

Last name: \_\_\_\_\_ First name: \_\_\_\_\_

Student Number: \_\_\_\_\_

Lab Day and Section(1 point) : \_\_\_\_\_

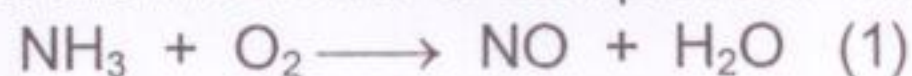
Lab TA name(1 point) : \_\_\_\_\_

# Chemistry 1311 D -2

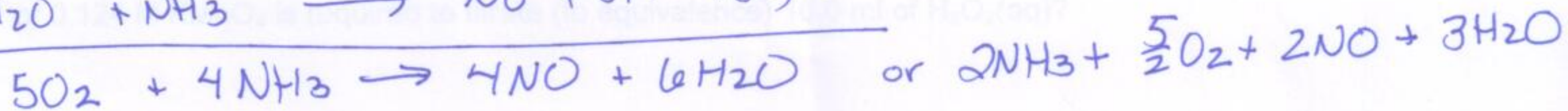
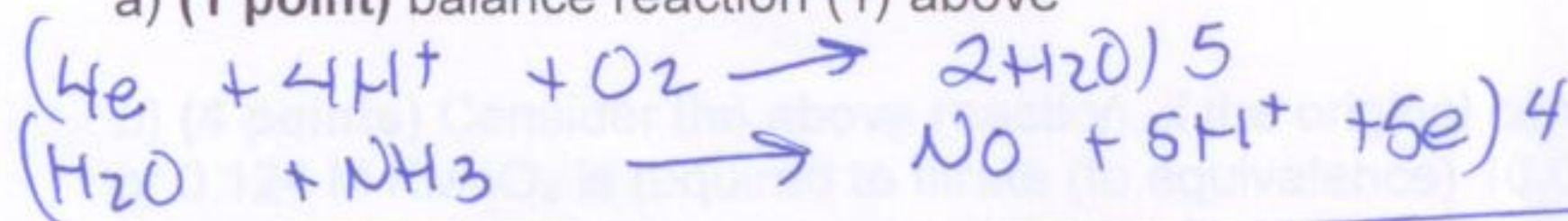
## Midterm October 2012

1. Nitric acid is used in the production of fertilizers and explosives. One step in the industrial production of nitric acid is the reaction of ammonia with molecular oxygen to form nitrogen monoxide.

In a study of this reaction, a chemist mixed 220 g ammonia ( $\text{NH}_3$ ) with 100 L of oxygen at  $30^\circ\text{C}$  and 201.5 kPa and allowed them to react to completion. The unbalanced chemical reaction is:



a) (1 point) balance reaction (1) above



b) (4 points) What masses of NO and  $\text{H}_2\text{O}$  are produced?

$$\text{NH}_3 \quad 220\text{g} \quad 17\text{g/mol} \quad n = \frac{220}{17} = 12.94 \text{ moles}$$

$$O_2 \quad n = \frac{PV}{RT} = \frac{201.5 \times 10^3 \times 100 \times 10^{-3} \text{ m}^3}{303 (8.314)} = 7.999 \text{ moles}$$

If I were to use all the  $\text{NH}_3$

$$n_{O_2} = n_{NH_3} \times \frac{5O_2}{4NH_3} = 16.18 \text{ moles } O_2 \text{ needed, only have } 8$$

