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.

$$n_{HNO_2} = 200.0 \text{ mL} \times \frac{0.1465 \text{ mol HNO}_2}{1000 \text{ mL HNO}_2} = 0.02930 \text{ mol}$$

$$n_{NO_2^{-1}} = 167.8 \text{ mL} \times \frac{0.1473 \text{ mol NO}_2^{-1}}{1000 \text{ mL NO}_2^{-1}} = 0.02472 \text{ mol}$$

$$pH = pK_a + \log\frac{n_{NO_2^{-1}}}{n_{HNO_2}} = 3.35 + \log\frac{0.02472}{0.02930} = 3.27$$

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).

$$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$$

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

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

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

$$0.02930 - 0.02473 = 4.58 \times 10^{-3} \text{mol H}^+$$

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

$$pH = -\log\left[H^+\right] = -\log 0.0125 = 1.90$$

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.

$$HNO_2 + OH^{-1} \rightarrow H_2O + NO_2^{-1}$$

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

$n_{HNO_2} = 0.02930$ mol

L.R. is OH^-.

$n'_{HNO_2} = 0.02930 - 0.02472 = 4.58\times10^{-3}$

$n'_{NO_2^{-1}} = 0.02472$

This is a buffer solution.

$$pH = pK_a + \log\frac{n_{NO_2^{-1}}}{n_{HNO_2}} = 3.35 + \log\frac{0.02472}{4.58\times10^{-3}} = 4.08$$

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	0.1465	10^{-7}	0
C	$-x$	$+x$	$+x$
E	$0.14650 - x$	$10^{-7} + x$	x

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{4.5\times10^{-4}} = 2.2\times10^{-11} = \frac{[OH^-][HNO_2]}{[NO_2^{-1}]} = \frac{(10^{-7}+x)(x)}{(0.1465-x)} \approx \frac{x^2}{0.1465}$$

$$x = \sqrt{(0.1465)(2.2\times10^{-11})} = 1.8\times10^{-6}$$

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

$$pOH = -\log[OH^-] = -\log 1.8\times10^{-6} = 5.74$$

$$pH = 14.00 - pOH = 8.26$$