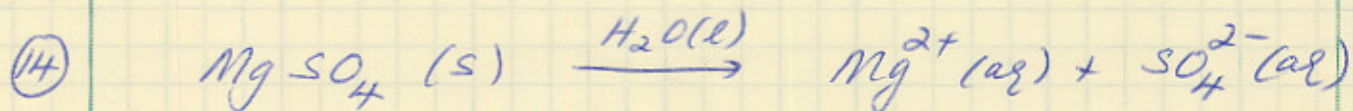


MATTER & MINERALS

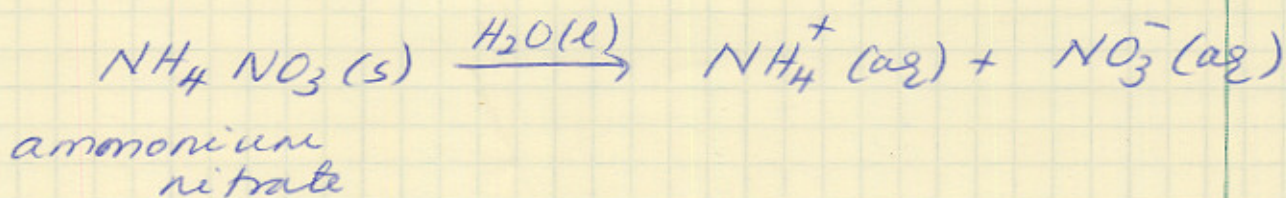
CHEMISTRY HOMEWORK - FALL - WEEK 5

Chapter 4

- (11) (a) $\text{Na}^+(\text{aq})$, $\text{Br}^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(b) $\text{Mg}^{2+}(\text{aq})$, $\text{Cl}^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(c) $\text{Al}^{3+}(\text{aq})$, $\text{NO}_3^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(d) $\text{NH}_4^+(\text{aq})$, $\text{SO}_4^{2-}(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(e) $\text{Na}^+(\text{aq})$, $\text{OH}^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(f) $\text{Fe}^{2+}(\text{aq})$, $\text{SO}_4^{2-}(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(g) $\text{K}^+(\text{aq})$, $\text{MnO}_4^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(h) $\text{H}^+(\text{aq})$, $\text{ClO}_4^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$
(i) $\text{NH}_4^+(\text{aq})$, $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$, $\text{H}_2\text{O}(\text{l})$



magnesium
sulfate



ammonium
nitrate

15 (a) 5.623 g NaHCO_3

$$\begin{aligned}\text{molar mass of } \text{NaHCO}_3 &= [22.99 + 1.008 + 12.01 + \\ &\quad 3(16.00)] \text{ g mol}^{-1} \\ &= 84.008 \text{ g mol}^{-1}\end{aligned}$$

$$5.623 \text{ g } \text{NaHCO}_3 \times \left(\frac{1 \text{ mol}}{84.008 \text{ g}} \right) = 6.693 \times 10^{-2} \text{ mol } \text{NaHCO}_3$$

$$\text{molarity} = \frac{\text{\# of moles of solute}}{\text{volume of solution (L)}}$$

$$= \frac{6.693 \times 10^{-2} \text{ mol}}{250.0 \times 10^{-3} \text{ L}} = \underline{\underline{0.2677 \text{ M}}}$$

$$\begin{aligned}\text{(b) molar mass of } \text{K}_2\text{Cr}_2\text{O}_7 &= [2(39.10) + 2(52.00) + \\ &\quad 7(16.00)] \text{ g mol}^{-1} \\ &= 294.2 \text{ g mol}^{-1}\end{aligned}$$

$$\begin{aligned}184.6 \text{ mg } \text{K}_2\text{Cr}_2\text{O}_7 &\times \left(\frac{1 \text{ g}}{10^3 \text{ mg}} \right) \times \left(\frac{1 \text{ mol}}{294.2 \text{ g}} \right) \\ &= 6.275 \times 10^{-4} \text{ mol}\end{aligned}$$

$$\text{molarity} = \frac{6.275 \times 10^{-4} \text{ mol}}{500.0 \times 10^{-3} \text{ L}} = \underline{\underline{1.255 \times 10^{-3} \text{ M}}}$$

$$(c) \quad 0.1025 \text{ g Cu} = 0.1025 \text{ g} \times \left(\frac{1 \text{ mol Cu}}{63.55 \text{ g}} \right) \\ = 1.6129 \times 10^{-3} \text{ mol Cu}$$

