

AS Level calculations involving Reacting Liquids

(A First Year Blog)

- Last week I showed you a number of calculations that involved solids. Now let me take you through some calculations involving liquids (and mainly *aqueous solutions* i.e. substances that have been dissolved in water).
- *I do not know where the questions that I have used originally came from (but the answers are mine) – and there are two excellent textbooks for calculations which I would strongly recommend to you viz **the I.L.P.A.C. series on Advanced Chemistry Calculations** i.e. ISBN 0-7195-7506-0, and also Jim Clark's book ISBN 0-582-41127-0.]*
- I should like to start by reminding you of some of the things that you will have learnt at GCSE Level (but it is just possible that some of you may be hearing them for the first time) – because by going over things that you are already supposed to know, I can make sure
 - (a) that you **do** in fact know the things that you are supposed to know, and
 - (b) that you have not forgotten them.
- I will therefore do two things in this Chapter. I will
 - talk about Titration and the rules that will help you to avoid mistakes, and
 - talk about some calculations with which you are already supposed to be familiar.

Please remember that 1 metre = 10 decimetres (dm) = 100 centimetres (cm) = 1,000 millimetres (mm), and (by definition)

- **1 litre** = 1dm x 1dm x 1dm = **1 cubic decimetre (1 dm³)**
and since 10 centimetres = 1 decimetre, then 1 litre also
= 10cm x 10cm x 10cm = **1,000 cm³ (1,000 cubic centimetres)**
= 1,000 millilitres (**1,000 ml**)
ALL these units represent EXACTLY the same volume (and 1 litre has a mass of 1kg).

$$\mathbf{1\ litre = 1\ dm^3 = 1,000\ cm^3 = 1,000\ ml.}$$

- If I had 5 goldfish swimming in 1 dm³ of Water, then the Concentration of goldfish in the Water would be 5 goldfish per cubic decimetre of Water or “5 goldfish dm⁻³”.
- Equally, if I had 25 goldfish swimming in 4.7 dm³ of Water, then the Concentration of goldfish in the Water = $\frac{25\ \text{goldfish}}{4.7\ \text{dm}^3\ \text{of water}} \approx 5.3\ \text{goldfish dm}^{-3}$.
- We can therefore say that

$$\text{Concentration} = \frac{\text{Number or Amount of things}}{\text{Volume of liquid in dm}^3}$$

$$C = \frac{N}{V\ \text{in dm}^3}, \quad \text{and} \quad N = C \times V\ \text{in dm}^3$$

NB The unit of concentration in Chemistry is “dm³”, therefore if you are given a volume in cm³ you **MUST convert it into dm³ by dividing the cm³ number by 1000**
e.g. 17.3 cm³ = $\frac{17.3}{1000}\ \text{dm}^3 = 0.0173\ \text{dm}^3$

- I would therefore strongly recommend that whenever you write out the formula for Concentration, you should always write "**V in dm³**" i.e. you should always write

$$C = \frac{N}{V \text{ in dm}^3}, \quad \text{therefore } N = C \times V \text{ in dm}^3$$

because by writing "**V in dm³**" you will always be reminded to convert numbers in cubic centimetres into their cubic decimetre equivalents.

Concentration (in the olden days Concentration was defined in terms of "Molarity")

- The Concentration of anything is always expressed as the number of moles of that substance for every cubic decimetre of solution or liquid, and the unit of Concentration is thus **moles per cubic decimetre** or **mol dm⁻³**.
- In the olden days, the Concentration of something was called its "Molarity", and it was written as "M", and it was easy to confuse the M for Mass with the M for Molarity, therefore Chemists in Europe stopped using the term "Molarity". [However, the term "Molarity" is still used occasionally, and if you ever read that the molarity of something is such and such, then you are being told the value of its Concentration.]

C or N or V questions

- In the exam you could be given any two of the three variables (C, N and V) and then asked to find the third one e.g.

Q) How many moles of Sulphuric acid are there in 17.3 cm³ of Sulphuric acid that has a strength or concentration of 0.050 mol dm⁻³?

$$\text{A) } C = \frac{N}{V \text{ in dm}^3} \quad \text{therefore } N = C \times V \text{ in dm}^3$$

Therefore Number of moles, $N = 0.050 \text{ mol dm}^{-3} \times 0.0173 \text{ dm}^3 = 0.000865 \text{ mol} = 8.656 \times 10^{-4} \text{ mol}$.
 "dm⁻³" above the line is equivalent to "dm³" below the line, therefore
 "dm⁻³" cancelled out "dm³", and **0.000865 mol ≈ 8.7 × 10⁻⁴ mol.**

Q) What is the concentration of 14.9 moles of a substance in 16.6 cm³ of liquid?

$$\text{A) } C = \frac{N}{V \text{ in dm}^3} \quad \text{therefore } C = \frac{14.9 \text{ mol}}{0.0166 \text{ dm}^3} \approx 897.6 \text{ mol dm}^{-3} \\ \approx 8.98 \times 10^2 \text{ mol dm}^{-3}$$

