Hydrogen ion concentration- *p*H scale and Buffer solution

181. A weak monoprotic acid of 0.1 *M*, ionizes to 1% in solution. What will be the pH of solution
(a) 1 (b) 2
(c) 3 (d) 11

182. H_4O^+ of a solution is 4. The hydroxide ion concentration of the solution would be
(a) 10^{-4} (b) 10^{-10}
(c) 10^{-2} (d) 10^{-12}

183. The pH of an aqueous solution containing $[H^+] = 3 \times 10^{-3} M$ is
(a) 2.471 (b) 2.523
(c) 3.0 (d) – 3

184. pH of blood is maintained constant by mechanism of
(a) Common ion effect
(b) Buffer
(c) Solubility
(d) All of these

185. The *pH* of normal *KOH* is
(a) 1 (b) 0
(c) 14 (d) 7

186. The concentration of hydrogen ion $[H^+]$ in 0.01 *M* HCl is
(a) 10^{12} (b) 10^{-2}
(c) 10^{-1} (d) 10^{-12}

187. A solution of weak acids is diluted by adding an equal volume of water. Which of the following will not change
(a) Strength of the acid
(b) The value of $[H_3O^+]$
(c) $pH$of the solution
(d) The degree of dissociation of acid

188. $Ka$of $H_2O_2$is of the order of
(a) 10^{-12} (b) 10^{-14}
(c) 10^{-16} (d) 10^{-10}

189. Equivalent weight of an acid
(a) Depends on the reaction involved
(b) Depends upon the number of oxygen atoms present
(c) Is always same
(d) None of the above

190. pHscale was introduced by
(a) Arrhenius (b) Sorensen
(c) Lewis (d) Lowry

191. Buffer solution is prepared by mixing
(a) Strong acid + its salt of strong base
(b) Weak acid + its salt of weak base
(c) Strong acid + its salt of weak base
base

(d) Weak acid + its salt of strong base

192. The pH of millimolar HCl is]
(a) 1 (b) 3
(c) 2 (d) 4

193. Which of the following is a Lewis base
(a) NaOH (b) NH_3
(c) BCl_3 (d) All of these

194. What will be the pHvalue of 0.05 M $Ba(OH)_2$solution
(a) 12 (b) 13
(c) 1 (d) 12.96

195. In a mixing of acetic acid and sodium acetate the ratio of concentration of the salts to the acid is increased ten times. Then the pH of the solution
(a) Increase by one
(b) Decreases by one
(c) Decrease ten fold
(d) Increases ten fold

196. The rapid change of pH near the stoichiometric point of an acid-base titration is the basis of indicator detection. pH of the solution is related to ratio of the concentrations of the conjugate acid (HIn) and base (In^-) forms of the indicator by the expression
(a) $log \frac{[HIn]}{[In^-]} = pH - pK_{In}$
(b) $log \frac{[In^-]}{[HIn]} = pH - pK_{In}$
(c) $log \frac{[In^-]}{[HIn]} = pK_{In} - pH$
(d) $log \frac{[HIn]}{[In^-]} = pK_{In} - pH$

197. Which of the following statement(s) is(are) correct
(a) The pH of $1.0 \times 10^{-8} M$ solution of HCl is 8
(b) The conjugate base of $H_2PO_4^-$ is HPO_4^{2-}
(c) Autoprotolysis constant of water increases with temperature
(d) When a solution of a weak monoprotic acid is titrated against a strong base, at half neutralization point $pH = \frac{1}{2}pK_a$

198. An aqueous solution of sodium carbnate has a pH greater than 7 becaue
(a) It contains more carbonate ions than H_2O molecules
(b) Contains more hydroxide ions than carbonate ions
(c) Na^+ions react with water
(d) Carbonate ions react with H_2O

199. A pH of 7 signifies

(a) Pure water

(b) Neutral solution

(c) Basic solution

(d) Acidic solution

200. Assuming complete dissociation, which of the following aqueous solutions will have the same pHvalue

(a) NH_3 of $0.01 MHCl$

(b) $10\frac{1}{K_a}$ of $0.01 MH_2SO_4$

(c) $50 ml$ of $10\frac{1}{K_a}$

(d) Mixture of 50 ml of $0.02 MH_2SO_4$ and $50 ml$ of $0.02 MNaOH$

201.