

Name: _____

Date: _____

Period: _____

IB Chem 1 Stoichiometry Lesson 1 - Guided Notes

Slide 3: Write the balanced symbol equation for the reaction of aluminium burning in air to form aluminium oxide.

Write the word equation

Aluminium + oxygen → aluminium oxide

Write the correct molecular formulae

$\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$

Is this balanced?

- **Balancing Equations**

Slide 4: Write the balanced symbol equation for the reaction of aluminium burning in air to form aluminium oxide.

3. Balance the Al atoms on the left and right

$2\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$

4. Balance the O atoms on the left and right

$2\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$

Is this balanced?

- **Balancing Equations**

Slide 5: Write the balanced symbol equation for the reaction of aluminium burning in air to form aluminium oxide.

5. Re-balance the Al atoms on the left and right



Is this balanced?

- **Balancing Equations**

Slide 6: Write the balanced symbol equation for the reaction of aluminium burning in air to form aluminium oxide.



- **Balancing Equations**
- These numbers are called coefficients.
- What do these represent?

Slide 7: $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$

- **Balancing Equations**
- Coefficients:
- Tells you the molar ratio or the “_____” of the reaction.

Slide 8: $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

Write this in terms of all of the ions present.

2. Cross out the “spectator ions”.

3. Check that the charges balance.

- **Ionic Equations**

Slide 9: NEVER change the formulae (subscript numbers) to make it balance.
Leave elements in multiple places for last. Start on elements with bigger masses.
Too many steps might be a hint that you have the wrong formulae.
For ionic equations, the total charge must balance.
Double check balancing for ALL elements when you are done.

- Tips for _____ Equations

Slide 10: Ar takes into account all the stable isotopes of that element

- It is the average mass of that element. It allows for the relative mass and relative _____ (amount) of that element.
- Chlorine - 35
- Chlorine - 37
- _____ atomic mass (R.A.M or Ar)

Slide 11: Mr is just all of the relative atomic masses (Ar) added together.

- So if you wanted to work out how heavy a compound was, you would do the _____...
- **Relative _____ mass (RMM or Mr)**

Slide 12: Let's calculate some Mr

- **This is CO₂**
- **Made up from...**

Slide 13: Now you need to add up the mass of your sheep...

- 1 carbon = 1 x 12 = 12
- 2 oxygen = 2 x 16 = +32
- _____ 44
- **And that is simply the _____ formula mass of CO₂**

Slide 14: Calculating molar masses (Mr)

- **Who can answer these fastest? Calculate the Mr of:**
- Li I
- CH₄
- CO₂
- HCl
- H₂SO₄

- NaCl
- NH₃
- MgCl₂
- TiO
- 10. C₆H₁₂O₆
- 11. Fe₂O₃
- 12. C₄H₁₀
- 13. MnO₄
- 14. C₆H₁₂
- 15. KMnO₄
- 16. Co₂(SO₄)₃
- 17. Ni(CH₃CO₂)₂
- 18. Na₂HPO₄

Slide 15: $7+127 = 134$

$$12+(4 \times 1) = 16$$

$$12+16+16 = 44$$

$$1+35.5 = 36.5$$

$$2+32+(16 \times 4) = 98$$

$$23+35.5 = 58.5$$

$$14+(3 \times 1) = 17$$

$$24+(35.5 \times 2) = 95$$

$$48+16 = 64$$

- 10. $(6 \times 12) + (12 \times 1) + (16 \times 6) = 180$
- 11. $(2 \times 56) + (3 \times 16) = 160$
- 12. $(4 \times 12) + (10 \times 1) = 58$
- 13. $55 + (4 \times 16) = 119$
- 14. $(6 \times 12) + (12 \times 1) = 84$
- 15. $39 + 55 + (16 \times 4) = 158$
- 16. $(2 \times 59) + (32 \times 3) + (16 \times 12) = 406$
- 17. $59 + (12 \times 2) + (3 \times 2) + (12 \times 2) + (16 \times 4) = 177$
- 18. $(23 \times 2) + 1 + 31 + (16 \times 4) = 142$

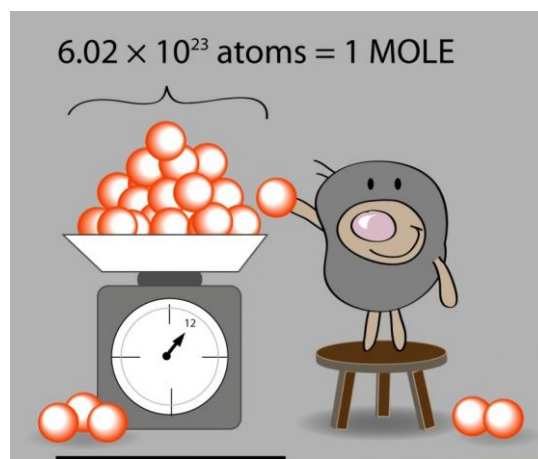
Slide 16: Theory of knowledge

- Assigning numbers to the masses of the chemical elements has allowed chemistry to develop into a physical science. Why

is mathematics so effective in describing the natural world?

Slide 17: Avogadro's constant

- This is the number of atoms (or ions, _____ or particles) in one mole of any substance (given in the data booklet/paper 1 q's).
- **Atoms are too small and light to weigh on a _____**
- **But if you know the mass of a very large number of atoms, then we can _____ the mass on one atom.**



Slide 18: Moles triangle

- Or g/mol



Slide 19: Stoichiometry - The branch of chemistry concerned with relative quantities (moles) of reactants and products in a chemical reaction.

Relative atomic mass (RAM/Ar) - The average mass of naturally occurring atoms of an element compared to 1/12th ^{12}C .

Relative molecular mass/molar mass (RMM/Mr) - The sum of all the relative atomic masses.

The mole - The amount of a substance containing 6.02×10^{23} particles (atoms, molecules, ions).

Avogadro's constant (or number) - The number of particles in one mole of a substance (6.02×10^{23}).

- Moles definitions

Slide 20: UNITS?

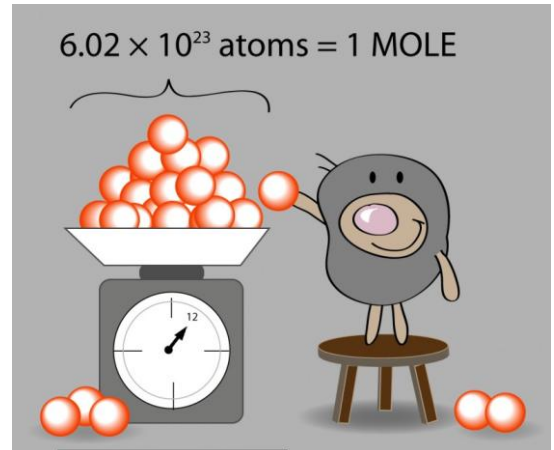
- THE AVOGADRO CONSTANT
- 6.022×10^{23}
- I do not need to write units after this number... why not?
- Since saying this whole number can be quite a mouthful, we call it a Mole.
- Watch the MOLE song to get an idea of just how big this number is!!!



Slide 22: Calculating numbers of particles

Slide 23: Avogadro's constant

- The magnitude of Avogadro's constant is beyond the scale of our everyday experience. How does our everyday experience limit our intuition?



Slide 24: Moles questions

Slide 25: Moles answers

- $0.100 \times 6.9 = 0.69 \text{ g}$
- 2.00 mol
- 1.00 mol

MOLE CALCULATIONS – EASIER - ANSWERS

- a) i) $\text{mass} = \text{mol} \times \text{RAM} = 0.100 \times 2 = 0.20\text{g}$
ii) $\text{mass} = \text{mol} \times \text{RAM} = 0.200 \times \text{NaBr} = 0.200 \times (23+80) = 0.200 \times 103 = 20.6\text{g}$
b) i) $\text{moles} = \text{mass} / \text{RAM} = 0.150 / \text{Fe} = 0.150 / 56 = 0.003 \text{ mol}$
ii) $\text{moles} = \text{mass} / \text{RAM} = 1.23 / [(2 \times 23) + 32 + (4 \times 16)] = 1.23 / 142 = 0.009 \text{ mol}$
c) i) $\text{RAM} = \text{mass}/\text{mol} = 13.59 / 0.107 = 127\text{g}/\text{mol}$ therefore it must be Iodine
ii) $\text{RAM} = \text{mass}/\text{mol} = 0.0232 / 0.00033 = 70\text{g}/\text{mol}$ therefore it must be Gallium
- 2) a)
b)
- 3) How many moles are present in 34 grams of $\text{Cu}(\text{OH})_2$? 0.35 moles
- 4) How much does 4.2 moles of $\text{Ca}(\text{NO}_3)_2$ weigh? 689 grams
- 5) What is the molar mass of MgO ? 40.3 g/mol
- 6) How are the terms "molar mass" and "atomic mass" different from one another? "Molar mass" is used to describe the mass of one mole of a chemical compound, while "atomic mass" is used to describe the mass of one mole of an element or the mass of one atom of an element.

