NOTE: some of the Ka values here may be a little different from the ones on your pKa table

Acid Base #2
Ka, Kb, Kw

→ handout on canvas TRUST + USE yours, method is the same.

1) a) Determine the pH of a 0.500 M solution of nitrous acid.

$$HNO_{2\ (aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + NO_2^-_{(aq)}$$

	HNO_2		H_3O^+	NO_2^-
I	0.5 M	-	-	-
C	-x		+x	+x
E	0.5 - x	assume x << 0.5	x	x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 5.6 \times 10^{-4} = \frac{x^2}{0.5}$$

↓ value from your table for HNO_2

$2.3 \times 10^{-4} = x^2$

$1.52 \times 10^{-2}\ M = x = [H_3O^+]$
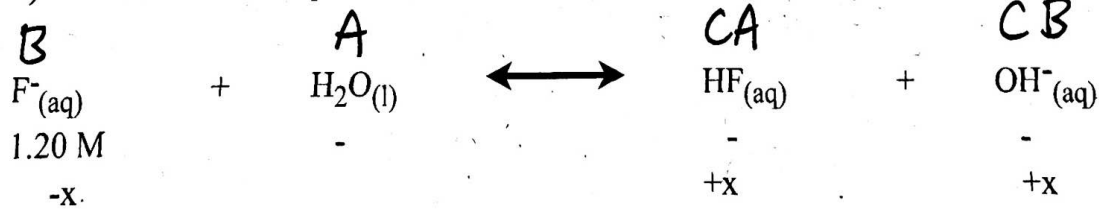
$$pH = -log[H_3O^+] = -log(1.52 \times 10^{-2})$$
$$pH = 1.82$$

b) Calculate the % ionization of this solution.

$$\% \text{ ionization} = \frac{1.52 \times 10^{-2}\ M}{0.500\ M} = 3.03\ \%$$

2) a) Determine the pH of 1.20 M solution of sodium fluoride.

$$\underset{B}{F^-_{(aq)}} + \underset{A}{H_2O_{(l)}} \longleftrightarrow \underset{CA}{HF_{(aq)}} + \underset{CB}{OH^-_{(aq)}}$$

	F^-		HF	OH^-
I	1.20 M	-	-	-
C	-x		+x	+x
E	1.20 - x	assume x <<1.20	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-4}} = \frac{[HF][OH^-]}{[F^-]} = 2.86 \times 10^{-11} = \frac{x^2}{1.20}$$

$3.43 \times 10^{-11} = x^2$

6.6×10^{-4}
↑ value from your canvas page for HF

$$[OH^-] = 5.86 \times 10^{-6}\ M = x$$
$$pOH = -\log[OH^-] = 5.23$$
$$pH = 14.00 - pOH = 14.00 - 5.23$$
$$pH = 8.77$$

b)   Calculate the % ionization of this solution.

$$\% \text{ ionization} = \frac{5.86 \times 10^{-6}\ M}{1.20\ M} = 0.000488\ \%\ \text{ionization}$$

3)   a)   Calculate the concentrations of **all** species in the equilibrium formed when 1.00 M vinegar is dissolved in water.

  b)   Draw the concentration vs time graph for all species from the moment that the vinegar is added until several minutes after equilibrium is established.

  c)   Continue the concentration vs time graphs after a small amount of HCl is added

  d)   Continue the concentration vs time graphs after a small amount of NaOH is added

a) $HOAc_{(aq)}$ + $H_2O_{(l)}$ $\longleftrightarrow$ $H_3O^+_{(aq)}$ + $OAc^-_{(aq)}$

