Pearson Edexcel

STUDY PACK 1.3

FORMULAE & AMOUNTS OF SUBSTANCES

Since 2005
Chembase
School of Chemistry

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Chemical Formulae

1. Explain in detail an experiment to find the formula of Copper (II) Oxide.

2. Explain how you can modify the above apparatus to determine the formula of water?

3. What is Empirical formula & Molecular formula? Hence write an equation to show the relationship between the two.

4. The empirical formula of a compound X is C_2H_5. It has a molar mass of 58 $gmol^{-1}$, what is the molecular formula of X?

5. The empirical formula of a hydrocarbon is CH. It has a RMM of 78. What is the molecular formula?

6. The empirical formula of a compound is C_3H_6O. It has a molar mass of 58$gmol^{-1}$ what is the molecular formula?

7. Find the percentage composition of **each** element in following compounds

$Co(NH_2)_2$	$(NH_4)_2SO_4$	$Fe_2(SO_4)_3$
$CuSO_4.5H_2O$	$FeSO_4.7H_2O$	$Na_2CO_3.10H_2O$
$CH_3CH_2CH_2CH_3$	CH_3CH_2OH	CH_3CH_2COOH

8. **Find the empirical formula of the following compounds,**

a. 4g of "Hg" forms 4.64g of Mercury Sulphide (HgS)

b. 0.62g of "P" forms 1.10g of Phosphorus oxide

c. 5.60g of "Fe" forms 16.25g of Iron Chloride

9. Calculate the empirical formula of a compound with the following % composition.

C 39.13% **O** 52.17% **H** 8.70%

10. A gaseous compound of "C" & "H" contain 80% "C" by mass. It's molecular mass is 30.2 g/mol, Calculate it's empirical formula & molecular formula.

11. Sodium burns in excess oxygen to give a yellow solid oxide that contains 58.97% of sodium. What is the empirical formula? If the oxide has a molar mass of 78 $gmol^{-1}$ what it's molecular formula?

12. A compound **P** contains 73.47% carbon & 10.20% hydrogen by mass, the remainder being oxygen. It's found from other sources that **P** has a Relative Molecular Mass of 98. Calculate the molecular formula of **P.**

13. 24.6g of a hydrated salt of $MgSO_4.\mathbf{X}H_2O$ gives 12g of anhydrous $MgSO_4$ on heating. What is the value of **X**?

14. An 18.3g sample of a hydrated compound contained 4.0g of Ca, 7.1g of Cl & 7.2g of H_2O only, Calculate the empirical formulae?

15. A 22.3 grams of an oxide of Lead produced 20.7 grams of metallic lead on reduction with hydrogen. Calculate the empirical formulae of the oxide. (A; PbO)

16. A hydrocarbon containing 92.3% of carbon has a relative molecular mass of 26 $gmol^{-1}$. What is the molecular formula of the hydrocarbon? (Answer: C_2H_2)

17. Determine the formula of the minerals with the following mass compositions.

Na 12.1%, Al 14.2% Si 22.2% O 42.1% H_2O 9.48

18. When 1.335g of a chloride of Aluminium is added to excess silver nitrate solution 4.305 g of silver chloride is produced. Calculate the empirical formula of the chloride of aluminum. *Hint: you will need to work out how much chlorine is in 4.305g of AgCl. This will be the amount of chlorine in the initial 1.335g of Aluminium chloride* (Answer: $AlCl_3$)

19. A 16g of a hydrocarbon burn in excess oxygen to produce 44g of carbon dioxide. What is the empirical formula of the hydrocarbon?. *Hint: you will need to work out the mass of carbon in 44g of CO_2.* (Answer: CH_4)

20. When a chloride of phosphorus containing 85.1% chlorine is heated a second chloride containing 77.5% chlorine is produced. Find the formulae of the chlorides. (Answers: PCl_5, PCl_3)

Answers

1. The formula of Copper (II) oxide can be found by passing a stream of dry hydrogen gas over a known mass of Copper (II) oxide & weighing the copper formed.

