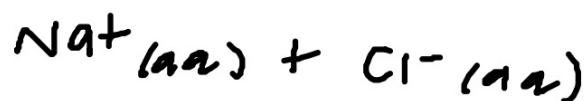


Study session, Dr. Chemist

Nots

Strong electrolytes = ions (aq) in net ionic eqn.

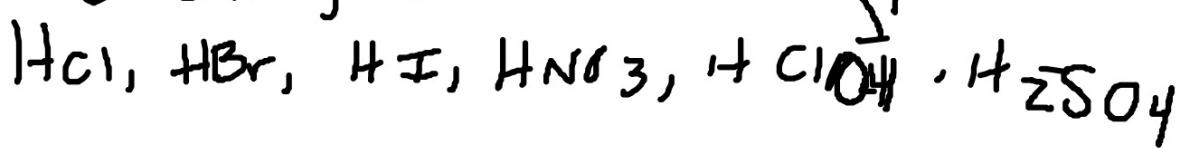
① Ionic (aq) $\text{NaCl}(\text{aq})$



② Strong acids



6 Strong acids



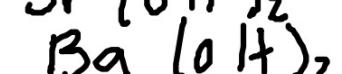
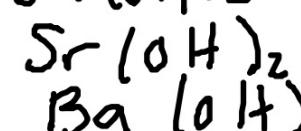
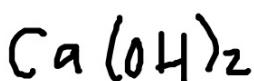
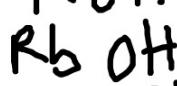
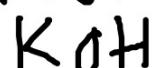
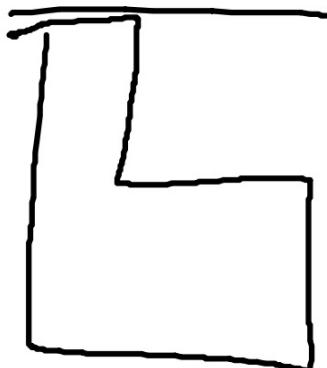
③ Strong bases



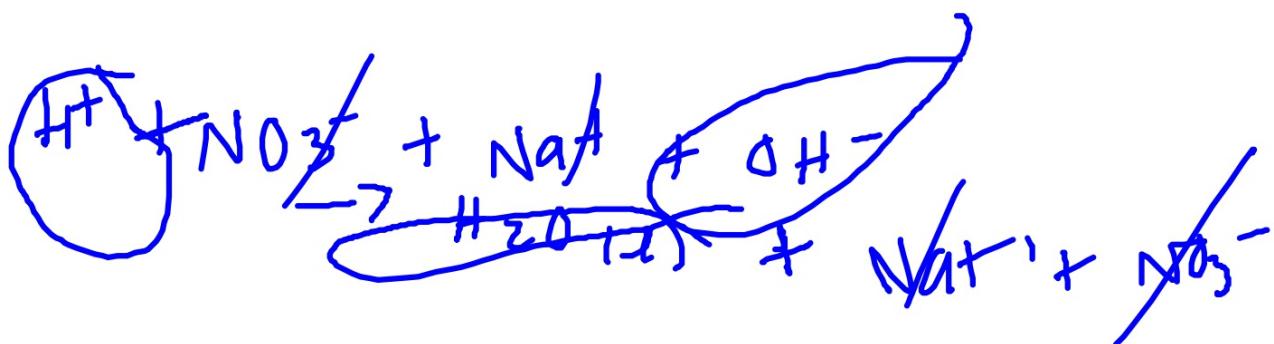
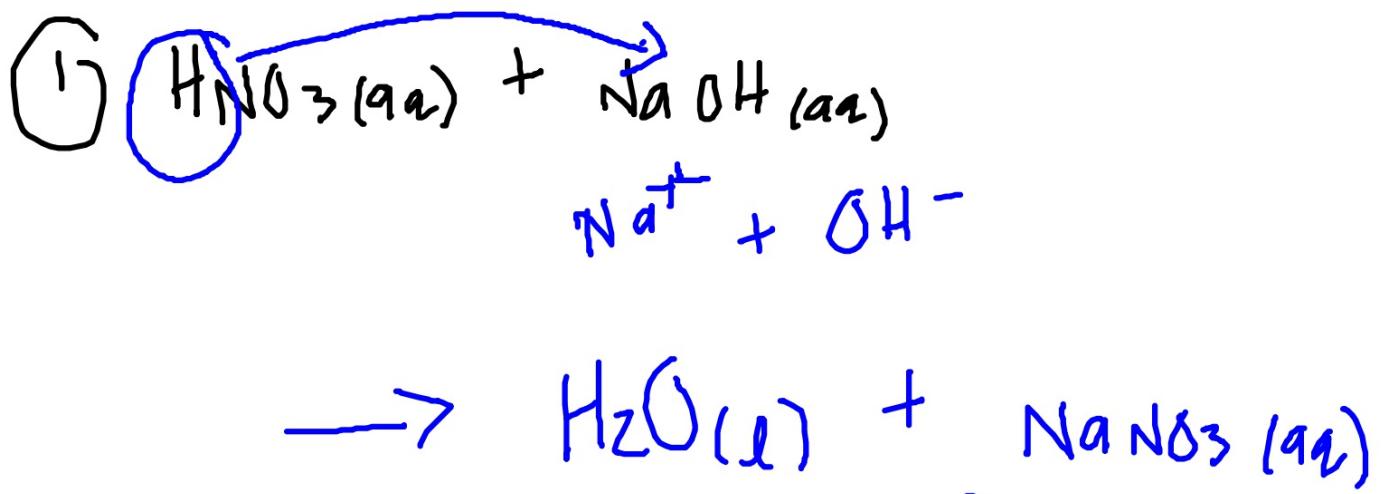
weak base
 $\text{NH}_3\text{(aq)}$