Q) What volume of liquid would there be if a solution of concentration 0.005 mol dm⁻³ had 1.739 × 10⁻³ moles in it?

$$\text{A) } V \text{ in dm}^3 = \frac{N}{C}, \quad \text{therefore } V = \frac{1.739 \times 10^{-3} \text{ mol}}{0.005 \text{ mol dm}^{-3}} \approx 0.348 \text{ dm}^3$$

- I hope that you can see that the "mol" above the line and the "mol" below the line have cancelled each other out, and that because the "dm⁻³" is going from below the line to above the line, it changes from "dm⁻³" to "dm³" (which is the same as "dm⁺³").
- OK, those are the basic GCSE calculations involving C, N, and V, so let us now do some slightly more difficult stuff.
- One of the most fundamental reactions in Chemistry is the neutralisation of an Acid by a Base – and (hardly surprisingly) this is called an Acid-Base Reaction. You may know it as
Acid + Base → Salt + Water
 and the neutralisation of an acid by a base is one of the most simple of all titrations.

- However, there are many other titrations that can be executed (e.g. the titration of a Thiosulphate solution against a solution of Iodine, and so on), but in all the titrations that you will do, what you will be trying to do is **to find out the EXACT point at which the full extent of the reaction (that is involved) has *just* taken place (this being called the “equivalence point”)**. (NB The “**end-point**” of the reaction is where the indicator that has been used changes colour. The ‘end-point’ indicates the reaching of the ‘equivalence point’.)
- As we saw in “solid” calculations, one of the IMMUTABLE rules of Chemistry is that **you cannot do ANY calculation involving a reaction without first knowing what the REACTION EQUATION** is therefore, whenever a reaction equation is involved, always start by writing down the reaction equation (and please make sure that your reaction equations are correctly balanced). In “solid” calculations the reaction equation is very often given, whereas in “liquid” calculations this is **not** the case.
- If you have forgotten how to balance an equation, then please go back to Chapter 27 and see how it is done; and, if you have forgotten your state (or phase) symbols then here they are again

(aq)	=	aquated / aqueous / dissolved in water
(s)	=	solid
(l)	=	liquid
(g)	=	gaseous

The Procedure in Titration

- “**Titration**” is the process by which an **unknown characteristic** of a solution (e.g. its Concentration) can be ascertained by “**titrating**” that solution against (or dripping small amounts of that solution into) a solution whose characteristics ARE known until the exact point is reached where one solution has just COMPLETELY reacted with the other solution.
- The following is the procedure for doing Titrations. Follow it, and you will never make a mistake in Titration.
 - 1) Write down the Equation for the Reaction of A with B.
 - 2) Derive the **MOLE Reaction Ratios** from the Reaction Equation.
 - 3) Wash out your conical flask carefully¹ with some of the solution whose characteristics are known (let’s call this solution A), and then place a given volume (e.g. 25 cm³) of solution A in the conical flask that you will place under your burette.
 - 4) Add three drops of the chosen **Indicator** to solution A in the conical flask.
 - 5) Wash out your burette carefully with some of solution B,² and then almost fill your burette with solution B. **DO NOT FILL THE BURETTE ABOVE THE ZERO MARK!**³ If you do so, then the invigilator in your Assessed Practical (i.e. your official State Practical exam) **will almost certainly have to fail you!**
 - 6) Now take the reading of the burette (this being your “**Initial Reading**”).
 - 7) Then titrate B against A until you obtain the exact point at which the reaction has fully taken place (this being indicated by the change in colour of the Indicator which you added to the conical flask). The last few drops of solution B **must** be dropped very slowly one drop at a time into the conical flask, while the flask is being shaken gently.
 - 8) Take the reading of the burette again (this now being your “**Final Reading**”), and the difference between the initial reading and the final reading is the exact amount of solution B that was needed to react with the 25 (or whatever) cm³ of solution A.
 - 9) Repeat steps (6) and (7) until you obtain two readings that are within 0.2 cm³ of each other. [You should be able to achieve this within three trials i.e. one “ranging” shot, and two accurate ones, and these latter two trials should give answers that are within 0.2 cm³ of each other – and preferably within 0.1 cm³ of each other.]
 - 10) From the known Concentration of A work out how many **moles of A** there are in the volume of solution A (i.e. in the 25 cm³).
 - 11) From the Reaction Ratio, work out how many **moles of solution B** there must have been in the average volume of B, established by the Titration.

¹ This reduces the risk of the flask being contaminated.

² This reduces the risk of the burette being contaminated.

³ If the liquid rises above the zero mark, then you cannot calculate how much will have been used in the titration.