Slide 26: Moles questions

Slide 27: Moles questions

Slide 28: Moles questions

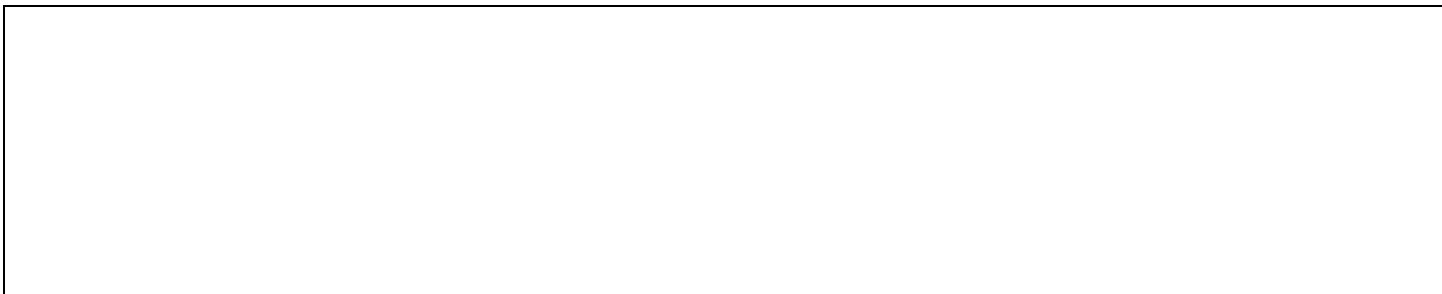
Slide 29: Moles questions

Slide 30: Moles questions

Slide 31: Moles questions



Slide 32: Moles questions

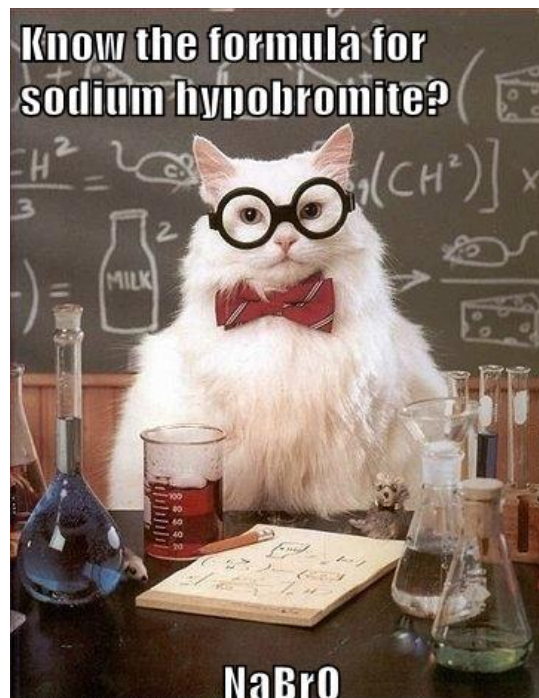


Slide 33: Moles - check your work

- C
- D
- B
- A
- D
- C
- B
- A
- D
- D
- A
- D
- D

Slide 34: What is the same within each set? What is different?

- **Empirical Formulae**



Slide 35: Empirical formula - simplest whole number ratio of atoms of each element present in the compound.

Molecular formula - is the actual number of atoms of each element in a compound

- **Definitions**

Empirical	Molecular
Simplest whole number ratio of elements	Actual whole number ratio Multiple of Empirical
CO_2	CO_2
H_2O	H_2O
NO_2	N_2O_4
P_2O_5	P_4O_{10}

Slide 36: CH2

H2O

CH2O

- C₆H₁₂
- EF or MF?

Slide 37: Calculate the EF of a compound with 1.26g of titanium and 3.74g of chlorine

Empirical formula is the lowest possible ratio of elements in a compound

Divide each mass by the Ar for that element

Divide these values by the smallest to find the ratio

Use the ratio to write the formula

- Worked Example

Slide 38: 18.39g of oxygen and 81.61g of chlorine are combined in 100g of a compound. Calculate the EF of the compound.

Divide each mass by the Ar for that element

Divide these values by the smallest to find the ratio

Use the ratio to write the formula

- Try it yourself

Slide 39: 28g of iron and 12g of oxygen are combined in of a compound. Calculate the EF of the compound.

- Non-integer example (.5)

Slide 40: 28g of iron and 18.8g of oxygen are combined in of a compound. Calculate the empirical formula of the compound.

- Non-integer example (.35)

Slide 41: A compound contain 16.4% potassium, 30.0% chlorine and 53.6% iodine by mass. Calculate the EF of the compound.

Divide each mass (%) by the Ar for that element

Divide these values by the smallest to find the ratio

Use the ratio to write the formula

- Using _____ mass

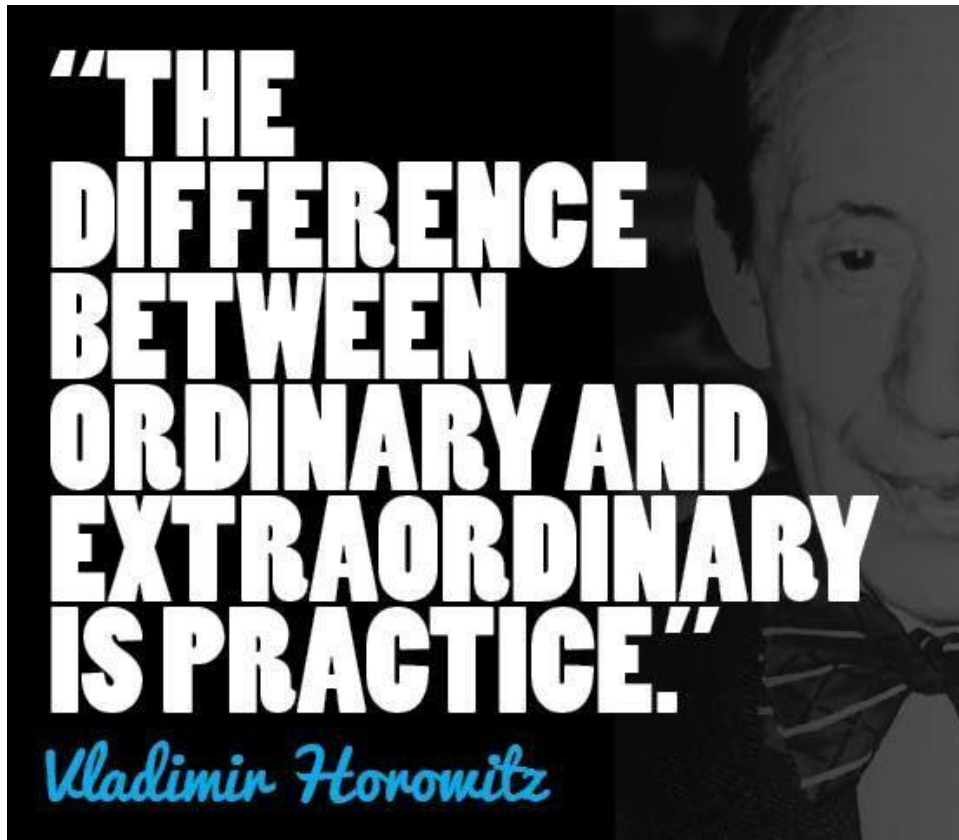
Slide 42: P: 0.282 g, O:0.218 g Mr:220

1) First calculate EF:

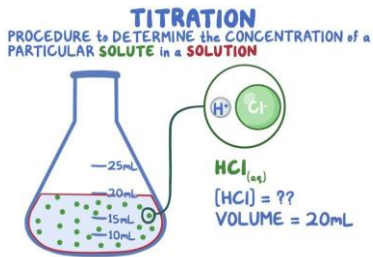
- 2) Now calculate MF
- $\text{Mr}(\text{EF P}_2\text{O}_3) = (2 \times 31) + (3 \times 16) = 110$
- $\text{Mr}(\text{compound}) = 220$
- So $\text{Mr} / \text{EF} = 220/110 = 2$ _____ MF is P_4O_6
- _____ MF from EF

Slide 44: Exam question

Slide 48: Practice makes perfect



Slide 49: Amount of Chemical Change



Slide 50: Conc =

Mol =

Conc =

- Calculating _____
- 1 mol in 1 dm³
- 1 mol dm⁻³
- 58.5g NaCl in 1 dm³ water
- 58.5/1 = 58.5 g/dm³
- 58.5g/58.5 g mol⁻¹ = 1 mol
- IB units of _____ : mol dm⁻³

$$5 \text{ mol in } 1 \text{ dm}^3$$

$$5 \text{ mol dm}^{-3}$$

$$292.5 \text{ g NaCl in } 1 \text{ dm}^3 \text{ water}$$

$$292.5/1 = 292.5 \text{ g/dm}^3$$

$$292.5 \text{ g} / 58.5 \text{ g mol}^{-1} = 5 \text{ mol}$$

Slide 51: Calculating concentration (mol dm⁻³)

- Moles
- (mol)
- Volume
- (dm³)
- Concentration
- (mol dm⁻³) .