Apparatus:

Dry hydrogen gas is passed over copper oxide before applying heat until all of air is sent out of the apparatus. This is done because a mixture of hydrogen & oxygen may explode when heated. Therefore, it is important to make sure that there is no air in the apparatus before applying heat. Once all the air has been sent out heating is done in the presence of hydrogen gas until a constant mass of copper is obtained inorder to ensure that all of copper oxide has turned to copper. At the end of the experiment record the mass of pure copper.

Results:

Mass of the apparatus :	100g
Mass of apparatus + copper (II) oxide :	105g
Therefore mass of copper (II) oxide :	5g
Mass of Copper :	4g
Mass of Oxygen :	1g (5g-4g)

Calculation:

2. Apparatus:

The apparatus of the previous experiment may be modified to determine the formula of water. A 'U' tube containing a known mass of anhydrous $CaCl_2$ is connected to the end of the combustion tube. Anhydrous $CaCl_2$ absorbs water & mass of the 'U' is determined at the beginning & at the end of the experiment. The increase in mass of the 'U' tube equals the mass of water.

Results :

Mass of water : 1.125g

Mass of 'O' : 1g (from the previous experiment)

Therefore mass of 'H' : 0.125g (1.125g-1g)

Calculation:

3. Empirical formula is the formula showing the simplest ratio of elements in a compound. The mass of the substance of this formula will be the smallest mass.

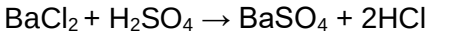
Molecular formula is the multiple of empirical formula.

Molecular formula mass = Empirical formula mass x n (n= whole number)

Answers

Calculations of products/reactants based on equations

1. What mass of barium sulfate would be produced from 10 g of barium chloride in the following reaction?



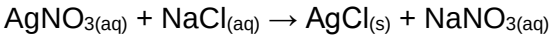
$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

2. What mass of potassium chloride would be produced from 20 g of potassium carbonate?

3. What masses of ethanol and ethanoic acid would need to be reacted together to give 1 g of ethyl ethanoate?

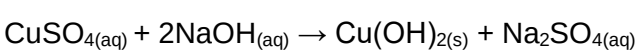
4. What mass of iron(III) oxide would need to be reduced to produce 100 tonnes of iron in a blast furnace?

5. What mass of silver nitrate as a solution in water would need to be added to 5 g of sodium chloride to ensure complete precipitation of the chloride?



$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

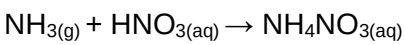
6. A solution of copper sulphate reacts with sodium hydroxide solution to produce a precipitate of copper hydroxide according to the following equation:



$CuSO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + Na_2SO_{4(aq)}$

What mass of sodium hydroxide would be needed to convert 15.95 g of copper sulfate to copper hydroxide and what mass of copper hydroxide would be produced?

7. What volume of ammonia gas would be needed to produce 40 g of ammonium nitrate in the following reaction?



$NH_{3(g)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

8. In the reaction between calcium carbonate and nitric acid what mass of calcium nitrate and what volume of carbon dioxide would be produced from 33.3 g of calcium carbonate?

9. What would be the total volume of gas produced by the action of heat on 33.1 g of lead(II) nitrate ?

10. .Sulfur dioxide can be removed from the waste gases of a power station by passing it through a slurry of calcium hydroxide. The equation for this reaction is:

$$SO_{2(g)} + Ca(OH)_{2(aq)} \rightarrow CaSO_{3(aq)}) + H_2O(l)$$

What mass of calcium hydroxide would be needed to deal with 1000 dm3 of Sulfur dioxide?

11. In a fermentation reaction glucose is converted to alcohol and carbon dioxide according to the following equation:

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

What mass of alcohol and what volume of carbon dioxide would be produced from 10 g of glucose?