- 12) You now know (a) the volume of B and (b) the number of moles of B, therefore you can now calculate the **Concentration of B** because $C = N \div V$ where

$$\begin{aligned} \text{Concentration} &= \frac{\text{The AMOUNT of something, } N \text{ (expressed in moles)}}{\text{The VOLUME involved, } V \text{ (expressed in dm}^3\text{)}} \\ &= X \text{ mol dm}^{-3} \quad [X \text{ mol dm}^{-3} = X \text{ moles per dm}^3] \end{aligned}$$

Burette and Conical flask

- In a Titration, in terms of the Mathematics involved, it really does not matter which substance is put in the conical flask and which is put in the burette – and, as far as the indicator is concerned, some people find it easier to spot a change from a colourless solution to the very first faint appearance of colour, and others find it easier to spot the disappearance of colour.
- The choice of what to titrate into what is thus very much a personal one – but in practical terms, it is very difficult to wash out some substances from a burette, therefore you should always put highly coloured or highly viscous substances in the conical flask and **NOT in the burette**.⁴

A) Calculations about Concentrations

- For example,
 - if you have a 0.50 mol dm³ solution of HCl, then this means that in 1 dm³ (or 1,000 cm³) of solution there are 0.50 moles of HCl; or,
 - if you have a 1.50 mol dm³ solution of H₂SO₄, then this means that in 1 dm³ (or 1,000 cm³) of solution there are 1.50 moles of H₂SO₄; or,
 - if you have a 6.00 mol dm³ solution of H₃PO₄, then this means that in 1 dm³ (or 1,000 cm³) of solution there are 6.00 moles of H₃PO₄.
- I have used acids in the above examples, but I could equally have said
 - if you have a 0.50 mol dm³ solution of NaOH, then this means that in 1 dm³ (or 1,000 cm³) of solution there are 0.50 moles of NaOH; or,
 - if you have a 1.50 mol dm³ solution of Mg(OH)₂, then this means that in 1 dm³ (or 1,000 cm³) of solution there are 1.50 moles of Mg(OH)₂.

Acid/Base Reactions

- If we look at the reaction $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NaCl(aq)}$, the reaction ratio tells us that at the point of exact neutralisation, 1 mole of HCl(aq) will have reacted with exactly 1 mole of NaOH(aq) to produce 1 mole of H₂O(l), and 1 mole of NaCl(aq).
- Therefore if we had calculated that we had (say) 0.47 moles of HCl(aq), then we would know with absolute certainty that at the point of exact neutralisation⁵ 0.47 moles of NaOH(aq) will have been used up (or neutralised). **It is the ratios that are important!** All the following ratios are equally valid.

HCl	NaOH	→	H₂O	NaCl
1 mole	1 mole		1 mole	1 mole
987654321 moles	987654321 moles		987654321 moles	987654321 moles
1.23456789 moles	1.23456789 moles		1.23456789 moles	1.23456789 moles
X moles	X moles		X moles	X moles

⁴ Throughout your life, always try to be aware of the things that other people have to do to clear up after you. In life, irrespective of the accident of one's birth, the true aristocrat/the noble person is the person who is always considerate of other people.

⁵ The point at which one solution exactly reacts with the other is called the "equivalence point".

Differing volumes of Acids and Bases

- If you put into a conical flask a given **Volume** (e.g. 25 cm³) of an Acid (or a Base) whose **Concentration** is known, then you can calculate exactly how many moles of the substance are in the conical flask.
- For example, if you put 25.0 cm³ of 0.500 mol dm⁻³ solution of NaOH into a conical flask, then the number of moles of NaOH in the conical flask can be calculated by

$$\text{Concentration, } C = \frac{\text{The Amount of something, } N \text{ (expressed in moles)}}{\text{The volume involved, } V \text{ (expressed in dm}^3)} = \frac{N \text{ moles}}{V \text{ in dm}^3}$$

$$N = C \times V \text{ in dm}^3, \text{ therefore } N \text{ (number of moles)} = 0.500 \text{ mol dm}^{-3} \times \left(\frac{25.0}{1000} \right) \text{ dm}^3 \\ = 0.0125 \text{ moles of NaOH}$$

NB1) Volume must always be expressed in dm³ therefore 25.0 cm³ **has** to be converted into dm³ by dividing it by 1,000.

NB2) The answer is in **moles** because dm⁻³ cancelled out dm³ to leave just **moles**.

