MATTER & MINERALS

CHEMISTRY HOMEWORK - FALL - WEEK 5

Chapter 4

(11) (a) $Na^+(aq)$, $Br^-(aq)$, $H_2O(l)$

(b) $Mg^{2+}(aq)$, $Cl^-(aq)$, $H_2O(l)$

(c) $Al^{3+}(aq)$, $NO_3^-(aq)$, $H_2O(l)$

(d) $NH_4^+(aq)$, $SO_4^{2-}(aq)$, $H_2O(l)$

(e) $Na^+(aq)$, $OH^-(aq)$, $H_2O(l)$

(f) $Fe^{2+}(aq)$, $SO_4^{2-}(aq)$, $H_2O(l)$

(g) $K^+(aq)$, $MnO_4^-(aq)$, $H_2O(l)$

(h) $H^+(aq)$, $ClO_4^-(aq)$, $H_2O(l)$

(i) $NH_4^+(aq)$, $C_2H_3O_2^-(aq)$, $H_2O(l)$

(14) $$MgSO_4(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + SO_4^{2-}(aq)$$

magnesium sulfate

$$NH_4NO_3(s) \xrightarrow{H_2O(l)} NH_4^+(aq) + NO_3^-(aq)$$

ammonium nitrate

(15) (a) 5.623 g $NaHCO_3$

molar mass of $NaHCO_3$ = [22.99 + 1.008 + 12.01 + 3(16.00)] $g\,mol^{-1}$

= 84.008 $g\,mol^{-1}$

$$5.623\,g\ NaHCO_3 \times \left(\frac{1\,mol}{84.008\,g}\right) = 6.693 \times 10^{-2}\,mol\ NaHCO_3$$

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

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

(b) Molar mass of $K_2Cr_2O_7$ = [2(39.10) + 2(52.00) + 7(16.00)] $g\,mol^{-1}$

= 294.2 $g\,mol^{-1}$

$$184.6\,mg\ K_2Cr_2O_7 \times \left(\frac{1\,g}{10^3\,mg}\right) \times \left(\frac{1\,mol}{294.2\,g}\right)$$

$$= 6.275 \times 10^{-4}\,mol$$

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

(C) $0.1025 g\ Cu = 0.1025 g \times \left(\frac{1\ mol\ Cu}{63.55 g}\right)$

$= 1.6129 \times 10^{-3}\ mol\ Cu$

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

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

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

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

(15) $75.0 mL\ ethanol \times \left(\frac{0.79 g}{cm^3}\right) = mass\ of\ ethanol$

[$density = \frac{mass}{volume} \Rightarrow mass = volume \times density$

also $mL = cm^3$]

$\therefore$ mass of ethanol $= 59.25 g$ $\quad C_2H_5OH$

~~# of moles of~~ Molar mass of ethanol } $= [2(12.01) + 6(1.008) + 16.00]\ g\ mol^{-1}$

$= 46.008\ g\ mol^{-1}$

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

$= 1.286\ mol$

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

(17) (a) $CaCl_2 (aq) \longrightarrow Ca^{2+} (aq) + 2 Cl^- (aq)$

0.15 M — 0.15 M — 2(0.15 M)

0.15 M — 0.30 M

(b) $Al(NO_3)_3 (aq) \longrightarrow Al^{3+} (aq) + 3 NO_3^- (aq)$

0.26M — 0.26M — 3(0.26 M)

0.26M — 0.78M

(c) $K_2Cr_2O_7 (aq) \longrightarrow 2 K^+ (aq) + Cr_2O_7^{2-} (aq)$

0.25M — 2(0.25M) — 0.25M

0.50M — 0.25M

(d) $Al_2(SO_4)_3 (aq) \longrightarrow 2 Al^{3+} (aq) + 3 SO_4^{2-}$

2.0×10^{-3} M — $2(2.0 \times 10^{-3} M)$ — $3(2.0 \times 10^{-3} M)$

$= 4.0 \times 10^{-3} M$ — $6.0 \times 10^{-3} M$

(20) NaOH 0.100M 100.0 mL

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

$\therefore$ # of moles of NaOH = molarity × volume of solution (L)

