(ii) the techniques and procedures used in experiments to measure volumes of solutions; the techniques and procedures used in experiments to prepare a standard solution from a solid or more concentrated solution and in acid–base titrations

Measuring volumes of solutions:

- Measuring cylinders and beakers may be used to measure out a **rough** volume of liquid.
- To determine a more **accurate** volume, **graduated pipettes** or **burettes** should be used.
- Pipettes:
  - Ensure pipette is clean by **rinsing with distilled water followed by** some of the **solution** to be pipetted.
  - Dip the end of the pipette into the solution and draw it up using a **pipette filler** until the desired volume is reached (when the **bottom of the meniscus** sits on the **graduation line** when viewed at **eye-level**).
  - Allow the liquid to run off into glassware **until it stops**. **Touch the end of the pipette to the side of the vessel** and remove the pipette.
  - **Do not force the last few drops from the pipette** as the apparatus is **calibrated to deliver** the required volume without adding the last drops.
- Burettes:
  - **Clean** the burette by **rinsing with water** and then a small volume of the **solution to be used**.
  - Ensuring the tap is closed, pour in the solution to be measured using a **funnel** so that the **meniscus is above the 0 line**.
  - Open the tap allowing some of the solution to run off into some glassware. Ensure the **jet is full** of the solution and there are **no air bubbles** present.
  - Allow the solution to run off until the **bottom of the meniscus is sitting on a graduation line**. Record the **initial burette reading to the nearest 0.05cm^3**.
  - Run the required volume of solution from the burette into a suitable transfer vessel.

Preparing a standard solution from solid:

- Calculate the required mass of solute. Accurately weigh the solid using a weighing bottle and balance to 2d.p, recording the mass of weighing bottle and weighing bottle + solid combined to determine the mass of solute weighed out.
- Pour 100cm^3 of deionised water into a 250cm^3 beaker, transferring all of the measured solute into it.
- Reweigh the weighing bottle to determine mass of solute added.
- Stir the mixture to completely dissolve the solute.
- Transfer the solution into a 250cm^3 volumetric flask. Wash the beaker and stirring rod with deionised water, transferring the washings to the volumetric flask.
- Add deionised water to the volumetric flask until the volume is within 1cm of the graduation line. Swirl at fixed intervals to ensure the solution is thoroughly mixed.
- Use a dropping pipette to add deionised water dropwise until the bottom of the meniscus is sitting on the graduation line at eye-level.

- Insert a stopper and invert the flask a few times to ensure thorough mixing.

Preparing a standard solution by dilution:

- Rinse a clean dry beaker with original standard solution (stock solution) and half fill it.
- Use a pipette filler to rinse a clean 25cm^3 pipette with the stock solution. Fill it to the graduation line (meniscus bottom on the line) and transfer it to a clean 250cm^3 volumetric flask.
- Add deionised water to make up the rest of the volume (same procedure as making a standard solution from a solid).

To prepare 50 m^3 of 1.0 mol dm^{-3} solution from a stock solution of concentration 2.0 mol dm^{-3}, what volume of stock solution is required?

Step 1: Use the equation $C_1V_1 = C_2V_2$ and insert known values.
$1.0\,\text{mol dm}^{-3} \times 0.05\,\text{dm}^3 = 2.0\,\text{mol dm}^{-3} \times Z\,\text{dm}^3$

Step 2: Rearrange the equation to calculate the required volume of stock solution.

$$Z = \frac{1.0\,\text{mol dm}^{-3} \times 0.05\,\text{dm}^3}{2.0\,\text{mol dm}^{-3}} = 0.025\,\text{dm}^3$$

Step 3: Convert dm^3 reading to cm^3 reading – 25 cm^3 of stock solution needed.

Acid-base titration:

- Rinse a burette with acid solution then fill it above the 0 line with the acid. Run off a little acid into a waste beaker to fill the jet and set the meniscus on a graduation line. Record initial reading to nearest 0.05cm^3.
- Fill a clean pipette with 25cm^3 alkaline solution and transfer it to a 250cm^3 conical flask.
- Add two or three drops of suitable indicator (phenolphthalein or methyl orange) to the conical flask and swirl to mix. Place on a white tile to make the end point more obvious.
- Run the acid into the conical flask until the first signs of a permanent colour change appear (when the end point is reached). This is a rough titration to give an idea of the volume required to neutralise the alkali. Record the final burette reading and subtract from the initial to give the titre.
- Refill the burette if necessary and record initial reading.
- Repeat step 3 and 4.
- Run the acid solution within 1cm^3 of the rough titre and add the rest dropwise until the first sign of a permanent colour change has occurred (end point).
- Repeat the accurate titrations until you have 3 concordant titres (within 0.1cm^3 of each other).

