

This set of review questions is designed to help you in preparation of the midterm exam, which will consist of the following three major topics: (1) **acids and bases** (K_a and K_b , acid-base titrations, titration curves, indicators, buffer and its calculation, and alkalinity); (2) **solubility of gases** (Henry's Law, BOD reduction in sewage, and solubility of oxygen and carbon dioxide in water); and (3) **solubility equilibria of ionic solids in water** (solubility rules, K_{sp} , common ion effect, partition equilibria and bioconcentration). These questions are intended to provide you with practice at solving quantitative problems. It is in your best interests to work through **all** these questions independently *before the exam*.

PART A Acids and Bases

- Consider aqueous solutions of the following solutes. For each solution listed below,
 - write formulas for the major solute species (ions or molecules) present
 - determine whether the solution is acidic, basic or neutral, justifying your answer with a balanced net ionic equation in the acidic and basic cases
 - write the equilibrium constant expression needed to calculate the pH of the solution (but do not calculate the pH)
 - NaCN
 - LiBr
 - KH_2PO_4
 - $\text{C}_2\text{H}_5\text{NH}_2$ (ethylamine)
 - $(\text{CH}_3)_3\text{NHCl}$ (trimethylammonium chloride)
- Calculate the pH of 0.10 M solutions of two weak acids, HA_1 and HA_2 with K_a values of 1.5×10^{-5} M and 4.6×10^{-3} M, respectively.
- Calculate K_c for each of the following reactions at 25°C. K_a for hypochlorous acid, HClO, is 3.0×10^{-8} M at 25°C.
 - $\text{H}_3\text{O}^+(\text{aq}) + \text{ClO}^-(\text{aq}) \rightleftharpoons \text{HClO}(\text{aq}) + \text{H}_2\text{O}(\ell)$
 - $\text{ClO}^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HClO}(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{HClO}(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\ell) + \text{ClO}^-(\text{aq})$
- Write a balanced net ionic equation and calculate the equilibrium constant (K_c) for the reaction that occurs when:
 - HBr(aq) is added to an aqueous Na_2CO_3 - NaHCO_3 buffer [$K_a(\text{HCO}_3^-)$: 5.6×10^{-11}]
 - aniline, $\text{C}_6\text{H}_5\text{NH}_2$, is titrated with hydrochloric acid [$K_b(\text{C}_6\text{H}_5\text{NH}_2)$: 3.8×10^{-10}]
 - $\text{Ca}(\text{OH})_2(\text{aq})$ is added to an aqueous $\text{CH}_3\text{NH}_3\text{Cl}$ - CH_3NH_2 buffer [$K_b(\text{CH}_3\text{NH}_2)$: 4.2×10^{-4}]
- The K_a of salicylic acid, $\text{HC}_7\text{H}_5\text{O}_3$, is 1.49×10^{-3} M at 25°C. Calculate the pH at the equivalence point in titration of a 2.00×10^{-2} M solution of salicylic acid with 2.00×10^{-2} M NaOH(aq) at 25°C. Choose an appropriate indicator for the titration with reference to the Table 11.6 in your textbook.
- Calculate the pH at half-equivalence point and at the equivalence point in the titration of 0.100 M $(\text{CH}_3)_3\text{N}(\text{aq})$ (trimethylamine) with 0.100 M $\text{HClO}_4(\text{aq})$. [K_b for $(\text{CH}_3)_3\text{N}(\text{aq})$ = 6.1×10^{-5} M]
- The following 0.10 M aqueous solutions are available:
formic acid sodium formate perchloric acid sodium hydroxide
Describe three ways to prepare a buffer, in each case using at least one different reagent, and write the net ionic equations to show any reactions occurred.
- In what ratio would you add 0.10 M HNO_2 and NaNO_2 solutions to obtain a buffer of pH = 3.0?
[$K_a(\text{HNO}_2)$ = 4.6×10^{-4}]
- Sketch a titration curve when 10 mL of 0.10 M NaA, a weak base with K_a for HA = 1.0×10^{-5} , is neutralized with 5.0 mL and 10.0 mL of 0.10 M HCl. Label the graph with appropriate axes, equivalence point, half-neutralization point, pK_a , major species at different regions and the approximate buffer region of HA.
- An acidic buffer is prepared from 250 mL of 0.10 M acetic acid, $\text{CH}_3\text{COOH}(\text{aq})$, with 25 mL of NaOH. Calculate the concentration of NaOH in order to make the acetate buffer at pH 4.90? [K_a for CH_3COOH is 1.8×10^{-5} .]
- Calculate the change in pH if 0.020 mol of HCl is added to 1.00 L of a solution that is 0.100 M in propionic acid, $\text{C}_2\text{H}_5\text{COOH}$, and 0.150 M in sodium propionate, $\text{C}_2\text{H}_5\text{COONa}$. [K_a for propionic acid is 1.34×10^{-5} M]
- Calculate the pH of a 0.25 M solution of sodium butyrate. [K_a for butyric acid is 1.5×10^{-5} M]
 - To a 100 mL of the above solution is added 20.0 mL of 1.00 M HCl. Write a net ionic equation and calculate the equilibrium constant for the reaction which occurs.
 - Calculate the pH of the solution after the HCl has been added.
 - Is the solutions prepared in (b) a buffer solution? Explain.
- The alkalinity of a 250 mL water sample is determined to be 4.31 mmol H^+ per litre. It requires 20.48 mL of HCl to reach the methyl orange endpoint. Calculate the concentration of HCl.

