PERIODIC TABLE OF THE ELEMENTS

☐ metals
☐ non-metals
☐ half-metals

*Average atomic mass for the naturally occurring mix of isotopes.

S	32,06
Zn	65,38
Ge	72,59
Tc	98

Example: Atomic number — 14; Matter — Si; Atomic mass (u) — 28,09

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1,008																	2 He 4,003
2	3 Li 6,941	4 Be 9,012											5 B 10,81	6 C 12,01	7 N 14,01	8 O 16,00	9 F 19,00	10 Ne 20,18
3	11 Na 22,99	12 Mg 24,31											13 Al 26,98	14 Si 28,09	15 P 30,97	16 S 32,07*	17 Cl 35,45	18 Ar 39,95
4	19 K 39,10	20 Ca 40,08	21 Sc 44,96	22 Ti 47,88	23 V 50,94	24 Cr 52,00	25 Mn 54,94	26 Fe 55,85	27 Co 58,93	28 Ni 58,69	29 Cu 63,55	30 Zn 65,39*	31 Ga 69,72	32 Ge 72,61*	33 As 74,92	34 Se 78,96	35 Br 79,90	36 Kr 83,80
5	37 Rb 85,47	38 Sr 87,62	39 Y 88,91	40 Zr 91,22	41 Nb 92,91	42 Mo 95,94	43 Tc 98,91*	44 Ru 101,1	45 Rh 102,9	46 Pd 106,4	47 Ag 107,9	48 Cd 112,4	49 In 114,8	50 Sn 118,7	51 Sb 121,8	52 Te 127,6	53 I 126,9	54 Xe 131,3
6	55 Cs 132,9	56 Ba 137,3	57 La 138,9	72 Hf 178,5	73 Ta 180,9	74 W 183,9	75 Re 186,2	76 Os 190,2	77 Ir 192,2	78 Pt 195,1	79 Au 197,0	80 Hg 200,6	81 Tl 204,4	82 Pb 207,2	83 Bi 209,0	84 Po 209	85 At 210	86 Rn 222
7	87 Fr 223	88 Ra 226,0	89 Ac 227,0	104 Db or Unq 261	105 Jl or Unp 262	106 Rf or Unh 263	107 Bh or Uns 262	108 Hn or Uno 265	109 Mt or Une 266									

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58 Ce 140,1	59 Pr 140,9	60 Nd 144,2	61 Pm 145	62 Sm 150,4	63 Eu 152,0	64 Gd 157,3	65 Tb 158,9	66 Dy 162,5	67 Ho 164,9	68 Er 167,3	69 Tm 168,9	70 Yb 173,0	71 Lu 175,0
90 Th 232,0	91 Pa 231,0	92 U 238,0	93 Np 237,0	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 254	100 Fm 257	101 Md 258	102 No 259	103 Lr 260

1 Techniques and Measurements

Name: Dua'a Khaled Abu Jamé Section ...6...

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. A metal sphere weighing 15.45 g is added to 21.27 mL water in a graduated cylinder. The water level rises to 24.78 mL. Calculate the density of the metal.

the volume of sphere = V after − V before = 24.78 − 21.27 = 3.51 ml

density = mass / volume ⇒ d = 15.45 / 3.51 = 4.40 g/mL

2. An empty beaker weighs 32.4257 g. A 10.00 mL sample of unknown liquid is transferred to the beaker. The total mass of the beaker and liquid sample was 40.1825 g. Calculate the density of the unknown liquid.

the mass of unknown liquid = 40.1825 − 32.4257 = 7.7568 g

density = mass / volume ⇒ density = 7.7568 / 10.00 = 0.77568 g/ml

3. A term that is easily confused with density is **specific gravity**. What is meant by specific gravity? What are the units of specific gravity?

Specific gravity = density / water density unit Less

1 Techniques and Measurements

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculation

A. Determination of the Density of Pure Liquid

	Trial (1)	Trial (2)
Mass of beaker	30.84 ±0.01 g	30.85±0.01 g
Mass of beaker + water	40.72±0.01 g	40.78±0.01 g
Mass of water	9.88±0.01 g	9.93±0.01 g
Volume of water	10 mL	10 mL
Temperature of water	24 °C	25° °C
Density	0.988 g/mL	0.993 g/mL
Average density	0.9905 g/mL	
Handbook density	0.99707 g/mL	
ID number of unknown liquid		
Mass of beaker	30.83±0.01 g	30.85±0.01 g
Mass of beaker + unknown liquid	40.03±0.01 g	40.05±0.01 g
Mass of unknown liquid	9.2 ±0.01 g	9.2±0.01 g
Volume of unknown liquid	10 mL	10 mL
Density of unknown liquid	0.92 g/mL	0.92 g/mL

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B. Density of Solutions

	Trial (1)	Trial (2)
Mass of beaker	g	g
Volume of solution	mL	mL
Mass of beaker + solution	g	g
Mass of solution	g	g
Temperature of solution	°C	°C
Density of solution	g/mL	g/mL

C. Density of Solids

	Trial (1)	Trial (2)
Mass of beaker	60.65±0.01 g	60.65±0.01 g
Mass of beaker + solid pieces	63.83±0.01 g	65.79±0.01 g
Mass of solid pieces	3.18±0.01 g	5.14±0.01 g
Initial water level in the graduated cylinder	50 mL	50 mL
Final water level in the graduated cylinder	51 mL	52 mL
Volume of solid pieces	1 mL	2 mL
Density of the solid substance	3.18 g/mL	2.57 g/mL

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Questions:

1. What error would be introduced into the density of the metal pellets if you had not shaken the pellets to remove adhering air bubbles? Would the density be too high or too low?

If we had not shaken the pellets, there would be air bubbles → the volume increases → the density would be too low than the true density ($density = \frac{mass}{volume}$)

2. The density of diamond is 3.51 g/cm^3 and the density of lead is 11.3 g/cm^3. If equal masses of diamond and lead were transferred to equal volumes of water in graduated cylinders, which graduated cylinder would have the highest volume reading? Explain.

$d_d = 3.51\ g/cm^3$ $d_L = 11.3\ g/cm^3$ $\{ mass_d = mass_L$

$d_L > d_d$ V_w the same for both d & L

$\frac{mass_L}{V_{w+L}} > \frac{mass_d}{V_{w+d}}$ but $mass_d = mass_L$

$V_{w+L} < V_{w+d}$

$V_w + V_L < V_w + V_d$

$V_{lead} < V_{diamond}$

2 Formula of a Hydrate

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What are alums? Give examples other than potassium alum.

Alums are a class of hydrates, consist of a double salt with the general formula ($M'M''(SO_4)_2 \cdot xH_2O$) M': a univalent cation M''': a trivalent cation

2. What are the hydrates? Give few examples of metal salt hydrates.

Hydrate is a compound that has crystallized from aqeous solution with weekly bound water molecules contained in the crystal. They could be transition metal salts as $CuSO_4 \cdot 5H_2O$ or non-transition metal salt as $NaC_2H_3O_2 \cdot 3H_2O$ or a covalent compound.

