2. Spot ink on the line (not touching solvent initially) & place sheet in solvent (ethanol or water).
3. Place watch glass on top to prevent solvent evaporating.
4. Ink will dissolve in solvent as they move up paper together.
5. Different dyes move up at different rates, so they separate & form spots in different places.
6. Point where solvent reaches as it reaches top is known as the solvent front (draw line in pencil).

☆ Amount of time spent in each phase (& hence the distance the molecules travel) is based on how soluble it is in the solvent & how attracted they are to the paper. If solubility is higher, more time in mobile phase, if attraction is higher, more time in stationary phase.

☆ Some chemicals are colourless & require a locating agent sprayed on it for it to show up.

☆ The R_f value is the ratio between the distance travelled by solute (baseline to spot) & distance travelled by solvent (baseline to solvent front).

☆ CALCULATING R_f VALUE (2 decimal places between 0 & 1):
R_f value = distance travelled by solute / distance travelled by solvent

☆ Paper chromatography is normally carried out to see if a certain chemical is present in a substance. This is investigated by testing the pure sample & unknown mixture & looking for a spot with the same R_f value (checked by repeating with different solvent).

☆ Pure substances would not be separated by a paper chromatography, unlike a mixture.

☆ In the UK, there are many sources of water that can be purified to potable water (safe to drink), including: surface water (lake, rivers & reservoirs), ground water (aquifers), & waste water (contaminated by human processes).

☆ HOW TO PURIFY WATER FOR DRINKING:
1. Filtration: using wire, mesh screen to sieve large objects.
2. Filtration: using gravel & sand beds to filter out solid bits.
3. Sedimentation: tank of iron/ aluminium sulfate to make fine particles clump together & settle.
4. Chlorination: chlorine gas bubbled through to kill microorganisms.

☆ Sea water can be distilled to provide potable water by simple distillation; however, it is very expensive as it requires a lot of energy.

☆ Deionised water is water with the normal ions found in tap water removed (calcium, iron & copper) – this prevents false results.

- ☆ While acid strength (strong or weak) tells you about the proportion of acid molecules that ionise in water, the concentration (concentrated or dilute) is very different. Concentration is a measure of how much acid there is in a litre (g/dm^3 or mol/dm^3) of water – ratio of dissolved acid molecules to volume of water.
- ☆ pH is dependent on the acid's concentration of hydrogen ions. If the concentration of hydrogen ions increases by a factor of 10, the pH decreases by 1 (e.g. concentration increase by factor of 100 (10 x 10) causes pH decrease by 2 (1+1)). Hence, if concentration decreases by factor of 10, the pH increases by 1.
- ☆ All metal oxides & hydroxides are bases; soluble metal hydroxides are alkalis because they dissolve in water to form hydroxide ions. As bases, metal oxides/ hydroxide reacts with acids to form salt & water (general equation: **acid + metal oxide/ hydroxide –> salt + water**).
- ☆ The salt formed are ionic compound with the metal the start of the name & acid at the end (e.g. chloride, nitrate, sulfate, & ethanoate).
- ☆ When other metals react with acid, they produce metal salt & hydrogen gas (**general equation acid + metal –> metal salt + hydrogen**). Some metals react so violently, they produce enough heat to ignite the hydrogen gas formed. This can be proven with a hydrogen test – where you hold a lit splint in the test tube until you hear a 'squeaky pop' sound.
- ☆ The squeaky pop is generated from the hydrogen react quickly with oxygen to form H_2O.
- ☆ When metal carbonates react with acid, they also produce a salt, water, & carbon dioxide gas (general equation: **acid + metal carbonate –> salt + carbon dioxide + water**). This can be proven by testing for carbon dioxide – where you bubble the gas through limewater (calcium hydroxide in water) & it turns cloudy.
- ☆ The cloudiness is a result of the formation of calcium carbonate.
- ☆ Salts are either soluble or insoluble – each one is obtained in a different way.
- ☆ WHETHER A SALT IS SOLUBLE OR INSOLUBLE
  common salts of sodium, potassium & ammonium: soluble
  nitrates: soluble
  common chlorides: soluble (except with silver & lead)
  common sulfates: soluble (except with lead, barium & calcium)
  common carbonates & hydroxides: insoluble (except with sodium, potassium & ammonium)
- ☆ Insoluble salt will form as a precipitate when two solution are mixed.
- ☆ HOW TO MAKE INSOLUBLE SALTS FROM SOLUBLE SALTS (precipitation & filtration):
  1. Add one spatula of soluble salt & some deionised water to dissolve it (shake).
  2. Repeat in a different test tube with another soluble salt.
  3. Tip both solution into a small beaker & stir. This causes the insoluble salt to precipitate.
  4. Place a funnel into the neck of a conical flask & add filter paper into a funnel.
  5. Pour contents into funnel (not down the side).
  6. Swill out beaker with more deionised water to ensure all precipitate is taken from beaker.
  7. Rinse contents of filter paper with water to rinse away any soluble salts.
  8. Scrape insoluble salt (residue) from filter paper & leave to dry.
- ☆ Soluble salts are made by reacting an acid with a metal or insoluble base (metal oxide/ hydroxide/ carbonate). You can choose the right reagents from the salt's name.
- ☆ HOW TO MAKE SOLUBLE SALTS FROM ACID & INSOLUBLE BASE (e.g. copper sulfate) (filtration & crystallisation):
  1. Put sulfuric acid in a beaker & gently heat in a water bath (in a fume cupboard).
  2. Add copper oxide (base) & stir – they will react to form copper sulfate & water.