All this Cu is converted to Cu^{2+} by adding HNO_3 .

$$\therefore \# \text{ of moles of } \text{Cu}^{2+} = 1.6129 \times 10^{-3} \text{ mol Cu}^{2+}$$

$$\text{Volume of solution} = 200.0 \text{ mL} = 200.0 \times 10^{-3} \text{ L}$$

$$\text{molarity} = \frac{1.6129 \times 10^{-3} \text{ mol}}{200.0 \times 10^{-3} \text{ L}} = \underline{\underline{8.065 \times 10^{-3} \text{ M}}}$$

$$(16) \quad 75.0 \text{ mL ethanol} \times \left(\frac{0.79 \text{ g}}{\text{cm}^3} \right) = \text{mass of ethanol}$$

$$\left[\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow \text{mass} = \text{volume} \times \text{density} \right]$$

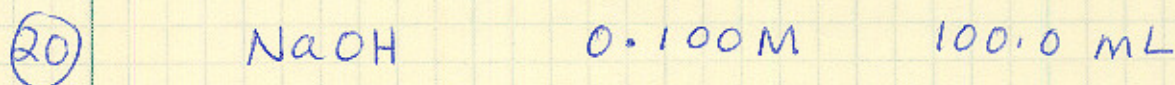
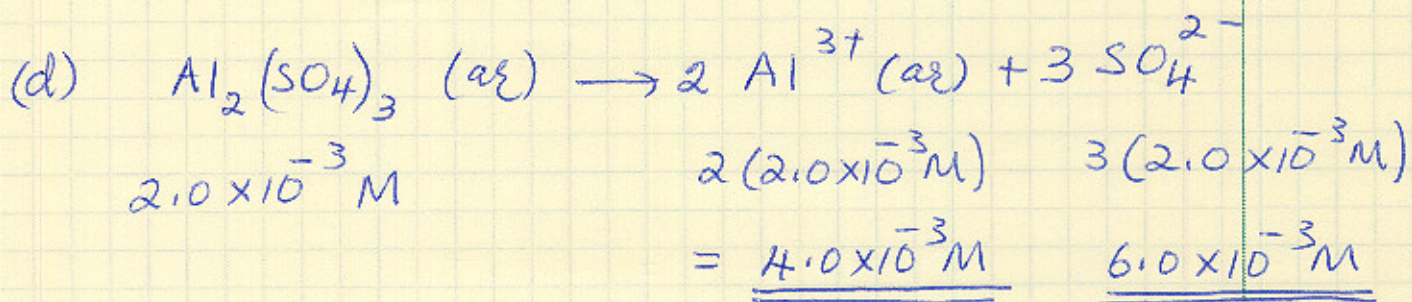
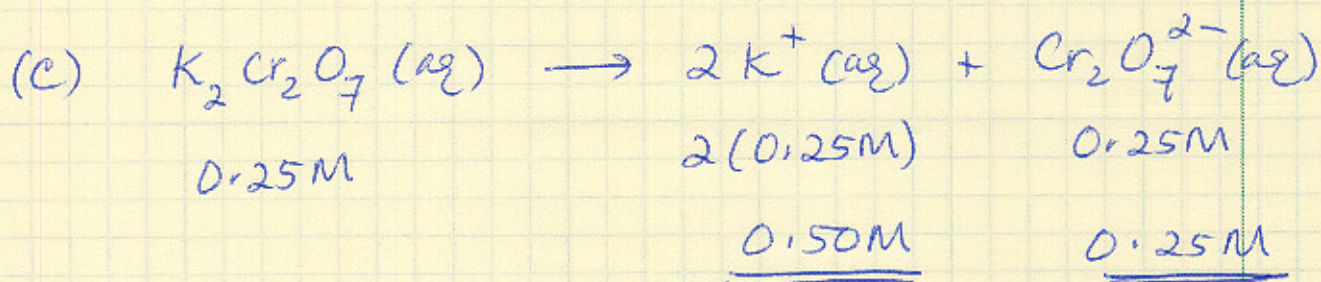
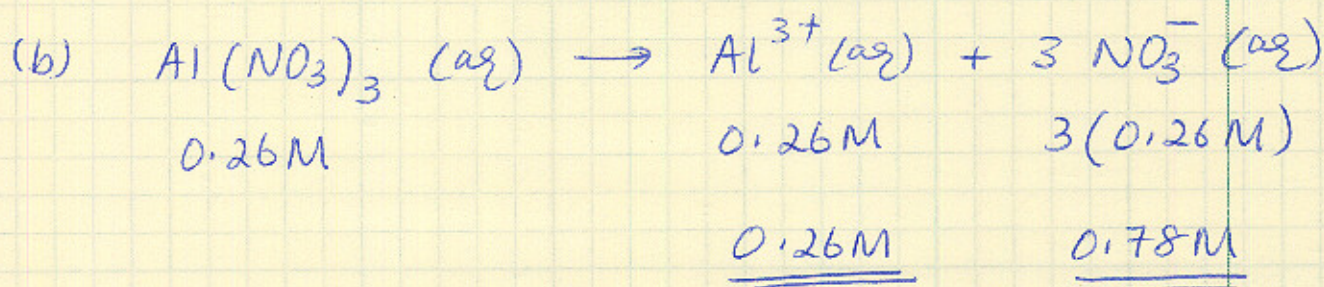
also $\text{mL} = \text{cm}^3$