3. Potassium ferric alum has the formula: $KFe(SO_4)_2 .xH_2O$. A sample of 1.26 g of this alum was heated in a crucible to get a constant mass. The mass of the anhydrous salt produced ($KFe(SO_4)_2$) was 0.72g. Calculate the value "x" in the formula of the alum.

mass of water = 1.26 − 0.72 = 0.54 g $\Rightarrow$ # mole of $H_2O = \frac{0.54}{18} = 0.03$ mol

mole of anhydrous = $\frac{0.72}{287}$ = 2.5×10^{-3} mol

$x = \frac{\text{\# mole of water}}{\text{\# moles of anhydrate}} = \frac{0.03}{0.0025} = 12$

2 Formula of a Hydrate

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

A. Potassium Alum:

Mass of empty crucible (m_1)	18.73±0.01 g
Mass of crucible and the alum (m_2)	19.76±0.01 g
Mass of crucible and anhydrous salt (m_3)	19.30 ±0.01 g
Mass of alum ($m_2 - m_1$)	1.03±0.01 g
Mass of anhydrous salt ($m_3 - m_1$)	0.57±0.01 g
Mass of water lost upon heating ($m_2 - m_3$)	0.46 ±0.01 g
Number of moles of water lost upon heating $\frac{m_2 - m_3}{18}$	0.03 mol
Number of moles of anhydrous salt ($KAl(SO_4)_2$) $= \frac{m_3 - m_1}{258}$	2.2×10^{-3} mol
Percentage of water of crystallization, by mass $\frac{m_2 - m_3}{m_2 - m_1} \times 100\%$	44.66 %
The value 'x" in the formula, (number of moles of water of crystallization / number of moles of anhydrous salt) $\frac{\#\ mole\ of\ H_2O}{\#\ mole\ anhydrous}$	11.63 ≃ 12

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B. Unknown Hydrate:

Unknown number: --------------

Mass of crucible (m_4) 18.72 ± 0.01 g

Mass of crucible and the hydrate (m_5)	19.75 ± 0.01 g
Mass of crucible and anhydrous salt (m_6)	19.32 g
Mass of ~~anhydrous~~ hydrate salt (m_5 – m_4)	1.03 ± 0.01 g
Mass of water lost upon heating (m_5 – m_6)	0.43 ± 0.01 g
Percentage of water of crystallization, by mass $\frac{m_5 - m_6}{m_5 - m_4}$	41.74 %

14.30

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QUESTIONS:

1. What is the effect on the calculated value of "x" if the dehydration of the alum is not complete

If dehydration is not complete ⇒ mass of water loss ↓
⇒ # moles of H_2O ↓ ⇒ X ↓ (decrease)
$X = \frac{\#\ mole\ of\ H_2O}{\#\ mole\ of\ anhydrate}$

88.7 $NiSO_4$

2. A student heated 1.40 g of hydrated nickel(II) sulfate in a crucible to get 0.77 g of anhydrous salt. What is the formula of the salt? (show your work)

mass of H_2O = 1.40 − 0.77 = 0.63 g ⇒ # mole = $\frac{0.63}{18}$ =
mole of anhydrate = $\frac{0.77}{154.7}$ = 4.9×10^{-3} mol

$X = \frac{\#\ mole\ of\ H_2O}{\#\ mole\ of} = \frac{0.035}{4.9\times10^{-3}} = 7.14 \approx 7$

⇒ $NiSO_4 \cdot 7H_2O$

3 The Empirical Formula of an Oxide

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What is the mole and molar mass? What are the molar masses of magnesium and atomic oxygen?

— The mole is the quantity of a given substance that contains as many molecules or formula units as the number of atoms in exactly 12 g carbon-12 = Avogadro's number = 6.022×10^{23}

— The molar mass of a substance is the mass of one mol of the substance.

2. When 0.192 g of phosphorus is burned, 0.440 g of a white oxide is obtained. (a) Write a balanced chemical equation for the reaction of Phosphorus with molecular oxygen based on this empirical formula. (b) Determine the empirical formula of the oxide.

$4P + 5O_2 \xrightarrow{\Delta} 2P_2O_5$

mass of O_2 = 0.440 − 0.192 = 0.248 g | # moles of O_2 = $\frac{0.248}{32}$

mole of (P) = $\frac{0.192}{30.9}$ = 6.21×10^{-3} mol

$\frac{0.0155}{0.00621}$ $\frac{0.00621}{0.00621}$

(2.5 : 1) 2

5 : 2

P_2O_5

3 The Empirical Formula of an Oxide

Name: Dua'a Khaled Abu-Jame' Section ....

Lab. Instructor .... Date ....

Results and Calculations

Mass of empty crucible (after first heating) m_1'	19.25 ± 0.01 g
Mass of empty crucible (after second heating) m_1''	19.25 ± 0.01 g
Mass of empty crucible (average) m_1	19.25 ± 0.01 g
Mass of crucible and Mg m_2	19.45 ± 0.01 g
Mass of Mg $m_2 - m_1$	0.2 ± 0.01 g
Moles of Mg $\frac{m_2 - m_1}{24.3}$	8.3×10^{-3} mol
Mass of crucible and Mg-oxide (after first heating) m_3'	19.57 ± 0.01 g
Mass of crucible and Mg-oxide (after second heating) m_3''	19.55 ± 0.01 g
Mass of crucible and Mg-oxide (average) m_3	19.56 ± 0.01 g
Mass of Mg-oxide produced $m_3 - m_1$	0.31 ± 0.01 g
Mass of oxygen gained $(m_3-m_1)-(m_2-m_1) = m_3 - m_2$	0.11 ± 0.01 g
Moles of oxygen $\frac{m_3 - m_2}{15.99}$	6.8×10^{-3} Mol
Empirical formula of the oxide	MgO
Mass percent of Mg in the oxide (x_1) (Experimentally) $\frac{m_2 - m_1}{m_3 - m_1} \times 100\%$	64.52 %
Mass percent of Mg in the oxide (x_2) (calculated for MgO) $\frac{24}{40}$	60 %
Percentage error = $(x_2-x_1)/x_2$ X 100%	7.53 %

0.00675 0.0083

$\%Mg = \frac{mass\ Mg}{mass\ MgO}$

QUESTIONS:

1. If water had not been added to your initial product, what error in the determined percentage of magnesium would have resulted (that is, if part of the product has been magnesium nitride)?Explain.

If water had not been added → a part of the product would be Mg_3N_2 → mass of MgO would be bigger → percentage of magnesium would ~~decreased~~ ↓ ~~be~~ increase.

$\%Mg = \frac{mass\ Mg}{mass\ MgO \downarrow}$ ⇒ % Mg ↑

2. If large amount of magnesium oxide had been lost during the heating of the crucible, would this have made the calculated %Mg in the product too high or too low? Explain.

MgO ↓ Mg $\%Mg = \frac{Mass\ of\ Mg \downarrow}{Mass\ of\ MgO \downarrow}$

⇒ % Mg ~~too high~~ would be constant because decreasing in MgO include decreasing in Mg.

4 Limiting Reactant

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. $Na_3PO_4.12H_2O$ and $BaCl_2.2H_2O$ are examples of compounds known as hydrates. What are the hydrates? Give other examples of hydrates.

Hydrate is a compound that has crystallized from aqeous solution with weakly bound water molecules contained in the crystal.

examples: $NaC_2H_3O_2 . 3H_2O$ / $KAl(SO_4)_2 . 12H_2O$

2. A mixture of 0.65 g $BaCl_2.\ 2H_2O$ and 0.35 g $Na_3PO_4.12\ H_2O$ was dissolved in water. Which of the two reactants is the limiting reactant? Calculate the mass of $Ba_3(PO_4)_2$ precipitate formed.

