VULCAN

Challenge Exam Key

Multiple Choice (1-50)

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

Detailed Answers (1-25)

1. (a) $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

   (b) $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

   (c) $CH_3CH_2OH + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$

   (d) $P_4 + 5\ O_2 \rightarrow 2\ P_2O_5$

   (e) $N_2O_5 + H_2O \rightarrow 2\ HNO_3$

2. (a) reaction of soluble species producing a solid precipitate. Example AgCl

   (b) reaction where there is an increase in the oxidation number of one or more elements. Example $H_2 + O_2 \rightarrow H_2O$

   (c) rapid reaction that produces a flame, usually involves O_2.
   Example $CH_4 + O_2 \rightarrow CO_2 + H_2O$

   (d) reaction involving the exchange of anions between cations.
   Example $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

   (e) reactions that show only those species that undergo a change of state.
   Example $Ag^+ + Cl^- \rightarrow AgCl$

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3. (a) $6\ \mathrm{Li} + \mathrm{N_2} \rightarrow 2\ \mathrm{Li_3N}$

(b) $\mathrm{NaI} + \mathrm{AgNO_3} \rightarrow \mathrm{AgI} + \mathrm{NaNO_3}$ net ionic equation okay

(c) $2\ \mathrm{KClO_3} \rightarrow 2\ \mathrm{KCl} + 3\ \mathrm{O_2}$

(d) $\mathrm{C_2H_4} + 3\ \mathrm{O_2} \rightarrow 2\ \mathrm{CO_2} + 2\ \mathrm{H_2O}$

(e) $\mathrm{Al(OH)_3} + 3\ \mathrm{HClO_4} \rightarrow \mathrm{Al(ClO_4)_3} + 3\ \mathrm{H_2O}$

4. Only need to answer five of eight

(a) $\mathrm{Ag^+} + \mathrm{Br^-} \rightarrow \mathrm{AgBr}$

(b) $\mathrm{CaCO_3} \rightarrow \mathrm{CaO} + \mathrm{CO_2}$

(c) $\mathrm{Cu(OH)_2} + 2\ \mathrm{H^+} \rightarrow \mathrm{Cu^{2+}} + 2\ \mathrm{H_2O}$

(d) $\mathrm{Pb^{2+}} + 2\ \mathrm{OH^-} \rightarrow \mathrm{Pb(OH)_2}$

(e) $\mathrm{Na_2O} + \mathrm{H_2O} \rightarrow 2\ \mathrm{Na^+} + \mathrm{OH^-}$

(f) $\mathrm{CH_3NH_2} + \mathrm{H_2O} \rightarrow \mathrm{CH_3NH_3^+} + \mathrm{OH^-}$

(g) $\mathrm{Cu^{2+}} + \mathrm{H_2S} \rightarrow \mathrm{CuS} + 2\ \mathrm{H^+}$

(h) $2\ \mathrm{Li} + \mathrm{Zn^{2+}} \rightarrow 2\ \mathrm{Li^+} + \mathrm{Zn}$

5. (a) precipitate is $\mathrm{Sn(OH)_2}$

$\mathrm{Sn^{2+}} + 2\ \mathrm{OH^-} \rightarrow \mathrm{Sn(OH)_2}$

(b) No reaction

(c) precipitate is CuS $\mathrm{Cu^{2+}} + \mathrm{S^{2-}} \rightarrow \mathrm{CuS}$

6. (a) 0 0 2 -1 oxidation numbers
$\mathrm{Ni} + \mathrm{Cl_2} \rightarrow \mathrm{NiCl_2}$
Ni is oxidized and $\mathrm{Cl_2}$ is reduced

(b) 2 5 -2 0 0 3 5 -2 oxidation numbers
$3\ \mathrm{Fe(NO_3)_2} + 2\ \mathrm{Al} \rightarrow 3\ \mathrm{Fe} + 2\ \mathrm{Al(NO_3)_3}$
$\mathrm{Fe^{2+}}$ is reduced and Al is oxidized

(c) 0 1 -1 0 1 -1 oxidation numbers
$\mathrm{Cl_2} + 2\ \mathrm{NaI} \rightarrow \mathrm{I_2} + 2\ \mathrm{NaCl}$
$\mathrm{Cl_2}$ is reduced and $\mathrm{I^-}$ is oxidized

(d) 2 -2 1 -1 2 6 -2 1 -2 oxidation numbers

$PbS + 4\ H_2O_2 \rightarrow PbSO_4 + 4\ H_2O$

S is oxidized and O is reduced

7. Activity series: X1 is most active species

X1
X3
X2

(b) The most reactive is X1

(c) Most readily undergoes oxidation: X1

8. (a) Acetone is a non-electrolyte so the only species present in solution is the acetone molecule.

