Common Laboratory Techniques

DVHS Chemistry – Mrs. Farmer

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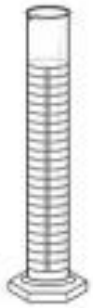
COMMON LAB EQUIPMENT

Safety Splash Goggles 	Beaker 	Erlenmeyer Flask 	Graduated Cylinder 
Distilled Water Wash Bottle 	Beaker Tongs 	Crucible Tongs 	Test Tube Tongs 

Common lab equipment
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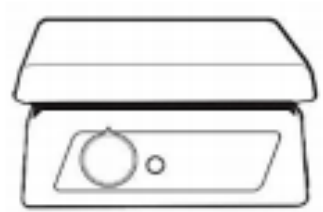
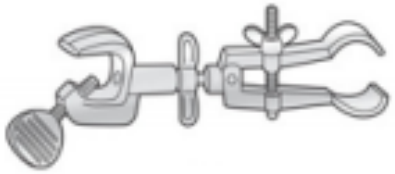
Hot Plate 	Spatulas and Scoopulas 	Disposable Pipette 	Rubber Policeman 
Forceps 	Ring Stand 	Iron Support Ring 	Utility Clamp 

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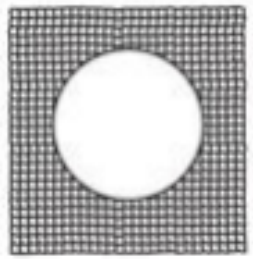
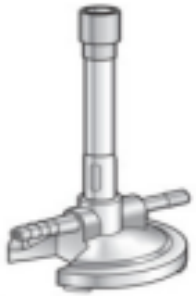
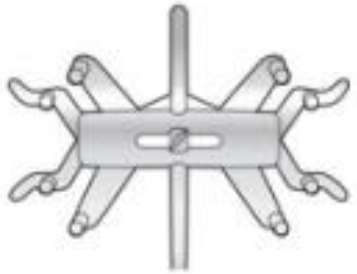
Wire Gauze with Clay Center 	Bunsen Burner 	Flint Striker 	Clay Triangle 
Crucible with Lid 	Evaporating Dish 	Burette Clamp 	Burette 

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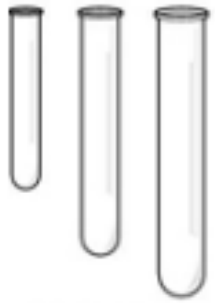
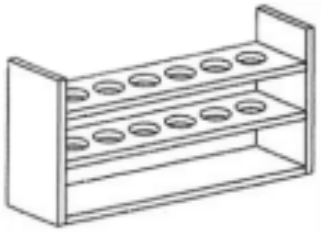
Filter Flask 	Buchner Funnel 	Aspirator for Sink 	Glass Funnel 
Test Tube Brush 	Test Tubes 	Test Tube Rack 	Morter and Pestle 

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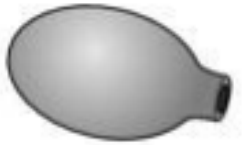
Volumetric Flask 	Glass Watch Glass 	Volumetric Pipette 	Rubber Pipette Bulb 
Rubber Stoppers 			

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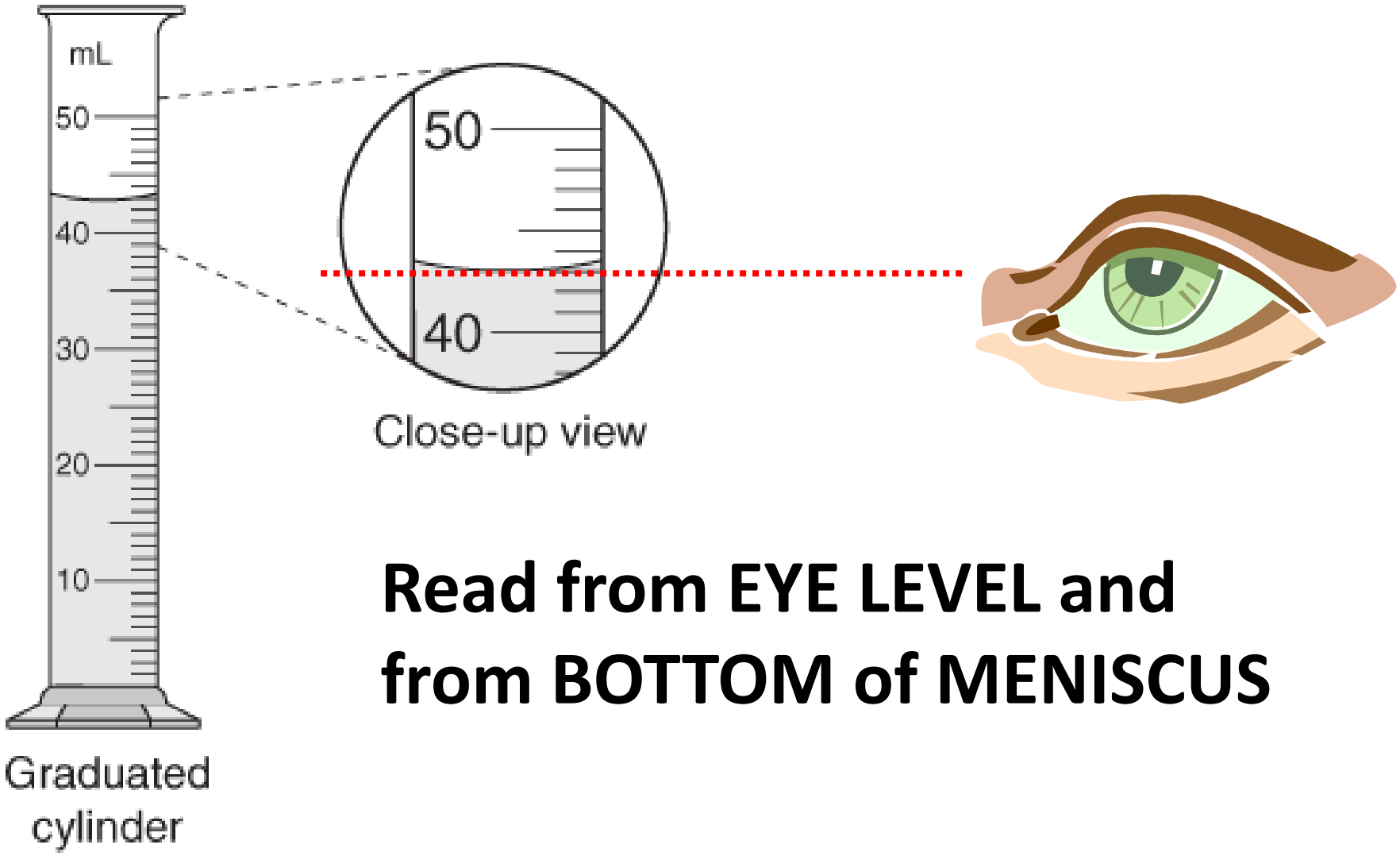
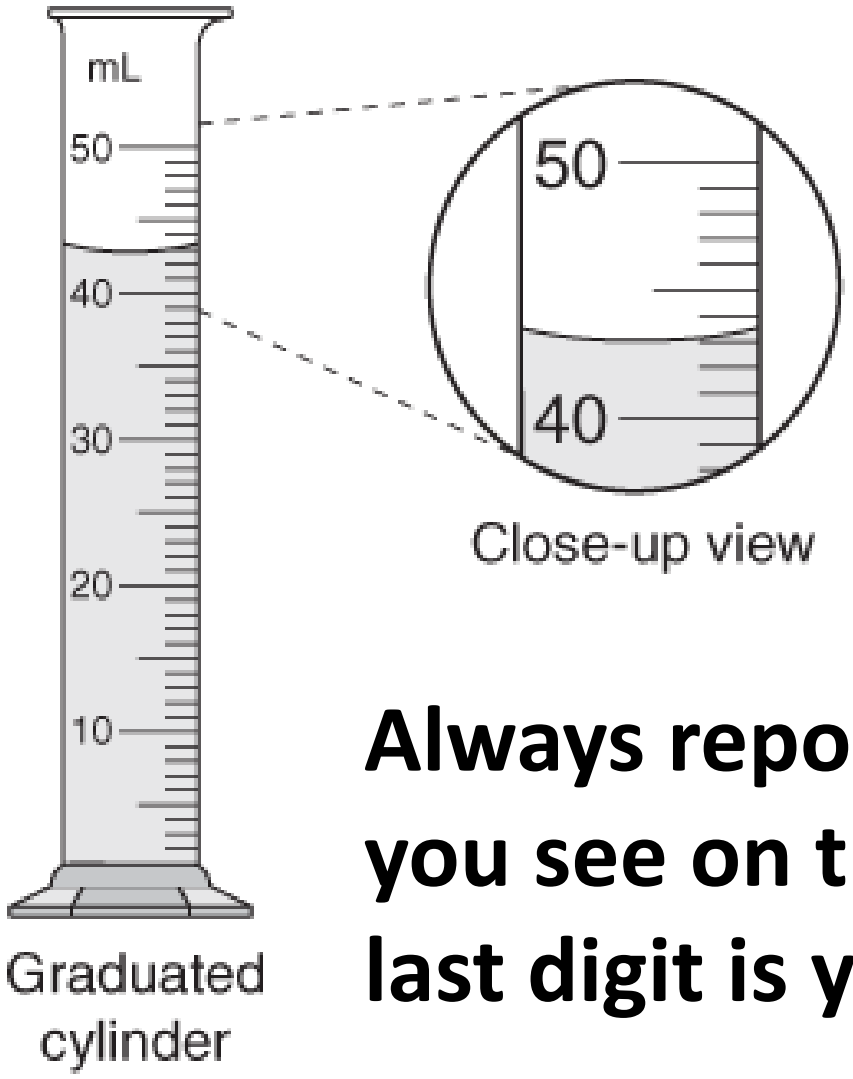
USING A GRADUATED CYLINDER

Using a graduated cylinder



Read from EYE LEVEL and from BOTTOM of MENISCUS

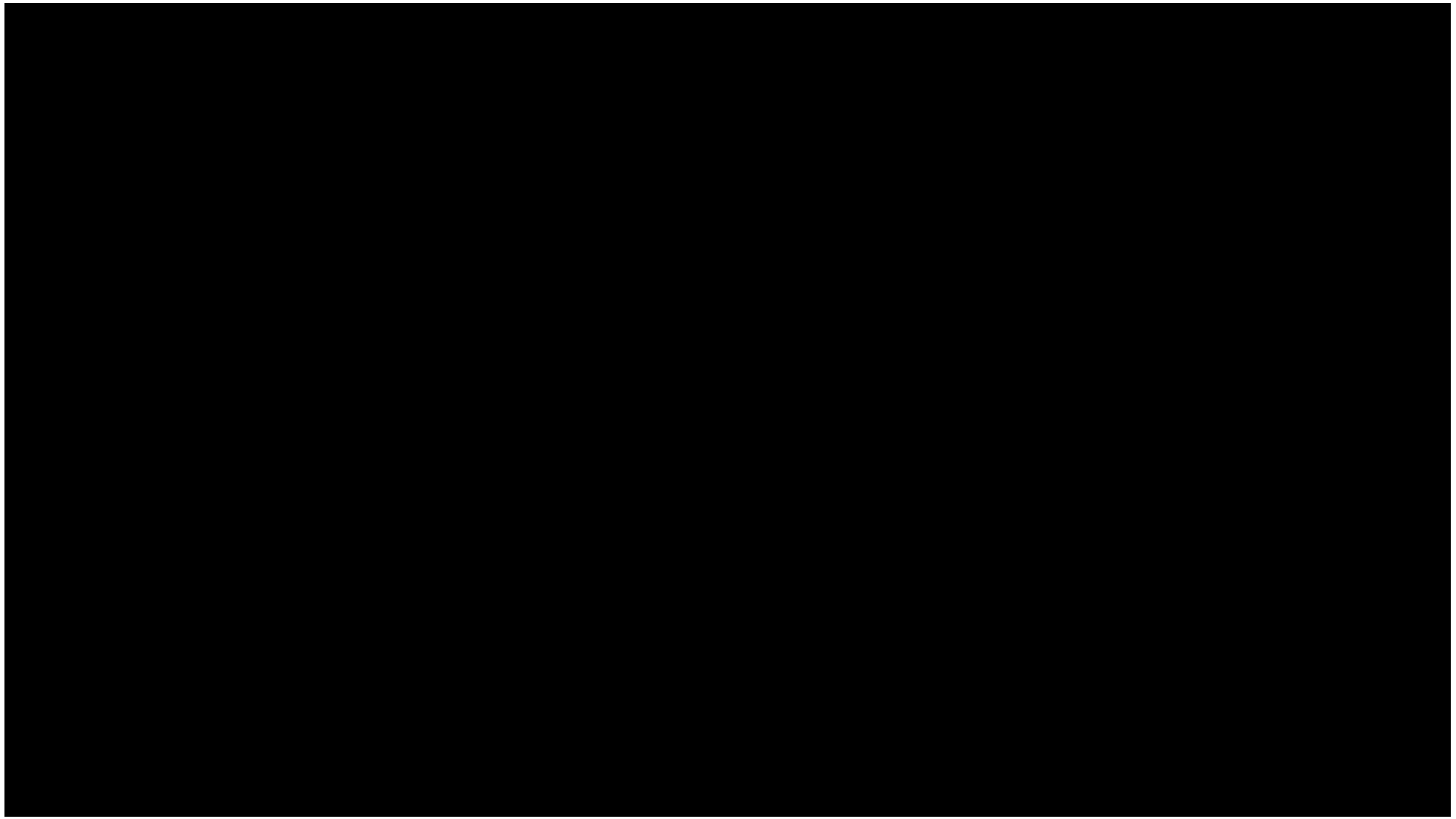
Using a graduated cylinder



This would be: 42.9 mL
It is somewhere between 42 and 43, my best guess is that it is 42.9 which means the 9 is my uncertain digit.

Always report one more digit than what you see on the cylinder markings – that last digit is your estimated, uncertain digit.

Using a graduated cylinder



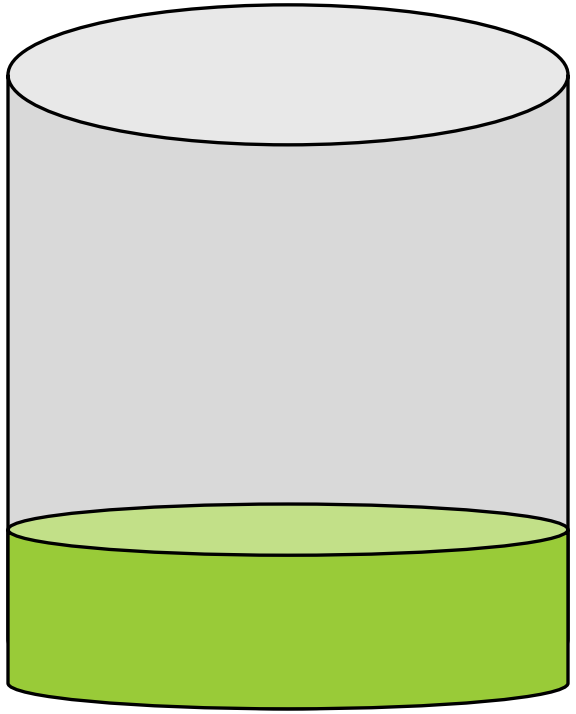
https://youtu.be/eUIcebi1GVc

VOLUME OF AN IRREGULAR OBJECT VIA WATER DISPLACEMENT

Volume of an irregular object

Measure volume by **water displacement**

10 mL



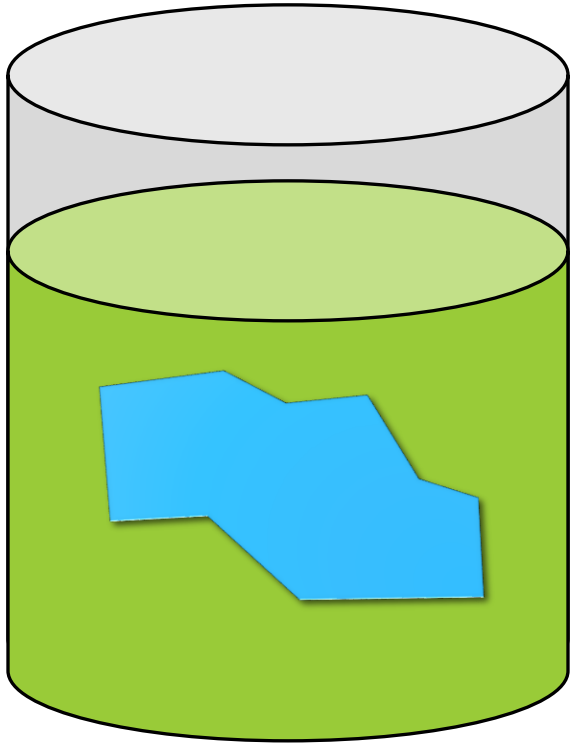
30 mL

- 10 mL

20 mL

= object volume!

30 mL



Volume of an irregular object

Measure volume by **water displacement**

https://youtu.be/eUIcebi1GVc

(start at minute 3:09)

USING A DIGITAL SCALE

Using a digital scale

- Make sure to hit the TARE button on the scale! Sometimes it is called the ZERO button

- NEVER weigh chemicals directly on the plate. Make sure to use some container or piece of weigh paper.

- Added too much of the chemical to the scale? DO NOT PUT BACK IN THE JAR!!! NO CONTAMINATION!

Using a digital scale

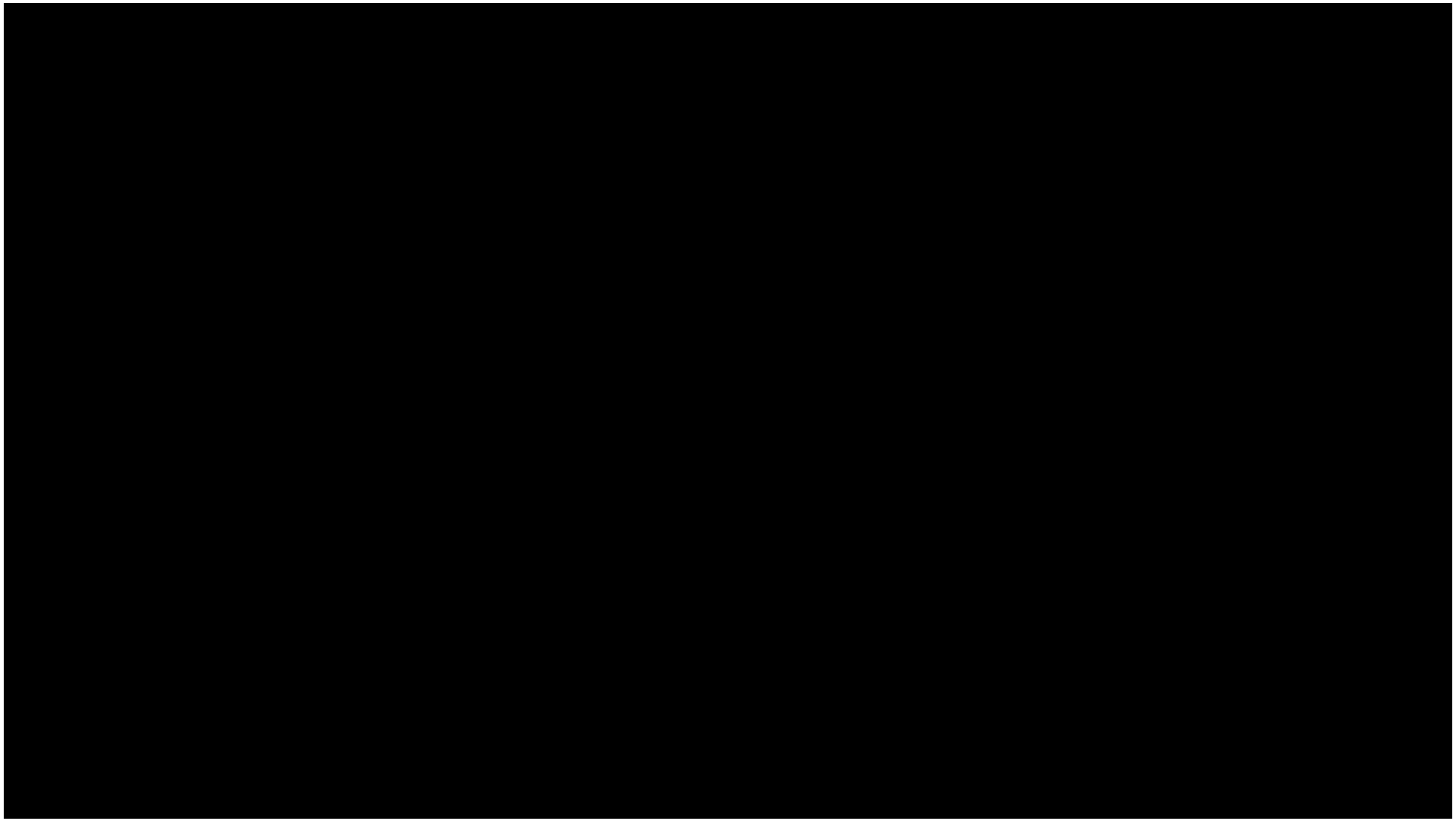
https://youtu.be/1vXGrTDN8sc

WEIGHING BY DIFFERENCE

Weighing by difference

- This time do NOT hit Tare on the scale.
- This time weigh the object/container BEFORE you do something (like add or remove chemicals)
- Then weight the object/container AFTER you do something.
- Subtract the two numbers to find how much you added or removed.

Weighing by difference

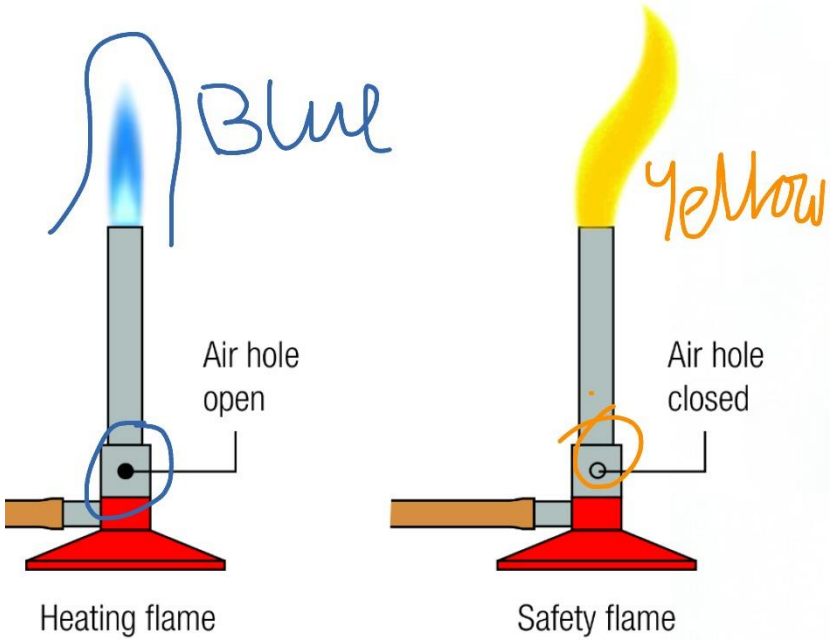


https://youtu.be/1vXGrTDN8sc

LIGHTING A BUNSEN BURNER

Lighting a Bunsen Burner

- Connect the gas hose to the valve on the lab station.
- Make sure the air regulator on the bottom of the Bunsen burner is open a little bit but not all the way.
- Get your match or striker ready
- Turn on the gas
- Bring the match or striker close to the Bunsen burner
  - Do NOT have your hand up above the Bunsen burner…the flame goes up!
    Have your hand approaching from down below.
- Gradually open the air regulator so that the flame will turn from a yellow to more blue.
  - A blue flame is hotter.

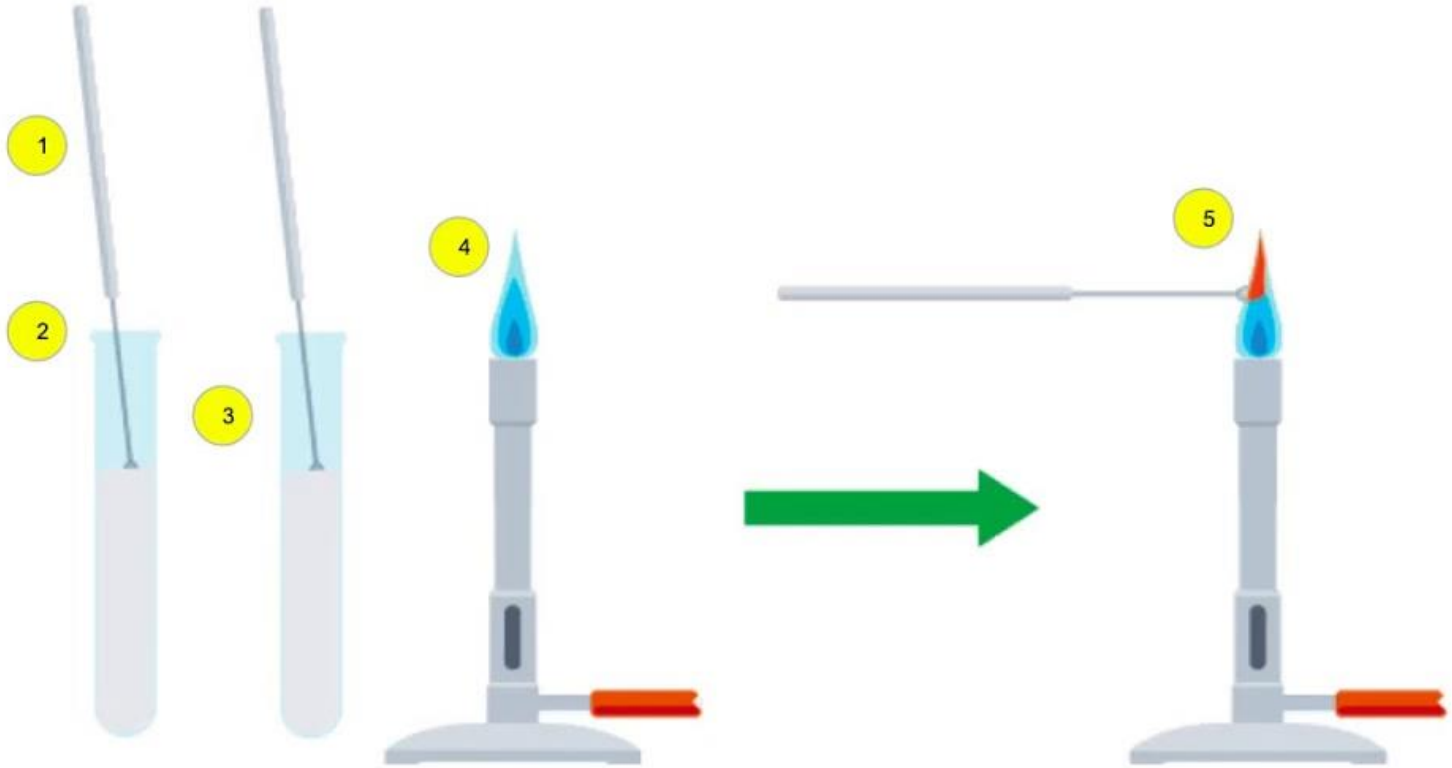


Lighting a Bunsen Burner

https://youtu.be/N7ssCM3qM3U

FLAME TESTS

Flame Tests

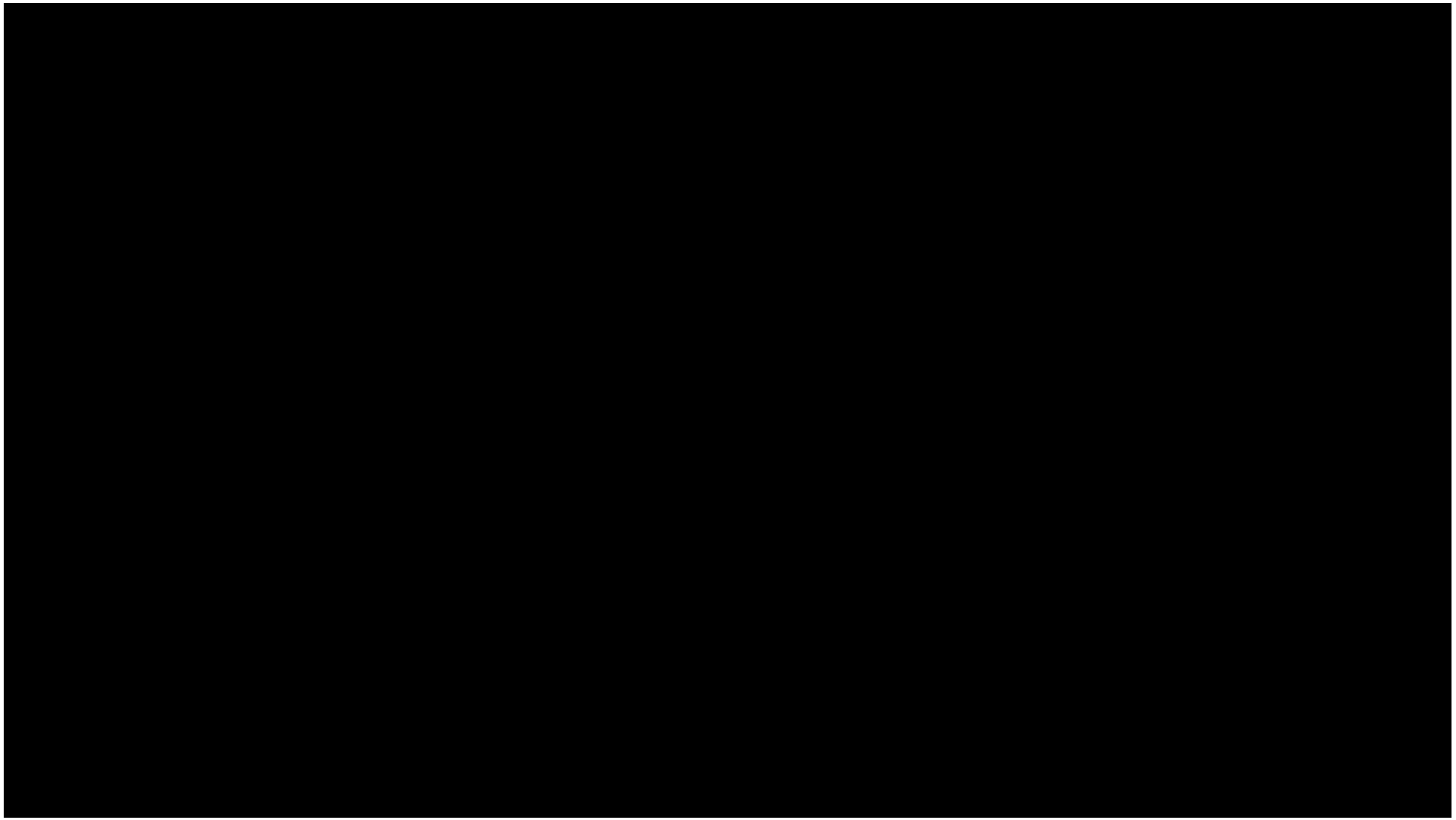


1. Platinum flame test wire - because Pt does not give a colour in a flame
2. Dip in conc. HCl to clean all existing ions off the wire
3. Dip in solid or in a solution of the unknown ion
4. Use blue flame which allows any colours to be seen
5. Place flame test wire in blue flame and check for colour produced

***Important to use a blue flame on the Bunsen burner because it is hotter than an orange/yellow flame!**

Images from: http://www.bbc.co.uk/staticarchive/a92cd29b94e92063b88c2aace62ec09de509bde5.gif, https://images.fineartamerica.com/images-medium-large/1-lithium-flame-test-.jpg, https://farm1.staticflickr.com/8/8633272_1ae3659a40.jpg, https://www.thoughtco.com/thmb/oGdwnydAeD3e2e4RdEB6vIS10BI=/3605x2764/filters:no_upscale():max_bytes(150000):strip_icc()/136820595-56a1338b5f9b58b7d0bcfcab.jpg, https://www.flinnsci.ca/globalassets/flinn-scientific/all-product-images-rgb-jpegs/ap5607_2.jpg?v=a8caa55374b04fb1b4e0000866490af7; https://render.fineartamerica.com/images/rendered/default/print/8.000/8.000/break/images-medium-5/1-copper-flame-test-science-photo-library.jpg

Flame Tests

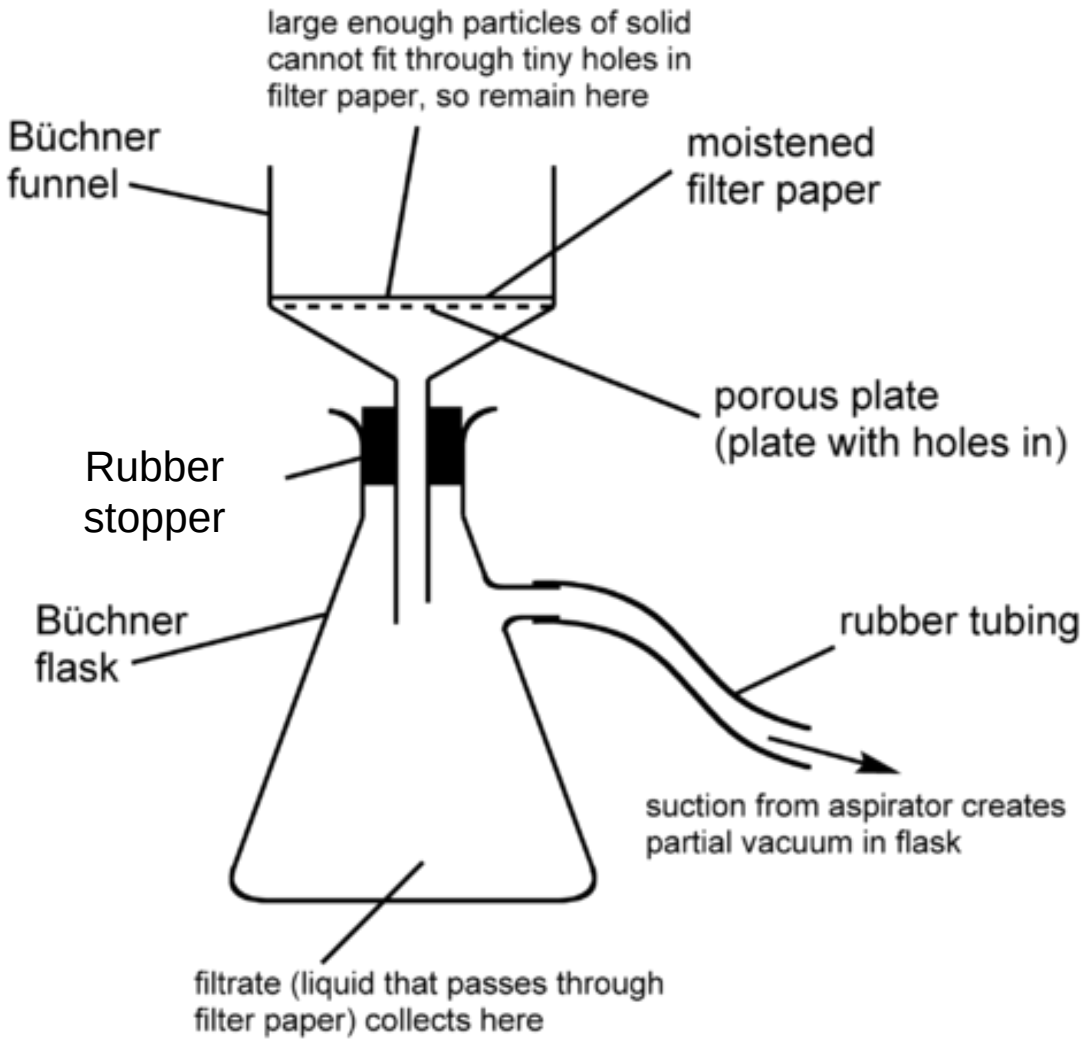


https://youtu.be/1EXr_L7Ojqg

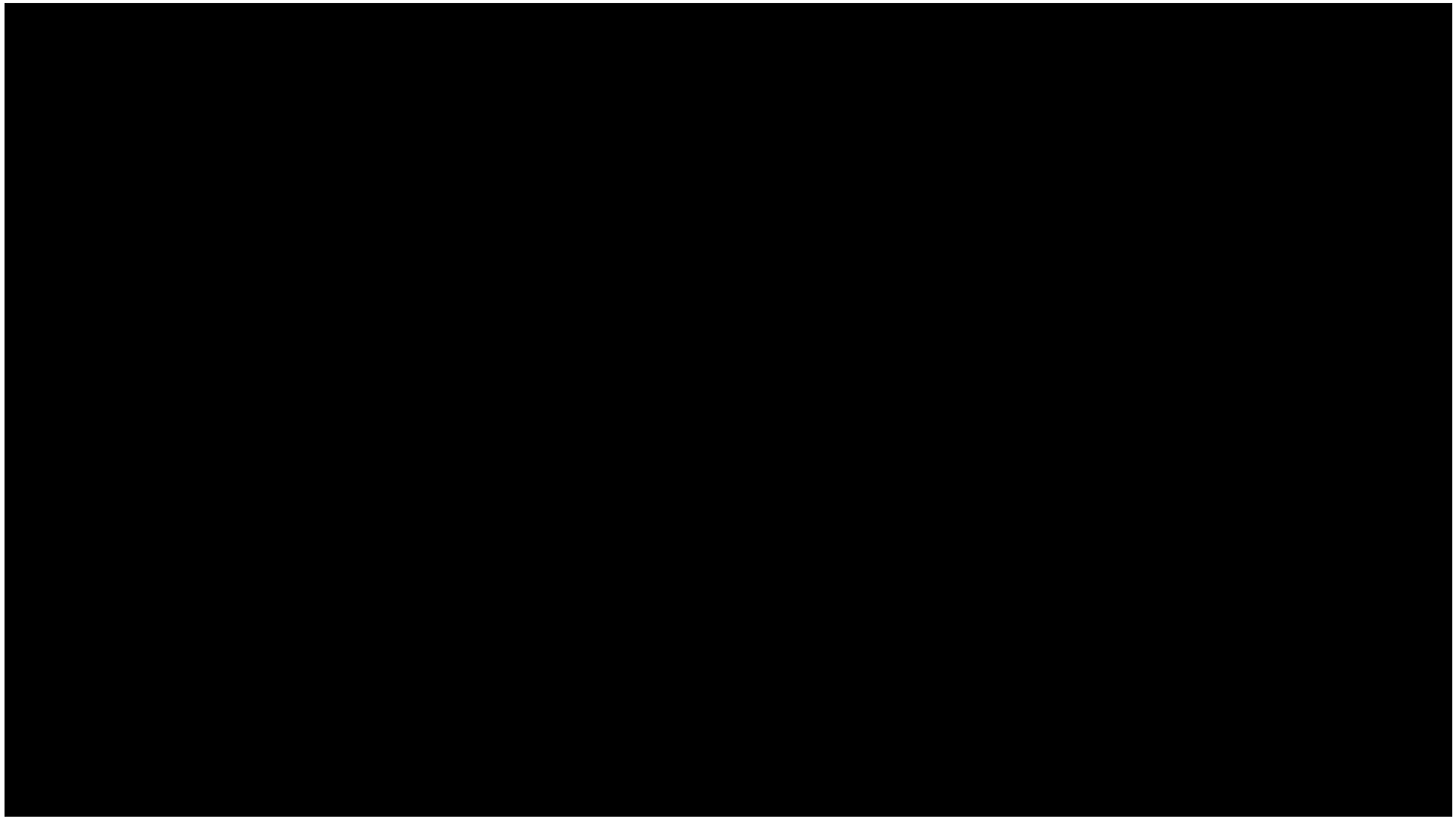
VACUUM FILTRATION

Vacuum Filtration

- Vacuum filtration helps us filter a solid out from a liquid a lot faster than using just a normal funnel.
- We attach the funnel to a Buchner flask which attaches to a hose, and the hose is attached to an aspirator on the sink faucet.
- When you turn the water to the sink on, the water flows through the aspirator into the sink. This causes suction and a partial vacuum inside the flask.
- The pressure in the flask ends up lower than in the room and so the higher air pressure helps push the liquid down through the filter paper into the flask.



Vacuum Filtration



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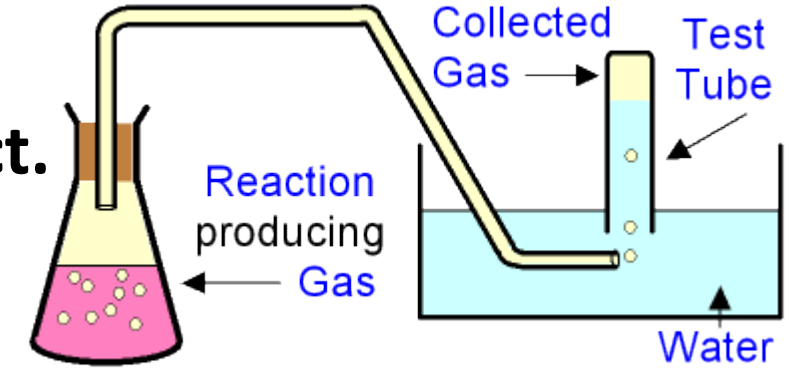
COLLECTING GAS OVER WATER

Collecting Gas Over Water



- When you collect gas over water there is water vapor mixed in with the gas you are trying to collect.
- You need to subtract out the pressure that is coming from the water vapor so you can find the pressure that is just from the gas you are interested in.
- You find the water vapor pressure from a chart. Take the temperature of the water and you find that row on the chart.

- Gas + water vapor = "wet gas"
- Gas with no water vapor = "dry gas"

- $P_{dry\ gas} = P_{wet\ gas} - P_{H2O\ Vapor}$

Water Vapor Pressure at Various Temperatures

Calculator for Water Vapor at Various Temperatures:
https://tinyurl.com/watervaporpressure

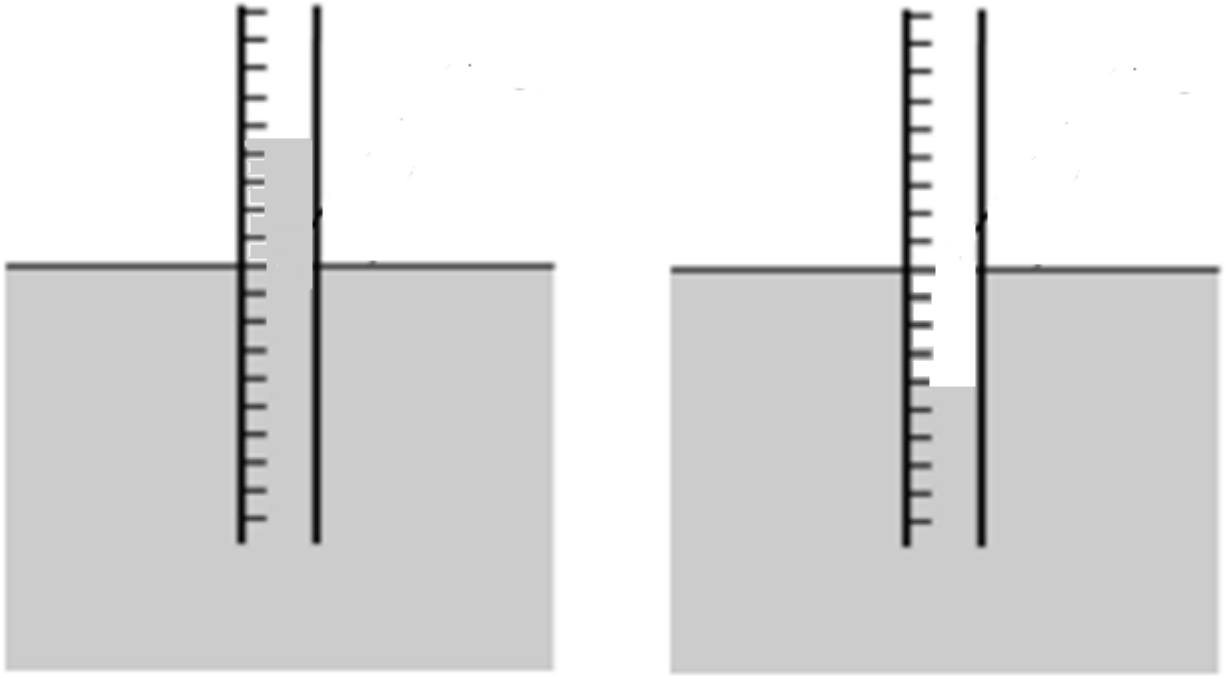
Walkthrough about how to collect a gas over water:
https://tinyurl.com/collectinggasoverwater

T (°C)	P (mmHg)	T (°C)	P (mmHg)	T (°C)	P (mmHg)	T (°C)	P (mmHg)
1	4.8853	26	25.1370	51	96.9771	76	300.7990
2	5.2511	27	26.6642	52	101.8561	77	313.5406
3	5.6409	28	28.2715	53	106.9439	78	326.7347
4	6.0559	29	29.9623	54	112.2477	79	340.3943
5	6.4975	30	31.7402	55	117.7751	80	354.5323
6	6.9673	31	33.6089	56	123.5336	81	369.1619
7	7.4667	32	35.5723	57	129.5310	82	384.2966
8	7.9973	33	37.6344	58	135.7755	83	399.9502
9	8.5608	34	39.7993	59	142.2751	84	416.1368
10	9.1588	35	42.0711	60	149.0384	85	432.8706
11	9.7932	36	44.4543	61	156.0740	86	450.1661
12	10.4659	37	46.9533	62	163.3906	87	468.0383
13	11.1787	38	49.5729	63	170.9974	88	486.5021
14	11.9337	39	52.3178	64	178.9036	89	505.5729
15	12.7330	40	55.1928	65	187.1186	90	525.2664
16	13.5787	41	58.2032	66	195.6521	91	545.59.85
17	14.4732	42	61.3541	67	204.5142	92	566.5854
18	15.4189	43	64.6509	68	213.7147	93	588.2434
19	16.4180	44	68.0992	69	223.2643	94	610.5894
20	17.4733	45	71.7046	70	233.1733	95	633.6405
21	18.5872	46	75.4730	71	243.4526	96	657.4138
22	19.7626	47	79.4105	72	254.1137	97	681.9270
23	21.0023	48	83.5232	73	265.1667	98	707.1980
24	22.3092	49	87.8175	74	276.6242	99	733.2450
25	23.684	50	92.2999	75	288.4977	100	764.2602

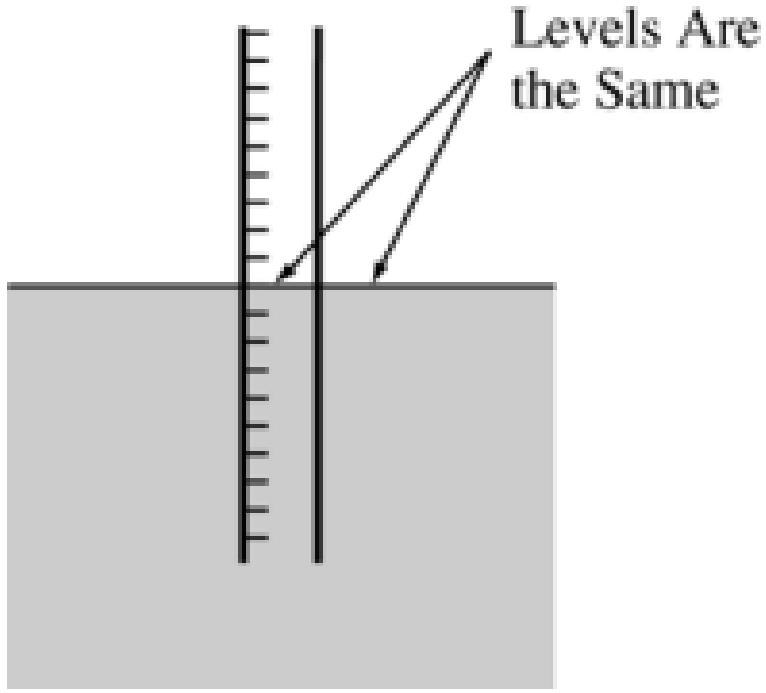
Collecting Gas Over Water

- If you line up the water level line of the collection container, with the water level line of the water bath, then the pressure inside the collection container will be the SAME as the pressure in the room!

Water lines don't match, pressure inside not the same as pressure outside!



Water lines match, pressure inside is the same as pressure outside!



Collecting Gas Over Water

https://youtu.be/_C4tjmekAs0

*you should use a deep enough water basin so you can make the water line in the cylinder line up with the water in the basin – this makes the pressure inside and outside the cylinder equal. Important if you are doing calculations!

Collecting Gas Over Water

Common Mistakes

- **Not lining up water lines**
  - Pressure inside tube will not match the atmospheric pressure in the room
- **Using a soluble gas**
  - Results in less gas being collected since some will be dissolved in the water
- **Forgetting to take the temperature of the water bath**
  - Won’t have the temperature to plug into calculations
- **Not subtracting the pressure of the water vapor**
  - Results in pressure of desired gass *appearing* too high

Collecting Gas Over Water

Common Applications

- **Collecting gases that form in reactions like**
  $Mg + 2HCl \rightarrow MgCl_2 + H_2$ (g)
  $2KClO_3 \rightarrow 2KCl + 3O_2$ (g)

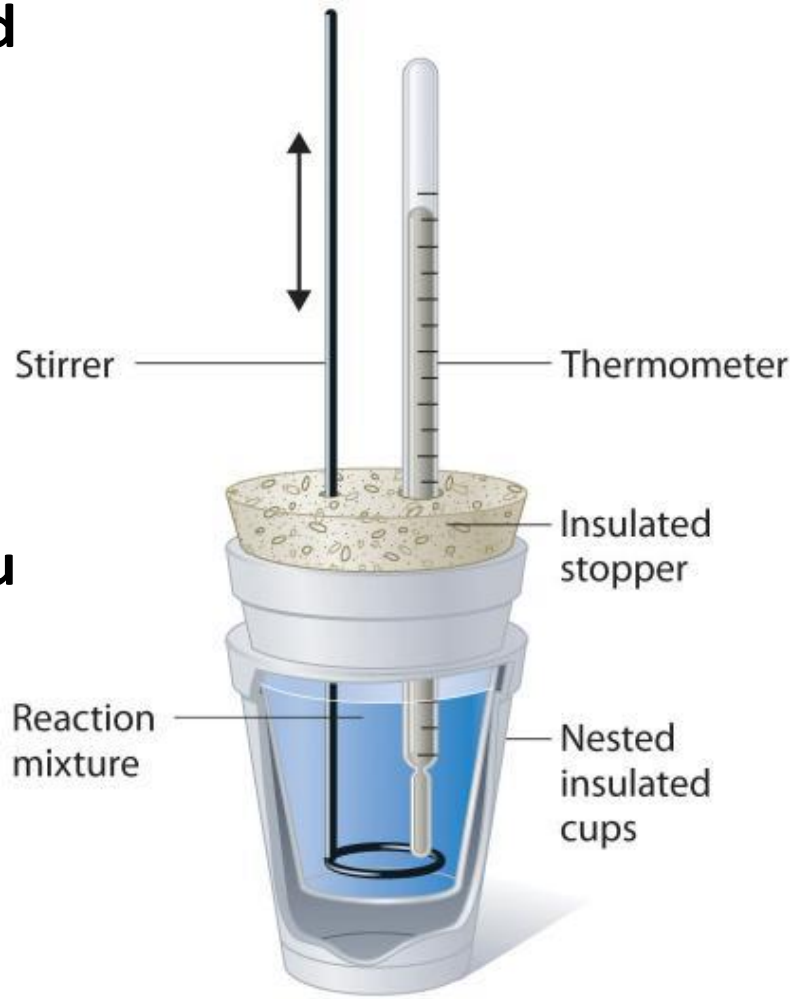
Important to Remember

- **Take the temperature of the bath to get the temperature of the gas.**
- **Use room temperature water for the bath.**
  - Gas solubility is minimized if you do not use cold water.
- **The volume of the gas must be read where the volume inside the eudiometer is at the same level as the water outside the bath.**
  - Allows the pressure inside to be equal to the atmospheric pressure
- **Pressure of atmosphere = Pressure of gas + Pressure of water vapor**

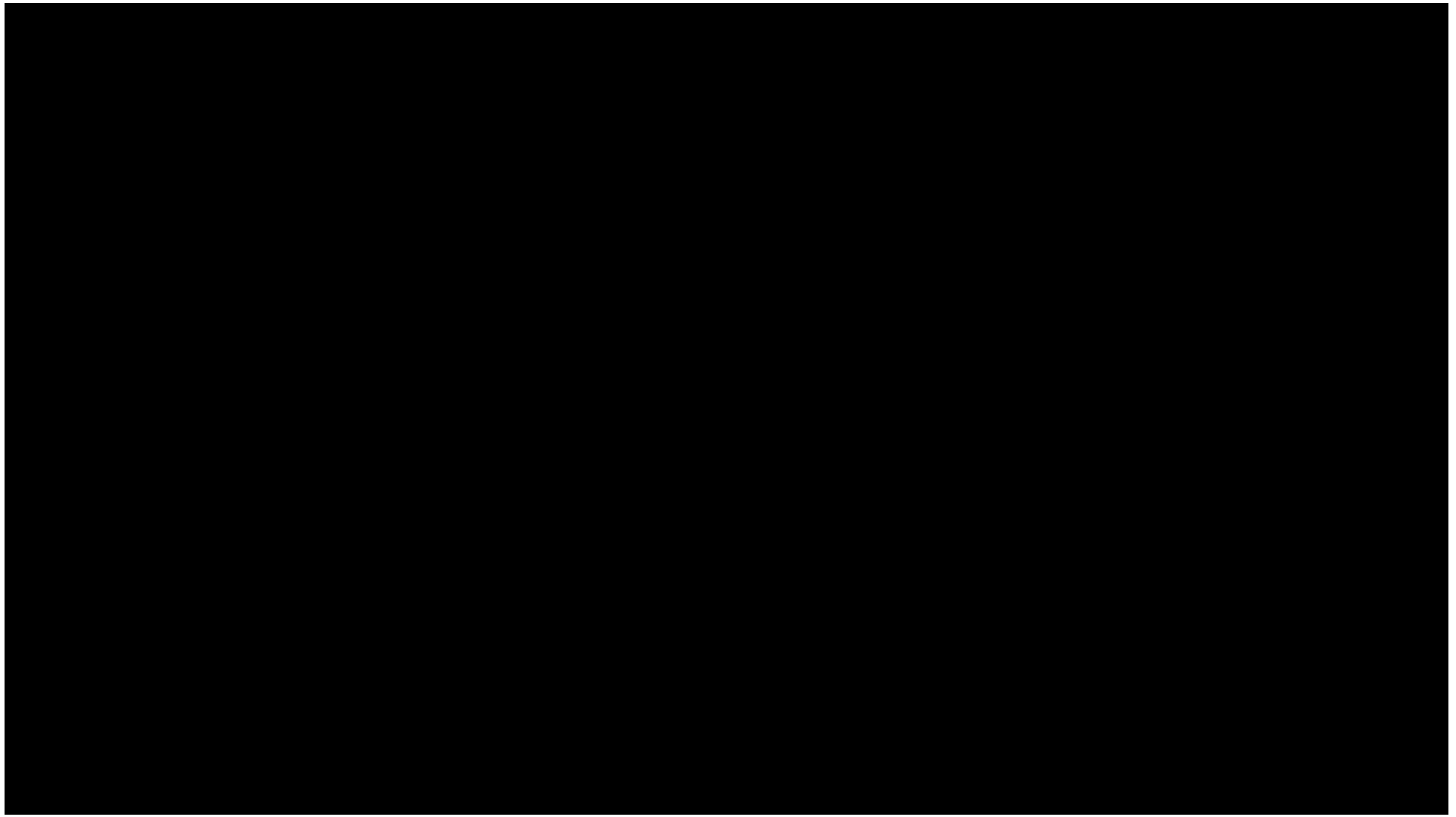
CALORIMETRY

Calorimetry

- Calorimetry exploits the fact that energy cannot be created or destroyed, it is only transferred. So if one object cools down, something else must have absorbed that energy.
- We say $Q_{absorbed} = -Q_{released}$
  Same magnitude but opposite in sign because one is absorbing and one is releasing.
- If you put your object/substance in water, you can easily measure the change in temp of the water which allows you to calculate the energy the water absorbed or released.
- Once you know how much energy the water absorbed or released you now know how much the object/substance released to the water or absorbed from the water.
- We use insulated containers so the water traps as much of the energy as possible, we don't want to lose energy to the surroundings!



Calorimetry



https://youtu.be/-VQEeqLVpGo

Calorimetry

Common Mistakes

- **Mixing up T_{final} and $T_{initial}$**
  - The final temperature is the highest (for exothermic) or lowest (for endothermic) temperature recorded during the rxn/process
- **Not stirring enough** - hotter/colder in some parts of solution
- **Endothermic reaction: temperature doesn't change enough**
  - Heat was absorbed by reaction from calorimeter/surroundings
  - Lid not sealed tightly on calorimeter
- **Exothermic reaction: temperature doesn't change enough**
  - Heat absorbed by calorimeter or lost to surroundings
  - Lid not sealed tightly on calorimeter

Common Applications

- **Mixtures of solids – what % of the mixture/alloy is a specific substance**
- **Determining the amount of a particular ion in a solution**

Calorimetry

Common Applications

- **Solving for the specific heat of a metal or the heat of reaction**

Important to Remember

- **Endothermic processes** have a decrease in temperature.
- **Exothermic processes** have an increase in temperature.
- The water is not part of the system. It is part of the surroundings.
- **q = mCΔT**
  - q = heat in Joules or calories
  - m = mass of entire solution (reactants + water) OR object, grams or kg
  - C = specific heat capacity, J/goC (or a variation of the above)
  - ΔT = Tfinal – Tinitial
- To calculate heat of solution: q/moles of salt
- To calculate heat of reaction: ----------> $\frac{q}{mol\ reactant\ used} = \frac{\Delta Hrxn}{coefficient\ from\ equation}$

USING A VOLUMETRIC FLASK

Using a Volumetric Flask

Don't forget – you want the bottom of the meniscus to be at the line mark!

(a) An amount of solute is weighed out on an analytical balance and then transferred to a volumetric flask.

(b) A portion of the solvent is added to the volumetric flask.

(c) The mixture is swirled until *all* of the solute is dissolved.

(d) Additional solvent is added up to the mark on the volumetric flask.

Using a Volumetric Flask

https://youtu.be/hrvXuX0Ow3s

MAKING A SOLUTION

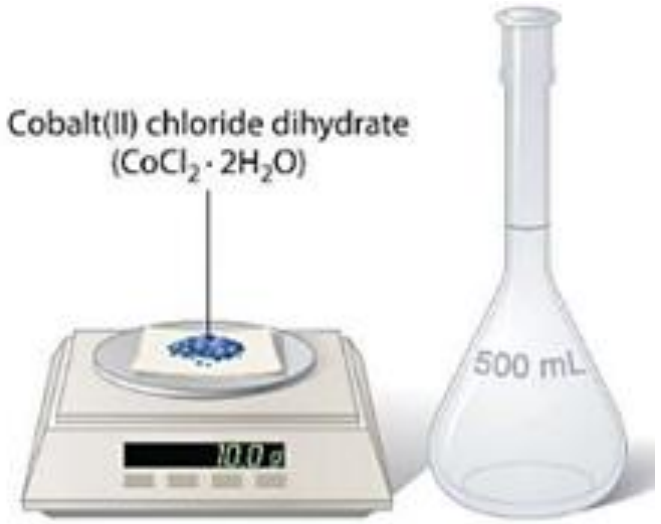
Making a solution

$$\text{Molarity} = \frac{moles\ solute}{liters\ of\ solution}$$

- We use molarity as our unit for concentration
- Our scales do not measure in moles – we need to measure a certain # of grams on our scale and convert that to moles.
- We need to make sure we are using the # of Liters of SOLUTION not just the water we added – the solute takes up some space!

- If you know the molarity you want and the volume you want:
  - Solve for moles solute needed.
  - Then convert moles to grams using molar mass.
  - Weigh that many grams out on your scale.
  - Add the solute to your container.
  - Fill with water until you reach the desired volume of solution

Making a solution



(a) An amount of solute is weighed out on an analytical balance and then

(b) The solid is added to a beaker along with a portion of the solvent. (If diluting an acid make sure to add the acid to the water and not the other way around! Also make sure you are cautious – if it is exothermic the container can get hot. Make sure to swirl to ensure all the solid is dissolved.

(c) Carefully pour the dissolved mixture into a volumetric flask, using a funnel.

The AP Readers would prefer that you use a pipet at this point to make sure you hit the line perfectly.

(d) Additional solvent is added up to the mark on the volumetric flask.

(e) Put a stopper in the flask and invert several times to ensure it is well mixed. Pour into a proper storage bottle.

Making a solution

Common Mistakes

- **Solid gets stuck in the neck of the flask**
  - Use a beaker to dissolve solute in some solvent, then transfer to volumetric flask
- **Overfilling the volumetric flask**
  - Results in a dilute solution
- **Not using distilled water**
  - Other ions present in the tap water could affect the experiment the solution is eventually used for.
- **Not using a volumetric flask (using a beaker or Erlenmeyer instead)**
  - Loss of precision in concentration of prepared solution

Making a solution

Common Applications

- Making solutions to dissolve substances for analysis, particularly titrations

Important to Remember

- Molarity = (moles solute) / (L solution)

Making a solution

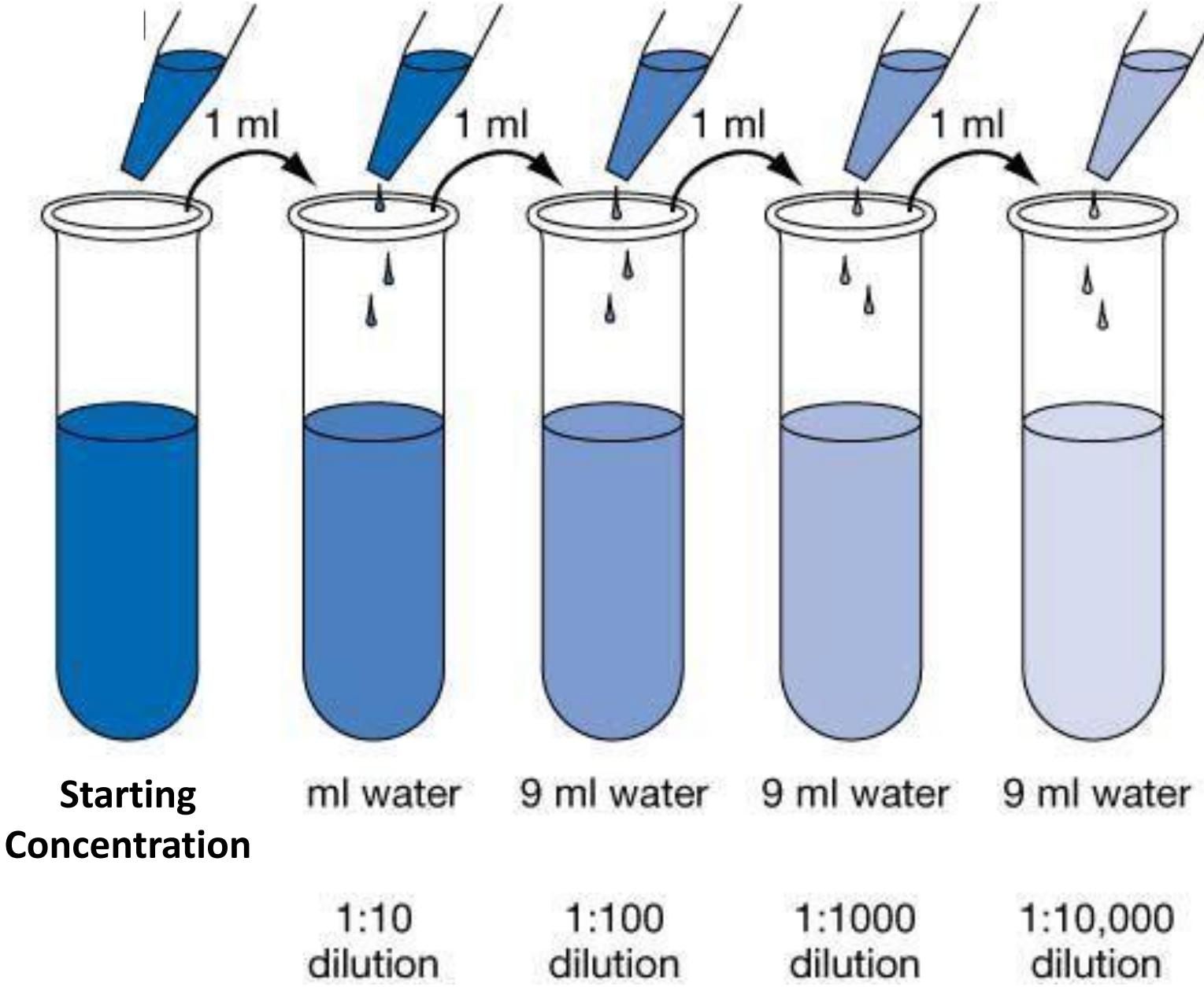
https://youtu.be/A2Yylo8vSCA

SERIAL DILUTION

Serial Dilution

- Serial dilutions do not need to be 10 mL each time, this is just an example image.
- The idea is that you take a small amount of the strong stuff and water it down.
- Then you take some of that new solution and water it down again.
- Etc etc etc…

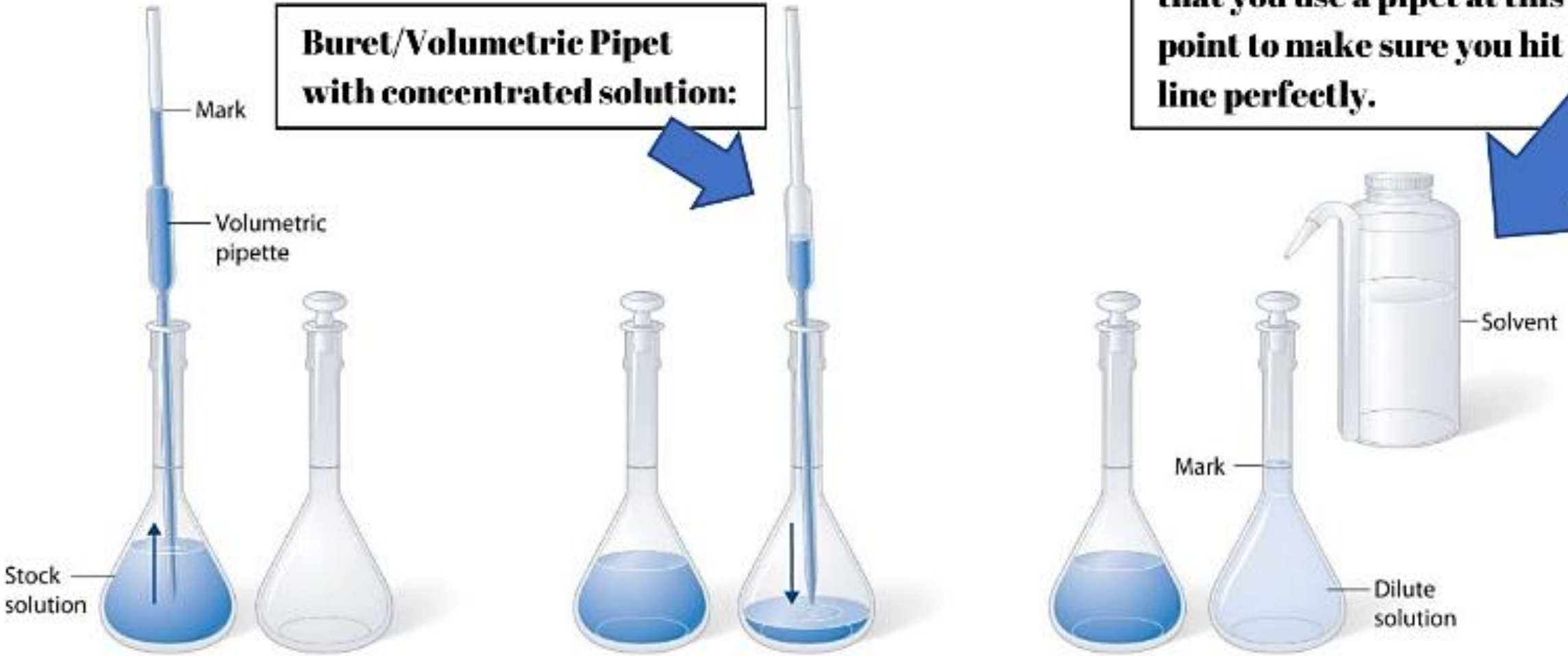
- $M_1V_1 = M_2V_2$



Serial Dilution

Buret/Volumetric Pipet with concentrated solution:

The AP Readers would prefer that you use a pipet at this point to make sure you hit the line perfectly.



(a) A volume (V_s) containing the desired moles of solute (M_s) is measured from a stock solution of known concentration.

(b) The measured volume of stock solution is transferred to a second volumetric flask.

(c) The measured volume in the second flask is then diluted with solvent up to the volumetric mark [$(V_s)(M_s) = (V_d)(M_d)$].

Serial Dilution

https://youtu.be/A2Yylo8vSCA

***you don't have to do a 1mL + 9mL dilution but it is a very classic ratio to use.**

Serial Dilution

Common Mistakes

- **Not adding the acid into the water (adding in reverse order)**
  - Solution can bubble up, steam from heat released, splattering could occur.
- **Overfilling the volumetric flask**
  - Results in a dilute solution
- **Not using distilled water.**
  - Other ions could affect the experiment for which the solution is used
- **Not using a volumetric flask (beaker or Erlenmeyer instead)**
  - Loss of precision in concentration of prepared solution

Common Applications

- **Making solutions to dissolve substances for analysis, particularly in titrations.**

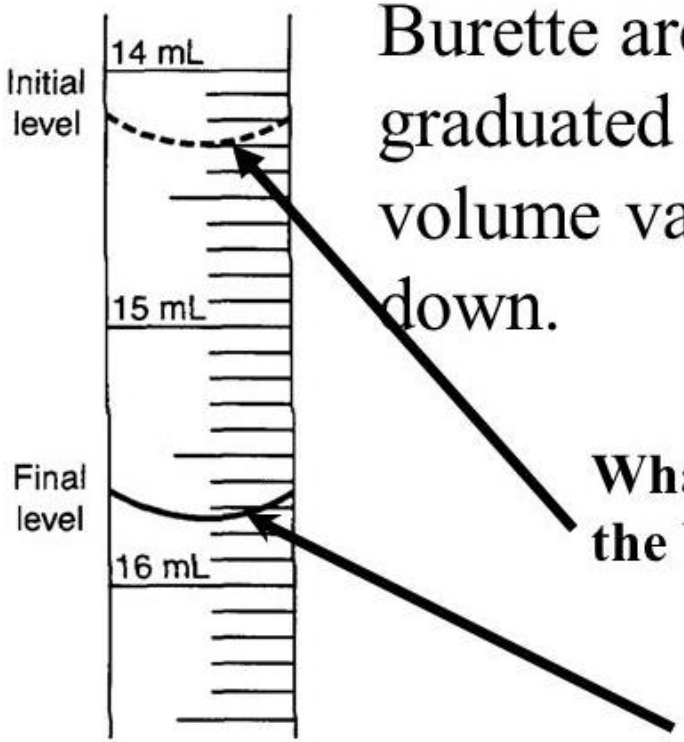
Important to Remember

- **Molarity = moles solute/L of solution**

SETTING UP AND READING A BURETTE

Setting up and Reading a Burette

- Remember to estimate one decimal place past where the tick line marks are on the burette, just like on a graduated cylinder!



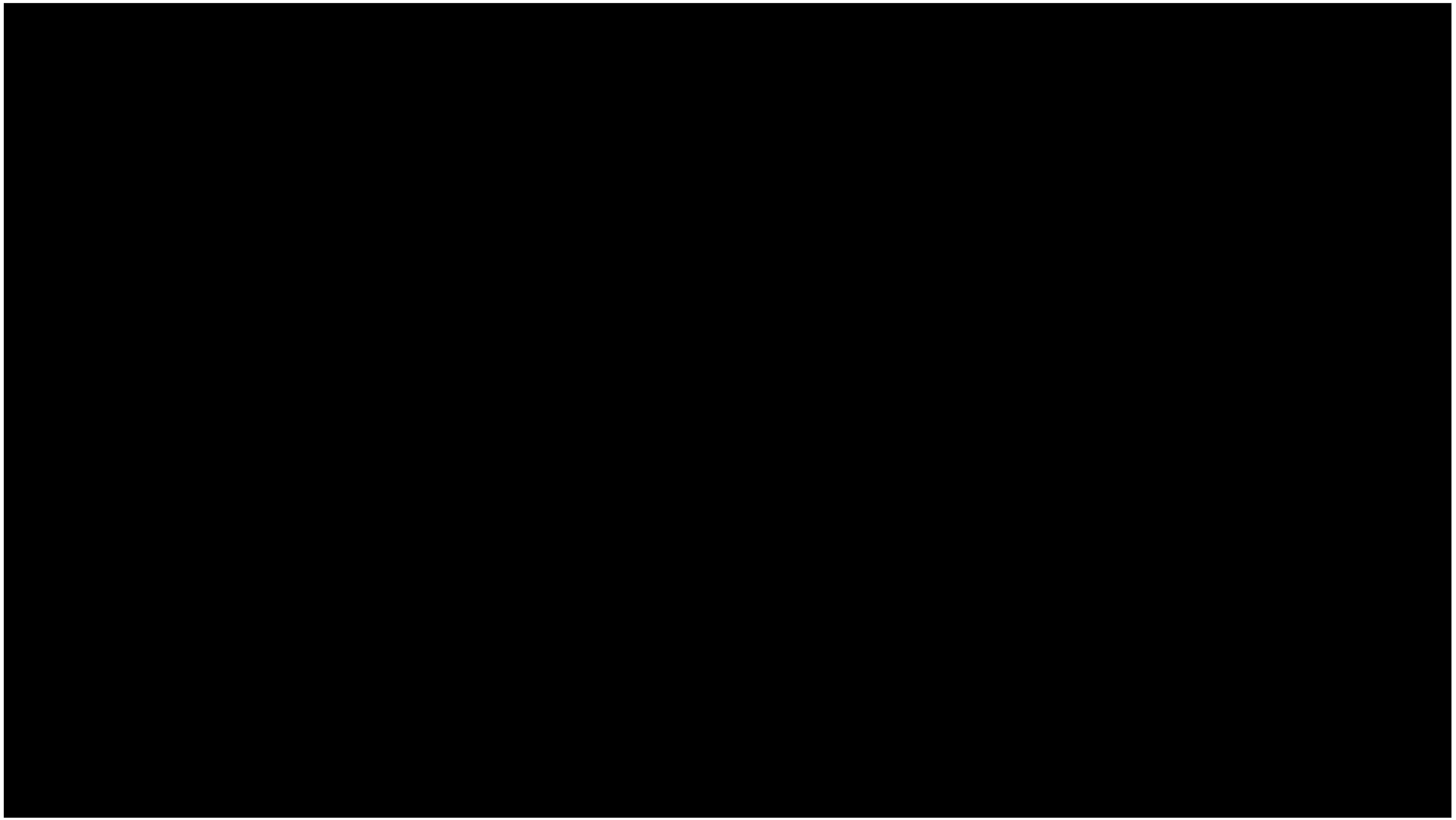
Burette are read the same way as a graduated cylinder except the volume value increases as you go down.