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

# of moles of NaOH = 0.01 mol

$$NaOH(aq) \longrightarrow Na^{+}(aq) + OH^{-}(aq)$$

0.01 mol | 0.01 mol | 0.01 mol

total ions = 0.02 mol

$BaCl_2$ 0.200M 50.0 mL = 50.0×10^{-3} L

# of moles of $BaCl_2$ = $(0.200M)(50.0 \times 10^{-3} L)$
= 0.01 mol

$$BaCl_2(aq) \longrightarrow Ba^{2+}(aq) + 2Cl^{-}(aq)$$

0.01 mol | 0.01 mol | 2(0.01 mol) = 0.02 mol

total ions = 0.03 mol

Na_3PO_4 0.150M 75.0 mL = 75.0×10^{-3} L

# of moles of Na_3PO_4 = $(0.150M)(75.0 \times 10^{-3} L)$
= 0.01125 mol

$$Na_3PO_4(aq) \longrightarrow 3Na^{+}(aq) + PO_4^{3-}(aq)$$

0.01125 mol | 3(0.01125 mol) = 0.03375 mol | (0.1125 mol)

total ions = 0.045 mol

Therefore ~~...~~ Na_3PO_4 solution has the highest number of ions.

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

molarity = 0.100M

weight = 0.350 g $NaHCO_3$

Molar mass of $NaHCO_3$ = $[22.99 + 1.008 + 12.01 + 3(16.00)]\ g\ mol^{-1}$

$= 84.008\ g/mol$

# of moles of $NaHCO_3$ = $0.350\,g \times \left(\frac{mol}{84.008\,g}\right)$

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

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

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

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

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

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

(24) (a)

Concentrated H_2SO_4	diluted H_2SO_4
Volume = V	volume = 1.00 L
molarity = 18 M	molarity = 0.50 M

moles H_2SO_4 in concentrated solution } = # moles of H_2SO_4 in diluted solution

molarity × volume = (molarity × volume)

$(18M)\ V = (0.50M)(1.00L)$

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

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

$= \underline{27.8\ mL}$

(b)

Concentrated HCl	diluted HCl
volume = V	volume = 1.00 L
molarity = 12 M	molarity = 0.50 M

$(12M) \times V = (0.50M) \times (1.00L)$

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

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

$V = \underline{41.7\ mL}$

2 (c) Solid $NiCl_2 \cdot 6H_2O$

Volume of solution = 1.00 L
molarity of solution = 0.50 M.

$$\text{Molarity} = \frac{\#\ \text{moles of } NiCl_2 \cdot 6H_2O}{\text{volume of solution}}$$

$$0.50\,M = \frac{\#\ \text{moles } NiCl_2 \cdot 6H_2O}{1.00\,L}$$

$$\therefore \#\ \text{moles of } NiCl_2 \cdot 6H_2O = (0.50\,M)(1.00\,L)$$
$$= 0.50\ \text{mol}$$

$$\text{molar mass of } NiCl_2 \cdot 6H_2O = [58.69 + 2(35.5) + 12(1.008) + 6(16)]\ g\ mol^{-1}$$
$$= 237.786\ g\ mol^{-1}$$

$$0.50\ \text{mol } NiCl_2 \cdot 6H_2O \times \left(\frac{237.786\ g}{1\ \text{mol}}\right)$$

$$= 118.893\ g\ NiCl_2 \cdot 6H_2O$$

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

Weigh out 119 g of $NiCl_2 \cdot 6H_2O$. Add volumetically to a 1.00 L volumetric flask and add enough water to make up to the mark.

a)

volume	=	V	$1.00 L$
molarity	=	$16 M$	$0.50 M$
# moles	=	$(16 M) V$	$= (0.50 M)(1.00 L)$
		V	$= \frac{(0.50 M)(1.00 L)}{16 M}$

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

$$V = 31.25 mL$$

(e) Na_2CO_3

moles of Na_2CO_3 = (molarity) (volume)

$= (0.50 M)(1.00 L) = 0.50$ mol.

molar mass of $Na_2CO_3 = [2(22.99) + (12.01) + 3(16.00)]$ g/mol

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

$$0.50 \text{ mol } Na_2CO_3 \times \left(\frac{105.99 g}{1 mol}\right) = 52.995 g.$$

$= 53 g.$

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

(25) **First solution = stock solution**

$10.8\,g\ (NH_4)_2SO_4$

volume of stocksolution $= 100.0\,mL = 100.0 \times 10^{-3}\,L$

molar mass of $(NH_4)_2SO_4$ $= [2(14.01) + 8(1.008) + 32.07 + 4(16.00)]\ g\,mol^{-1}$