I	1.00 M	-	-	-
C	-x		+x	+x
E	1.00 - x	assume x << 1.0	x	x
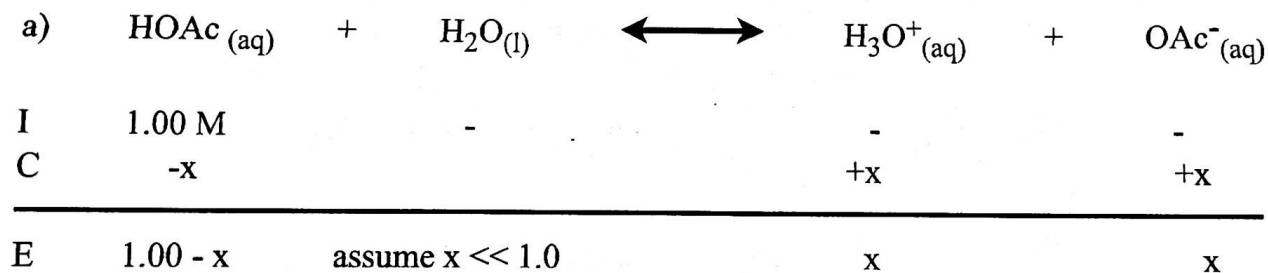
$$Ka = \frac{[H_3O^+][OAc^-]}{[HOAc]} = 1.8 \times 10^{-5} = \frac{x^2}{1.0}$$

$$1.8 \times 10^{-5} = x^2$$
$$4.24 \times 10^{-3}\ M = x$$

$$[HOAc] = 1.00 - x = 0.99576\ M = 1.00\ M$$

$$[H_3O^+] = [OAc^-] = x = 4.2 \times 10^{-3}\ M$$

4)   Calculate the $[H_3O^+]$ for each solution;

  a)   1.0 M H_3PO_4

  b)   1.0 M $H_2PO_4^-$

  c)   1.0 M HPO_4^{2-}

  d)   Add up all the concentrations of a), b) and c). What pattern do you notice?

a) $H_3PO_{4\,(aq)}$ + $H_2O_{(l)}$ $\longleftrightarrow$ $H_3O^+_{\,(aq)}$ + $H_2PO_4^-_{\,(aq)}$

I	1.0 M	-	-	-
C	-x		+x	+x

E   1.0 - x   assume x << 1.0   x   x
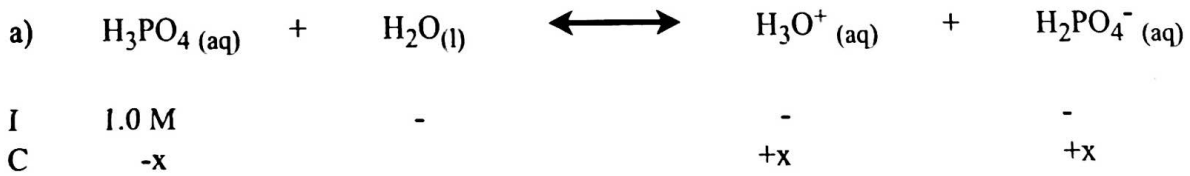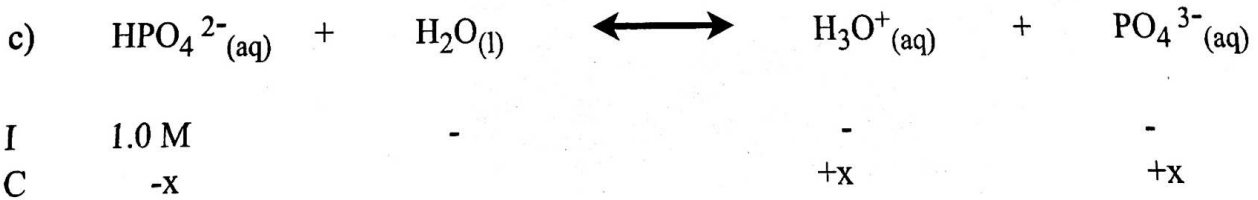
$$Ka = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]} = 7.5 \times 10^{-3} = \frac{x^2}{1.0} \qquad 7.5 \times 10^{-3} = x^2$$

$$8.7 \times 10^{-2}\,M = x = [H_3O^+]$$

\

b) $H_2PO_4^-_{\,(aq)}$ + $H_2O_{(l)}$ $\longleftrightarrow$ $H_3O^+_{\,(aq)}$ + $HPO_4^{2-}_{\,(aq)}$