3. Keep adding copper oxide until in excess & reaching bottom (neutralisation is done).
4. Filter out excess copper oxide using filter paper & funnel.
5. Crystallise copper sulfate solution (filtrate) by heating in evaporating dish.
6. Stop heating & leave solution to cool. Filter out crystals & dry.

copper oxide + sulfuric acid –> copper sulfate + water

$CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

☆ Soluble salts can be also be made using titration with an acid & an alkali – where you see the volume of reactant needed to react completely with another reactant. Instead, the alkali reactant is soluble (unlike the base).

☆ HOW TO MAKE SOLUBLE SALTS USING ACID-ALKALI REACTION (titration & crystallisation):

1. Measure out set amount of acid into conical flask using a pipette, add some indicator & place on a white tile.
2. Fill the burette with alkali & record the volume.
3. Slowly add alkali & swirl each time until indicator changes colour (acid has been neutralised).
4. Read final volume left in burette to calculate volume needed to neutralise acid.
5. Then repeat reaction with exact same volume of acid & alkali (& no indictor) to produce an uncontaminated salt.
6. Perform a crystallisation reaction to the solution of salt & water.

☆ The indicator used must have a clear colour change between acid/ alkali & neutral (e.g. methyl orange or phenolphthalein).

☆ Electrolysis is when you break down a substance using electricity. You pass an electric current through an ionic substance (molten or in solution) & the ions move towards the electrodes. At the electrodes, they react, & the ionic substance starts to decompose (e.g. aluminium oxide into aluminium & oxygen). It requires a liquid to conduct the electricity – known as electrolytes (contain free ions & allow it to work).

☆ The ionic compound cannot be solid because the ions are in fixed positions, but molten or dissolved compounds have ions that can move freely & conduct electricity. Also, electrodes must be made from inert material (e.g. graphite or platinum) so they do not react.

☆ HOW ELECTROLYSIS WORKS:

1. Two electrodes (electricity- conducting solid) is submerged into the electrolyte (molten/ in solution) to create an electrical circuit.
2. Ions move between electrodes – allowing the conduction of electricity through the circuit.
3. Positive ions (cations) move towards the negative electrode (cathode) & gain electrons.
4. Negative ions (anions) move towards the positive electrode (anode) & lose electrons.
5. As they gain or lose electrons (& become atoms/ molecules), they are discharged from the electrode – these are the products of electrolysis.

☆ Electrolysis always involves a redox (reduction & oxidation) reaction. Reduction occurs at the negative electrode as the cations are gaining electrons; oxidation occurs at the positive electrode as the anions are losing electrons.

☆ Binary compounds are ionic compounds containing two elements (cation & anion) that are ions. Electrolysis of a binary compound produces neutral metal & non-metal elements from losing/ gaining electrons at the electrodes.

☆ Half-equations show how electrons are transferred during the reactions at the electrodes.

☆ HOW TO WRITE A HALF EQUATION (at cathode/ negative electrode):

