

# مختبر أساسيات الكيمياء

م.فادية

للطالب المبدع  
محمد المومني

إرادة - ثقة - تغيير

بِسْمِ اللّٰهِ الرَّحْمٰنِ الرَّحِیْمِ  
اللّٰهُمَّ صَلِّ وَسَلِّمْ عَلٰی سَيِّدِنَا مُحَمَّدٍ

دَفْتَرِ مَخْتَبِرِ

اساسيات الكيمياء

(مصر، فاديه)

عمل الطالب:-

أ. محمد المومني

# Experiment (1)

## Safety and Equipment

### Safety rules

1. Always wear laboratory coat and safety glasses (goggles).
2. Do not work in a laboratory if no lecturer or technician is present.
3. Never chew gum, eat, drink, or smoke in the laboratory.
4. Never taste chemicals. Wash your hands well before leaving the laboratory.
5. Do not wear scarves because they may accidentally be ignited in the flame of a Bunsen burner.
6. Girls with long hair should tie it back before entering the lab, it, too, may accidentally catch fire.
7. Contact lenses should not be worn in the lab.
8. Note the position of **safety equipment** like fire extinguishers, eye washers, emergency showers, fire blanket and first aid boxes. Report all accidents immediately to a staff member or technician.



Safety goggles



fire extinguisher



eye washer with  
emergency shower

fire blanket



first aid box



9. Read **container labels** carefully.
10. Do not insert droppers into reagent bottles. Pour a small amount of the chemical into a beaker.
11. Do not return any excess chemicals to the reagent bottles.
12. Be careful about discarding away **wastes**. Always follow instructions.
  - a. Do not dispose of solids into the sink.
  - b. Do not leave glassware or any other solid materials, including filter papers, in the sink. Put broken glassware into the labeled buckets.
  - c. Some waste liquids must be stored into special bottles, not disposed of in the sink.
13. Do not use your mouth to fill a pipette. There are special bulbs for this purpose.
14. Use the **fume hood** when handling strong-smelling or irritating chemicals.
15. When heating anything in a test tube, do not point the mouth of the test tube towards yourself or towards any other person.
16. Notify your lab instructor if there is a mercury spill due to a broken mercury thermometer.
17. Notify your lab instructor immediately of any injury, spill, fire, or explosion.
18. Before leaving the laboratory turn off any water taps and burners and dispose of solid waste in the correct container. Also, wash all apparatus used and clean up the bench top.



**Read the following safety rules before going into the laboratory and abide by them.**

**Wear safety glasses at all times. Prescription glasses are adequate but contact lenses are not adequate.**

- 1- Always wear your laboratory coat. Do not wear clothing that hinders free movement of your hands or hangs loose outside your laboratory coat.
- 2- Do not work in a laboratory if no lecturer or technician is present. Read the experimental instructions carefully before starting the work. Especially note any precautions that must be taken.
- 3- Never eat, drink, or smoke in the laboratory. Never taste chemicals. Wash your hands well before leaving the laboratory. Also, wash your hands or any part of your body immediately with water when it comes in contact with chemicals.
- 4- Do not use your mouth to fill a pipette. There are special bulbs for this purpose.
- 5- Note the position of safety equipment like fire extinguishers, eye washers, and first aid boxes. Report all accidents immediately to a staff member or technician.
- 6- Use the fume hood when handling strong-smelling or irritating chemicals.
- 7- Be careful about discarding away wastes. Always follow instructions.

Do not dispose of solids into the sink. Do not leave glassware or any other solid materials, including filter papers, in the sink. Put broken glassware into the labeled buckets.

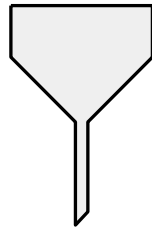
Some waste liquids must be stored into special bottles, not disposed of in the sink. A staff member or technician will help you.

