

- D All three species have unbonded pairs of electrons and could therefore serve as Lewis bases.
- C  $K_a$  increases with increasing dissociation of the acid. Therefore, the smaller  $K_a$  the weaker the acid.

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

- B Le Chatelier's principle gives insight here, stating that when a change is applied to a dynamic system in equilibrium, the system opposes the change in working to restore equilibrium. When we increase the temperature of our water, we are disturbing the equilibrium state of autoprotolysis. Our experiments show that the direction toward a restoration of equilibrium is toward increasing product. This must be taking up heat (endothermic).
- A Examine the math here. Notice that I make use of the fact that  $\log(ab) = \log(a) + \log(b)$  and that the square root of ten is about 3.2.

$$\begin{aligned} pK_a &= -\log K_a \\ &= -(\log(3.2 \times 10^{-8})) \\ &= -(\log 3.2 + \log(10^{-8})) \\ &= -\left(\frac{1}{2} - 8\right) \\ &= 7.5 \end{aligned}$$

- C Our mass of HCN is equal to one mole, which we are putting into 5 liters of water, giving us approximately a 0.2 M solution. If we call our solution hydrogen ion concentration " $x$ ", then the amount of acid that doesn't dissociate will be  $0.2 - x$ . Let's make some approximations. First of all, in our solution there will be only one type of positive charge, the  $H^+$  ions, while there are two types of negative charge, the cyanide ions,  $CN^-$  (our conjugate base) and  $OH^-$  ions. Because the net charge of our solution is zero, we can say that the total number of positives must equal the total number of negatives. If we assume that the hydroxide ion concentration is very small (usually a good assumption in acidic solution) then we can say that  $[CN^-] = [H^+] = x$ .

our original hydrocyanic acid dissociates (usually a good assumption with a *weak* acid), then we can call our equilibrium concentration of HCN merely 0.2 (rather than  $0.2 - x$ ). We can then express  $K_a$  in a very convenient form to find our pH:

$$\begin{aligned} K_a &= \frac{[H^+][CN^-]}{[HCN]} = \frac{x^2}{.2} = 5 \times 10^{-10} \\ [H^+] &= 1 \times 10^{-5} \\ \text{pH} &= -\log(1 \times 10^{-5}) = 5 \end{aligned}$$

- B From our pH, we can find our hydrogen ion concentration:

$$\begin{aligned} \text{pH} &= -\log([H^+]) = 2 \\ [H^+] &= 1 \times 10^{-2} \text{ M} \end{aligned}$$

Because this is a strong acid, we know that *every molecule of acid donates a proton to solution*. Therefore, our solution contains 1/100 mol of acid.

$$\frac{1.0 \text{ g}}{1 \times 10^{-2} \text{ mol}} = 100 \frac{\text{g}}{\text{mol}}$$

- D In aqueous solution (at 298 K) it is always true that

$$\text{pH} + \text{pOH} = 14$$

- C Titration of an acid with a base always produces a salt (in the case of titration of a strong base with a strong acid, the salt is neither an acid nor a base: NaCl here). 5.9 g NaCl is 0.1 mol. This is exactly equal to the number of moles of NaOH

9. A

10. D

11. A First we must determine the  $pK_a$  of carbonic acid:

$$\begin{aligned} pK_a &= -\log K_a \\ &= -(\log(4.5 \times 10^{-7})) \\ &= -(\log 4.5 + \log(10^{-7})) \\ &= -\left(\frac{3}{4} - 7\right) \\ &= 6.25 \end{aligned}$$

Then, using the Henderson-Hasselbalch equation:

$$\begin{aligned} \text{pH} &= pK_a + \log \frac{[\text{base}]}{[\text{acid}]} \\ \text{pH} &= 6.25 + \log \frac{1}{10} \end{aligned}$$

12. D Choice **d** is characteristic of the titration of weak acid with strong base:

