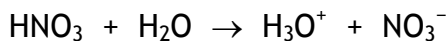


Module 5 Lesson 5 Exercises Answer Key

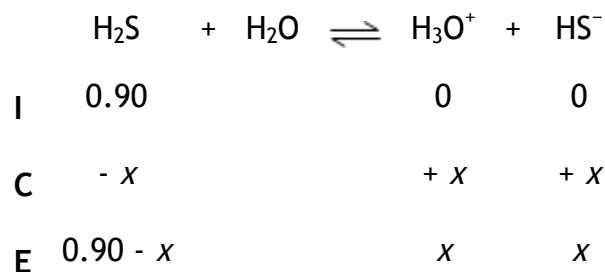
1. HNO₃ is a strong acid. It dissociates 100%.



$$[\text{H}_3\text{O}^+] = [\text{NO}_3^-] = 0.70 \text{ mol/L}$$

$$[\text{HNO}_3] = 0 \text{ mol/L}$$

- 2.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HS}^-]}{[\text{H}_2\text{S}]}$$

$$1.0 \times 10^{-7} = \frac{x^2}{0.90 - x} \quad K_a \text{ is very small assume } x \text{ to be negligible}$$

$$(1.0 \times 10^{-7})(0.90) = \left(\frac{x^2}{0.90}\right)(0.90)$$

$$\sqrt{9.0 \times 10^{-8}} = \sqrt{x^2}$$

$$3.0 \times 10^{-4} = x$$

$$[\text{H}_3\text{O}^+] = x = 3.0 \times 10^{-4} \text{ mol/L}$$

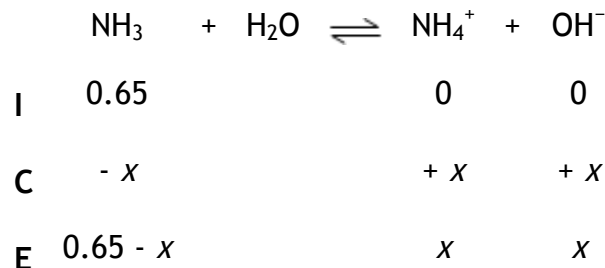
3. NaOH is a strong base, therefore complete dissociation.



$$[\text{Na}^+] = [\text{OH}^-] = 0.10 \text{ mol/L}$$

$$[\text{NaOH}] = 0 \text{ mol/L}$$

4. Set up an ICE table



Substitute into equilibrium law and solve for x.

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.0 \times 10^{-7} = \frac{x^2}{0.65 - x} \quad K_b \text{ is very small assume } x \text{ to be negligible}$$

$$(1.0 \times 10^{-7})(0.65) = \left(\frac{x^2}{0.65}\right)(0.65)$$

$$\sqrt{6.5 \times 10^{-8}} = \sqrt{x^2}$$

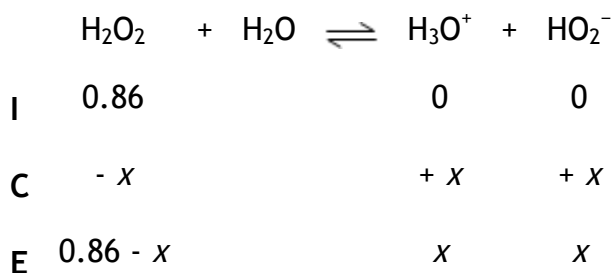
$$2.55 \times 10^{-4} = x$$

$$[\text{NH}_4^+] = [\text{OH}^-] = x = 2.6 \times 10^{-4} \text{ mol/L}$$

$$\begin{aligned} [\text{NH}_3] &= 0.65 - x \\ &= 0.65 - 2.6 \times 10^{-4} \text{ mol/L} \\ &= 0.6497 \text{ mol/L} \end{aligned}$$

$$[\text{NH}_3] = 0.65 \text{ mol/L (this rounds to 0.65 so our assumption was correct)}$$

5. Set up an ICE table



5. (con't)

Substitute into equilibrium law and solve for x.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HO}_2^-]}{[\text{H}_2\text{O}_2]}$$

$$2.4 \times 10^{-12} = \frac{x^2}{0.86 - x} \quad K_a \text{ is very small assume } x \text{ to be negligible}$$

$$(2.4 \times 10^{-12})(0.86) = \left(\frac{x^2}{0.86}\right)(0.86)$$

$$\sqrt{2.064 \times 10^{-12}} = \sqrt{x^2}$$

$$1.44 \times 10^{-6} = x$$

$$[\text{H}_3\text{O}^+] = x = 1.4 \times 10^{-6} \text{ mol/L}$$

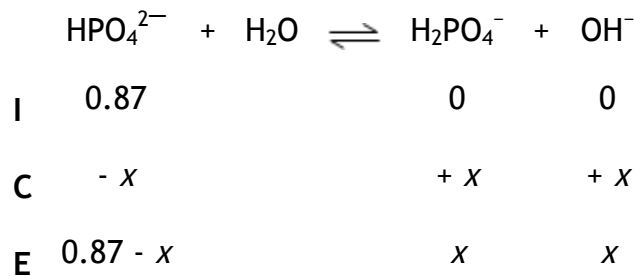
6. $\% \text{ dissociation} = \frac{[\text{ionized}]}{[\text{acid}]} \times 100\%$