MULTIPLE CHOICE QUESTIONS

- Which of the following ionic salts forms an acidic solution when dissolved in water?
(a) LiClO_4 (b) K_2S (c) KF (d) NaHCO_3 (e) NaHSO_4
- Which of the following solutes, when dissolved in water to 1.0 M, gives the solution with the highest pH?
(a) HNO_2 (b) NaNO_2 (c) HI (d) NaI (e) NH_4I
- The pH of a 0.0025 M aqueous solution of calcium hydroxide is:
(a) 11.70 (b) 12.40 (c) 11.10 (d) 12.00 (e) 2.30
- The final pH of the solution prepared by mixing 100 mL of 0.10 M HCl and 100 mL of 0.10 M $\text{Ba}(\text{OH})_2$ is:
(a) 1.30 (b) 1.60 (c) 7.00 (d) 12.00 (e) 12.70
- For an aqueous solution labelled "0.10 M HNO_2 ", which of the following statements is true?
(a) $[\text{HNO}_2(\text{aq})] = K_a$ (b) $[\text{HNO}_2(\text{aq})] = [\text{H}^+(\text{aq})]$ (c) $[\text{HNO}_2(\text{aq})] = [\text{NO}_2^-(\text{aq})]$
(d) $[\text{H}_3\text{O}^+(\text{aq})] = 0.10 \text{ M}$ (e) $[\text{HNO}_2(\text{aq})]$ does not equal $[\text{H}^+(\text{aq})]$.
- The conjugate acid of hydrazine, N_2H_4 , is:
(a) H_2O (b) N_2H_3^- (c) N_2H_5^+ (d) H_3O^+ (e) N_2H_4
- Given that K_a for HClO and HCN are 3.0×10^{-8} and 4.8×10^{-10} , respectively. The equilibrium constant for the reaction
 $\text{HClO}(\text{aq}) + \text{CN}^-(\text{aq}) \rightleftharpoons \text{ClO}^-(\text{aq}) + \text{HCN}(\text{aq})$ is:
(a) 6.3×10^1 (b) 1.6×10^{12} (c) 1.6×10^{-2} (d) 1.4×10^{-3} (e) 1.4×10^{-17}
- The pH of a 0.0200 M solution of chloroacetic acid is 2.35. The K_a of chloroacetic acid is:
(a) 1.3×10^{-3} (b) 1.0×10^{-3} (c) 4.5×10^{-3} (d) 2.0×10^{-5} (e) 3.1×10^{-7}
- The pH of 0.20 M $\text{CH}_3\text{NH}_2(\text{aq})$ is 11.93. What is the value of K_b for CH_3NH_2 ?
(a) 8.5×10^{-3} (b) 1.4×10^{-5} (c) 7.2×10^{-5} (d) 1.7×10^{-3} (e) 3.8×10^{-4}
- Calculate the pH of a 0.136 M solution of aqueous sodium fluoride. $K_a(\text{HF}) = 3.5 \times 10^{-4}$.
(a) 2.16 (b) 5.71 (c) 8.30 (d) 8.73 (e) 11.84
- The pH of 1.0 M $\text{HCOOH}(\text{aq})$ is 1.88. What is the percent ionization of HCOOH ?
(a) 1.3% (b) 1.8% (c) 13% (d) 50% (e) 94%
- Which of the following is the strongest acid?
(a) HClO_2 ($\text{p}K_a = 2.00$) (b) HF ($\text{p}K_a = 3.46$) (c) HNO_2 ($\text{p}K_a = 3.34$)
(d) HCN ($\text{p}K_a = 9.32$) (e) CH_3COOH ($\text{p}K_a = 4.75$)
- Which of the following 0.10 M aqueous solution gives the highest pH?
(a) $\text{C}_5\text{H}_5\text{N}$ ($\text{p}K_b = 8.74$) (b) HF ($\text{p}K_a = 3.46$) (c) H_3PO_4 ($\text{p}K_a = 2.12$)
