

TOPIC 1 PROBLEM SET – QUANTITATIVE CHEMISTRY

Many of these questions involve concepts you should be familiar with from MYP chemistry. They progress from simple to more challenging. After completing 1 – 58, a selection of prior exam questions follows. It is important that you develop a feel for what prior exam questions are like, as being familiar with exam questions will be quite helpful when you take your IB chemistry exam.

1. Practice your fluency in inorganic nomenclature with these charts.

formula	# and type of atoms and A_r of each atom	Formula mass (M_r)
$Ba_3(PO_4)_2$	3Ba 2P 8O	601.93
$Fe(NO_3)_3$	1Fe 3N 9O	241.88
$(NH_4)_2SO_4$	2N 8H 1S, 4O	132.10

Ionic Compounds:

cations	anions	formula	name	Formula mass (M_r)
Al^{3+}	S^{2-}	Al_2S_3	Aluminum Sulfide	149.96
K^+	O^{2-}	K_2O	Potassium oxide	94.20
Ca^{2+}	NO_3^-	$Ca(NO_3)_2$	calcium nitrate	164.10
Zn^{2+}	CO_3^{2-}	$ZnCO_3$	Zinc (II) carbonate	125.40
Cu^{2+}	F^-	CuF_2	Copper (II) fluoride	101.55

Molecular (Covalent) Compounds:

Name	Formula	Formula mass (M_r)
Sulfur trioxide	SO_3	80.00
diphosphorus pentoxide	P_2O_5	141.94
Carbon tetrachloride	CCl_4	118.36
Trinitrogen hexasulfide	N_3S_6	234.03
nitrogen dioxide	NO_2	46.01

2. Give the name for the following ions:

- a) NH_4^+ ammonium ion e) OH^- hydroxide ion i) PO_4^{3-} phosphate ion
 b) SO_4^{2-} sulfate ion f) $Cr_2O_7^{2-}$ dichromate ion j) CH_3COO^- acetate ion
 c) CrO_4^{2-} chromate ion g) CO_3^{2-} carbonate ion

3. d) NO_3^- nitrate ion h) HCO_3^- bicarbonate ion (hydrogen carbonate)

3. How many moles are there in 4.2 g aluminum chloride? $AlCl_3 = 133.33$

$$\frac{4.2g}{133.33g/mol} = 0.0315 \text{ mol } AlCl_3$$

4. How many moles are there in 42.5 g copper (II) fluoride? $CuF_2 = 101.54$

$$\frac{42.5g}{101.54g/mol} = 0.419 \text{ mol } CuF_2$$

5. What is the mass of 2.25 moles zinc (II) phosphate? $Zn_3(PO_4)_2 = 386.17$

$$2.25 \text{ mol} \times 386.17g/mol = 868.9g \text{ } Zn_3(PO_4)_2$$

6. How many moles are there in 27.5 g aluminum nitrate? $AlN = 40.99$

$$\frac{27.5g}{40.99g/mol} = 0.671 \text{ mol } AlN$$

7. How many molecules of sulfur dioxide are there in 1.25 g of sulfur dioxide? $SO_2 = 64.06$

$$\frac{1.25g}{64.06g/mol} \times 6.02 \times 10^{23} \text{ molecules/mol} = 1.17 \times 10^{22} \text{ molecules } SO_2$$

8. What is the mass of 3.29×10^{25} molecules of water? H_2O

$$\frac{3.29 \times 10^{25} \text{ molecules}}{6.02 \times 10^{23} \text{ molecules/mol}} \times 18.02g/mol = 984.8g \text{ } H_2O$$

9. How many moles of hydrogen are there in 64.9 g water?

$$\frac{64.9g}{18.02g/mol} \times 2 \text{ mol } H = 7.20 \text{ mol } H$$

10. What mass of zinc is in 14.4 g zinc (II) fluoride? $ZnF_2 = 103.4$

$$\frac{14.4g}{103.4g/mol} \times 1 \text{ mol } Zn = 0.139 \text{ mol } Zn$$

11. How many atoms of bromine are there in 0.155 g magnesium bromide? $MgBr_2 = 184.11$

$$\frac{0.155g}{184.11g/mol} \times 2 \text{ mol } Br \times 6.02 \times 10^{23} \text{ atoms/mol} = 1.01 \times 10^{21} \text{ atoms } Br$$