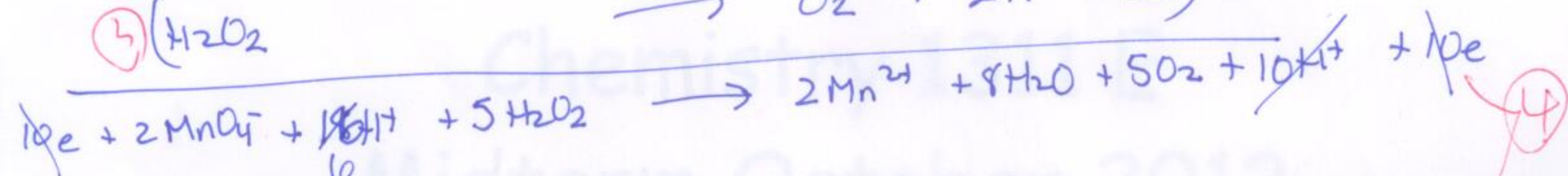
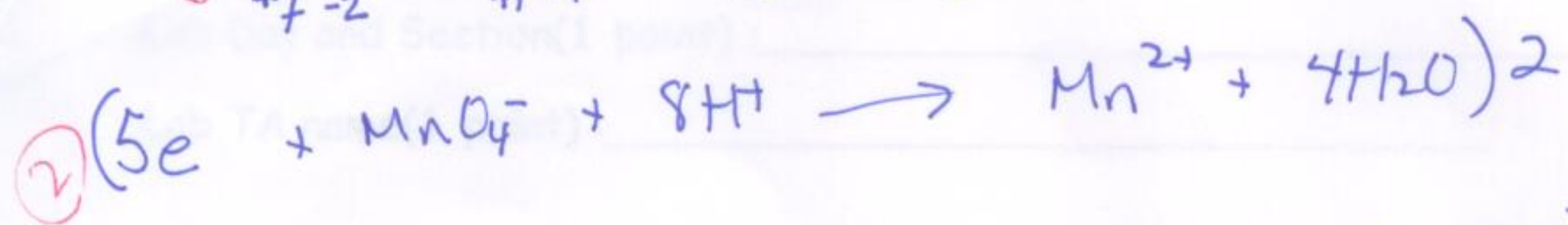
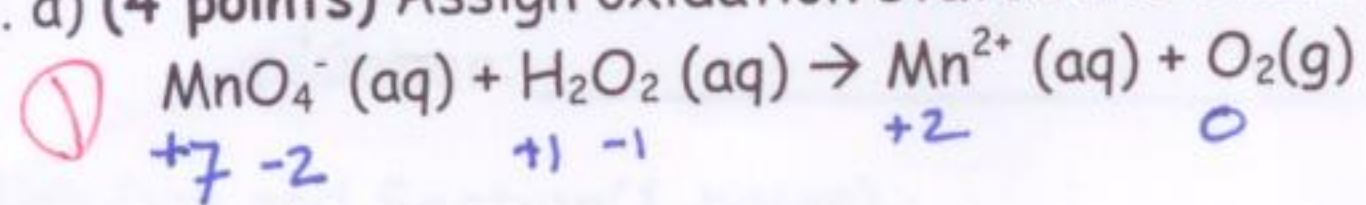
$O_2$  limiting.

$$n_{NO} = n_{O_2} \times \frac{4NO}{5O_2} = 7.999 \times \frac{4}{5} = 6.4 \text{ moles NO}$$

$$\text{mass}_{NO} = n_{NO} \times MM_{NO} = 6.4 \times 30 \text{ g/mol} = 192 \text{ g NO}$$

$$\text{mass}_{H_2O} = n_{H_2O} \times MM_{H_2O} = n_{O_2} \times \frac{6H_2O}{5O_2} \times 18 \text{ g/mol} = 8 \times \frac{6}{5} \times 18 = 173 \text{ g H}_2\text{O}$$

2. a) (4 points) Assign oxidation states and balance the following reaction in acid:

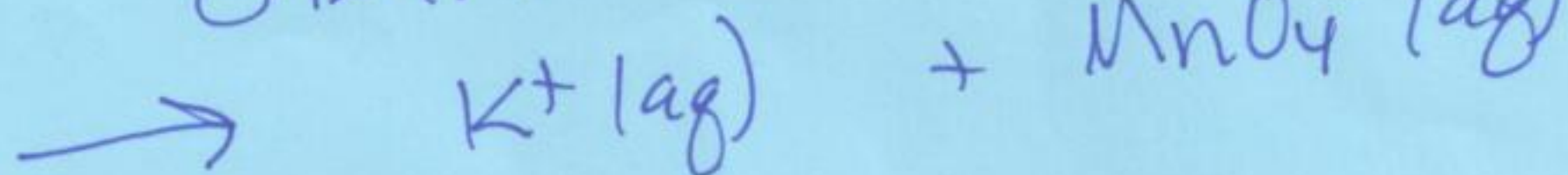
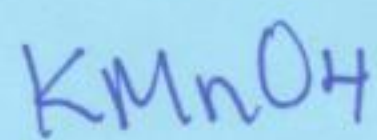


b) (4 points) Consider the above reaction, if the original concentration of  $\text{H}_2\text{O}_2$  was 0.52 M, what volume of 0.124 M  $\text{KMnO}_4$  is required to titrate (to equivalence) 10.0 ml of  $\text{H}_2\text{O}_2(\text{aq})$ ?

10ml  $\text{H}_2\text{O}_2$       0.52 M

V?

0.124 M



$n_{\text{MnO}_4^-}$  ?

$n_{\text{H}_2\text{O}_2} = 0.52 \text{ M} \times 10 \times 10^{-3} \text{ l} = 5.2 \times 10^{-3} \text{ moles}$

$n_{\text{MnO}_4^-} = n_{\text{H}_2\text{O}_2} \times \frac{2\text{MnO}_4^-}{5\text{H}_2\text{O}_2}$   
 $= 5.2 \times 10^{-3} \times 2/5 = 2.08 \times 10^{-3} \text{ moles}$

$[\text{MnO}_4^-] = 2.08 \times 10^{-3} = 0.124 \text{ M} \times V$   
 $V = 0.0168 \text{ l} ; 16.8 \text{ ml}$

3. Combustion analysis of a 1.70 g of a compound formed as a by-product of the industrial synthesis of polymers produced 1.32 g of  $\text{CO}_2$  and 0.631 g of  $\text{H}_2\text{O}$ . The mass percent of iodine in the compound was determined by converting the iodine in a 0.850 g sample of the compound into 1.15 g of lead(II)iodide,  $\text{PbI}_2$ .

a) (2 points) What is the % by mass of C in the sample?

$$n_{\text{CO}_2} = \frac{1.32}{44} = 0.03 \text{ moles} ; \text{ mass C} = 0.03 \times 12 = 0.36 \text{ g}$$

$$\% \text{ C} = 21.18\%$$

b) (2 points) What is the % by mass of H in the sample?

$$n_{\text{H}_2\text{O}} = \frac{0.631}{18 \text{ g/mol}} = 0.03505 \text{ moles} ; \text{ mass H} = 0.03505 \times 1.008 \times 2$$

$$\% \text{ H} = \cancel{2.08\%} \quad 4.12\%$$

$$= 0.0353 \text{ g} \times 2 = 0.0701$$

c) (2 points) What is the % by mass of I in the sample?

$$\text{PbI}_2 = 207.2 + 2(126.9) = 461$$

$$n_{\text{PbI}_2} = 1.15 \text{ g} / 461 \text{ g/mol} = 2.495 \times 10^{-3}$$

$$\text{mass I} = 4.989 \times 10^{-3} \times 126.9 = 0.633 \text{ g}$$

$$\% \text{ I} = 0.633 / 0.850 = 74.48\%$$

$$n_{\text{I}} = 2 \times 2.495 \times 10^{-3} = 4.989 \times 10^{-3}$$

d) (1 point) Could this compound also contain oxygen? Explain your answer.

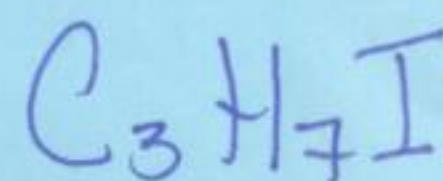
$$100 - 21.18 - \cancel{4.12} - 74.48 =$$

$$100 - 99.78 = 0.22\%$$

No  $\rightarrow$   $\% \text{ C} + \% \text{ H} + \% \text{ I} \approx 100$ .

