Chapter 15: 2, 6, 8, 10, 18, 26, 40, 52, 66, 72

15.2 Identify each substance as an acid or a base and write a chemical equation showing how it is an acid or a base according to the Arrhenius definition.

a. NaOH(aq)

$$NaOH(aq) \rightarrow Na^{+} (aq) + OH^{-} (aq)$$

b. H_2SO_4 (aq)

$$H_2SO_4 (aq) \rightarrow 2 H^{+} (aq) + SO_4^{2-} (aq)$$

c. HBr (aq)

$$HBr (aq) \rightarrow H^{+} (aq) + Br^{-} (aq)$$

d. Sr $(OH)_2$ (aq)

$$Sr(OH)_2 (aq) \rightarrow Sr^{2+} (aq) + 2 OH^{-} (aq)$$

15.6 Write the formula for the conjugate acid of each base.

a. NH_3
b. ClO_4^-
c. HSO_4^-
d. CO_3^{2-}

SOLUTION:

The conjugate acid is always just the base with a proton added.

e. NH_4^+
f. $HClO_4$
g. H_2SO_4
h. HCO_3^-

15.8 HCO_3^- is amphoteric. Write equations that demonstrate both its acidic nature and its basic nature in aqueous solution.

SOLUTION:

If you are an acid in water, you give a proton to water. If you are a base in water, you get a proton from water.

HCO_3^- (aq) + H_2O (l) ↔ H_2CO_3 (aq) + OH^- (aq)
HCO_3^- (aq) + H_2O (l) ↔ CO_3^{2-} (aq) + H_3O^+ (aq)

15.10 Classify each acid as strong or weak. If the acid is weak, write an expression for the acid ionization constant (K_a)

a. HF
b. $HCHO_2$
c. H_2SO_4
d. H_2CO_3

SOLUTION:

The only strong acid in the list is H_2SO_4. The other 3 are weak.

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]}$$

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}$$

15.18 Calculate $[H_3O^+]$ and $[OH^-]$ for each solution.

a. pH=8.55
pOH = 14-pOH = 14-8.55 = 5.45

So:

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

$$pOH = -\log[OH -]$$

$$8.55 = -\log[H_3O^+]$$

$$[H_3O^+] = 10^{-8.55} = 2.82 \times 10^{-9}$$

$$5.45 = -\log[OH -]$$
$$[OH -] = 10^{-5.45} = 3.55 \times 10^{-6}$$

b. pH=11.23
pOH = 14-11.23=2.77

$$11.23 = -\log[H_3O^+]$$

$$[H_3O^+] = 10^{-11.23} = 5.89 \times 10^{-12}$$

$$2.77 = -\log[OH-]$$
$$[OH-] = 10^{-2.77} = 1.70 \times 10^{-3}$$

c. pH=2.87
pOH=14-2.87=11.13

$$2.87 = -\log[H_3O^+]$$

$$[H_3O^+] = 10^{-2.87} = 1.35 \times 10^{-3}$$

$$11.13 = -\log[OH-]$$
$$[OH-] = 10^{-11.13} = 7.41 \times 10^{-12}$$

15.26 Determine the pH of each solution.

a. 0.028 M HI

HI is a strong acid, it completely dissociates.

$$pH = -\log(H_3O^+) = -\log(0.028) = 1.55$$

b. 0.115 M $HClO_4$

Perchloric acid is a STRONG acid. It completely dissociates.