Common indicators used for acid-base titrations:

- Indicators that change colour abruptly over a small pH range are needed to find the exact moment that the end point has been reached:
    - Phenolphthalein - changes from colourless to pink when adding alkali to acid
    - Methyl orange - changes from red to yellow when adding alkali to acid
- Indicators such as universal indicator are too gradual to accurately determine end point.
- A white tile should be used to give the colour change more prominence to judge the end point more accurately.

(d) balanced full and ionic chemical equations, including state symbols

1) You can also write an **ionic equation** for any reaction involving **ions** that happens **in solution**.
2) In an ionic equation, only the **reacting particles** (and the **products** they form) are included.

Example: Here is the **full balanced equation** for the reaction of **nitric acid** with **sodium hydroxide**:

$$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$$

The **ionic** substances in this equation will **dissolve**, breaking up into ions in solution. You can rewrite the equation to show all the **ions** that are in the reaction mixture:

$$H^+ + NO_3^- + Na^+ + OH^- \rightarrow Na^+ + NO_3^- + H_2O$$

Leave anything that isn't an ion in solution (like the H_2O) as it is.

To get from this to the ionic equation, just cross out any ions that appear on **both sides** of the equation — in this case, that's the sodium ions (Na^+) and the nitrate ions (NO_3^-).

An ion that's present in the reaction mixture, but doesn't get involved in the reaction is called a spectator ion.

So the **ionic equation** for this reaction is:

$$H^+ + OH^- \rightarrow H_2O$$

3) When you've written an ionic equation, check that the charges are **balanced**, as well as the atoms — if the charges don't balance, the equation isn't right.

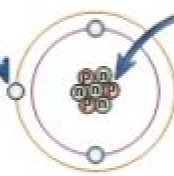
In the example above, the **net charge** on the left hand side is +1 + –1 = **0** and the net charge on the right hand side is **0** — so the charges balance.

s = solid
l = liquid
g = gas
aq = aqueous (solution in water)

(e) conventions for representing the distribution of electrons in atomic orbitals; the shapes of s- and p-orbitals.

All elements are made of **atoms**. Atoms are made up of 3 types of particle — **protons**, **neutrons** and **electrons**.

Electrons
1) Electrons have **–1** charge.
2) They whizz around the nucleus in **shells**. These shells take up most of the **volume** of the atom.



Nucleus
1) Most of the **mass** of the atom is concentrated in the nucleus.
2) The **diameter** of the nucleus is rather titchy compared to the whole atom.
3) The nucleus is where you find the **protons** and **neutrons**.

The mass and charge of these subatomic particles is **tiny**, so **relative mass** and **relative charge** are used instead.

Subatomic particle	Relative mass	Relative charge
Proton	1	+1
Neutron	1	0
Electron, e^-	$\frac{1}{2000}$	–1

The mass of an electron is negligible compared to a proton or a neutron — this means you can usually ignore it.

(f) the electronic configuration, using sub-shells and atomic orbitals, of:

(i) atoms from hydrogen to krypton

Mg: $1s^2 2s^2 2p^6 3s^2$ or $[Ne]3s^2$

Kr: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ or $[Ar]3d^{10} 4s^2 4p^6$

Cr: $[Ar]3d^5 4s^1$

Cu: $[Ar]3d^{10} 4s^1$

- Chromium and copper are exceptions. The 4s sub-shell **does not completely fill** before the 3d sub-shell begins to fill. This is thought to be because this configuration is more **electrically stable**.

(ii) ions of the s- and p-block of Periods 1 to 4

Mg^{2+}: $1s^2 2s^2 2p^6$ or $[Ne]$

Co: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$

Co^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7$

- Ions in the d-block lose their electrons in the 4s sub-shell before 3d, even though 4s fills before 3d.

(g) how knowledge of the structure of the atom developed in terms of a succession of gradually more sophisticated models; interpretation of these and other examples of such developing models

<u>Dalton's atom</u>

- Ancient greeks believed matter consisted of **indivisible particles** . Atom comes from the greek word **atomos**, meaning indivisible.
- John **Dalton** described atoms as **solid spheres** where **different spheres make up different elements**

<u>Plum pudding model</u>

- JJ Thomson did a series of experiments to conclude atoms **were not solid and indivisible**.
- Measurements of charge and mass revealed that atoms must contain smaller negatively charged particles which he named **corpuscles (electrons)**.
- The '**plum pudding**' model was created - a **positively charged sphere** with **negatively charged electrons** embedded within it.