BC Science Chemistry 12

Chapter 4 – Acid - Base Equilibrium Answer Key

September 20, 2012

4.1 Identifying Acids and Bases

Warm Up

Observation	Acid	Base
Turns phenolphthalein pink		√
Feels slippery		√
Has pH = 5.0	√	
Tastes sour	√	
Conducts electricity	√	√
Reacts with metal to produce a gas	√	

Practice Problems - Identifying Arrhenius Acids and Bases

1.

a) H_2SO_4 acid
b) XeF_6 molecular
c) CH_3COOH acid
d) $NaCH_3COO$ salt
e) KOH base
f) NH_3 molecular (but also a base)

2.

a) $CH_3COOH + LiOH \rightarrow LiCH_3COO + H_2O$
salt

b) $2HI + Ca(OH)_2 \rightarrow 2H_2O + CaI_2$
salt

c) $3Mg(OH)_2 + 2H_3PO_4 \rightarrow 6H_2O + Mg_3(PO_4)_2$
salt

3.

	Parent Acid	Parent Base
a) KNO_2	HNO_2	KOH
b) NH_4Cl	HCl	NH_4OH (or NH_3)
c) CuC_2O_4	$H_2C_2O_4$	$Cu(OH)_2$

d) $NaCH_3COO$ CH_3COOH $NaOH$

Practice Problems – Identifying Conjutate Acid-Base Pairs

1.

a) HIO_3 + NO_2^- <-> HNO_2 + IO_3^-
acid base acid base

b) HF + $HC_2O_4^-$ <-> $H_2C_2O_4$ + F^-
acid base acid base

c) $Al(H_2O)_6^{3+}$ + SO_3^{2-} <-> HSO_3^- + $Al(H_2O)_5OH^{2+}$
acid base acid base

2.

Conjugate acid	Conjugate base
H_2O_2	HO_2^-
H_3BO_3	$H_2BO_3^-$
HCOOH	$HCOO^-$
$HC_6H_5O_7^{2-}$	$C_6H_5O_7^{3-}$

3..

a) HNO_2 + NH_3 <-> NO_2^- + NH_4^+
acid base base acid
conjugate pairs: HNO_2/ NO_2^- and NH_3/NH_4^+

b) $H_3C_6H_5O_7$ + CN^- <-> $H_2C_6H_5O_7^-$ + HCN
acid base base acid
conjugate pairs: $H_3C_6H_5O_7$/ $H_2C_6H_5O_7^-$ and CN^-/ HCN

c) PO_4^{3-} + H_2S <-> HPO_4^{2-} + HS^-
base acid acid base
conjugate pairs: HPO_4^{2-}/ PO_4^{3-} and H_2S / HS^-

Quick Check

1.

HCO_3^- + CN^- <-> CO_3^{2-} + HCN

2. Write an equation for a reaction between HCO_3^- and H_2O where HCO_3^- acts as a base.

$HCO_3^- + H_2O \leftrightarrow H_2CO_3 + OH^-$

3. Circle substances that are amphiprotic in the following list:

a) CH_3COOH b) $H_2PO_4^-$ c) PO_4^{3-} d) $H_2C_2O_4$ e) $HC_2O_4^-$

b) and e) are amphiprotic

4.1 Activity – Conjugate Pairs Memory Game

Results and Discussion:

1. 2 substances that differ by a H^+ ion or proton

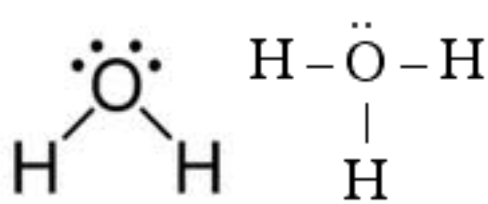
2. They differ by 2 protons. H_2SO_3 and HSO_3^- are conjugate pairs, and SO_3^{2-} and HSO_3^- are conjugate pairs.

Review Questions

1. Both Arrhenius and Bronsted Lowry acids contain H^+ ions. Definitions of a base differ. Arrhenius base contains OH^- ions. Bronsted Lowry base accepts H^+ ion.

2. A Hydrogen atom contains 1 proton and 1 electron. When a H^+ ion is formed, the electron is given away, leaving a single proton.

3.



One of the lone pairs of electrons on the oxygen atom of water will attract the H^+ ion from an acid.

4.

a) $NH_3 + H_3PO_4 \leftrightarrow NH_4^+ + H_2PO_4^-$

base acid acid base

b) $H_2PO_4^- + SO_3^{2-} \leftrightarrow HSO_3^- + HPO_4^{2-}$

acid base acid base

c) $CH_3NH_2 + CH_3COOH \leftrightarrow CH_3COO^- + CH_3NH_3^+$

base acid base acid

5..

$HCOOH + H_2O \leftrightarrow HCOO^- + H_3O^+$

Acid base base acid

6.

$C_5H_5N + H_2O \leftrightarrow C_5H_5NH^+ + OH^-$

Base acid acid base

7.

$ClO^- + NH_4^+ \leftrightarrow HClO + NH_3$

Base acid acid base

b) HClO

8. a) Add a H^+ ion. Example: CN^- becomes HCN

b) Remove a H^+ ion. Example: HNO_2 becomes NO_2^-

9. a) HO_2^-

b) $N_2H_5^+$

c) $OC_6H_5^-$

d) $C_6H_5NH_3^+$

10. A substance that is able act as both an acid or a base depending on the other substance present. Examples: HSO_3^-, $H_2PO_4^-$, HPO_4^{2-}, H_2O

11.

a)

As an acid: HCO_3^- + H_2O $\leftrightarrow$ CO_3^{2-} + H_3O^+

As a base: HCO_3^- + H_2O $\leftrightarrow$ H_2CO_3 + OH^-

Use litmus paper to test the solution. If it is acidic, the litmus will turn red. If it is basic, the litmus will turn blue.

b) As a base. H_2O and CO_2 can be written as H_2CO_3.

12.

H_2O + H_2O $\leftrightarrow$ H_3O^+ + OH^-

Acid base acid base

4.2 The Strengths of Acids and Bases

Warm-Up

1. Ionic compounds contain a metal ion, or NH_4^+ ions. Molecular compounds contain only non-metals. (Some acids are ionic…you will learn this in this section)

2. ions

3. 2 substances that differ by a proton or H^+ ion. Ex. HF/F^-

Quick Check

1. Concentrated acids have a high molarity (ex. 6M acid) whereas a strong acid is one that ionizes completely (ex. HCl). You can have a concentrated weak acid – such as 6 M HF.

2. HNO_3 is a strong acid so ionizes completely:

$HNO_3(aq)$ + $H_2O(l)$ $\rightarrow$ $H_3O^+(aq)$ + $NO_3^-(aq)$
1.0 M 1.0 M

$CH_3COOH(aq)$ + $H_2O(l)$ $\leftrightarrow$ $H_3O^+(aq)$ + $CH_3COO^-(aq)$
1.0 M very few very few

Because both acids are the same concentration, but HNO_3 is strong and CH_3COOH is weak, the HNO_3 will have more ions in solution.

3. In 0.1 M NaOH: a bright light In 0.1 M NH_3: very dim light or no light visible

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$ $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$

4. No, Cl^- will not accept an H^+ ion. HCl is a strong acid and completely ionizes.

Quick Check

1. $HF(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + F^-(aq)$ $K_a = [H_3O^+][F^-] / [HF]$

2. NaOH is a strong base, which means that it ionizes completely. There is no equilibrium present, so no K_b.

3. $CH_3CH_2NH_2(aq) + H_2O(l) \leftrightarrow OH^-(aq) + CH_3CH_2NH_3^+(aq)$

$K_b = [OH^-][CH_3CH_2NH_3^+] / [CH_3CH_2NH_2]$

4. Acid: $HC_2O_4^-(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_2O_4^{2-}(aq)$

$K_a = [H_3O^+][C_2O_4^{2-}] / [HC_2O_4^-]$

Base: $HC_2O_4^-(aq) + H_2O(l) \leftrightarrow OH^-(aq) + H_2C_2O_4(aq)$
$K_b = [OH^-][H_2C_2O_4] / [HC_2O_4^-]$

Quick Check

1

a) strong acid

b) strong base

c) weak base

d) weak acid

e) weak base

f) weak acid

2.

c) weak base: $NH_3(aq)$ + $H_2O(l)$ <-> $NH_4^+(aq)$ + $OH^-(aq)$
K_b= $[NH_4^+][OH^-] / [NH_3]$

d) weak acid: $C_6H_5COOH(aq)$ + $H_2O(l)$ <-> $H_3O^+(aq)$ + $C_6H_5COO^-(aq)$
K_a= $[H_3O^+][C_6H_5COO^-] / [C_6H_5COOH]$

e) weak base: $CN^-(aq)$ + $H_2O(l)$ <-> $HCN(aq)$ + $OH^-(aq)$
K_b= $[HCN][OH^-] / [CN^-]$

f) weak acid: $HNO_2(aq)$ + $H_2O(l)$ <-> $H_3O^+(aq)$ + $NO_2^-(aq)$
K_a= $[H_3O^+][NO_2^-] / [HNO_2]$

3. Methanoic acid > carbonic acid > boric acid > phenol

$HCOOH > H_2CO_3 > H_3BO_3 > C_6H_5OH$

Practice Problems – Predicting Whether Reactants or Products Will Be Favoured in a Bronsted-Lowry Acid Base Equilibrium

1

a) H_2O_2 + HSO_3^- <-> H_2SO_3 + HO_2^- reactants favored

b) $H_3C_6H_5O_7$ + NH_3 <-> $H_2C_6H_5O_7^-$ + NH_4^+ products favored

c) HCO_3^- + $H_2PO_4^-$ <-> H_2CO_3 + HPO_4^{2-} reactants favored

2. H_3AsO_4 + SO_4^{2-} <-> $H_2AsO_4^-$ + HSO_4^-

HSO_4^- is stronger

3.