- OK, so now we know that there are **0.0125 moles of NaOH(aq)** in the conical flask, and if (let us say) that it takes exactly 17.3 cm³ of HCl to neutralise the NaOH, then we can work out the Concentration of HCl because
 - 1) from the reaction equation **HCl(aq) + NaOH(aq) → H₂O(l) + NaCl(aq)** the mole reaction ratio of HCl to NaOH is 1 : 1, therefore there must have been 0.0125 moles of HCl in the 17.3 cm³ of HCl, and
 - 2) since $C = \frac{N \text{ mol}}{V \text{ in dm}^3}$, then the Concentration of the acid solution is given by
$$C = \frac{0.0125 \text{ mol}}{0.0173 \text{ dm}^3},$$
and $C \approx 0.723 \text{ mol dm}^{-3}$.
- As you can see, **we have been able to work out the unknown Concentration of the Acid by titrating the Acid against a known volume (i.e. 25 cm³) of an Alkali that has a known Concentration.**

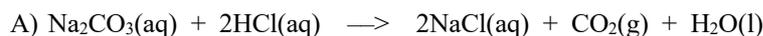
Other reactions besides Acid/Base reactions

- We will learn that there are many other types of reactions besides Acid/Base reactions; and, while I have presented Titration so far only in terms of Acid/Base reactions, nevertheless, Titration can be applied to the Volumetric analysis of **any** two solutions that react together (e.g. the titration of a Thiosulphate solution against a solution of Iodine, and so on).

Let us do another example.

- **Question : If 25.0 cm³ of 0.100 mol dm⁻³ Na₂CO₃(aq) react with 35.0 cm³ of HCl, what is the Concentration of the Acid?**
- Always
 - (A) Start with the **reaction equation**.
 - (B) Work out the **MOLE Reaction Ratio**.
 - (C) Work out the **number of moles used in the known solution**.
 - (D) Use the reaction ratio to work out the **number of moles of the unknown solution**.
 - (E) Use the **volume** of the unknown solution and the **number of moles** to work out its concentration.

You MUST start with a balanced reaction equation. Write it out!
(Hint: what happens when you react a Carbonate with an Acid?)



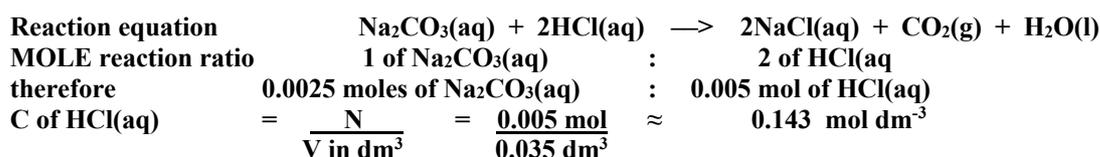
B) Reaction Ratio = 1 of $\text{Na}_2\text{CO}_3(\text{aq})$: 2 of $\text{HCl}(\text{aq})$

C) No. of moles of $\text{Na}_2\text{CO}_3(\text{aq})$ = Concentration (in mol dm^{-3}) x Volume (in dm^3)
 = $0.100 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3$
 = $0.0025 \text{ mol of Na}_2\text{CO}_3(\text{aq})$

D) Reaction Ratio = 1 : 2 , \therefore No of moles of HCl = $0.0025 \text{ mol} \times 2 = 0.005 \text{ mol of HCl}$

E) There are 0.005 moles of HCl in 35.0 cm^3 of solution, therefore
 Concentration = $N (\text{in mol}) \div V (\text{in dm}^3)$, $\therefore C = 0.005 \text{ mol} \div 0.035 \text{ dm}^3$
 $\approx 0.14286 \text{ mol dm}^{-3}$
 = $0.143 \text{ mol dm}^{-3}$ (to 3 significant figures).

- Let me now show you a way of setting out the whole calculation – and I would advise you to ALWAYS set out your liquid calculations in this manner.



- That's it! You can do the whole calculation in FOUR lines, and in your exams that is exactly what you should do.
- I have chosen this example deliberately because the reaction equation says “2HCl” rather than “1HCl” as in the equation” $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l}) + \text{NaCl}(\text{aq})$ ”.
- The numbers in front of any of the substances (i.e. 1/2/3/4/etc) will affect your MOLE calculations, **but you will NOT need them after that!**

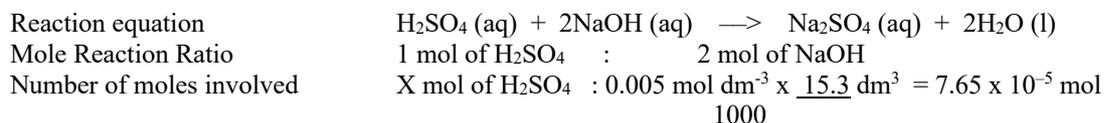
**Once you have calculated the MOLE Reaction Ratio,
 the number in front of the substance
 does not come into the calculation again!**

- OK, I hope that you now see that the first thing that you need in any Titration is the Reaction Equation because it is from THIS that we derive the Reaction Ratio!**
- I shall do one more example for you, and this time I will do an Acid/Base reaction with a Diprotic Acid⁶ (just in case you are asked to do so in the exam). A diprotic acid dissociates TWO hydrogen ions per molecule of acid, and we will talk about this in detail when we get to Thermodynamics or as it is also called “How Far, How Fast”.]

⁶ “Diprotic” means that it disassociates two protons. In the olden days a diprotic acid used to be called a “dibasic” acid.