Slide 52: Calculating concentrations

- $M_r \text{ NaOH: } 23 + 16 + 1 = 40$
- Mol NaOH: 80 g/40 g mol⁻¹
- = 2 mol NaOH
- $M = 2 \text{ mol NaOH} / 1 \text{ dm}^3 = 2 \text{ mol dm}^{-3}$

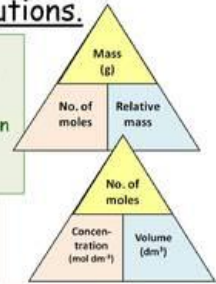
Concentrations of solutions.

Example 1:

80 g of sodium hydroxide (NaOH) are dissolved to make a 1 litre solution. What is the concentration of sodium hydroxide in mol dm⁻³?
(A_r: H = 1, O = 16, Na = 23,)

Step 1: (using earlier formula)

No. of moles of NaOH = 80 / 40
= 2 moles.



Slide 53: You-Try Calculating concentration

- **Exercise 1 – Calculate the _____ of the following solutions (show your working).**
- a) 10 mol of _____ chloride in 1dm³ of solution
- b) 1.5 mol of _____ iodide in 150 cm³ of solution
- c) 2.3 mol of lithium _____ in 500 cm³ of solution
- **Exercise 2 – _____ the volume required to make the following solutions (show your working).**
- a) 0.5 mol of a 2 mol dm⁻³ _____ of silver nitrate
- b) 250 mol of a 1.5 mol dm⁻³ _____ of sodium chloride
- c) 10 mol of a 0.4 mol dm⁻³ solution of sodium _____
- Pick 2 options from each _____ to complete

Slide 54: Calculating concentration

- **Exercise 3 – Calculate the _____ (mol dm⁻³) of the following solutions (show all your working).**
Hint - There are 2 steps!
- a) 10g of _____ chloride (MgCl₂) in 1dm³ of solution
- b) 1.5g of _____ iodide (KI) in 150cm³ of solution
- c) 2.3g of lithium _____ (LiCl) in 500cm³ of solution
- **Exercise 4 – _____ the mass required to make the following solutions (show all your working).**
Hint - There are 2 steps!
- a) 0.5dm³ of a 2 mol dm⁻³ _____ of silver nitrate (AgNO₃)
- b) 250cm³ of a 1.5 mol dm⁻³ _____ of sodium chloride (NaCl)
- c) 10cm³ of a 0.4 mol dm⁻³ solution of sodium _____ (NaCO₃)
- d) 500cm³ of a 5 mol dm⁻³ solution of _____ sulphate (MgSO₄)
- Pick 2 options from each _____ to complete

Slide 55: Check Your Work

- **Exercise 1 –**
- 10 mol dm⁻³
- 10 mol dm⁻³
- 4.6 mol dm⁻³
- **Exercise 2 –**
- 0.25 dm³
- 166.7 dm³
- 25 dm³
- **Exercise 3 –**
- 0.1 g

- 0.06 g
- 0.1 g
- **Exercise 4 –**
- 169.87 g
- 21.92 g
- 0.33 g
- 300.95 g

Slide 56: Diluting solutions

- **Starting volume = V_1**
- **Total volume = V_2**
- **Water added = $V_2 - V_1$**
- **Start conc = M_1**
- **Final conc = M_2**
- $M_1V_1 = M_2V_2$



Slide 57: Worked example

- We have a 10cm^3 of a solution of 1.5 mol dm^{-3} HCl. We need to make a solution that is only 0.1 mol dm^{-3} . What would the **final volume** of the solution be?
- $M_1V_1 = M_2V_2$

Slide 58: Worked example - final volume

- We have a 10 cm^3 of a stock solution of 1.5 mol dm^{-3} HCl. We need to make a solution that is only 0.1 mol dm^{-3} . What would the **final volume** of the solution be?
- $1.5 \times 0.010 = 0.1 \times V_2$
- **Rearrange**
- $1.5 \times 0.010 = V_2 \times 0.1$
- 0.1
- $0.15\text{ dm}^3 = V_2$ (final volume)
- $M_1V_1 = M_2V_2$
- Make sure your units are consistent!

Slide 59: Worked example - How much water to add?

- We have a 10cm^3 of a stock solution of 1.5 mol/dm^3 HCl. We need to make a solution that is only 0.1 mol/dm^3 .
What volume of water would we need to add?
- $1.5 \times 0.010 = 0.1 \times V_2$
- $1.5 \times 0.010 = V_2$
- 0.1
- $0.15\text{ dm}^3 = V_2$ (final volume)
- We are adding water to the starting solution (stock solution), so we need to subtract this from our final volume. Water added = $V_2 - V_1$
- $V_2 - V_1 = 0.15\text{dm}^3 - 0.01\text{dm}^3 = 0.14\text{ dm}^3$
- $M_1V_1 = M_2V_2$

Slide 60: Let's practice this!

- You add 200cm^3 of water to 300cm^3 of a 2 mol dm^{-3} solution of NaCl. Calculate the new concentration.
- What volume of water do we need to add to 0.5 mol dm^{-3} of HCl to make a solution of 400cm^3 of 0.1 mol dm^{-3} ?
- Show your work!

Slide 61: Check your answers

- 1.2 mol dm^3
- 0.32 dm^3

Slide 62: Making solutions

Slide 63: Making standard solutions

- Using a precise mass balance (to at least 2 decimal places) _____ measure the mass of the solute.
- Dissolve the solute in a small volume of _____ water in a beaker and stir with a stirring rod.
- Use a funnel to transfer the solution to a _____ flask.
- Rinse the beaker, stirring rod and funnel with distilled water and pour this water into the volumetric flask. This ensures that all the solute is _____ to the flask.
- Make up to the required volume in the _____ flask with distilled water.
- Add the stopper to the flask and invert several times to ensure _____ mixing.

Slide 64: This ensures that any solid left in the weighing boat is accounted for (alternatively the weighing boat could be rinsed into the beaker).

- **Stage 1: _____ a known mass of the solid**
- Weigh the solid into a weighing boat recording the mass (to _____ precision)
- _____ the solid to a beaker
- Re-weigh the weighing boat and record the _____ in mass
- **Stage 2: _____ the solid in distilled water**
- Add the minimum _____ water to the beaker to dissolve
- Stir the beaker with the glass rod until the solid has _____
- **Stage 3: rinsing all _____ and making up the solution**
- Transfer the solution to a (250cm³) _____ flask (using a funnel)
- Rinse all used glassware into the _____ flask (glass rod, beaker and funnel)
- Add distilled water up to the _____ line (on the volumetric flask)
- Invert the _____ times
- Adding too much distilled water at this stage may result in not enough left to rinse the equipment without _____ the final volume.
- All glassware that has been in contact with the solution must be rinsed to ensure every bit of solute finds its way into the _____ flask
- The final solution needs to be fully mixed before use (do not do this stage before the solution has been made up to the _____).
- _____ - **Making up a solution**

Slide 65: Turn and Talk

- Why is it not advised to pour the solid directly into a volumetric flask when making a solution?

Slide 66: You Try! Exam-Style Question

- A scientist is planning to make up a 1.75 mol dm^{-3} standard solution of sodium carbonate for a titration. When weighing the solid sodium carbonate out they spill a small amount of solid onto the balance pan. Furthermore when adding distilled water to the volumetric flask, they go slightly over graduation, but use dropping pipette to remove some of the solution to bring it back onto the graduation.
- 1. If the scientist is to make up 250 cm^3 of standard solution, calculate the mass of sodium carbonate required.
- 2. The scientist makes two clear errors during their method, explain how each error will affect the concentration of their solution.
- 3. **Challenge** - The scientist then wants to make a new standard solution of 0.09 mol dm^{-3} , using their original solution. Describe how they would prepare the new solution.