- 8- Do not leave a lit burner unattended. Always stay clear form the flame.
- 9- When heating anything in a test tube, do not point the mouth of the test tube towards yourself or towards any other person.
- 10- Before leaving the laboratory turn off any water taps and burners and dispose of solid waste in the correct container. Also, wash all apparatus used and clean up the bench top.
- 11- Keep your bench clean and tidy while you are working. Clean up any spills or broken glass immediately. Keep your books and papers away from water, chemicals, and flames. Position your apparatus on the bench so that it is convenient and comfortable to use. Keep unused equipment out of the way, so that you do not knock it over.
- 12- If you are in any doubt about anything, ask the staff member or technician for advice.
- 13- Request the assistance of your instructor/technician if and when you suffer a cut or a burn **or face a dangerous** situation.

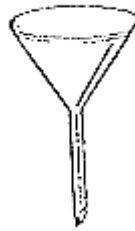
## SOME LABORATORY APPARATUS



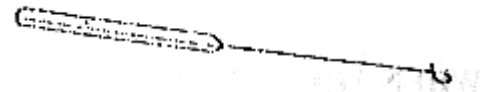
Beaker



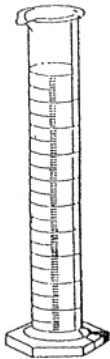
Büchner funnel



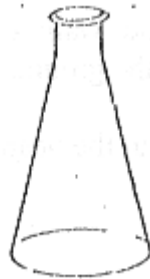
Funnel



Glass rod with platinum wire  
(for flame testing)



Graduated cylinder



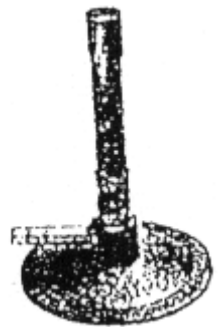
Erlenmeyer flask



Dropping bottle



Test tubes



Bunsen burner



Stirring rod



Dropper



Safety glasses  
(Goggles)



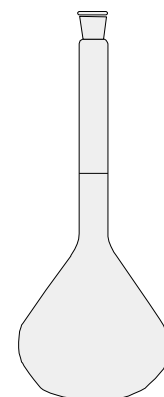
Test tube rack



Plastic wash bottle



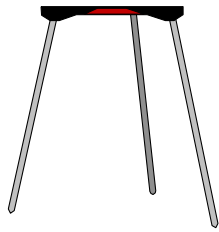
Clamp holder



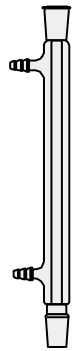
Volumetric flask



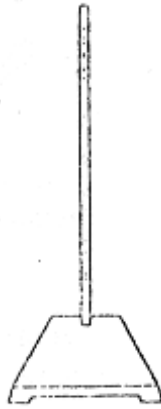
Burette



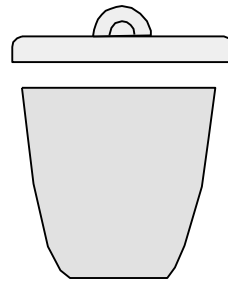
Tripod



Condenser



Stand



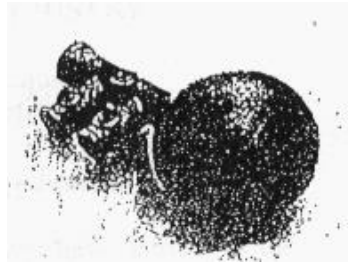
Crucible



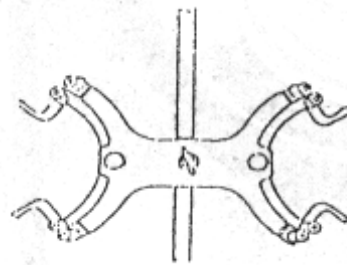
Volumetric pipette



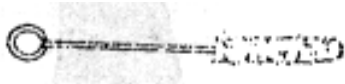
Reagent bottle



Pipette filler



Burette clamp



Test tube brush



Metal spatula



Test tube holder



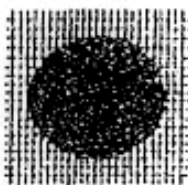
Ring clamp



Ceramic square



Forceps



Wire gauze


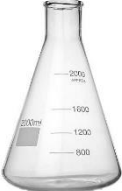
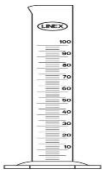
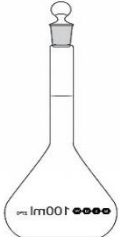
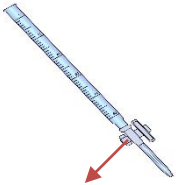