12. In the following reactions calculate the mass of precipitate formed from 20 g of the metal salt in each case.

(i) $ZnSO_{4(aq)} + 2NaOH \rightarrow Zn(OH)_{2(s)} + Na_2SO_{4(aq)}$

(ii) $Al_2(SO_4)_{3(aq)} + 6NaOH \rightarrow 2\ Al(OH)_{3(s)} + 3Na_2SO_{4(aq)}$

(iii) $MgSO_{4(aq)} + 2NaOH \rightarrow Mg(OH)_{2(s)} + Na_2SO_{4(aq)}$

13. What volume of hydrogen would be produced by 1 g of calcium in its reaction with water?

14. What mass of magnesium would be needed to produce 100 cm^3 of hydrogen?

Answers

Reactions involving gases

Whenever gases are involved in a reaction you need to remember that they have both mass and volume and that **1 mole of any gas has the same volume, 24 000 cm^3, at room temperature and 1 atmosphere pressure as 1 mole of any other gas**. (See *Section 4* for more details)

This means:

2 g of hydrogen, H_2, has a volume of 24 000 cm^3

32 g of oxygen, O_2, has a volume of 24 000 cm^3

81 g of hydrogen bromide, HBr, has a volume of 24 000 cm^3

The effect of this is to make calculations involving gas volumes much easier than you might expect.

Consider the following reaction:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

This says:

2 moles of NO(g) react with 1 mole of O_2(g) to give 2 moles of NO_2(g)

$\therefore$ (2 x 24 000) cm^3 of NO react with (1 x 24 000) cm^3 of oxygen to give (2 x 24 000) cm^3 of NO_2

2 cm^3 of NO react with 1 cm^3 of oxygen to give 2 cm^3 of NO_2

ie for gases only the reacting volume ratios are the same as the reacting mole ratios in the equation.

Example 1

What volume of sulphur trioxide would be produced by the complete reaction of 100 cm^3 of sulphur dioxide with oxygen? What volume of oxygen would be needed to just react with the sulphur dioxide?

	$2SO_2(g)$	+	$O_2(g)$	$\rightarrow$	$2SO_3(g)$
Ratios	2		1		2
ie	1		½		1
	100 cm^3		50 cm^3		100 cm^3

Thus 100 cm^3 of sulphur dioxide will need 50 cm^3 of oxygen and produce 100 cm^3 of sulphur dioxide.

Example 2

What would be the composition of the final product in Example 1 if 100 cm^3 of oxygen had been used rather than 50 cm^3?

Since 100 cm^3 of the sulphur dioxide needs only 50 cm^3 of oxygen there must be 50 cm^3 of oxygen unused. Thus the final volume is:

100 cm^3 of sulphur dioxide plus 50 cm^3 of excess oxygen = 150 cm^3

Example 3

What volume of ammonia would be produced if 10 cm^3 of nitrogen was reacted with 20 cm^3 of hydrogen?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

You need to think before you start this question. The reacting volumes given in the question are not the same as those in the reaction. You must have excess of one of the gases.

From the equation:

10 cm^3 of nitrogen needs 30 cm^3 of hydrogen. You only have 20 cm^3 of hydrogen so the nitrogen is in excess**.**

In this case you will need to use the hydrogen volume in the calculation.

	$N_2(g)$	+	$3H_2(g)$	$\rightarrow$	$2NH_3(g)$
Ratios	1		3		2
	1/3		1		2/3
	1/3 x 20		20		2/3 x 20
	6.67 cm^3		20 cm^3		13.33 cm^3

Thus 20 cm^3 of hydrogen will react to give **13.33 cm^3 of ammonia** and there will be 3.33 cm^3 of hydrogen left over.

Calculations based on equations involving only gases

Section a

In *Section a*. assume that you have 10 cm^3 of the first named reactant and then calculate the volumes of all the gases involved in the equation. In these examples the reactions are being carried out at above 100°C and you should assume the water is present as a gas and therefore has a volume.

1	CH_4	+	$2O_2$	$\rightarrow$	CO_2	+	$2H_2O$
2	C_2H_4	+	$3O_2$	$\rightarrow$	$2CO_2$	+	$2H_2O$
3	$2C_2H_2$	+	$5O_2$	$\rightarrow$	$4CO_2$	+	$2H_2O$
4	$2C_8H_{18}$	+	$25O_2$	$\rightarrow$	$16CO_2$	+	$18H_2O$
5	N_2	+	$3H_2$	$\rightarrow$	$2NH_3$		

Section b

In *Section b.* you are asked to find the total volume of gas produced at room temperature and pressure. You should ignore the volume of water produced as this will have condensed as a liquid. Be careful in some cases, as there is an excess of one of the reactants.