SO_3^{2-} + $Cr(H_2O)_6^{3+}$ <-> HSO_3^- + $Cr(H_2O)_5OH^{2+}$ products favored so $K_{eq} > 1$

Activity – Determining the Relative Strengths of 6 Acids

Procedure:

In the table below, compare the strength of each pair of acids HA to HIn. The first one has been filled in for you from the discussion above.

	HIn_1 / In_1^-	**HIn_2 / In_2^-**	**HIn_3 / In_3^-**
HA_1 / A_1^-	$HA_1 > HIn_1$	$HA_1 > HIn_2$	$HA_1 < HIn_3$

HA_2 / A_2^-	$HA_2 > HIn_1$	$HA_2 > HIn_2$	$HA_2 > HIn_3$
HA_3 / A_3^-	$HA_3 < HIn_1$	$HA_3 > HIn_2$	$HA_3 < HIn_3$

Results and Discussion

1. Rank the 6 unknown acids in order from strongest to weakest:

$HA_2 > HIn_3 > HA_1 > HIn_1 > HA_3 > HIn_2$

2. Construct a table similar to the Relative Strengths of Bronsted-Lowry Acids using the 6 unknown acids. Be sure to include ionization equations and an arrows on each side of the table labelled: increasing strength of acid or base.

Stronger acid $HA_2 \leftrightarrow H^+ + A_2^-$ weaker base

$HIn_3 \leftrightarrow H^+ + In_3^-$

$HA_1 \leftrightarrow H^+ + A_1^-$

$HIn_1 \leftrightarrow H^+ + In_1^-$

$HA_3 \leftrightarrow H^+ + A_3^-$

Weaker acid $HIn_2 \leftrightarrow H^+ + In_2^-$ stronger base

Review Questions

1.

a) strong base

b) weak acid

c) Strong acid

d) weak base

2.

$H_3BO_3 + H_2O \leftrightarrow H_2BO_3^- + H_3O^+$ $K_a = [H_2BO_3^-][H_3O^+] / [H_3BO_3]$

$PO_4^{3-} + H_2O \leftrightarrow HPO_4^{2-} + OH^-$ $K_b = [HPO_4^{2-}][OH^-] / [PO_4^{3-}]$

3.

Both acids have the same concentration. Oxalic acid is a weak acid and hydriodic acid is strong. Conductivity depends on ions in solution. Since hydriodic acid is a strong aci, it will ionize completely. Oxalic acid is weak, so it ionizes to a lesser extent, forming fewer ions in solution.

4. HNO_3 is a strong acid so ionizes completely:

$HNO_3(aq)$ + $H_2O(l)$ → $H_3O^+(aq)$ + $NO_3^-(aq)$
2.0 M 2.0 M total [ion] = 4.0 M

HNO_2 is a weak acid, so does ionize completely.

5. a) 6 M NH_3

b) 0.001 M HCl

6. a) water < carbonic acid < citric acid < sulphurous acid < sulphuric acid

b) carbonate > ammonia > monohydrogen phosphate > fluoride > nitrite > water

7.

$HC_6H_5O_7^{2-}$ + H_2O <-> $C_6H_5O_7^{3-}$ + H_3O^+
$K_a = [C_6H_5O_7^{3-}][H_3O^+] / [HC_6H_5O_7^{2-}]$

$H_2C_6H_5O_7^-$ + H_2O <-> $HC_6H_5O_7^{2-}$ + H_3O^+
$K_a = [HC_6H_5O_7^{2-}][H_3O^+] / [H_2C_6H_5O_7^-]$

$Al(H_2O)_6^{3+}$ + H_2O <-> $Al(H_2O)_5OH^{2+}$ + H_3O^+
$K_a = [Al(H_2O)_5OH^{2+}][H_3O^+] / [Al(H_2O)_6^{3+}]$

H_2O_2 + H_2O <-> HO_2^- + H_3O^+
$K_a = [HO_2^-][H_3O^+] / [H_2O_2]$

8.

NH_3 + H_2O <-> NH_4^+ + OH^-
$K_b = [NH_4^+][OH^-] / [NH_3]$

$C_6H_5COO^-$ + H_2O <-> C_6H_5COOH + OH^-
$K_b = [C_6H_5COOH][OH^-] / [C_6H_5COO^-]$

CH_3COO^- + H_2O <-> CH_3COOH + OH^-
$K_b = [CH_3COOH][OH^-] / [CH_3COO^-]$

$HC_6H_5O_7^{2-}$ + H_2O ←> $C_6H_5O_7^-$ + OH^-
$K_b = [C_6H_5O_7^-][OH^-] / [HC_6H_5O_7^{2-}]$

C_5H_5N + H_2O ←> $C_5H_5NH^+$ + OH^-
$K_b = [C_5H_5NH^+][OH^-] / [C_5H_5N]$

9.

a) $Fe(H_2O)_6^{3+}(aq)$ + $HO_2^-(aq)$ $\rightleftharpoons$ H_2O_2 + $Fe(H_2O)_5OH^{2+}$ products

b) $H_2SO_3(aq)$ + $IO_3^-(aq)$ $\rightleftharpoons$ HSO_3^- + HIO_3 reactants

c) $CN^-(aq)$ + $H_2PO_4^-(aq)$ $\rightleftharpoons$ HCN + HPO_4^{2-} products

10. $H_2C_2O_4$ + HCO_3^- ←> H_2CO_3 + $HC_2O_4^-$

11. HClO > HBrO > H_2SiO_3

12. All are strong acids so completely ionize in solution to form H_3O^+:

HCl + H_2O → H_3O^+ + Cl^-

HBr + H_2O → H_3O^+ + Br^-

HI + H_2O → H_3O^+ + I^-

4.3 The Ionization of Water

Warm Up

1. A weak acid only partially ionizes. A strong acid ionizes completely.

2. Both acids have the same concentration, so we can compare strength. HCl is a strong acid, so completely ionizes. CH_3COOH is a weak acid so produces fewer ions. Using the K_a Table, water is the weakest acid, so would have the least ions present. From least to most conductive: water, CH_3COOH , HCl

3. H_2O ←> H^+ + OH^- $K_a = [H^+][OH^-]$

Quick Check

1. energy + $2H_2O$ <-> H_3O^+ + OH^-

If temperature increases, equilibrium shifts right and the concentrations of H_3O^+ and OH^- both increase.

$K_w = [H_3O^+][OH^-]$

Therefore, K_w will increase as well.

2. $[H_3O^+] = [OH^-]$

3.

$K_w = [H_3O^+][OH^-]$

$2.9 \times 10^{-15} = x^2$

$x = 5.4 \times 10^{-8}$ M $= [H_3O^+] = [OH^-]$

Practice Problems – Calculating $[H_3O^+]$ and $[OH^-]$ in Solutions of a Strong Acid or Strong Base (Assume the temperature in each case is 25°C.)

1.

$[H_3O^+] = 0.15$ M because $HClO_4$ is a strong acid

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = (0.15)\,[OH^-]$

$[OH^-] = 6.67 \times 10^{-14}$ M

acidic because $[H_3O^+] > [OH^-]$

2.

$Mg(OH)_2(s)$ <-> $Mg^{2+}(aq)$ + $2OH^-(aq)$

s 2s

$K_{sp} = [Mg^{2+}][OH^-]^2$

$5.6 \times 10^{-12} = 4s^3$

$s = 1.1 \times 10^{-4}$ M

$[OH^-] = 2s = 2.2 \times 10^{-4}$ M

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = [H_3O^+]\,(2.2 \times 10^{-4})$

$[H_3O^+] = 4.5 \times 10^{-11}$ M

basic because $[H_3O^+] < [OH^-]$

3.

$$[NaOH] = \frac{1.42\ \cancel{g}}{0.250\ L} \times \frac{1\ mol}{40.0\ \cancel{g}} = 0.142\ M = [OH^-]$$

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = [H_3O^+]\ (0.142)$

$[H_3O^+] = 7.04 \times 10^{-14}$ M

basic because $[H_3O^+] < [OH^-]$

Practice Problems – What Happens When a Strong Acid is Added to a Strong Base?

1.

$$[HNO_3] = \frac{1.5\ \cancel{mol}}{L} \times \frac{0.150\ L}{0.400\ \cancel{L}} = 0.56\ M = [H_3O^+]$$

$$[KOH] = \frac{0.80\ \cancel{mol}}{L} \times \frac{0.250\ L}{0.400\ \cancel{L}} = 0.50\ M = [OH^-]$$

Excess $[H_3O^+] = 0.56\ M – 0.50\ M = 0.06$ M

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} =(0.06)\ [OH^-]$

$[OH^-]= 2 \times 10^{-13}$ M

2.

Excess $[H_3O^+] = 0.12\ M =$ initial $[H_3O^+]$ - initial $[OH^-]$

$0.12\ M = 0.20\ M$ - initial $[OH^-]$

initial $[OH^-] = 0.08$ M

$$\text{mass NaOH} = \frac{0.08\ \cancel{mol}}{\cancel{L}} \times \frac{40.0\ g}{1\ \cancel{mol}} \times 0.500\ \cancel{L} = 2\ g$$

3.

$$[HBr] = \frac{0.105\ \cancel{mol}}{L} \times \frac{0.0184\ L}{0.0407\ \cancel{L}} = 0.0475\ M = [H_3O^+]$$

$$[HCl] = \underline{0.256\ \cancel{mol}} \times \underline{0.0223\ L} = 0.140\ M = [H_3O^+]$$

L 0.0407 ~~L~~

Total $[H_3O^+]$ = 0.0475 M + 0.140 M = 0.188M

$K_w = [H_3O^+][OH^-]$

1.00×10^{-14} =(0.188) $[OH^-]$

$[OH^-]$= 5.32×10^{-14} M

Activity 4.3 – Counting Water Molecules and Hydronium Ions

Question: How many water molecules does it take to produce one hydronium ion?