Q) What volume of Sulphuric Acid (H₂SO₄) would it require to neutralise 15.3 cm³ of Sodium Hydroxide (NaOH) if the NaOH had a strength or concentration of 0.005 mol dm⁻³ and the H₂SO₄ had a strength of 0.020 mol dm⁻³?

- The two things that you must do are
 - 1 write out the balanced reaction equation and**
 - 2 calculate the number of moles for the substance where you have been given both the concentration and the volume.**



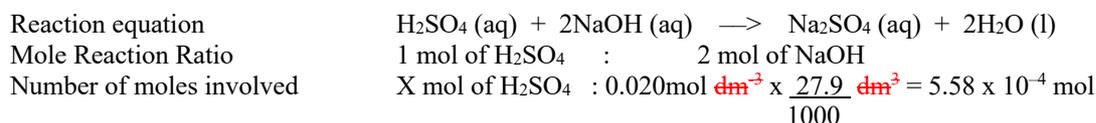
Where $X = \frac{1 \text{ mol} \times 7.65 \times 10^{-5} \text{ mol}}{2 \text{ mol}} = 3.825 \times 10^{-5} \text{ mol of H}_2\text{SO}_4 \text{ (aq)}$

∴ Volume of H₂SO₄ = V in dm³ = $\frac{N}{C} = \frac{3.825 \times 10^{-5} \text{ mol}}{0.020 \text{ mol dm}^{-3}} \approx 1.91 \times 10^{-3} \text{ dm}^3$

- Or, they might ask you

Q) What concentration of Sulphuric Acid (H₂SO₄) would it require to neutralise 27.9 cm³ of Sodium Hydroxide (NaOH) if the NaOH had a strength or concentration of 0.020 mol dm⁻³ and 36.9 cm³ of the H₂SO₄ were used?

I'll say it again because it so important : **you cannot do ANY calculation involving a reaction without first knowing what the reaction equation is, therefore start with the reaction equation**



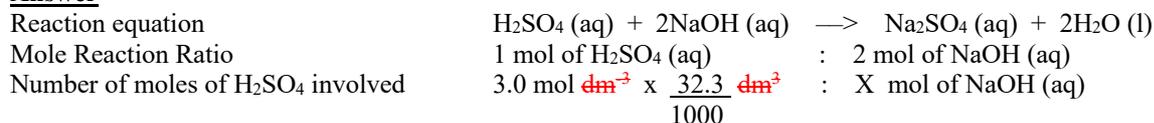
Where $X = \frac{1 \text{ mol} \times 5.58 \times 10^{-4} \text{ mol}}{2 \text{ mol}} = 2.79 \times 10^{-4} \text{ mol of H}_2\text{SO}_4 \text{ (aq)}$

∴ C of H₂SO₄ = $\frac{N}{V \text{ in dm}^3} = \frac{2.79 \times 10^{-4} \text{ mol}}{0.0369 \text{ dm}^3} \approx 7.56 \times 10^{-3} \text{ mol dm}^{-3}$

- Or, they may ask you

Q) What volume of Sodium Hydroxide (NaOH) would it require to neutralise 32.3 cm³ of Sulphuric Acid (H₂SO₄) if the NaOH had a strength/concentration of 5.0 mol dm⁻³ and the H₂SO₄ had a strength of 3.0 mol dm⁻³ ?

Answer



Where $X = \frac{2 \text{ mol} \times .0969 \text{ mol}}{1 \text{ mol}} = 0.1938 \text{ mol of NaOH (aq)}$

∴ V of NaOH = $\frac{N}{C} = \frac{0.1938 \text{ mol}}{5.0 \text{ mol dm}^{-3}} \approx 3.88 \times 10^{-2} \text{ dm}^3$

Relationship between Mass and Concentration

• Concentration is the AMOUNT in a given VOLUME, and it can be stated either
 i) in **moles per dm³**, where $C \text{ (in mol dm}^{-3}\text{)} = \frac{N \text{ (in moles)}}{V \text{ (in dm}^3\text{)}}$,

or ii) in **g per dm³** where $C = \frac{\text{Mass (in g)}}{V \text{ (in dm}^3\text{)}}$;

• In the second instance i.e. (ii) above, we can say that

therefore $\text{Mass} = (N \times \text{RAM}) \text{ g}$,
 $C = \frac{(N \times \text{RAM}) \text{ g}}{V \text{ (in dm}^3\text{)}}$ for an element,

and $\text{Mass} = (N \times \text{RMM}) \text{ g}$ for a molecular compound
 therefore $C = \frac{(N \times \text{RMM}) \text{ g}}{V \text{ (in dm}^3\text{)}}$ for a molecular compound

and $\text{Mass} = (N \times \text{RFM}) \text{ g}$ for an ionic compound
 therefore $C = \frac{(N \times \text{RFM}) \text{ g}}{V \text{ (in dm}^3\text{)}}$ for an ionic compound

• There are many calculations that can be done by titrating liquids, and I would strongly urge you to work through as many calculations involving liquids as you can from “Advanced Chemistry Calculations” by I.L.P.A.C. (the ISBN No for it is 0 7195 75606 0) and also Jim Clark’s book ISBN 0-582-41127-0.