Watch glass
















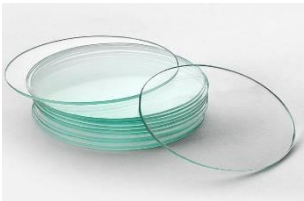
Crucible tongs




# Some Laboratory Apparatus

Name	picture	use
1. Beaker		Container, Running reactions, mixing chemicals, heating chemicals
2. Erlenmeyer flask		For mixing, heating, cooling, filtration, useful for titrations
3. Graduated cylinder		For measuring the volume of liquids.
4. Volumetric flask		For precise dilutions and preparation of standard solutions.
5. Burette	 <p data-bbox="444 1415 558 1446">Stopcock</p>	For accurately dispensing known amounts of liquid
6. Pipette	 <p data-bbox="396 1696 542 1728">Pipette filler</p>	To transport a measured volume of liquid

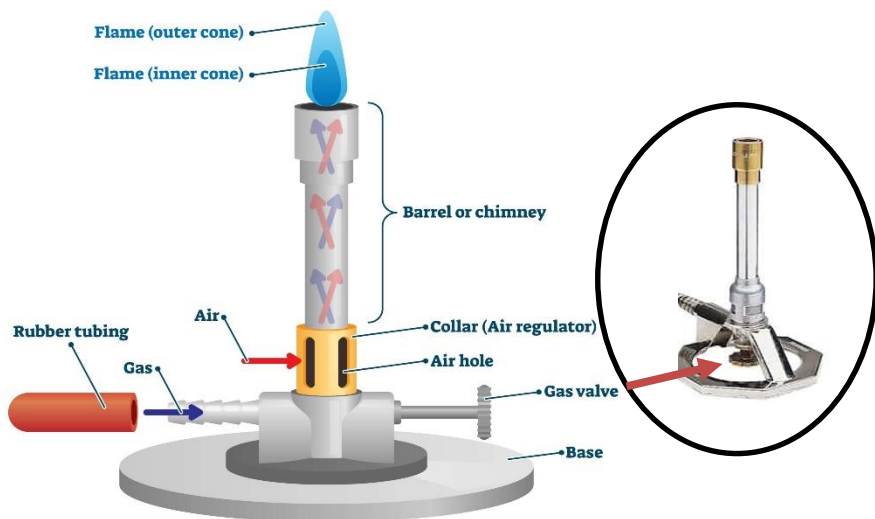
7. Bunsen burner		Heat source in the chemistry lab.
8. Stand		To hold various objects.
9. Ring		When attached to the ring stand, this iron ring is used to support glassware.
10. Wire gauze		Used in conjunction with iron ring, supports the beaker flasks, or other glassware during heating
11. Clamp	 <p>Clamp holder</p>	When attached to the ring stand, this clamp is used to hold a test tube or flask or burette
12. Funnel		For filtering and for adding chemicals without spilling

13. Buchner funnel		Funnel used for vacuum filtration
14. Glass rod/ stirring rod		For stirring chemicals
15. Test tubes		For mixing and heating chemicals and running reactions– smaller quantities than beakers and flasks.
16. Test tube holder		For transport, holding test tube in a flame or placing test tube in hot water bath
17. Dropper		For adding small amounts of chemicals – usually by drops
18. Dropper bottle		For the accurate dispensing of reagents one drop at a time
19. Reagent bottle		Contain chemicals in liquid and stored in cabinets or on shelves

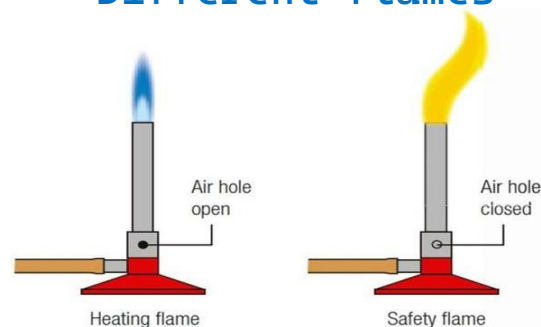
20. Sample vial		For storing small amounts of chemicals
21. Crucible	 <p>Crucible lid</p>	To strongly heat substances above a flame
22. Clay triangle		To hold a crucible above a flame – used in conjunction with iron ring
23. Crucible tongs		To transport hot crucibles and to remove their covers
24. Spatula		To Transfer materials and samples from one place to another.
25. Weighing boat		To weigh substances
26. Watch glass		To hold solids while being weighed, as a cover for a beaker,