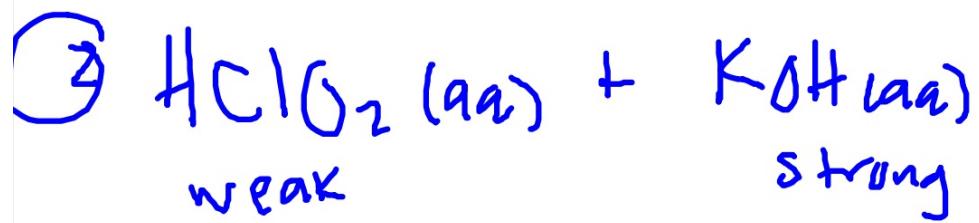


$$S_{SB} = \text{Grp 1} + 2$$

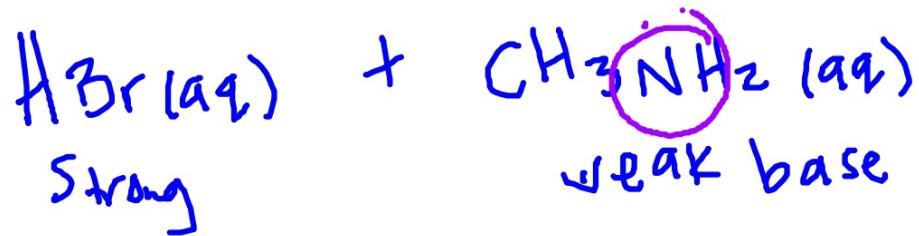


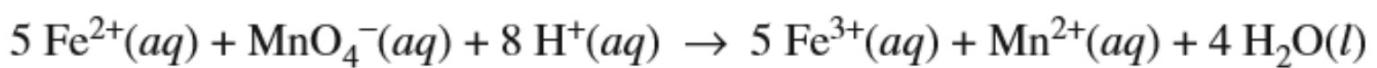
Examples





3





The student titrates a 10.0 mL sample of the $\text{Fe}^{2+}(aq)$ solution. Calculate the value of $[\text{Fe}^{2+}]$ in the solution if it takes 17.48 mL of added 0.0350 M $\text{KMnO}_4(aq)$ to reach the equivalence point of the titration.

$$[\text{Fe}^{2+}] = ? \text{ M} = \frac{\frac{? \text{ mol Fe}^{2+}}{0.01 \text{ L}}}{0.01 \text{ L}} = ? \text{ mol Fe}^{2+}$$

react completely

$$(0.035 \text{ M } \text{KMnO}_4)(0.01748 \text{ L}) \\ = 6.118 \times 10^{-4} \text{ mol } \text{KMnO}_4$$

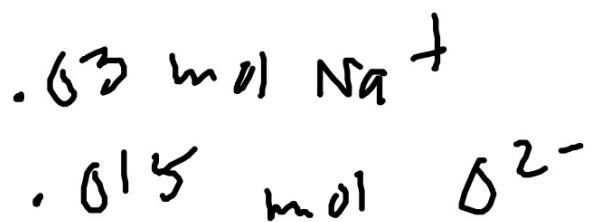
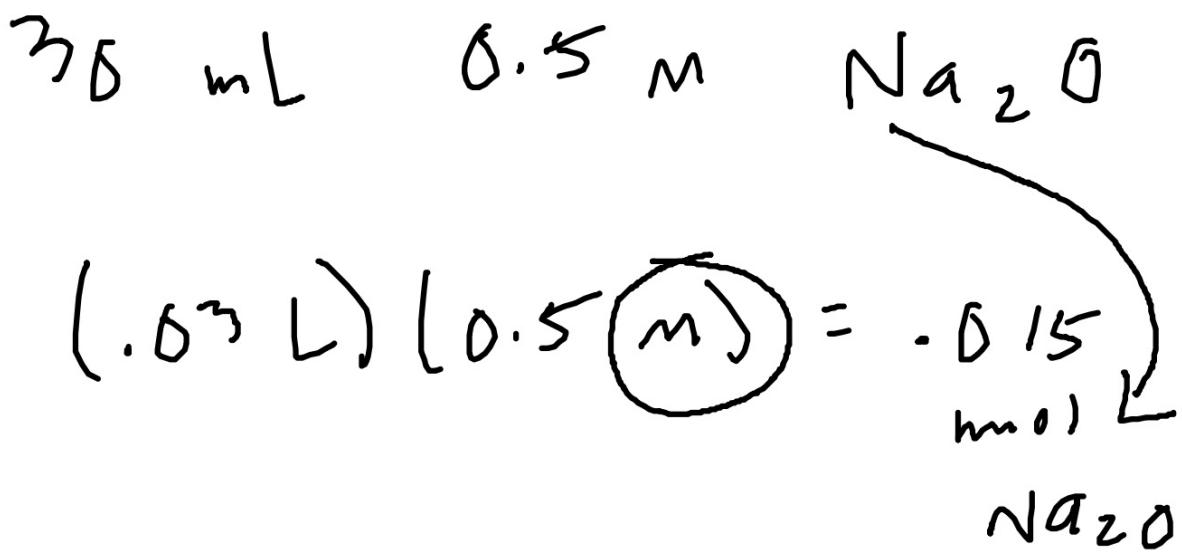
$$6.118 \times 10^{-4} \text{ mol } \text{KMnO}_4 \quad | \quad \begin{matrix} 1 \text{ mol MnO}_4^- \\ \cancel{1 \text{ mol KMnO}_4} \end{matrix}$$

$$\frac{6.118 \times 10^{-4} \text{ mol MnO}_4^-}{1 \text{ mol MnO}_4^-} \left| \begin{array}{l} 5 \text{ mol Fe}^{2+} \\ \hline \end{array} \right.$$