Procedure:

1. Calculate the number of water molecules present in 1.0 L of water using the density of water (1.00 g/mL), and its molar mass.

$$1.0\ \cancel{L} \times \frac{100.\ \cancel{mL}}{1\ \cancel{L}} \times \frac{1.00\ \cancel{g}}{\cancel{mL}} \times \frac{1\ \cancel{mol}}{18.0\ \cancel{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}} = 3.3 \times 10^{25}\ \text{molecules}$$

2. Calculate the number of hydronium ions in 1.0L of water. (HINT: you need the $[H_3O^+]$ in pure water from above)

$$1.0\ \cancel{L} \times \frac{1.0 \times 10^{-7}\ \cancel{mol}}{1\ \cancel{L}} \times \frac{6.02 \times 10^{23}\ \text{ions}}{1\ \cancel{mol}} = 6.0 \times 10^{16}\ \text{ions}$$

3. Using the above answers, calculate the ratio of ions/molecules. This is the percentage ionization of water.

$$\frac{6.0 \times 10^{16}\ \text{ions}}{3.3 \times 10^{25}\ \text{molecules}} = 1.8 \times 10^{-9}$$

4. Using the ratio above, calculate the number of water molecules required to produce 1 hydronium ion.

$$1\ \cancel{ion} \times \frac{3.3 \times 10^{25}\ \text{molecules}}{6.0 \times 10^{16}\ \cancel{ions}} = 5.5 \times 10^{8}\ \text{molecules}$$

Results and Discussion

1. From the ratio of ions to molecules, it is evident that an enormously small percentage of water molecules actually ionize. Because there are an enormously large number of molecules present in the solutions we use, a reasonable number of hydronium and hydroxide ions are present. What volume of water contains only one hydronium ion?

$$5.5 \times 10^{8}\ \cancel{molecules} \times \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \cancel{molecules}} \times \frac{18.0\ \cancel{g}}{1\ \cancel{mol}} \times \frac{1\ \text{mL}}{1.0\ \cancel{g}} = 1.6 \times 10^{-14}\ \text{mL}$$

Review Questions

1. $NH_3 + NH_3 \leftrightarrow NH_4^+ + NH_2^-$

2. Complete the following table:

[H_3O^+]	**[OH^-]**	**Acidic, basic or neutral ?**
1.7×10^{-15}	$6.0M$	basic
$3.2 \times 10^{-4}M$	3.1×10^{-11}	acidic
1.1×10^{-3}	$9.2 \times 10^{-12}M$	acidic
$2.5M$	4.0×10^{-15}	acidic
2.1×10^{-10}	$4.7 \times 10^{-5}\ M$	basic

3. 57.1 kJ + $2H_2O \leftrightarrow H_3O^+ + OH^-$

As temperature increases, equilibrium shifts right and [H_3O^+] and [OH^-] increase, so K_w increases.

4. HI is a strong acid, so [H_3O^+] = 0.20 M

$K_w = [H_3O^+][OH^-]$

$1.0 \times 10^{-15} = (0.20)\ [OH^-]$

$[OH^-] = 5.0 \times 10^{-15}$ M

5.

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = (6.3 \times 10^{-7})\ [OH^-]$

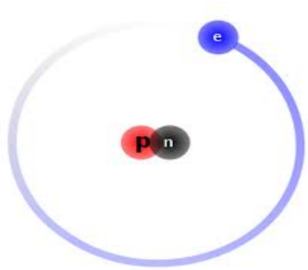
$[OH^-] = 1.6 \times 10^{-8}$ M

acidic because $[H_3O^+] > [OH^-]$

6. Complete the table:

Temperature	**K_w**	**[H_3O^+]**	**[OH^-]**	**Acidic, basic, or neutral ?**
50° C	5.5×10^{-14}	2.3×10^{-7} M	2.3×10^{-7} M	neutral
100° C	5.1×10^{-13}	7.1×10^{-7} M	7.1×10^{-7} M	neutral

7.



$2D_2O \leftarrow> D_3O^+ + OD^-$

$K_w = [D_3O^+][OD^-]$

$K_w = (8.9 \times 10^{-8})(8.9 \times 10^{-8})$

$K_w = 7.9 \times 10^{-15}$

8.

$Ca(OH)_2(s) \leftarrow> Ca^{2+}(aq) + 2OH^-(aq)$

s 2s

$K_{sp} = [Ca^{2+}][OH^-]^2$

$4.7 \times 10^{-6} = 4s^3$

$s = 0.0106$ M

$[OH^-] = 2s = 0.0211$ M

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = [H_3O^+](0.0211)$

$[H_3O^+] = 4.7 \times 10^{-13}$ M

9. A student combines the following solutions: (draw 3 graduated cylinders and put the level of solution in the first cylinder at 21 mL, the second cylinder 15 mL and the third cylinder 25 mL. Label the first cylinder 0.80*M* HCl, the second cylinder 1.2*M* NaOH and the third cylinder 0.60*M* HI) Calculate the $[H_3O^+]$ and $[OH^-]$ in the resulting solution.

Cylinder 1: $\frac{0.80 \text{ mol}}{\cancel{L}} \times 0.021 \cancel{L} = 0.017$ mol HCl = $[H_3O^+]$

Cylinder 2: $\frac{1.2 \text{ mol}}{\cancel{L}} \times 0.015 \cancel{L} = 0.018$ mol NaOH = $[OH^-]$

Cylinder 3: $\frac{0.60 \text{ mol}}{\cancel{L}} \times 0.025 \cancel{L} = 0.015$ mol HI = $[H_3O^+]$

$[H_3O^+]$ excess = $\frac{(0.017 \text{ mol} + 0.015 \text{ mol}) - 0.018 \text{ mol}}{0.061 \text{ L}} = 0.23$ M

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = (0.23)[OH^-]$

$[OH^-] = 4.3 \times 10^{-14}$ M

10.

Excess $[OH^-] = 0.010$ M = initial $[OH^-]$ - initial $[H_3O^+]$

0.010 M = initial $[OH^-]$ - 0.250 M

initial $[OH^-]$ = 0.350 M

$$\text{mass } Sr(OH)_2 = 0.150\ \cancel{L} \times \frac{0.350\ \cancel{\text{mol } OH^-}}{\cancel{L}} \times \frac{1\ \cancel{\text{mol } Sr(OH)_2}}{2\ \cancel{\text{mol } OH^-}} \times \frac{121.6\ g}{1\ \cancel{\text{mol } Sr(OH)_2}} = 3.2\ g$$

Section 4.4 pH and pOH

Warm Up

1.

Solution	$[H_3O^+]$	$[OH^-]$
A. 1.0 *M* NaOH	**1.0×10^{-14} *M***	**1.0 *M***
B. 1.0 *M* HCl	**1.0 *M***	**1.0×10^{-14} *M***

2. The moles of hydronium ions present in solution A are equal to the moles of hydroxide ions present in solution B. Therefore, upon mixing, neither will be in excess as a result of the following net ionic equation for the neutralization reaction:

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$$

The solution will therefore be neutral with $[H_3O^+] = [OH^-] =$ **1.0×10^{-7} *M***

3. $$1.0\ \cancel{\text{mol } H_3O^+} \times \frac{1.0\ \text{L solution}}{1.0 \times 10^{-14}\ \cancel{\text{mol } H_3O^+}} = \mathbf{1.0 \times 10^{14}\ L\ solution}$$

Quick Check

1.

Solution	$[H_3O^+]$	pH
orange juice	3.2×10^{-4} *M*	**3.49**
milk of magnesia	2.52×10^{-11} *M*	**10.598**

stomach acid	0.031 *M*	**1.51**

2. The $[H_3O^+]$ increased because the pH decreased. A **5** unit pH decrease corresponds to a $10^{\mathbf{5}}$ times increase in hydronium concentration.

3. Initial pH (prior to dilution) = - log (0.10) = **1.00**

After the dilution, the $[H_3O^+]$ = 0.010 *M*. So final pH (after dilution) = - log (0.010) = **2.00**

(Note 2 decimal places in the answer corresponding to 2 significant digits in each concentration)

Practice Problems – Converting pH to $[H_3O^+]$

1.

$[H_3O^+]$	No. of sig. digits	pH
5.00×10^{-5} *M*	**3**	**4.301**
4.6×10^{-4} *M*	**2**	3.34
6.4×10^{-11} *M*	**2**	**10.19**
8.81×10^{-7} *M*	**3**	6.055
0.003 45 *M*	**3**	**2.462**

2. Just before the equivalence point was reached: $[H_3O^+] = 10^{-4.35} =$ **4.5×10^{-5} *M***

Just after the equivalence point was reached: $[H_3O^+] = 10^{-9.65} =$ **2.2×10^{-10} *M***

3. Because pH is defined as a negative logarithm, converting a negative pH value to a concentration involves the following: $[H_3O^+] = 10^{-(-1.20)} = 10^{1.20} =$ **16 *M*** (2 sig. digits)

The pH scale is not necessary to conveniently express concentrations this large.

Practice Problems – pH and pOH

1.

Solution	$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic/Basic/Neutral?
Orange Juice	**3×10^{-4} *M***	**3×10^{-11} *M***	**3.5**	10.5	**Acidic**
Tears	3.98×10^{-8} *M*	**2.51×10^{-7} *M***	**7.40**	**6.60**	**Basic**
Blood	**4.0×10^{-8} *M***	**2.5×10^{-7} *M***	7.40	**6.60**	**Basic**
Milk	**3.16×10^{-7} *M***	3.16×10^{-8} *M*	**6.500**	**7.500**	**Acidic**

2.

a. As a solution becomes more acidic, both $[H_3O^+]$ and pOH **increase** and both $[OH^-]$ and pH **decrease**.

b. A basic solution has a pOH value which is **less** than 7, and $[H_3O^+]$ which is **greater** than 10^{-7} *M*.

c. If the pH of a solution equals 14.0, the $[OH^-]$ equals **1** *M*. (1 sig. digit)

d. If the pOH of a solution decreases by 5, then the $[H_3O^+]$ has **decreased** by a factor of $\mathbf{10^5}$.

