Module 5 Lesson 5 Exercises Answer Key

1. HNO_3 is a strong acid. It dissociates 100%.

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$

$[H_3O^+] = [NO_3^-] = 0.70$ mol/L
$[HNO_3] = 0$ mol/L

2.

	H_2S	+ H_2O	$\rightleftharpoons$	H_3O^+	+ HS^-
I	0.90			0	0
C	$-x$			$+x$	$+x$
E	$0.90 - x$			x	x

$$K_a = \frac{[H_3O^+][HS^-]}{[H_2S]}$$

$$1.0\times10^{-7} = \frac{x^2}{0.90-x} \quad K_a \text{ is very small assume x to be negligible}$$

$$\left(1.0\times10^{-7}\right)(0.90) = \left(\frac{x^2}{0.90}\right)(0.90)$$

$$\sqrt{9.0\times10^{-8}} = \sqrt{x^2}$$

$$3.0\times10^{-4} = x$$

$[H_3O^+] = x = 3.0 \times 10^{-4}$ mol/L

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

$$NaOH \rightarrow Na^+ + OH^-$$

$[Na^+] = [OH^-] = 0.10$ mol/L
$[NaOH] = 0$ mol/L

4. Set up an ICE table

	NH_3	+	H_2O	$\rightleftharpoons$	NH_4^+	+	OH^-
I	0.65				0		0
C	- x				+ x		+ x
E	0.65 - x				x		x

Substitute into equilibrium law and solve for x.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

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

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

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

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

$[NH_4^+] = [OH^-] = x = 2.6 \times 10^{-4}$ mol/L

$[NH_3]$ = 0.65 - x
= 0.65 - 2.6 x 10^{-4} mol/L
= 0.6497 mol/L

$[NH_3]$ = 0.65 mol/L(this rounds to 0.65 so our assumption was correct)

5. Set up an ICE table

	H_2O_2	+	H_2O	$\rightleftharpoons$	H_3O^+	+	HO_2^-
I	0.86				0		0
C	- x				+ x		+ x
E	0.86 - x				x		x

5. (con't)

Substitute into equilibrium law and solve for x.

$$K_a = \frac{[H_3O^+][HO_2^-]}{[H_2O_2]}$$

$$2.4\times10^{-12} = \frac{x^2}{0.86 - x} \quad K_a \text{ is very small assume x to be negligible}$$

$$\left(2.4\times10^{-12}\right)(0.86) = \left(\frac{x^2}{0.86}\right)(0.86)$$

$$\sqrt{2.064\times10^{-12}} = \sqrt{x^2}$$

$$1.44\times10^{-6} = x$$

$[H_3O^+]$ = x = 1.4 x 10^{-6} mol/L

6.

$$\% \text{ dissociation} = \frac{[\text{ionized}]}{[\text{acid}]}\times100\%$$

$$= \frac{[H_3O^+]}{[HX]}\times100\%$$

$$= \frac{4.5\times10^{-6}}{0.45}\times100\%$$

$$= \mathbf{1.0\times10^{-3}\%}$$

7.

	HPO_4^{2-}	+ H_2O	$\rightleftharpoons$	$H_2PO_4^-$	+ OH^-
I	0.87			0	0
C	- x			+ x	+ x
E	0.87 - x			x	x

$$K_b = \frac{[H_2PO_4^-][OH^-]}{[HPO_4^{2-}]}$$

$$1.6\times10^{-7} = \frac{x^2}{0.87 - x} \quad K_b \text{ is very small assume x to be negligible}$$

$$\left(1.6\times10^{-7}\right)(0.87) = \left(\frac{x^2}{0.87}\right)(0.87)$$

$$\sqrt{1.392\times10^{-7}} = \sqrt{x^2}$$

$$3.73\times10^{-4} = x$$

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

$$\% \text{ dissociation} = \frac{[OH^-]}{[HPO_4^{2-}]} \times 100\%$$
$$= \frac{3.73 \times 10^{-4}}{0.87} \times 100\%$$
$$\mathbf{= 0.043\% \text{ dissociated}}$$

8. Rearrange the % dissociation equation to solve for $[H_3O^+]$

$$[H_3O^+] = \frac{(\%\text{diss})[\text{acid}]}{100\%}$$
$$= \frac{(0.12\%)(0.38 \text{mol/L})}{100\%}$$
$$\mathbf{[H_3O^+] = 4.56 \times 10^{-4} \text{ mol/L}}$$

9. $HA + H_2O \rightleftharpoons H_3O^+ + A^-$

$$[H_3O^+] = \frac{(\%\text{diss})[HA]}{100\%}$$
$$= \frac{(0.025\%)(0.45 \text{ mol/L})}{100\%}$$
$$= 0.01125 \text{ mol/L}$$

$[H_3O^+] = [A^-] = 0.01125$ mol/L

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$
$$= \frac{(0.01125)(0.01125)}{0.45}$$
$$\mathbf{K_a = 2.8 \times 10^{-4}}$$

10.