(d) HClO ($\text{p}K_a = 7.52$) (e) Since all are acids, the pH is the same for all solutions.
- For the titration of 25.0 mL of 0.100 M $\text{HF}(\text{aq})$ ($\text{p}K_a = 3.46$) with 0.100 M $\text{NaOH}(\text{aq})$, the main species in solution after addition of 12.5 mL of base are:
(a) $\text{HF}(\text{aq})$, $\text{F}^-(\text{aq})$ and $\text{H}^+(\text{aq})$ (b) $\text{HF}(\text{aq})$, $\text{OH}^-(\text{aq})$ and $\text{Na}^+(\text{aq})$
(c) $\text{F}^-(\text{aq})$ and $\text{Na}^+(\text{aq})$ (d) $\text{HF}(\text{aq})$, $\text{F}^-(\text{aq})$ and $\text{Na}^+(\text{aq})$
(e) $\text{F}^-(\text{aq})$, $\text{OH}^-(\text{aq})$ and $\text{Na}^+(\text{aq})$
- The solution at the equivalence point in a titration of 0.10 M $\text{NH}_3(\text{aq})$ with 0.10 M $\text{HCl}(\text{aq})$ is best described as:
(a) a solution of pH = 7.00 (b) a 0.10 M solution of NH_4Cl
(c) a 0.050 M solution of NH_4Cl (d) a 1:1 buffer solution
(e) a solution containing equal concentrations of NH_3 and NH_4Cl
- A 25.0 mL volume of 0.100 M weak acid, HA , is titrated with 0.100 M NaOH to the equivalence point. If the pH at the equivalence point is 8.28, calculate K_a for this acid.
(a) 1.4×10^{-4} (b) 7.3×10^{-4} (c) 5.2×10^{-9} (d) 1.9×10^{-6} (e) 1.0
- If one were to prepare a buffer solution using pyridine ($\text{C}_5\text{H}_5\text{N}$, $K_b = 1.8 \times 10^{-9} \text{ M}$) and pyridinium chloride ($\text{C}_5\text{H}_5\text{NHCl}$), its pH would be in the approximate range:
(a) 7.7 to 9.7 (b) 4.3 to 6.3 (c) 2.9 to 4.9 (d) 9.3 to 11.3 (e) 11.1 to 13.1
- Which pair of reagents, mixed in the appropriate proportion, could make a buffer in pH 3-4 range:
(a) NH_3 and NH_4Cl (K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$)
(b) HNO_2 and NaOH (K_a for $\text{HNO}_2 = 4.6 \times 10^{-4}$)
(c) $\text{C}_5\text{H}_5\text{N}$ and HCl (K_b for $\text{C}_5\text{H}_5\text{N} = 1.8 \times 10^{-9}$)
(d) HClO and NaClO (K_a for $\text{HClO} = 3.0 \times 10^{-8}$)
- Which of the following indicators would be most suitable for the titration of 0.10 M $\text{C}_5\text{H}_5\text{NHCl}(\text{aq})$ with 0.10 M $\text{KOH}(\text{aq})$? [K_b for $\text{C}_5\text{H}_5\text{N}$ is 1.8×10^{-9}]
(a) methyl red ($\text{p}K_{\text{In}} = 5.0$)
(b) thymol blue ($\text{p}K_{\text{In}} = 1.7$)
(c) phenolphthalein ($\text{p}K_{\text{In}} = 9.4$)
(d) bromothymol blue ($\text{p}K_{\text{In}} = 7.1$)