$= 132.154\ g\,mol^{-1}$

$$10.8\,g\ (NH_4)_2SO_4 \times \left(\frac{mol}{132.154\,g}\right) = 8.172 \times 10^{-2}\,mol$$

$$\text{molarity of stock solution} = \frac{8.172 \times 10^{-2}\,mol}{100.0 \times 10^{-3}\,L}$$

$$= 0.8172\,M$$

Final solution = diluted solution

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

\# of moles of $(NH_4)_2SO_4$ in the stock solution used $= (0.8172\,M)(10.00 \times 10^{-3}\,L)$

$= 8.172 \times 10^{-3}\,mol$

$\therefore$ 8.172×10^{-3} mol of $(NH_4)_2SO_4$ was added and the final volume of diluted solution was 60.00 mL = 60.00×10^{-3} L

$$\text{molarity of diluted solution} = \frac{8.172 \times 10^{-3}\ \text{mol}}{60.00 \times 10^{-3}\ \text{L}}$$

$$= 0.1362\ \text{M}$$

$$= \underline{\underline{1.36 \times 10^{-1}\ \text{M}}}$$

(28) <u>Stock solution</u>

$$1.584\text{g Mn} = 1.584\text{g} \times \left(\frac{\text{mol}}{54.94\text{g}}\right) \approx$$

$$= 2.883 \times 10^{-2}\ \text{mol Mn}$$

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

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

volume of stock solution = 1.000 L

$$\text{molarity of stock solution} = \frac{2.883 \times 10^{-2}\ \text{mol}}{1.000\ \text{L}}$$

$$= \underline{\underline{2.883 \times 10^{-2}\ \text{M}}}$$

Solution A

	stock sol^n taken	sol^n A prepared
molarity	$= 2.883 \times 10^{-2}$ M	M_A (say)
volume	$= 50.00$ mL $= 50.00 \times 10^{-3}$ L	1000.0 mL $= 1.0000$ L

$$\left.\begin{array}{l}\text{\# of moles of}\\ \text{solute}\end{array}\right\} = \underbrace{(2.883 \times 10^{-2}\,M)(50.00 \times 10^{-3}\,L)} = M_A\,(1.0000\,L)$$

$= ~~1.4416 mol~~$

$= 1.4415 \times 10^{-3}$ mol

$$\therefore \quad M_A = \frac{1.4415 \times 10^{-3}\ \text{mol}}{1.0000\ \text{L}}$$

$$= \underline{\underline{1.442 \times 10^{-3}\ M}}$$

Solution B

	solution A taken	sol^n B prepared
molarity	1.4415×10^{-3} M	M_B
volume	10.00 mL $= 10.00 \times 10^{-3}$ L	250.0 mL $= 250.0 \times 10^{-3}$ L

$$\left.\begin{array}{l}\text{\# moles of}\\ \text{solute}\end{array}\right\} = \underbrace{(1.4415 \times 10^{-3}\,M)(10.00 \times 10^{-3}\,L)} = M_B\,(250.0 \times 10^{-3}\,L)$$

$$= 1.4415 \times 10^{-5}\ \text{mol} = M_B\,(250.0 \times 10^{-3}\,L)$$

250.0×10^{-3} L

Solution C

	Solution B taken	Solution C prepared
molarity	5.766×10^{-5} M	M_c
volume	10.00 mL $= 10.00 \times 10^{-3}$ L	500.0 mL $=$ [illegible] $\times 10^{-3}$ L

$$\left.\begin{array}{l}\text{\# moles of} \\ \text{solute}\end{array}\right\} = \underbrace{(5.766 \times 10^{-5}\,\text{M})(10.00 \times 10^{-3}\,\text{L})} = M_C\,(500.0 \times 10^{-3}\,\text{L})$$

$$= 5.766 \times 10^{-7}\,\text{mol} = M_C\,(500.0 \times 10^{-3}\,\text{L})$$

$$M_C = \frac{5.766 \times 10^{-7}\,\text{mol}}{500.0 \times 10^{-3}\,\text{L}} = \underline{\underline{1.153 \times 10^{-6}\,\text{M}}}$$