12. How many atoms are there in 12 g of K_2O ? $= 94.20$

$$\frac{12g}{94.20g/mol} \times 6.02 \times 10^{23} \text{ molecules/mol} \times 3 \text{ atoms in } K_2O = 2.3 \times 10^{23} \text{ atoms } K_2O$$

13. Ionic compounds form solid crystals that are held together by the attractive force between positive cations and negative anions. A hydrated ionic compound contains small amounts of water distributed throughout the crystal. A hydrated ionic compound has a unique formula, in which the number of moles of water is denoted after the formula of the compound. For example, in calcium chloride dihydrate, there are two moles of water to every one mole of calcium chloride, and the formula is written: $CaCl_2 \cdot 2H_2O$.

A hydrated copper (II) sulfate crystal is dehydrated by heating above a flame. Hydrated copper (II) sulfate has a characteristic light blue color, while dehydrated copper (II) sulfate is gray/white.

Consider the data

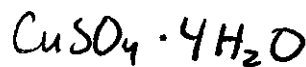
Object	Mass (± 0.1 g)
evaporating dish	47.4 g
evaporating dish and $CuSO_4 \cdot X H_2O$	52.5g 52.5g
mass of dish and $CuSO_4$ remaining after heating	50.9 g

Calculate the formula of the hydrate.

$$CuSO_4 = 50.9 - 47.4 = 3.5g \text{ } CuSO_4$$

$$H_2O = 52.5 - 50.9 = 1.6g \text{ } H_2O$$

$$CuSO_4 = \frac{3.5g}{159.61g/mol} = 0.0219 \text{ mol } CuSO_4$$

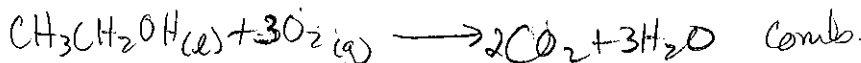


$$H_2O = \frac{1.6g}{18.02g/mol} = 0.08879 \text{ mol } H_2O = 4$$

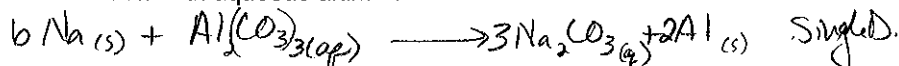
For each of the following (14 – 22):

- classify the reaction
- write chemical equation (include state symbols: s, l, g, aq)
- balance the equation

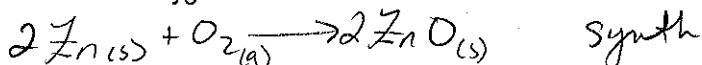
14. liquid ethyl alcohol (CH₃CH₂OH) burns in the presence of oxygen



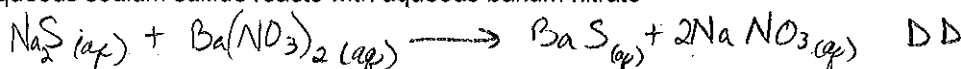
15. sodium reacts with aqueous aluminum carbonate



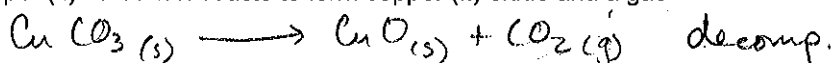
16. zinc reacts with oxygen



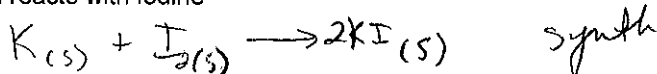
17. aqueous sodium sulfide reacts with aqueous barium nitrate



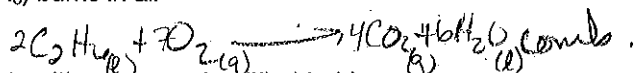
18. copper (II) carbonate reacts to form copper (II) oxide and a gas



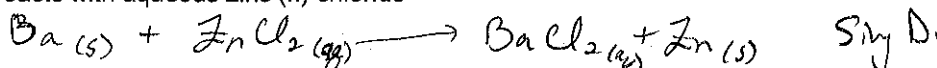
19. potassium reacts with iodine



20. ethane (C₂H₆) burns in air



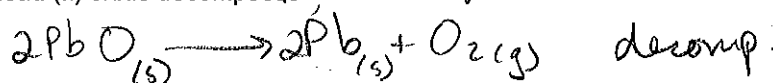
21. barium reacts with aqueous zinc (II) chloride



22. aqueous sodium carbonate reacts with aqueous zinc (II) fluoride



23. lead (II) oxide decomposes



For each of the following (24 – 27), determine the empirical formula from the percentage composition or mass data given. An empirical formula displays the **lowest whole number mole ratio of elements in a compound**. A molecular formula shows this same ratio, only it isn't always in the most simplified form (for example, glucose has a molecular formula of C₆H₁₂O₆, but when simplified, its empirical formula is CH₂O).