20. Which of the following aqueous solutions is a buffer with a pH greater than 7.0? For HNO_2 , $K_a = 4.6 \times 10^{-4}$, and for NH_3 , $K_b = 1.8 \times 10^{-5}$.
- (a) 10 mL of 0.1 M $\text{NH}_3(\text{aq})$ + 10 mL of 0.1 M $\text{HCl}(\text{aq})$
 (b) 10 mL of 0.1 M $\text{HNO}_2(\text{aq})$ + 5.0 mL of 0.1 M $\text{NaOH}(\text{aq})$
 (c) 10 mL of 0.1 M $\text{HNO}_2(\text{aq})$ + 10 mL of 0.1 M $\text{NaOH}(\text{aq})$
 (d) 10 mL of 0.1 M $\text{NH}_3(\text{aq})$ + 5.0 mL of 0.1 M $\text{HCl}(\text{aq})$
 (e) 10 mL of 0.1 M $\text{NH}_3(\text{aq})$ + 10 mL of 0.1 M $\text{HNO}_2(\text{aq})$
21. Calculate the pH of a buffer solution consisting of 0.40 M $\text{NH}_4\text{Cl}(\text{aq})$ and 0.20 M $\text{NH}_3(\text{aq})$. [K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$]
- (a) 8.95 (b) 9.26 (c) 9.56 (d) 5.05 (e) 4.44
22. A sample of 100 mL of 0.10 M weak acid, HA ($K_a = 1.0 \times 10^{-5}$), is titrated with standard 0.10 M KOH. How many mL of KOH will have been added when the pH in the titration flask is 5?
- (a) 0 (b) 10 (c) 25 (d) 50 (e) 100
23. What volume of 0.40 M NH_3 solution must be added to 1.0 L of 0.10 M NH_4Cl solution to give a buffer having pH of 10.00? For NH_3 , $K_b = 1.8 \times 10^{-5}$
- (a) 1.4 L (b) 1.1 L (c) 0.97 L (d) 0.61 L (e) 0.29 L
24. In the titration of 25.0 mL of 0.100 M aqueous acetic acid ($K_a = 1.8 \times 10^{-5}$) with 0.100 M $\text{NaOH}(\text{aq})$, the pH after addition of 10.0 mL of titrant is:
- (a) 4.34 (b) 4.57 (c) 4.76 (d) 4.92 (e) 5.14
25. What is the change in pH that results from the addition of 20 mL of 0.10 M NaOH to a buffer made by combining 200 mL of 0.070 M benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$, $K_a = 6.4 \times 10^{-5}$) with 100 mL of 0.070 M sodium benzoate?
- (a) +0.31 (b) +0.20 (c) +0.18 (d) -0.20 (e) -0.31

PART B Solubility of Gases; BOD

1. A carbonated drink is bottled at 25°C and contains CO_2 gas at a pressure of 5.0 atm over the liquid. If the partial pressure of CO_2 in the atmosphere is 3.5×10^{-4} atm, calculate the equilibrium concentration of CO_2 in the drink (a) before and (b) after the bottle is opened. [K_H for $\text{CO}_2(\text{g}) = 3.4 \times 10^{-2} \text{ M atm}^{-1}$]
2. (a) Given that K_H for $\text{O}_2(\text{g})$ is $1.3 \times 10^{-3} \text{ M atm}^{-1}$ and the partial pressure of $\text{O}_2(\text{g})$ in the atmosphere is 0.21 atm at 25°C. Calculate the amount of dissolved oxygen, O_2 , in natural waters at 25°C in:
 (i) mol L^{-1}
 (ii) ppm
 (b) The organic matter, ethyl alcohol $\text{C}_2\text{H}_6\text{O}$, was added to the unpolluted lake. Total concentration of organic matter in the lake reached 5.00 ppm.
 (i) What is the BOD for the lake in mol L^{-1} and in ppm?
 (ii) Is there enough dissolved oxygen in the lake?
3. A 200 mL sample of oxygenated water is equilibrated with the atmosphere at 4°C and then is analyzed by the Winkler method (see equations below).
- $$\text{Mn}^{2+}(\text{aq}) + 2 \text{OH}^{-}(\text{aq}) + \frac{1}{2} \text{O}_2(\text{aq}) \rightarrow \text{MnO}_2(\text{s}) + \text{H}_2\text{O}(\ell)$$
- $$\text{MnO}_2(\text{s}) + 4 \text{H}^{+}(\text{aq}) + 2 \text{I}^{-}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{I}_2(\text{aq}) + 2 \text{H}_2\text{O}(\ell)$$
- $$\text{I}_2(\text{aq}) + 2 \text{S}_2\text{O}_3^{2-}(\text{aq}) \rightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2 \text{I}^{-}(\text{aq})$$
- The Henry's law constant for the dissolution of $\text{O}_2(\text{aq})$ in water at 4°C is determined to be $2.02 \times 10^{-3} \text{ mol L}^{-1} \text{ atm}^{-1}$. Calculate the concentration, in ppm, of $\text{O}_2(\text{aq})$ at equilibrium, knowing the atmospheric partial pressure of $\text{O}_2(\text{g})$ is 0.209 atm. What volume of 0.02468 M $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ is needed to reduce all the I_2 ?
4. Discuss the following topics:
- Various ways of measuring status of oxygen in natural water
 - BOD reduction of sewage and industrial waste water
 - Phosphate removal from sewage