Slide 67: Mark your work

Slide 68: Primary standard solutions

- A standard solution is a solution with an accurately known concentration. A primary standard solution is made using a primary standard. A primary standard is a substance that has the following properties:
- High purity (99.9 %).
- High molar mass.
- Low reactivity.
- Does not change composition in contact with air.
- Examples:
- Sodium carbonate (Na_2CO_3)
- Oxalic Acid ($\text{H}_2\text{C}_2\text{O}_4$)
- Potassium hydrogen iodate ($\text{K}(\text{HIO}_3)_2$)
- Potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$)

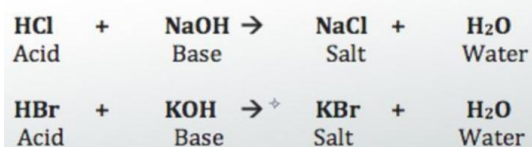
Slide 69: Secondary standard solutions

- A **secondary standard solution** is a solution that has been standardised against a primary standard solution. For example, a sodium hydroxide solution can be standardised against a primary standard solution, and can then be used as a secondary standard solution.
- The use of standard solutions with accurately known concentrations is particularly important in **volumetric analysis**.
- uses the technique of titration, in which the concentration of a solution is determined by reacting it with a standard solution of known concentration.

Slide 70: Volumetric analysis

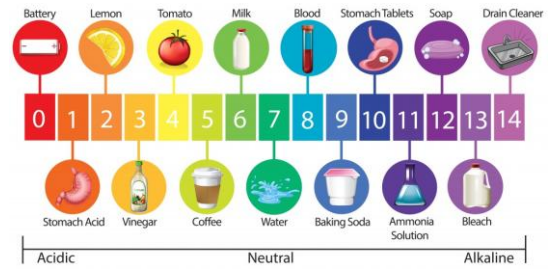
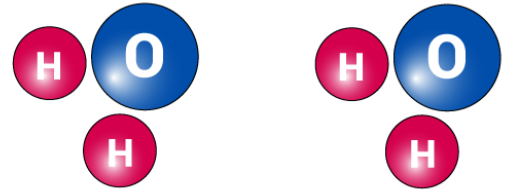
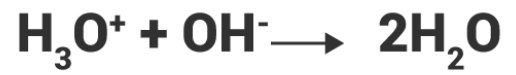
Slide 71: What is a titration?

- a technique that uses a reaction to calculate the unknown concentration (mol dm^{-3}) of a solution
- This is most commonly a neutralisation reaction between an acid and an alkali.

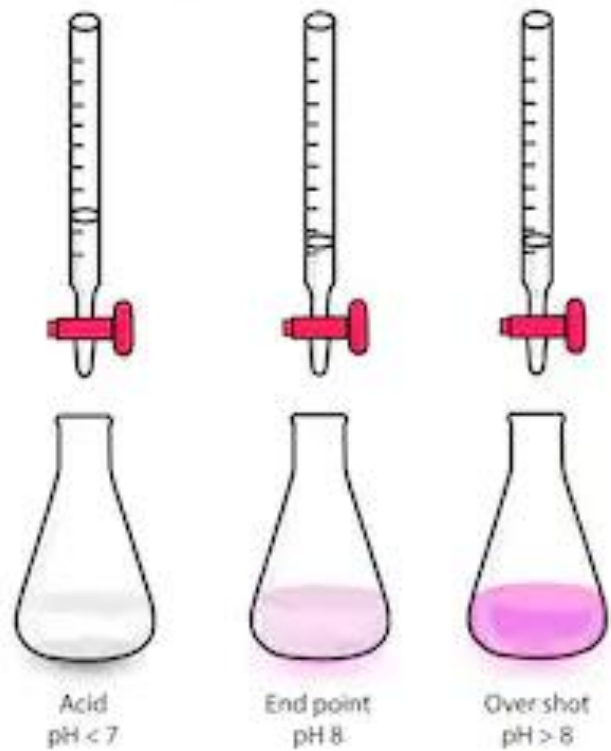


Slide 72: Neutralisation

- Occurs when an acid and base (or alkali - a soluble base) react to form a neutral _____ of salt and water.

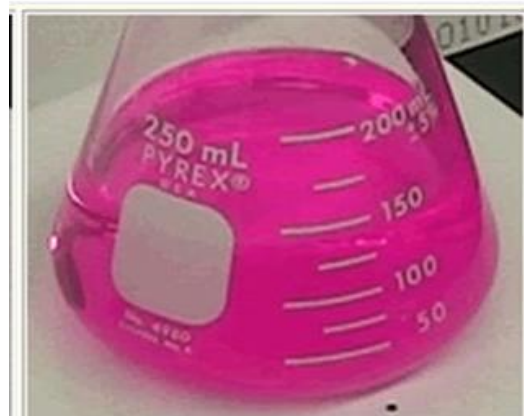


Slide 73: Acid-base titration

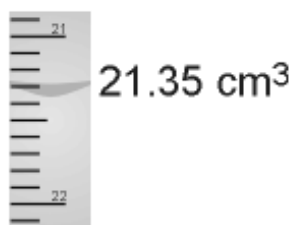
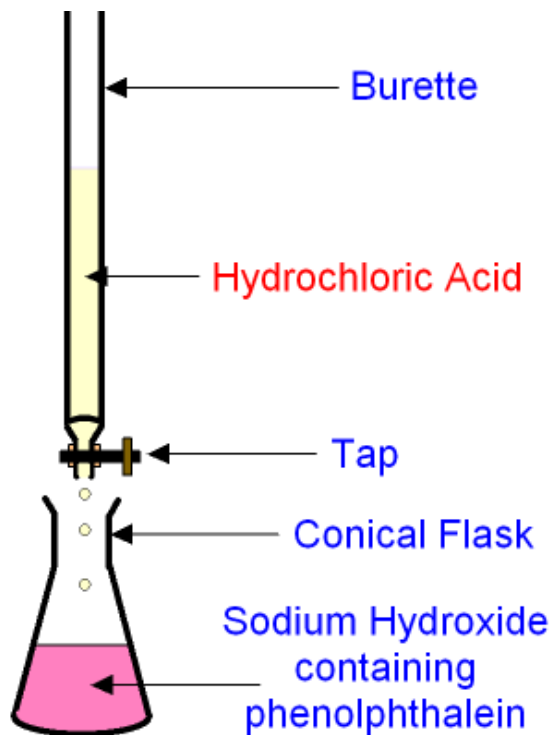


Slide 74: Using a pipette and burette correctly

- BEFORE AT AFTER
- _____ endpoint endpoint
- TOP OF _____ AT EYE LEVEL OR BELOW!!
- The known _____ in the burette is called the titrant
- The unknown _____ in the flask is called the analyte



Slide 75: RECORD TITRATIONS TO 2dp



	Rough	Accurate			
		1st	2nd	3rd	etc.
Final Reading (cm ³)					
Initial Reading (cm ³)					
Titre (cm ³)					

Slide 76: Carry out a titration

- **Stage 1: _____ the conical flask**
- Using a glass pipette, transfer solution with **unknown** _____ into the conical flask.
- Add several drops of an _____ indicator to the conical flask.
- Rinse any solution on the sides of the flask into the flask with _____ water.
- **Stage 2: _____ the burette**
- Fill the burette up with the solution of **known** _____ using a funnel.
- Run the solution through (into waste _____) ensuring that there are no air bubbles present.
- Remove the funnel from the _____.
- **Stage 3: _____ the titration**
- Place the _____ flask on a white piece of paper beneath the burette.
- Record the _____ volume of the burette (often this will be 0.00cm³).

- Titrate the _____, swirling the conical flask until a colour change is observed.
- Record the end volume in the burette and _____ the titre (volume of solution it takes to reach the end-point) .
- Repeat until _____ results (within 0.20 cm³) are achieved (at least 4 runs) and calculate a mean titre.

Slide 77: How to calculate your unknown concentration

- **Draw a sketch**, including down all the volumes and _____ you know.
- Write down a _____ **chemical equation for the reaction**.
- Calculate the **moles of the solution** whose _____ you did know.
- Use your balanced equation to _____ the **moles of the reacting substance**.
- Using moles and volume, you can now work out the _____ **of the unknown substance**.