1. Write chemical symbol for positive ion on left hand side of equation.

TOPIC 5: Separate Chemistry 1

➢ METALS, ALLOYS & CORROSION

- ☆ Transition metals are found in the centre of the periodic table. (e.g. gold, silver, copper, iron, zinc & platinum). They have the typical properties of a metal, including: strong, hard, dense, shiny, malleable, & ductile. Also, they have a high melting points (except mercury) & high densities.
- ☆ They have many uses, including: gold for jewellery & electrical components; copper for water pipes & electrical wires, iron as a catalyst (in the Haber process for ammonia); & vanadium pentoxide as a catalyst (V_2O_5 - in the Contact process for sulfuric acid)
- ☆ The compounds of transition metals are colourful (depending on the metal ion contained), including: blue (Cu^{2+}), green (Fe^{2+}), orange brown (Fe^{3+} - rust).
- ☆ Alloys are mixtures of a metal & either a metal or non-metal element. Adding another metal, alters the pattern of the lattice, so the layer cannot easily slide over each other anymore (not malleable). This is because different elements have different size atoms. Hence, alloys are stronger & harder than pure metal.
- ☆ Shape memory alloys can return to their original shape after deformation.
- ☆ Pure iron is too soft & easily shaped for most uses, so it is mostly produced to make alloys of steel. Different types of steel have different properties & uses, including: Low-carbon steel with 0.1-0.3% carbon (which is easily shaped & used for car bodies), High-carbon steel with 0.22-2.5% carbon (which is very strong, flexible, & brittle & used for cutting tool blades & bridges) & Stainless steel with chromium or nickel sometimes added (which is corrosion-resistant & hard & used for cutlery & corrosive substance containers). Steel is less likely to rust/corrode.
- ☆ Brass is an example of a copper alloy. It uses copper & zinc & has a gold-like appearance. It is more malleable than bronze & ideal with moving parts that require low friction (e.g. water taps & door fittings). Another example is bronze – which is an alloy of copper & tin. It is harder & good for making statues, decorative objects & medals.
- ☆ Gold alloys contain zinc, copper or silver to harden the gold, when used in jewellery. Pure gold is described as 24 carat – the lower the carat, the lower the percentage of pure gold in the alloy.
- ☆ CALCULATING MASS OF GOLD IN AN ALLOY:
  mass of gold in alloy = $^{\text{carat}}/_{24}$ x mass of alloy
- ☆ Pure aluminium is not strong or dense – however, the low density is useful for aircraft manufacture. When it is mixed with magnesium, it makes the alloy Magnalium. If the alloy has a small amount of magnesium (5%), it is strong, light & less easily corroded, so it is used for making aeroplane parts. If the alloy has large amount of magnesium (50%), it is reactive & burns brightly like magnesium (but it is more stable than magnesium), so it is used for fireworks.
- ☆ Metals can corrode in the presence of oxygen & water to form their metal oxides (becomes oxidised) – this can be proven in the rust experiment. Rusting is specifically the corrosion of iron (**iron + oxygen + water –> hydrated iron (III) oxide**).
- ☆ HOW TO PERFORM THE RUST EXPERIMENT:
  1. Water with no air: iron nail in boiling tube with water & oil.
  2. Air with no water: iron nail in boiling tube with calcium chloride (drying agent).
  3. CORROSION (air & water): iron nail with water & air.
  4. CONTROL (nothing): iron nail in Vaseline with calcium chloride.
- ☆ The mass of the rust nail will increase as iron atoms bond with the oxygen & water molecules – this results in a compound heavier than iron by itself.

- ☆ HOW PHYSICAL PROPERTIES OF HYDROCARBONS DIFFER:
  1. Boiling points: intermolecular forces are bigger in bigger molecules, so more energy is required to break it; they also have more points along its length to be attracted to another molecule (so they are many places the forces must be broken.
  2. Ease of ignition/ flammability: shorter molecules are easier to ignite; since they have lower boiling points, they tend to be gases at room temperature. The gas molecules mix with oxygen in the air – this gas mixture ignites if it comes in contact with a spark.
  3. Viscosity/ thickness (how easily a substance flows): stronger the intermolecular forces are, harder it is to flow; longer chain hydrocarbons have higher viscosity.
- ☆ If you burn hydrocarbons, the carbon & hydrogen atoms react with oxygen to form carbon dioxide & water (general equation: **hydrocarbon + oxygen –> carbon dioxide + water**) – this them great fuels. Energy is released to the surroundings – making it an exothermic reaction. Carbon dioxide & water are the only products when the reaction has sufficient oxygen.
- ☆ However, if the reaction has insufficient oxygen in the air, they undergo incomplete combustion (this happens in some appliances, like boiler that use carbon compounds as fuel). The products contain less oxygen; carbon dioxide & water are still produced, but so can carbon monoxide (CO) & carbon in the form of soot.
- ☆ Carbon monoxide can bind to the haemoglobin in red blood cells – which normally carries oxygen; this means less oxygen can be transported around your body; a lack of oxygen in the blood supply can lead to fainting, coma, or death. Soot is tiny particles of carbon; it makes buildings look dirty, reduces air quality, worsen/ cause respiratory problems (damage to cells lining lungs) or cause global dimming (acts a barrier against sunlight when carried in the air). If a household appliance is producing carbon monoxide or soot, it is inefficient & a serious health hazard.
- ☆ Another pollutant is sulphur dioxide – which is released when fossil fuels are burnt along with carbon dioxide & various nitrogen oxides, due to sulfur impurities in the fossil fuels. When sulfur dioxide mixes with water in clouds, they react to form dilute sulfuric acid – which falls as acid rain. Acid rain causes a lot of problems, including: killing trees & aquatic life, damaging buildings, & statues (mostly limestone), breathing problems in humans & corrosion of metal.
- ☆ When fuels are burnt in the internal combustion engines of a car, they release a lot of energy in the form of heat; these high temperatures cause the nitrogen & oxide in the air to react & form nitrogen oxides (reverted back using catalytic converters). These are harmful pollutants, since they contribute to acid rain & photochemical smog – which is a type of air pollution that can cause breathing difficulties, tiredness, & headaches (mostly due to traffic in large cities).
- ☆ Instead of using petrol to power vehicles, hydrogen can be used; it can also be used for fuel cells.
- ☆ HOW HYDROGEN AS FUEL HAS ADVANTAGES & DISADVANTAGES:
  1. Advantages: fuel cell produce energy with waste product of water, no harmful pollutants released, obtaining from water (renewable resource), & can be obtained from waste product of fuel cell.
  2. Disadvantages: special, expensive energy required to use hydrogen as a fuel, must be manufactured – which is expensive & requires energy from fossil fuels releasing pollutants (normally) for electrolysis of water, highly flammable, hard to store safely, & expensive as it is not widely available.
- ☆ Cracking is when longer saturated hydrocarbons (alkanes) are broken down into smaller, more useful ones; smaller hydrocarbon chains are flammable (makes a good fuel), alkenes are used to make polymers, & both are in higher demand, but longer chain are thick, gloopy liquids. It produces alkanes (saturated) & alkenes (unsaturated).

