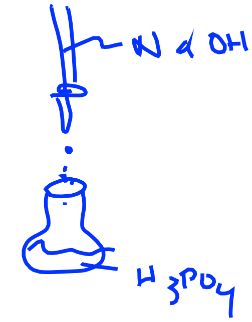

Chem E-1a
Friday Review Problems
Chapter 4

1. Your lab TF gives you a 50.0 mL sample of a phosphoric acid solution of unknown concentration. You are asked to determine the concentration of this solution by titrating it with a 0.731-molar sodium hydroxide solution.

$0.731 \frac{mol}{L}$ $\qquad \rightarrow H_3PO_4$

   a) Write a complete, balanced equation for the reaction of phosphoric acid with sodium hydroxide.

NaOH

$$H_3PO_4(aq) + 3NaOH(aq) \longrightarrow 3H_2O(l) + Na_3PO_4(aq)$$

   b) Write a net ionic equation for this reaction.

COMPLETE IONIC:

$$H_3PO_4(aq) + \cancel{3Na^+(aq)} + 3OH^-(aq) \longrightarrow 3H_2O(l) + \cancel{3Na^+(aq)} + PO_4^{3-}(aq)$$

NET IONIC:

$$H_3PO_4(aq) + 3OH^-(aq) \longrightarrow 3H_2O(l) + PO_4^{3-}(aq)$$

   c) You add 200. mL of distilled water and a couple of drops of an appropriate indicator to the phosphoric acid solution and then begin titrating. The endpoint is reached when 38.3 mL of sodium hydroxide has been added. Calculate the molarity of the unknown phosphoric acid solution.

USE MOLARITY AS CONVERSION FACTOR

$$38.3 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.731 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol } H_3PO_4}{3 \text{ mol NaOH}} = 0.00933 \text{ mol } H_3PO_4 \text{ INITIALLY}$$

$$\text{INITIAL MOLARITY OF } H_3PO_4 = \frac{\text{mol } H_3PO_4}{\text{L } H_3PO_4}$$

50 ML IN L

$$= \frac{0.00933 \text{ mol } H_3PO_4}{0.050 \text{ L } H_3PO_4}$$

$$= 0.187 \text{ mol/L}$$

$$= \boxed{0.187 \text{ M } H_3PO_4}$$

"H_2A"

2. When 0.50 grams of an unknown diprotic acid is dissolved in 100.0 mL of water and titrated with 0.10M NaOH, it takes 52.3 mL of base to completely neutralize the acid. What is the molar mass of the acid?

CALCULATE g H_2A = 0.50 g

UNITS OF MOLAR MASS = $\frac{g}{mol}$

CALCULATE MOL H_2A

BALANCED REACTION:

$$H_2A\,(aq) + 2\,NaOH\,(aq) \longrightarrow 2\,H_2O\,(l) + Na_2A\,(aq)$$

$$52.3 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.100 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol } H_2A}{2 \text{ mol NaOH}} = 0.002615 \text{ mol } H_2A$$

$$\text{MOLAR MASS OF } H_2A = \frac{0.50 \text{ g}}{0.002615 \text{ mol}}$$

$$= \boxed{191.2 \text{ g/mol}}$$

3. Write the oxidation state of each atom in the following substances:

a) $\overset{+1}{H}\overset{0}{O}\overset{-1}{F}$

b) $\boxed{\overset{+1}{H}}\boxed{\overset{+4}{C}}\boxed{\overset{-2}{O_3}}^{-}$

$H = +1$

$C = +4$

$O = -2$

$H + C + 3(O) = -1$

$1 + C + 3(-2) = -1$

$C = 4$

c) $\overset{+1}{H_3}\overset{+5}{P}\overset{-2}{O_4}$

$3(H) + P + 4(O) = 0$

$3(1) + P + 4(-2) = 0$

$P = 5$

d) $\overset{+2}{Fe}\,\underbrace{\overset{+4}{S}\overset{-2}{O_3}}$

Fe^{2+} SO_3^{2-} = SULFITE ION

$$2\overset{0}{Cu}(s) + \overset{0}{O_2}(g) \rightarrow 2\overset{+2}{Cu}\overset{-2}{O}(s)$$

4. Consider the following balanced equation:

$$Cr_2O_3 + 2\ CCl_4 \longrightarrow 2\ CrCl_3 + CO_2 + COCl_2$$

A reaction vessel is filled with 2.3 moles of Cr_2O_3 and 3.8 moles of CCl_4 and this reaction proceeds to completion. Determine the number of moles of each reactant and product present once this reaction is complete.