$$= .003059 \text{ mol Fe}^{2+}$$

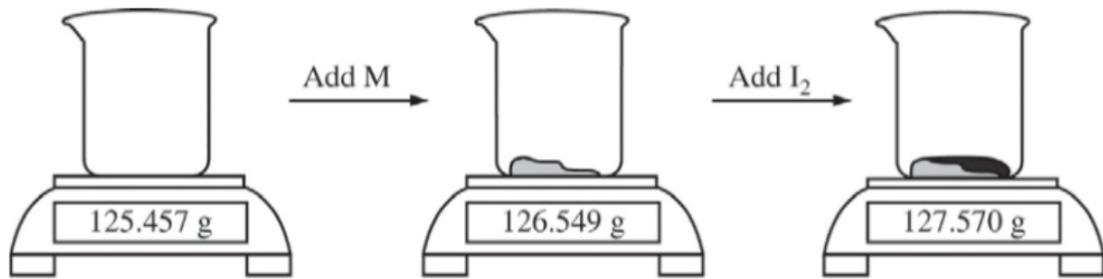
$$[\text{Fe}^{2+}] = .003059 \text{ mol Fe}^{2+}$$

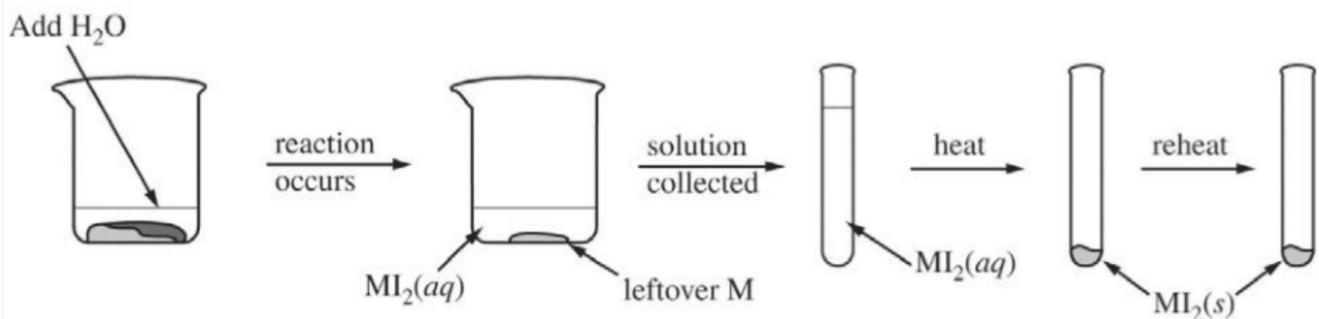
$$= \boxed{.3059 \text{ M}}$$





To determine the molar mass of an unknown metal, M, a student reacts iodine with an excess of the metal to form the water-soluble compound MI_2 , as represented by the equation above. The reaction proceeds until all of the I_2 is consumed. The $MI_2(aq)$ solution is quantitatively collected and heated to remove the water, and the product is dried and weighed to constant mass. The experimental steps are represented below, followed by a data table.





Data for Unknown Metal Lab	
Mass of beaker	125.457 g
Mass of beaker + metal M	126.549 g
Mass of beaker + metal M + I ₂	127.570 g
Mass of MI ₂ , first weighing	1.284 g
Mass of MI ₂ , second weighing	1.284 g

(a) Given that the metal M is in excess, calculate the number of moles of I₂ that reacted.

$$127.570 \text{ g} - 126.549 \text{ g} = 1.021 \text{ g I}_2$$

$$\frac{1.021 \text{ g I}_2}{253.82 \text{ g}} = 0.004023 \text{ mol I}_2$$

0.004023 mol I₂

(b) Calculate the molar mass of the unknown metal M.

$$M_m M_M = g M$$

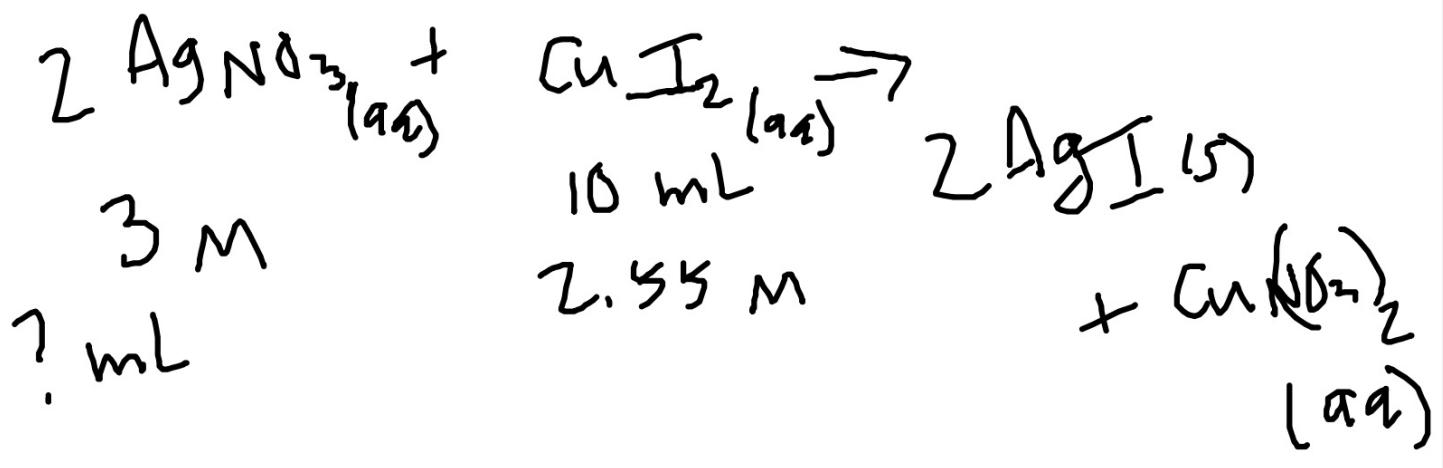
$$\frac{m \text{ol } M}{m \text{ol } I_2}$$

$$\frac{0.04023 \text{ mol } I_2}{1 \text{ mol } I_2} \times \frac{1 \text{ mol } M}{1 \text{ mol } M} = 0.04023 \text{ mol } M$$