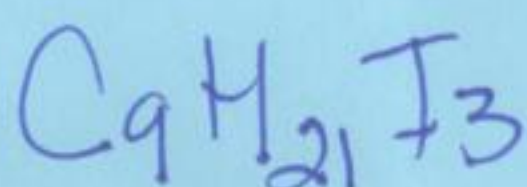
e) (2 points) What is the empirical formula of the compound?

100g	C: 21.18g	= 1.765 moles	$\div 0.5869 = 3.007$
	H 4.12g	= 4.087	$\div 0.5869 = 6.96$
	I 74.48g	= 0.5869	$\div 0.5869 = 1$

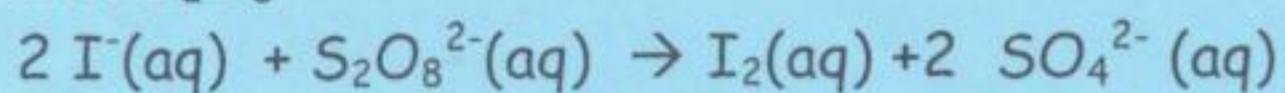


f) (1 point) If the molar mass of the compound was found experimentally to be 527 g what is its molecular formula.

$$\frac{\text{MM}}{\text{EM}} = \frac{527 \text{ g/mol}}{170} = 3.1 \sim 3$$



5. (7 points) The following initial rate information was collected at 25 °C for the reaction of I<sup>-</sup> with S<sub>2</sub>O<sub>8</sub><sup>2-</sup>:



[I <sup>-</sup> ] <sub>0</sub> (M)	[S <sub>2</sub> O <sub>8</sub> <sup>2-</sup> ] <sub>0</sub> (M)	Initial Rate M/min	
0.125	0.150	4.4 × 10 <sup>-2</sup>	1
0.375	0.150	2.3 × 10 <sup>-1</sup>	2
0.125	0.050	1.5 × 10 <sup>-2</sup>	3

Determine the rate law, order and evaluate the rate constant for this reaction.

$$\frac{R_1}{R_2} = \left( \frac{[I^-]_1}{[I^-]_2} \right)^x = \left( \frac{4.4 \times 10^{-2}}{2.3 \times 10^{-1}} \right)^x$$

$$(\cdot 333)^x = (0.191)^x$$

$$x \log \cdot 333 = \log \cdot 191$$

$$x(-0.4815) = -0.7189$$

$$x = 1.5$$

$$\frac{R_1}{R_3} = \left( \frac{[S_2O_8^{2-}]_1}{[S_2O_8^{2-}]_3} \right)^y$$

$$\frac{4.4}{1.5} = \left( \frac{0.150}{0.050} \right)^y$$

$$2.93 = 3^y$$

$$x = 1$$

$$R = k [\text{S}_2\text{O}_8^{2-}]^1 [\text{I}^-]^{1.5}$$

M<sup>2.5</sup> (1.5)

$$k = \frac{R}{[\text{S}_2\text{O}_8^{2-}][\text{I}^-]^{1.5}} = \frac{4.4 \times 10^{-2}}{0.15(\cdot 125)^{1.5}} = 6.637$$