I	1.0 M	-	-	-
C	-x		+x	+x

E   1.0 - x   assume x << 1.0   x   x
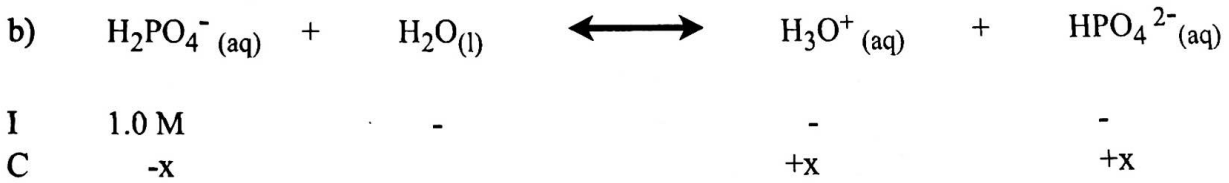
$$Ka = \frac{[H_3O^+][HPO_4^{2-}]}{[H_2PO_4^-]} = 6.2 \times 10^{-8} = \frac{x^2}{1.0} \qquad 6.2 \times 10^{-8} = x^2$$

$$2.5 \times 10^{-4}\,M = x = [H_3O^+]$$

c) $HPO_4^{2-}_{\,(aq)}$ + $H_2O_{(l)}$ $\longleftrightarrow$ $H_3O^+_{\,(aq)}$ + $PO_4^{3-}_{\,(aq)}$

I	1.0 M	-	-	-
C	-x		+x	+x

E   1.0 - x   assume x << 1.0   x   x

$$Ka = \frac{[H_3O^+][PO_4^{3-}]}{[HPO_4^{2-}]} = 2.2 \times 10^{-13} = \frac{x^2}{1.0} \qquad 2.2 \times 10^{-13} = x^2$$

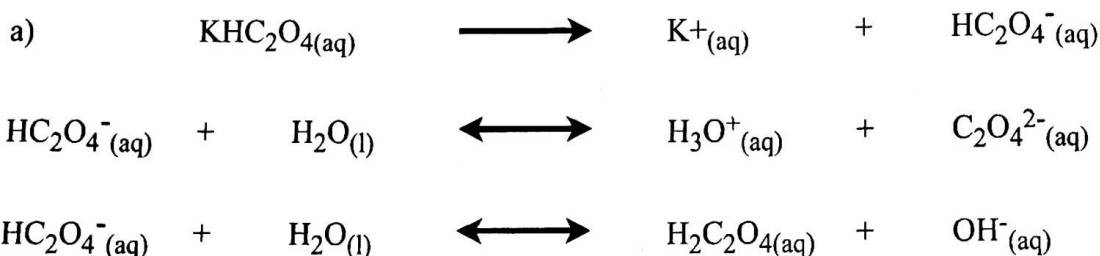$$4.7 \times 10^{-7}\,M = x = [H_3O^+]$$

**Note: There are hydroxides also forming in this sequence, but we are only interested in the hydronium production.

d) $[H_3O^+] = 8.66 \times 10^{-2} + 2.49 \times 10^{-4} + 4.69 \times 10^{-7}$

$[H_3O^+] = 0.086\ 6 + 0.000\ 249 + 0.000\ 000\ 469 = 0.086849 = 0.087\ M$

The total (to 2 sig figs) is exactly the same as the $[H_3O^+]$ for the first ionization.

5) A solution of 0.6500 M potassium binoxalate is formed.
   a) Is this solution acidic or basic? Prove with calculations.
   b) Calculate the pH and pOH.

a) $KHC_2O_{4(aq)} \longrightarrow K^+_{(aq)} + HC_2O_4^-{}_{(aq)}$

$HC_2O_4^-{}_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + C_2O_4^{2-}{}_{(aq)}$

$HC_2O_4^-{}_{(aq)} + H_2O_{(l)} \longleftrightarrow H_2C_2O_{4(aq)} + OH^-_{(aq)}$