$3BaCl_2 + 2Na_3PO_4 \longrightarrow Ba_3(PO_4)_2 + 6NaCl$

$0.65 g\ BaCl_2 \times \frac{1\ mol\ BaCl_2 \cdot 2H_2O}{244.2 g\ BaCl_2} \times \frac{1\ mol\ Ba_3(PO_4)_2}{3\ mol\ BaCl_2} = 8.9 \times 10^{-4}\ mol\ (Ba_3(PO_4)_2)$

$0.35 g\ Na_3PO_4 \times \frac{1\ mol\ Na_3PO_4 \cdot 12H_2O}{380.2 g\ Na_3PO_4 \cdot 12H_2O} \times \frac{1\ mol\ Ba_3(PO_4)_2}{2\ mol\ Na_3PO_4} = 4.6 \times 10^{-4}\ mol\ Ba_3(PO_4)_2$

~~$BaCl_2 . 2H_2O$~~ Na_3PO_4 is the limiting reactant.

the $Ba_3(PO_4)_2$ mass $= 4.6 \times 10^{-4}\ mol \times \frac{601.9\ g}{1\ mol} = 0.277 g$

4 Limiting Reactant

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

A. Precipitation of $Ba_3(PO_4)_2$:

Unknown Number: 15

Mass of salt mixture (m_1)	1.00 ± 0.01 g
Mass of filter paper (m_2)	0.40 ± 0.01 g
Mass of filter paper and $Ba_3(PO_4)_2$ (m_3)	0.62 ± 0.01 g

B. Determination of the Limiting Reactant:

Limiting reactant in salt mixture is ...$BaCl_2$...

Excess reactant in salt mixture is ...Na_3PO_4...

Mass of $Ba_3(PO_4)_2$ precipitated ($m_3 - m_2$)	0.22 ± 0.01 g
Number of moles of $Ba_3(PO_4)_2$ (n_1)	3.66×10^{-4} mol

244.2

(a) If the limiting reactant is $BaCl_2$. $2H_2O$: $n_2 = 3n_1$

Number of moles of $BaCl_2.2H_2O$ reacted (n_2)	1.098×10^{-3} mol
Mass of $BaCl_2.2H_2O$ reacted (m_4)	0.27 ± 0.01 g

$n_2 \times 244.2$

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(b) If the limiting reactant is $Na_3PO_4.12H_2O$.

Number of mole of $Na_3PO_4.12\ H_2O$ reacted(**n_3**) $n_3 = 2n_1$	moles
Mass of $Na_3PO_4.12H_2O$ reacted(**m_5**)	± 0.01 g
Mass of salt mixture(**m_1**)	1.00 ± 0.01 g

(c) Mass percent calculation.

Mass of excess reactant $m_1 - m_4$ (Depending on which reactant is the limiting reactant)	0.73 ± 0.01 g
Mass percentage of $BaCl_2.2H_2O$. in the mixture $\frac{m_4}{m_1} \times 100\%$	27 %
Mass percentage of $Na_3PO_4 12H_2O$ in the mixture. $\frac{m_1 - m_4}{m_1} \times 100$	73 %

$\frac{m_5}{m_1} \times 100\% \rightarrow \frac{m_1 - m_5}{100}$ Na_3PO_4 $BaCl_2$

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QUESTIONS:

1. Calculate the mass of $Ba_3(PO_4)_2$ produced from the reaction of 0.85 g $Na_3PO_4.12H_2O$ with excess $BaCl_2.2H_2O$. What is the purpose of heating the mixture in step 3 for 20 minutes?

$2Na_3PO_4 + 3BaCl_2 \rightarrow 6NaCl + Ba_3(PO_4)_2$

$0.85 g\ Na_3PO_4.12H_2O \times \frac{1\ mol\ Na_3PO_4.12H_2O}{380\ g} \times \frac{1\ mol\ Ba_3(PO_4)_2}{2\ mol\ Na_3PO_4} \times \frac{601}{1\ mol}$

$= 0.67\ g$

The purpose of heating is increase the size of precipitation particles so that they can be separated of the solution in the filteration.

2. What is the purpose of washing the precipitate with hot water in step 4? How would the reported percentage of the excess reactant be affected if the precipitate was not washed in this step?

To collect the whole amount of the precipitate of the beaker and the funnel.
If the precipitate was not washed → the mass of $Ba_3(PO_4)_2$ will be less than the true one → n ↓ → limiting reactant → the excess reactant ↑ → the excess percentage increases

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5 Determination of Acetic Acid in Vinegar

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Why was the standard NaOH solution not prepared by calculating the amount of solid NaOH needed for 100 mL of solution, weighing it accurately, and making it up to exactly 100 mL of total volume?

NaOH reacts with water that will changes

because there would be an excess mass of water in NaOH so we use the end point which is more accurate

2. Why not simply rinse the buret with distilled water rather than the solution to be used in it?

to avoid that there is distilled water from the buret wall

3. Why does the volume of water added to potassium hydrogen phthalate not have to be measured carefully?

because we don't need to calculate the volume of water the most important is the presence of KHP to limit the end point then the volume of NaOH we need (the number of moles is the same, KHP is very pure) doesn't change by changing the volume phenolphtalien

5 Determination of Acetic Acid in Vinegar

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

A. Standardization of NaOH Solution

	Trial (I)	Trial (II)
Mass of flask + KHP	86.95 ± 0.01 g	~~[illegible]~~ 87.27 ± 0.01 g
Mass of Flask	86.76 ± 0.01 g	87.08 ± 0.01 g
Mass of KHP	0.19 ± 0.01 g	0.19 ± 0.01 g
Molar mass of KHP	204.22 g/mol	204.22 g/mol
Moles of KHP	9.30×10^{-4} mol	9.30×10^{-4} mol
Initial volume of NaOH	1.85 mL	9.5 mL
Final volume of NaOH	9.5 mL	16.2 mL
Moles of NaOH	9.30×10^{-4} mol	9.30×10^{-4} mol
Molarity of NaOH	0.122 x ~~10^{-4}~~ M	0.139 x ~~10^{-4}~~ M
Average Molarity of NaOH	~~1.305×10^{-4}~~ 0.1305 M	

$M_1 V_1 = M_2 V_2$

$M_1 \times 7.65 = 9.30 \times 10^{-4}$

$M_1 = 1.22 \times 10^{-4}$

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B. Mass percent of acetic acid in vinegar.

	Trial (I)	Trial (II)
Volume of vinegar	10.00 mL	10.00 mL
Initial volume of NaOH	16.2 mL	31.3 mL
Final volume of NaOH	30.8 mL	46.0 mL
Volume of NaOH used	14.6 mL	14.7 mL
Average Molarity of NaOH, From Part A:	1.305×10^{-4} M	
Moles of NaOH used	1.91×10^{-3} mol	1.92×10^{-3} mol
Moles of CH_3COOH in vinegar Reacted with the NaOH	1.91×10^{-3} mol (0.191)	1.92×10^{-3} mol
Molarity of CH_3COOH in vinegar	~~1.91×10^{-4}~~ 0.1911 M	~~$1.92 \times 10^{-}$~~ 0.192 M
Average molarity of acetic acid in vinegar		~~1.91×10^{-4}~~ 0.191 M
Molar Mass of acetic acid, CH_3COOH		60.0 g/mol
Mass of CH_3COOH per liter of vinegar (= M × molar mass; 1000 ml, 1000 g)		~~0.0115~~ 11.5 g
Mass of CH_3COOH per 100 g of vinegar (assume that vinegar has a density of 1.00 g/mL) (=100 ml = 0.1 L)		~~$1.15 \times 10^{-}$~~ 1.15 g
Mass percent acetic acid in vinegar		1.15 %

mass of CH_3COOH per 1L of vinegar × 100 %

1L = 1000 ml = 1000 g ⇐ density = 1.00

$$\%Mass = \frac{0.0115}{1000} \times 100\% =$$

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QUESTIONS:

calculated $\frac{|x_2 - x_1|}{x_2} \times 100\%$ experimently

1. Calculate the percent error would have been in a titration that used 26.65 mL of a solution if a bubble with a volume of 0.30 mL had been swept out of the tip during the titration.