24. A ~~3.99~~^{4.00} gram sample of an unknown compound that contains 1.71 g carbon and 2.29 g oxygen

$$\text{C: } \frac{1.71\text{g}}{12.01\text{g}} = \frac{0.1424}{0.1424} = 1$$

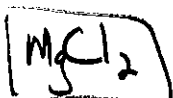
$$\text{O: } \frac{2.29\text{g}}{15.99\text{g}} = \frac{0.1432}{0.1424} = 0.99$$



25. A 7.95 gram sample of an unknown compound that contains 2.03 g magnesium and 5.92 g chlorine

$$\text{Mg: } \frac{2.03\text{g}}{24.31\text{g}} = \frac{0.0835}{0.0835}$$

$$\text{Cl: } \frac{5.92\text{g}}{35.45\text{g}} = \frac{0.167}{0.0835}$$



26. An unknown sample that contains 31.9 % potassium, 29.0 % chlorine, and 39.1 % oxygen, by mass.

$$\text{K: } \frac{31.9\text{g}}{39.10\text{g}} = \frac{0.8159}{0.816}$$

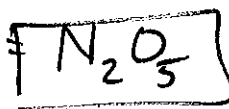
$$\text{Cl: } \frac{29.0\text{g}}{35.45\text{g}} = \frac{0.818}{0.816}$$

$$\text{O: } \frac{39.1\text{g}}{15.99\text{g}} = \frac{2.45}{0.816}$$



27. Deduce the empirical formula of a compound that contains 25.9 % nitrogen and 74.1% oxygen, by mass.

$$\text{N: } \frac{25.9\text{g}}{14.01\text{g}} = \frac{1.849}{1.849} = 1 \times 2$$



$$\text{O: } \frac{74.1\text{g}}{15.99\text{g}} = \frac{4.634}{1.849} = 2.5 \times 2$$

28. A compound is 80% carbon and 20% hydrogen, by mass.

a) What is the empirical formula?

$$C: \frac{80g}{12.01g} \times \frac{1 \text{ mol}}{12.01g} = 6.66 = 1 \quad H: \frac{20g}{1.01g} \times \frac{1 \text{ mol}}{1.01g} = 19.8 = 2.97 \quad CH_3$$

b) If the molar mass is about 30, what is the molecular formula?

$$CH_3 = 15.04g/mol \quad \frac{30}{15.04} \approx 2 \quad \boxed{C_2H_6}$$

29. Determine the empirical and molecular formulas of a compound with a molecular weight of 42 that contains 85.9% C and 14.1% H.

$$C: \frac{85.9g}{12.01g} \times \frac{1 \text{ mol}}{12.01g} = 7.152 = 1 \quad H: \frac{14.1g}{1.01g} \times \frac{1 \text{ mol}}{1.01g} = 13.96 = 1.95 \quad CH_2 = \frac{42}{14.03} = 3 \quad \boxed{C_3H_6}$$

30. A compound is analyzed and found to contain: 4.80g carbon, 0.80g hydrogen, 3.20g oxygen and 2.80g nitrogen.

What is the empirical formula? If the molar mass is about 174, what is the molecular formula?

$$C: \frac{4.80g}{12.01g} \times \frac{1 \text{ mol}}{12.01g} = 0.3997 \quad O: \frac{3.20g}{15.99g} \times \frac{1 \text{ mol}}{15.99g} = 0.2001 \quad C_2H_4NO = 58.06$$

$$H: \frac{0.80g}{1.01g} \times \frac{1 \text{ mol}}{1.01g} = 0.792 \quad N: \frac{2.80g}{14.01g} \times \frac{1 \text{ mol}}{14.01g} = 0.2000 \quad \frac{174}{58.06} = 3 \quad \boxed{C_6H_{12}O_3N_3}$$