27. Thermometer		For measuring temperature. Alcohol have red bulbs, mercury have a silver colored bulb
28. Wash bottle		Squeeze sides to dispense distilled water Ex. for rinsing solids out of a container when filtering
29. Brushes		Used to clean and scrub glassware

## Bunsen burner

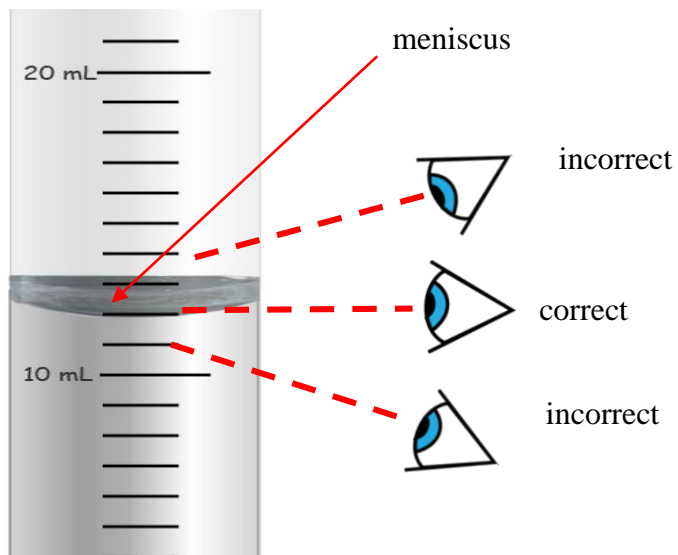


### Different flames



## Using the graduated cylinder

When water and other liquids are poured into a narrow glass tube such as a graduated cylinder, the surface of the water is curved, not flat. This curve is called a **meniscus**. As shown below:



It is important when you read glassware used for measuring volumes of liquids that you keep your eye level with the meniscus, and use the **bottom of the meniscus** for your readings.

## Trends of accuracy of tools:

Pipette > burette > volumetric flask > graduated cylinder > beaker



**More accurate**

## Uncertainty

### 1. Digital equipment (example: digital balances)

**Uncertainty in a digital measuring device =  $\pm$  the smallest increment**

One decimal:  $0.0 \pm 0.1$

Example:  $3.5 \text{ g} \pm 0.1$

Two decimals:  $0.00 \pm 0.01$

Example:  $3.29 \text{ g} \pm 0.01$

Three decimals  $0.000 \pm 0.001$

Example:  $3.128 \text{ g} \pm 0.001$



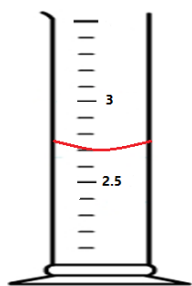
### 2. Graduated tools (example: graduated cylinder)

Step 1: find the smallest division

Step 2: find the uncertainty

**Uncertainty in a scale measuring device =  $\frac{1}{2} \pm$  the smallest increment (division)**

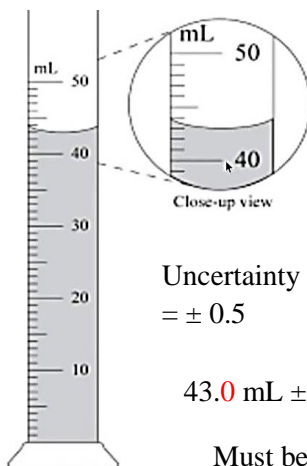
Step 3: the number of decimal numbers in the measured volume should match the number of decimal numbers in the uncertainty.