$$\therefore \text{mass of ethanol} = 59.25 \text{ g} \quad \text{C}_2\text{H}_5\text{OH}$$

$$\left. \begin{array}{l} \# \text{ of moles of} \\ \text{of ethanol} \end{array} \right\} \text{Molar mass} = [2(12.01) + 6(1.008) + 16.00] \text{ g mol}^{-1} \\ = 46.008 \text{ g mol}^{-1}$$

$$\# \text{ of moles of ethanol} = 59.25 \text{ g} \times \left(\frac{\text{mol}}{46.008 \text{ g}} \right) \\ = 1.286 \text{ mol}$$

$$\text{molarity} = \left(\frac{1.286 \text{ mol}}{250.0 \times 10^{-3} \text{ L}} \right) = \underline{\underline{5.14 \text{ M}}}$$



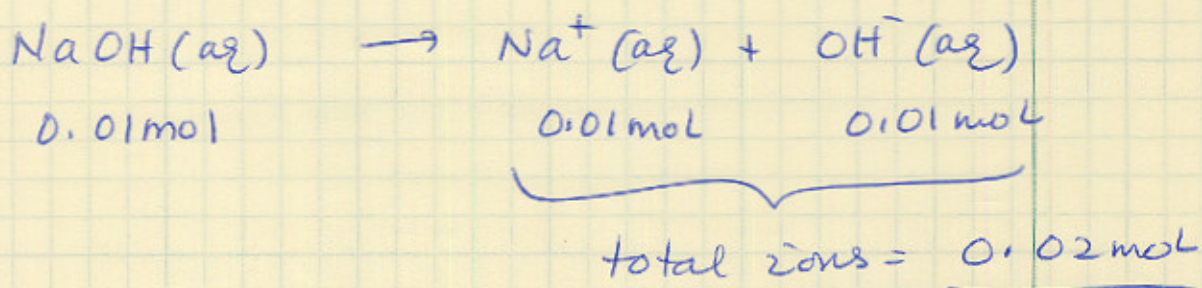
$$\text{molarity} = \frac{\# \text{ of moles of solute}}{\text{volume of solution (L)}}$$

$$\therefore \# \text{ of moles of NaOH} = \text{molarity} \times \text{volume of solution (L)}$$

$$= (0.100 \text{ M}) (100.0 \times 10^{-3} \text{ L})$$

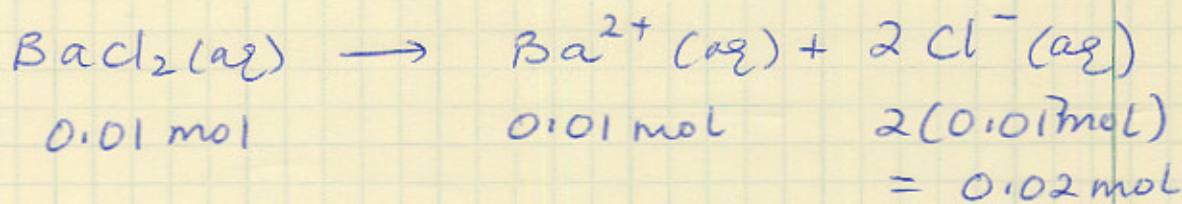
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$$\# \text{ of moles of NaOH} = 0.01 \text{ mol}$$



$$\underline{\text{BaCl}_2} \quad 0.200 \text{ M} \quad 50.0 \text{ mL} = 50.0 \times 10^{-3} \text{ L}$$

