **1 mark**

$[H_3O^+] = 5.43 \times 10^{-8}$ ← **1 mark**

pH = 7.265

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

b) *State whether water at this temperature is acidic, basic or neutral, and explain.* **(1 mark)**

Solution:

Since $[H_3O^+] = [OH^-]$, the water is neutral. ← **1 mark**

3. *Calculate the pH of 0.50 M H_2S.* **(4 marks)**

Solution:

	H_2S	+ H_2O ⇄	H_3O^+	+ HS^-
[I]	0.50		0	0
[C]	$-x$		$+x$	$+x$
[E]	$0.50-x$		x	x

$\left(1\frac{1}{2}\ \textbf{marks}\right)$

(assume x is negligible)

$$\left.\begin{aligned} K_a = 9.1\times10^{-8} &= \frac{[H_3O^+][HS^-]}{[H_2S]} \\ 9.1\times10^{-8} &= \frac{(x)(x)}{(0.50)} \end{aligned}\right\} \leftarrow \textbf{1 mark}$$

$x = [H_3O^+] = 2.13\times10^{-4}$ ← **1 mark**

pH = 3.67 ← $\frac{1}{2}$ **mark**

4. a) *Write an equation to represent the predominant reaction when $HC_2O_4^-$ is mixed with HSO_4^-.* **(1 mark)**

Solution:

Example:

$HC_2O_4^- + HSO_4^- \rightleftarrows H_2C_2O_4 + SO_4^{2-}$ **(1 mark)**

b) *Justify your statement by comparing K_a values.* **(1 mark)**

Solution:

Example:

$K_a\, HSO_4^- > K_a\, HC_2O_4^-$ **(1 mark)**

c) *Identify a conjugate acid-base pair.* **(1 mark)**

Solution:

Example:

$$\left.\begin{aligned} &HSO_4^- \text{ and } SO_4^{2-} \\ &\textbf{OR} \quad HC_2O_4^- \text{ and } H_2C_2O_4 \end{aligned}\right\} \leftarrow \textbf{1 mark}$$

d) *Predict whether the equilibrium will favour the formation of reactants or products. Explain.* **(2 marks)**

Solution:

Example:

Reactants are favoured since $K_a HSO_4^- < K_a H_2C_2O_4$ ← **2 marks**

5. *Write an equation representing the ionization of water and state both ion concentrations that exist for pure water to have a pH = 7.20.* **(3 marks)**

Solution:

Example:

$H_2O_{(l)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + OH^-_{(aq)}$ ←**1 mark**

Since pH = 7.20, $[H_3O^+] = 6.3\times10^{-8}$ M ←**1 mark**

$[H_3O^+] = [OH^-] = 6.3\times10^{-8}$ M ←**1 mark**

6. *Calculate the pH of 0.25 M $NaHCO_3$, a basic salt.* **(5 marks)**

Solution:

Example:

	HCO_3^- + H_2O ⇄	H_2CO_3 +	OH^-
[I]	0.25	0	0
[C]	$-x$	$+x$	$+x$
[E]	$0.25-x$	x	x

←**1 mark**

(assume x is negligible)

$$K_b = \frac{1.0\times10^{-14}}{4.3\times10^{-7}} = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]}$$ ←**1 mark**

$$2.33\times10^{-8} = \frac{x^2}{0.25}$$ ←**1 mark**

←**1 mark**

$x = [OH^-] = 7.62\times10^{-5}$ M

pOH = 4.12

pH = 9.88

←**1mark**

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

7. *Explain why the action of a buffer solution is limited.* **(2 marks)**

Solution:

Example:

Buffer action depends on the presence of sufficient amounts ←**1 mark** of weak acid and conjugate base in the buffer solution. ←**1 mark**

8. *Using calculations, show why the electrical conductivity of 1.0 M H_2CO_3 will be less than that for 0.10 M HCl.* **(4 marks)**

For Example:

For 0.10 M HCl, a strong acid:

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

total concentration = 0.10 M + 0.10 M = 0.20 M **(1 mark)**

For 1.0 M H_2CO_3, a weak acid:

	$H_{2}CO_{3(aq)}$ + $H_2O_{(l)}$ ⇄	$HCO_3^-{}_{(aq)}$ +	$H_3O^+{}_{(aq)}$
[I]	1.0	0	0
[C]	$-x$	$+x$	$+x$
[E]	$1.0-x$	x	x

$\left(2\frac{1}{2}\ \textbf{marks}\right)$

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = 4.3\times10^{-7}$$

$$\frac{(x)(x)}{1.0-x} = 4.3\times10^{-7}$$

$x = 6.6\times10^{-4}$ M $= [H_3O^+] = [HCO_3^-]$

$\left(\frac{1}{2}\ \textbf{mark}\right)$

Total ion concentration = 6.6×10^{-4} M + 6.6×10^{-4} M = 1.3×10^{-3} M

Therefore, smaller ion concentration, lower conductivity.

9. a) *Write an equation representing the ionization of water. Include the heat of reaction (57.1 kJ) in the equation.* **(2 marks)**

For Example:

$2H_2O_{(l)} + 57.1\text{kJ} \rightleftarrows H_3O^+_{(aq)} + OH^-_{(aq)}$

OR

$H_2O_{(l)} + 57.1\text{kJ} \rightleftarrows H^+_{(aq)} + OH^-_{(aq)}$

Note: Endothermic can be deduced from the data provided.

1 mark for the equation.
1 mark for determining endothermic.

b) *If a small amount of NaOH is added to water, what happens to the value of K_w?* **(1 mark)**

For Example:

K_w remains unchanged.

10. *Calculate the pH of 3.0 M Na_2CO_3. Start by writing the equation for the predominant equilibrium reaction.* **(5 marks)**

For Example: **(1mark)**

	$CO_3^{2-}{}_{(aq)}$ + $H_2O_{(l)}$ $\rightleftarrows$	$HCO_3^-{}_{(aq)}$ +	$OH^-_{(aq)}$
[I]	3.0	0	0
[C]	$-x$	$+x$	$+x$
[E]	$3.0-x$	x	x

(1 mark)

(assume x is negligible)

$$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{5.6\times10^{-11}} = 1.79\times10^{-4}$$

$$= \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$$

$$1.79\times10^{-4} = \frac{(x)(x)}{(3.0)}$$

$$x = [OH^-] = 0.0232\text{M}$$

$$\text{pOH} = 1.64$$

$$\text{pH} = 12.36$$

(1 mark)

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)