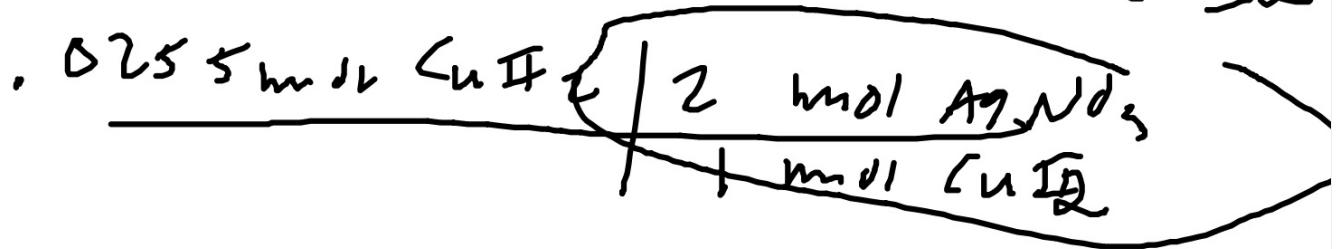
$$1.284 - 1.021 = 0.263 \text{ g M}$$

$$\frac{0.263 \text{ g M}}{0.04023 \text{ mol M}}$$

65.37
g/mol



$$(2.55 \text{ M CuI}_2) \cdot 0.1 \text{ L} = .0255 \text{ mol CuI}_2$$



, 051 mol AgNO₃

$$3 \text{ M} = \frac{-0.51 \text{ mol}}{x}$$

$$x = .617 \text{ L}$$

= 17 mL

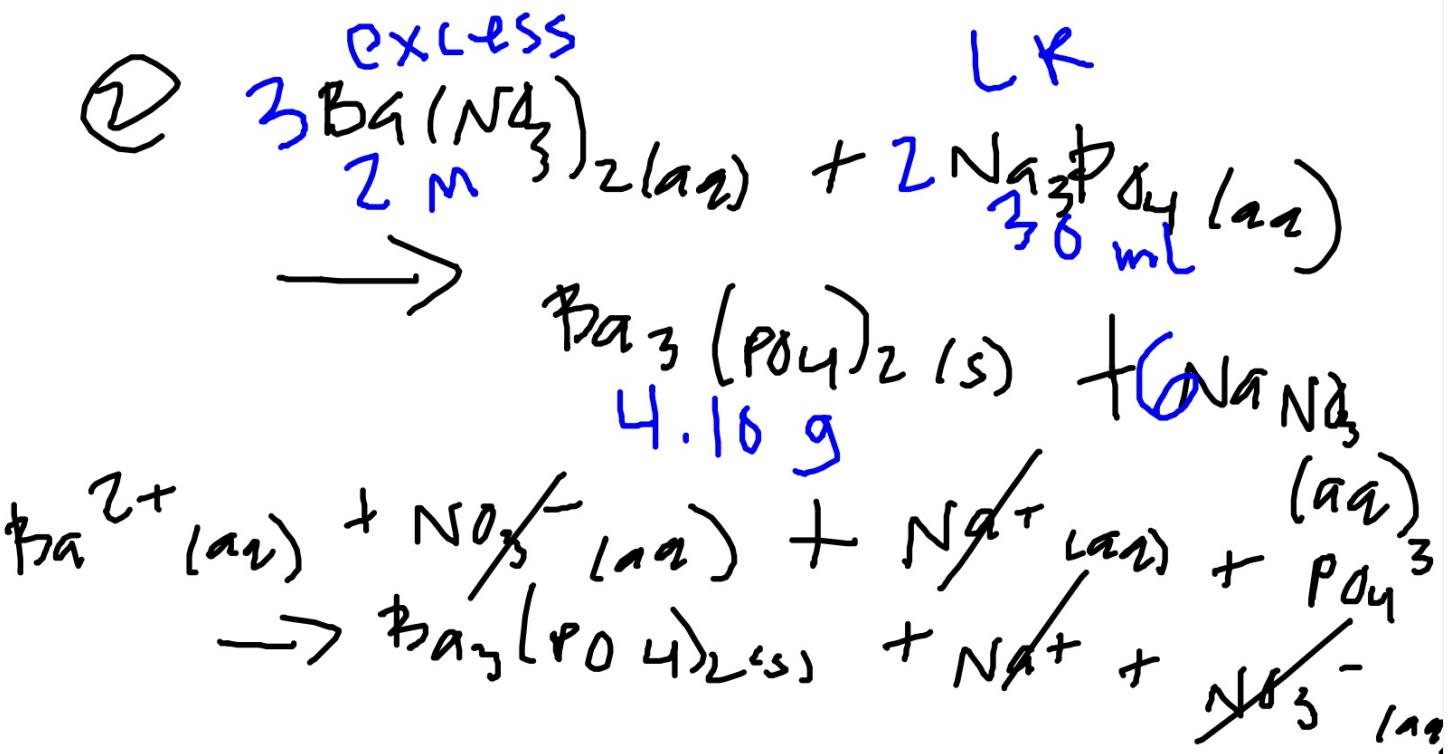
224 g mol $MgCl_2$

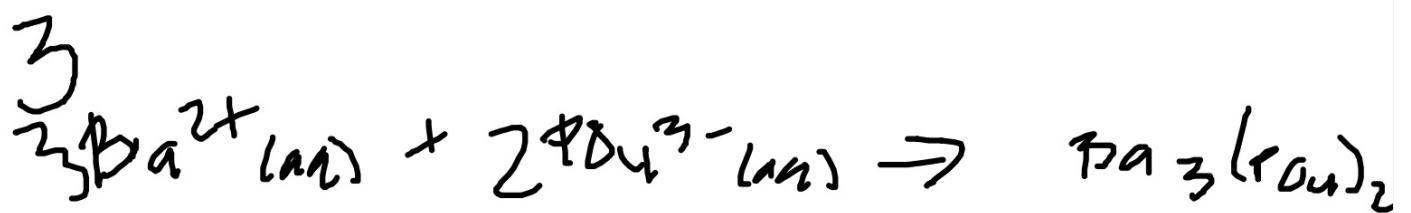
$$[Cl^-] = \frac{\text{mol } Cl^-}{\text{L soln}}$$