3. We would expect the K_w for water at 10°C to be less than 10^{-14} because the autoionization of water is endothermic. Regardless of the temperature, pure water is always neutral and so:

For pure water at any temperature: $[H_3O^+] = [OH^-]$ and pH = pOH

Therefore, because pH + pOH = pK_w, when we calculate the value of pK_w, the pH (and also the pOH) must be *half the value of* pK_w.

$$pK_w = -\log(2.55 \times 10^{-15}) = 14.593 \quad \text{Therefore, pH (and pOH)} = \frac{14.5934}{2} = \mathbf{7.297}$$

Practice Problems – Mixing Strong Acids and Bases

1. $$[H_3O^+]_{ST} = 0.40\,M \times \frac{25.0\ \cancel{mL}}{40.0\ \cancel{mL}} = 0.250\,M \text{ (in excess)}$$

$$[OH^-]_{ST} = 0.30\text{ M} \times \frac{15.0\ \cancel{mL}}{40.0\ \cancel{mL}} = 0.112\,M$$

$[H_3O^+]_{XS} = 0.250\,M - 0.112\,M = 0.138\,M$ so pH = - log (0.138) = **0.86** (2 sig digits)

2. $[H_3O^+]_{ST} = [HCl] = 10^{-0.60} = 0.251\,M$ (in excess)

$$[OH^-]_{ST} = \frac{1.0\ \cancel{g\ NaOH}}{0.2500\text{ L}} \times \frac{1\text{mol NaOH}}{40.0\ \cancel{g}} = 0.100\,M$$

$[H_3O^+]_{XS} = 0.251\ M - 0.100\ M = 0.151\ M$

pH = - log (0.151) = **0.82** (2 sig. digits)

3. The first solution has a pH = 2.00 and is therefore acidic with hydronium ions being the major species in the solution. The second solution has a pOH = 3.00 and is therefore basic with hydroxide ions as the major species.

In the first solution: $[H_3O^+] = 10^{-2.00} = 0.010\ M$

In the second solution: $[OH^-] = 10^{-3.00} = 0.0010\ M$

Therefore:

$$[H_3O^+]_{ST} = 0.010\ M \times \frac{25.0\ \cancel{\text{mL}}}{70.0\ \cancel{\text{mL}}} = 3.57 \times 10^{-3}\ M \text{ (in excess)}$$

$$[OH^-]_{ST} = 0.0010\ M \times \frac{45.0\ \cancel{\text{mL}}}{70.0\ \cancel{\text{mL}}} = 6.43 \times 10^{-4}\ M$$

$$[H_3O^+]_{XS} = 3.57 \times 10^{-3}\ M - 6.43 \times 10^{-4}\ M = 2.93 \times 10^{-3}\ M$$

pH = - log (2.93×10^{-3}) = **2.53** (2 sig. digits)

4. The final solution is basic so the equation which applies is: $[OH^-]_{XS} = [OH^-]_{ST} - [H_3O^+]_{ST}$

As both $[OH^-]_{XS}$ and $[OH^-]_{ST}$ are available to us, we must solve for $[H_3O^+]_{ST}$. Therefore:

$$[H_3O^+]_{ST} = [OH^-]_{ST} - [OH^-]_{XS}$$

Initial pOH = 14.000 - 11.176 = 2.824. So $[OH^-]_{ST} = 10^{-2.824} = 1.500 \times 10^{-3}\ M$

Final pOH = 14.000 - 10.750 = 3.250. So $[OH^-]_{XS} = 10^{-3.250} = 5.623 \times 10^{-4}\ M$

$[H_3O^+]_{ST} = (1.500 \times 10^{-3}\ M) - (5.623 \times 10^{-4}\ M) = 9.377 \times 10^{-4}\ M$

$9.377 \times 10^{-4}\ \underline{\text{mol } H_3O^+} = 9.377 \times 10^{-4}\ \underline{\cancel{\text{mol HCl}}} \times 36.5\ \underline{\text{g HCl}} \times 1.50\ \cancel{\text{L}}$ = **0.0513 g HCl**

L ~~L~~ ~~mol~~ (3 sig. digits)

4.4 Activity: *pH*inding Acidic and Basic Common Solutions

Solution	pH	pOH	$[H_3O^+]$ *M*	$[OH^-]$ *M*	Acidic/Basic/Neutral
Unpolluted Rainwater	5.5	**8.5**	**3 x 10^{-6}**	**3 x 10^{-9}**	**Acidic**
Saliva	**6.7**	7.3	**2 x 10^{-7}**	**5 x 10^{-8}**	**Basic**
Stomach Acid	**1.51**	**12.49**	0.031	**3.2 x 10^{-13}**	**Acidic**
Tears	**7.40**	**6.60**	**4.0 x 10^{-8}**	2.5 x 10^{-7}	**Basic**
Vinegar	2.9	**11.1**	**1 x 10^{-3}**	**8.2 x 10^{-12}**	**Acidic**
Milk	**6.4**	7.6	**4 x 10^{-7}**	**3 x 10^{-8}**	**Acidic**
Milk of Magnesia	**10.49**	**3.51**	3.2 x 10^{-11}	**3.1 x 10^{-4}**	**Basic**
Lemon Juice	**2.30**	**11.70**	**5.0 x 10^{-3}**	2.0 x 10^{-12}	**Acidic**
Tomato Juice	4.2	**9.8**	**6 x 10^{-5}**	**2 x 10^{-10}**	**Acidic**
Orange Juice	**3.5**	10.5	**3 x 10^{-4}**	**3 x 10^{-11}**	**Acidic**
Grapefruit Juice	**3.00**	**11.00**	0.0010	**1.0 x 10^{-11}**	**Acidic**
Liquid Drain Cleaner	**14.00**	**0.00**	**1.0 x 10^{-14}**	1.0	**Basic**
Black Coffee	5.1	**8.9**	**8 x 10^{-6}**	**1 x 10^{-9}**	**Acidic**
Urine	**6.0**	8.0	**1 x 10^{-6}**	**1 x 10^{-8}**	**Acidic**
Blood	**7.40**	**6.60**	4.0 x 10^{-8}	**2.5 x 10^{-7}**	**Basic**
Laundry Bleach	**12.00**	**2.00**	**1.0 x 10^{-12}**	0.010	**Basic**
Windex™	10.7	**3.3**	**2 x 10^{-11}**	**5 x 10^{-4}**	**Basic**
Pepto Bismol	**5.8**	8.2	**2 x 10^{-6}**	**6 x 10^{-9}**	**Acidic**
Household Ammonia	**11.89**	**2.11**	1.3 x 10^{-12}	**7.7 x 10^{-3}**	**Basic**
Red Wine	**2.51**	**11.49**	**3.1 x 10^{-3}**	3.2 x 10^{-12}	**Acidic**

Results and Discussion

1. None of the solutions are neutral, although some are only slightly acidic or basic.

2. The majority of household solutions used for cleaning purposes are quite basic. In the table above, the solutions having hazardous warning labels are *italicized*.

3. The consumable solutions are usually acidic.

4.4 Review Questions

1. The pH of a solution is defined as the negative logarithm of the concentration of hydronium ions. Equation: $pH = -\log [H_3O^+]$

The pOH of a solution is defined as the negative logarithm of the concentration of hydroxide ions. Equation: $pOH = -\log [OH^-]$

2. Sorenson's logarithmic pH and pOH scales are a convenient and compact way of expressing the typically very small concentrations of hydronium and hydroxide ions respectively and the extent by which they can change in aqueous solutions

3.

$[H_3O^+]$	**pH**	**Acidic/Basic/Neutral**
3.50×10^{-6} *M*	**5.456**	**Acidic**
3.1×10^{-12} *M*	11.51	**Basic**
0.00550 *M*	**2.260**	**Acidic**
1.0 *M*	0.00	**Acidic**
6.8×10^{-9} *M*	**8.17**	**Basic**

4.

$[OH^-]$	**pOH**	**Acidic/Basic/Neutral**
7.2×10^{-9} *M*	**8.14**	**Acidic**
2.8×10^{-10} *M*	9.55	**Acidic**
4.88×10^{-4} *M*	**3.312**	**Basic**
1.0×10^{-14} *M*	14.00	**Acidic**
0.000625 *M*	**3.204**	**Basic**

5.

a) As a solution's pOH value and $[H_3O^+]$ both decrease, the solution becomes more **basic**.

b) As a solution's pH value and $[OH^-]$ both decrease, the solution becomes more **acidic**.

c) The **product** of the $[H_3O^+]$ and $[OH^-]$ equals K_w.

d) The **sum** of pH and pOH equals pK_w.

6.

$[H_3O^+]$	**pOH**	**Acidic/Basic/Neutral**
0.0342 *M*	**12.534**	**Acidic**
2.51×10^{-6} *M*	8.400	**Acidic**
7.2×10^{-12} *M*	**2.86**	**Basic**
1.64×10^{-11} *M*	3.215	**Basic**

7. Pure water is always **neutral** at any temperature and so pH = pOH.

Therefore pH $= \frac{pK_w}{2} =$ **6.51** (2 sig. digits)

8. The dissociation equation is: $Sr(OH)_2(s) \rightarrow Sr^{2+}(aq) + 2\ OH^-(aq)$

   Therefore: $[OH^-] = 2 \times 0.30\ M = 0.60\ M$

   $pOH = -\log(0.60) = 0.222$

   $pH = 14.000 - 0.222 =$ **13.78** (2 sig. digits)

9. $[OH^-] = \frac{2.00\ \cancel{g\ NaOH}}{0.5000\ L} \times \frac{1\ mol\ NaOH}{40.0\ g} = 0.100\ M$

   $pOH = -\log(0.100) = 1.000$

   $pH = 14.000 - 1.000 =$ **13.000** (3 sig. digits)

10. $[H_3O^+] = 10^{-2.50} = 3.16 \times 10^{-3}\ M = [HI]$ because HI is a strong acid.

    $3.16 \times 10^{-3} \frac{\cancel{mol\ HI}}{\cancel{L}} \times \frac{127.9\ g\ HI}{\cancel{mol}} \times 2.0\ \cancel{L} =$ **0.81 g** (2 sig. digits)

11.