➢ ALCOHOLS & CARBOXYLIC ACIDS

- ☆ Alcohols are a homologous series containing the -OH functional group. They end with the suffix 'ol'; their general formula is '$C_nH_{2n+1}OH$'. It begins with methanol (CH_3OH), ethanol, propanol, then butanol. Like alkenes, most alcohol are isomers – which are molecules with the same molecular formula, but different structure (this can affect the reactivity).
- ☆ It can be represented as a molecular formula (ratio of atoms), structural formula (shows functional group), or displayed formula (2D representation of covalent bonds). In displayed formula, the hydrogen atoms surround the carbon atoms – which is bonded to the functional group.
- ☆ When you heat a mixture an alcohol with an acid catalyst, an alkene & water are formed – this is known as a dehydration reaction.
  alcohol (heat+ acid catalyst) –> alkene & water
- ☆ Fermentation is the process of using yeast to produce alcohols from carbohydrates – it is anaerobic respiration; sugars (e.g. glucose) are converted to ethanol & carbon dioxide. The carbohydrates mostly come from sugar cane or beet plants; the yeast cells contain enzymes (zymase) – which speed up the reaction.
  glucose (yeast) –> ethanol + carbon dioxide
  $C_6H_{12}O_6$ (yeast) –> $2C_2H_5OH + 2CO_2$
- ☆ HOW TO PRODUCE ETHANOL BY FERMENTATION:
  1. Mix yeast & carbohydrate solution, seal container & leave in warm place.
  2. Keep at optimum temperature of 30-40^{o}C for enzymes (too low, too slow; too high causes it to denature & stop reaction); keep anaerobic conditions (or ethanol becomes ethanoic acid when oxidised & aerobic respiration to produce water, instead of ethanol).
  3. When concentration reaches 10-20%, fermentation eventually stops – this is because the alcohol kills off the yeast.
  4. Collect ethanol solution from the top once the yeast will fall to the bottom.
- ☆ Fermentation can only produce a dilute concentration of alcohol (e.g. beer at 4%); the solution can be distilled to produce a high concentration (e.g. spirits at 40%).
- ☆ To increase the concentration, fractional distillation is used; this is because ethanol has a lower boiling point at 78^{o}C compared to water at 100^{o}C. The fermentation mixture is heated, so ethanol evaporates & rises up the fractional column; then, it condenses back into liquid in the Liebig condense & collecting in a separate container.
- ☆ It is best to use an electric heater as ethanol is flammable.
- ☆ When alcohols burn, they release energy; this can be used to power cars, heat homes, or generate electricity. The amount of energy is gives out determines how good of a fuel it is (more is better).
- ☆ HOW TO INVESTIGATE ALCOHOLS AS A FUEL:
  1. Put some alcohol into a spirit burner, then measure the mass of the burner & fuel on a mass balance.
  2. Measure 100 cm^3 of distilled water into a copper calorimeter.
  3. Insulate the calorimeter with a draught excluder & cover with insulating lid – then add a thermometer.
  4. Take the initial temperature of the water, then put the burner under the calorimeter & light.
  5. Continue to stir water using the thermometer (blow out burner when temperature reaches 20^{o}C).
  6. Find mass of fuel that has been used (subtract final mass from initial).
  7. Repeat with different alcohols.
- ☆ To compare your results, find the alcohol that used the least mass to heat the water to the given temperature; this means more energy per gram is released.