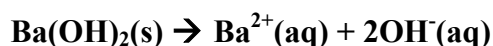


- Calculate the pH of a 0.020 M solution of Ba(OH)<sub>2</sub>.

Marks  
1

Ba(OH)<sub>2</sub> is a strong base so it will completely dissociate in solution:



As each Ba(OH)<sub>2</sub> dissociates to make 2OH<sup>-</sup>, a 0.020 M solution has [OH<sup>-</sup>(aq)] = 0.040 M.

By definition, pOH = -log<sub>10</sub>[OH<sup>-</sup>(aq)] so pOH = -log<sub>10</sub>(0.040) = 1.40.

As pH + pOH = 14.00,

$$\text{pH} = 14.00 - 1.40 = 12.60$$

$$\text{pH} = 12.60$$

- Calculate the pH of a 0.150 M solution of HNO<sub>2</sub>. The pK<sub>a</sub> of HNO<sub>2</sub> is 3.15.

3

As HNO<sub>2</sub> is a weak acid, [H<sub>3</sub>O<sup>+</sup>] must be calculated using a reaction table:

	HNO <sub>2</sub>	H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	NO <sub>2</sub> <sup>-</sup>
initial	0.150	large		0	0
change	-x	negligible		+x	+x
final	0.150 - x	large		x	x

The equilibrium constant K<sub>a</sub> is given by:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{x^2}{0.150 - x}$$

As pK<sub>a</sub> = -log<sub>10</sub>K<sub>a</sub>, K<sub>a</sub> = 10<sup>-3.15</sup> and is very small, 0.150 - x ~ 0.150 and hence:

$$x^2 = 0.150 \times 10^{-3.15} \quad \text{or} \quad x = 1.03 \times 10^{-2} \text{ M} = [\text{H}_3\text{O}^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(1.03 \times 10^{-2}) = 1.987$$

$$\text{pH} = 1.987$$

ANSWER CONTINUES ON THE NEXT PAGE

- Calculate the pH of a solution that is 0.080 M in acetic acid and 0.160 M in sodium acetate. The pK<sub>a</sub> of acetic acid is 4.76.

**The solution contains both a weak acid (acetic acid) and its conjugate base (the acetate ion). This is a buffer and its pH can be calculate using the Henderson-Hasselbalch equation. With [acid] = 0.080 M and [base] = 0.160 M,**

$$\text{pH} = \text{p}K_{\text{a}} + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right) = 4.76 + \log\left(\frac{0.160}{0.080}\right) = 5.06$$

$$\text{pH} = 5.06$$