HClO is a weak electrolyte so the prevalent species is HClO, however, a small amount of H_3O^+ and ClO^- are also present.

NH_4Cl is a strong electrolyte so the only species present are NH_4^+ and Cl^-.

(b) Acetone solution will contain 0.20 moles of acetone

HClO solution will contain 0.200 moles of HClO and 7.7×10^{-5} moles of H_3O^+ and ClO^- . Acceptable answer would be zero ions.

NH_4Cl solution would only contain 0.200 moles of NH_4^+ and Cl^-.

9. (a) $2\ H^+ + Na_2O \rightarrow 2\ Na^+ + H_2O$

(b) $Ca^{2+} + CO_3^{2-} \rightarrow CaCO_3$

(c) $2\ H^+ + Mg \rightarrow Mg^{2+} + H_2$

(d) $Cu^{2+} + S^{2-} \rightarrow CuS$

10. Acid strength is based on the proton donating ability of a species. Strong acids will donate a proton to water, the base, 100%. To differentiate we need to change the nature of the solvent. Use a more acidic solvent to differentiate the proton donating strength of these acids.

11. (a) $[H^+] = 2.13 \times 10^{-3}$ M

(b) 0.61% ionized

(c) $[C_2H_5COO^-]/[C_2H_5COOH] = 1.03$

(d) pH would increase since you are neutralizing some of the acid with the

NaOH and increasing the moles of conjugate base.

12. (a) pH = 11.92 pOH = 2.08

(b) $(CH_3)_2NH_2^+ + H_2O \rightarrow (CH_3)_2NH + H_3O^+$

$K_b = 1.85 \times 10^{-11}$

(c) pH = 11.28

(d) The pH of the buffer is very basic so you would need to add base to the solution to neutralize some or all of the acid. Thus the choice is $HClO_4$.

13. (a) pH = 6.12

(b) pH = 9.06

(c) NaOH will neutralize some of the boric acid and increase the concentration of the conjugate base, $H_2BO_3^-$ so the system becomes more basic.

(d) HCl will neutralize some of the borate ions, a base, and increase the moles of boric acid so the system becomes more acidic.

14. (a) $[H_3O^+]$ = [benzoate ion] = 1.77×10^{-3} M
[benzoic acid] = 0.05 M

(b) $C_6H_5COO^-$ or $C_7H_5O_2^-$ $K_b = 1.59 \times 10^{-10}$

(c) pH = 2.10

(d) % dissociation = 0.013

15. (a) pH = 2.32 pOH = 14.92-2.32 = 12.60

(b) pH = 3.35

(c) Maximum buffer capacity occurs when the ratio between acid and its conjugate base is 1.0.

concentration of conjugate base, NO_2^- = 0.250 moles/0.50L = 0.50 M
concentration of HNO_2 given as 0.50 M

ratio of the two is 0.50/0.50 = 1.0 so we have maximum buffer capacity

16. (a) Use the Henderson-Hasselbalch equation for buffers

$pH = pK_a + \log [A^-]/[HA]$

ratio of HCO_3^- to H_2CO_3 is 10.8

(b) When the pH of the buffer is 7.1, then the ratio of HCO_3^- to H_2CO_3 is 5.41.

17. (a) $H_2CO_3 + 2\ NaOH \rightarrow Na_2CO_3 + 2\ H_2O$

(b) $H_2CO_3 + H_2O \rightarrow HCO_3^- + H_3O^+$

$HCO_3^- + H_2O \rightarrow CO_3^{2-} + H_3O^+$

$K_{a1} = 4.3 \times 10^{-7}$

(c) pH = 3.48 $[HCO_3^-] = 3.28 \times 10^{-4}$ M $[CO_3^-] = 5.6 \times 10^{-11}$ M

(d) $K_b = 1.78 \times 10^{-4}$

18. (a) $K_{a2} = 1.6 \times 10^{-12}$

(b) pH = 3.52

19. (a) $K = 5.0 \times 10^{-15}$

(b) $[H_2C_6H_5O_7^-] = 1.43 \times 10^{-2}$ M

$[HC_6H_5O_6^-] = 1.7 \times 10^{-5}$ M

$[C_6H_5O_7^-] = 4.76 \times 10^{-10}$

20. Acids containing the same number of oxygen atoms and OH units increase acid with increasing electronegativity-see table 16.6

Thus, acid strength is HOI < HOBr < HOCl < HOF

21. (a) pH = 2.3

(b) The reason the Cu^{2+} is less acidic than Fe^{3+} is due to the decrease in charge and size of the Cu^{2+} ion. This reduces the polarity of the metal-O bond and thus the O-H bond is stronger than for the iron case. Thus, it is more difficult to break the o-H bond in the hydrated copper ion than for the hydrated iron.

