Calculating pH Answer Key

1. **(a)** pH = –log [H_3O^+(aq)]
pH = –log(5.32 × 10^{-7} mol/L)
pH = 6.274

(b) pH = –log [H_3O^+(aq)]
pH = –log(6.1 × 10^{-5} mol/L)
pH = 4.21

(c) pH = –log [H_3O^+(aq)]
pH = –log(2.679 × 10^{-14} mol/L)
pH = 13.5720

(d) pH = –log [H_3O^+(aq)]
pH = –log(0.23 mol/L)
pH = 0.64

2. Solutions (a), (b), and (d) would be considered acidic.

3.

[H_3O^+(aq)] (mol/L)	**pH**
1.37 × 10^{-2}	1.863
2.38 × 10^{-5}	4.623
2.38 × 10^{-6}	5.623
1.00 × 10^{-7}	7.000
3.45 × 10^{-9}	8.462
3.45 × 10^{-10}	9.462
3.45 × 10^{-11}	10.462
5.33 × 10^{-12}	11.273

4. As the concentration of H_3O^+(aq) decreases, pH values increase.

5. As acidity decreases, pH increases.

6. As basicity decreases, pH decreases.

7. For every pH increase of one unit, the concentration of H_3O^+(aq) decreases by a factor of ten.

8. For every pH increase of two units, the concentration of H_3O^+(aq) decreases by a factor of 10^2, or 100.

9. You would expect the pH to increase by three units if an acidic solution is diluted by a factor of 10^3.

10. **(a)** pH = –log [H_3O^+(aq)]
[H_3O^+(aq)] = 0.563 mol/L
pH = 0.249

(b) pH = –log[H_3O^+(aq)]
[H_3O^+(aq)] = 2.3 × 10^{-4} mol/L
pH = 3.64

(c) pH = $-\log[H_3O^+(aq)]$
$[H_3O^+(aq)] = 9.342 \times 10^{-5}$ mol/L
pH = 4.0296

11. (a) $M_{HClO_4} = 1.01\frac{g}{mol} + 35.45\frac{g}{mol} + 4\left(16.00\frac{g}{mol}\right) = 100.46\frac{g}{mol}$

$n = \frac{m}{M}$

$n_{HClO_4} = \frac{3.62\,g}{100.46\frac{g}{mol}}$

$n_{HClO_4} = 0.036034\,mol$

$c = \frac{n}{V}$

$c_{HClO_4} = \frac{0.036034\,mol}{2.0\,L}$

$c_{HClO_4} = 0.01802\frac{mol}{L}$

Ratio of $\frac{H_3O^+}{HClO_4} = \frac{1}{1}$

$c_{H_3O^+} = 0.018\frac{mol}{L}$

$pH = -\log[H_3O^+]$

$pH = -\log\left(0.01802\frac{mol}{L}\right)$

$pH = 1.74$

(b) $M_{HBr} = 1.01\frac{g}{mol} + 79.90\frac{g}{mol} = 80.91\frac{g}{mol}$

$n = \frac{m}{M}$

$n_{HBr} = \frac{2.357\,g}{80.91\frac{g}{mol}}$

$n_{HBr} = 0.029131\,mol$

$c = \frac{n}{V}$

$c_{HBr} = \frac{0.029131\,mol}{50.0\,L}$

$c_{HBr} = 5.83 \times 10^{-4}\frac{mol}{L}$

Ratio of $\frac{H_3O^+}{HBr} = \frac{1}{1}$

$c_{H_3O^+} = 5.83 \times 10^{-4}\frac{mol}{L}$

$pH = -\log[H_3O^+]$

$pH = -\log\left(5.83 \times 10^{-4}\frac{mol}{L}\right)$

$pH = 3.234$

Calculating pH Answer Key

(cont'd)

(c) $M_{HNO_3} = 1.01\frac{g}{mol} + 14.01\frac{g}{mol} + 3\left(16.00\frac{g}{mol}\right) = 63.02\frac{g}{mol}$

$$n = \frac{m}{M}$$

$$n_{HNO_3} = \frac{8\times10^{-6}\,g}{63.02\frac{g}{mol}}$$

$$n_{HNO_3} = 1.269\times10^{-7}\,mol$$

$$c = \frac{n}{V}$$

$$c_{HNO_3} = \frac{1.269\times10^{-7}\,mol}{0.0200\,L}$$

$$c_{HNO_3} = 6.345\times10^{-6}\frac{mol}{L}$$

Ratio of $\frac{H_3O^+}{HNO_3} = \frac{1}{1}$

$$c_{H_3O^+} = 6.345\times10^{-6}\frac{mol}{L}$$

$$pH = -\log[H_3O^+]$$

$$pH = -\log\left(6.345\times10^{-6}\frac{mol}{L}\right)$$

$$pH = 5.2$$