1 What volume of oxygen would be needed to convert 1000 cm^3 of nitrogen monoxide, NO, to nitrogen dioxide, NO_2? (Assume all volumes are measured at the same temperature and pressure.)

2 In the production of sulphuric acid sulphur dioxide is converted to sulphur trioxide by reaction with the oxygen in the air. What volume of air (assume 20% of the air is oxygen) would be needed to produce 150 cm^3 of sulphur trioxide? Assume complete conversion of sulphur dioxide to sulphur trioxide.

3 What is the final volume of gas produced at room temperature when 10 cm^3 of methane is burnt with 30 cm^3 of oxygen?

4 What is the final volume of gas produced at room temperature if 5 cm^3 of octane are burnt with 100 cm^3 of oxygen?

5 In a reaction between methane and oxygen 60 cm^3 of methane was burnt with 60 cm^3 of oxygen. What is the composition of the gas mixture produced?

6 What volume of ammonia would be produced if 10 cm^3 of nitrogen was reacted with 60 cm^3 of hydrogen?

7 What would be the final volume of gas produced in the reaction between 10 cm^3 of hydrogen and 10 cm^3 of oxygen?

The percentage yield

A reaction may not always yield the total amount of product predicted by the chemical equation. The loss may be due to: -

- An incomplete reaction
- Errors in weighing the reactants & products
- Loss of material due to transferring & separating the product

The percentage yield can be calculated by using the below equation.

$$\textit{Percentage yield} = \frac{\textbf{Actual yield}}{\textbf{Expected yield}} \textbf{X 100 \%}$$

Example:

Heating 12.4 g of copper (II) carbonate in a crucible produced 7.0 g of copper (II) oxide. What is the percentage yield of copper (II) oxide?

$CuCO_3 \longrightarrow CuO + CO_2$

1mol	1mol	1mol
(64+12+48)	(64+16)	
=124g	=80g	

Therefore, heating 12.4g of $CuCO_3$ should produce 8.0g of CuO

Expected yield = 8.0g
Actual yield = 7.0g

Percentage yield = $\frac{7.0}{8.0}$ x 100 % = 87.5 %

Questions

1. Reacting 216g of Iron (III) Oxide with Carbon monoxide in the blast furnace produce 140 g of Iron. What is the percentage yield of Iron?

2. Heating 66.20 g of Lead (II) nitrate produced 40 g of Lead (II) oxide. What is the percentage yield of Lead (II) oxide?

3. Heating 20 g of Calcium Carbonate in a crucible produced 10.5 g of Calcium oxide. What is the percentage yield of Calcium oxide?

Concentrations in Parts Per Million (PPM)

1. **Explain what is meant by the term parts per million (ppm)?**

   Parts per million (ppm) is one part if a substance dissolved in a million parts of the solvent.

Parts per million (ppm) is used for expressing concentrations by mass. These units turn out to be convenient when the solute concentrations are very small (almost trace amounts). For example, if a solution has 1 ppm solute this would mean that 1 g of solvent would have one "millionth" gram of solute. Equivalently, 1 kg of this solvent will have 1 mg of solute etc.. By definition we have:

$$\frac{\text{Mass of Solute}}{\text{Mass of Solvent}} \times 10^6 = \text{Concentration in ppm}$$

For Liquids:
One milligram of particle in 1 litre of water equals to 1ppm
1 Litre of water = 1000grams
1 gram = 1000mg
Therefore 1 litre = 1000x1000mg =$1x10^6$mg
So 1mg in 1 litre of water = 1 mg /l OR 1ppm

For Solids:
1 kilogram (kg) = 1 million milligrams (mg)
so: 1 mg/kg = 1 part per million

For Gases:
A concentration of 1ppm means $1cm^3$ in 1 x 10^6 cm^3 of air.

$$\frac{\text{Volume of Gas}}{\text{Volume of Air}} \times 10^6 = \text{Concentration in ppm}$$

2. State the importance of ppm in daily analysis?

Parts per million (ppm) is most commonly used to describe very small amounts of contaminants in our environment. They are measures of concentration, the amount of one material in a larger amount of another material;

For example, the weight of a toxic chemical in a certain weight of food. They are expressed as concentrations rather than total amounts so we can easily compare a variety of different environmental situations.