$$= 0.241 \text{ mol/L}$$

$$= \frac{0.241 \text{ mol}}{0.35 \text{ L}} = 0.685 \text{ M Cl}^-$$

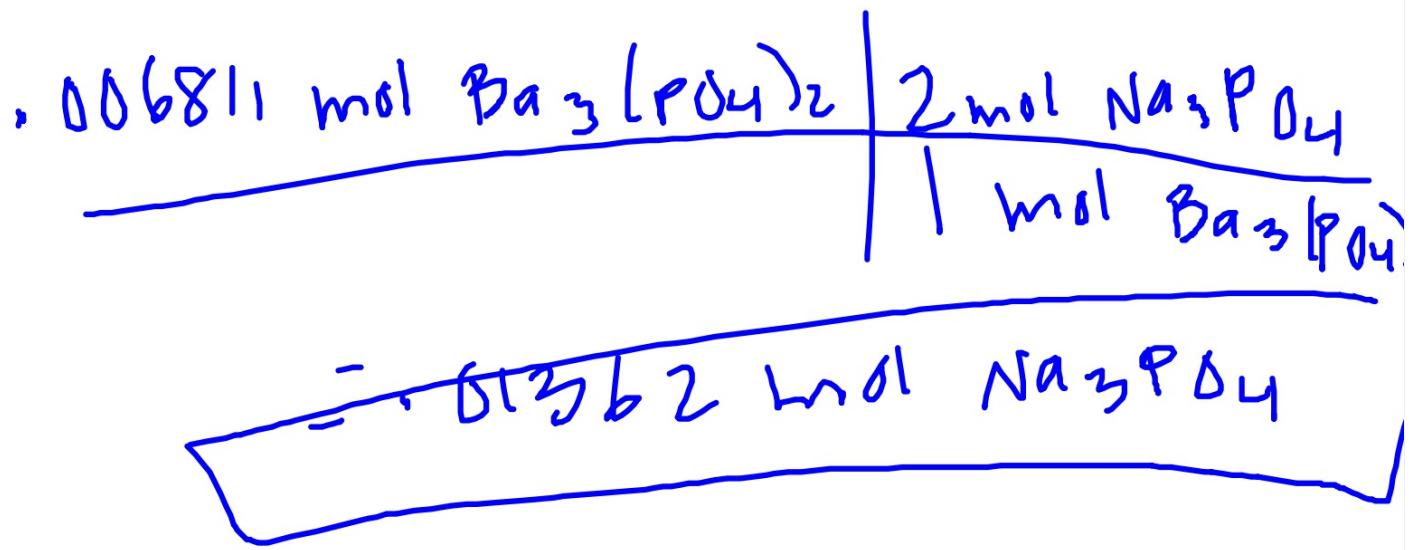
Study Guide



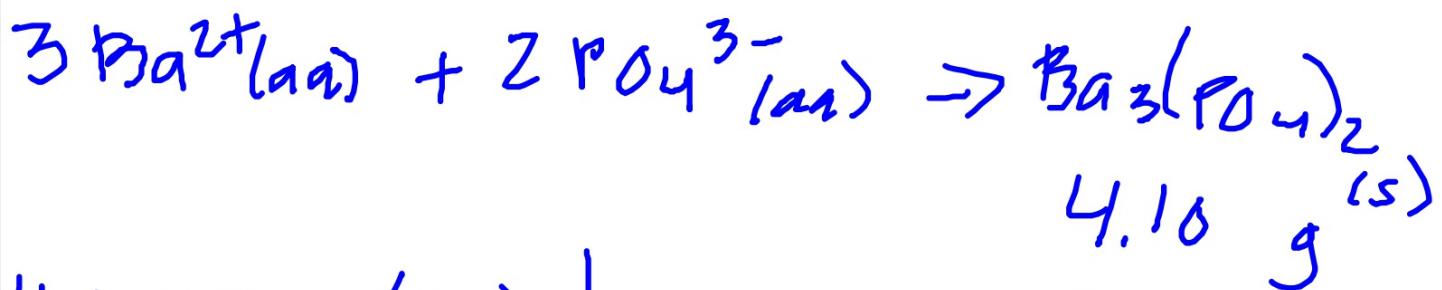


$$\frac{4.10 \text{ g Ba}_3(\text{PO}_4)_2}{601.93 \text{ g}} \mid \text{l mol Ba}_3(\text{PO}_4)_2$$