$[H_3O^+]$	$[OH^-]$	pOH	pH	Acidic/Basic/Neutral
$5.620 \times 10^{-5}\ M$	**$1.779 \times 10^{-10}\ M$**	**9.7497**	**4.2503**	**Acidic**
$2.22 \times 10^{-11}\ M$	$0.000450\ M$	**3.347**	**10.653**	**Basic**
$3.2 \times 10^{-2}\ M$	**$3.2 \times 10^{-13}\ M$**	12.50	**1.50**	**Acidic**
$3 \times 10^{-11}\ M$	**$3 \times 10^{-4}\ M$**	**3.5**	10.5	**Basic**

12. $[H_3O^+]_{ST} = 0.50\ M \times \frac{75.0\ \cancel{mL}}{200.0\ \cancel{mL}} = 0.188\ M$ (in excess)

$[OH^-]_{ST} = \frac{0.20\ \cancel{g\ NaOH}}{0.125\ L} \times \frac{1\ mol}{40.0\ \cancel{g}} \times \frac{125.0\ \cancel{mL}}{200.0\ \cancel{mL}} = 0.025\ M$

$[H_3O^+]_{XS}$ = 0.188 M - 0.0250 M = 0.163 M

pH = - log (0.163) = **0.79** (2 sig. digits)

13.

Initial $[H_3O^+]$ in acidic solution = $10^{-1.50}$ = 0.0316 M

$$[H_3O^+]_{ST} = 0.0316\text{ M} \times \frac{200.0\ \cancel{\text{mL}}}{500.0\ \cancel{\text{mL}}} = 0.0126\ M$$

Initial $[OH^-]$ in basic solution = $10^{-1.50}$ = 0.0316 M

$$[OH^-]_{ST} = 0.0316\ M \times \frac{300.0\ \cancel{\text{mL}}}{500.0\ \cancel{\text{mL}}} = 0.0190\ M \text{ (in excess)}$$

$[OH^-]_{XS}$ = 0.0190 M - 0.0126 M = 0.0064 M (note 1 sig. digit after subtraction)

pOH = - log (0.0064) = 2.19 so pH = 14.00 - 2.19 = **11.8** (1 sig. digit)

14.

[HI] = $[H_3O^+]_{ST}$ because HI is a strong acid.

$$[H_3O^+]_{ST} = \frac{3.2\ \cancel{\text{g HI}}}{0.5000\text{ L}} \times \frac{1\text{ mol HI}}{127.9\ \cancel{\text{g}}} = 0.050\ M$$

Initial pOH = 14.00 - 13.00 = 1.00 so $[OH^-]_{ST}$ = $10^{-1.00}$ = 0.10 M (in excess)

$[OH^-]_{XS}$ = 0.10 M - 0.050 M = 0.05 M (note 1 sig. digit after subtraction)

So pOH = - log (0.050) = 1.30 so pH = 14.00 - 1.30 = **12.7** (1 sig. digit)

15.

One solution method is to combine both acidic solutions (call them "A" and "B") to determine initial $[H_3O^+]$ and then use the initial $[H_3O^+]$ to determine $[H_3O^+]_{ST}$.

Combining the acidic solutions gives: $[H_3O^+]_A = 0.20\ M \times \frac{25.0\ \cancel{mL}}{60.0\ mL} = 0.0833\ M$

$[H_3O^+]_B = 0.15\ M \times \frac{35.0\ \cancel{mL}}{60.0\ \cancel{mL}} = 0.0875\ M$

So initial $[H_3O^+] = 0.0833\ M + 0.0875\ M = 0.1708\ M$

$[H_3O^+]_{ST} = 0.1708\ M \times \frac{60.0\ \cancel{mL}}{100.0\ \cancel{mL}} = 0.1025\ M$

$[OH^-]_{ST} = 0.30\ M \times \frac{40.0\ \cancel{mL}}{100.0\ \cancel{mL}} = 0.120\ M$ (in excess)

$[OH^-]_{XS} = 0.120\ M - 0.1025\ M = 0.0175\ M$

So pOH = - log (0.0175) = 1.756 so pH = 14.000 - 1.756 = **12.24** (2 sig. digits)

An alternative approach would be to calculate the total number of moles of hydronium ions from the acid solutions and then subtract that total from the moles of hydroxide ions from the basic solution. The excess moles of hydroxide ions present in the final 100.0 mL are then used to calculate final pH.

16.

The final solution is basic so the equation which applies is: $[OH^-]_{XS} = [OH^-]_{ST} - [H_3O^+]_{ST}$

As both $[OH^-]_{XS}$ and $[OH^-]_{ST}$ are available to us, we must solve for $[H_3O^+]_{ST}$. Therefore:

$$[H_3O^+]_{ST} = [OH^-]_{ST} - [OH^-]_{XS}$$

$[OH^-]_{ST} = [KOH] = 0.0350\ M$

Final pOH = 14.000 - 11.750 = 2.250. So $[OH^-]_{XS} = 10^{-2.250} = 0.005623\ M$

$[H_3O^+]_{ST} = (0.0350\ M) - (0.005623\ M) = 0.02938\ M$

$[H_3O^+]_{ST} = [HCl] = 0.02938 \frac{\cancel{mol\ HCl}}{\cancel{L}} \times 36.5 \frac{g\ HCl}{\cancel{mol}} \times 0.4500\ \cancel{L} =$ **0.483 g HCl** (3 sig. digits)

17.

The final solution is acidic so the equation which applies is: $[H_3O^+]_{XS} = [H_3O^+]_{ST} - [OH^-]_{ST}$

As both $[H_3O^+]_{XS}$ and $[H_3O^+]_{ST}$ are available to us, we must solve for $[OH^-]_{ST}$. Therefore:

$$[OH^-]_{ST} = [H_3O^+]_{ST} - [H_3O^+]_{XS}$$

$[H_3O^+]_{ST}$ = [HCl] = 0.0125 *M*

Final pH = 2.75 so $[H_3O^+]_{XS}$ = $10^{-2.75}$ = 0.00178 *M*

$[OH^-]_{ST}$ = 0.0125 *M* - 0.00178 *M* = 0.01072

$[OH^-]_{ST}$ = [LiOH] = 0.0107 ~~mol LiOH~~ / ~~L~~ x 23.9 g Li / ~~mol~~ x 0.5000 ~~L~~ = **0.13 g LiOH**

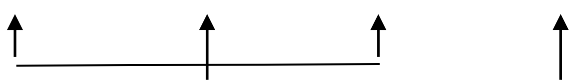
(2 sig. digits)

Section 4.5 Calculations Involving K_a and K_b

Warm Up

1. a) HCl(*aq*) + H_2O(*l*) → Cl^-(*aq*) + H_3O^+(*aq*) (*note* ***1-way*** *arrow*)

Acid 1 ***Base 2*** ***Base 1*** ***Acid 2***

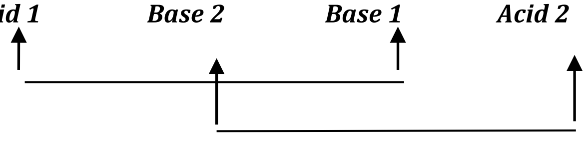


b) As HCl is a strong acid and thus 100 % ionized, the $[H_3O^+]$ = [HCl] = 0.50 *M*

and pH = - log (0.50) = 0.30

2. a) HF(*aq*) + H_2O(*l*) $\rightleftharpoons$ F^-(*aq*) + H_3O^+(*aq*) (*note* ***2-way*** *arrow*)

Acid 1 ***Base 2*** ***Base 1*** ***Acid 2***



The reactants are favoured in this equilibrium because the HF is a weaker acid than H_3O^+ (and H_2O is a weaker base than F^-) and acid-base equilibria always favour the weaker species.

b) As HF is a weak acid and thus ionizes only to a slight extent, the $[H_3O^+]_{eq}$ << $[HF]_{initial}$ Therefore, we would expect the pH to be higher than 0.30.

3.

a. K_a for $HNO_2 = \dfrac{[NO_2^-][H_3O^+]}{[HNO_2]}$	b. K_b for $C_2O_4^{2-} = \dfrac{[HC_2O_4^-][OH^-]}{[C_2O_4^{2}]}$
c. K_a for $HC_2O_4^- = \dfrac{[C_2O_4^{2-}][H_3O^+]}{[HC_2O_4^-]}$	d. K_b for $NH_3 = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$

Quick Check

1. The amount of water that a weak acid will convert to hydronium ions is insignificant compared to the magnitude of water's concentration (55.6 *M*). The concentration of water is thus assumed to be constant and so is not included in an ICE table.

2. Because the K_a of a weak acid is normally so much greater than K_w, we can assume that the initial $[H_3O^+]$ resulting from the autoionization of water is insignificant compared to the equilibrium $[H_3O^+]$ resulting from the ionization of this weak acid.

3. The value of K_a is small enough compared to the initial concentration of the acid that the percent of the acid which actually ionizes will not significantly change that original concentration. Remember that this assumption is only valid if the percent ionization of the weak acid is ≤ 5 %.

Practice Problems - Calculating pH given K_a and $[HA]_{initial}$

1.

Let $x = [H_3O^+]_{eq}$

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

I	0.50		0	0
C	- x		+ x	+ x
E	0.50 - x		x	x

Assume $0.50 - x \approx 0.50$

$$K_a = \frac{[HCOO^-][H_3O^+]}{[HCOOH]} = \frac{x^2}{0.50} = 1.8 \times 10^{-4}$$

$$x = \sqrt{(1.8 \times 10^{-4})(0.50)} = 9.49 \times 10^{-3}\ M$$

pH = - log (9.49 x 10^{-3}) = **2.02** (2 sig. digits)

2.

Let $x = [H_3O^+]_{eq}$

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

I	0.850		0	0
C	- x		+ x	+ x
E	0.850 - x		x	x

Assume $0.850 - x \approx 0.850$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = \frac{x^2}{0.850} = 1.8 \times 10^{-5}$$

$$x = \sqrt{(1.8 \times 10^{-5})(0.850)} = 3.91 \times 10^{-3}\ M$$

pH = - log (3.91 x 10^{-3}) = **2.41** (2 sig. digits)

3. $[H_3O^+] = 10^{-2.27} = 5.37 \times 10^{-3}\ M$

$$\text{Percent ionization} = \frac{[H_3O^+]_{eq}}{[HA]_{initial}} \times 100\ \% = \frac{5.37 \times 10^{-3}\ M}{0.10\ M} \times 100\ \% = \mathbf{5.4\ \%}$$

The simplifying assumption would not be valid in this case because the percent ionization is greater than 5 %

Practice Problems - Calculating $[HA]_{initial}$, given K_a and pH

1.