$$= \frac{[\text{H}_3\text{O}^+]}{[\text{HX}]} \times 100\%$$

$$= \frac{4.5 \times 10^{-6}}{0.45} \times 100\%$$

$$= 1.0 \times 10^{-3}\%$$

7.



$$K_b = \frac{[\text{H}_2\text{PO}_4^-][\text{OH}^-]}{[\text{HPO}_4^{2-}]}$$

$$1.6 \times 10^{-7} = \frac{x^2}{0.87 - x} \quad K_b \text{ is very small assume } x \text{ to be negligible}$$

$$(1.6 \times 10^{-7})(0.87) = \left(\frac{x^2}{0.87}\right)(0.87)$$

$$\sqrt{1.392 \times 10^{-7}} = \sqrt{x^2}$$

$$3.73 \times 10^{-4} = x$$

7. (con't) $[\text{OH}^-] = 3.73 \times 10^{-4} \text{ mol/L}$

$$\begin{aligned}\% \text{ dissociation} &= \frac{[\text{OH}^-]}{[\text{HPO}_4^{2-}]} \times 100\% \\ &= \frac{3.73 \times 10^{-4}}{0.87} \times 100\% \\ &= \mathbf{0.043\% \text{ dissociated}}\end{aligned}$$

8. Rearrange the % dissociation equation to solve for $[\text{H}_3\text{O}^+]$

$$\begin{aligned}[\text{H}_3\text{O}^+] &= \frac{(\% \text{diss})[\text{acid}]}{100\%} \\ &= \frac{(0.12\%)(0.38 \text{ mol/L})}{100\%}\end{aligned}$$

$$[\text{H}_3\text{O}^+] = \mathbf{4.56 \times 10^{-4} \text{ mol/L}}$$

9. $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$

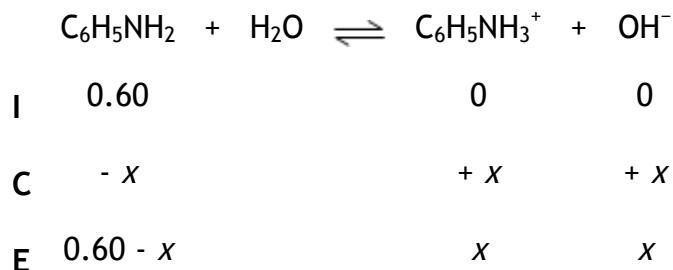
$$\begin{aligned}[\text{H}_3\text{O}^+] &= \frac{(\% \text{diss})[\text{HA}]}{100\%} \\ &= \frac{(0.025\%)(0.45 \text{ mol/L})}{100\%} \\ &= 0.01125 \text{ mol/L}\end{aligned}$$

$$[\text{H}_3\text{O}^+] = [\text{A}^-] = 0.01125 \text{ mol/L}$$

$$\begin{aligned}K_a &= \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \\ &= \frac{(0.01125)(0.01125)}{0.45}\end{aligned}$$

$$K_a = \mathbf{2.8 \times 10^{-4}}$$

10.



$$K_b = \frac{[\text{OH}^-][\text{C}_6\text{H}_5\text{NH}_3^+]}{[\text{C}_6\text{H}_5\text{NH}_2]}$$

$$3.8 \times 10^{-10} = \frac{x^2}{0.60 - x} \quad K_b \text{ is very small assume } x \text{ to be negligible}$$

$$(3.8 \times 10^{-10})(0.60) = \left(\frac{x^2}{0.60} \right)(0.60)$$

$$\sqrt{2.28 \times 10^{-10}} = \sqrt{x^2}$$

$$1.51 \times 10^{-5} = x$$

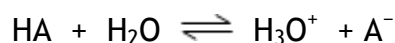
$$[\text{OH}^-] = x = 1.51 \times 10^{-5} \text{ mol/L}$$

$$\% \text{ dissociation} = \frac{[\text{OH}^-]}{[\text{C}_6\text{H}_5\text{NH}_2]} \times 100\%$$

$$= \frac{1.51 \times 10^{-5}}{0.60} \times 100\%$$

$$= \mathbf{2.5 \times 10^{-3} \% \text{ dissociated}}$$

11.



