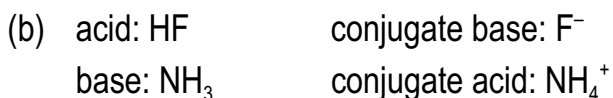
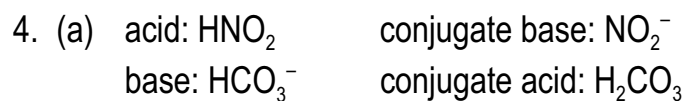
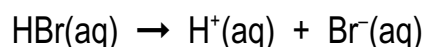


Equilibrium Practice 3: Acids and Bases

ANSWERS



7. (a) strong acid



$$[\text{H}^+] = 0.050 \text{ mol/L}$$

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

$$= -\log (0.050)$$

$$= 1.3010\dots$$

Therefore, the pH of the solution is 1.30

(b) weak base

	$\text{C}_6\text{H}_5\text{NH}_2(\text{aq})$	+	$\text{H}_2\text{O}(\text{l})$	$\rightleftharpoons$	$\text{C}_6\text{H}_5\text{NH}_3^+(\text{aq})$	+	$\text{OH}^-(\text{aq})$
I	0.050		—		0		~0
C	-x		—		+x		+x
E	0.050-x		—		x		x

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

$$4.1 \times 10^{-10} = \frac{(x)(x)}{0.050-x}$$

$$4.1 \times 10^{-10} = \frac{x^2}{0.050-x}$$

$$4.1 \times 10^{-10} = \frac{x^2}{0.050} \quad \text{small } K_a ; \text{ assume } 0.050-x = 0.050$$

$$2.05 \times 10^{-11} = x^2 \quad \{\text{multiply by } 0.050\}$$

$$4.5276 \dots \times 10^{-6} = x \quad \{\text{square root}\}$$

$$[\text{OH}^-]_{\text{eq}} = x \text{ mol/L}$$

$$= 4.5276 \times 10^{-6} \text{ mol/L}$$

$$K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$$

$$[\text{H}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]}$$

$$= \frac{1.0 \times 10^{-14}}{4.5276 \times 10^{-6}}$$

$$= 2.2086 \times 10^{-9} \text{ mol/L}$$

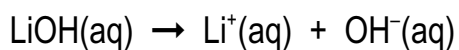
$$\text{pH} = -\log [\text{H}^+]$$

$$= -\log (2.2086 \times 10^{-9})$$

$$= 8.6558 \dots$$

Therefore the pH of the solution is 8.66.

(c) strong base



$$[\text{OH}^-] = 0.050 \text{ mol/L}$$

$$[\text{H}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]}$$

$$= \frac{1.0 \times 10^{-14}}{0.050}$$

$$= 2 \times 10^{-13} \text{ mol/L}$$

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

$$= -\log (2 \times 10^{-13})$$

$$= 12.6989 \dots$$

Therefore the pH of the solution is 12.70.

(d) weak acid

	$\text{C}_6\text{H}_5\text{COOH(aq)}$	$\rightleftharpoons$	$\text{C}_6\text{H}_5\text{COO}^-$	+	$\text{H}^+(\text{aq})$
I	0.050		0		~0
C	-x		+x		+x
E	0.050-x		x		x

$$K_a = \frac{[\text{C}_6\text{H}_5\text{COO}^-(\text{aq})][\text{H}^+(\text{aq})]}{[\text{C}_6\text{H}_5\text{COOH(aq)}]}$$

$$6.3 \times 10^{-5} = \frac{(x)(x)}{0.050-x}$$

$$6.3 \times 10^{-5} = \frac{x^2}{0.050-x}$$

$$6.3 \times 10^{-5} = \frac{x^2}{0.050} \quad (\text{small } K_a ; \text{ assume } 0.050-x = 0.050)$$

$$3.15 \times 10^{-6} = x^2 \quad \{\text{multiply by } 0.050\}$$

$$1.7748 \times 10^{-3} = x \quad \{\text{square root}\}$$

$$\begin{aligned} [\text{H}^+]_{\text{eq}} &= x \text{ mol/L} \\ &= 1.7748 \times 10^{-3} \text{ mol/L} \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log (1.7748 \times 10^{-3}) \\ &= 2.7508 \dots \end{aligned}$$

Therefore, the pH of the solution is 2.75.

(e) weak base

	$\text{NH}_3(\text{aq})$	+	$\text{H}_2\text{O}(\text{l})$	$\rightleftharpoons$	$\text{NH}_4^+(\text{aq})$	+	$\text{OH}^-(\text{aq})$
I	0.050		—		0		~0
C	-x		—		+x		+x
E	0.050-x		—		x		x

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

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.050-x}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.050-x}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.050} \quad (\text{small } K_b ; \text{ assume } 0.050-x = 0.050)$$

$$9 \times 10^{-7} = x^2 \quad \{\text{multiply by } 0.050\}$$

$$9.4868 \times 10^{-4} = x \quad \{\text{square root}\}$$

$$\begin{aligned} [\text{OH}^-]_{\text{eq}} &= x \text{ mol/L} \\ &= 9.4868 \times 10^{-4} \text{ mol/L} \end{aligned}$$

$$K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$$

$$[\text{H}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]}$$

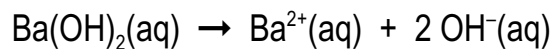
$$= \frac{1.0 \times 10^{-14}}{9.4868 \times 10^{-4}}$$

$$= 1.05409 \times 10^{-11} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log (1.05409 \times 10^{-11}) \\ &= 10.9771 \dots \end{aligned}$$

Therefore the pH of the solution is 10.98.

(f) strong base



$$[\text{OH}^{-}] = 2(0.050 \text{ mol/L}) = 0.10 \text{ mol/L}$$

$$\begin{aligned} [\text{H}^{+}(\text{aq})] &= \frac{K_w}{[\text{OH}^{-}(\text{aq})]} \\ &= \frac{1.0 \times 10^{-14}}{0.10} \\ &= 1 \times 10^{-13} \text{ mol/L} \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^{+}] \\ &= -\log (1 \times 10^{-13}) \\ &= 13 \end{aligned}$$

Therefore the pH of the solution is 13.00.

(g) weak acid

	$\text{HCN}(\text{aq})$	$\rightleftharpoons$	H^{+}	+	$\text{CN}^{-}(\text{aq})$
I	0.050		~ 0		0
C	$-x$		$+x$		$+x$
E	$0.050-x$		x		x

$$K_a = \frac{[\text{H}^{+}(\text{aq})][\text{CN}^{-}(\text{aq})]}{[\text{HCN}(\text{aq})]}$$

$$6.2 \times 10^{-10} = \frac{(x)(x)}{0.050-x}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.050-x}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.050} \quad (\text{small } K_a ; \text{ assume } 0.050-x = 0.050)$$

$$3.1 \times 10^{-11} = x^2 \quad \{\text{multiply by } 0.050\}$$

$$5.5677 \dots \times 10^{-6} = x \quad \{\text{square root}\}$$

$$\begin{aligned}[\text{H}^+]_{\text{eq}} &= x \text{ mol/L} \\ &= 5.5677 \times 10^{-6} \text{ mol/L}\end{aligned}$$

$$\begin{aligned}\text{pH} &= -\log [\text{H}^+] \\ &= -\log (5.5677 \times 10^{-6}) \\ &= 5.2543 \dots\end{aligned}$$

Therefore, the pH of the solution is 5.25.