What is the value on the burette? 14.30

What is the value on the burette? 15.74

Final reading – Initial reading = volume that was dispensed

Setting up and Reading a Burette



Video #1

https://youtu.be/Lr1nLTCqZvM

Setting up and Reading a Burette

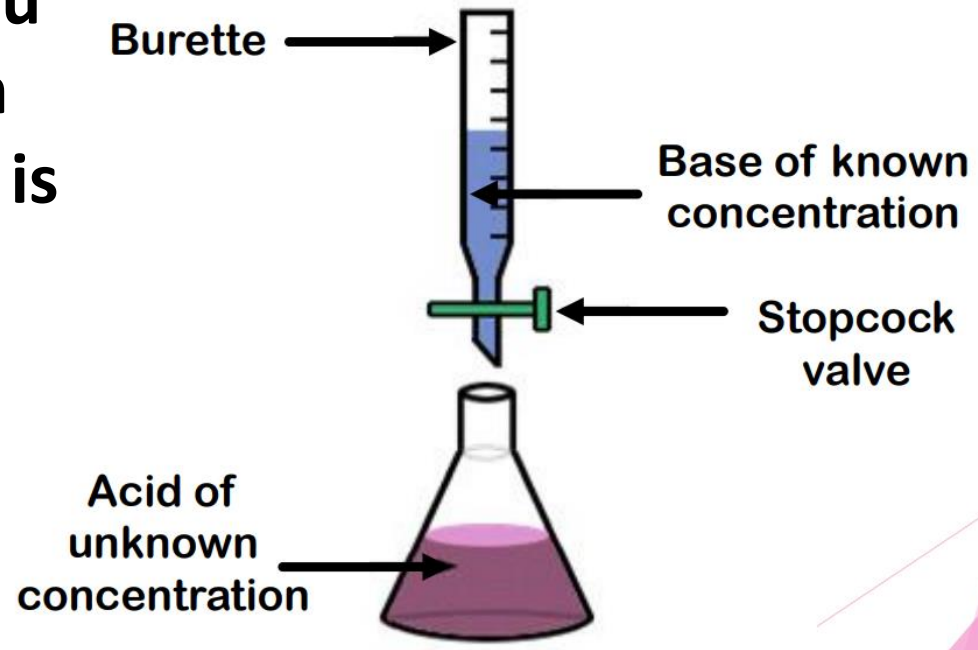
Video #2

https://youtu.be/qdmp4_Nwd-Q

PERFORMING A TITRATION

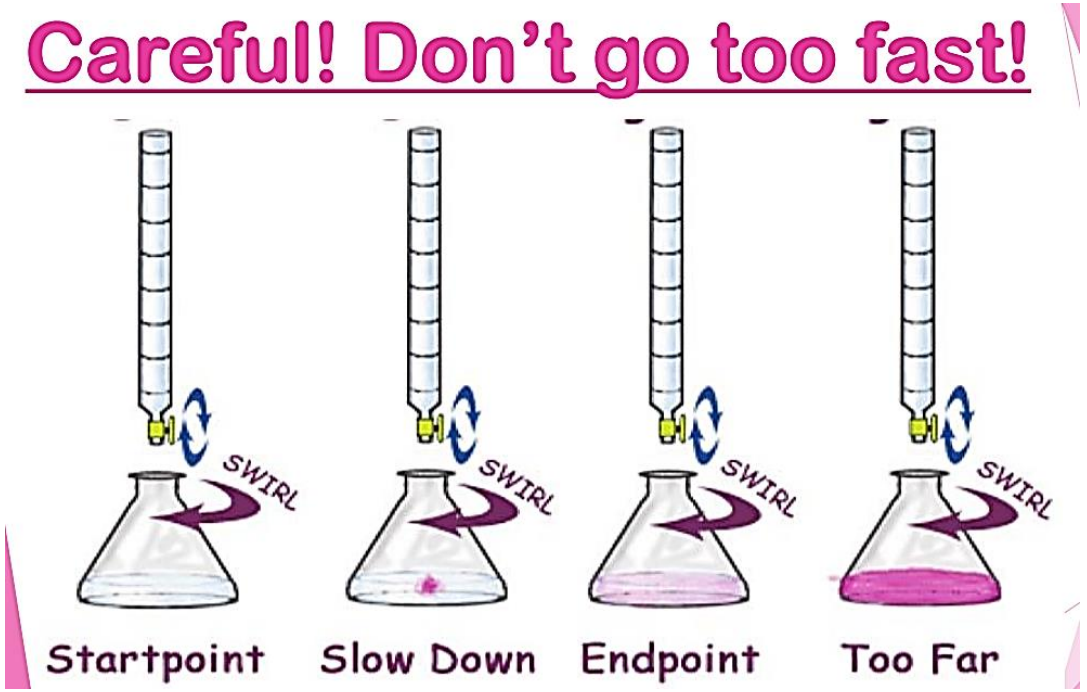
Preforming a Titration

- A titration is used to help determine the unknown concentration of a substance, usually an acid or a base.
- If you have an unknown acid concentration you slowly add a base with a known concentration until you reach the “equivalence point” which is when the moles of acid = moles of base.
- You can then use stoichiometry to help you convert from the known moles of base added into the unknown concentration of the acid.
- Or vice versa if you have an unknown base concentration and a known acid concentration.

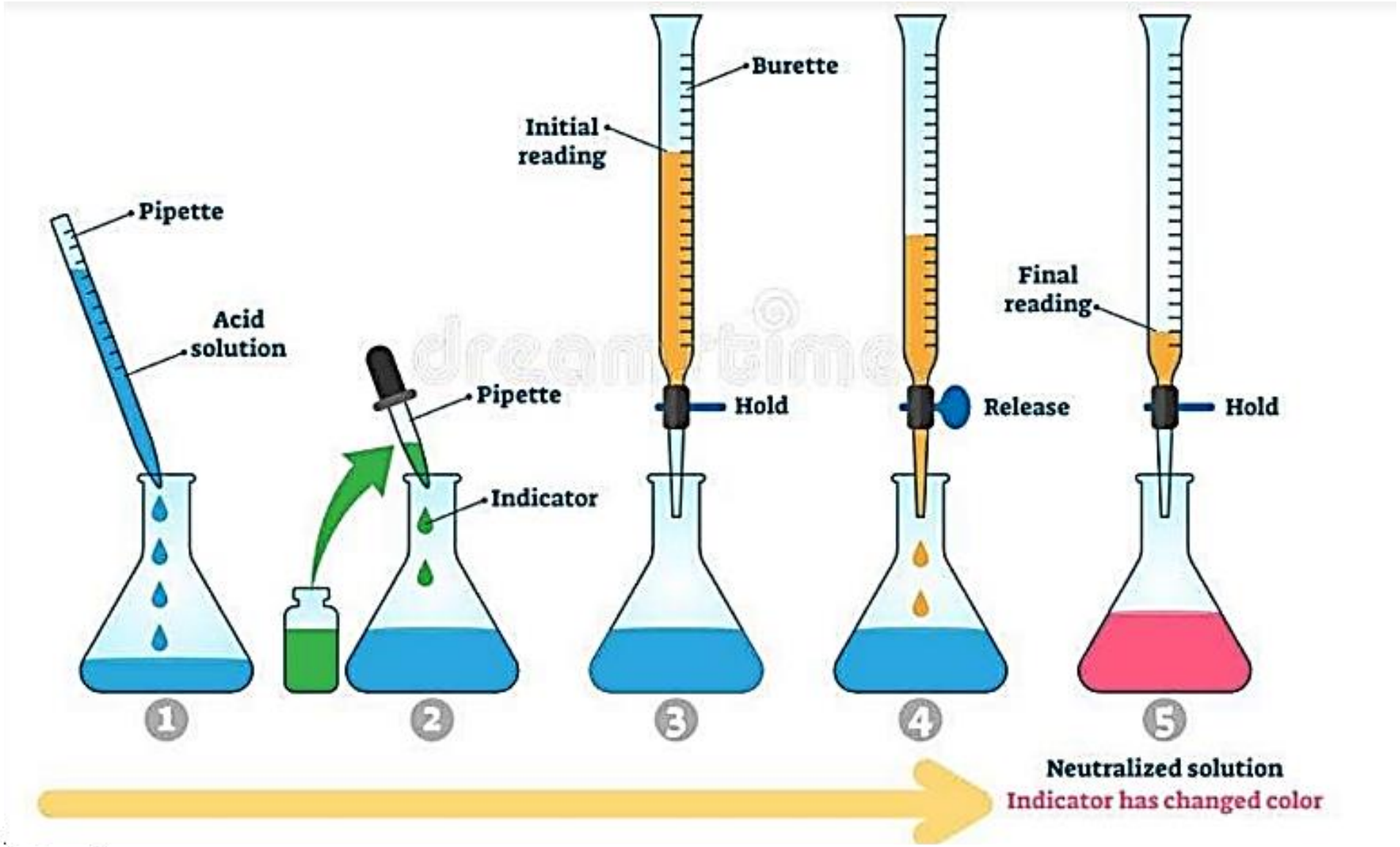


Preforming a Titration

- You often use an “indicator” to determine when you have reached the equivalence point.
- If you pick the right indicator, it will turn colors when the
  moles acid = moles of base.
- That tells you when to stop adding your known concentration acid or base



Preforming a Titration



Preforming a Titration

Video #1

https://youtu.be/tIbD8MG1qMM

Preforming a Titration

Video #2

https://youtu.be/sFpFCPTDv2w

Preforming a Titration

Video #3

https://youtu.be/YqfvRBJ-iPg

Performing a Titration

Common Mistakes

- **Overshooting the titration (too dark of a color at the end)**
  - Results in the concentration of the unknown solution in the flask *appearing* to be higher than it actually is, since more titrant was added.
- **Not using indicator. -** No perceivable endpoint.
- **Using incorrect indicator. -** pH at the equivalence point should be approximately equal to the pKa of the indicator.
- **Cleaning and preparing the buret incorrectly.**
  - Rinse buret with distilled water, add a small amount of titrant to buret, swirl, and let it out through the stem. Repeat again with more titrant. Be sure to rinse with titrant twice.
  - Consequence of improper cleaning = a titrant that is more dilute, which will result in an analyte that *appears* to be more concentrated than it is
- **Reading buret incorrectly**
  - It should be read from the bottom of the meniscus. If on the line, add a 0, if inbetween, estimate the final digit

Performing a Titration

Common Applications

- **Solving for the concentration of an unknown substance (analyte).**
- **Acid/Base, Redox**

Important to Remember

- **Molarity = moles solute/L of solution**
- **Analyte:** substance in flask
- **Titrant:** substance in buret
- **Standard solution:** sol'n of known concentration, usually goes into the buret.
- **M1V1 = M2V2** is helpful for solving for the concentration of the analyte solution at the equivalence point **(if the acid is monoprotic)**
- **For polyprotic acids** use stoich to determine concentration of unknown
- **Endpoint:** point in titration where flask solution changes color
- **Equivalence point:** point in the titration where the moles acid = moles of base

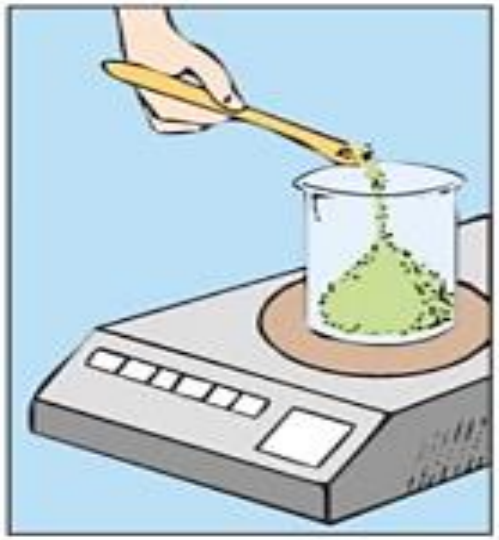
GRAVIMETRIC ANALYSIS

Gravimetric Analysis

**your ion of interest doesn't have to be the cation, this is just an example*

A technique to determine the quantity of an ion present by making it precipitate out of solution and then using stoichiometry to determine how much of the ion was originally present by comparing it to the mass of your new precipitated compound

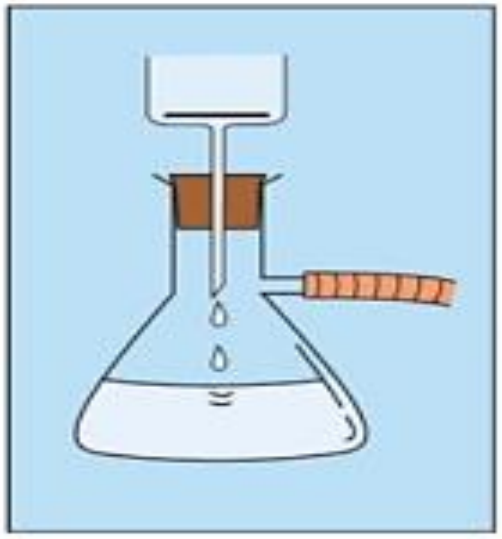
Weigh compound AX



Dissolve comp. AX in an appropriate solvent

Add compound BZ – pick one that will make an insoluble product with A

Filter out new solid product AZ. Rinse ppt to remove soluble ions and excess reagent.



Weigh AZ and do stoich to determine how much A ion you started with *

Gravimetric Analysis

https://youtu.be/kMeJj2YgwZw

Gravimetric Analysis

Common Mistakes

- **Precipitate is not dry when you take the final mass**
    - Results in the *appearance* of more precipitate than was actually produced because some mass is actually water

Common Applications

- **Mixtures of solids – what % of the mixture/alloy is a specific substance**
- **Determining the amount of a particular ion in a solution**

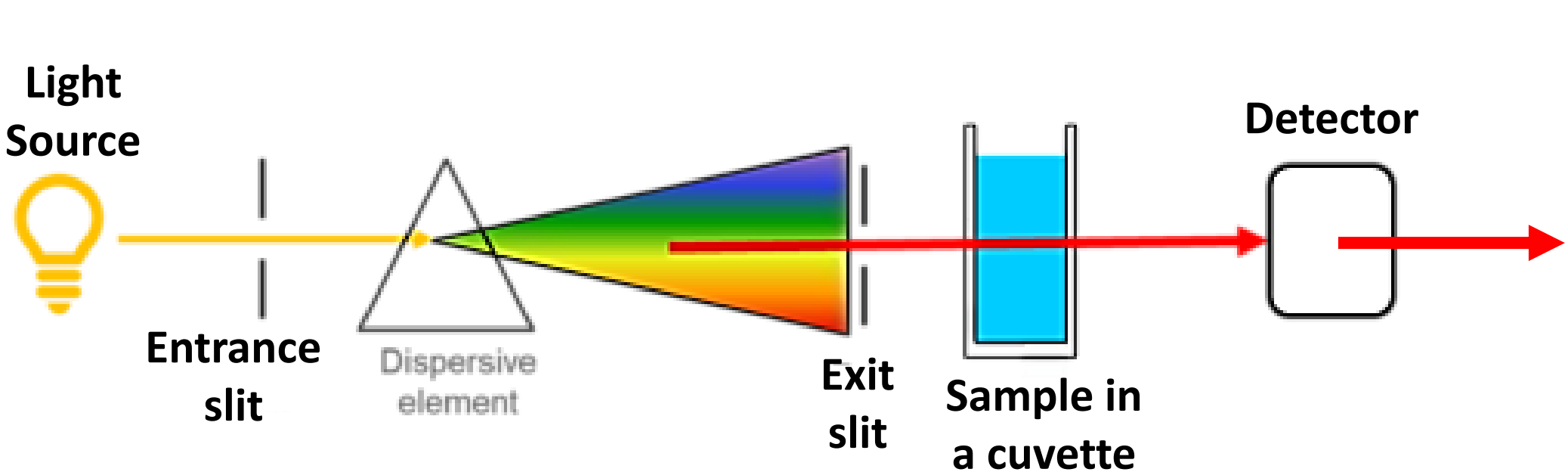
Important to Remember

- **All sodium, nitrate, ammonia, potassium compounds are soluble – SNAP!**
- **Net ionic equations should not include these spectator ions!**

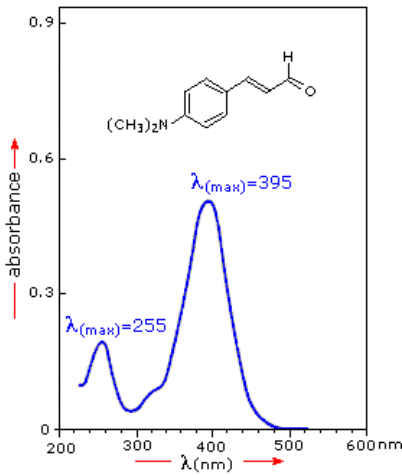
UV-VIS SPECTROSCOPY

UV-VIS-Spectroscopy

- **Different types of bonds absorb different wavelengths of light.**
- **If you pass light through a sample you can detect which wavelengths pass through the sample and which are absorbed.**
- **You can use the wavelengths of light absorbed/detected to determine the structure, or the concentration of the sample.**
- **In this class we will use Beer's Law to find concentrations using UV-Vis Spectroscopy.**
  Beers Law: A = εbC
  (A = absorbance, = Molar absorptivity L/mol*cm, b = path length of cuvette , C = concentration