   26.65 − 0.30 = 26.35

   percent error = $\frac{0.30}{26.35} \times 100\%$ = 1.14 %

2. If a solution is diluted, the same number of moles of reagent is still present. Hence, it is possible to calculate a new volume or a new molarity from a known greater molarity. The equation is: $M_1V_1 = M_2V_2$. If you wish to make 500 mL of a 0.20 M solution of a reagent from 1.5 M solution, how much of the latter solution would you use?

   $M_1V_1 = M_2V_2$

   $0.20 \times 0.5 = 1.5 \times V_2$

   $V_2 = 0.0667$ L = 66.7 ml

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6 The Neutralizing Capacity of Antacid Tablets

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Write a balanced equation for the reaction of HCl and $NaHCO_3$

   $HCl + NaHCO_3 \rightarrow NaCl + H_2O + CO_2$

2. How many moles of HCl are needed to react with 0.35 g of $NaHCO_3$?

   $0.35 g\ NaHCO_3 \times \frac{1\ mol\ NaHCO_3}{84.01\ g\ NaHCO_3} \times \frac{1\ mol\ HCl}{1\ mol\ NaHCO_3} = 4.2$

3. Why we do not use strong bases as active ingredients of antacid tablet?

   * because the decreasing in PH wouldn't be so large so we add weak bases. Also using strong bases → PH increases more than 7 but we need PH = 5.4 approximately in the stomatch

   Because NaOH is caustic

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0.331

bromothymol acid-base indicator

6 The Neutralizing Capacity of Antacid Tablets

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

Name of antacid tablet: ....

Ingredients:...

	Trial (I)	Trial (II)
Mass of antacid sample	0.20 ±0.01 g	0.20 ±0.01 G
Volume of 0.150 M HCl solution	50 mL	50 mL
Moles of HCl (used to dissolve antacid)	$\frac{50}{1000} \times 0.15 = 7.5\times10^{-3}$ mol	7.5×10^{-3} Mol
Initial buret reading	14.8 mL	20.6 mL
Final buret reading	20.6 mL	26.8 mL
Volume of NaOH added	5.8 mL	6.2 mL
Moles of NaOH (used to titrate the excess acid)	8.7×10^{-4} mol	9.3×10^{-4} Mol
Moles of excess HCl	8.7×10^{-4} mol	9.3×10^{-4} Mol
Moles of HCl(needed to neutralize the antacid tablet)	6.63×10^{-3} mol	6.57×10^{-3} Mol
Neutralizing capacity of antacid	0.0332 mol HCl/g	0.0329 mol HCl/g
Average (mol HCl/g antacid)	0.0331 mol HCl/g antacid	

10^{-3} mol HCl

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QUESTIONS:

1. Assume the concentration of stomach acid is 0.10M and the neutralizing capacity of the acid is 0.011mol/g. Calculate the volume of stomach acid needed to neutralize 1.0 g antacid tablet.

$\frac{\# mols\ HCl}{g\ antacid} = 0.011$ | $M = \frac{\# mols}{V}$

$\# mols\ HCl = \frac{0.011}{1} = 0.011\ mol\ HCl$ | $0.10 = \frac{0.011}{V}$

$V = 0.11\ L$

0.04 L

2. A student dissolves 0.353 g sample of antacid in 40.00 mL of 0.141 M HCl solution, then back-titrated to the end-point with 8.92 mL of a 0.203 M NaOH solution.

a) Calculate number of moles of acid in the original 40.00 mL of HCl?

$\# mols\ of\ HCl = M \times V$

$= \frac{0.141 \times 40.0}{1000}$

$= 5.64 \times 10^{-3}\ mols\ HCl$

b) How many moles of base were used in the back-titration of excess HCl?

$\# moles\ of\ NaOH = M \times V$

$= \frac{0.203 \times 8.92}{1000} = 1.81 \times 10^{-3}$

c) How many moles of excess HCl I?

$\# moles\ of\ excess\ HCl = \# moles\ of\ NaOH$

$= 1.81 \times 10^{-3}\ mol\ HCl$

98.5

7 Vitamin C Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Explain why cooked fruits and vegetables have lower vitamin C content than fresh fruits and vegetables.

   Because cooking leaches water soluble vitamin C from the vegetable ~~by air~~
   high temprature accelerates its degredation by air oxidation

2. What are the oxidizing agents in this experiment?

   IO_3^- / I_3^- ⇒ KIO_3

3. What will be the color change of the starch indicator that indicates the end of the titration in this experiment?

   dark blue → colourless

4. Vitamin C is an acid (ascorbic acid) and a reducing agent. Which property is utilized for this analysis in this experiment?

   Reducing agent

5. If 29.57 ml of a vegetable juice contains 35% of the recommended daily allowance of vitamin C (equal to 60 mg). How many milliliters of the vegetable juice will provide 100% of the recommended daily allowance?

   29.57 ml → 35%

   ?? ml → 100%

   $\frac{29.57 \text{ ml} \times 100\%}{35\%}$ = 84.5 ml

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7 Vitamin C Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

	Trial (I)	Trial (II)
Mass of sample	0.12 ± 0.01 g	0.12 ± 0.0[illegible] g
Volume of 0.01 M KIO_3 added	25.0 ± 0.01 mL	25.0 ± 0.01 mL
Moles of IO_3^- added	~~0.25~~ 0.25 x 10^-3 mol	0.25 x 10^-3 mol
Moles of I_3^- generated, total	~~0.75~~ 0.75 x 10^-3 mol	0.75 x 10^- mol
Initial buret reading	1.0 ± 0.1 mL	5.2 mL
Final buret reading	4.8 ± 0.1 mL	9.2 mL
Volume of 0.04 M $Na_2S_2O_3$ added	3.8 ± 0.1 mL	4.0 ± 0.1 mL
Molar concentration of $Na_2S_2O_3$ (V x M 0.0[illegible])	0.04 mol/L	0.04 mol/L
Moles of $S_2O_3^{2-}$ added (V x M)	~~0.152~~ 1.52 x 10^-4 mol	1.6 x 10^-4 mol
Moles of I_3^- reduced by $S_2O_3^{2-}$ (excess)	~~7.6~~ 7.6 x 10^-5 mol	8 x 10^-5 mol
Moles of I_3^- reduced by $C_6H_8O_6$	~~0.706~~ 6.74 x 10^-4 mol	6.7 x 10^-4 mol
Moles of $C_6H_8O_6$ in sample	~~0.706~~ 6.74 x 10^-4 mol	6.7 x 10^-4 mol
Mass of $C_6H_8O_6$ in sample (176.08)	~~124.3~~ 0.118 g	0.1179 g
Mass percent of $C_6H_8O_6$ in sample	98.8 %	98.3 %
Average percent of $C_6H_8O_6$ in sample	98.55 %	

QUESTIONS:

1. A 25.0 mL volume of 0.010 M KIO_3, containing an excess of KI, is added to a 0.246 g sample of a lemon solution containing vitamin C. The red brown solution, caused by the presence of excess I_3^-, is titrated to a colorless starch end point with 10.7 mL of 0.100 M $Na_2S_2O_3$.

a. How many moles of KIO_3 were added to the lemon solution?

moles of $KIO_3 = V \times M$

$= \frac{25.0}{1000} L \times 0.010 \frac{mol}{L} = 2.50 \times 10^{-4}$ mol

b. Calculate the moles of I_3^- that are generated from KIO_3.