Slide 78: How to calculate your unknown concentration

- $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
- In this experiment, an unknown _____ of 100 cm³ NaOH is titrated with 0.120 mol dm⁻³ HCl

Slide 79: How to calculate your unknown concentration

- $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
- Note: you can calculate standard deviation for the _____ of your mean titre, but it is not required

Slide 80: How to calculate your unknown concentration

- $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
- _____ moles of titrant (HCl)
- $0.120 \text{ mol dm}^{-3} \times 0.02082 \text{ dm}^3 = 0.0024984 \text{ mol HCl}$
- Using _____, 1 mol of HCl reacts with 1 mol of NaOH, so there are also 0.0024984 mol NaOH that reacted in the flask
- There was originally 100 cm³ of NaOH in the flask. This means the _____ of NaOH is $0.0024984 \text{ mol} / 0.100 \text{ dm}^3 = 0.024984 \approx 0.0250 \text{ mol dm}^{-3}$

Slide 81: Propagation of uncertainty

- $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
- Assume the uncertainty of the HCl was 0.002 mol dm⁻³. The uncertainty of the volume of titrant was 0.10 cm³. Since these two numbers were multiplied, add **percent** _____.
- % uncertainty HCl _____: $0.002/0.120 = 0.01667 = 1.667\%$
- % _____ volume HCl: $0.0001 \text{ dm}^3/0.02082 \text{ dm}^3 = 0.004803 = 0.4803\%$
- Add percent _____: 2.15%
- **This is the _____ for the mol of NaOH as well**

Slide 82: Propagation of uncertainty

- $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$

- 2. To calculate the uncertainty of the final _____ of NaOH, percent uncertainties of the moles of NaOH and the volume of NaOH also have to be added. Assume the absolute uncertainty of the volume of NaOH was $\pm 0.05 \text{ cm}^3$.
- % _____ volume NaOH: $0.00005/0.100 \text{ dm}^3 = 0.0005 = 0.05\%$
- % uncertainty for NaOH _____: add % uncertainty moles NaOH and volume NaOH:
- $2.15\% + 0.05\% = 2.20\%$ _____
- Convert to absolute _____: $0.0250 \text{ mol dm}^{-3} \times 0.0220 = 0.00055 \text{ mol dm}^{-3}$
- _____ NaOH: $0.0250 \text{ mol dm}^{-3} \pm 0.0006 \text{ mol dm}^{-3}$

Slide 83: Titration results table - use every time

- _____ **SHOULD ALWAYS BE RECORDED TO 2DP**
E.g. 26.30 or 26.35 or 26.40

Slide 84: Understanding questions

- A teacher noticed that a student had not cleared a large air bubble from below the burette tap in preparing the burette for use before starting the titration. This air bubble was ejected during the first titration of the volumetric flask mixture:
- 1. State the effect that this would have on the first titration result.
- 2. State and explain the effect this would have on the average titre.
- In a titration experiment, a good technique is essential for an accurate result to be obtained:
- 3. Suggest a reason for removing the funnel after filling the burette
- 4. Suggest one other source of error in using a burette to carry out a titration
- During the titration, the inside of the conical flask is rinsed with distilled water:
- 5. Suggest why rinsing improves the accuracy of the titre.
- 6. Explain why this extra volume of water does not change the volume of titre required.
- 7. Suggest why repeating a titration can increase reliability.
- 8. Suggest why a 250cm^3 conical flask is preferred over a 250cm^3 beaker during a titration.

Slide 85: The value of the titre would be higher (than the true value)

2. It should have no effect as the first titration can be ignored / subsequent titrations would be accurate
 3. As a droplet from the funnel could enter the burette affect volume / readings / titre
 4. Air bubble in jet. NOT: misreading burette or overshooting endpoint
 5. (Returns) reagent on the sides of the flask to the reaction mixture (to ensure that all of the acid / alkali reacts)
 6. Water is not a reagent and does not change the number of moles of reagents
 7. Idea that a single titration could be flawed / anomalous
 8. Less chance of losing liquid on swirling, NOT: 'easier to swirl'
- Mark your work

Slide 86: Parts Per Million (ppm) is a measurement of the concentration of a solution. For very dilute solutions, concentrations are sometimes expressed in parts per million.

- Parts per _____ (ppm)
- $\text{ppm} = \text{mg dm}^{-3}$

Slide 87: Question 1. A solution has a concentration of 1.25 g dm^{-3} . What is its concentration in ppm?

Convert the mass in grams to a mass in milligrams:

$$1.25 \text{ g} = 1.25 \times 1000 \text{ mg} = 1250 \text{ mg}$$

Re-write the concentration in mg dm^{-3}

$$\text{concentration in } \text{mg dm}^{-3} = 1250 \text{ mg dm}^{-3} = 1250 \text{ ppm}$$

Question 2. A solution has a concentration of 0.5 mg cm^{-3} . What is its concentration in ppm?

Convert the volume in cm^3 to dm^3 :

$$\text{volume : } 1000 \text{ cm}^3 = 1 \text{ dm}^3$$

Re-write the concentration in mg dm^{-3}

$$\text{concentration in } \text{mg dm}^{-3} = 0.5 \text{ mg} \times 1000 \text{ cm}^3 / 1 \text{ dm}^3 = 500 \text{ mg dm}^{-3} = 500 \text{ ppm}$$

- Parts per million (ppm)

Slide 88: Question 1. A solution has a concentration of 2.50 g dm^{-3} . What is its concentration in ppm?

Question 2. A solution has a concentration of 0.32 mg cm^{-3} . What is its concentration in ppm?

- You Try! (ppm)

Slide 89: Question 1. 2,500 ppm

Question 2. 320 ppm

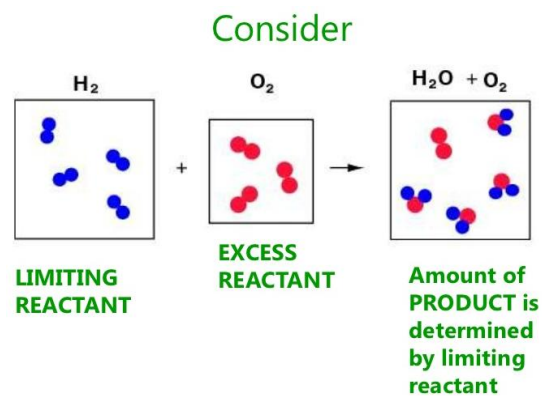
- Answers (ppm)

Slide 90: In some chemical reactions we have an too much of one reactant.
This means one reactant is in excess (too much) and one is a limiting reactant (limits the reaction).

- Limiting reactants

Slide 91: In order to work out which reactant is limiting, we will need to calculate the number of moles of each reactant and compare this to the molar ratio (big number/coefficient).

- Limiting reactants



Slide 92: Moles triangle

- g/mol
- RAM: _____ atomic mass (Ar)
- RFM: Relative formula mass (aka relative _____ mass, Mr)



Slide 93: First balance the equation!



- $\text{FeCl}_3 + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 + \text{Cl}_2$
- How many moles of chlorine gas can be produced if 6g of FeCl_3 react with 4g of O_2 ? What is the limiting reactant?
- **Example 1**

Slide 94: Example 1

- How many moles of chlorine gas can be produced if 6g of FeCl_3 react with 4g of O_2 ? What is the limiting reactant?
- **Don't use the coefficient to calculate Mr!!**

Slide 95: Example 1

- How many moles of chlorine gas can be produced if 6g of FeCl_3 react with 4g of O_2 ? What is the limiting reactant?
- **What should the molar ratio of FeCl_3 to O_2 be?**
- **4:3=1.33**
- **What do we actually have?**

Slide 96: Example 1

- How many moles of chlorine gas can be produced if 6g of FeCl_3 react with 4g of O_2 ? What is the limiting reactant?
- Molar ratio $0.0369/0.125 = 0.295$
- This is LESS than we need, so FeCl_3 is LIMITING!
- **What molar ratio is needed? 4:3 = 1.33**
- **What do we have?**

Slide 97: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2
- We have to work out which reactant is limiting BEFORE we can now which moles to use for the product.

Slide 98: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2

Slide 99: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2

Slide 100: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2
- Molar ratio $0.499/1.41 = 0.354$
- This is MORE than we need, so C_2H_6 is EXCESS! We use these moles of limiting reactant (O_2) for further calculations
- What molar ratio is needed? $2:7=0.286$
- What is it?

Slide 101: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2
- $1.41 \text{ mol O}_2 \times 6 \text{ mol H}_2\text{O} = 1.21 \text{ mol H}_2\text{O}$
- 7 mol O_2

Slide 102: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2

Slide 103: $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

If 15 g of C_2H_6 react with 45 g of O_2 , how many grams of water will be produced?

- Example 2

Slide 104: Example 3

- How much ZnCl_2 can be produced from the reaction of 2g of Zn and 3g of HCl? What is the limiting reactant?
- $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$
- Molar ratio $0.0308/0.0822 = 0.375$
- This is LESS than we need, so Zn is the limiting reactant.
- What molar ratio do we need? $1:2=0.5$
- What do we have?

Slide 105: Example 3

- How much ZnCl_2 can be produced from the reaction of 2g of Zn and 3g of HCl? What is the limiting reactant?
- $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$
- Molar ratio 1:1

Slide 106: Try it yourself

- $2\text{NaCl} + \text{Pb}(\text{NO}_3)_2 \rightarrow 2\text{NaNO}_3 + \text{PbCl}_2$
- How many grams of lead chloride are produced from the reaction of 15.3 g of NaCl and 60.8 g of $\text{Pb}(\text{NO}_3)_2$?
- What is the limiting reactant?
- How many grams of excess is left over?