$$= 0.006811 \text{ mol Ba}_3(\text{PO}_4)_2$$



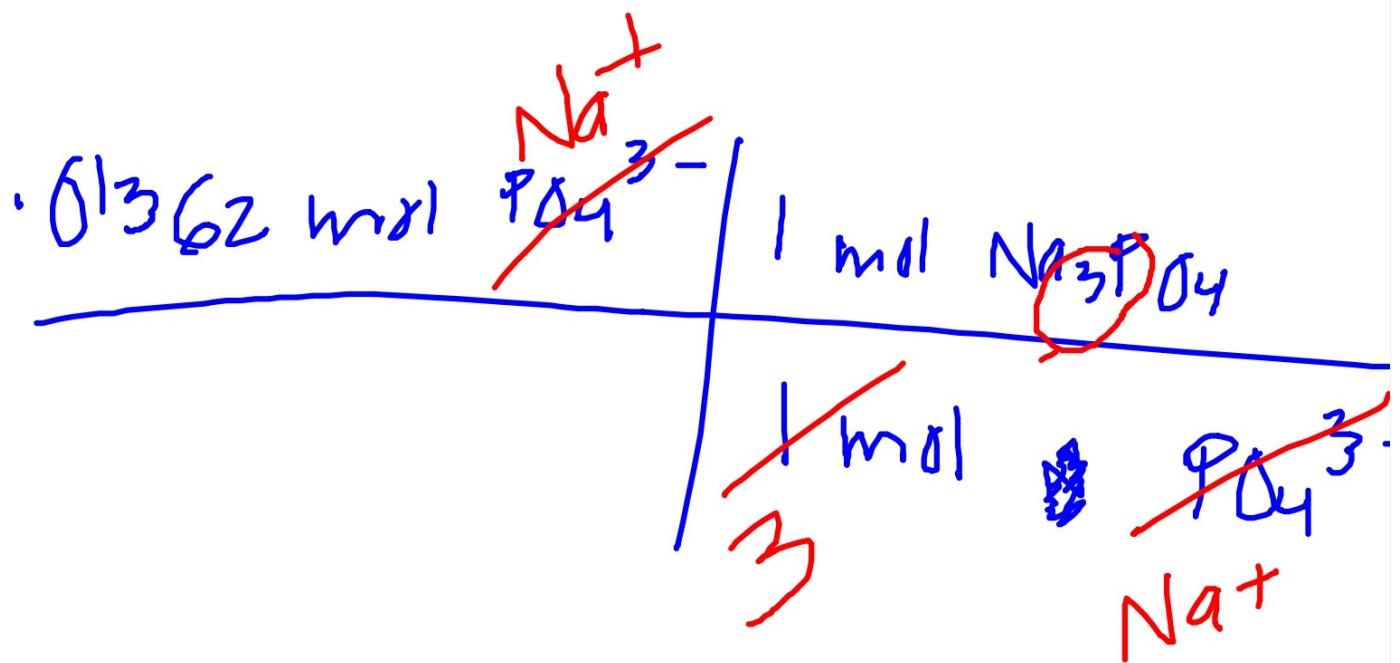
Net ? mol Na_3PO_4



$$\begin{array}{c|c} 4.10 \text{ g } \text{Ba}_3(\text{PO}_4)_2 & 1 \text{ mol } \text{Ba}_3(\text{PO}_4)_2 \\ \hline & 601.93 \text{ g} = .006811 \text{ mol } \text{Ba}_3(\text{PO}_4)_2 \end{array}$$

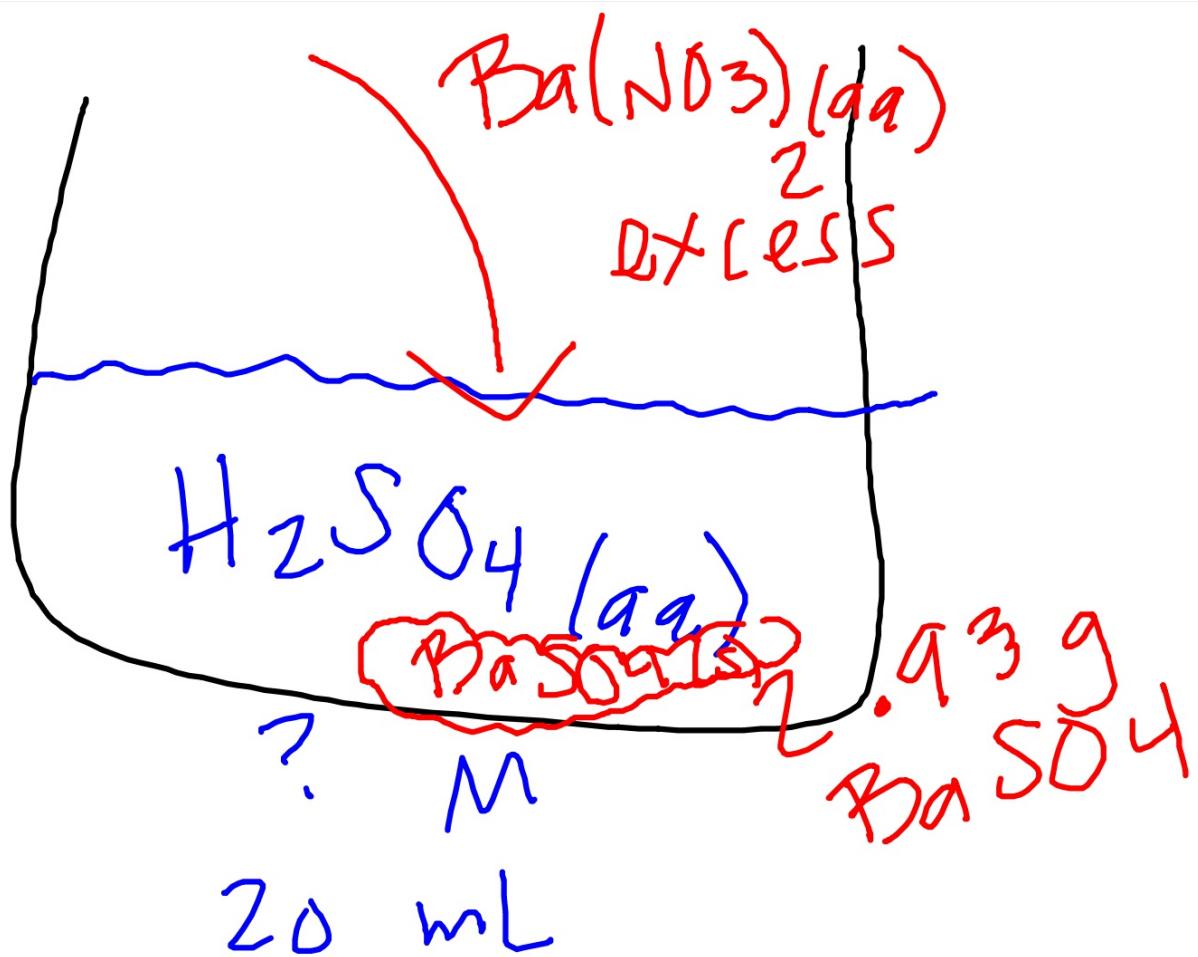
$$\begin{array}{c|c} .006811 \text{ mol } \text{Ba}_3\text{PO}_4 & 2 \text{ mol } \text{PO}_4^{3-} \text{ Ba}_3(\text{PO}_4)_2 \\ \hline & 1 \text{ mol } \text{Ba}_3(\text{PO}_4)_2 \end{array}$$