$IO_3^- + 8I^- + 6H^+ \longrightarrow 3I_3^- + 3H_2O$

2.5×10^{-4} ??

2.5×10^{-4} mol $IO_3^- \times \frac{3 \text{ mol } I_3^-}{1 \text{ mol } IO_3^-} = 7.50 \times 10^{-4}$ mol I_3^-

c. How many moles of I_3^- reacted with the 0.100 M $Na_2S_2O_3$ in the titration?

$2S_2O_3^{2-}$ + excess $I_3^- \longrightarrow 3I^- + S_4O_6^{2-}$

mols of $Na_2S_2O_3 = M \times V$

$= 0.100 \frac{mol}{L} \times \frac{10.7}{1000} L = 1.07 \times 10^{-3}$ mol $Na_2S_2O_3$

1.07×10^{-3} mol $S_2O_3^{2-} \times \frac{1 \text{ mol } I_3^-}{2 \text{ mol } S_2O_3^{2-}} = 5.35 \times 10^{-4}$ mol I_3^-

0.535 ×

d. How many moles of I_3^- had reacted with the vitamin C in the lemon sample?

# of mols of I_3^- (reacted) $= 7.5 \times 10^{-4} - 5.35 \times 10^{-4}$

$= 2.15 \times 10^{-4}$ mol I_3^- reacted with the vitamin C

e. Calculate the moles and grams of vitamin C in the sample.

$C_6H_8O_6 + I_3^- + H_2O_{(l)} \longrightarrow C_6H_8O_7 + 3I^- + 2H^+$

2.15×10^{-4} mol $I_3^- \times \frac{1 \text{ mol } C_6H_8O_6}{1 \text{ mol } I_3^-} = 2.15 \times 10^{-4}$ mol $C_6H_8O_6$

$^{-4}$ ... 176 g $C_6H_8O_6$ ... = 0.0378 g $C_6H_8O_6$

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f. Calculate the percent (by mass) of vitamin C in the lemon sample.

% vitamin C = $\frac{0.0378}{0.246}$ x 100% = 15.4%

2. If the blue color does not appear when the starch solution is added during the titration, should you continue titrating or discard the sample? Explain.

discar the sample because If the blue color doesn't appear this mean that there is no I_3^- in the solution which we need to determine the amount of vitamin C in the sample

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8 Bleach Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Write balanced equation(s) for the reactions involved in standardization of sodium thiosulfate.

$Cr_2O_7^{2-} + 6I^- + 14H^+ \rightarrow 3I_2 + 2Cr^{3+} + 7H_2O$

$ClO^- + 2S_2O_3^{2-} + H_2O \rightarrow Cl^- + S_4O_6^{2-} + 2OH^-$

$I_3^- + 2S_2O_3^{2-} \rightarrow 3I^- + S_4O_6^{2-}$

2. In today's chemical analysis of bleach solution, what substance is oxidized by the hypochlorite ion?

the stains (It removes the "excitable" electrons from the substance) by the energy of the visible light

3. What are the formulas of sodium thiosulfate and hypochlorite ion?

$Na_2S_2O_3$ ClO^-

4. Does sodium thiosulfate serve as an oxidizing agent or as a reducing agent? What does sodium thiosulfate oxidize or (reduce)?

As a reducing agent

It reduce I_2 To I^-

It reduces ClO^- To Cl^-

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8 Bleach Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

A. Standardization of thiosulfate solution:

	Trial(I)	Trial(II)
Volume of 0.010 M $K_2Cr_2O_7$ solution	25 ± 0.01 mL	25 ± 0.01 mL
Initial buret reading	10.9 ± 0.01 mL	25.9 ± 0.01 mL
Final buret reading	26.0 ± 0.01 mL	40.8 mL
Volume of $Na_2S_2O_3$ solution	15.1 ± 0.01 mL	14.9 mL
Number of moles of $K_2Cr_2O_7$	2.5×10^{-4} mol	2.5×10^{-4} mol
Number of moles of $Na_2S_2O_3$	1.5×10^{-3} mol	1.5×10^{-3} mol
Molarity of $Na_2S_2O_3$	0.0993 M	0.101 M
Average molarity of $Na_2S_2O_3$	0.100 M	

Seen

B. Analysis of bleach solution:

	Trial(I)	Trial(II)
Volume of bleach solution	25 ± 0.01 mL	25 ± 0.01 mL
Initial buret reading	17.9 ± 0.01 mL	26.8 ± 0.01 mL
Final buret reading	26.6 ± 0.01 mL	35.7 ± 0.01 mL
Volume of $Na_2S_2O_3$ solution	8.7 ± 0.01 mL	8.9 ± 0.01 mL
Number of moles of $Na_2S_2O_3$ oxidized M×V	8.70×10^{-4} mol	8.90×10^{-4} mol
Number of moles of ClO^- reduced ÷2	4.35×10^{-4} mol	4.45×10^{-4} mol
Molarity of diluted bleach solution	0.0174 M	0.0178 M
Molarity of original bleach solution	0.348 M	0.356 M
Average molarity of original bleach solution	0.352 M	

QUESTIONS:

$\frac{100}{10} = 10$ مرات

1. A 10.0 mL bleach sample is diluted to 100 mL in a volumetric flask. A 25 mL of this solution is analyzed according to the procedure in this experiment. If 11.3 mL of 0.30 M $Na_2S_2O_3$ is needed to reach the stoichiometric point, calculate the mass percent of NaClO in the original sample? (Assume the density of bleach solution is 1.084 g/mL).

\# mols of $Na_2S_2O_3$ = M x V

$= 0.30 \times \frac{11.3}{1000} = 3.39 \times 10^{-3}$ mol

$ClO^- + 2S_2O_3^{-2} + H_2O \rightarrow Cl^- + S_4O_6^{2-} + 2OH^-$

\# mols of $ClO^- = \frac{3.39 \times 10^{-3}}{2} = 1.695 \times 10^{-3}$ mol

Density = $\frac{m}{V}$ | ~~# mols of $ClO^- = 1.695 \times 10^{-3} \times 10 = 1.6$~~

~~$1.084 = \frac{m}{25}$~~ | $M = \frac{\#\text{mols}}{V} = \frac{1.695 \times 10^{-3}}{25 \times 10^{-3}} = 0.0678$

~~mass of bleach = 27.1 g~~

~~1 mol NaClO x $\frac{74.5 \text{ g NaClO}}{1 \text{ mol NaClO}}$ = 74.5 g NaClO~~

M → (10 ml) = 0.0678 x 10 = 0.678 M

density = $\frac{m}{V}$

$1.084 = \frac{m}{10.0 \text{ ml}} \Rightarrow m = 10.84$ g = 10.84 g

\# mols of (10.0 ml) = $M_{10ml} \times V$

$= 0.0678 \times \frac{10}{1000} = 6.78 \times 10^{-3}$ mol x $\frac{74.5 \text{ g NaClO}}{1 \text{ mol NaClO}}$

= 0.0505 g NaClO

% NaClO = $\frac{0.0505}{10.84}$ x 100% = 0.466%

9 Molar Mass of a Volatile Liquid

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Dumas method assumes the gas behaves ideally, in general, when gases behave ideally?

   when we have low Pressure and high temprcature

2. The vapor from an unknown volatile liquid occupies a 269 mL Erlenmeyer flask at 98.7 °C and 748 torr. The mass of vapor is 0.791 g.