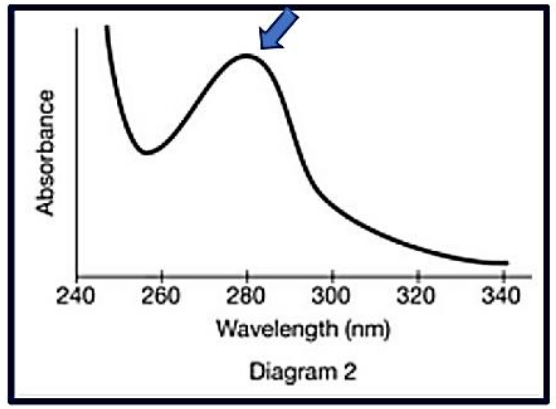


UV-Vis Spectrum

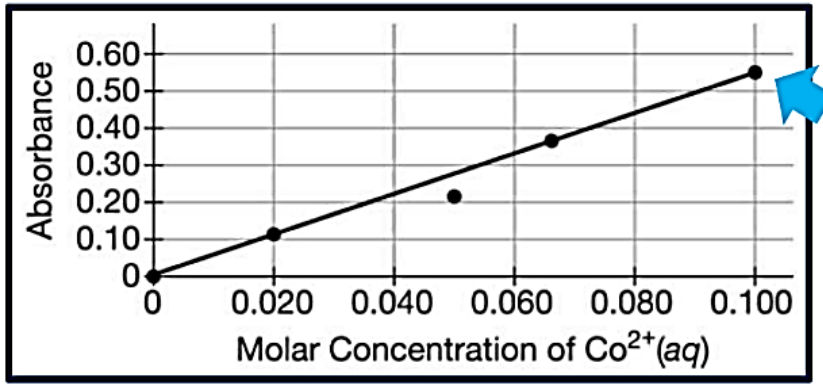


UV-VIS-Spectroscopy

1. **Pick the wavelength for the solution where absorbance is highest for the solute. Complementary colors are usually best.**

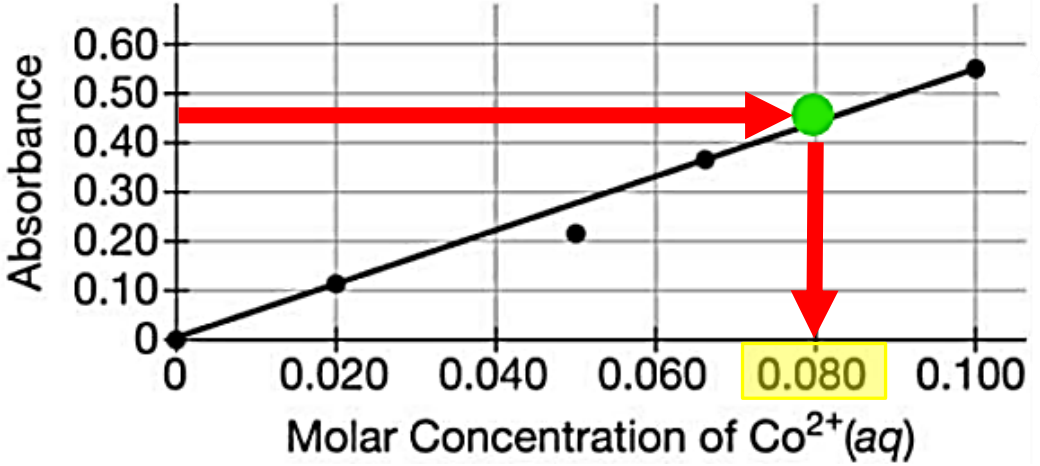


2. **Measure absorbance for different concentrations at that wavelength. Graph results. This is making your "standard curve."**



Concentration and Absorbance are directly proportional.

3. **Measure the absorbance for your unknown concentration solution. You can compare it's absorption to the concertation using your graph.**



UV-VIS-Spectroscopy

Video #1

https://youtu.be/-et7jDXOLB4

UV-VIS-Spectroscopy

Video #2

https://youtu.be/dSEWkypbhLk

UV-VIS Spectroscopy

Common Mistakes

- **Absorbance is lower than it should be (point falls below the line)**
  - Cuvette was cleaned with distilled water and then immediately filled with solution, creating a more dilute solution
  - Too little solute in the prepared solution
- **Absorbance is higher than it should be (point falls above the line)**
  - Cuvette is dirty with fingerprints/dust, etc.
  - Too much solute in the prepared solution
  - Contamination with a more concentrated solution
  - Used a cuvette with a longer path for one data point
  - Used frosted/ridged side of cuvette instead of the clear side
- **Did not use the correct wavelength of maximum absorbance for the solute.**
  - Absorbances could be too low especially for dilute solutions
- **Overfilled the cuvette -** Should not have an impact on data
- **Picked a wavelength where it is high absorbance for the solvent**
  - Won't be able to distinguish absorbance due to solvent vs. solute

UV-VIS Spectroscopy

Common Applications

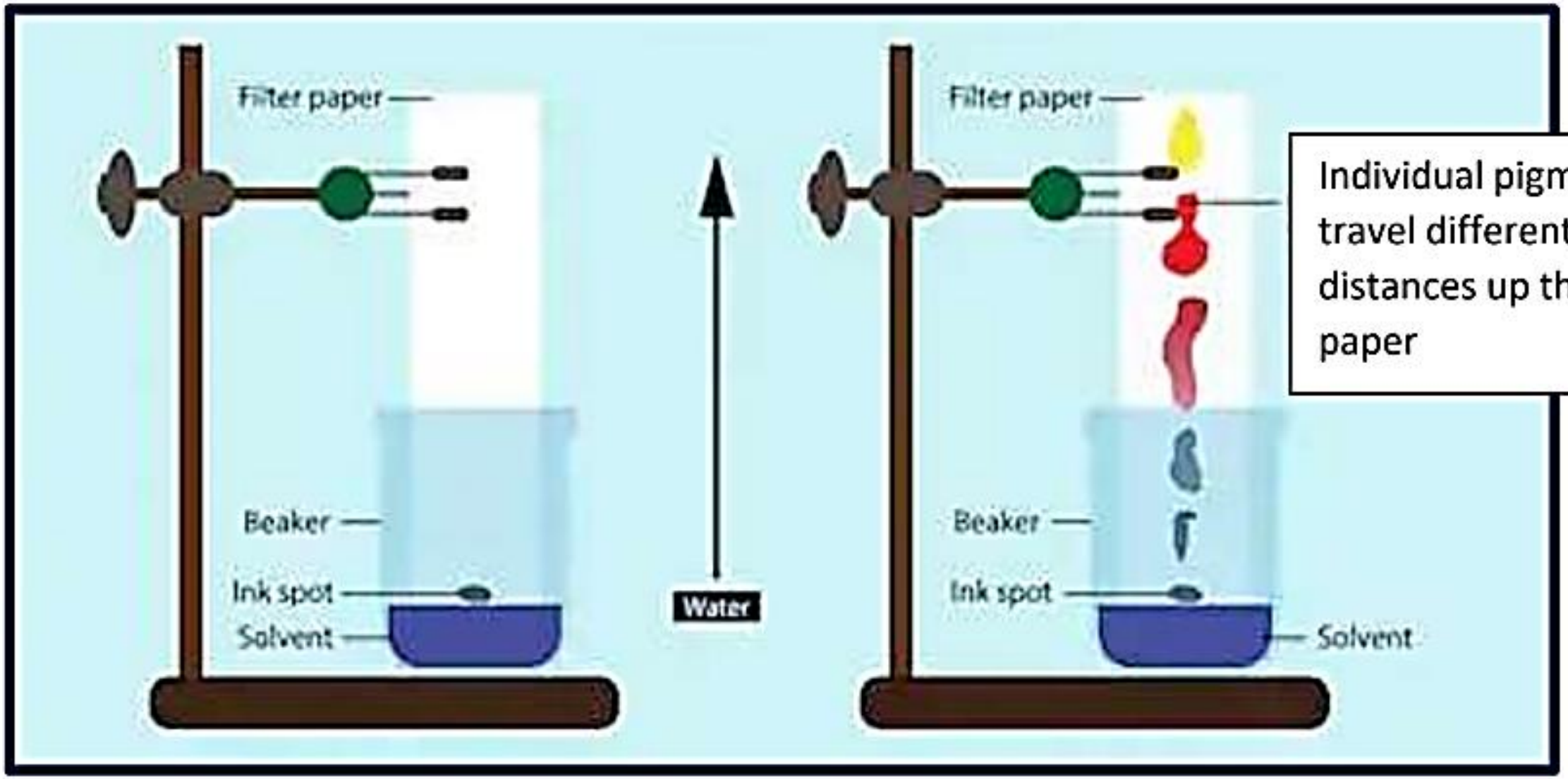
- **Determining the concentration of a solution of unknown concentration using solutions of known concentration**
- **Kinetics reactions** (like bleach + blue food dye)

Important to Remember

- **Before using, you need to calibrate the spectrophotometer with a blank of just solvent** (in order to account for any absorbance due to solvent and cuvette itself)
- **Molarity = moles solute/L of solution**
- **Absorbance = amount of light the solution absorbs at a specific wavelength**
- **Molar absorptivity (1/M*cm) describes how intensely a sample absorbs light at a specific wavelength** (constant unique to the substance at a specific wavelength)
- **Path length** of sample = length of the cuvette where the light will travel (cm)
- **Concentration is molarity**

CHROMATOGRAPHY

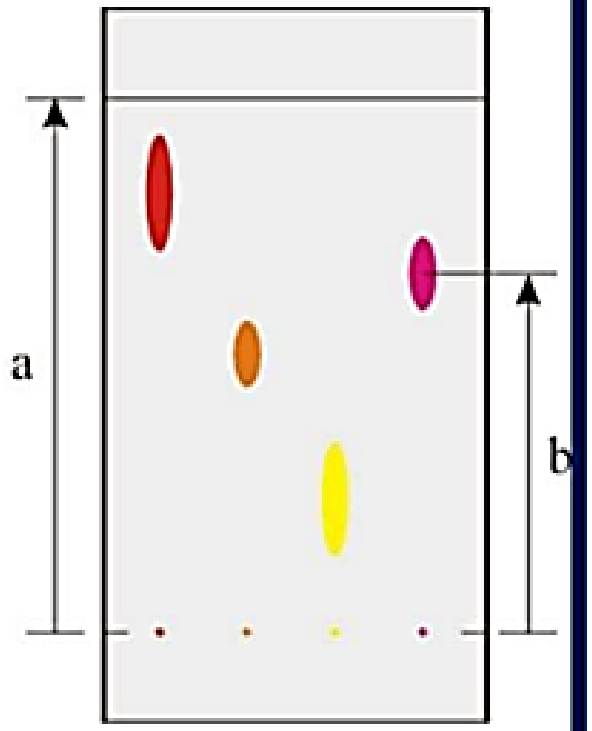
Chromatography



Chromatography

The R_f value for each dye is then worked out using the formula:

$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$$



Chromatography

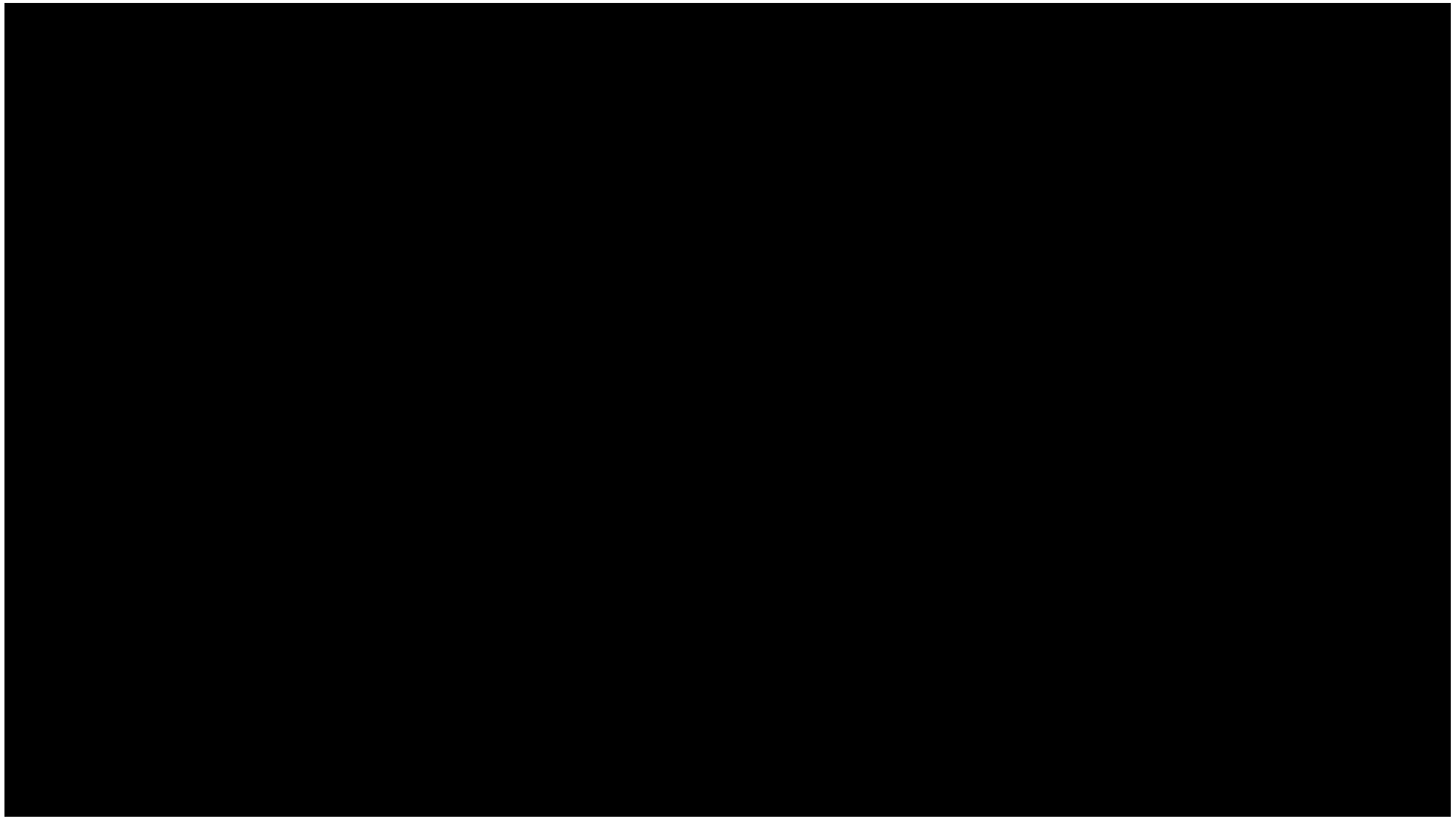
Common Mistakes

- **Solvent reaches the top of the paper strip.**
  - Rf values cannot be calculated as we do not know how far the solvent would have traveled had there been more paper.
- **No major difference in polarity between paper and solvent**
  - Substances cannot be adequately separated
- **No major differences in polarity of components of mixture**
  - Substances cannot be adequately separated

