

1. Calculate the pH of a solution that is obtained by mixing 200.0 mL of 0.1465 M Nitrous acid with 167.8 mL of 0.1473 M Sodium Nitrite.

This solution is a weak acid (Nitrous acid) with its conjugate base (Nitrite ion). It is a buffer.

$$n_{\text{HNO}_2} = 200.0 \text{ mL} \times \frac{0.1465 \text{ mol HNO}_2}{1000 \text{ mL HNO}_2} = 0.02930 \text{ mol}$$

$$n_{\text{NO}_2^-} = 167.8 \text{ mL} \times \frac{0.1473 \text{ mol NO}_2^-}{1000 \text{ mL NO}_2^-} = 0.02472 \text{ mol}$$

$$\text{pH} = \text{p}K_a + \log \frac{n_{\text{NO}_2^-}}{n_{\text{HNO}_2}} = 3.35 + \log \frac{0.02472}{0.02930} = 3.27$$

2. Calculate the pH of a solution that is obtained by mixing 200.0 mL of 0.1465 M Nitric acid with 167.8 mL of 0.1473 M Sodium Hydroxide.

We have a strong base and a strong acid present so we must do the neutralization first (Chem 101).



$$n_{\text{H}^+} = 200.0 \text{ mL} \times \frac{0.1465 \text{ mol H}^+}{1000 \text{ mL}} = 0.02930 \text{ mol H}^+$$

$$n_{\text{OH}^-} = 167.8 \text{ mL} \times \frac{0.1473 \text{ mol OH}^-}{1000 \text{ mL}} = 0.02472 \text{ mol OH}^-$$

OH^- is the limiting reactant. The moles of H^+ remaining are:

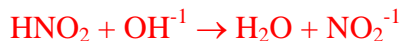
$$0.02930 - 0.02472 = 4.58 \times 10^{-3} \text{ mol H}^+$$

$$[\text{H}^+] = \frac{4.58 \times 10^{-3} \text{ mol H}^+}{367.8 \text{ mL} \times \frac{10^{-3} \text{ L}}{1 \text{ mL}}} = 0.0125 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -\log 0.0125 = 1.90$$

3. Calculate the pH of a solution that is obtained by mixing 200.0 mL of 0.1465 M Nitrous acid with 167.8 mL of 0.1473 M Sodium Hydroxide.

We have a strong base and a weak acid. We do the neutralization first and see what remains.



$$n_{\text{OH}^{-}} = 0.02472 \text{ mol} \quad (\text{see calcs above})$$

$$n_{\text{HNO}_2} = 0.02930 \text{ mol}$$

L.R. is OH^{-} .

$$n'_{\text{HNO}_2} = 0.02930 - 0.02472 = 4.58 \times 10^{-3}$$

$$n'_{\text{NO}_2^{-1}} = 0.02472$$

This is a buffer solution.

$$\text{pH} = \text{p}K_a + \log \frac{n_{\text{NO}_2^{-1}}}{n_{\text{HNO}_2}} = 3.35 + \log \frac{0.02472}{4.58 \times 10^{-3}} = 4.08$$

4. Calculate the pH of a solution that is 0.1465 M Sodium Nitrite.

This is a weak base solution (Nitrite is the conjugate base of Nitrous acid). Solve accordingly.

	NO_2^{-1}	OH^{-}	HNO_2
<i>I</i>	0.1465	10^{-7}	0
<i>C</i>	$-x$	$+x$	$+x$
<i>E</i>	$0.14650 - x$	$10^{-7} + x$	x

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{4.5 \times 10^{-4}} = 2.2 \times 10^{-11} = \frac{[\text{OH}^{-}][\text{HNO}_2]}{[\text{NO}_2^{-1}]} = \frac{(10^{-7} + x)(x)}{(0.1465 - x)} \approx \frac{x^2}{0.1465}$$

$$x = \sqrt{(0.1465)(2.2 \times 10^{-11})} = 1.8 \times 10^{-6}$$

10^{-7} is about 5% of x so we can neglect it. x is also negligible compared to 0.1465.

$$\text{pOH} = -\log [\text{OH}^{-}] = -\log 1.8 \times 10^{-6} = 5.74$$

$$\text{pH} = 14.00 - \text{pOH} = 8.26$$