22. (a) If you determine the pH of the solution from the data you will get a pH of 7.44. This indicates that the phenol solution is basic and not acidic. The problem is that the water is a significant contributor of hydrogen ions and that source has been ignored. The calculation is thus much more complex than the usual weak acid calculations.

23. (a) pH = 10.44

(b) $H_2CO_3 + CO_3^{2-} \rightarrow HCO_3^-$

Thus, adding carbonic acid neutralizes some of the carbonate ion, therefore reducing the amount of base in the solution and decreasing the pH so the solution is more acidic.

24. (a) NaCl pH = 7.00

(b) ClO^- pH = 7.48

(c) $CH_3NH_3^+$ pH = 5.82

(d) IO_3^- pH = 6.88 solution should be basic, however, the K_b is so small and the contribution of water to the $[H_3O^+]$ has been ignored so you get an incorrect answer.

(e) $CH_3CH_2NH_3^+$ pH = 5.90

25. $HClO_4$ is a strong acid, dissociating 100%. There is no equilibrium between the acid and its conjugate base so you cannot form a buffer. Carbonic acid is a weak acid in equilibrium with its conjugate base and thus forms buffers.

Free Form (1-25)

1. The subscript 3 in HNO_3 refers to the number of oxygen atoms in one molecule of HNO_3 while the coefficient 3 in front of HNO_3 in a balanced equation refers to the number of moles of nitric acid in the balanced chemical reaction.

2. Pure water is free of ions and thus a non-electrolyte. This means it is an insulator. Tap water contains ions such as Mg^{2+}, Na^+, etc. and thus is a weak electrolytic solution and will conduct slightly.

3. Most readily oxidized metals are locate in the far left portion of the periodic table while the least easily oxidized are located on the right side near the interface with the non-metals.

4. (a) $Zn + Cd^{2+} \rightarrow Zn^{2+} + Cd$

$Cd + Ni^{2+} \rightarrow Cd^{2+} + Ni$

(b) Cd is between Zn and Ni in the activity series.

(c) $Ca + Cd^{2+} \rightarrow Cd + Ca^{2+}$

Thus, Ca is above Cd in the activity series.

5. Nitric acid is a strong acid and a strong oxidizing agent while HCl is only a strong acid. The reaction of copper with HNO_3 is an oxidation-reduction reaction requiring a reasonably strong oxidizing agent.

6. The addition of silver sulfate to the barium solution precipitates the barium as $BaSO_4$ and the silver as AgCl. Thus, you have significantly reduced the ion content of the solution and the solution is a very weak electrolyte and will no conduct.

7. Test a portion of the solution with Na_2CO_3. If calcium acetate is present the calcium will precipitate as calcium carbonate.

   $Ca^{2+} + CO_3^{2-} \rightarrow CaCO_3$

   Test a portion of the solution with Na_2SO_4. If lead nitrate is present lead sulfate will precipitate.

   $Pb^{2+} + SO_4^{2-} \rightarrow PbSO_4$

   Test a portion of the solution with NaCl. If silver acetate is present silver chloride will precipitate. Lead will also precipitate but the sulfate test will show whether you have lead or silver.

   $Ag^{+} + Cl^{-} \rightarrow AgCl$

8. Both acids react with water to produce hydronium ions. Both react with bases to produce salts. Both are electrolytes: nitric is a strong electrolyte and phosphoric is a weak electrolyte.

   The common behavior is due to the presence of acid protons in both species.

9. Autoionization of water $2\ H_2O \leftrightarrows H_3O^{+} + OH^{-}$

   (a) $NH_3 + NH_3 \leftrightarrows NH_4^{+} + NH_2^{-}$

   (b) $SO_2 + SO_2 \leftrightarrows SO_3^{2-} + SO^{2+}$

   (c) $HCN + HCN \leftrightarrows H_2CN^{+} + CN^{-}$

10. The concentration of water remains essentially constant in equilibrium due to its high concentration, 55.6 moles of water per liter. A 1.0 M solution of acetic acid would contain 4.2×10^{-3} moles of acetate ion and hydronium ion per liter. Thus, you would have lost 4.2×10^{-3} moles of water out of 55.6 moles, or 7.6×10^{-3} % of the water. Therefore, $[H_2O]$ is basically unchanged and not included in the equilibrium expression since the equilibrium expression is a ratio of those species that do change significantly during equilibrium.

11. The definition of pH is $-\log[H^{+}]$. Thus, if you increase the $[H^{+}]$ by a factor of 1000 the difference in the pH of the two solutions is 3.0.

12. The contribution of water to the hydronium concentration is usually a very small fraction of the total hydronium ion produce from the equilibrium of a weak acid in solution. However, if you have a very weak acid, with a very small K_a, or a very low concentration then water may contribute a significant fraction of hydronium ions to the solution and you would need to include both equilibrium processes in the determination of pH.