	$C_6H_5NH_2$	+ H_2O $\rightleftharpoons$	$C_6H_5NH_3^+$	+ OH^-
I	0.60		0	0
C	- x		+ x	+ x
E	0.60 - x		x	x

$$K_b = \frac{[OH^-][C_6H_5NH_3^+]}{[C_6H_5NH_2]}$$

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

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

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

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

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

$$\% \text{ dissociation} = \frac{[OH^-]}{[C_6H_5NH_2]} \times 100\%$$

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

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

11. $HA + H_2O \rightleftharpoons H_3O^+ + A^-$

You must find the $[H_3O^+]$ and $[A^-]$ first:

$$[H_3O^+] = \frac{(\%\text{diss})[HA]}{100\%}$$

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

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

11. (con't)

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$= \frac{\left(1.125\times10^{-4}\right)\left(1.125\times10^{-4}\right)}{0.750}$$

$$\mathbf{K_a = 1.7\times10^{-8}}$$

12. $HB + H_2O \rightleftharpoons H_3O^+ + B^-$

$$K_a = \frac{[H_3O^+][B^-]}{[HB]}$$

$$= \frac{\left(4.5\times10^{-10}\right)\left(4.5\times10^{-10}\right)}{0.80}$$

$$\mathbf{K_a = 2.5\times10^{-19}}$$

13. To calculate pH, we need to calculate $[H_3O^+]$. We must set up an "ICE" table

	HOCl	+	H_2O	$\rightleftharpoons$	H_3O^+	+	OCl^-
I	0.10				0		0
C	- x				+ x		+ x
E	0.10 - x				x		x

$$K_a = \frac{[H_3O^+][OCl^-]}{[HOCl]}$$

$$3.5\times10^{-8} = \frac{x^2}{0.10-x}$$ K_a is very small, assume x is negligible

$$3.5\times10^{-8} \approx \frac{x^2}{0.10}$$

$$\left(3.5\times10^{-8}\right)(0.10) = \left(\frac{x^2}{0.10}\right)(0.10)$$

$$\sqrt{3.5\times10^{-9}} = x$$

$$5.92\times10^{-5} = x$$

$[H_3O^+] = x = 5.92 \times 10^{-5}$ mol/L

pH = $-\log[H_3O^+] = -\log(5.92 \times 10^{-5})$ = **4.23**

14. $HPO_4^{2-} + H_2O \rightleftharpoons OH^- + H_2PO_4^-$

pOH = 14.00 – 9.00 = 5.00

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

$$K_b = \frac{[H_2PO_4^-][OH^-]}{[HPO_4^{2-}]} = \frac{(1.0\times10^{-5})^2}{0.20}$$

K_b = **5.0×10^{-10}**

15. $HX + H_2O \rightleftharpoons H_3O^+ + X^-$

$[H_3O^+] = 10^{-pH} = 10^{-4.40} = 3.98 \times 10^{-5}$ mol/L

$$K_a = \frac{[H_3O^+][X^-]}{[HX]} = \frac{(3.98\times10^{-5})^2}{0.25}$$

K_a = **6.3×10^{-9}**

16. $NH_3 + H_2O \rightleftharpoons OH^- + NH_4^+$

$$[OH^-] = \frac{(\%\text{ dissociation})[NH_3]}{100\%} = \frac{(4.3\%)(0.010\text{ mol/L})}{100\%} = 4.3\times10^{-4}\text{ mol/L OH}^-$$

pOH = -log[OH⁻] = -log(4.3 x10^{-4}) = **3.37**

pH = 14 – pOH = 14 – 3.37 = **10.63**

17. We need to set up an "ICE" table to find $[OH^-]$. We can use $[OH^-]$ 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)

$$K_b = \frac{[H_2NNH_3^+][OH^-]}{[H_2NNH_2]}$$

$$3.0\times10^{-6} = \frac{x^2}{2.0 - x}$$ K_b is very small, assume x is negligible

$$3.0\times10^{-6} \approx \frac{x^2}{2.0}$$

$$\left(3.0\times10^{-6}\right)(2.0) = \left(\frac{x^2}{2.0}\right)(2.0)$$

$$\sqrt{6.0\times10^{-6}} = x$$

$$2.45\times10^{-3} = x$$

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

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

pH = 14 – pOH = 14 – 2.61 = **11.39**

18. $HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$

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

$[H_3O^+] = 10^{-pH} = 10^{-4.20} = 6.3 \times 10^{-5}$ mol/L

$$\%\ \text{dissociation} = \frac{[H_3O^+]}{[HNO_2]}\times100\% = \frac{6.3\times10^{-5}}{0.20}\times100\%$$

% dissociation = 0.0315 %