$Ka = \dfrac{[H_3O^+][C_2O_4^{2-}]}{[HC_2O_4^-]} = 6.4 \times 10^{-5}$

$Kb = \dfrac{Kw}{Ka} = \dfrac{[H_2C_2O_4][OH^-]}{[HC_2O_4^-]} = \dfrac{1.0 \times 10^{-14}}{5.9 \times 10^{-2}}$

$Ka = 6.4 \times 10^{-5}$

$Kb = 1.69 \times 10^{-13}$

Since Ka > Kb the ionization reaction happens more frequently than the hydrolysis reaction, and the resulting solution contains more H_3O^+ than OH^-. Since $[H_3O^+] > [OH^-]$, this system is acidic.

b) $HC_2O_4^-{}_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + C_2O_4^{2-}{}_{(aq)}$
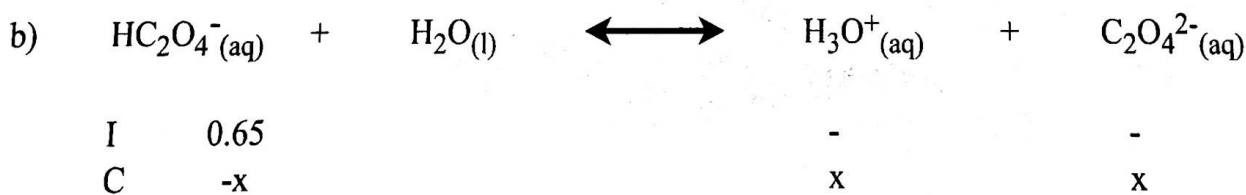
	$HC_2O_4^-$	H_2O	H_3O^+	$C_2O_4^{2-}$
I	0.65		-	-
C	-x		x	x
E	0.65-x	assume x << 0.65	x	x

$Ka = \dfrac{[H_3O^+][C_2O_4^{2-}]}{[HC_2O_4^-]} = \dfrac{x^2}{0.65} = 6.4 \times 10^{-5}$

$x^2 = 4.16 \times 10^{-5}$

$x = 6.45 \times 10^{-3}\ M = [H_3O^+]$

Assumption not valid:

$x^2 = (6.4 \times 10^{-5})(0.65 - .00645)$

$x^2 = 4.12 \times 10^{-5}$

$x_2 = 0.00642\ M$

$pH = -\log[H_3O^+] = 2.19$

$pOH = 14 - pH = 11.81$

6) 110.00 g of lithium bicarbonate is dissolved in 2.00 L of water.
   a)   Is this solution acidic or basic? Prove with calculations.
   b)   Calculate the pH and pOH.

a) $LiHCO_{3(aq)}$ $\longrightarrow$ $Li^+_{(aq)}$ $+$ $HCO_3^-_{(aq)}$

$HCO_3^-_{(aq)}$ $+$ $H_2O_{(l)}$ $\longleftrightarrow$ $H_3O^+_{(aq)}$ $+$ $CO_3^{2-}_{(aq)}$

$HCO_3^-_{(aq)}$ $+$ $H_2O_{(l)}$ $\longleftrightarrow$ $H_2CO_{3(aq)}$ $+$ $OH^-_{(aq)}$

$Ka = \dfrac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]} = 5.6 \times 10^{-11}$

$Kb = \dfrac{Kw}{Ka} = \dfrac{[H_2CO_3][OH^-]}{[HCO_3^-]} = \dfrac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}}$

$Ka = 5.6 \times 10^{-11}$

$Kb = 2.33 \times 10^{-8}$

Since $Ka < Kb$ the hydrolysis reaction happens more frequently than the ionization reaction, and the resulting solution contains more H_3O^+ than OH^-. Since $[H_3O^+] < [OH^-]$, this system is basic.

b) $1 \times Li = 6.9$ amu $(110.0 \text{ g } LiHCO_3)\dfrac{(1 \text{ mol})}{(67.9 \text{ g})} = 1.62$ mol of $LiHCO_3$
     $1 \times H = 1.0$ amu
     $1 \times C = 12.0$ amu
     $\underline{3 \times O = 48.0}$ amu $[HCO_3^-] = 1.62 \text{ mol}/2.00 \text{ L} = 0.81 \text{ M } HCO_3^-$
     67.9 amu