Examples where ppm is used,

1. Concentration of a Toxic chemical in a sample of water
2. Amount of pollutant gases in atmosphere
3. Amounts of pollutants in exhaust fumes

3. A 155.3 g sample of pond water is found to have 1.7×10^{-4}g of phosphates. What is the concertation of phosphates in ppm?

$$\frac{1.7 \times 10^{-4} \text{ g Phosphates}}{1.553 \times 10^{2} \text{ g Solution}} \times 10^{6} = 1.1 \text{ ppm}$$

4. Now suppose we have 400 g sample of pond water and it has a concentration of 3.5 ppm dissolved nitrates. What is the mass of dissolved nitrates in this sample?

$$400 \text{ g solution} \times \frac{3.5 \times 10^{-6} \text{ g nitrates}}{1 \text{ g of solution}} = 1.4 \text{ mg nitrates}$$

5. Calculate the concentration in ppm of solution which contains 0.250g of solute dissolved in 675g of solvent?

6. Calculate the concentration in ppm of KNO_3 solution where 45mg of KNO_3 is dissolved in 950g of water. (note the units of solvent is in grams and solute in mg)

7. A sample of river water contains nitrate ions of concentration 23ppm. What is the mass of nitrate ions in 750g of river water?

8. Calculate the concentration in ppm of SO_2 gas, when 2.8dm^3 of SO_2 gas is mixed with 12000 dm^3 of air.

9. Calculate the volume of CO, when concentration of 89ppm is found in 6000 dm^3 of air.

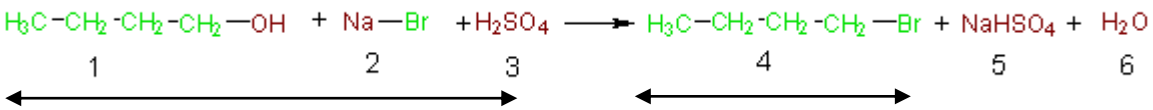
10. 500 dm^3 of air contains 75 cm^3 of Chlorine gas. What is the concentration in ppm?

11. Show by calculation which sample of has a higher concentration of NO_2. Sample 'A' contains 55 cm^3 of NO_2 in 750 dm^3 whereas Sample 'B' contains 1.8 dm^3 in 3500 dm^3.

Atom Economy

As indicated previously most chemists have traditionally measured the efficiency of a reaction by the percentage yield, however this only tells part of the story. If one considers the below reaction where a total of 4.13 g of reactants (0.8 g of 1-butanol, 1.33 g of NaBr and 2.0 g of H_2SO_4) was used, and that at best this reaction will only yield 1.48 g of the desired product, the question might be asked "what happens to the bulk (4.13 g -1.48 g = *2.7 g*) of the mass of reactants?". The answer is they end up in side products ($NaHSO_4$ and H_2O) that may be unwanted, unused, toxic and/or not recycled/reused. The side products are often times treated as wastes and must be disposed of or otherwise treated. At best only 36% (1.48 g/4.13 g X 100) of the mass of the reactants end up in the desired product. If the actual yield is 81% then only 29% (.81 X .36 X 100) of the mass of the reactants actually ends up in the desired product!

In an effort to foster awareness of the atoms of reactants that are incorporated into the desired product and those that are wasted (incorporated into undesired products), concept of ***atom economy*** was developed.



Reactants **main product**

Formula weight: 74 103 98 137

The percentage atom economy.

This can be calculated by taking the ratio of the mass of the main desired product (137) to the total mass of the atoms of all the reactants (275) and multiplying by 100. As shown below this reaction has only 50% atom economy.

% Atom Economy = (FW of the main desired product **/** FW of all reactants) X 100
= (137/275) X 100 = 50%

Thus at best (if the reaction produced 100% yield) then only half of the mass of the reactants would be incorporated into the desired product while the rest would be wasted in unwanted side products.

N.B : Reaction which has no bi products therefore would give 100% atom economy. Grater the amount of bi-products lower will be the atom economy.

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