31. A sample of a hydrocarbon burns completely in oxygen to form 13.2g carbon dioxide and 5.4g water. What is the empirical formula?

$$C: \frac{13.2g CO_2}{44.00g} \times \frac{1 \text{ mol}}{44.00g} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.30 \text{ mol C} = 1 \quad \boxed{CH_2}$$

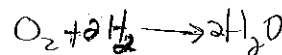
$$H: \frac{5.4g H_2O}{18.02g} \times \frac{1 \text{ mol}}{18.02g} \times \frac{2 \text{ mol H}}{2 \text{ mol } H_2O} = 0.599 \text{ mol H} = 2$$

For 32 - 37:

- Write correct formulas for products and reactants.
- Write the equation
- Balance the equation
- Perform the conversion, using the following T-chart template as a guide if you like:

mass given	mol given	unknown from equation	formula mass unknown
	Formula mass given	given from equation	mol unknown

32. How many grams of oxygen react with excess hydrogen to form 13.5 g water?



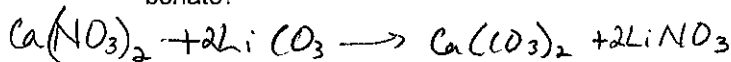
$$\frac{13.5g H_2O}{18.02g} \times \frac{1 \text{ mol } H_2O}{18.02g} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} \times \frac{32.00g O_2}{1 \text{ mol } O_2} = \boxed{11.99g O_2}$$

33. How many grams of silver are formed as 125 g zinc reacts with excess aqueous silver nitrate?



$$\frac{125g Zn}{65.41g} \times \frac{1 \text{ mol}}{65.41g} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Zn}} \times \frac{107.87g Ag}{1 \text{ mol Ag}} = \boxed{412.3g Ag}$$

34. What mass of aqueous lithium carbonate reacts with excess calcium nitrate to form 87 g of aqueous calcium carbonate?



$$\frac{87g CaCO_3}{100.09g} \times \frac{1 \text{ mol}}{100.09g} \times \frac{2 \text{ mol } Li_2CO_3}{1 \text{ mol } CaCO_3} \times \frac{66.92g Li_2CO_3}{1 \text{ mol } Li_2CO_3} = \boxed{72.8g Li_2CO_3}$$

35. How many grams of oxygen form as 50.0 g of Hg₂O decomposes?



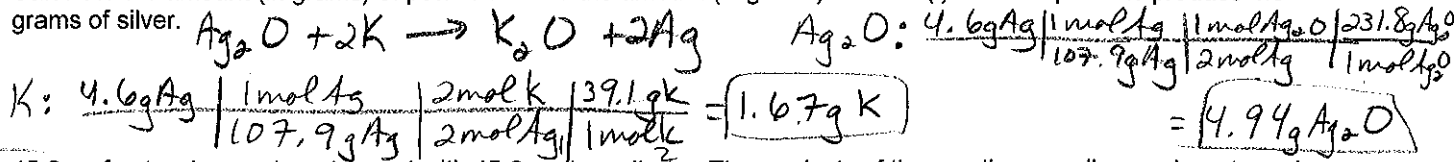
$$\frac{50.0g Hg_2O}{417.17g} \times \frac{1 \text{ mol}}{417.17g} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } Hg_2O} \times \frac{32.0g O_2}{1 \text{ mol } O_2} = \boxed{1.92g O_2}$$



36. How many grams of potassium oxide form as 15.0g potassium burns in the presence of oxygen?

$$\frac{15.0g K}{39.10g} \times \frac{1 mol K}{1 mol K} \times \frac{2 mol K_2O}{4 mol K} \times \frac{94.2g K_2O}{1 mol K_2O} = 18.07g K_2O$$

37. Calculate the amount (in grams) of potassium and the amount (in grams) of silver (I) oxide required to produce 4.6 grams of silver.



38. 45.0 g of potassium carbonate react with 45.0 g silver nitrate. The products of the reaction are silver carbonate and potassium nitrate.

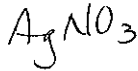
a) What is the theoretical yield of silver (I) carbonate (i.e. what is the maximum mass of silver (I) carbonate that can be produced)?

$$45.0g K_2CO_3 \times \frac{1 mol}{138.21g} = 0.32559 mol K_2CO_3$$

$$0.1325 mol Ag_2CO_3 \times \frac{275.75g}{1 mol} = 36.52g Ag_2CO_3$$

$$45.0g AgNO_3 \times \frac{1 mol}{169.87g} = 0.2649 mol AgNO_3 \times \frac{0.1325 mol Ag_2CO_3}{2 mol AgNO_3} = 0.1325 mol Ag_2CO_3$$

b) What is the limiting reactant?