• Let me finish by once again giving you the rules for doing a titration calculation i.e.

- (A) Start with the **reaction equation**.
- (B) Work out the **MOLE Reaction Ratio**.
- (C) Work out the **number of moles used in the known solution**.
- (D) Use the reaction ratio to work out the **number of moles of the unknown solution**.
- (E) Use the **volume** of the unknown solution and the **number of moles** to work out its concentration (or you can work out the volume if you were given the concentration).

• Actually, even at University, titration calculations do not get much more difficult than this, because all that your tutors can do is to give you multiple titrations with one being dependent upon another, but the principles involved will be exactly the same and thus the work will conceptually **not** be more difficult.

- Eileen Ramsden (a former Prof of Chemistry at Leeds University) used to give the following examples of Titration exercises in her excellent book "A Level Chemistry" – and I have talked about these reactions when I was telling you about Ionic Half-Equations and the balancing of Ionic Equations.

1) **Reaction between acidified Potassium Manganate (VII), KMnO_4 and $\text{Fe}^{2+}(\text{aq})$**

- Acidified Potassium Permanganate is a very powerful Oxidising agent. [When you get to the Oxidisation of Alcohols please remember that you must not use \$\text{KMnO}_4\$ because it is too powerful.](#) You must use acidified Potassium Dichromate (Reaction 2 on this page).
- $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \longrightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) + 5\text{Fe}^{3+}(\text{aq})$
- At the equivalence point the purple Potassium Permanganate turns to such a pale pink that it is almost a colourless solution – and thus no indicator is needed.

2) **Potassium Dichromate (VI) and Ethanedioate**

- This is the gentler oxidising agent that you must use with Alcohols.
- $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 3\text{C}_2\text{O}_4^{2-}(\text{aq}) \longrightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 6\text{CO}_2(\text{g})$
- In the Acidified⁷ Potassium Dichromate (VI), $\text{K}_2\text{Cr}_2\text{O}_7(\text{aq})$, the Cr^{6+} ions are reduced to Cr^{3+} ions and you get the characteristic "orange to green" colour change.

3) **Iodine and Sodium Thiosulphate**

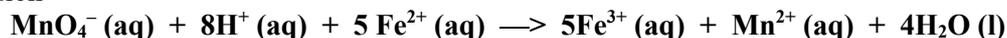
- There is a lovely colour change in this titration (just as there is with Potassium Permanganate). Just as the brown colour of the Iodine starts to fade at the end of the titration, just add a tiny amount of starch and the iodine will go a bright blue (and might even go to a very dark blue black) – and at the equivalence point "bang" the colour will go disappear just like that!
- $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{aq}) \longrightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2\text{I}^-(\text{aq})$

There is a very nice 2 minute video of this titration at https://www.youtube.com/watch?v=_tZYYZ9F7fM

⁷ The addition of a concentrated strong acid (such as concentrated Sulphuric Acid) speeds up the reaction.

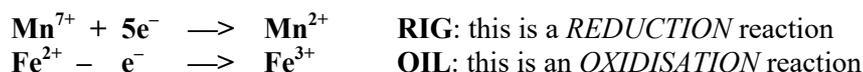
- OK, let me conclude this note by doing a couple of examples from the excellent ILPAC textbook “Advanced Chemistry Calculations” by Lainchbury Stephens and Thompson (LST). (I believe that it is out of print, but sometimes you can get hold of an old copy for 1p (plus package and postage from Amazon or whoever). It might cost £3-4 in total, but it is the best £3-4 that you will spend at ‘A’ Level.

- This is a question that LST set on page 27 of their book. I quote it because we will come across the equation



on a number of occasions in the course of the next two years.

- Please work out the Oxidisation numbers for Mn and Fe and then the aquated ionic half-equations for this Ionic equation are as follows



- If you multiply the Iron half-equation by 5 to equate the number of electrons, and add up the two half-equations to get rid of the electrons, then you get the Full Ionic Equation that I have quoted.
- OK, this is what LST said

“9.80g of pure (hydrated) Ammonium Iron(II) Sulphate [$\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$] was made up to 250 cm³ with dilute Sulphuric Acid (H_2SO_4). 25 cm³ of this solution was reacted completely with 24.6 cm³ of Potassium Manganate(VII) (Potassium Permanganate) solution. Calculate the concentration of Potassium Manganate(VII) solution.”

- Please do not let the Ammonium Iron(II) Sulphate [$\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$] put you off. All that it is doing is providing us with the Fe^{2+} that we need!
- Could **YOU** calculate the Relative Formula Mass of Ammonium Iron(II) Sulphate and you will see that the RFM or M_r is 392.0 g mol⁻¹ (i.e. 1 mole of it has a Mass of 392.0 g). It thus becomes a matter of simple proportions to calculate the number of moles in 9.8g viz.