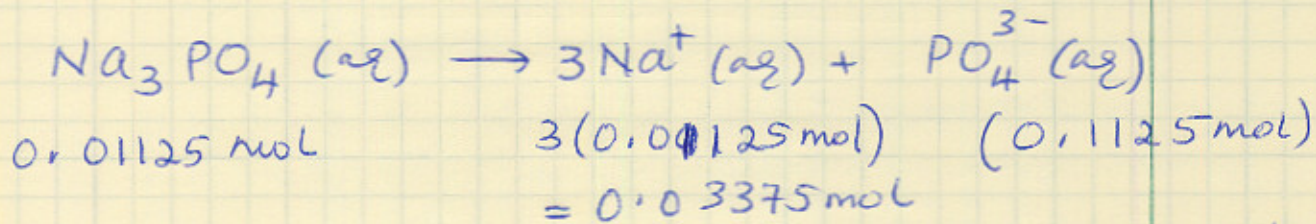
$$\begin{aligned} \# \text{ of moles of BaCl}_2 &= (0.200 \text{ M})(50.0 \times 10^{-3} \text{ L}) \\ &= 0.01 \text{ mol} \end{aligned}$$



total ions = 0.03 mol

$$\underline{\text{Na}_3\text{PO}_4} \quad 0.150 \text{ M} \quad 75.0 \text{ mL} = 75.0 \times 10^{-3} \text{ L}$$

$$\begin{aligned} \# \text{ of moles of Na}_3\text{PO}_4 &= (0.150 \text{ M})(75.0 \times 10^{-3} \text{ L}) \\ &= 0.01125 \text{ mol} \end{aligned}$$



total ions = 0.045 mol

6/ Therefore ~~Base~~ Na_3PO_4 solution has the highest number of ions.

(21) volume = ~~volume~~ V (say)

molarity = 0.100M

weight = 0.350g NaHCO_3

Molar mass of $\text{NaHCO}_3 = [22.99 + 1.008 + 12.01 + 3(16.00)] \text{g mol}^{-1}$

= 84.008g/mol

of moles of $\text{NaHCO}_3 = 0.350\text{g} \times \left(\frac{\text{mol}}{84.008\text{g}} \right)$

= $4.166 \times 10^{-3} \text{mol}$

molarity = $\frac{\text{moles of solute}}{\text{volume of solution}}$

∴ volume of solution = $\frac{\text{moles of solute}}{\text{molarity}}$

$V = \frac{4.166 \times 10^{-3} \text{mol}}{0.100 \text{mol L}^{-1}}$

= $0.04166 \text{L} \times \frac{10^3 \text{mL}}{\text{L}}$

= 41.7 mL

(24) (a) Concentrated H_2SO_4

$$\text{Volume} = V$$

$$\text{molarity} = 18M$$

moles H_2SO_4 in
concentrated solution

molarity \times volume

$$(18M) V$$

$$\therefore V$$

diluted H_2SO_4

$$\text{volume} = 1.00L$$

$$\text{molarity} = 0.50M$$

moles of H_2SO_4 in
diluted solution

(molarity \times volume)

$$(0.50M) (1.00L)$$

$$= \frac{(0.50M)(1.00L)}{(18M)}$$

$$V = 0.02778L \times \left(\frac{10^3 mL}{L}\right)$$

$$= \underline{\underline{27.8 mL}}$$

(b) Concentrated HCl

$$\text{volume} = V$$

$$\text{molarity} = 12M$$

$$(12M) \times V$$

$$V$$

diluted HCl

$$\text{volume} = 1.00L$$

$$\text{molarity} = 0.50M$$

$$(0.50M) \times (1.00L)$$

$$= \frac{(0.50M)(1.00L)}{12M}$$

$$= 4.167 \times 10^{-2} L$$

$$V = \underline{\underline{41.7 mL}}$$

(c) solid $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$

$$\begin{aligned} \text{Volume of solution} &= 1.00\text{L} \\ \text{molarity of solution} &= 0.50\text{M} \end{aligned}$$