PART C Solubility Equilibria of Ionic Compounds

1. Calculate the solubility of the following ionic salts in M (mol L^{-1}) and in g L^{-1} .
- (a) CaC_2O_4 $K_{\text{sp}} = 2.0 \times 10^{-9} \text{ M}^2$ (molar mass: 128 g mol^{-1})
 (b) Ag_2SO_4 $K_{\text{sp}} = 1.2 \times 10^{-5} \text{ M}^3$ (molar mass: 311.7 g mol^{-1})
 (c) $\text{Al}(\text{OH})_3$ $K_{\text{sp}} = 5.0 \times 10^{-33} \text{ M}^4$ (molar mass: 78 g mol^{-1})
2. Calculate K_{sp} from the following solubility data:
- (a) lead (II) fluoride $S = 2.1 \times 10^{-3} \text{ M}$
 (b) silver (I) sulfate $S = 4.5 \text{ g L}^{-1}$ (molar mass: 312 g mol^{-1})
 (c) silver (I) phosphate $S = 1.6 \times 10^{-5} \text{ M}$

3. Which of the ionic salts listed below has greater solubility (in M, mol L⁻¹)?
 Ag_2CrO_4 $K_{\text{sp}} = 1.1 \times 10^{-12} \text{ M}^3$
 AgCl $K_{\text{sp}} = 1.8 \times 10^{-10} \text{ M}^2$
4. K_{sp} for $\text{BaF}_2(\text{s})$ is $1.0 \times 10^{-6} \text{ M}^3$.
 (a) Calculate $[\text{Ba}^{2+}]$ in a saturated solution of BaF_2 .
 (b) Calculate $[\text{Ba}^{2+}]$ in a 0.10 M solution of NaF .
 (c) How many grams of BaF_2 will dissolve in a 0.50 L solution of 0.10 M NaF ? [$M(\text{BaF}_2) = 175 \text{ g mol}^{-1}$]
5. Calculate the solubility of $\text{Zn}(\text{OH})_2(\text{s})$ in a solution buffered at $\text{pH} = 8.0$. [K_{sp} for $\text{Zn}(\text{OH})_2 = 1.0 \times 10^{-15} \text{ M}^3$]
6. (a) At what pH will $\text{Pb}(\text{OH})_2$ precipitate from a $\text{Pb}(\text{NO}_3)_2$ solution which has a Pb^{2+} concentration of 660 ppm?
 (b) Will higher in pH reduce or increase solubility of $\text{Pb}(\text{OH})_2$? Explain.
 [K_{sp} for $\text{Pb}(\text{OH})_2 = 2.8 \times 10^{-16} \text{ M}^3$; molar mass of $\text{Pb} = 207.2 \text{ g mol}^{-1}$]
7. Show by calculation whether precipitation occurs when 0.10 L of $1.0 \times 10^{-4} \text{ mol L}^{-1} \text{ NaF}$ is added to 0.10 L of $5.0 \times 10^{-5} \text{ mol L}^{-1} \text{ CaCl}_2$. [K_{sp} for $\text{CaF}_2 = 1.6 \times 10^{-10} \text{ M}^3$]
8. Some barium chloride is added to a solution that contains both K_2SO_4 (0.050 M) and Na_3PO_4 (0.010 M). Which precipitate begins to form first? [K_{sp} for barium sulfate and barium phosphate are 8.7×10^{-11} and 6.0×10^{-39} .]
9. Which of the following changes listed below will shift the ZnCO_3 solubility equilibrium to the right?
 (a) increase $[\text{Zn}^{2+}]$ (c) addition of NaOH
 (b) addition of HCl (d) addition of solid ZnCO_3
10. Review "Partition equilibria - Extraction of non-polar compounds from water" from the textbook.

