

**PART A
MC #1**

Refer to the Table:
“Acid-Base Properties of Some common Ions in Aqueous solution”

<u>solution</u>	<u>cation</u>	<u>anion</u>	<u>RESULTANT</u>
(a) LiClO ₄	Li ⁺ (neutral)	ClO ₄ ⁻ (neutral)	neutral
(b) K ₂ S	K ⁺ (neutral)	S ²⁻ (<i>basic</i>)	<i>basic</i>
(c) KF	K ⁺ (neutral)	F ⁻ (<i>basic</i>)	<i>basic</i>
(d) NaHCO ₃	Na ⁺ (neutral)	HCO ⁻ (<i>basic</i>)	<i>basic</i>
(e) NaHSO ₄	Na ⁺ (neutral)	HSO ₄ ⁻ (<i>acidic</i>)	ACIDIC

**PART A
MC #2**

Highest pH → MOST BASIC (i.e. least acidic)

<u>solution</u>	<u>cation</u>	<u>anion</u>	<u>RESULTANT</u>
(a) HNO ₂	H ⁺ (<i>acidic</i>)	NO ₂ ⁻ (<i>basic</i>)	<i>weak acid</i>
(b) NaNO ₂	Na ⁺ (neutral)	NO ₂ ⁻ (<i>basic</i>)	BASIC
(c) HI	H ⁺ (<i>acidic</i>)	I ⁻ (neutral)	<i>strong acid</i>
(d) NaI	Na ⁺ (neutral)	I ⁻ (neutral)	neutral
(e) NH ₄ I	NH ₄ ⁺ (<i>acidic</i>)	I ⁻ (neutral)	<i>weak acid</i>

Decreasing order of acidity: HI > HNO₂ > NH₄I > NaI > NaNO₂

**PART A
MC #3**

Ca(OH)₂(aq) is a *strong base*.
1 mole of Ca(OH)₂(aq) contains **1 mole** of Ca²⁺(aq) and **2 moles** of OH⁻(aq).
 $\therefore c(\text{Ca}(\text{OH})_2) = 0.0025 \text{ M} = c(\text{Ca}^{2+})$

$$\begin{aligned} \therefore c(\text{OH}^-) &= c(\text{Ca}(\text{OH})_2) \times (2 \text{ moles OH}^- \div 1 \text{ mole Ca}(\text{OH})_2) \\ &= (0.0025 \text{ M}) \times (2) \\ &= 0.0050 \text{ M} \end{aligned}$$

pOH = - log [OH⁻]

$$\text{pOH} = - \log (0.0050 \text{ M}) = 2.30$$

$$\therefore \text{pH} = 14.00 - \text{pOH} = 14.00 - 2.30 = \underline{\underline{11.70}}$$

PART A
MC #4

Given: $c(\text{HCl}) = 0.10 \text{ M}$ $V(\text{HCl}) = 100 \text{ mL}$
 $c(\text{Ba}(\text{OH})_2) = 0.10 \text{ M}$ $V(\text{Ba}(\text{OH})_2) = 100 \text{ mL}$

To find: final pH

The complete reaction is: $2 \text{HCl}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{BaCl}_2(\text{aq}) + 2 \text{H}_2\text{O}(\ell)$

- ① $n(\text{HCl}) = c(\text{HCl}) \times V(\text{HCl}) = (0.10 \text{ mol L}^{-1}) \times (100 \times 10^{-3} \text{ L}) = 0.010 \text{ mol} = n(\text{H}^+)$
② $n(\text{Ba}(\text{OH})_2) = c(\text{Ba}(\text{OH})_2) \times V(\text{Ba}(\text{OH})_2) = (0.10 \text{ mol L}^{-1}) \times (100 \times 10^{-3} \text{ L}) = 0.010 \text{ mol}$

- ③ **1 mole** $\text{Ba}(\text{OH})_2(\text{aq})$ contains **2 moles** of $\text{OH}^-(\text{aq})$

$$\begin{aligned} n(\text{OH}^-) &= n(\text{Ba}(\text{OH})_2) \times (2 \text{ moles OH}^- \div 1 \text{ mole Ba}(\text{OH})_2) \\ &= (0.010 \text{ mol}) \times (2) \\ &= 0.020 \text{ mol} \end{aligned}$$

2 moles HCl reacts with **1 mole** $\text{Ba}(\text{OH})_2$ (or 2 moles $\text{OH}^-(\text{aq})$)

For H^+ , ratio = $(0.010 \text{ mol}) \div 2 \text{ moles} = 0.0050$

For OH^- , ratio = $(0.020 \text{ mol}) \div 2 \text{ moles} = 0.010$

$\therefore \text{OH}^-(\text{aq})$ is *in excess* and H^+ is the *limiting reactant*.

