

181. (c) $[H^+] = C \cdot \alpha$

$$[H^+] = 0.1 \times \frac{1}{100} = 10^{-3}$$

$$pH = -\log [H^+] = -\log 10^{-3} = 3$$

182. (b) $pH = 4$

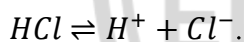
$$pH + pOH = 14 ; pOH = 14 - pH$$

$$pOH = 14 - 4 = 10 ; [OH^-] = 10^{-10} M$$

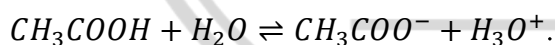
183. (b) $pH = \log \frac{1}{[H^+]} = \log \frac{1}{[3 \times 10^{-3}]} = 2.523$

185. (c) It is a strong base.

186. (b) $[H^+]$ Concentration in $0.01 M HCl$ is $10^{-2} M$ because $0.01 M HCl$ have only H^+



187. (b) The value of H_3O^+ ions will not changed.



188. (a) H_2O_2 (Hydrogen peroxide) is a corrosive volatile liquid. It is slightly acidic in nature. Its pK_a value is approximately 10^{-12} .

189. (a) Depends on the reaction involved

Explanation:



The equivalent weight of an acid is the **mass of acid that supplies 1 mole of H^+ ions** in a reaction.

It **varies depending on how the acid reacts** (e.g., H_2SO_4 can donate 1 or 2 protons depending on the reaction).

190. (b) Sorensen

Explanation:

pH scale (power of hydrogen) was introduced by **S.P.L. Sorensen in 1909** to express the acidity or basicity of a solution.

191. (d) Weak acid + its salt of strong base

Explanation:

A **buffer solution** resists pH changes.

Acidic buffer: Weak acid + its salt with strong base

Example: Acetic acid + Sodium acetate \rightarrow pH < 7

Basic buffer: Weak base + its salt with strong acid

Example: Ammonia + Ammonium chloride \rightarrow pH > 7

192. (b) $pH = \log \frac{1}{[H^+]}$; $pH = \log \frac{1}{10^{-3}}$; $pH = 3$.

193. (b) NH_3 is Lewis base because of one lone pair of electron.

194. (b) $Ba(OH)_2 \rightarrow Ba^{+2}_{.05M} + 2OH^{-}_{2 \times 0.5M}$

$$pOH = \log \frac{1}{[OH]^-} = \log \frac{1}{.1} = 1$$

$$pH + pOH = 14; pH + 1 = 14; pH = 14 - 1 = 13$$

195. (a) If concentration of acid increases ten times in a buffer then **pH** of the solution increases by one.





196. (b) $\log \left(\frac{[In^-]}{[HIn]} \right) = pH - pK_{In}$

Explanation (Word-friendly):

Henderson-Hasselbalch equation for an indicator: $pH = pK_{In} + \log \left(\frac{[In^-]}{[HIn]} \right)$

Rearranging gives: $\log \left(\frac{[In^-]}{[HIn]} \right) = pH - pK_{In}$

197. (b), (c), (d)

Explanation (Word-Friendly):

Statement (a) incorrect

Very dilute HCl (1×10^{-8} M) \rightarrow contribution of water autoionization is significant.

pH \approx 7, not 8.

Statement (b) Correct

$H_2PO_4^-$ loses a proton \rightarrow forms HPO_4^{2-} , which is its conjugate base.

Statement (c) Correct

The **autoprotolysis constant of water (K_w)** increases with temperature.

Statement (d) Correct

For weak monoprotic acid titrated with strong base:

At **half-neutralization**, $[HA] = [A^-] \rightarrow pH = pK_a$ (from Henderson-Hasselbalch equation).

198. (b) $pH > 7$ = Basic

It means contain more hydroxide ions than carbonate ions.

199. (a) At 7pH the concentration of OH^- and H^+ are equal.

200. (a,d) M.eq. of 0.01 $MHCl = \frac{.01 \times 100}{1000} = 1 \times 10^{-3}$

$pH = 3$

M.eq. of .02 $MH_2SO_4 = \frac{.04 \times 50}{1000} = 2 \times 10^{-3}$

M.eq. of .02 $MNaOH = \frac{0.02 \times 50}{1000} = 1 \times 10^{-3}$



$$\text{Left } [H^+] = 2 \times 10^{-3} - 1 \times 10^{-3} = 1 \times 10^{-3}; pH = 3$$