1 mole of Ammonium Iron(II) Sulphate has a Mass of 392.0 g, therefore
X moles 9.8g

$$\text{Where } X = \frac{1 \text{ mol} \times 9.8 \text{ g}}{392.0 \text{ g}} \approx 0.025 \text{ mol}$$

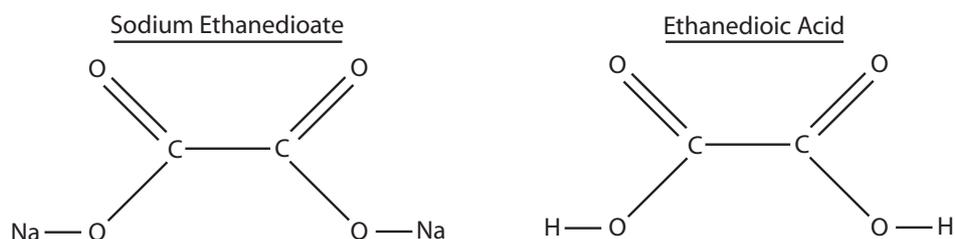
- If there are 0.025 mol of Ammonium Iron(II) Sulphate in 250 cm³,
then there are 0.0025 mol of Ammonium Iron(II) Sulphate in 25.0 cm³.

- From the Full Ionic Equation between Fe^{2+} and Mn^{7+} above we know that the Reaction Ratio is
therefore $5\text{Fe}^{2+} : 1\text{Mn}^{7+}$
 $0.0025 \text{ mol} : 5 \times 10^{-4} \text{ mol}$

- and since Concentration = Number of moles ÷ Volume in dm³,

$$\text{Concentration of the Potassium Permanganate solution} = \frac{5 \times 10^{-4} \text{ mol}}{24.6 \times 10^{-3} \text{ dm}^3} = 0.0203 \text{ mol dm}^{-3}.$$

- Let me do another one from Lainchbury Stephens and Thompson (LST) for you.
- You will need to remember that in Chapter 33 I told you that in the analysis of **the reaction between an Oxidising agent and a Dioate ion**, the ion that will be oxidised into Carbon Dioxide is called the "Ethane-dioate ion" (or "Ethanedioate ion") and it is written $(C_2O_4)^{2-}$, and it can be found for example in Sodium Ethanedioate $Na_2C_2O_4$, and in Ethanedioic acid $H_2C_2O_4$. In the example that I am about to do, MnO_4^- (the Permanganate(VII) ion) is the oxidising agent. You do need to know that when a dioate ion is oxidised, it forms Carbon Dioxide. *[In fact, in "Chains & Rings" you will learn that on Oxidisation, ALL carbon compounds form either Carbon Monoxide and this is called "incomplete oxidisation", or they will form Carbon Dioxide and this is called "complete oxidisation" – but we will get to that in due course.]*
- Just in case you are starting to forget how to draw bond line diagrams, I shall draw both Sodium Ethanedioate $Na_2C_2O_4$, and Ethanedioic acid $H_2C_2O_4$ for you *(and please do remember that we are talking about a "Dioate" and a "Dioic Acid" – and in Greek "di" means "two")*.

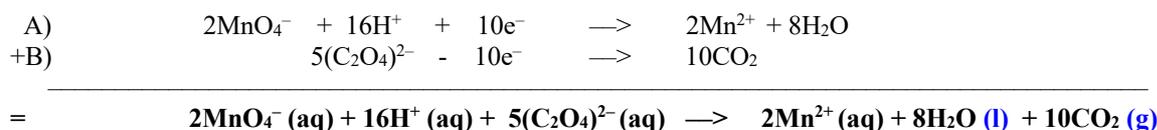


- You should by now be able to recognise the MnO_4^- Manganate(VII) half-equation that you require!



and it is obvious that equation (A) has to be multiplied by "2" and equation (B) by "5" in order to equate the two half-equations. *[NB Please could you derive Equation (B) for yourself from the equation given in the SEP Table. I'll give you a hint : the equation in the SEP table includes the H^+ ion from the Ethanedioic acid and also H^+ on the other side of the equation. You will have to multiply by 2 to create a full molecule of acid, cancel out H^+ ions and switch things around – and then Bob's your uncle!]*

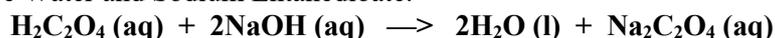
- If we do this and then add equations (A) and (B) together we get



- OK, this is what LST said

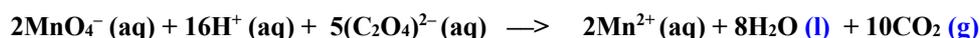
"25.0 cm^3 of a solution containing Ethanedioic Acid ($H_2C_2O_4$), and Sodium Ethanedioate ($Na_2C_2O_4$) required 14.8 cm^3 of 0.100 mol dm^{-3} Sodium Hydroxide solution for neutralisation. A further 25.0 cm^3 of the $H_2C_2O_4 / Na_2C_2O_4$ solution required 29.5 cm^3 of 0.0200 mol dm^{-3} Potassium Manganate(VII) solution for oxidisation under acidic conditions at about 60°C. Calculate the mass of each anhydrous constituent per dm^3 of the solution."