	$HClO_4$ (aq)	+ H_2O (l)	↔	$HClO_3^-$ (aq)	+H_3O^+ (aq)
I	0.115 M	-		0	0
C	-0.115	-		+0.115	+0.115
E	0 M	-		0.115	0.115

$$pH = -\log(H_3O^+) = -\log(0.115) = 0.939$$

c. A solution that is 0.055 M in $HClO_4$ and 0.028 M in HCl

These are both strong acids. They both completely dissociate. They can be treated consecutively. First the perchloric acid dissociates:

	$HClO_4$ (aq)	+ H_2O (l)	↔	$HClO_3^-$ (aq)	+H_3O^+ (aq)
I	0.055 M	-		0	0
C	-0.055	-		+0.055	+0.055
E	0 M	-		0.055	0.055

Then the HCl

	HCl (aq)	+ H_2O (l)	↔	Cl^- (aq)	+H_3O^+ (aq)
I	0.028 M	-		0	0.055
C	-0.028	-		+0.028	+0.028
E	0 M	-		0.028	0.083

$$pH = -\log(H_3O^+) = -\log(0.083) = 1.08$$

d. A solution that is 1.85% HCl by mass. (Assume a density of 1.01 g/ml for the solution.)

HCl is a strong acid, it completely dissociates. But to get the pH, we need M not % by mass.

$$\frac{1.85\ g\ HCl}{100\ g\ solution}\frac{1\ mol\ HCl}{35.453\ g\ HCl}\frac{1.01\ g\ HCl\ solution}{1\ mL\ solution}\frac{1000\ mL}{1L} = 0.527\ M\ HCl$$

$$pH = -\log(H_3O^+) = -\log(0.527) = 0.278$$

15.40 Determine the percent ionization of a 0.225 M solution of benzoic acid.

SOLUTION: It's a weak acid, $K_a = 6.5x10^{-5}$ M.

Equilibrium problems have how many parts…? 1? 2? No, no, I'm pretty sure it's 3. 3 parts!

1. Balanced equation
   $HC_7H_5O_2$ (aq) + H_2O (l) $\leftrightarrow C_7H_5O_2^-$ (aq) + H_3O^+ (l)

2. Equilibrium constant expression

$$K_a = 6.5 \times 10^{-5} M = \frac{[C_7H_5O_2^-][H_3O^+]}{[HC_7H_5O_2]}$$

3. ICE chart

	$HC_7H_5O_2$ (aq)	+ H_2O (l)	$\leftrightarrow$	$C_7H_5O_2^-$ (aq)	+H_3O^+ (aq)
I	0.225 M	-		0	0
C	-x	-		+x	+x
E	0.225 - x	-		x	x

$$6.5 \times 10^{-5} M = \frac{[C_7H_5O_2^-][H_3O^+]}{[HC_7H_5O_2]} = \frac{(x)(x)}{(0.225 - x)}$$

Might as well try the assumption: x<<0.225

$$6.5 \times 10^{-5} M = \frac{(x)(x)}{(0.225)}$$

$$x^2 = 1.4625 \times 10^{-5}$$
$$x = 3.82 \times 10^{-3}$$

I check the assumption:

$$\frac{0.00382}{0.225} \times 100 = 1.7\% - good\ assumption!$$

	$HC_7H_5O_2$ (aq)	+ H_2O (l)	$\leftrightarrow$	$C_7H_5O_2^-$ (aq)	+H_3O^+ (aq)
I	0.225 M	-		0	0
C	-0.00382	-		+0.00382	+0.00382
E	0.221	-		0.00382	0.00382

Percent ionization is just what it sounds like: what percent of the original acid dissociated (ionized)?

$$\frac{0.00382\ M\ ionization}{0.225\ M\ benzoic\ acid} \times 100 = 1.7\%\ ionization$$

15.52 For each strong base solution, determine $[OH^-]$, $[H_3O^+]$, pH, and pOH

a. 8.77×10^{-3} M LiOH
b. 0.0112 M $Ba(OH)_2$
c. 1.9×10^{-4} M KOH
d. 5.0×10^{-4} M $Ca(OH)_2$

SOLUTION:

Since these are all strong bases, it takes all the equilibrium fun out of it. They each completely dissociate.