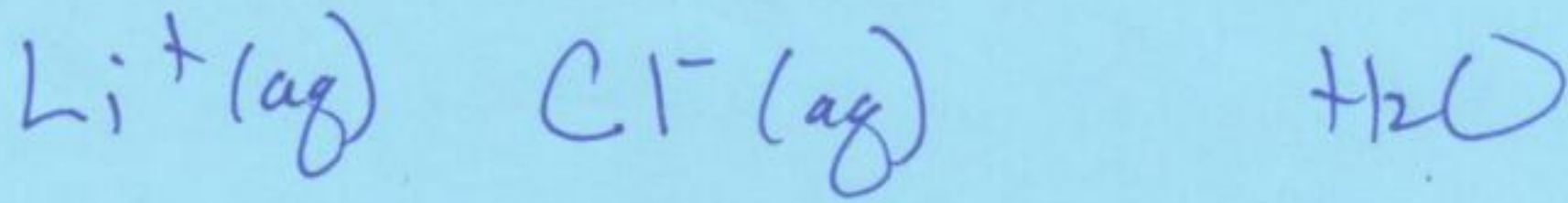
$$= \frac{2.3 \times 10^{-1}}{0.15(\cdot 375)^{1.5}} = 6.677$$

$$R = 6.65 \text{ M}^{-1.5} \text{ s}^{-1}$$

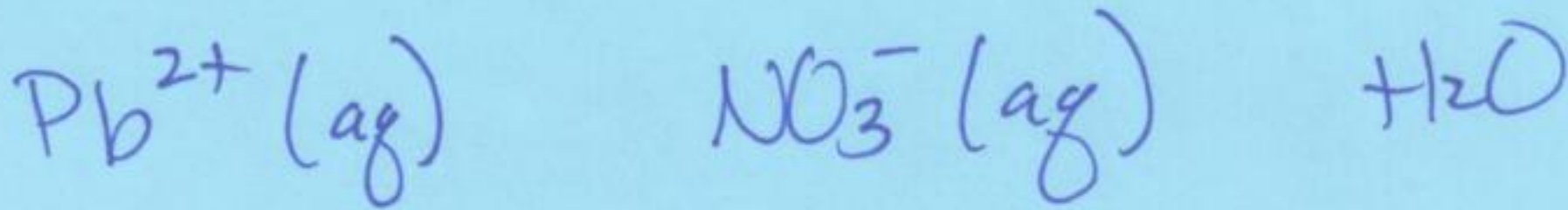
order 2.5 (1.5)

4. You have three beakers each containing a 200 ml of water. To the first you add 850 mg of solid LiCl, to the second you add 6.8 g of solid  $\text{Pb}(\text{NO}_3)_2$  and to the third, you add 2.4 g of solid  $\text{MgSO}_4$ .

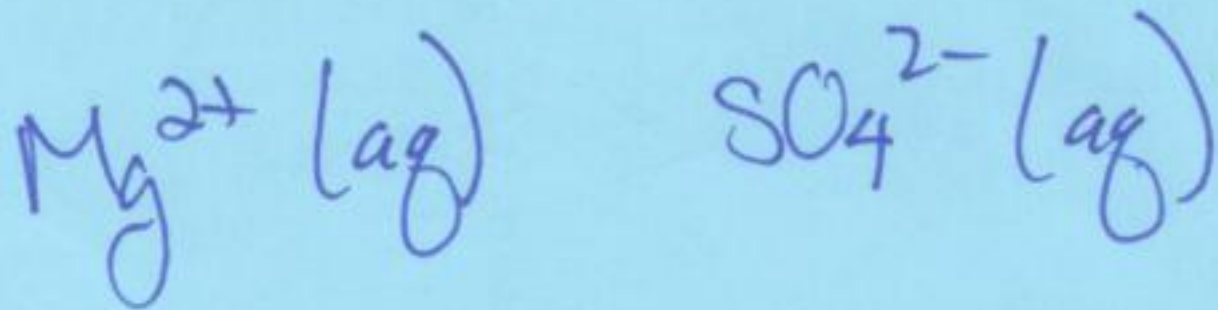
a) (1 point) What species will you find in the first beaker?