- **Acid/Base neutralisation:** Before you start doing any calculations, you ought to stop and think the thing through. The Sodium Hydroxide is an Alkali and it will therefore react with the Ethanedioic Acid to produce Water and Sodium Ethanedioate.



That means that the number of moles of $\text{Na}_2\text{C}_2\text{O}_4$ will have risen. The question may require the "before" or "after" mass. As it happens, the syntax in the question was poor. What LST meant to say was that A further 25.0 cm^3 of the $\text{H}_2\text{C}_2\text{O}_4 / \text{Na}_2\text{C}_2\text{O}_4$ solution required 29.5 cm^3 from a fresh lot of $0.0200 \text{ mol dm}^{-3}$ Potassium Manganate(VII) solution for oxidation under acidic conditions at about 60°C . In other words, LST wanted to know how much $\text{Na}_2\text{C}_2\text{O}_4$ there was in the **original mixture**.

- **MnO_4^- acting as an Oxidising agent:** The MnO_4^- will now Oxidise the Carbon in $(-\text{COO}^-)_2$ in both the Acid and the Ethanedioate into CO_2 . I have just derived the equation for you on page 12 viz.



Now we can start doing the Maths.

A) Calculation of the Mass of Ethanedioic Acid

- Number of moles of Sodium Hydroxide = $C \times V \text{ in dm}^3$
 $= 0.100 \text{ mol dm}^{-3} \times (14.8 \times 10^{-3}) \text{ dm}^3$
 $= 1.48 \times 10^{-3} \text{ mol}$.



Therefore Mole RR: $\frac{1}{7.4 \times 10^{-4}} \quad \frac{2}{1.48 \times 10^{-3}}$

Therefore No. of mol $7.4 \times 10^{-4} \quad 1.48 \times 10^{-3}$
 Therefore Concentration of Ethanedioic Acid = $N \div V \text{ in dm}^3$
 $= \frac{7.4 \times 10^{-4} \text{ mol}}{(25.0 \times 10^{-3}) \text{ dm}^3}$
 $= 0.0296 \text{ mol dm}^{-3}$

- Mass of Ethanedioic Acid = Relative Molecular Mass (in g mol^{-1}) x Number of moles (in mol dm^{-3})
 RMM or M_r of Ethanedioic Acid = $[(2 \times 1.0) + (2 \times 12.0) + (4 \times 16.0)] \text{ g mol}^{-1} = 90.0 \text{ g mol}^{-1}$
Therefore Mass of Ethanedioic Acid = $90.0 \text{ g mol}^{-1} \times 0.0296 \text{ mol dm}^{-3} = 2.664 \text{ g dm}^{-3}$.

B) Calculation of the Mass of Sodium Ethanedioate

- Number of moles of Potassium Manganate(VII) = $C \times V \text{ in dm}^3$
 $= 0.0200 \text{ mol dm}^{-3} \times (29.5 \times 10^{-3}) \text{ dm}^3$
 $= 5.9 \times 10^{-4} \text{ mol}$.



Therefore Mole RR: $\frac{2}{5.9 \times 10^{-4}} \quad \frac{5}{1.475 \times 10^{-3}}$

Therefore No. of mol $5.9 \times 10^{-4} \quad 1.475 \times 10^{-3}$
 Therefore Concentration of Sodium Ethanoate = $N \div V \text{ in dm}^3$
 $= \frac{1.475 \times 10^{-3} \text{ mol}}{(25.0 \times 10^{-3}) \text{ dm}^3}$
 $= 0.059 \text{ mol dm}^{-3}$

- However, $0.0296 \text{ mol dm}^{-3}$ came from the Ethanedioic Acid,
therefore $(0.059 - 0.0296) \text{ mol dm}^{-3}$ came from the Sodium Ethanedioate = $0.0294 \text{ mol dm}^{-3}$
- Mass of Sodium Ethanedioate = M_r (in g mol^{-1}) x Number of moles (in mol dm^{-3})
RMM or M_r of Sodium Ethanedioate = $[(2 \times 23.0) + (2 \times 12.0) + (4 \times 16.0)] \text{ g mol}^{-1} = 134.0 \text{ g mol}^{-1}$
Therefore Mass of Ethanedioic Acid = $134.0 \text{ g mol}^{-1} \times 0.0294 \text{ mol dm}^{-3} = 3.9396 \text{ g dm}^{-3}$
 $\approx 3.94 \text{ g dm}^{-3}$

OK, that has given you a lot of practice of calculations involving Liquids. I will look at Gases next.