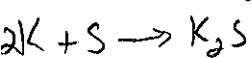
c) What is the reactant present in excess?



d) Determine the percentage yield if only 30.0 grams of silver (I) carbonate are produced experimentally.

$$\frac{30.0g Ag_2CO_3 \text{ actual}}{36.52g Ag_2CO_3 \text{ theory}} \times 100 = 82.2\%$$

39. 15.0 g of potassium reacts with 5.0 g sulfur.



a) What is the theoretical yield of potassium sulfide (i.e. what is the maximum mass of potassium sulfide that can be produced)?

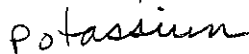
$$\frac{15.0g K}{39.1g K} \times \frac{1 mol K}{1 mol K} \times \frac{1 mol K_2S}{2 mol K} \times \frac{94.19g K_2S}{1 mol K_2S} = 18.11g K_2S$$

$$\frac{5.0g S}{32.0g S} \times \frac{1 mol S}{1 mol S} \times \frac{1 mol K_2S}{1 mol S} \times \frac{94.19g K_2S}{1 mol K_2S} = 14.72g K_2S$$

b) What is the limiting reactant?



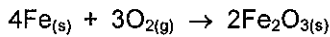
c) What is the reactant present in excess?



d) Determine the percentage yield if only 10.9 grams of potassium sulfide are produced experimentally.

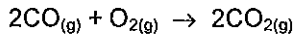
$$\frac{10.9g}{14.72g} \times 100 = 74.1\% \text{ yield}$$

40. What is the maximum mass of iron (III) oxide produced when 10.0 grams of iron are allowed to completely oxidize?



$$\frac{10.0\text{g Fe}}{55.85\text{g}} \times \frac{1\text{mol Fe}}{1\text{mol Fe}} \times \frac{2\text{mol Fe}_2\text{O}_3}{4\text{mol Fe}} \times \frac{159.7\text{g Fe}_2\text{O}_3}{1\text{mol Fe}_2\text{O}_3} = 14.3\text{g Fe}_2\text{O}_3$$

41. What volume of carbon dioxide is released when 10 dm³ of carbon monoxide are combined with 10 dm³ of oxygen?



$$\frac{10\text{ dm}^3\text{O}_2}{20.7\text{ dm}^3} \times \frac{1\text{ mol O}_2}{1\text{ mol O}_2} \times \frac{2\text{ mol CO}_2}{1\text{ mol O}_2} \times \frac{22.7\text{ dm}^3\text{CO}_2}{1\text{ mol CO}_2} = 20\text{ dm}^3\text{CO}_2$$

$$\frac{10\text{ dm}^3\text{CO}}{22.7\text{ dm}^3} \times \frac{1\text{ mol CO}}{1\text{ mol CO}} \times \frac{2\text{ mol CO}_2}{2\text{ mol CO}} \times \frac{22.7\text{ dm}^3\text{CO}_2}{1\text{ mol CO}_2} = 10\text{ dm}^3\text{CO}_2$$

42. 1.00 moles of NaCl are dissolved in 500 ml of solution. What is the concentration in mol dm⁻³?

$$\frac{1.00\text{ mol}}{0.500\text{ dm}^3} = 2\text{ mol dm}^{-3}\text{ NaCl}$$

43. 3.5 moles of KF are dissolved in 2.0 L of solution. What is the concentration in mol dm⁻³?

$$\frac{3.5\text{ mol}}{2.0\text{ dm}^3} = 1.75\text{ mol dm}^{-3}\text{ KF}$$

44. 133.0 g of NaCl are dissolved in 1.00 L of solution. What is the concentration in mol dm⁻³?

$$\frac{133.0\text{g}}{58.45\text{g NaCl}} \times \frac{1\text{ mol NaCl}}{1\text{ mol NaCl}} = 2.28\text{ mol NaCl} \rightarrow \frac{2.28\text{ mol NaCl}}{1.00\text{ dm}^3} = 2.28\text{ mol dm}^{-3}\text{ NaCl}$$