Neutralization of **SA + SB:** $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$

1 mole $\text{H}^+(\text{aq})$ reacts with **1 mole** $\text{OH}^-(\text{aq})$

After mixing, $n(\text{OH}^-)$ *leftover* = $n(\text{OH}^-) - n(\text{H}^+)$
= $(0.020 \text{ mol}) - (0.010 \text{ mol})$
= 0.010 mol

$$\text{total volume} = 100 \text{ mL HCl} + 100 \text{ mL Ba}(\text{OH})_2 = 200 \text{ mL} = 200 \times 10^{-3} \text{ L}$$

$$\begin{aligned} \text{final } c(\text{OH}^-) &= n(\text{OH}^-) \text{ leftover} \div V(\text{total volume}) \\ &= (0.010 \text{ mol}) \div (200 \times 10^{-3} \text{ L}) \\ &= 0.050 \text{ mol L}^{-1} \end{aligned}$$

$$\text{pOH} = -\log [\text{OH}^-]$$

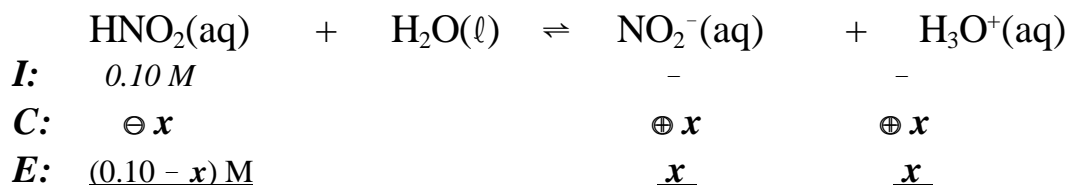
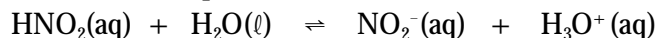
$$\text{pOH} = -\log (0.050 \text{ mol L}^{-1}) = 1.301$$

$$\therefore \text{pH} = 14.00 - \text{pOH} = 14.00 - 1.301 = 12.699 = \underline{\underline{12.70}}$$

PART A
MC #5

Given: 0.10 M HNO₂(aq)

- ① HNO₂(aq) is a *weak acid*, which only partially dissociates in water.
 ② The acid dissociation equilibrium is:



(where $x = [\text{H}_3\text{O}^+] \text{ produced} = [\text{NO}_2^-] \text{ produced} = [\text{HNO}_2] \text{ dissociated}$)

$$K_c = \frac{[\text{NO}_2^-] [\text{H}_3\text{O}^+]}{[\text{HNO}_2]} = K_a (\text{HNO}_2)$$

- (a) $[\text{HNO}_2(\text{aq})] = K_a$ **UNTRUE**
 (∴ K_a should be equal to the above K_c expression.)
- (b) $[\text{HNO}_2(\text{aq})] = [\text{H}^+(\text{aq})]$ **UNTRUE**
 (∴ total $[\text{HNO}_2(\text{aq})] = \text{dissociated } [\text{H}^+(\text{aq})] + \text{undissociated } [\text{HNO}_2(\text{aq})]$
 Only the strong acids (SA) being completely dissociated in water
 will have $[\text{SA}] = [\text{H}^+(\text{aq})]$ (e.g. $[\text{HCl}] = [\text{H}^+(\text{aq})]$)
- (c) $[\text{HNO}_2(\text{aq})] = [\text{NO}_2^-(\text{aq})]$ **UNTRUE**
 (∴ **At equilibrium,** $[\text{NO}_2^-(\text{aq})] \text{ produced} = [\text{H}^+(\text{aq})] \text{ produced} = x$
 $[\text{HNO}_2(\text{aq})] \text{ leftover} = (0.10 - x)$)
- (d) $[\text{H}_3\text{O}^+(\text{aq})] = 0.10 \text{ M}$ **UNTRUE**
 (∴ HNO₂(aq) is a *weak acid*, which only partially dissociates in aqueous solution.)
 This statement is correct only for the strong acids
 (e.g. HCl, HBr, HI, H₂SO₄, HNO₃ and HClO₄).
- (e) **$[\text{HNO}_2(\text{aq})]$ does not equal $[\text{H}^+(\text{aq})]$** **TRUE**
 (∴ **At equilibrium,** $[\text{HNO}_2(\text{aq})] = (0.10 - x)$ and
 $[\text{H}^+(\text{aq})] \text{ produced} = x = [\text{NO}_2^-(\text{aq})] \text{ produced}$)