Slide 107: Check Your Work

- NaCl limiting
- _____ 36.43g PbCl₂
- Excess Pb(NO₃)₂ 17.55g

Slide 108: Practice makes perfect

Slide 109: The yield of a chemical reaction

- **Do now:**
- Why did I ask you to calculate the MAXIMUM mass of the product?
- What name do we give this?
- What do you think a 'percentage yield' is?
- *(This is what we were doing last lesson)*

Slide 110: Definitions

- *(This is what we were doing last lesson)*
- *(Normally given in the question, or calculated by rearranging the equation below)*

Slide 112: Few reactions give a 100% yield.

Come up with 3 reasons why this is the case.

Incomplete reactions – not all reactions completely finish (there is some reactant left)

Losses during the practical e.g. purification, gas escaping

Unwanted by-products being produced

- **Why is % yield less than theoretical yield?**

Slide 113: Percentage yield in industry

- The concept of _____ yield is vital in monitoring the efficiency of industrial processes.

Slide 114: Easy Example

Slide 115: A reaction between 384g solid sulfur and excess oxygen produced 680 g sulfur dioxide. Calculate the percentage yield.

- **Harder Example**

Slide 116: A reaction between 384g solid sulfur and excess oxygen produced 680 g sulfur dioxide. Calculate the percentage yield.

- **Harder Example**

Slide 117: A reaction between 384g solid sulfur and excess oxygen produced 680 g sulfur dioxide. Calculate the percentage yield.

- Molar ratio S : SO₂ = 1:1
- **Harder Example**

Slide 118: A reaction between 384g solid sulfur and excess oxygen produced 680 g sulfur dioxide. Calculate the percentage yield.

- Molar ratio S : SO₂ = 1:1
- **Harder Example**
- **Percent Yield: (680/767.1) x 100= 88.6%**

Slide 119: You Try!

- Q - What is the % yield of NH₃ if 40.5 g NH₃ is produced from 20.0 mol H₂ and excess N₂?
- Step 1: Balance the chemical equation
- N₂ + H₂ → NH₃
- Step 2: Determine actual and theoretical yield. Actual is given, theoretical is calculated:

- Step 3: Calculate the % yield

Slide 120: Check your work

- Q - What is the % yield of NH_3 if 40.5 g NH_3 is produced from 20.0 mol H_2 and excess N_2 ?
- Step 1: Write the balanced chemical equation
- $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
- Step 2: Determine actual and theoretical yield. Actual is given, theoretical is calculated:
- Step 3: Calculate the % yield
- Molar ratio 2:3
- % yield =
- $\times 100\%$
- $\times 100\%$

Slide 121: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

What is the % yield of H_2O if 58 g H_2O are produced by combining 60 g O_2 and 7.0 g H_2 ?

Hint: determine limiting reactant first (the reactant where there isn't enough)

- Challenging question
- Answer = 92.9%

Slide 122: More Percent Yield Questions

- The electrolysis of water forms H_2 and O_2 . $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
- What is the % yield of O_2 if 12.3 g of O_2 is produced from the decomposition of 14.0 g H_2O ?

- 107 g of oxygen is produced by heating 300 grams of potassium chlorate. Calculate % yield.

$$2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$$
- What is the % yield of ferrous sulphide if 3.00 moles of Fe reacts with excess sulfur to produce 220 grams of ferrous sulphide?
- $\text{Fe} + \text{S} \rightarrow \text{FeS}$

Slide 123: Check your work

- The electrolysis of water forms H_2 and O_2 . $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
- What is the % yield of O_2 if 12.3 g of O_2 is produced from the decomposition of 14.0 g H_2O ?
- **Answer = 98.9%**
- 107 g of oxygen is produced by heating 300 grams of potassium chlorate. Calculate % yield.

$$2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2 \quad \mathbf{91.1\%}$$
- What is the % yield of ferrous sulphide if 3.00 moles of Fe reacts with excess sulfur to produce 220 grams of ferrous sulphide? **83.4%**
- $\text{Fe} + \text{S} \rightarrow \text{FeS}$

Slide 124: Iron pyrites (FeS_2) reacts with oxygen according to the following equation:



If 300 g of iron pyrites is burned in 200 g of O_2 , 143 grams of ferric oxide is produced. What is the percent yield of ferric oxide?

70 grams of manganese dioxide is mixed with 3.5 moles of hydrochloric acid. How many grams of Cl_2 will be produced from this reaction if the % yield for the process is 42%?



- **More Percent Yield Questions**

Slide 125: Iron pyrites (FeS₂) reacts with oxygen according to the following equation:



If 300 g of iron pyrites is burned in 200 g of O₂, 143 grams of ferric oxide is produced. What is the percent yield of ferric oxide? 78.8%

70 grams of manganese dioxide is mixed with 3.5 moles of hydrochloric acid. How many grams of Cl₂ will be produced from this reaction if the % yield for the process is 42%? 24g



- Check your work

Slide 126: Exam question

Slide 127: Mark scheme

Slide 128: Exam question

Slide 129: Mark scheme

Slide 130: Percent Yield

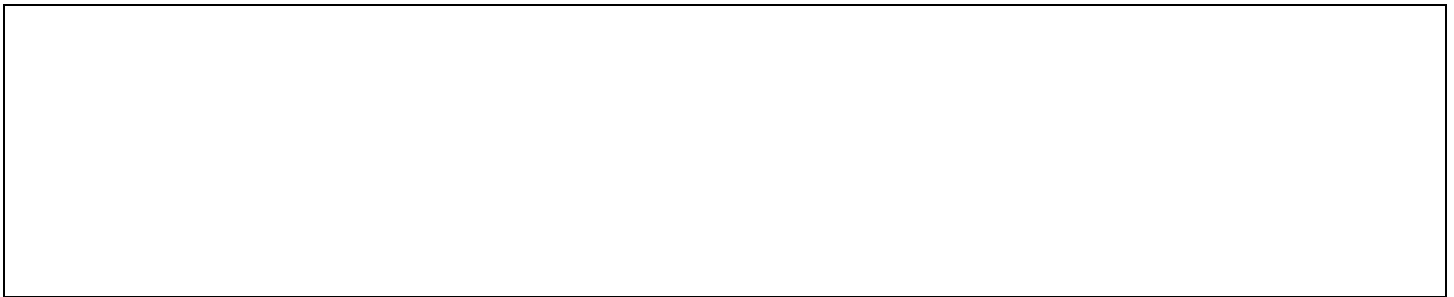
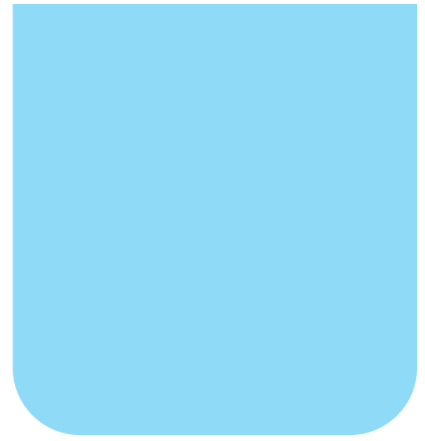
Slide 131: Limiting Reactants and Theoretical Yield

- The heat produced from an acetylene torch is produced by the burning of acetylene (C_2H_2) in oxygen.
- $2C_2H_2 + 5O_2 \rightarrow$
- How many **grams of water** can be produced by the reaction of **2.40 mol C_2H_2** and **7.40 mol O_2** ?
- **Acetylene**
- (*Systematic Name: Ethyne*)
- $4CO_2 + 2H_2O$
- $2C_2H_2 + 5O_2$



Slide 132: Theoretical and Actual Yield

- The answer calculated in the previous problem is called the *theoretical yield*.
- The **theoretical yield** is the *maximum* amount of product that could be formed from given amounts of reactants.
- In contrast, the amount of product that *actually* forms when the reaction is carried out in the laboratory is called the **actual yield**.
- Theoretical Yield
- Actual Yield



Slide 133: Percent Yield

- The **percent yield** is a measure of the efficiency of a reaction. It is a measure that indicates how successful a reaction has been.
- The **percent yield** is the ratio of the actual yield to the theoretical yield expressed as a percent.
- If actual yield of water in the first problem was **25.0 g**, what is the **percent yield** of the reaction?
- actual yield
- theoretical yield



Slide 134: Practice Problem

- **Calculating the Theoretical Yield of a Reaction**
- Calcium carbonate, which is found in seashells, is decomposed by heating. The balanced equation for this reaction is:
- What is the *theoretical yield* of CaO if **24.8 g CaCO₃** is heated?
- $\text{CaCO}_{3(s)} \rightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$

Slide 135: Practice Problems

- **Calculating the Percent Yield of a Reaction**
- What is the *percent yield* if **13.1 g CaO** is *actually* produced when **24.8 g CaCO₃** is heated?
- $\text{CaCO}_{3(s)} \rightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$

Slide 136: Atom Economy

- Even if a reaction has a 100% yield, it may not be _____ in another sense – its **atom economy**.
- Many _____ produce by-products that are of little or no use. The atom economy is a measure of the *mass of useful products* compared to the *mass of reactants*.
- _____ mass of useful products
- _____ mass of reactants
- Mass of Reactants
- Mass of Useful _____

Slide 137: Atom Economy

- Mass of Reactants
- Mass of Useful _____

- Useful Product
- Reactants
- 72.92

214.96

72.92

Slide 138: Atom Economy

- Useful Product
- Reactants
- 72.92
- x 100

214.96

72.92

Slide 139: Atom Economy

- The production of _____ acid can be achieved by the following reaction:
- Its atom _____ would be:
- (58.44 g mol^{-1})
- 2
- +
- (98.08 g mol^{-1})
- 2
- (36.46 g mol^{-1})
- $\text{NaCl}_{(s)}$
- 2

- $\text{H}_2\text{SO}_4(l)$
- $\text{HCl}(g)$
- **2**
- **MM**_{useful product(s)}
- **MM**_{reactants}

Slide 140: Atom Economy Practice

- Calculate the atom economy for the following reaction, assuming hydrochloric acid is the useful product:

Slide 141: Atom Economy Practice

- If we wanted to produce sodium chloride, which reaction below would have be most efficient in terms of their atom economy?

