

Acid #9 **Answer Key**

- $$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$$

Since the reaction is endothermic, then, lowering the temperature from 25°C to 5.0°C shifts the equilibrium to favour reactants. This shift lowers the $[\text{H}_3\text{O}^+]$ subsequently raising the pH.
- Colour changes occur when $[\text{H}^+]$ in the solution equals the K_a of the indicator. Since different indicators have different K_a values, then they must change colour at different pH.
- Moles diprotic acid = $2.68\text{g} / 126\text{g mol}^{-1} = 0.02127 \text{ mol}$
 $[\text{acid}] = \text{mol} / \text{vol} = 0.02127 / 0.250 = 0.08508 \text{ M}$
 Moles acid used = $(0.08508\text{M})(25.0\text{ml}) = 2.127\text{mmol}$
 Moles H^+ used = $(2.127\text{mmol}) \times 2 = 4.254 \text{ mmol}$
 At the stoichiometric point, $\text{mol H}^+ = \text{mol OH}^-$, then: moles $\text{OH}^- = 4.254 \text{ mmol}$
 $[\text{NaOH}] = 4.254\text{mmol} / 19.32\text{ml} = 0.220\text{M}$
- The K_b need to be compared:

NO_2^-	$K_b = 2.0 \times 10^{-11}$ (2.17×10^{-11})
$\text{C}_6\text{H}_5\text{COO}^-$	$K_b = 1.5 \times 10^{-10}$ (1.54×10^{-10})

Since K_b for $\text{C}_6\text{H}_5\text{COO}^-$ is larger, then $\text{C}_6\text{H}_5\text{COO}^-$ is the stronger base.
- $\text{Al}(\text{NO}_3)_3 + \text{H}_2\text{O} \rightarrow \text{Al}(\text{H}_2\text{O})_6^{3+} + \text{NO}_3^-$
 The NO_3^- do not hydrolyze
 The hydrated aluminum ions do hydrolyze and produce H_3O^+ which give the $\text{Al}(\text{NO}_3)_3$ solution its acid property.

$$\text{Al}(\text{H}_2\text{O})_6^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+} + \text{H}_3\text{O}^+$$
- Let the acid be HA.

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

Then: $[\text{H}_3\text{O}^+] = [\text{A}^-] = (20\%)(0.100) = 0.0200 \text{ M}$
 $[\text{HA}] = 0.100 - 0.0200 = 0.0800\text{M}$
 $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$
 $K_a = (0.0200)^2 / 0.0800$
 $K_a = 5.00 \times 10^{-3}$
- $\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$
 Let $X = [\text{H}_3\text{O}^+] = [\text{HS}^-]$
 Then $[\text{H}_2\text{S}] = 0.0500 - X \sim 0.0500$ (if X is small)
 $K_a = [\text{H}_3\text{O}^+][\text{HS}^-] / [\text{H}_2\text{S}]$
 $(9.1 \times 10^{-8})(1.0 \times 10^{-7}) = X^2 / 0.0500$
 $[\text{H}_3\text{O}^+] = X = 7.07 \times 10^{-5}$ (6.75×10^{-5})
 $\% \text{dissociation} = [\text{H}_3\text{O}^+] / \text{initial}[\text{H}_2\text{S}] \times 100$
 $\% = 7.07 \times 10^{-5} / 0.0500 \times 100 = 0.14\% \quad 0.13\%$

8. Compare the K_a with the K_b for the ion:

$$K_a \text{ for } \text{HPO}_4^{2-} = 4.4 \times 10^{-13} \quad (2.2 \times 10^{-13})$$

$$K_b \text{ for } \text{HPO}_4^{2-} = 1.6 \times 10^{-7}$$

Since the K_b is far greater than the K_a , the ion and its solution is basic.

9. $[\text{Acid}] = \text{moles} / \text{vol} = ((100.0\text{g}) / 120\text{gmol}^{-1}) / 0.600 = 1.389 \text{ M}$

$$[\text{H}^+] = 2 \times [\text{Acid}] = 2 \times 1.389 = 2.778 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -4.437 = -0.444$$

10. (a) In acid solution, the acid form of an indicator predominates; H₂Sg is present in larger concentration.

(b) Because the acid form is green, the indicator would colour the acid solution green.

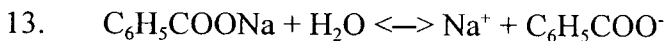
11. Moles $\text{OH}^- = (50.0)(0.20) = 10.0\text{mmol}$

at the stoichiometric point, mol $\text{H}^+ = \text{mol } \text{OH}^-$, then moles $\text{H}^+ = 10.0 \text{ mmol}$

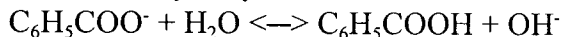
moles oxalic acid = $10.0 \text{ mmol} \times 1/2 = 5.00 \text{ mmol acid}$

$$\text{mass of acid} = (5.00 \text{ mmol})(126\text{g/mol}) = 0.63 \text{ g}$$

12. $K_b = K_w / K_a = 1.00 \times 10^{-14} / 6.6 \times 10^{-5} = 1.5 \times 10^{-10}$



Na^+ do not hydrolyze

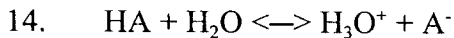


Let $x = [\text{OH}^-] = [\text{C}_6\text{H}_5\text{COOH}]$, then $[\text{C}_6\text{H}_5\text{COO}^-] = 0.10 - x \sim 0.10\text{M}$

$$K_b = [\text{C}_6\text{H}_5\text{COOH}][\text{OH}^-] / [\text{C}_6\text{H}_5\text{COO}^-]$$

$$1.5 \times 10^{-10} = X^2 / 0.10 \quad X = [\text{OH}^-] = 3.87 \times 10^{-6} \text{ M}$$

$$\text{pH} = 8.59$$



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

$$[\text{H}_3\text{O}^+] = [\text{A}^-] = 3.0 \times 10^{-3} \text{ M}$$

$$[\text{HA}] = 0.0100 - 3.0 \times 10^{-3} = 7.00 \times 10^{-3} \text{ M}$$

$$K_a = (3.0 \times 10^{-3})^2 / 7.00 \times 10^{-3}$$

$$K_a = 1.3 \times 10^{-3}$$