   $\frac{748 \text{ mmHg}}{760 \text{ mmHg}} = 0.984$ atm

   a) How many moles of vapor are present?

   $PV = nRT$ $T = 372$ K $P = 748 \text{ torr} \times \frac{101.325 \text{ kPa}}{\text{torr}}$ $\frac{748 \text{ mmHg}}{\text{atm}}$ $= 0.984$ atm

   $n = \frac{PV}{RT} = \frac{0.984 \text{ atm} \times 269 \times 10^{-3} \text{ L}}{0.0821 \frac{\text{L·atm}}{\text{K·mol}} \times 372 \text{ K}} = 8.66 \times 10^{-3}$ mol

   b) What is the molar mass of the volatile liquid?

   $m = M \times n$

   $M = \frac{m}{n} = \frac{0.791 \text{ g}}{8.66 \times 10^{-3}} = \text{~~0.120 g/mol~~}$ 91.3 g/mol

   c) What is the density of vapor at STP?

   $P = 1$ atm $T = 0°C = 273$ K

   $PV = nRT \Rightarrow \frac{V}{n} = \frac{RT}{P} = \frac{0.0821 \frac{\text{atm·L}}{\text{K·mol}} \times 273 \text{ K}}{1 \text{ atm}} = 22.4 \frac{\text{L}}{\text{mol}}$

   molar mass = 0.120 g/mol

   molar volume = 22.4 L/mol

   density = $\frac{\text{mass}}{\text{vol}} = 91.3$ g/ml $= 4.08 \times 10^{-3}$ g/ml

9 Molar Mass of a Volatile Liquid

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

	Trial (I)	Trial (II)
Boiling point of water	90 ± 0.5 °C	°C
Atmospheric pressure	760 mmHg	mmHg
Mass of empty flask	77.9 ± 0.01 g	g
Mass of flask and condensed vapor	79.02 g	g
Mass of condensed vapor	1.18 g	g
Volume of flask	170 ± 0.01 mL	mL
Boiling point of water	363 ± 0.5 K	K
Atmospheric pressure	1.00 atm	atm
Volume of flask	0.17 L	L
Gas constant (R)	0.082 ~~8.314 L.atm /mol.K~~	
Molar mass of unknown	206.7 g/mol	g/mol
Average molar mass		g/mol

$PV = nRT$

$1_{atm} \times 0.17 L = n \times 0.082 \frac{atm.L}{K.mol} \times 363 K$

$n = 5.71 \times 10^{-3} mol$

Seen

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QUESTIONS:

1. If the outside of the flask is not dried after vaporizing the liquid, will the calculated molar mass be too high or too low? Explain.

If the outside of the flask is not dried after vaporizing the mass of flask and condensed vapour ↑

n is the same $n = \frac{mass}{molar\ mass}$

So that the molar mass will increase

2. Consider the following data: Mass of condensed vapor = 0.495 g at 98 ºC and 745 mm Hg occupies 127 mL. What is the molar mass of the liquid?

$T = 98 + 273 = 371\ K$

$P = 745\ mmHg \times \frac{1\ atm}{760\ mmHg} = 0.98\ atm$

V =

$PV = nRT$

$0.98\ atm \times \frac{127}{1000}\ L = n \times 0.082\ \frac{L.atm}{mol.K} \times 371\ K$

$n = 4.09 \times 10^{-3}\ mol$

$n = \frac{mass}{molar\ mass}$

$4.09 \times 10^{-3} = \frac{0.495\ g}{molar\ mass}$

$molar\ mass = 121.02 = 121\ g/mol$

seen

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10 Thermochemistry and Hess's Law

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What is the definition of the *joule* in terms of the basic SI units?

$w = F \cdot S$

$= N \cdot m = Joul$

2. In a calorimeter calibration experiment, a sampe of 51.203 g of water at 55.2 °C is added to a calorimeter containing 49.783 g of water at 23.5 °C. After stirring and waiting for the system to equilibrate, the final temperature reached is 37.6 °C. Calculate the calorimeter constant.

$q_{warm} = q_{cold} + q_{calorimeter}$

$m \times s \times \Delta t_{warm} = m \cdot s \cdot \Delta t_{cold} + C \cdot \Delta t_{cold}$

$51.203 \times 4.184 \times (55.2 - 37.6) = 49.783 \times 4.184 \times (37.6 - 23.5) + C \times (37.6 - 23.5)$

$3770.5 = 2936.9 + C \times 14.1 \rightarrow C = 59.1\ J/C°$

3. Give chemical equations for the reactions that will occur during this Experiment

$NaOH + HCl \longrightarrow NaCl + H_2O$

4. The acids and bases to be used in this experiment are classified as strong acids or bases. Use your textbook to find what is meant by the word strong in this context.

completely Ionized

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10 Thermochemistry and Hess's Law

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

63 °C

A. Determination of a Calorimeter Constant.

	Trial (I)	Trial (II)
Mass (or volume) of cold water	75.0±0.01 g	g
Temperature of cold water	20 ±0.2 °C	°C
Mass (or volume) of hot water	75.0±0.01 g	g
Temperature of hot water	63 °C	°C
Final temperature reached	39 °C	°C
Temperature change, $\Delta T_{(cold\ water)}$	19 °C	°C
Temperature change, $\Delta T_{(hot\ water)}$	24 °C	°C
Calorimeter constant (C)	82.6 J/°C	J/°C
Average value of calorimeter constant	82.6	J/°C

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B. Heat of Acid/Base Reactions

	Trial (I)	Trial (II) (If time permits)
Volume of 2 M NaOH used	50±0.01 mL	mL
Initial temperature of NaOH	22 °C	°C
Volume of 2 M HCl used	50±0.01 mL	mL
Initial temperature of HCl	22 °C	°C
Final temperature reached	27 °C	°C
Total mass of (volume) of mixture	100 g	g
Temperature change, ΔT	5 °C	°C
Average value of the Calorimeter constant (C)	82.6 J/°C	
Heat flow (q)*	413 J	J
Moles of NaOH	0.05 mol	mol
Moles of HCl reacted	0.05 mol	mol
Moles of water produced	0.05 mol	mol
ΔH*	-48.96 kJ/mol H_2O	kJ/mol H_2O
Average value of ΔH	-48.96 kJ/mol H_2O	
Literature value of ΔH	kJ/mol H_2O	

3

* Assume density of solution = 1.0 g/mL and specific heat of solution = 4.07 J/g. °C

$q = m\, s\, \Delta T$

100 4.07 $T_f - T_i$; T_{high} ; $\frac{T_{base} + T_{acid}}{2}$

$\Delta H_{rxn} = \Delta H_{solution} + \Delta H_{calorimeter}$

$m s \Delta t + C \Delta t$

$= 100 \times 4.07 \times 5 + 82.6 \times 5$

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Questions:

1. What effect on the calculated calorimeter constant would be observed if the calorimeter cup were made of conducting material (such as metal) rather than plastic foam?