$$\text{molarity} = \frac{\# \text{ moles of } \text{NiCl}_2 \cdot 6\text{H}_2\text{O}}{\text{Volume of solution}}$$

$$0.50\text{M} = \frac{\# \text{ moles } \text{NiCl}_2 \cdot 6\text{H}_2\text{O}}{1.00\text{L}}$$

$$\begin{aligned} \therefore \# \text{ moles of } \text{NiCl}_2 \cdot 6\text{H}_2\text{O} &= (0.50\text{M})(1.00\text{L}) \\ &= 0.50\text{mol} \end{aligned}$$

$$\begin{aligned} \text{molar mass of } \text{NiCl}_2 \cdot 6\text{H}_2\text{O} &= [58.69 + 2(35.5) + \\ &\quad 12(1.008) + 6(16)] \text{g mol}^{-1} \\ &= 237.786 \text{g mol}^{-1} \end{aligned}$$

$$0.50\text{mol } \text{NiCl}_2 \cdot 6\text{H}_2\text{O} \times \left(\frac{237.786\text{g}}{1\text{mol}} \right)$$

$$= 118.893\text{g } \text{NiCl}_2 \cdot 6\text{H}_2\text{O}$$

$$= \underline{\underline{119\text{g}}}$$

Weigh out 119g of $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$. Add volumetrically to a 1.00L volumetric flask and add enough water to make up to the mark.

a1

$$\text{volume} = V$$

$$\text{molarity} = 16\text{M}$$

$$\# \text{ moles} = (16\text{M})V$$

$$V$$

$$1.00\text{L}$$

$$0.50\text{M}$$

$$(0.50\text{M})(1.00\text{L})$$

$$\frac{(0.50\text{M})(1.00\text{L})}{16\text{M}}$$

$$V = 0.03125\text{L} \times \frac{1000\text{mL}}{\text{L}}$$

$$V = \underline{\underline{31.25\text{mL}}}$$

(e) Na_2CO_3

$$\text{moles of } \text{Na}_2\text{CO}_3 = (\text{molarity})(\text{volume})$$

$$= (0.50\text{M})(1.00\text{L}) = 0.50\text{mol}$$

$$\text{molar mass of } \text{Na}_2\text{CO}_3 = [2(22.99) + (12.01) + 3(16.00)]$$

$$= 105.99\text{g mol}^{-1} \quad \text{g/mol}$$

$$0.50\text{mol } \text{Na}_2\text{CO}_3 \times \left(\frac{105.99\text{g}}{1\text{mol}} \right) = 52.995\text{g}$$

$$= 53\text{g}$$

Weigh 53g of Na_2CO_3 , add volumetrically to a 1.00L volumetric flask. Add enough water to make up to the 1.00L mark.

(25) First solution = stock solution

$$10.8 \text{ g } (\text{NH}_4)_2 \text{SO}_4 \quad \left. \begin{array}{l} \text{volume of} \\ \text{stock solution} \end{array} \right\} = 100.0 \text{ mL} \\ = 100.0 \times 10^{-3} \text{ L}$$

$$\left. \begin{array}{l} \text{molar mass of} \\ (\text{NH}_4)_2 \text{SO}_4 \end{array} \right\} = [2(14.01) + 8(1.008) + 32.07 \\ + 4(16.00)] \text{ g mol}^{-1} \\ = 132.154 \text{ g mol}^{-1}$$

$$10.8 \text{ g } (\text{NH}_4)_2 \text{SO}_4 \times \left(\frac{\text{mol}}{132.154 \text{ g}} \right) = 8.172 \times 10^{-2} \text{ mol}$$

$$\begin{aligned} \text{molarity of stock solution} &= \frac{8.172 \times 10^{-2} \text{ mol}}{100.0 \times 10^{-3} \text{ L}} \\ &= 0.8172 \text{ M} \end{aligned}$$

Final solution = diluted solution

	stock solution used	dilute solution prepared
volume =	10.00 mL = $10.00 \times 10^{-3} \text{ L}$	50.00 mL = $60.00 \times 10^{-3} \text{ L}$
Concentration =	0.8172 M	

$$\left. \begin{array}{l} \# \text{ of moles of } (\text{NH}_4)_2 \text{SO}_4 \\ \text{in the stock solution} \\ \text{used} \end{array} \right\} = (0.8172 \text{ M}) (10.00 \times 10^{-3} \text{ L}) \\ = 8.172 \times 10^{-3} \text{ mol}$$