Common Applications

- **Determining the components of a mixture**

Chromatography



https://youtu.be/eUIcebi1GVc

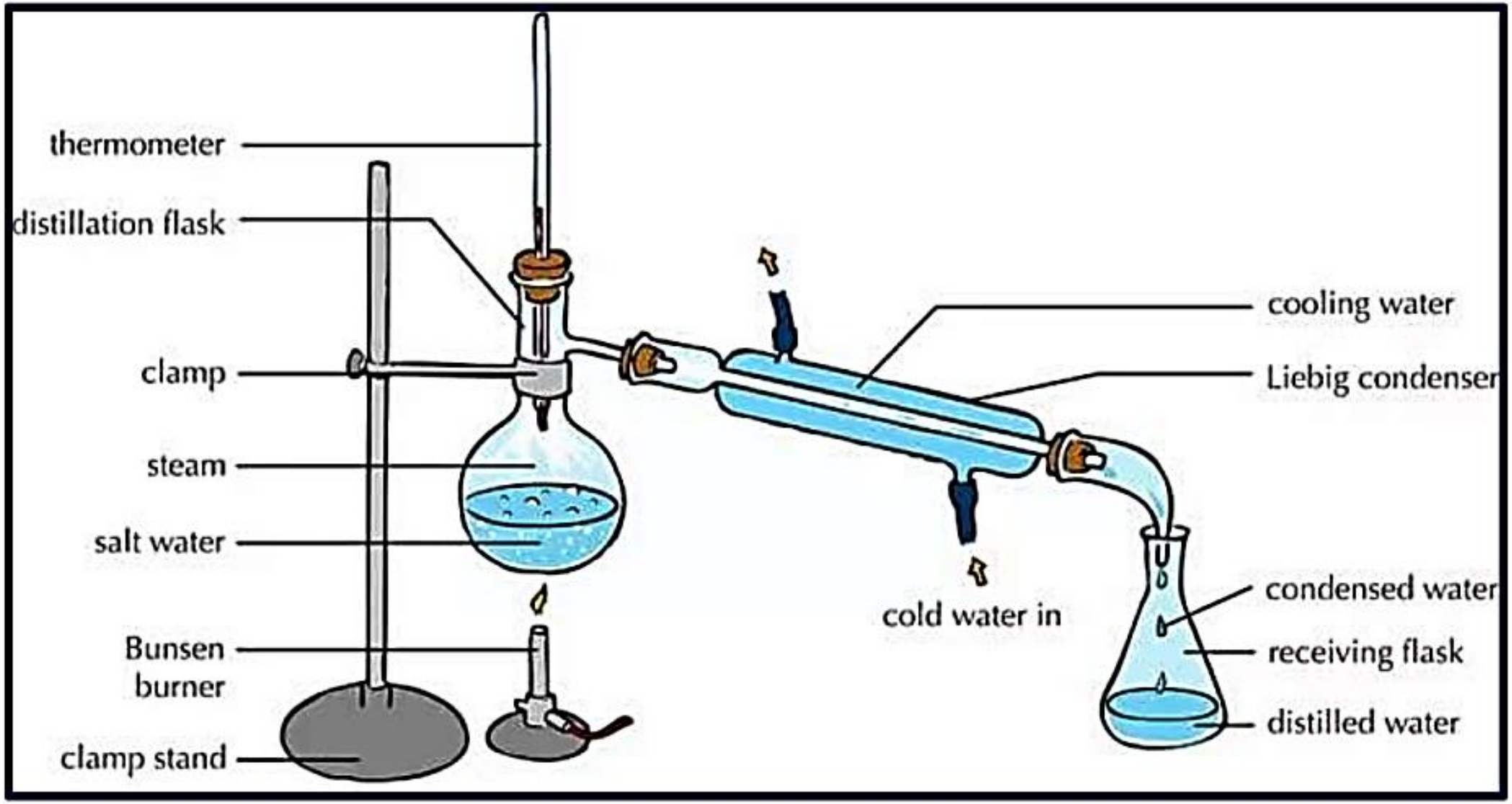
Chromatography

Important to Remember

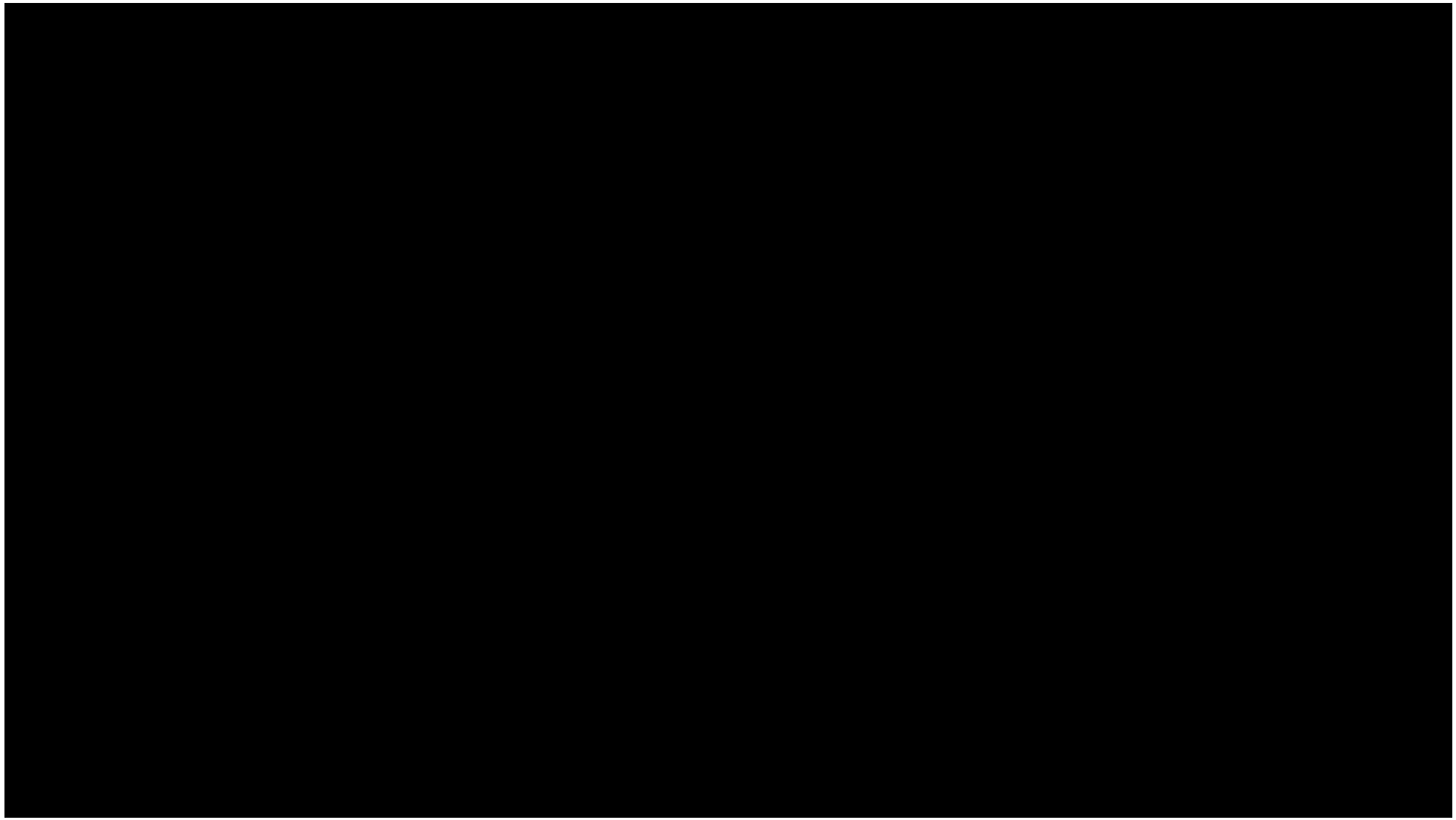
- **The substance that travels further up** the paper is more attracted to solvent.
- **The substance that travels the least** is most attracted to the paper.
- **Paper** is usually relatively nonpolar in comparison to the solvent.
- **Silica plates** are usually relatively polar in comparison to the solvent.
- **If multiple trials are run,** compare Rf values, not relative heights.
- **Polar substances** tend to lack symmetry, have polar bonds, and have lone pairs on the central atom. They are most soluble in other polar substances.
- **Nonpolar substances** tend to be symmetrical, have identical bonds, and have no lone pairs on the central atom. They are most soluble in other nonpolar substances.

FRACTIONAL DISTILLATION

Fractional Distillation



Fractional Distillation



https://youtu.be/eUIcebi1GVc

Fractional Distillation

Common Mistakes

- **Components of the mixture have the same or very similar boiling points.**
  - Same boiling points – could not be separated
  - Similar boiling points – difficult to control temp enough to separate
- **Thermometer touched bottom or side of the flask**
  - Solution will appear hotter than it actually is

Common Applications

- **Separating components in a solution/mixture based on different boiling points.**

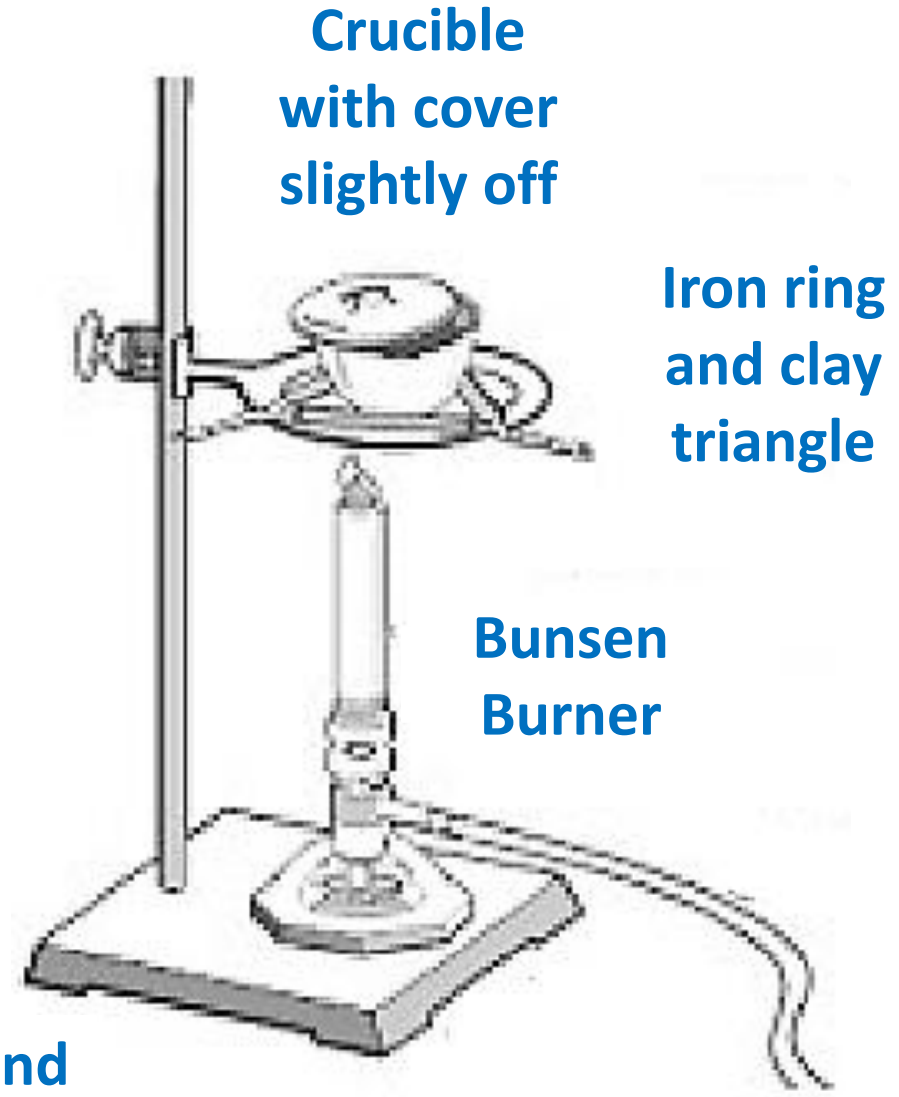
Important to Remember

- **Distillate -** substance collected in the flask at the (substance with lower boiling point)
- **Lower boiling point substance** has a greater vapor pressure and weaker IMFs
- **Higher boiling point substance** has a lower vapor pressure and stronger IMFs
- **Temperature of the solution will remain constant while a component is boiling off.**

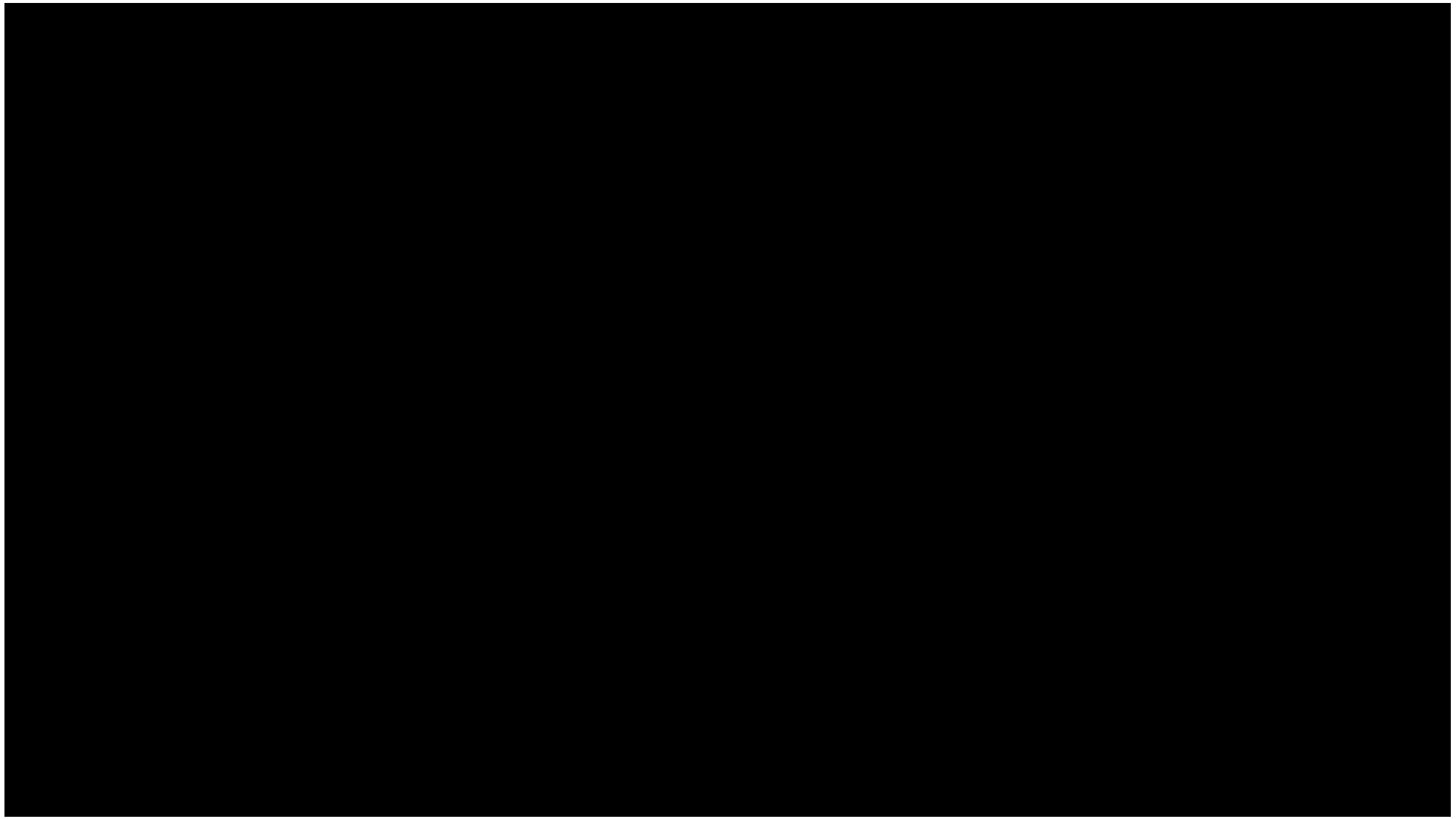
PERCENT COMPOSITION OR FORMULA OF A HYDRATE

Percent Comp/Formula of a Hydrate

1. Take mass of hydrate.
2. Heat until all water has been driven off.
3. Cool, then weigh.
4. Heat again for a couple more minutes.
5. Cool, then weigh.
6. If constant mass has been reached, experiment is complete.



Percent Comp/Formula of a Hydrate



https://youtu.be/eUIcebi1GVc

Percent Comp/Formula of a Hydrate

Common Mistakes

- **Not heating the hydrate enough**
  - Ratio of anhydrous salt: water will not be accurate, as water will remain in the sample
  - Appears fewer moles of water and more moles of salt will be in the hydrate
- **Overheating the hydrate**
  - Anhydrous salt could decompose in the heat
  - It will appear as though the salt is composed of more water than it is
- **Salt sticks to spatula or is spilled in the process of the lab**
  - It will appear as if there is more water in the sample than there actually is; more moles of water will appear to be in sample than there actually are
- **Crucible is weighed while still warm**
  - Inaccurate mass will be obtained

Percent Comp/Formula of a Hydrate

Common Applications

- **Empirical formula of hydrates, percent composition of hydrates**

Important to Remember

- **Hydrated salt:** before heating
- **Anhydrous salt:** after heating
- **Moles of anhydrous salt:** moles of water = ratio for hydrate