It will increase $q_{warm} = q_{cold} + q_{calorimeter}$

conducting → $q_{warm} = q_{cold} + q_{calorimeter} + q$

→ $q_{calorimeter}$ will be higher than the ...

$q_{calorimeter} = C \times \Delta T$ → $C_{calorimeter}$ ↑

2. Why is water typically used as the heat-absorbing liquid in calorimeter?

① Because its high specific heat which

② constant for the temperature of reaction

....

3. Given that ΔH for the reaction: NH_3 + HCl → NH_4Cl, is -12.4 kJ/mol and the measured value of ΔH for the NaOH/HCl reaction is -55.8 kJ, use Hess"s law to calculate ΔH for the reaction:

$$NaOH + NH_4Cl \rightarrow NaCl + NH_3 + H_2O$$

$NH_3 + HCl \rightarrow NH_4Cl$ $\Delta H = -12.4$ kJ/mol

$NaOH + HCl \rightarrow NaCl + H_2O$ $\Delta H = -55.8$ kJ

~~$NH_4Cl \rightarrow NH_3 + HCl$~~ ~~$\Delta H = -12.4$ kJ/mol~~

~~$NH_3 + HCl \rightarrow NH_4Cl$~~

~~$NaOH + HCl \rightarrow NaCl + H_2O$~~ ~~$\Delta H = 55.8$ kJ/mol~~

$NH_4Cl \rightarrow NH_3 + HCl$ $\Delta H = 12.4$ kJ/mol

(+) $NaOH + HCl \rightarrow NaCl + H_2O$ $\Delta H = -55.8$ kJ/mol

$NH_4Cl + NaOH \rightarrow NH_3 + NaCl + H_2O$ $\Delta H = -43.4$ kJ

11 Molar Mass from Freezing Point Depression

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

3. Define what we mean by the colligative property

a solution characteristic depends only on the number of solute molecules in a given quantity of solvent and don't depend on the nature or type of the solute

4. A 0.2436-g sample of an unknown substance was dissolved in 20.0 mL of cyclohexane. The density of cyclohexane is 0.779 g/mL. The freezing-point depression was 2.5 °C. Calculate the molar mass of the unknown substance. [K_f (cyclohexane) = 20.5 °C/m]

mass of cyclohexane $= 20.0\ ml \times 0.779\ \frac{g}{ml} \times \frac{10^{-3}\ Kg}{1g} = 0.01558\ Kg$

$\Delta T_f = 2.5\ °C = K_f \times m$

$2.5\ C° = 20.5\ \frac{C°}{m} \times \frac{0.2436\ g / \text{molar mass}}{0.01558\ Kg}$

molar mass $= \frac{20.5\ \frac{C°}{m} \times 0.2436\ g}{2.5\ C° \times 0.01558\ Kg} = 128.2\ g/mol$

seen

11 Molar Mass from Freezing Point Depression

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

(A) Freezing point of cyclohexane

Density of cyclohexane = 0.779 g/mL

	Trial (I)	Trial (II)
Volume of cyclohexane	10.0 ± 0.01 mL	mL
Mass of cyclohexane	7.79[illegible] g	g

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Trial (I)		Trial (II)	
Time (s)	Temp.(°C)	Time (s)	Temp.(°C)
0.5	15°C		
1	14°C		
1.5	13°C		
2	11.5°C		
2.5	10.3°		
3	9.5°		
3.5	9		
4	8		
4.5	7.5°C		
5	7.5		
5.5	7.5		
6	7.5		
6.5	7.5		
7	7.5		
7.5	7.5		
8	7.5		
8.5	7.5		
9	7.5		7.0
Freezing point of cyclohexane	7.5 ± 0.2 °C [Trial (I)]		7.0 ± 0.2 °C [Trial (II)]
Average value	7.25 ± 0.2		°C

B-Freezing point of solutions

	Trial (I)	Trial (II)
Mass of paper + solute	mL	mL
Mass of paper	g	
Mass of solute (m_1)	0.2 ± 0.01 g	

Trial (I)		Trial (II)	
Time (s)	Temp.(ºC)	Time (s)	Temp.(ºC)

7.25

Freezing point of solution	3.75 ±0.2 ºC[Trial (I)]	3.5 ±0.2 ºC [Trial (II)]
ΔT_f	3.5 ± 0.2 ºC	3.75 ±0.2 ºC
Molar Mass	150.4 g/mol	140.38 g/mol
Average value	145.4 g/mol	
Total mass of solute ($m_1 + m_2$)	g	

145

good

Excellent

3.25 = 20.5 × m

3.25 = 20.5 × 0.2 / molar mass

7.79 × 10⁻³ Kg molar mass

molar mass = 0.2 × 20.5

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QUESTIONS:

1. What would be the effect of each of the following on the calculated molar mass of the solute?

   a) Some cyclohexane evaporated while the freezing point of pure cyclohexane was being measured.

   no effect on the freezing point

   molar mass of the solute increases $\Delta T_f = K_f \times \frac{\text{mols of solute}}{\text{kg of solvent}}$

   b) Some cyclohexane evaporated after the solute was added.

   ~~no effect~~ decreases

   c) A foreign solute was already present in the cyclohexane.

   no change

   d) The thermometer is not calibrated correctly. It gives a temperature that is 1.5 °C too low at all temperatures.

   no effect

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12 Solubility Rules: Solubilities within a Family

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Give the names and symbols of the alkaline earth metals that you will encounter in this experiment.

Magnisium Mg – Calcium Ca – Sontrinshium Sr
Barium Ba

2. What is the general electron configuration of the alkaline earth Metals and lead?

alkaline earth metals : ~ ns^2
lead : $[Xe]_{54}\ 6s^2 4f^{14} 5d^{10} 6p^2$

$_{12}Mg : [Ne]_{10}\ 3s^2$
$_{20}Ca : [Ar]_{18}\ 4s^2$
$_{38}Sr : [Kr]_{36}\ 5s^2$
$_{56}Ba : [Xe]_{54}\ 6s^2$

3. Compare the formulas of the oxides formed by the alkaline earth metals with those of lead.

alkaline earth metals :- MO
lead :- PbO_2

4. Give general formulas for the following compounds with alkaline earth metals or lead:

(a) A sulfate MSO_4

(b) A carbonate MCO_3

(c) A chromate $MCrO_4$

(d) An oxalate MC_2O_4

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12 Solubility Rules: Solubilities within a Family

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

Record the color and amount (large, small, traces, none) of precipitates.

	$Mg(NO_3)_2$	$Ca(NO_3)_2$	$Sr(NO_3)_2$	$Ba(NO_3)_2$	$Pb(NO_3)_2$
NaOH					
NaCl					
NaBr					
NaI					
Na_2SO_4					
Na_2CO_3					
$Na_2C_2O_4$					
Unknown					

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Questions:

1. What trends in the solubility's of the compounds of the alkaline earth metals can be concluded?

As the atomic number of the Group IIA metal increase the solubility's may decrease

2. (a) Compare the solubility's of the lead compounds with those of the alkaline earth metals. How are the solubility's similar, and how do they differ?

   (b) As you have seen, lead can form the same kinds of compounds as the alkaline earth metals. Nevertheless, the solubility's may differ markedly. Give a convincing reason.