Uncertainty =  $\pm \frac{1}{2} * 0.1$   
=  $\pm 0.05$

2.70 mL  $\pm 0.05$

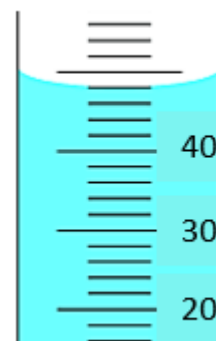
Must be added



Uncertainty =  $\pm \frac{1}{2} * 1$   
=  $\pm 0.5$

43.0 mL  $\pm 0.5$

Must be added



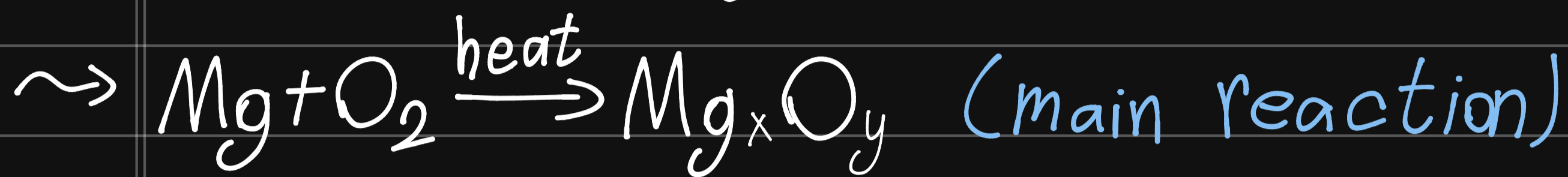
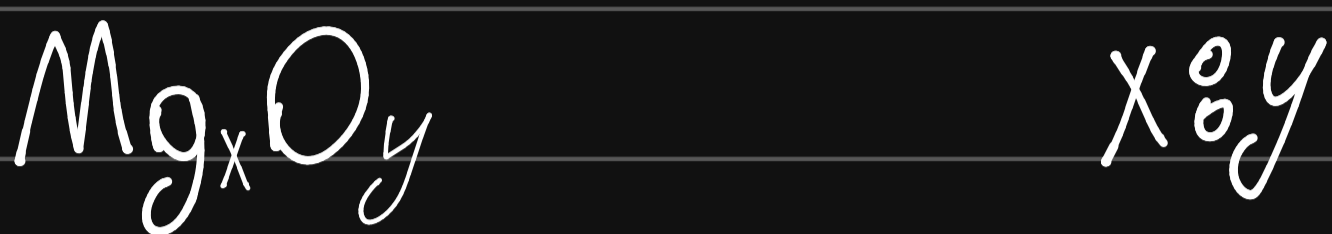
Uncertainty =  $\pm \frac{1}{2} * 2 = \pm 1$

48 mL  $\pm 1$

# Ex. 2<sup>o</sup> - empirical formula

الهدف

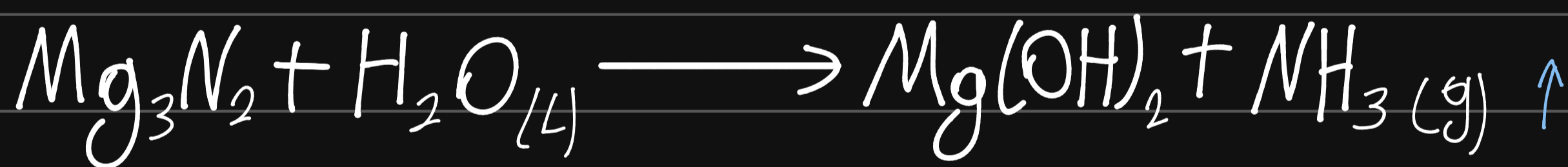
Aim<sup>o</sup> - empirical formula for Magnesium oxide.



↳ Magnesium nitride (side product)

↳ to remove it<sup>o</sup> -

① add water<sup>o</sup> -



② heat<sup>o</sup> -



\* to find Empirical formula of  $\text{X}_n\text{Y}_m$ .

- ① find the masses of them in  $\text{X}_n\text{Y}_m$ .
- ② find their no. of moles in  $\text{X}_n\text{Y}_m$ .
- ③ divide by the lowest no. of moles.
- ④ make sure the ratio is an integer.