	$HCO_3^-_{(aq)}$	$+$	$H_2O_{(l)}$	$\longleftrightarrow$	$H_2CO_{3(aq)}$	$+$	$OH^-_{(aq)}$
I	0.81				-		-
C	-x				x		x
E	0.81-x		assume x << 0.81		x		x
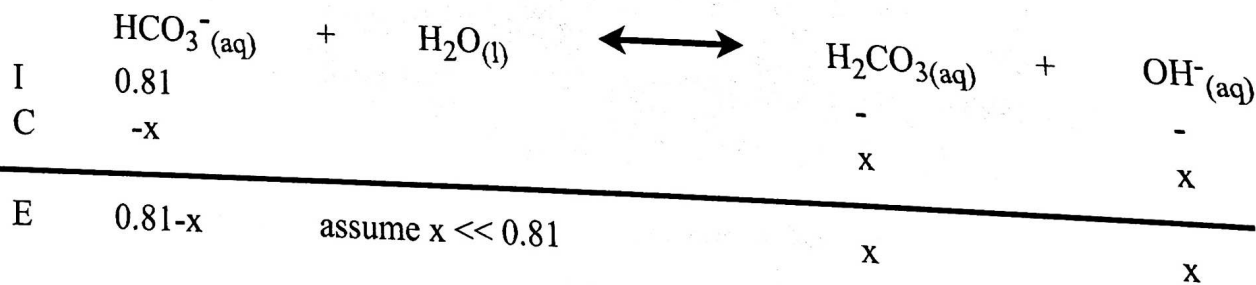
$Kb = \dfrac{[H_2CO_3][OH^-]}{[HCO_3^-]} = \dfrac{x^2}{0.81} = 2.33 \times 10^{-8}$ $x^2 = 1.89 \times 10^{-8}$

$x = 1.37 \times 10^{-4} \text{ M} = [OH^-]$

$pOH = -\log[OH^-] = 3.86$

$pH = 14 - pOH = 14 - 3.86 = 10.14$

7) Explain the "leveling effect". What is the consequence of the leveling effect?

The leveling effect refers to the approximation that "strong" acids and bases dissociate completely in solution. Since none of the molecular strong acids/bases remain in solution, only their dissociation products, H_3O^+ and OH^- respectively and their conjugates actually exist in the solution. Consequently, you can put an acid that is stronger than H_3O^+ in water, but the result is the formation of H_3O^+. Similarly, you can place a very strong base in water and the result will be OH^-. Therefore, all strong acids and bases get "leveled" and the strongest acid that can exist in water is hydronium and the strongest base that can continue to exist in water is hydroxide.

8) Determine the pH of a saturated solution of magnesium carbonate. → **Ksp**

$$MgCO_{3(s)} \longrightarrow Mg^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$$

y   y   y

$$Ksp = [Mg^{2+}][CO_3^{2-}] = 6.8 \times 10^{-6} = y^2$$
$$[Mg^{2+}] = [CO_3^{2-}] = y = 2.61 \times 10^{-3} M$$

$$CO_3^{2-}_{(aq)} + H_2O_{(l)} \longleftrightarrow HCO_3^-_{(aq)} + OH^-_{(aq)}$$

	CO_3^{2-}	H_2O	HCO_3^-	OH^-
I	0.00261			
C	-x		+x	+x
E	0.00261-x		x	x

$$Kb = \frac{Kw}{Ka} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-11}} = 1.8 \times 10^{-4} = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]} = \frac{x^2}{0.00261-x}$$

Assume x << 0.00261 $x^2 = 4.69 \times 10^{-7}$
The assumption is NOT valid

therefore progressively approximate

$x_1 = 6.85 \times 10^{-4}$
$x_2 = 5.88 \times 10^{-4}$
$x_3 = 6.03 \times 10^{-4}$
$x_4 = 6.01 \times 10^{-4}$

$[OH^-] = 6.01 \times 10^{-4} M$ $pOH = -\log[OH^-] = 3.22$

$$pH = 14 - pOH = 14 - 3.22$$

$$pH = 10.78$$