45. How many moles of sodium carbonate are dissolved in 500 ml of 0.055 mol dm⁻³ solution?

$$\frac{0.500\text{ dm}^3}{1\text{ dm}^3} \times 0.055\text{ mol dm}^{-3} = 0.0275\text{ mol Na}_2\text{CO}_3$$

46. What is the concentration in mol dm⁻³ of 38 g lead (II) iodide dissolved in 500 ml of solution?

$$\frac{38\text{g PbI}_2}{461.2\text{g PbI}_2} \times \frac{1\text{ mol}}{1\text{ mol}} = 0.0824\text{ mol PbI}_2 \rightarrow \frac{0.0824\text{ mol}}{0.500\text{ dm}^3} = 0.165\text{ mol dm}^{-3}\text{ PbI}_2$$

47. How many grams of NaCl are there in 500 ml of a 1.3 mol dm⁻³ solution NaCl?

$$\frac{0.500\text{ dm}^3}{1\text{ dm}^3} \times 1.3\text{ mol dm}^{-3} \times 58.45\text{g NaCl} = 38.0\text{g NaCl}$$

48. What mass of copper will react with 500 mL of 2 mol dm⁻³ magnesium chloride? $\text{Cu} + \text{MgCl}_2 \rightarrow \text{CuCl}_2 + \text{Mg}$

$$\frac{0.500\text{ dm}^3}{1\text{ dm}^3} \times 2\text{ mol MgCl}_2 = 1.0\text{ mol MgCl}_2 \times \frac{1\text{ mol Cu}}{1\text{ mol MgCl}_2} \times 63.55\text{g Cu} = 63.55\text{g Cu}$$

49. Methanol, CH₃OH, can be prepared according the equation: $\text{CO} + 2\text{H}_2 \rightarrow \text{CH}_3\text{OH}$

When 175 g of carbon monoxide is mixed with 36.5 g hydrogen gas and allowed to react, 112 g of product forms. Determine a) theoretical yield b) percent yield c) limiting reactant (CO is limiting)

$$\frac{175\text{g CO}}{28.01\text{g}} \times \frac{1\text{ mol CO}}{1\text{ mol CO}} \times \frac{1\text{ mol CH}_3\text{OH}}{1\text{ mol CO}} \times 32.05\text{g CH}_3\text{OH} = 200.2\text{g CH}_3\text{OH} = \text{theoretical yield}$$

$$\frac{36.5\text{g H}_2}{2.02\text{g H}_2} \times \frac{1\text{ mol H}_2}{2\text{ mol H}_2} \times \frac{1\text{ mol CH}_3\text{OH}}{1\text{ mol CH}_3\text{OH}} \times 32.05\text{g CH}_3\text{OH} = 289.1\text{g CH}_3\text{OH}$$

$$\% \text{ yield} = \frac{112\text{g}}{200.2\text{g}} \times 100 = 55.9\% \text{ yield}$$

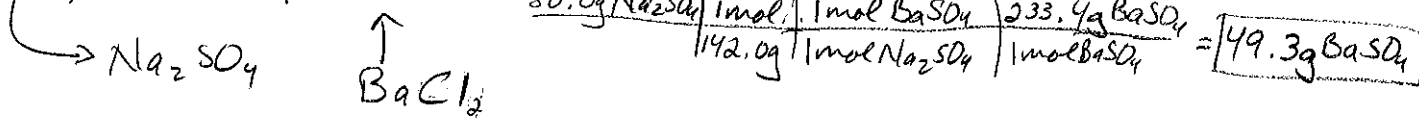
$$0.150 \text{ dm}^3 \left| \frac{1.5 \text{ mol}}{1 \text{ dm}^3} \right| \frac{208.2 \text{ g BaCl}_2}{1 \text{ mol}} = 46.85 \text{ g BaCl}_2 \text{ or } 0.225 \text{ mol BaCl}_2$$

50. 150 mL of 1.5 mol dm⁻³ BaCl₂ and 30.0g Na₂SO₄ are placed in a beaker to react. $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow$



b) Which reactant is the limiting reactant?

c) Which reactant is present in excess?



d) If 20 g of BaSO₄ is isolated, what is the percent yield?