Let $x = [H_3C_6H_5O_7]_{initial}$

$[H_3O^+]_{eq} = 10^{-2.50} = 0.00316\ M$

$$H_3C_6H_5O_7 + H_2O \rightleftharpoons H_2C_6H_5O_7^- + H_3O^+$$

I	x		0	0
C	- 0.00316		+ 0.00316	+ 0.00316
E	x – 0.00316		0.00316	0.00316

$$K_a = \frac{[H_3C_6H_5O_7^-]\,[H_3O^+]}{[H_3C_6H_5O_7]} = \frac{(0.00316)^2}{(x - 0.00316)} = 7.1 \times 10^{-4}$$

$$x - 0.00316 = \frac{(0.00316)^2}{7.1 \times 10^{-4}}$$

x = **0.017 *M*** (2 sig. digits)

2.

Let $x = [H_2C_2O_4]_{initial}$

$[H_3O^+]_{eq} = 10^{-0.55} = 0.282\ M$

$$H_2C_2O_4 + H_2O \rightleftharpoons HC_2O_4^- + H_3O^+$$

I	x		0	0
C	- 0.282		+ 0.282	+ 0.282
E	x – 0.282		0.282	0.282

$$K_a = \frac{[HC_2O_4^-]\,[H_3O^+]}{[H_2C_2O_4]} = \frac{(0.282)^2}{(x - 0.282)} = 0.059$$

$$x - 0.282 = \frac{(0.282)^2}{0.059}$$

x = **1.6 *M*** (2 sig. digits)

3.

Let $x = [HNO_2]_{initial}$

$[H_3O^+]_{eq} = 10^{-1.85} = 0.0141\ M$

$$HNO_2 + H_2O \rightleftharpoons NO_2^- + H_3O^+$$

I	x		0	0
C	- 0.0141		+ 0.0141	+ 0.0141
E	x – 0.0141		0.0141	0.0141

$$K_a = \frac{[NO_2^-]\,[H_3O^+]}{[HNO_2]} = \frac{(0.0141)^2}{(x - 0.0141)} = 4.6 \times 10^{-4}$$

$$x - 0.0141 = \frac{(0.0141)^2}{4.6 \times 10^{-4}}$$

x = **0.45 *M*** (2 sig. digits)

Practice Problems - Calculating K_a, given $[HA]_{initial}$ and pH

1.

$[H_3O^+]_{eq} = 10^{-3.000} = 0.00100\ M$

$H_2C_6H_6O_6 + H_2O \rightleftharpoons HC_6H_6O_6^- + H_3O^+$

I	0.**100**		0	0
C	- 0.00**100**		+ 0.00100	+ 0.00100
E	0.0**99**9 (2 s.d.)		0.00100	0.00100

$K_a = \frac{[HC_6H_6O_6^-][H_3O^+]}{[H_2C_6H_6O_6]} = \frac{(0.00100)^2}{0.0999} = \mathbf{1.0 \times 10^{-5}}$ (2 sig. digits)

2.

$[H_3O^+]_{eq} = 10^{-2.75} = 0.00178\ M$

$C_2H_5OCOOH + H_2O \rightleftharpoons C_2H_5OCOO^- + H_3O^+$

I	0.025		0	0
C	- 0.00178		+ 0.00178	+ 0.00178
E	0.0232		0.00178	0.00178

$K_a = \frac{[C_2H_5OCOO^-][H_3O^+]}{[C_2H_5OCOOH]} = \frac{(0.00178)^2}{0.0232} = \mathbf{1.4 \times 10^{-4}}$ (2 sig. digits)

3.

$[H_3O^+]_{eq} = 10^{-2.005} = 0.00989\ M$

$HA + H_2O \rightleftharpoons A^- + H_3O^+$

I	1.50		0	0
C	- 0.00989		+ 0.00989	+ 0.00989
E	1.49		0.00989	0.00989

$K_a = \frac{[A^-][H_3O^+]}{[HA]} = \frac{(0.00989)^2}{1.49} = \mathbf{6.56 \times 10^{-5}}$ (The acid is most likely benzoic acid)

(3 sig. digits)

Quick Check

1.

Weak Base	Conjugate Acid	Appropriate K_a Value	pK_a Value	Calculated K_b	pK_b Value
$HC_2O_4^-$	$H_2C_2O_4$	5.9 x 10^{-2}	**1.23**	1.7 x 10^{-13}	**12.77**
$H_2PO_4^-$	H_3PO_4	7.5 x 10^{-3}	**2.12**	1.3 x 10^{-12}	**11.88**
HPO_4^{2-}	$H_2PO_4^-$	6.2 x 10^{-8}	**7.21**	1.6 x 10^{-7}	**6.79**
NO_2^-	**HNO_2**	**4.6 x 10^{-4}**	**3.34**	**2.2 x 10^{-11}**	**10.66**
$HC_6H_5O_7^{2-}$	**$H_2C_6H_5O_7^-$**	**1.7 x 10^{-5}**	**4.77**	**5.9 x 10^{-10}**	**9.23**
HCO_3^-	**H_2CO_3**	**4.3 x 10^{-7}**	**6.37**	**2.3 x 10^{-8}**	**7.63**
CN^-	**HCN**	**4.9 x 10^{-10}**	**9.31**	**2.0 x 10^{-5}**	**4.69**

Practice Problems - Calculating [OH$^-$] and pH using K_b and [B]$_{initial}$

1.

$$\frac{12.0 \text{ g } N_2H_4}{0.5000 \text{ L}} \times \frac{1 \text{ mol}}{32.0 \text{ g } N_2H_4} = 0.750\ M$$

Let x = [OH$^-$]$_{eq}$

$$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + OH^-$$

I	0.750		0	0
C	- x		+ x	+ x
E	0.750 - x		x	x

Assume 0.750 – x ≈ 0.750

$$K_b = \frac{[N_2H_5^+]\,[OH^-]}{[N_2H_4]} = \frac{x^2}{0.750} = 1.7 \times 10^{-6}$$

x = 1.13 x 10^{-3} M OH$^-$ so pOH = - log (1.13 x 10^{-3}) = 2.947

pH = 14.000 - 2.947 = **11.05** (2 sig. digits)

2.

$$K_b \text{ for } HCOO^- = \frac{K_w}{K_a \text{ for } HCOOH} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-4}} = 5.56 \times 10^{-11}$$

Let $x = [OH^-]_{eq}$

$HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$

	$HCOO^-$	H_2O	$HCOOH$	OH^-
I	0.60		0	0
C	- x		+ x	+ x
E	0.60 - x		x	x

Assume 0.60 – x ≈ 0.60

$$K_b = \frac{[HCOOH][OH^-]}{[HCOO^-]} = \frac{x^2}{0.60} = 5.56 \times 10^{-11}$$

$x = [OH^-] =$ **5.8×10^{-6} *M*** (2 sig. digits) so pOH = - log (5.8×10^{-6}) = **5.24**

pH = 14.00 - 5.24 = **8.76** and $[H_3O^+] = 10^{-8.76}$ = **1.7×10^{-9} *M***

3. Calculate the K_a and pK_a values for

(a) Methylammonium, $CH_3NH_3^+$	(b) Hydrazinium, $N_2H_5^+$
$K_a = \frac{K_w}{K_b \text{ for } CH_3NH_3} = \frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} =$ **2.3×10^{-11}**	$K_a = \frac{K_w}{K_b \text{ for } N_2H_4} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-6}} =$ **5.9×10^{-9}**
pK_a = - log (2.3×10^{-11}) = **10.64**	pK_a = - log (5.9×10^{-9}) = **8.23**

Practice Problems - Calculating $[B]_{initial}$, given K_b and pH (or pOH)

1.

pOH = 14.00 - 11.80 = 2.20

$[OH^-] = 10^{-2.20} = 0.00631\ M$

Let $x = [C_2H_5NH_2]_{initial}$

$$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$$

I	x		0	0
C	- 0.00631		+ 0.00631	+ 0.00631
E	x – 0.00631		0.00631	0.00631

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{(0.00631)^2}{x - 0.00631} = 5.6 \times 10^{-4}$$

$$x - 0.00631 = \frac{(0.00631)^2}{5.6 \times 10^{-4}} \quad \text{so } x = 0.0774\ M$$

$$\frac{0.0774 \text{ mol } C_2H_5NH_2}{\text{L}} \times \frac{45.0 \text{ g } C_2H_5NH_2}{\text{mol}} \times 0.250 \text{ L} = \mathbf{0.87\ g}\ C_2H_5NH_2$$ (2 sig. digits)

2.

$$K_b \text{ for } CN^- = \frac{K_w}{K_a \text{ for HCN}} = \frac{1.00 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.04 \times 10^{-5}$$

pOH = 14.00 - 11.50 = 2.50

$[OH^-] = 10^{-2.50} = 0.00316\ M$

Let x = $[CN^-]_{initial}$

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

I	x		0	0
C	- 0.00316		+ 0.00316	+ 0.00316
E	x - 0.00316		0.00316	0.00316

$$K_b = \frac{[HCN][OH^-]}{[CN^-]} = \frac{(0.00316)^2}{x - 0.00316} = 2.04 \times 10^{-5}$$

$$x - 0.00316 = \frac{(0.00316)^2}{2.04 \times 10^{-5}}$$ so x = **0.49 *M*** (2 sig. digits)

3.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-10}} = 1.0 \times 10^{-4}\ M$$

Let x = $[N_2H_4]_{initial}$

$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + OH^-$

I	x		0	0
C	$- 1.0 \times 10^{-4}$		$+ 1.0 \times 10^{-4}$	$+ 1.0 \times 10^{-4}$
E	$x - 1.0 \times 10^{-4}$		1.0×10^{-4}	1.0×10^{-4}

$$K_b = \frac{[N_2H_5^+][OH^-]}{[N_2H_4]} = \frac{(1.0 \times 10^{-4})^2}{x - 1.0 \times 10^{-4}} = 1.7 \times 10^{-6}$$

$$x - 1.0 \times 10^{-4} = \frac{(1.0 \times 10^{-4})^2}{1.7 \times 10^{-6}}$$ so x = **6.0×10^{-3} *M*** (2 sig. digits)

Practice Problems - Calculating K_b, given $[B]_{initial}$ and pH (or pOH)

1.

pOH = 14.00 - 12.90 = 1.10 $[OH^-]_{eq} = 10^{-1.10} = 0.0794\ M$

$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$

I	0.400		0	0
C	- 0.0794		+ 0.0794	+ 0.0794
E	0.321		0.0794	0.0794

$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = \frac{(0.0794)^2}{0.321} = 0.0196 = \mathbf{0.020}$ (2 sig. digits)

Percent ionization $= \frac{[OH^-]_{eq}}{[CH_3NH_2]_{initial}} \times 100\ \% = \frac{0.0794}{0.400} \times 100\ \% = 19.8\ \% = \mathbf{20.\ \%}$ (2 sig. digits)

This calculated K_b value is greater than the K_b given in the sample problem (4.4×10^{-4}). This indicates that the temperature of the solution is higher than 25° C.