You must find the $[\text{H}_3\text{O}^+]$ and $[\text{A}^-]$ first:

$$[\text{H}_3\text{O}^+] = \frac{(\% \text{diss})[\text{HA}]}{100\%}$$

$$= \frac{(0.015\%)(0.750 \text{ mol/L})}{100\%}$$

$$= 1.125 \times 10^{-4} \text{ mol/L}$$

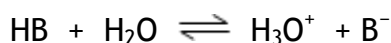
11. (con't)

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

$$= \frac{(1.125 \times 10^{-4})(1.125 \times 10^{-4})}{0.750}$$

$$K_a = 1.7 \times 10^{-8}$$

12.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{B}^-]}{[\text{HB}]}$$

$$= \frac{(4.5 \times 10^{-10})(4.5 \times 10^{-10})}{0.80}$$

$$K_a = 2.5 \times 10^{-19}$$

13. To calculate pH, we need to calculate $[\text{H}_3\text{O}^+]$. We must set up an "ICE" table

	HOCl	+	H ₂ O	\rightleftharpoons	H ₃ O ⁺	+	OCl ⁻
I	0.10				0		0
C	-x				+x		+x
E	0.10 - x				x		x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OCl}^-]}{[\text{HOCl}]}$$

$$3.5 \times 10^{-8} = \frac{x^2}{0.10 - x} \quad K_a \text{ is very small, assume } x \text{ is negligible}$$

$$3.5 \times 10^{-8} \approx \frac{x^2}{0.10}$$

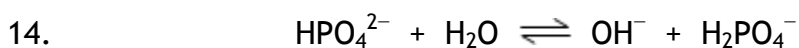
$$(3.5 \times 10^{-8})(0.10) = \left(\frac{x^2}{0.10}\right)(0.10)$$

$$\sqrt{3.5 \times 10^{-9}} = x$$

$$5.92 \times 10^{-5} = x$$

$$[\text{H}_3\text{O}^+] = x = 5.92 \times 10^{-5} \text{ mol/L}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(5.92 \times 10^{-5}) = 4.23$$



$$\text{pOH} = 14.00 - 9.00 = 5.00$$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.00} = 1.0 \times 10^{-5} \text{ mol/L}$$

$$K_b = \frac{[\text{H}_2\text{PO}_4^-][\text{OH}^-]}{[\text{HPO}_4^{2-}]} = \frac{(1.0 \times 10^{-5})^2}{0.20}$$

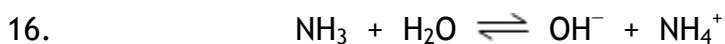
$$K_b = \mathbf{5.0 \times 10^{-10}}$$



$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.40} = 3.98 \times 10^{-5} \text{ mol/L}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{X}^-]}{[\text{HX}]} = \frac{(3.98 \times 10^{-5})^2}{0.25}$$

$$K_a = \mathbf{6.3 \times 10^{-9}}$$



$$[\text{OH}^-] = \frac{(\% \text{ dissociation})[\text{NH}_3]}{100\%} = \frac{(4.3\%)(0.010 \text{ mol/L})}{100\%} = 4.3 \times 10^{-4} \text{ mol/L OH}^-$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(4.3 \times 10^{-4}) = \mathbf{3.37}$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.37 = \mathbf{10.63}$$

17. We need to set up an "ICE" table to find $[\text{OH}^-]$. We can use $[\text{OH}^-]$ to find pH.

	H_2NNH_2	+	H_2O	\rightleftharpoons	OH^-	+	H_2NNH_3^+
I	2.0				0		0
C	- x				+ x		+ x
E	2.0 - x				x		x

17. (con't)

$$K_b = \frac{[\text{H}_2\text{NNH}_3^+][\text{OH}^-]}{[\text{H}_2\text{NNH}_2]}$$

$$3.0 \times 10^{-6} = \frac{x^2}{2.0 - x} \quad K_b \text{ is very small, assume } x \text{ is negligible}$$

$$3.0 \times 10^{-6} \approx \frac{x^2}{2.0}$$

$$(3.0 \times 10^{-6})(2.0) = \left(\frac{x^2}{2.0}\right)(2.0)$$

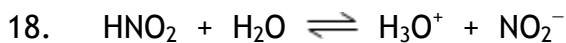
$$\sqrt{6.0 \times 10^{-6}} = x$$

$$2.45 \times 10^{-3} = x$$

$$[\text{OH}^-] = 2.45 \times 10^{-3} \text{ mol/L}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(2.45 \times 10^{-3}) = 2.61$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2.61 = \mathbf{11.39}$$



If we find the hydronium ion concentration, we can calculate percent dissociation:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.20} = 6.3 \times 10^{-5} \text{ mol/L}$$

$$\% \text{ dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{HNO}_2]} \times 100\% = \frac{6.3 \times 10^{-5}}{0.20} \times 100\%$$

$$\% \text{ dissociation} = \mathbf{0.0315 \%}$$