$$\frac{20 \text{ g}}{49.3 \text{ g}} \times 100 = 40.6\% \text{ yield}$$

$$0.042 \text{ mol Ca(NO}_3)_2 \left| \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{3 \text{ mol}} \right| \frac{342.15 \text{ g}}{1 \text{ mol}} = 4.79 \text{ g Al}_2(\text{SO}_4)_3$$

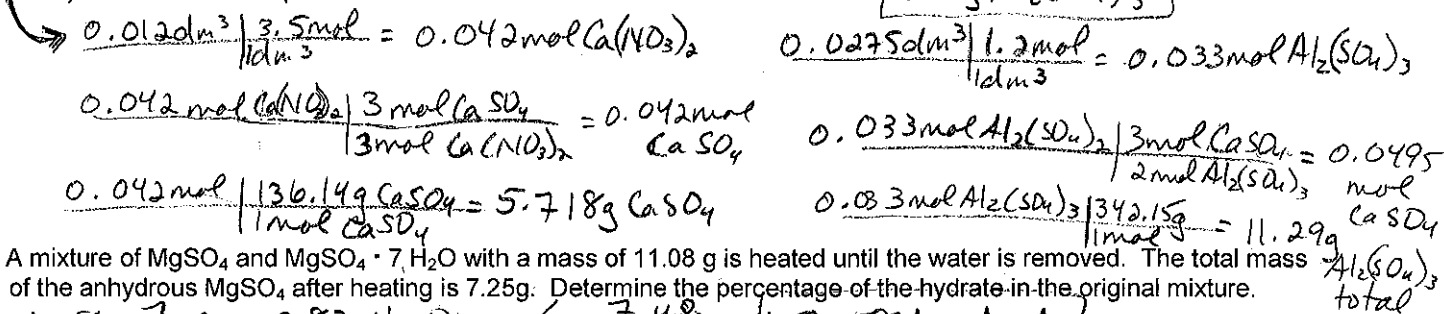
11.29 g - 4.79 g = 6.5 g Al₂(SO₄)₃ in excess

51. 12 mL of 3.5 mol dm⁻³ calcium nitrate and 27.5 ml of 1.2 mol dm⁻³ aluminum sulfate are placed in a beaker to react.

a) What is the theoretical yield of CaSO₄? 5.72 g CaSO_4

b) Which reactant is the limiting reactant? $\text{Ca(NO}_3)_2$

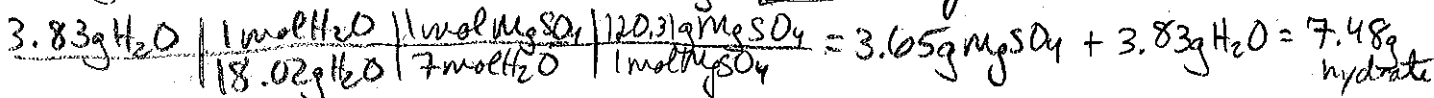
c) Which reactant is present in excess? How much will remain unreacted? $6.5 \text{ g Al}_2(\text{SO}_4)_3$



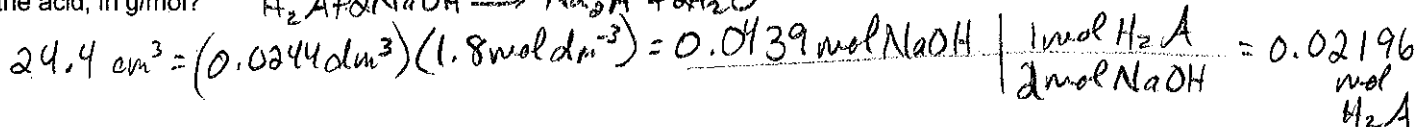
52. A mixture of MgSO₄ and MgSO₄ · 7H₂O with a mass of 11.08 g is heated until the water is removed. The total mass of the anhydrous MgSO₄ after heating is 7.25g. Determine the percentage of the hydrate in the original mixture.