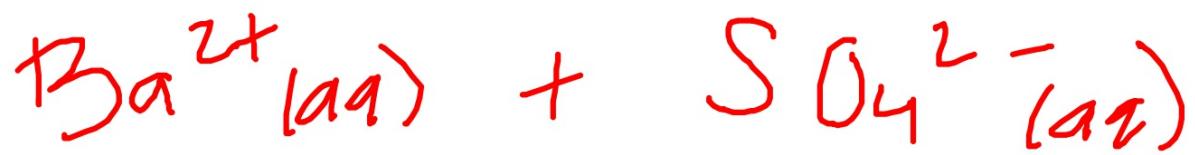
• 0,362 mol PO_4^{3-}



Filtrate = Spectators Na^+
 NO_3^-
+

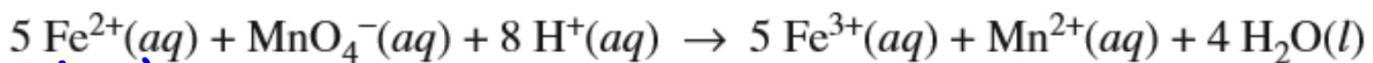
Excess Ba^{2+}





2.93 g

$$[\text{H}_2\text{SO}_4] = \frac{\text{mol H}_2\text{SO}_4}{\text{L soln}} = \frac{\text{---}}{0.02 \text{ L}}$$



10mL
reacts

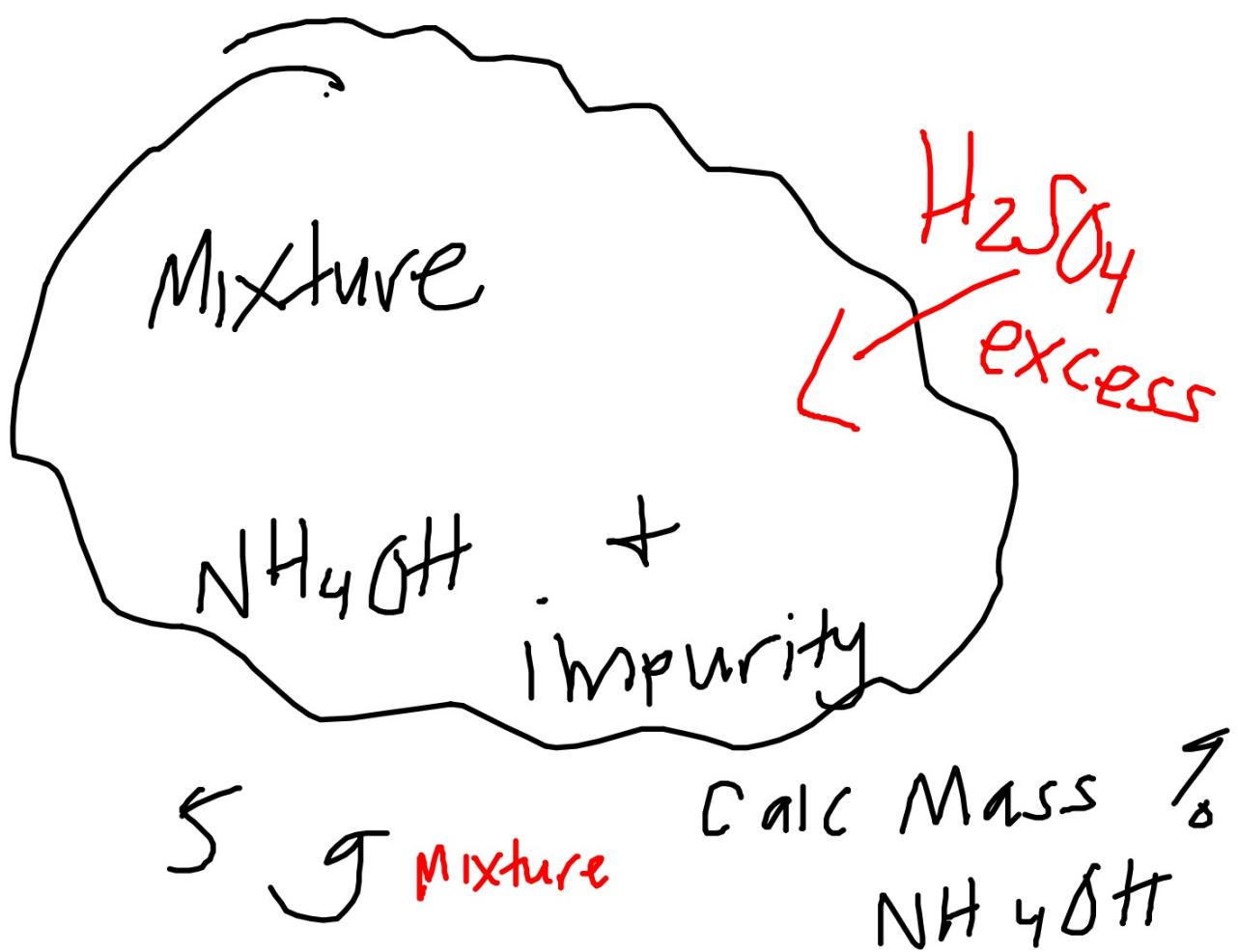
The student titrates a 10.0 mL sample of the $\text{Fe}^{2+}(aq)$ solution. Calculate the value of $[\text{Fe}^{2+}]$ in the solution if it takes 17.48 mL of added 0.0350 M $\text{KMnO}_4(aq)$ to reach the equivalence point of the titration.