Slide 142: Atom Economy Practice

- Iron (Fe) can be extracted from its ore in the following reaction.
- Calculate the atom economy of the reaction.

Slide 143: Gas Laws

Slide 144: Avogadro's Law

- enables the mole ratio of reacting gases to be _____ from volumes of the gases.
- equal amounts of moles of gases occupy the same volume under the same conditions of _____ and pressure.
- volumes have equal moles of separate particles (molecules or _____ atoms) in them.
- Equal volumes of gases under the same conditions of _____ and pressure contain the same number of molecules.

Slide 145: Calculating volumes of gases

Slide 146: Calculating volumes of gases

- Conditions of STP (standard _____ and pressure) are a temperature of 273 K and a pressure of 100 kPa.
- This is in your data _____ on section 1.

Slide 148: You Try!

- How many moles of gas are in:
 - 36 dm³ of CO₂
 - 10000dm³ of Hydrogen H₂
- What volume of gas would these occupy?
- 36 g of helium (He)
- 138 g of nitrogen dioxide (NO₂)
- What mass of gas is present in 48 cm³ of oxygen gas (O₂)? Units!

Slide 149: Extra Questions

Slide 150: For a fixed mass of gas at a constant temperature:

$PV = \text{constant (k)}$

When a gas changes pressure from P_1 to P_2 while undergoing a volume change from V_1 to V_2 :

$$P_1 \times V_1 = P_2 \times V_2$$

- **Boyle's law**
- **Robert Boyle** was an Anglo-Irish natural _____, chemist, physicist, and inventor.

Slide 151: How could you use marshmallows to test Boyle's law?

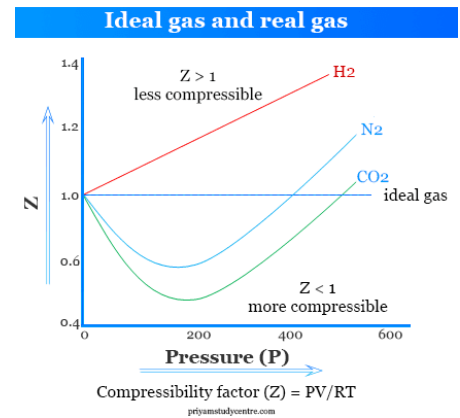
- Did you know that marshmallows originated in Egypt 4,000 years ago?
- Today, they are made by only three companies worldwide.
- Useless facts:
- <https://www.youtube.com/watch?v=1I93Y2AH0mw>
- **Boyle's law demonstration**

Slide 152: A gas has an initial volume of 300 m³ at standard atmospheric pressure (100 kPa). Calculate the final volume of this gas if its pressure is increased by 400 kPa at a constant temperature.

- Boyle's law: $P_1 \times V_1 = P_2 \times V_2$
- $100 \text{ kPa} \times 300 \text{ m}^3 = 500 \text{ kPa} \times V_2$
- $30\,000 = 500 V_2$
- $V_2 = 30\,000 / 500$
- **Final volume = 60 m³**
- **Boyle's law question**

Slide 153: Ideal vs. Real Gases

- Ideal Gas:
- Follow Boyle's Law at all _____
- Plot of PV/RT vs. P at constant _____ produces horizontal line
- No _____ attractions (can't be liquified)
- No Volume
- Real Gas:
- Follow Boyle's Law at high _____
- Plot of PV/RT vs. P at constant _____ produces Amagat curve
- _____ attractions (can be liquified)
- Have volume
- A gas touching the _____ line behaves like an ideal gas at that pressure
- As pressure increases, heavier gases are more _____ (due to IMF)
- As pressure increases, lighter gases are less _____ (due to weaker IMF & higher KE)



Slide 154: All of these introductory gas laws assume gases are IDEAL!!!

- **Gas Laws Assume Ideal Gases**

Slide 155: For a fixed mass of gas at a constant pressure:

$V = \text{constant (k)}$

T

When a gas changes volume from V_1 to V_2 while undergoing a temperature change from T_1 to T_2 :

$V_1 = V_2$

$T_1 \quad T_2$

- Graph of an ideal gas _____ Charles' law. The gas volume becomes zero at 0K.
- **Charles's law**
- **Jacques Alexandre César Charles** was a French inventor, scientist, _____, and balloonist.



Slide 156: A gas has an initial volume of 50m³ at a temperature of 127oC. Calculate the final temperature required in oC to decrease its volume to 20m³ at a constant pressure.

- Charles' law: $V_1 / T_1 = V_2 / T_2$
- Temperatures must be in kelvin, so: $T_1 = 400\text{K}$
- $50\text{m}^3 / 400\text{K} = 20\text{m}^3 / T_2$
- $T_2 = (20 \times 400) / 50$
- $T_2 = 160\text{K}$
- **Final temperature = - 113 °C**
- **Charles's law question**

Slide 157: Gases at 0 °C still have pressure and volume.

Absolute zero (in Kelvin) is the only value that assumes no molecular motion.

Using negative temperatures would mathematically predict negative pressures/volumes. This is IMPOSSIBLE!

- **Why Can't We Use Celsius?**

Slide 158: For a fixed mass of gas at a constant volume:

$P = \text{constant (k)}$

T

When a gas changes pressure from P_1 to P_2 while undergoing a temperature change from T_1 to T_2 :

$$P_1 = P_2$$

$$T_1 \quad T_2$$

- **Gay-Lussac's law**
- Joseph Louis **Gay-Lussac** was a French chemist and

_____.

Slide 159: A gas has an initial pressure of 100kPa at a temperature of 27°C. Calculate the final pressure of this gas if its temperature is increased by 300°C at a constant volume

- Pressure law: $P_1 / T_1 = P_2 / T_2$
- Temperatures must be in kelvin
- so: $T_1 = 300 \text{ K}$ and $T_2 = 600 \text{ K}$
- $100 \text{ kPa} / 300 \text{ K} = P_2 / 600 \text{ K}$
- $P_2 = (100 \times 600) / 300$
- **Final pressure = 200 kPa**
- **Gay-Lussac's law question**

Slide 160: Combining some of the laws we have just seen:

$$P \times V = \text{constant (k)}$$
$$T$$

Therefore:

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

- **Combined gas law**

Slide 161: Helpful video - how to use gas laws (maths)

Slide 162: $PV / T = \text{a constant (k)}$

can be written: $P_1V_1 / T_1 = P_2V_2 / T_2$

temperatures must be in kelvin, so: $T_1 = 293\text{K}$

$$(101000 \times 1) / 293 = (303 \times 6) / T_2$$

$$T_2 = (293 \times 303000 \times 6) / (101000 \times 1)$$

final temperature = 5274 K

- A fixed mass of gas has its pressure increased from 101 kPa to 303 kPa, its volume by 5 m³ from 1 m³ while its temperature is raised from 20°C.
- Calculate its final temperature.
- **Question 2**

Slide 163: Combining all three gas laws for a constant mass of gas gives:

$$PV = \text{a constant (k)}$$

T

the constant = nR and so:

$$PV = nRT \text{ – the ideal gas equation}$$

where:

n = number of moles of the gas

R = molar gas constant = 8.31 J K⁻¹ mol⁻¹

Unit T = K, P = Pa, V = m³

- What is an ideal gas?
- One that obeys Boyle's law
- **The ideal gas equation**
- This gas constant can have different values depending on the units of pressure and volume
- 1 J = 1 Pa m³

Slide 165: Calculate the volume of one mole an ideal gas at 0oC and 100 kPa (standard temperature and pressure)

- $PV = nRT$
- becomes: $V = nRT / P$
- temperatures must be in Kelvin, so: $T = 273\text{K}$
- = (1 mol x 8.31 J K⁻¹ mol⁻¹ x 273K) / 100 000 Pa
- = 0.0227 m³
- **volume = 22.7 dm³ (cubic decimetres OR litres)**
- This is also known as 'molar volume'.
- What is special about this volume?
- **Question 1**

Slide 166: Determining the Molar Mass of a Gas from Experimental Data

The ideal gas equation isn't just theoretical; it's a practical tool. Given the mass of a gas and its volume, pressure, and temperature, one can ascertain its molar mass.