ANSWERS

PART A Acids and Bases

1. (i) basic (ii) neutral (iii) acidic (iv) basic (v) acidic
2. pH for $\text{HA}_1 = 2.91$; pH for $\text{HA}_2 = 1.72$ 3.(a) $3.3 \times 10^7 \text{ M}^{-1}$ (b) $3.3 \times 10^{-7} \text{ M}$ (c) $3.0 \times 10^6 \text{ M}^{-1}$
4. (a) $\text{H}_3\text{O}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{H}_2\text{O}(\ell)$ $K_c = 1.8 \times 10^{10}$
 (b) $\text{C}_5\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightleftharpoons \text{C}_5\text{H}_5\text{NH}_3^+(\text{aq}) + \text{H}_2\text{O}(\ell)$ $K_c = 3.8 \times 10^4$
 (c) $\text{OH}^-(\text{aq}) + \text{CH}_3\text{NH}_3^+(\text{aq}) \rightleftharpoons \text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\ell)$ $K_c = 2.4 \times 10^3$
5. $\text{pH} = 7.413$; bromothymol blue and phenolphthalein
6. pH at half-equivalence point: 9.79; pH at the equivalence point: 5.54
7. (a) weak acid + strong base: $\text{HCOOH}(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{H}_2\text{O}(\ell)$
 (b) weak base + strong acid: $\text{HCOO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightleftharpoons \text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\ell)$
 (c) weak acid + weak base: $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
8. $[\text{NO}_2^-] : [\text{HNO}_2] = 0.46 : 1.0$ 10. 0.59 M 11. $\Delta\text{pH} = (-) 0.141 \text{ pH units}$
12. (a) $\text{pH} = 9.11$ (b) $6.7 \times 10^4 \text{ M}^{-1}$ (c) $\text{pH} = 4.22$ (d) yes
13. $5.26 \times 10^{-2} \text{ M}$

MULTIPLE CHOICE QUESTIONS

- | | | | | |
|------|-------|-------|-------|-------|
| 1. e | 6. c | 11. a | 16. a | 21. a |
| 2. b | 7. a | 12. a | 17. b | 22. d |
| 3. a | 8. a | 13. d | 18. b | 23. a |
| 4. e | 9. e | 14. d | 19. c | 24. b |
| 5. e | 10. c | 15. c | 20. d | 25. c |

PART B Solubility of Gases; BOD

1. (a) 0.17 mol L^{-1} (b) $1.2 \times 10^{-5} \text{ mol L}^{-1}$
2. (a) (i) $2.7 \times 10^{-4} \text{ mol L}^{-1}$ (ii) 8.7 ppm
 (b) (i) $3.26 \times 10^{-4} \text{ mol L}^{-1}$; 10.4 ppm (ii) not enough
3. 13.5 ppm; 13.7 mL

PART C Solubility Equilibria of Ionic Compounds

1. (a) $4.5 \times 10^{-5} \text{ mol L}^{-1}$; $5.7 \times 10^{-3} \text{ g L}^{-1}$ (b) $1.4 \times 10^{-2} \text{ mol L}^{-1}$; 4.5 g L^{-1}
 (c) $3.7 \times 10^{-9} \text{ mol L}^{-1}$; $2.9 \times 10^{-7} \text{ g L}^{-1}$
2. (a) $3.7 \times 10^{-8} \text{ M}^3$ (b) $1.2 \times 10^{-5} \text{ M}^3$ (c) $1.8 \times 10^{-18} \text{ M}^4$
3. Ag_2CrO_4 4. (a) $6.3 \times 10^{-3} \text{ M}$ (b) $1.0 \times 10^{-4} \text{ M}$ (c) 8.8 mg
5. $1.0 \times 10^{-3} \text{ M}$ 6. (a) $\text{pH} = 7.47$ (b) will be reduced
7. CaF_2 will not precipitate
8. $\text{Ba}_3(\text{PO}_4)_2$ will precipitate first 9. b