2.

$[OH^-]_{eq} = 10^{-3.90} = 1.26 \times 10^{-4}\ M$

$Mor + H_2O \rightleftharpoons HMor^+ + OH^-$

I	0.010		0	0
C	$- 1.26 \times 10^{-4}$		$+ 1.26 \times 10^{-4}$	$+ 1.26 \times 10^{-4}$
E	0.00987		1.26×10^{-4}	1.26×10^{-4}

$K_b = \frac{[HMor^+][OH^-]}{[Mor]} = \frac{(1.26 \times 10^{-4})^2}{9.87 \times 10^{-3}} = \mathbf{2 \times 10^{-6}}$ (1 sig. digits)

pK_b = - log (1.6 x 10^{-6}) = **5.79**

3.

pOH = 14.00 - 9.84 = 4.16 $[OH^-]_{eq}$ = $10^{-4.16}$ = 6.92 x 10^{-5} *M*

Qui + H_2O ⇌ $HQui^+$ + OH^-

I	0.0015		0	0
C	- 6.92 x 10^{-5}		+ 6.92 x 10^{-5}	+ 6.92 x 10^{-5}
E	1.43 x 10^{-3}		6.92 x 10^{-5}	6.92 x 10^{-5}

$$K_b = \frac{[HQui^+][OH^-]}{[Qui]} = \frac{(6.92 \times 10^{-5})^2}{1.43 \times 10^{-3}} = \mathbf{3.3 \times 10^{-6}} \text{ (2 sig. digits)}$$

Activity 4.5: An Organically Grown Table of Relative Acid Strengths

Compound Name	Formula for Conjugate Acid	Formula for Conjugate Base	Calculate *Ka* for Acid or Conjugate Acid given:
Acids chloroacetic acid	$ClCH_2COOH$	**$ClCH_2COO^-$**	**1.4 x 10^{-3}**
phenylacetic acid	C_7H_7COOH	**$C_7H_7COO^-$**	**4.9 x 10^{-5}**
propanoic acid	C_2H_5COOH	**$C_2H_5COO^-$**	**1.3 x 10^{-5}**
pyruvic acid	C_2H_3OCOOH	**$C_2H_3OCOO^-$**	**2.8 x 10^{-3}**
lactic acid	C_2H_5OCOOH	**$C_2H_5OCOO^-$**	**1.4 x 10^{-4}**
acetylsalicylic acid	$C_8H_7O_2COOH$	**$C_8H_7O_2COO^-$**	**3.6 x 10^{-4}**
glycolic acid	CH_3OCOOH	**CH_3OCOO^-**	**1.5 x 10^{-4}**
glyoxylic acid	$CHOCOOH$	**$CHOCOO^-$**	**3.5 x 10^{-4}**
glyceric acid	$C_2H_5O_2COOH$	**$C_2H_5O_2COO^-$**	**2.9 x 10^{-4}**
Bases			

pyridine	**$C_5H_5NH^+$**	C_5H_5N	**5.9×10^{-6}**
trimethylamine	**$(CH_3)_3NH^+$**	$(CH_3)_3N$	**1.5×10^{-10}**
piperidine	**$C_5H_{10}NH_2^+$**	$C_5H_{10}NH$	**7.8×10^{-12}**
tert-butylamine	**$(CH_3)_3CNH_3^+$**	$(CH_3)_3CNH_2$	**2.0×10^{-11}**
ethanolamine	**$C_2H_5ONH_3^+$**	$C_2H_5ONH_2$	**3.2×10^{-10}**
n-propylamine	**$C_3H_7NH_3^+$**	$C_3H_7NH_2$	**2.9×10^{-11}**

Relative Strengths of Some Organic Acids and Bases

Strength of Acid	Equilibrium Reaction With Water: Acid + H_2O $\rightleftharpoons$ H_3O^+ + Base	K_a Value	Strength of Base
Stronger ↑	$C_2H_3OCOOH + H_2O \rightleftharpoons H_3O^+ + C_2H_3OCOO^-$	2.8×10^{-3}	Weaker ↓
	$ClCH_2COOH + H_2O \rightleftharpoons H_3O^+ + ClCH_2COO^-$	1.4×10^{-3}	
	$C_8H_7O_2COOH + H_2O \rightleftharpoons H_3O^+ + C_8H_7O_2COO^-$	3.6×10^{-4}	
	$CHOCOOH + H_2O \rightleftharpoons H_3O^+ + CHOCOO^-$	3.5×10^{-4}	
	$C_2H_5O_2COOH + H_2O \rightleftharpoons H_3O^+ + C_2H_5O_2COO^-$	2.9×10^{-4}	
	$CH_3OCOOH + H_2O \rightleftharpoons H_3O^+ + CH_3OCOO^-$	1.5×10^{-4}	
	$C_2H_5OCOOH + H_2O \rightleftharpoons H_3O^+ + C_2H_5OCOO^-$	1.4×10^{-4}	
	$C_7H_7COOH + H_2O \rightleftharpoons H_3O^+ + C_7H_7COO^-$	4.9×10^{-5}	
	$C_2H_5COOH + H_2O \rightleftharpoons H_3O^+ + C_2H_5COO^-$	1.3×10^{-5}	
	$C_5H_5NH^+ + H_2O \rightleftharpoons H_3O^+ + C_5H_5N$	5.9×10^{-6}	
	$C_2H_5ONH_3^+ + H_2O \rightleftharpoons H_3O^+ + C_2H_5ONH_2$	3.2×10^{-10}	
	$(CH_3)_3NH^+ + H_2O \rightleftharpoons H_3O^+ + (CH_3)_3N$	1.5×10^{-10}	
	$C_3H_7NH_3^+ + H_2O \rightleftharpoons H_3O^+ + C_3H_7NH_2$	2.9×10^{-11}	
	$(CH_3)_3CNH_3^+ + H_2O \rightleftharpoons H_3O^+ + (CH_3)_3CNH_2$	2.0×10^{-11}	
Weaker	$C_5H_{10}NH_2^+ + H_2O \rightleftharpoons H_3O^+ + C_5H_{10}NH$	7.8×10^{-12}	Stronger

Results and Discussion

1. The K_a values of the weak monoprotic carboxylic acids are within approximately 10^2 times of each other.

2. The strongest acid is pyruvic acid and the strongest base piperidine.

3. The K_a values allow all of the above weak acids to be incorporated into the Table of Relative Acid Strengths from weakest to strongest acids.

4.5 Review Questions

1.

$$\frac{23.0\ \cancel{g\ HCOOH}}{0.5000\ L} \times \frac{1\ mol}{46.0\ \cancel{g\ HCOOH}} = 1.00\ M\ HCOOH$$

Let $x = [H_3O^+]_{eq}$

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

	$HCOOH$	H_2O	$HCOO^-$	H_3O^+
I	1.00		0	0
C	$-x$		$+x$	$+x$
E	$1.00 - x$		x	x

Assume $1.00 - x \approx 1.00$

$$K_a = \frac{[HCOO^-][H_3O^+]}{[HCOOH]} = \frac{x^2}{1.00} = 1.8 \times 10^{-4}$$

$$x = \sqrt{(1.8 \times 10^{-4})(1.00)} = 0.01342\ M$$

pH = - log (0.01342) = **1.87** (2 sig. digits) $[H_3O^+] = 10^{-1.87} =$ **0.013 M**

pOH = 14.00 - 1.87 = **12.13** $[OH^-] = 10^{-12.13} =$ **7.4×10^{-13} M**

2.

$\frac{5.6 \text{ ~~L H}_2\text{S~~}}{2.50 \text{ L}} \times \frac{1 \text{ mol}}{22.4 \text{ ~~L H}_2\text{S~~}} = 0.10\ M\ H_2S$

Let x = $[H_3O^+]_{eq}$

$H_2S + H_2O \rightleftharpoons HS^- + H_3O^+$

I	0.10		0	0
C	- *x*		+ *x*	+ *x*
E	0.10 - *x*		*x*	*x*

Assume $0.10 - x \approx 0.10$

$K_a = \frac{[HS^-][H_3O^+]}{[H_2S]} = \frac{x^2}{0.10} = 9.1 \times 10^{-8}$

$x = \sqrt{(9.1 \times 10^{-8})(0.10)} = 9.54 \times 10^{-5}\ M$

pH = - log (9.54×10^{-5}) = **4.02** (2 sig. digits)

Percent ionization = $\frac{9.54 \times 10^{-5}}{0.10} \times 100\%$ = **0.095 %**

3.