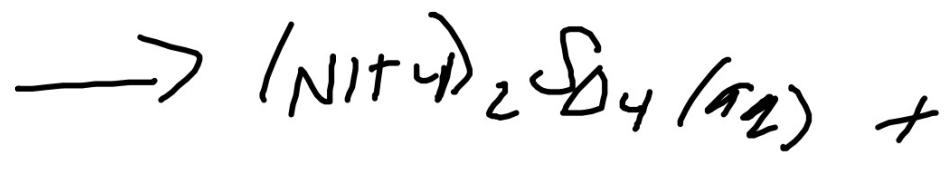
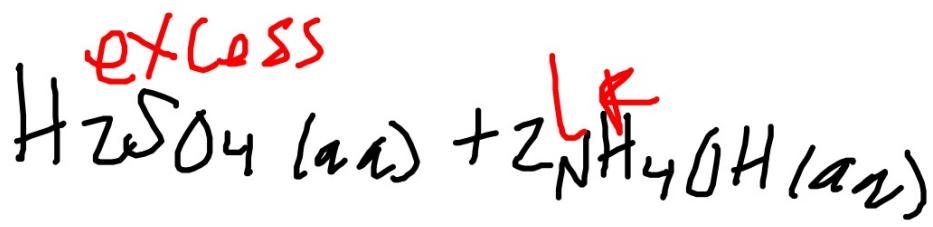
react completely

$$[\text{Fe}^{2+}] = ? \text{ M} = \frac{? \text{ mol Fe}^{2+}}{? \text{ L soln}} = \frac{? \text{ mol Fe}^{2+}}{0.01748 \text{ L}}$$
$$(0.0350 \text{ M } \text{KMnO}_4)(0.01748 \text{ L})$$
$$= 6.118 \times 10^{-4} \text{ mol } \text{KMnO}_4$$
$$1 : 1 = 6.118 \times 10^{-4} \text{ mol}$$
$$\underline{6.118 \times 10^{-4} \text{ mol MnO}_4^-} \quad \left| \begin{array}{c} \text{5 mol Fe}^{2+} \text{ MnO}_4^- \\ \hline 1 \text{ mol MnO}_4^- \end{array} \right.$$

$$= .003059 \text{ mol Fe}^{2+}$$

$$[\text{Fe}^{2+}] = \frac{.003059 \text{ mol Fe}^{2+}}{.01 \text{ L}}$$

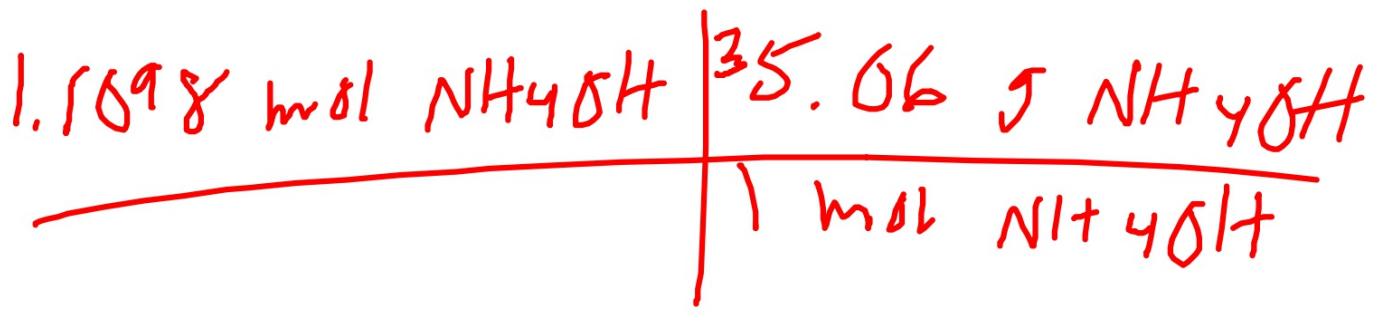




$$\begin{array}{c|cc} 10 \text{ g H}_2\text{O} & 1 \text{ mol H}_2\text{O} & 10 \text{ g} \\ \hline & 18.02 \text{ g H}_2\text{O} & \\ & = .5549 \text{ mol H}_2\text{O} & \end{array}$$



$$= 1.1098 \text{ mol NH}_4\text{OH}$$



$$= 38.9 \text{ g NH}_4\text{OH}$$

$$\text{Mass \%} = \frac{\text{g } \text{NH}_4\text{SOH}}{\text{g mix}} \times 100$$

$$= \frac{38.9 \text{ g}}{5 \text{ g}} \times 100$$

77.8%

9.52 g $MgCl_2$

dissolved in H_2O

→ total volume = 200 mL

Calculate $[MgCl_2]$

Molar mass $MgCl_2 = 95.21 \text{ g/mol}$

$$\frac{9.5\cancel{L}}{\cancel{95.5}} \underset{10}{=} \frac{1}{10} = .1$$

$$\frac{.1}{.2} = \frac{1}{2} = .50 M$$

$MgCl_2$

$$[Cl^-] = 0.5 \times 2 = 1 M$$

Cl^-

.50 mol Na^+

.25 L

$$\frac{.5}{.25} = .5 \text{ (u)} = \boxed{M}$$