13. The percent dissociation of a weak acid varies as the inverse square root of the acid concentration. Thus, as the concentration of the acid decreases the % dissociation increases. For a 1.0 M solution of acetic acid the % dissociation is 0.42% while that of a 0.01 M solution is 4.2%.

14. (a) $HCl > HF$ and $H_2S > H_2O$ Order of increasing acidity due to the differences in bond strengths. HF and HO bonds are stronger that HCl and HS bonds. The weaker the bond the easier it is to dissociate the proton, and therefore the stronger the acid.

    (b) $CH_4 < NH_3 << H_2O < HF$ Order of increasing acidity is due to the increasing electronegativity of the non-hydrogen atom. See p 645 of Brown.

15. (a) $C_7H_5O_2^-$ is a base so adding it to $HC_7H_5O_2$ will increase the pH since you are neutralizing some of the acid upon addition of base.

    (b) $C_5H_5NH^+$ is an acid so adding it to pyridine will decrease the pH since you are neutralizing some pf the base.

    (c) Adding ammonia to a solution of HCl will neutralize some of the acid and thus increase the pH.

    (d) Adding $NaHCO_3$ to carbonic acid will increase the pH since the bicarbonate ion is the conjugate base of carbonic acid.

    (e) Adding sodium perchlorate to NaOH will not change the pH since ClO_4^- is the anion of a strong acid and no reaction will occur.

16. $CH_3COOH + H_2O \leftrightarrows CH_3COO^- + H_3O^+$

    Acetic acid is a weak acid that produces hydronium ions when added to water

    $CH_3COO^- + H_2O \leftrightarrows CH_3COOH + OH^-$

    Acetate ions are weak bases that produce OH^- when added to water.

17. (a) An example of a strong electrolyte that is not an acid or base is any soluble salt, such as NaCl.

(b) Examples of Lewis acids include metal ions, BH_3, etc.

(c) Examples of weak acids are acetic, formic, benzoic, etc.

(d) An example of a buffer solution with a pH close to 7 is carbonic acid and sodium bicarbonate. K_a for the acid should be around 10^{-7}. If K_a is not close to this value you can still adjust the pH to 7.0, however, they may not have very good buffer capacity.

18. As you increase the number of oxygen atoms around the central atoms you increase the oxidation number of the central atom. This increases the aciditiy by weakening the O-H bond. The increase in oxidation number of the central atom causes it to shift charge from the O atoms increasing the polarity of the O-H bond.

19. Not all Lewis acids are Bronsted-Lowry (B-L) acids. B-L definition requires that the acid contain a proton capable of being remove, ionized, in the presence of a base, while the Lewis acid theory applies to any species that can accept an electron pair.

20. $HF + NH_3 \leftrightarrows F^- + NH_4^+$

$K = 1.4 \times 10^5$ so the equilibrium lies very far to the right. The reaction is between an acid, HF, and a base, NH_3. Based on this fact you would also predict that the reaction proceeds to the right.

21. An amphoteric species is one that can act as an acid or a base.

$HCO_3^- + H_2O \leftrightarrows H_2CO_3 + OH^-$

$HCO_3^- + H_2O \leftrightarrows CO_3^{2-} + H_3O^+$

22. We can use the approximation that $[A^-] - x$ can be approximated by $[A^-]$ because x is small, the % ionization is small, in comparison to the original concentration of the acid. We therefore do not lose a significant amount of acid during equilibrium. This approximation would not work if the original acid concentration were very low. For example, a 10^{-5} M acetic acid solution would give a proton concentration of 1.34×10^{-5} M (more that the original acid concentration which is impposible) using the approximation, and 7.16×10^{-6} M without the approximation.

23. The salt formed contains the conjugate acid of the weak base. Thus, hydrolysis results in a shift of the pH from 7.0 to a solution that has a pH less than 7.0

For example $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$ neutralization reaction

$NH_4^+ + H_2O \leftrightarrows NH_3 + H_3O^+$ hydrolysis of salt produced from neutralization.

24. Carbon-hydrogen bonds are essentially non-polar and thus will not release the H^+ to water. The only acid proton is on the polar O-H group and thus, benzoic acid, with a total of six hydrogen atoms only has one acid proton. See p 645

25. Before addition of 0.200 moles of sodium acetate to 0.700 M acetic acid the $[H_3O^+]$ is 3.55×10^{-3} M. After the addition of the salt the $[H_3O^+]$ is 6.3×10^{-5} M. $[H_3O^+]$ values were determined from the equilibrium expression.

Since the $[H_3O^+]$ is reduced when the salt is added, the reaction equilibrium must have shifted to the left.

$CH_3COOH + H_2O \leftrightarrows CH_3COO^- + H_3O^+$