

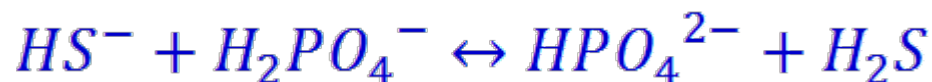
# Examples: Weak acid and weak base

20.0 mL of 0.30 M  $\text{NaH}_2\text{PO}_4$  was added to 20.0 mL of 0.30 M  $\text{NaHS}$ . What are the concentrations of all species at equilibrium? [The  $\text{pK}_a$  of  $\text{H}_2\text{PO}_4^-$  and  $\text{H}_2\text{S}$  are 7.2 and 6.9]

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Solution: Look up the  $K_a$  for each reaction involved in this acid-base equilibrium.



$$K = \frac{K_a(\text{H}_2\text{PO}_4^-)}{K_a(\text{H}_2\text{S})} = \frac{6.20 \times 10^{-8}}{1.26 \times 10^{-7}} = 0.49$$

$$K = \frac{[\text{HPO}_4^{2-}][\text{H}_2\text{S}]}{[\text{HS}^-][\text{H}_2\text{PO}_4^-]}$$

# Examples: Weak acid and weak base

Species	$HS^-$	$H_2PO_4^-$	$HPO_4^{2-}$	$H_2S$
Initial	0.15	0.15	0.0	0.0
Difference	-x	-x	x	x
Final	0.15-x	0.15-x	x	x

$$K = \frac{x^2}{(0.15 - x)^2}$$

$$K(0.15 - x)^2 - x^2 = 0$$

$$0.0225K - 0.3Kx - (1 - K)x^2 = 0$$

$$0.011 - 0.147x - 0.51x^2 = 0$$

$$x = \frac{0.147 \pm \sqrt{0.147^2 + 4(0.011)(0.51)}}{-2(0.51)}$$

$$x = 0.0636$$

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Initial	0.15	0.15	0.0	0.0
Difference	-x	-x	x	x
Final	0.15-x	0.15-x	x	x

Thus the concentrations are:

$$[HS^-] = [H_2PO_4^-] = 0.0864 M$$

$$[H_2S] = [HPO_4^{2-}] = 0.0636 M$$

Such mixtures form a double buffer. Both the  $[HS^-]/[H_2S]$  and the  $[HPO_4^{2-}]/[H_2PO_4^-]$  ratio can be used in the Henderson-Hasselbach equation to predict the pH. Regardless of which buffer we choose we will obtain the same answer for the pH.

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Difference	-x	-x	x	x
Final	0.15-x	0.15-x	x	x

Use the Hendersen-Hasselbach equation to predict the pH.

$$pH = 7.2 + \log_{10} \left( \frac{0.064}{0.086} \right) = 7.06$$

We can check using other  $pK_a$  for  $H_2S$ .

$$pH = 6.9 + \log_{10} \left( \frac{0.086}{0.064} \right) = 7.03$$

The small difference is due to rounding error.