a. 8.77×10^{-3} M will create 8.77×10^{-3} M OH^- ions.
pOH=-log(OH-)=-log(8.77×10^{-3}) = 0.826
pH = 14-pOH = 14-0.826= 13.17
$$[H_3O^+] = 10^{-13.17} = 6.76 \times 10^{-14} M$$

b. $0.0112\ M\ Ba(OH)_2 \frac{2\ mol\ OH-}{1\ mol\ Ba(OH)_2} = 0.0224\ M\ OH-$
pOH=-log(OH-)=-log(0.0224) = 1.64
pH = 14-pOH = 14-1.64= 12.36
$$[H_3O^+] = 10^{-12.36} = 4.37 \times 10^{-13} M$$

c. 1.9×10^{-4} M KOH will create 1.9×10^{-4} M OH-
pOH=-log(OH-)=-log(1.9×10^{-4}) = 3.72
pH = 14-pOH = 14-3.72= 10.28
$$[H_3O^+] = 10^{-10.28} = 5.25 \times 10^{-11} M$$

d. $5.0 \times 10^{-4}\ M\ Ca(OH)_2 \frac{2\ mol\ OH-}{1\ mol\ Ca(OH)_2} = 1.0 \times 10^{-3}\ M\ OH-$
pOH=-log(OH-)=-log(0.0010) = 3
pH = 14-pOH = 14-3= 11
$$[H_3O^+] = 10^{-11} = 1.0 \times 10^{-11} M$$

15.66 Classify each anion as basic or neutral. For the anions that are basic, write an equation that shows how the anion acts as a base.

a. $C_7H_5O_2^-$
Conjugate base of a weak acid (benzoic), so it is basic.

$C_7H_5O_2^-$ (aq) + H_2O(l)↔$HC_7H_5O_2$ (aq) + OH^- (aq)

b. I^-
Conjugate base of a weak acid (HI), so it is basic.
I^- (aq) + H_2O(l)↔HI (aq) + OH^- (aq)

c. NO_3^-

Conjugate base of a strong acid (HNO_3) so it is neutral.

d. F^-

Conjugate base of a weak acid (HF), so it is basic.

F^- (aq) + H_2O(l)↔HF (aq) + OH^- (aq)

15.72 Determine whether each salt will form a solution that is acidic, basic, or pH-neutral.

a. $Al(NO_3)_3$

Al^{3+} is the conjugate acid of $Al(OH)_3$ – a weak base
NO_3^- is the conjugate base of HNO_3 – a strong acid

The NO_3^- will do nothing, so the solution will be acidic due to the Al^{3+}

b. $C_2H_5NH_3NO_3$

$C_2H_5NH_3^+$ is the conjugate acid of ethylamine – a weak base
NO_3^- is the conjugate base of HNO_3 – a strong acid

The NO_3^- will do nothing, so the solution will be acidic due to the ethylamine

c. K_2CO_3

K^+ is the conjugate of a strong base (KOH) so we ignore it.

CO_3^{2-} is the conjugate of a weak acid (H_2CO_3) so the solution will be basic.

e. RbI

Rb^+ is the conjugate of a strong base (RbOH) so we ignore it.
I- Is the conjugate of a strong acid (HI), so we ignore it.

The resulting solution will, as a result, be pH-neutral.

f. NH_4ClO

NH_4^+ is the conjugate acid of a weak base, NH_3, $K_b = 1.76\text{x}10^{-5}$

ClO^- is the conjugate base of a weak acid, HClO, $Ka = 2.9\text{x}10^{-8}$

If you calculate the corresponding K_a of NH_4^+ $K_a = \frac{1\times10^{-14}}{1.76\times10^{-5}} = 5.68 \times 10^{-10}$

And the K_b of ClO^- $^+$ $K_b = \frac{1\times10^{-14}}{2.9\times10^{-8}} = 3.45 \times 10^{-7}$

You can see that ClO^- is a better base than NH_4^+ is an acid. The resulting solution should be basic.