

WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NAME _____

Panther ID _____

For problems involving calculations you must show your work for credit.

1) Find the pH for a 0.200 M aqueous solution of phenol (C₆H₅OH). The value for K_a for phenol at this temperature is K_a = 1.3 x 10⁻¹⁰. Also find the percent dissociation of phenol for these conditions.

The reaction is C₆H₅OH(aq) + H₂O(l) ⇌ H₃O⁺(aq) + C₆H₅O⁻(aq)

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_6\text{H}_5\text{O}^-]}{[\text{C}_6\text{H}_5\text{OH}]} = 1.3 \times 10^{-10}$$

	Initial	Change	Equilibrium
H ₃ O ⁺	0	x	x
C ₆ H ₅ O ⁻	0	x	x
C ₆ H ₅ OH	0.200	- x	0.200 - x

So $\frac{(x)(x)}{(0.200 - x)} = 1.3 \times 10^{-10}$

If we assume $x \ll 0.200$, then we get $\frac{x^2}{0.200} = 1.3 \times 10^{-10}$

$$x^2 = (1.3 \times 10^{-10})(0.200) = 2.6 \times 10^{-11}$$

$$x = (2.6 \times 10^{-11})^{1/2} = 5.1 \times 10^{-6} \quad \text{Since } x \text{ is small compared to } 0.200, \text{ our assumption was good}$$

$$[\text{H}_3\text{O}^+] = 5.1 \times 10^{-6} \text{ M} \quad \text{pH} = -\log_{10}(5.1 \times 10^{-6}) = 5.29$$

$$\text{percent dissociation} = \frac{[\text{C}_6\text{H}_5\text{O}^-]_{\text{eq}}}{[\text{C}_6\text{H}_5\text{OH}]_{\text{initial}}} \times 100\% = \frac{5.1 \times 10^{-6}}{0.200} \times 100\% = 2.6 \times 10^{-3}\%$$

2) A 0.100 M aqueous solution of which of the following ionic compounds would be expected to have a pH much less than 7.0?

- a) NaNO₃
- b) KF
- c) KOH
- d) Both a and b
- e) None of the above

_____E_____

NaNO₃ is the salt of a strong acid and a strong base, and so when added to water will give pH = 7.0. KF is the salt of a weak acid and a strong base, and so it contains a weak base (F⁻). When added to water it will give pH > 7.0. KOH is a strong soluble base, and so it will also give a pH > 7.0. So the correct answer is E. None of the solutions will give pH < 7.0

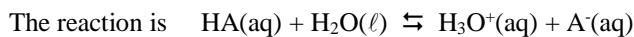
3) The pH of a 0.0400 M aqueous solution of a weak monoprotic acid HA, measured at T = 25.0 °C, is pH = 3.71.

a) Find the value for $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ present in the above solution.

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 3.71 = 10.29$$

$$[\text{H}_3\text{O}^+] = 10^{-3.71} = 1.95 \times 10^{-4} \text{ M} \quad [\text{OH}^-] = 10^{-10.29} = 5.1 \times 10^{-11} \text{ M}$$

b) Find the value for K_a for the above acid, at T = 25.0 °C.



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

If we set up an ICE table, we get

	Initial	Change	Equilibrium
H_3O^+	0	x	x
A^-	0	x	x
HA	0.0400	- x	0.0400 - x

But $x = [\text{H}_3\text{O}^+] = 1.95 \times 10^{-4} \text{ M}$, so $[\text{A}^-] = x = 1.95 \times 10^{-4} \text{ M}$
 $[\text{HA}] = 0.0400 - x = 0.0400 - 1.95 \times 10^{-4} = 3.98 \times 10^{-2} \text{ M}$

So $K_a = \frac{(1.95 \times 10^{-4})(1.95 \times 10^{-4})}{(3.98 \times 10^{-2})} = 9.6 \times 10^{-7}$

4) The value for the acid ionization constant for hypochlorous acid (HOCl) is $K_a = 3.5 \times 10^{-8}$ at T = 25. °C. What is the value for K_b for the hypochlorite ion (OCl^-) at this temperature?

For an acid/conjugate base pair we know (at T = 25. °C) that $K_a K_b = 1.0 \times 10^{-14}$

$$\text{So } K_b = \frac{1.0 \times 10^{-14}}{K_a} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-8}} = 2.9 \times 10^{-7}$$