UNITS: # OF MOLES

REACTION:	Cr_2O_3 +	$2\ CCl_4$ $\longrightarrow$	$2\ CrCl_3$ +	CO_2 +	$COCl_2$
INITIAL:	2.3	3.8	0	0	0
CHANGE:	$-x$	$-2x$	$+2x$	$+x$	$+x$
END:	$2.3-x$	$3.8-2x$	$2x$	x	x
	$=2.3-1.9$	$=3.8-2(1.9)$	$=2(1.9)$		
	$=0.4$	$=0$	$=3.8$	$=1.9$	$=1.9$

PLUG IN $x=1.9$

SOLVE FOR X: IF REACTION GOES TO COMPLETION

FOR EACH REACTANT CALCULATE: $\frac{\#\text{ OF MOLES}}{\text{STOICH. COEFF.}}$

SMALLEST # = X

Cr_2O_3: $\frac{2.3}{1} = \cancel{2.3}$

CCl_4: $\frac{3.8}{2} = 1.9$

SMALLEST

SO $x = 1.9$

AND CCl_4 IS LIMITING REAGENT

Pb^{2+} AsO_4^{3-}

5. Lead (II) arsenate, $Pb_3(AsO_4)_2$, is a poisonous white powder which was once commonly used as an insecticide. When 250. mL of 0.500-molar potassium arsenate is mixed with 100. mL of 0.750-molar lead (II) nitrate, a precipitate of lead arsenate is formed.

(s)

a) Write a complete, balanced equation for this reaction.

$$2K_3AsO_4\,(aq) + 3Pb(NO_3)_2\,(aq) \longrightarrow Pb_3(AsO_4)_2\,(s) + 6KNO_3\,(aq)$$

b) Write a net ionic equation for this reaction.

$$3Pb^{2+}(aq) + 2AsO_4^{3-}(aq) \longrightarrow Pb_3(AsO_4)_2\,(s)$$

c) Assuming this reaction proceeds to completion, determine the mass of lead (II) arsenate which would be formed and calculate the molarity of the lead, potassium, nitrate, and arsenate ions in this solution at the end of the reaction.

$$250\text{ mL } K_3AsO_4 \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.500\text{ mol}}{1\text{ L}} = 0.125\text{ mol } K_3AsO_4$$

$$0.125\text{ mol } K_3AsO_4 \times \frac{1\text{ mol } AsO_4^{3-}}{1\text{ mol } K_3AsO_4} = 0.125\text{ mol } AsO_4^{3-}$$

$$0.125\text{ mol } K_3AsO_4 \times \frac{3\text{ mol } K^+}{1\text{ mol } K_3AsO_4} = 0.375\text{ mol } K^+$$

$$100\text{ mL } Pb(NO_3)_2 \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.750\text{ mol}}{1\text{ L}} = 0.075\text{ mol } Pb(NO_3)_2$$

$$\Downarrow$$

0.075 mol Pb^{2+}

0.150 mol NO_3^-

FIND x:

$\frac{0.075}{3} = 0.025$ SMALLER

$x = 0.025$

$\frac{0.125}{2} = 0.0625$

5. c) (continued – space for additional work.)

\# OF MOLES

R:	$3Pb^{2+}(aq)$ +	$2AsO_4^{3-}(aq)$ $\longrightarrow$	$Pb_3(AsO_4)_2(s)$
I:	0.075	0.125	0
C:	$-3x$	$-2x$	$+x$
E:	$0.075-3x$	$0.125-2x$	x
	$= 0.075 - 3(0.025)$	$= 0.125 - 2(0.025)$	
	$= 0$	$= 0.075$	$= 0.025$

FIND FINAL CONCENTRATIONS OF IONS:

TOTAL VOLUME = 250 mL + 100 mL = 350 mL

= 0.350 L

$[Pb^{2+}] = \frac{0 \text{ mol}}{0.350 \text{ L}} = 0 \text{ M}$

$[AsO_4^{3-}] = \frac{0.075 \text{ mol}}{0.350 \text{ L}} = 0.214 \text{ M}$

$[K^+] = \frac{0.375 \text{ mol}}{0.350 \text{ L}} = 1.07 \text{ M}$

$[NO_3^-] = \frac{0.150 \text{ mol}}{0.350 \text{ L}} = 0.429 \text{ M}$

SPECTATOR IONS SO # OF MOLES DOES NOT CHANGE!

FIND MASS OF $Pb_3(AsO_4)_2(s)$:

$0.025 \text{ mol } Pb_3(AsO_4)_2(s) \times \frac{899.4 \text{ g } Pb_3(AsO_4)_2}{1 \text{ mol } Pb_3(AsO_4)_2} = 22.49 \text{ g } Pb_3(AsO_4)_2(s)$