% ionization = $\frac{[H_3O^+]_{eq}}{[HA]_{initial}} \times 100\ \%$ so $[H_3O^+]_{eq} = \frac{(\%\ \text{ionization})\ [HA]_{initial}}{100\ \%}$

$[H_3O^+]_{eq} = \frac{(1.34)(0.100)}{100} = 0.00134\ M$ so pH = - log (0.00134) = **2.873** (3 sig. digits)

As 1.34% < 5%, we can assume that $[HA]_{eq} \approx [HA]_{initial}$ and $[H_3O^+]_{eq} = [A^-]_{eq}$

This allows us to calculate K_a and thereby identify the unknown acid.

$$K_a = \frac{(0.00134)^2}{0.100} = \mathbf{1.80 \times 10^{-5}}$$ (The unknown acid is acetic acid.)

Note: We could have also chosen not to make the assumption because all of the equilibrium concentration values are available to us. The resulting K_a calculation would have yielded the following:

$$K_a = \frac{(0.00134)^2}{0.09866} = \mathbf{1.82 \times 10^{-5}}$$

(The unknown acid would still represent acetic acid given that the K_a values on the Table are given to 2 sig. digits.)

4.

Let $x = [HF]_{initial}$

$[H_3O^+]_{eq} = 10^{-2.00} = 0.0100\ M$

$$HF + H_2O \rightleftharpoons F^- + H_3O^+$$

I	x		0	0
C	- 0.0100		+ 0.0100	+ 0.0100
E	x – 0.0100		0.0100	0.0100

$$K_a = \frac{[F^-][H_3O^+]}{[HF]} = \frac{(0.0100)^2}{(x - 0.0100)} = 3.5 \times 10^{-4}$$

$$x - 0.0100 = \frac{(0.0100)^2}{3.5 \times 10^{-4}}$$ so $x = 0.296\ M$ HF

$$0.296 \frac{\cancel{\text{mol HF}}}{\cancel{\text{L}}} \times \frac{20.0 \text{ g HF}}{\cancel{\text{mol}}} \times 1.5\ \cancel{\text{L}} = \mathbf{8.9 \text{ g HF}}$$ (2 sig. digits)

5.

Let x = $[H_3PO_4]_{initial}$

pH = 14.00 - 12.50 = 1.50 so $[H_3O^+]_{eq} = 10^{-1.50}$ = 0.0316 M

$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

I	x		0	0
C	- 0.0316		+ 0.0316	+ 0.0316
E	x – 0.0316		0.0316	0.0316

$$K_a = \frac{[H_2PO_4^-][H_3O^+]}{[H_3PO_4]} = \frac{(0.0316)^2}{(x - 0.0316)} = 7.5 \times 10^{-3}$$

$$x - 0.0316 = \frac{(0.0316)^2}{7.5 \times 10^{-3}}$$ so x = **0.16 *M* H_3PO_4** (2 sig. digits)

6.

Let x = $[H_2C_2O_4]_{initial}$

$[H_3O^+]_{eq} = 10^{-2.35}$ = 0.00447 M

$H_2C_2O_4 + H_2O \rightleftharpoons HC_2O_4^- + H_3O^+$

I	x		0	0
C	- 0.00447		+ 0.00447	+ 0.00447
E	x – 0.00447		0.00447	0.00447

$$K_a = \frac{[HC_2O_4^-][H_3O^+]}{[H_2C_2O_4]} = \frac{(0.00447)^2}{(x - 0.00447)} = 5.9 \times 10^{-2}$$

$$x - 0.00447 = \frac{(0.00447)^2}{5.9 \times 10^{-2}}$$ so x = 0.00481 M $H_2C_2O_4$

$0.00481 \frac{\cancel{mol\ H_2C_2O_4}}{\cancel{L}} \times \frac{90.0\ g\ H_2C_2O_4}{\cancel{mol}} \times 0.250\ \cancel{L}$ = **0.11 g $H_2C_2O_4$** (2 sig. digits)

7.

$[H_3O^+]_{eq} = 10^{-4.62} = 2.40 \times 10^{-5}\ M$

$HClO + H_2O \rightleftharpoons ClO^- + H_3O^+$

I	0.020		0	0
C	$- 2.40 \times 10^{-5}$		$+ 2.40 \times 10^{-5}$	$+ 2.40 \times 10^{-5}$
E	≈ 0.020		2.40×10^{-5}	2.40×10^{-5}

$K_a = \frac{[ClO^-][H_3O^+]}{[HClO]} = \frac{(2.40 \times 10^{-5})^2}{0.020}$ = $\mathbf{2.9 \times 10^{-8}}$ (2 sig. digits)

8.

pH = 14.00 - 11.34 = 2.66 so $[H_3O^+]_{eq} = 10^{-2.66} = 0.00219\ M$

$C_6H_5CH_2COOH + H_2O \rightleftharpoons C_6H_5CH_2COO^- + H_3O^+$

I	0.100		0	0
C	- 0.00219		+ 0.00219	+ 0.00219
E	0.0978		0.00219	0.00219

$K_a = \frac{[C_6H_5CH_2COO^-][H_3O^+]}{[C_6H_5CH_2COOH]} = \frac{(0.00219)^2}{0.0978}$ = $\mathbf{4.9 \times 10^{-5}}$ (2 sig. digits)

9.

Conjugate Acid	**Conjugate Base**	**K_a for acid**	**pK_a**	**K_b for base**	**pK_b**
HNO_2	NO_2^-	**4.6×10^{-4}**	**3.34**	**2.2×10^{-11}**	**10.66**
H_2O_2	**HO_2^-**	**2.4×10^{-12}**	**11.62**	**4.2×10^{-3}**	**2.38**
C_6H_5OH	$C_6H_5O^-$	**1.3×10^{-10}**	**9.89**	**7.7×10^{-5}**	**4.11**
HSO_4^-	**SO_4^{2-}**	**1.2×10^{-2}**	**1.92**	**8.3×10^{-13}**	**12.08**

10.

Conjugate Acid	Conjugate Base	K_a for acid	pK_a	K_b for base	pK_b
$H_2PO_4^-$	HPO_4^{2-}	**6.2×10^{-8}**	**7.21**	**1.6×10^{-7}**	**6.79**
$H_2C_6H_5O_7^-$	**$HC_6H_5O_7^{2-}$**	**1.7×10^{-5}**	**4.77**	**5.9×10^{-10}**	**9.23**
H_3BO_3	$H_2BO_3^-$	**7.3×10^{-10}**	**9.14**	**1.4×10^{-5}**	**4.86**
HCO_3^-	**CO_3^{2-}**	**5.6×10^{-11}**	**10.25**	**1.8×10^{-4}**	**3.75**

11.

$$\frac{5.6 \text{ } \cancel{\text{L NH}_3}}{0.750 \text{ L}} \times \frac{1 \text{ mol}}{22.4 \text{ } \cancel{\text{L NH}_3}} = 0.333 \text{ } M \text{ NH}_3$$

Let $x = [OH^-]_{eq}$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

	NH_3	H_2O	NH_4^+	OH^-
I	0.333		0	0
C	$-x$		$+x$	$+x$
E	$0.333 - x$		x	x

Assume $0.333 - x \approx 0.333$

$$K_b \text{ for } NH_3 = \frac{K_w}{K_a \text{ for } NH_4^+} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{x^2}{0.333} = 1.8 \times 10^{-5}$$

$x = 2.45 \times 10^{-3}$ M OH^- so pOH = $-\log(2.45 \times 10^{-3}) = 2.611$

pH = 14.000 - 2.611 = **11.39** (2 sig. digits)

12.

Let x = $[OH^-]_{eq}$

$(CH_3)_2CHNH_2$ + H_2O $\rightleftharpoons$ $(CH_3)_2CHNH_3^+$ + OH^-

I	0.60		0	0
C	- x		+ x	+ x
E	0.60 - x		x	x

Assume 0.60 – x ≈ 0.60

$$K_b = \frac{[(CH_3)_2CHNH_3^+][OH^-]}{[(CH_3)_2CH\ NH_2]} = \frac{x^2}{0.60} = 4.7 \times 10^{-4}$$

x = 0.0168 M OH^- so pOH = - log (0.0168) = **1.77** (2 sig. digits)

pH = 14.00 - 1.77 = **12.23**

13.

pOH = 14.00 - 10.00 = 4.00

$[OH^-]$ = 1.00×10^{-4} M

Let x = $[SO_3^{2-}]_{initial}$

SO_3^{2-} + H_2O $\rightleftharpoons$ HSO_3^- + OH^-

I	x		0	0
C	- 1.00×10^{-4}		+ 1.00×10^{-4}	+ 1.00×10^{-4}
E	x - 1.00×10^{-4}		1.00×10^{-4}	1.00×10^{-4}

$$K_b \text{ for } SO_3^{2-} = \frac{K_w}{K_a \text{ for } HSO_3^-} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7}$$

$$K_b = \frac{[HSO_3^-][OH^-]}{[SO_3^{2-}]} = \frac{(1.00 \times 10^{-4})^2}{x - 1.00 \times 10^{-4}} = 1.0 \times 10^{-7}$$

$x - 1.00 \times 10^{-4} = \frac{(1.00 \times 10^{-4})^2}{1.0 \times 10^{-7}}$ so x = **0.10 *M*** (2 sig digits)

14.

$[OH^-] = 10^{-2.50} = 0.00316\ M$

Let $x = [(CH_3)_3N]_{initial}$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

I	x		0	0
C	- 0.00316		+ 0.00316	+0.00316
E	x - 0.00316		0.00316	0.00316

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]} = \frac{(0.00316)^2}{x - 0.00316} = 6.3 \times 10^{-5}$$

$x - 0.00316 = \frac{(0.00316)^2}{6.3 \times 10^{-5}}$ so $x = 0.162\ M\ (CH_3)_3N$

$2.5\ \cancel{L} \times 0.162\ \frac{\cancel{mol}}{\cancel{L}} \times \frac{22.4\ L\ (CH_3)_3N}{\cancel{mol}}$ = **9.1 L** (2 sig. digits)

15.

$[OH^-] = 10^{-2.20} = 0.00631\ M$

$$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$$

I	0.100		0	0
C	- 0.00631		+ 0.00631	+ 0.00631
E	0.0937		0.00631	0.00631

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{(0.00631)^2}{0.0937} = \mathbf{4.2 \times 10^{-4}}$$ (2 sig. digits)