Steps:

1. Utilise the ideal gas equation to determine the number of moles (n).
2. Calculate molar mass: Molar mass = given mass / n

Molar Mass (M) = m / n
 M = Molar mass of the gas (g mol^{-1})
 m = Mass of the gas (g)
 n = Amount of substance (mol)

Slide 167: Determining the Molar Mass of a Gas from Experimental Data

- $PV = nRT$
- $n = PV/RT$
- = $(101,000 \text{ Pa } (0.0056 \text{ m}^3) / 8.31 \text{ J mol}^{-1} \text{ K}^{-1} (273 \text{ K}))$
- = 0.25 mol
- $10\text{g}/0.25 \text{ mol} = 40. \text{ g mol}^{-1}$

Example: Suppose you have 10g of an unknown gas at 1 atm and 273 K, occupying 5.6 L. Using the Ideal Gas Equation, you can deduce the number of moles and consequently the gas's molar mass.

Slide 168: Assume that 1 mole of any gas occupies 22.7 dm³ at STP

- 2. Moles of $\text{CaC}_2 = 0.160 / 64 = 0.0025$ (1)
- Moles of $\text{C}_2\text{H}_2 = 0.00250$ (1:1) (1)
- Volume = $0.0025 \times 22.7 = 0.0568 \text{ dm}^3$ (1)
- Calculate the volume of hydrogen produced at room temperature and pressure by 3.25 g of zinc metal when reacted with excess acid.
- $\text{Zn}_{(s)} + \text{H}_2\text{SO}_{4(aq)} \rightarrow \text{ZnSO}_{4(aq)} + \text{H}_{2(g)}$
- 0.160 g of calcium carbide (CaC_2) was treated with excess water. Calculate the volume of ethyne (C_2H_2) gas produced at room temperature and pressure.
- $\text{CaC}_{2(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow \text{C}_2\text{H}_{2(g)} + \text{Ca}(\text{OH})_{2(s)}$
- 1. Moles of Zn = $3.25/65.0 = 0.050$ (1)
- Moles of $\text{H}_2 = 0.050$ (1:1) (1)
- Volume = $0.050 \times 22.7 = 1.14 \text{ dm}^3$ (1)
- **Further questions**

Slide 169: ideal gas constant R = 8.31 J K⁻¹ mol⁻¹

- 0.688 g of a hydrocarbon occupies 472 cm³ at 98 kPa and 67 °C.
- Calculate the:
 - (i) mol of hydrocarbon in 0.688 g
 - (ii) relative molecular mass of the hydrocarbon.
- a) (i) $PV=nRT$ (1)
- $n = 98000\text{Pa} \times 4.72 \times 10^{-4} \text{ m}^3$ (1)
- $8.31\text{JK}^{-1} \text{ mol}^{-1} \times 340\text{K}$ (1)
- =0.0164 moles (1)
- (ii) $0.688\text{g}/0.0164\text{moles} = 42.03$ (2)
- b) (i) Suggest the molecular formula and name of the hydrocarbon
- (ii) Give an equation for the complete combustion of the hydrocarbon
- (iii) Calculate the volume of oxygen needed to completely react with
- 0.21g of the hydrocarbon under the same conditions.
- **Further questions**

Slide 170: b) (i) C₃H₆ (1)

Propene (1)

(ii) $2\text{C}_3\text{H}_6 + 9\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$ (1)

(ii) $0.21 / (42.03) = 0.00500 \text{ mol propene}$ (1)

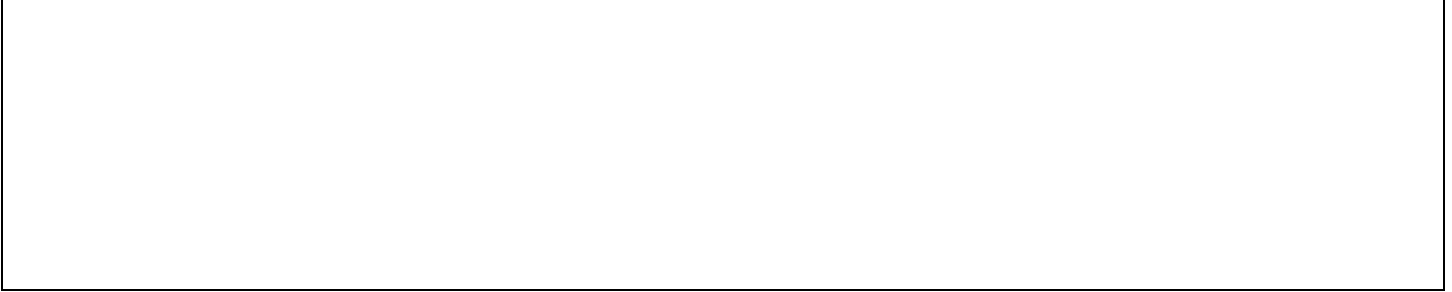
2 : 9 ratio (1)

$0.00500 \text{ mol} \times 2/9 = 0.0225 \text{ mol O}_2 \text{ needed}$ (1)

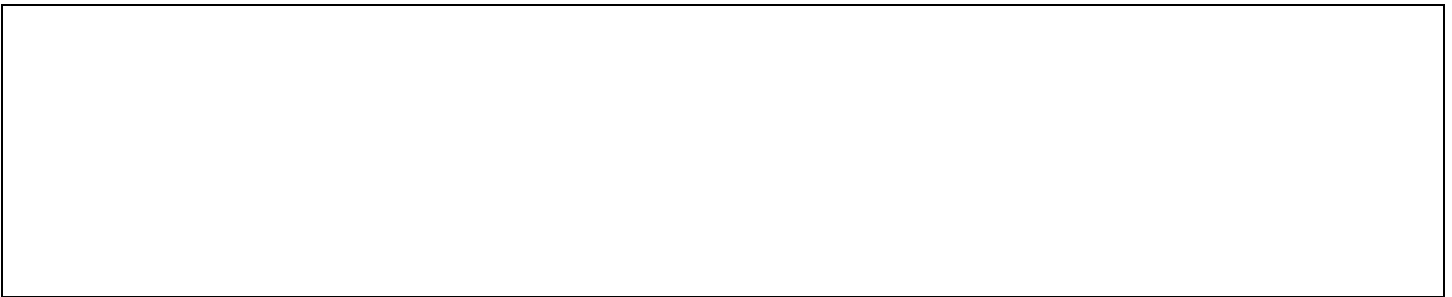
$0.0225 \times 22.7 = 0.51\text{dm}^3 \text{ oxygen gas needed}$ (1)

- b) (i) Suggest the molecular formula and name of the hydrocarbon
- (ii) Give an equation for the complete combustion of the hydrocarbon
- (iii) Calculate the volume of oxygen needed to completely react with
- 0.21g of the hydrocarbon under the same conditions.
- **Further questions**

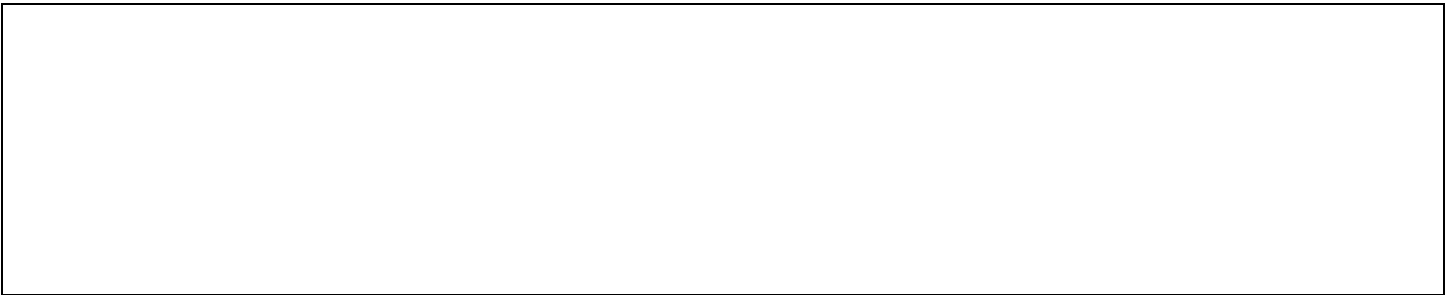
Slide 171: Further questions



Slide 172: Exam question – part a



Slide 173: Exam question – part b



Slide 174: Mark scheme

Slide 175: Mark scheme