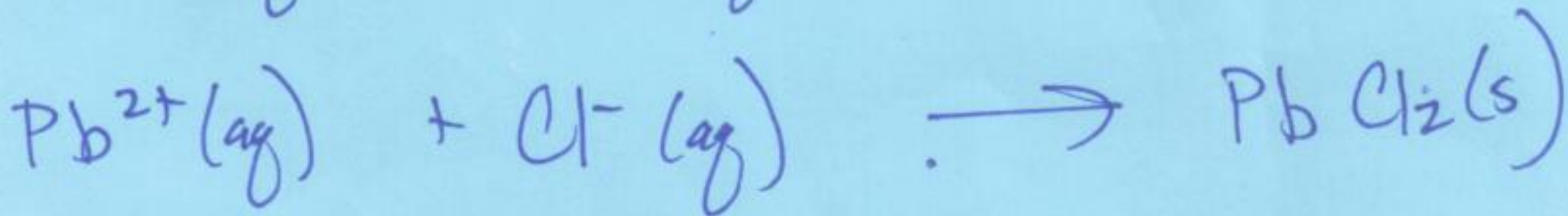
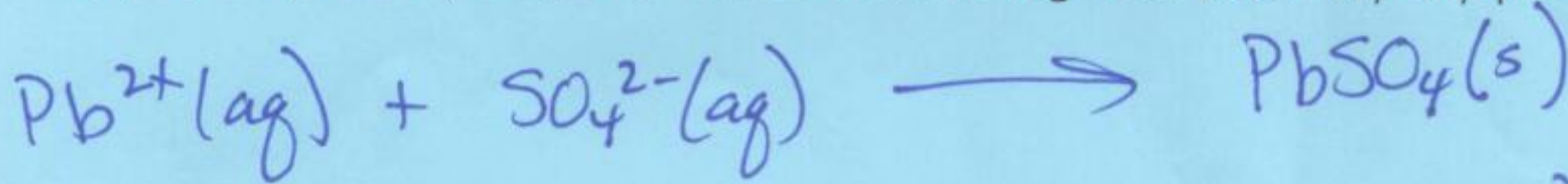
b) (1 point) in the second?



c) (1 point) in the third?



d) (3 point) Next you mix all three beakers together, identify any precipitates that could form.



e) BONUS: Next you throw in a handful of  $(\text{NH}_4^+)_3\text{PO}_4$ . What happens?

Puts  $\text{NH}_4^+(\text{aq})$  &  $\text{PO}_4^{3-}(\text{aq})$  into sol<sup>n</sup>

$$n \text{ LiCl} = \frac{0.850}{(7+35.5)} = 0.02 \text{ moles}$$

$$n \text{ Pb}^{2+} = \frac{6.8 \text{ g}}{(207.2 + 2(14 + 3(16)))} = 0.02 \text{ moles}$$

$$n \text{ SO}_4^{2-} = \frac{2.4 \text{ g}}{(32 + 4(16))} = 0.025 \text{ moles}$$

$\text{Cl}^-$  gets rid of  $\frac{1}{2} \text{ Pb}^{2+}$   
 $\text{SO}_4^{2-}$  gets of rest.

$\text{PO}_4^{3-}$  will ppt  $\text{Mg}^{2+}$   
 $\text{Mg}_3(\text{PO}_4)_2(\text{s})$

6. (4 points) Calculate the heat in kilojoules to vaporize 40 g  $\text{CCl}_4$  initially at  $25^\circ\text{C}$ .

You might be interested that:

the boiling point of  $\text{CCl}_4$  is  $77^\circ\text{C}$ ,

the melting point is  $-23^\circ\text{C}$

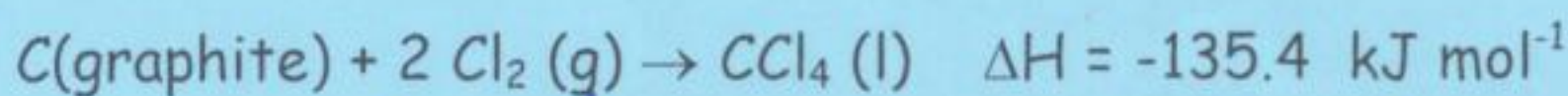
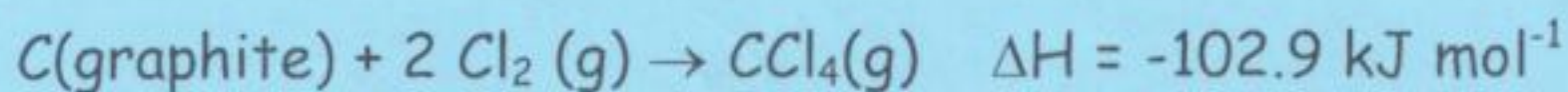
the molar heat capacity of  $\text{CCl}_4(\text{s})$  is  $44.22 \text{ J mol}^{-1}\text{C}^{-1}$ ;

the molar heat capacity of  $\text{CCl}_4(\text{l})$  is  $132.4 \text{ J mol}^{-1}\text{C}^{-1}$  ←

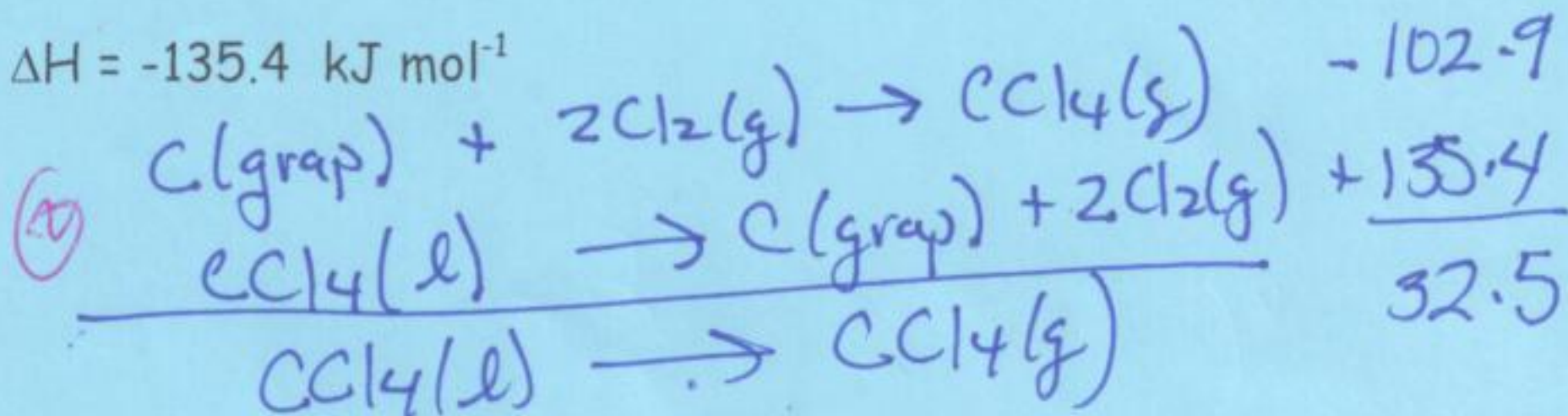
the molar heat capacity of  $\text{CCl}_4(\text{g})$  is  $82.7 \text{ J mol}^{-1}\text{C}^{-1}$ .

$$\text{MM} = 153.81 \text{ g/mol}$$

$$n_{\text{CCl}_4} = 40/153.8 = 0.260 \text{ moles}$$



$\text{CCl}_4$  is a liquid.



$$\text{heat} = n C_m \Delta T + n \Delta H$$

warm

$$= 0.260 (132.4 \text{ J mol}^{-1}\text{C}^{-1}) (77-25) + 0.260 (325 \times 10^3 / \text{mol})$$

$$= 1.79 \text{ kJ} + 8.45 \text{ kJ}$$

$$= \underline{\underline{10.24 \text{ kJ}}}$$

Question	Points Possible	Points Earned
Title page	2	
1	5	
2	8	
3	10	
4	6	
5	7	
6	4	
Total	40	