∴ 8.172×10^{-3} mol of $(\text{NH}_4)_2\text{SO}_4$ was added
and the final volume of diluted solution
was $60.00 \text{ mL} = 60.00 \times 10^{-3} \text{ L}$

$$\begin{aligned}\text{molarity of diluted solution} &= \frac{8.172 \times 10^{-3} \text{ mol}}{60.00 \times 10^{-3} \text{ L}} \\ &= 0.1362 \text{ M} \\ &= \underline{\underline{1.36 \times 10^{-1} \text{ M}}}\end{aligned}$$

28) Stock solution

$$\begin{aligned}1.584 \text{ g Mn} &= 1.584 \text{ g} \times \left(\frac{\text{mol}}{54.94 \text{ g}} \right) \\ &= 2.883 \times 10^{-2} \text{ mol Mn}\end{aligned}$$

all of this metal is converted to Mn^{2+}

$$\therefore \# \text{ moles } \text{Mn}^{2+} = 2.883 \times 10^{-2} \text{ mol}$$

$$\text{volume of stock solution} = 1.000 \text{ L}$$

$$\begin{aligned}\text{molarity of stock solution} &= \frac{2.883 \times 10^{-2} \text{ mol}}{1.000 \text{ L}} \\ &= \underline{\underline{2.883 \times 10^{-2} \text{ M}}}\end{aligned}$$

Solution A

	<u>Stock solⁿ taken</u>	<u>Solⁿ A prepared</u>
molarity	$2.883 \times 10^{-2} \text{ M}$	M_A (say)
Volume	50.00 mL $= 50.00 \times 10^{-3} \text{ L}$	$1000.0 \text{ mL} =$ 1.0000 L
# of moles of solute	$\left\{ = \underbrace{(2.883 \times 10^{-2} \text{ M})}_{\text{Stock sol}^n} \underbrace{(50.00 \times 10^{-3} \text{ L})}_{\text{Volume}} \right\} = M_A (1.0000 \text{ L})$ $= 1.4415 \times 10^{-3} \text{ mol}$	
$\therefore M_A =$	$\frac{1.4415 \times 10^{-3} \text{ mol}}{1.0000 \text{ L}}$ $= \underline{\underline{1.4415 \times 10^{-3} \text{ M}}}$	

Solution B

	<u>Solⁿ A taken</u>	<u>Solⁿ B prepared</u>
molarity	$1.4415 \times 10^{-3} \text{ M}$	M_B
Volume	$10.00 \text{ mL} = 10.00 \times 10^{-3} \text{ L}$	$250.0 \text{ mL} = 250.0 \times 10^{-3} \text{ L}$
# moles of solute	$\left\{ = \underbrace{(1.4415 \times 10^{-3} \text{ M})}_{\text{Sol}^n \text{ A}} \underbrace{(10.00 \times 10^{-3} \text{ L})}_{\text{Volume}} \right\} = M_B (250.0 \times 10^{-3} \text{ L})$ $= 1.4415 \times 10^{-5} \text{ mol} = M_B (250.0 \times 10^{-3} \text{ L})$	

$$250.0 \times 10^{-3} \text{ L}$$

Solution C

	<u>Solution B taken</u>	<u>Solution C prepared</u>
molarity	$5.766 \times 10^{-5} \text{ M}$	M_C
volume	$10.00 \text{ mL} = 10.00 \times 10^{-3} \text{ L}$	$500.0 \text{ mL} =$ $500.0 \times 10^{-3} \text{ L}$
# moles of solute	$(5.766 \times 10^{-5} \text{ M})(10.00 \times 10^{-3} \text{ L}) = M_C (500.0 \times 10^{-3} \text{ L})$ $= 5.766 \times 10^{-7} \text{ mol} = M_C (500.0 \times 10^{-3} \text{ L})$	
	$M_C = \frac{5.766 \times 10^{-7} \text{ mol}}{500.0 \times 10^{-3} \text{ L}} = \underline{\underline{1.153 \times 10^{-6} \text{ M}}}$	