$11.08 \text{ g} - 7.25 \text{ g} = 3.83 \text{ g H}_2\text{O}$

$\frac{7.48 \text{ g}}{11.08 \text{ g}} = 67.5\% \text{ hydrate}$

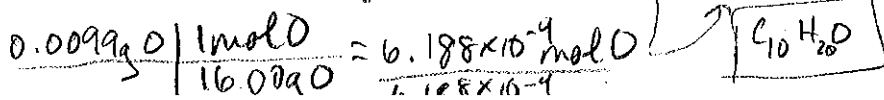
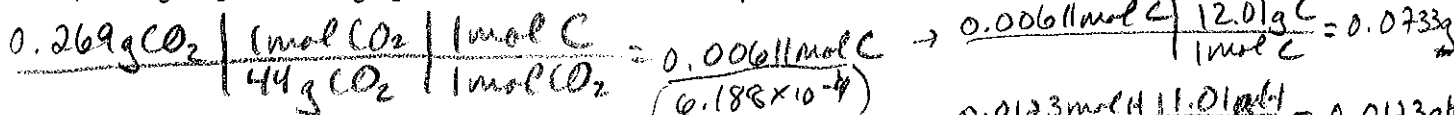


53. If 12.56 g of a solid dibasic acid reacts completely with 24.4 cm³ of 1.8 mol dm⁻³ NaOH, what is the molar mass of the acid, in g/mol?



$$\frac{12.56 \text{ g}}{0.02196 \text{ mol}} = 571.95 \text{ g/mol}$$

54. Menthol is a compound that contains only carbon, hydrogen and oxygen. When a 0.0956g sample of menthol burns in air, 0.269 g CO₂ and 0.110 g H₂O are formed. What is the empirical formula of menthol?



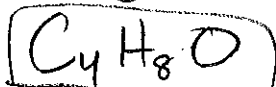
0.007321g

55. 7.321 mg of an organic compound containing carbon, hydrogen, and oxygen was analyzed by combustion. The amount of carbon dioxide produced was 17.873 mg and the amount of water produced was 7.316 mg. Determine the empirical formula of the compound.

$$C: \frac{17.873 \text{ mg}}{1000 \text{ mg}} \times \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 4.06 \times 10^{-4} \text{ mol C} \rightarrow \frac{4.06 \times 10^{-4} \text{ mol}}{1 \text{ mol}} \times 12.01 \text{ g} = 0.00488 \text{ g C}$$

$$H: \frac{7.316 \text{ mg}}{1000 \text{ mg}} \times \frac{1 \text{ mol H}_2\text{O}}{18.01 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 8.12 \times 10^{-4} \text{ mol H} \rightarrow \frac{8.12 \times 10^{-4} \text{ mol}}{1 \text{ mol}} \times 1.01 \text{ g} = 8.20 \times 10^{-4} \text{ g H}$$

$$0.007321 \text{ g} - (0.00488 \text{ g C} + 8.20 \times 10^{-4} \text{ g H}) = 0.001621 \text{ g O} \rightarrow \frac{0.001621 \text{ g O}}{16.00 \text{ g O}} = 1.01 \times 10^{-4} \text{ mol O}$$



56. 0.1101 gram of an organic compound containing carbon, hydrogen, and oxygen was analyzed by combustion. The amount of carbon dioxide produced was 0.2503 gram and the amount of water produced was 0.1025 gram. A determination of the molar mass of the compound indicated a value of approximately 115 grams/mol. Determine the empirical formula and the molecular formula of the compound.

$$C: \frac{0.2503 \text{ g CO}_2}{44.0 \text{ g CO}_2} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.005689 \text{ mol C} \rightarrow \frac{0.005689 \text{ mol C}}{1 \text{ mol C}} \times 12.01 \text{ g} = 0.06832 \text{ g C}$$

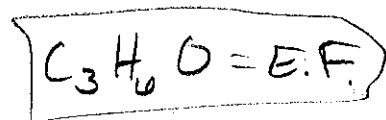
$$H: \frac{0.1025 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.01138 \text{ mol H} \rightarrow \frac{0.01138 \text{ mol H}}{1 \text{ mol H}} \times 1.01 \text{ g} = 0.01149 \text{ g H}$$

$$0.1101 \text{ g} - (0.06832 \text{ g C} + 0.01149 \text{ g H}) = 0.0303 \text{ g O}$$

$$\frac{0.0303 \text{ g O}}{16.0 \text{ g O}} = 0.001893 \text{ mol O} = 1$$

$$C: \frac{0.005689 \text{ mol C}}{0.001893} = 3$$

$$H: \frac{0.01138 \text{ mol H}}{0.001893} = 6$$



↳ mass = 58.06 g/mol

$$\frac{115 \text{ g mol}^{-1}}{58.06 \text{ g mol}^{-1}} = 1.98$$

$E.F. \times 2 = M.F.$