13 Solubility Product Constant and Common Ion Effect

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Write the mass action expression for these slightly soluble salt equilibria:

a) $CuS(s) \leftrightarrows Cu^{2+}(aq) + S^{2-}(aq)$, K_{sp} = $[Cu^{+2}][S^{2-}]$

b) $Al_2S_3(s) \leftrightarrows 2Al^{3+}(aq) + 3S^{2-}(aq)$, K_{sp} = $[Al^{3+}]^2[S^{2-}]^3$

c) $BaSO_4(s) \leftrightarrows Ba^{2+}(aq) + SO_4^{2-}(aq)$, K_{sp} = $[Ba^{2+}][SO_4^{2-}]$

d) $Sr_3(PO_4)_2(s) \leftrightarrows 3Sr^{2+}(aq) + 2PO_4^{3-}(aq)$, K_{sp} = $[Sr^{2+}]^3[PO_4^{3-}]^2$

2. Calculate the molar solubility and the solubility (in g/L) of $PbCl_2$, (K_{sp} for $PbCl_2 = 1.6x10^{-5}$).

$PbCl_2 \rightleftharpoons Pb^{+2} + 2Cl^-$ $[Pb^{+2}] = \frac{1}{2}[Cl]$

$K_{sp} = [Pb^{+2}][Cl^-]^2 = [Pb^{+2}](2[Pb^{+2}])^2$ ~~3.28 g/L~~

$[Pb^{+2}][Cl^-]^2 = 1.6x10^{-5}$

- molar solubility $= \sqrt[3]{1.6x10^{-3}} = 0.117$ g/L

- solubility $= 0.117 \frac{mol}{L} \times \frac{152.9}{1 mol PbCl_2} = 17.89$ g/L

3. Calculate the molar solubility of $PbCl_2$ in the presence of 0.10 M NaCl, (K_{sp} for $PbCl_2 = 1.6x10^{-5}$).

$NaCl \rightleftharpoons Na^+ + Cl^-$ (0.1M, 0.1M) 0.10 mol/L

$K_{sp} = [Pb^{+2}][Cl]^2$ $PbCl_2 \rightarrow Pb^{+2} + 2Cl^-$ (0.1)

$1.6x10^{-5} = [Pb^{+2}][0.1]^2$ ~~in the presence of 0.10 M~~

~~X = 0.117 mol/L → molar solubility = 0.117 + 0.10 = 0.217 mol/L~~

$[Pb^{+2}] = 1.6x10^{-3}$ mol/L

~~X = 0.016 M~~

molar solubility $= [Pb^{+2}] = 1.6x10^{-3}$ mol/L

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13 Solubility Product Constant and Common Ion Effect

Results Calculations

A- Molar solubility and K_{sp} for $Ca(OH)_2$

10.5

Concentration of standard HCl solution used = 0.05 mol/L

	Trial 1	Trial 2
Initial buret reading	11.00±0.001 mL	mL
Final buret reading	37.50±0.01 mL	mL
Volume of standard HCl added	26.5 ±0.01 mL	mL
Moles of HCl added	1.33×10^{-3} mol	mol
Moles of OH^- in saturated $Ca(OH)_2$ solution	1.33×10^{-3} mol	mol
Volume of saturated $Ca(OH)_2$ solution titrated	25 ±0.01 mL	mL
$[OH^-]$ at equilibrium (moles / volume)	0.0532 mol/L	mol/L
$[Ca^{2+}]$ at equilibrium (÷2)	0.0266 mol/L	mol/L
Molar solubility of $Ca(OH)_2$	0.0266 mol/L	mol/L
Average molar solubility of $Ca(OH)_2$	0.0266 mol/L	
Average solubility of $Ca(OH)_2$ ($S \times M$)	1.97 g/L	
K_{sp} of $Ca(OH)_2$	7.44×10^{-5}	

$Ca(OH)_2 \rightleftharpoons Ca^{+2} + 2OH^-$

seen

CIVILIUM

B- Solubility of $Ca(OH)_2$ in the Presence of Ca^{2+}

Concentration of standard HCl solution used = 0.05 mol/L

	Trial 1	Trial 2
Initial buret reading 52.1 ±0.01	~~37.5~~ ±0.01 mL	mL
Final buret reading 65.7 ±0.01	65.7 ±0.01 mL	mL
Volume of standard HCl added 13.6 ±0.01	28.2 ±0.01 mL	mL
Moles of HCl added 6.8×10^{-4}	1.41×10^{-3} mol	mol
Moles of OH^- in saturated $Ca(OH)_2/Ca^{2+}$ solution 6.85×10^{-4}	1.41×10^{-3} mol	mol
Volume of saturated $Ca(OH)_2/Ca^{2+}$ solution titrated	25 ±0.01 mL	mL
$[OH^-]$ at equilibrium 0.0274	~~0.0564~~ 0.0564 mol/L	mol/L
$[Ca^{2+}]$ at equilibrium	0.1 mol/L	mol/L
Molar solubility of $Ca(OH)_2$ in $Ca(OH)_2/Ca^{2+}$ solution	0.0137 mol/L	mol/L
Average solubility of $Ca(OH)_2$ in $Ca(OH)_2/Ca^{2+}$ solution 1.015	2.08	g/L
Average molar solubility of $Ca(OH)_2$ in $Ca(OH)_2/Ca^{2+}$ solution 0.0137	~~154.9~~	mol/L

Ca^{+2} excess $\frac{1}{2}[OH]$ excess

seen

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QUESTIONS:

1. A saturated solution of $Ca(OH)_2$ is prepared by adding enough $Ca(OH)_{2(s)}$ to distilled pre-boiled water, the solution is then filtered to remove any insoluble amount of $Ca(OH)_2$.

   a) Why pre-boiled water is used to prepare the saturated solution of $Ca(OH)_2$?

   to get rid of gases

   b) If some solid $Ca(OH)_2$ remained after filtration, how would it affects the calculated values of molar solubility of $Ca(OH)_2$ and the K_{sp} for $Ca(OH)_2$?

   ↑ increases $K_{sp} = [Ca^{+2}][OH^-]^2$

   $HCl + OH^- \rightarrow Cl^- + H_2O$ } titration

   mols of HCl ↑ → mols of OH⁻ ↑ → [OH⁻] ↑ → $[Ca^{+2}] = \frac{1}{2}[OH^-]$ ↑ → ↑ molar solubility → $K_{sp} = [Ca^{+2}][OH^-]^2$ ↑

2. How did the added $CaCl_2$ affect the molar solubility of $Ca(OH)_2$? Explain.

   $Ca(OH)_2 \rightarrow Ca^{+2}$ ↑ $+ 2OH^-$ ↓ $CaCl_2 \rightarrow Ca^{+2} + 2Cl^-$

   decreases $K_{sp} = [Ca^{+2}][OH^-]^2$ $[OH^-] = \sqrt{\frac{K_{sp}}{[Ca^{+2}]}}$

   ↓ molar solubility $= \frac{1}{2}[OH^-]$ ↓

3. Calculate the molar solubility of $Ca(OH)_2$ in 0.10 M NaOH solution. Use the value of K_{sp} obtained in part A. 5.529×10^{-3}

   $Ca(OH)_{2(s)} \rightarrow Ca^{+2} + 2OH^-$ $NaOH \rightarrow Na^+ + OH^-$ (0.1 M, 0.1 M, 0.1 M)

   $[Ca^{+2}]$ 0.1

   $K_{sp} = [Ca^{+2}][OH^-]^2$

   $7.44 \times 10^{-5} = [Ca^{+2}] \times (0.1)^2$

   $[Ca^{+2}] = 7.44 \times 10^{-3}$ mol/L

   molar solubility $= 7.44 \times 10^{-3}$ mol/L

   seen

   0.0266