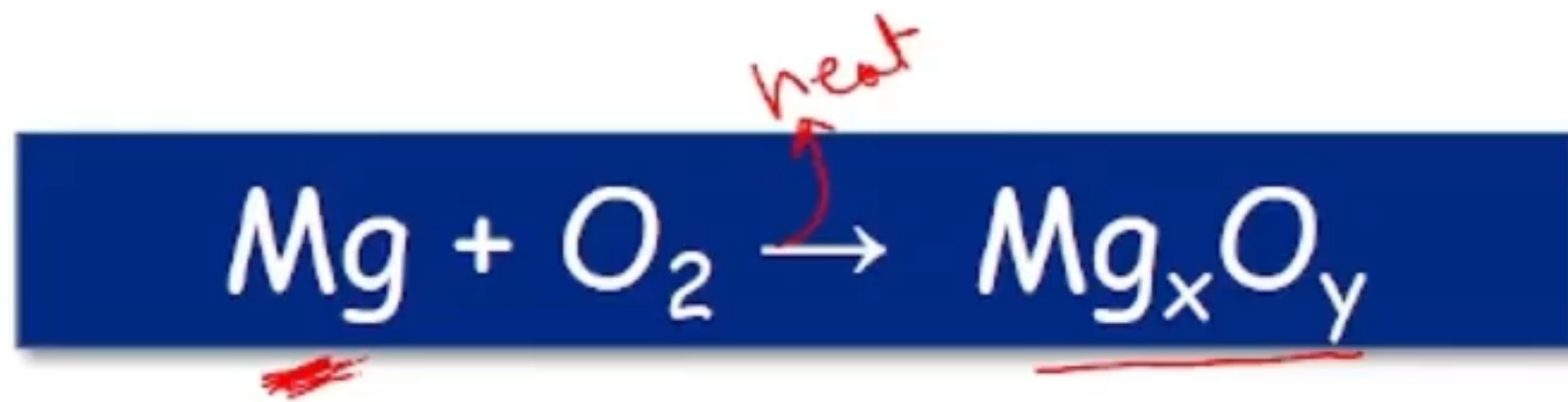
# Empirical Formula

Formula of a compound

Molecular Formula	Empirical Formula
is the <u>true</u> number of atoms in a compound	is simplest whole number ratio of atoms in compounds.
$C_2H_4, C_6H_{12}O_6$ $\frac{2:4}{2} \quad \frac{6:12:6}{6}$ $CH_2$	$CH_2, CH_2O$ $1:2 \quad 1:2:1$

# Objective

- Determination of the empirical formula of magnesium oxide ( $Mg_xO_y$ ).



$x, y \rightarrow$  integer no.  
 $Mg:O$



Magnesium metal

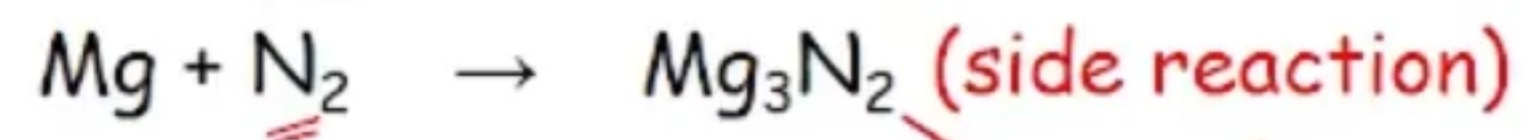
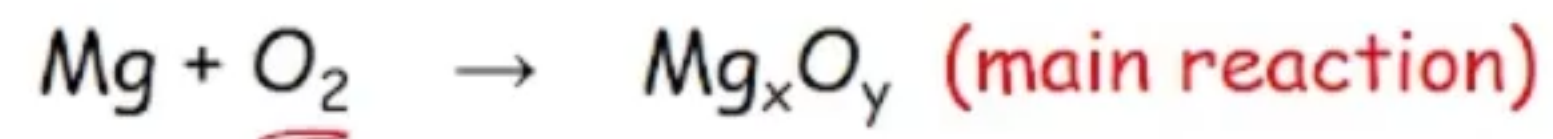
- Mg react with oxygen from the air
- Masses before and after the oxidation are measured.
- The resulting masses are used to calculate the experimental empirical formula.

$$\text{Mass O} = \text{mass } Mg_xO_y - \text{mass } Mg$$

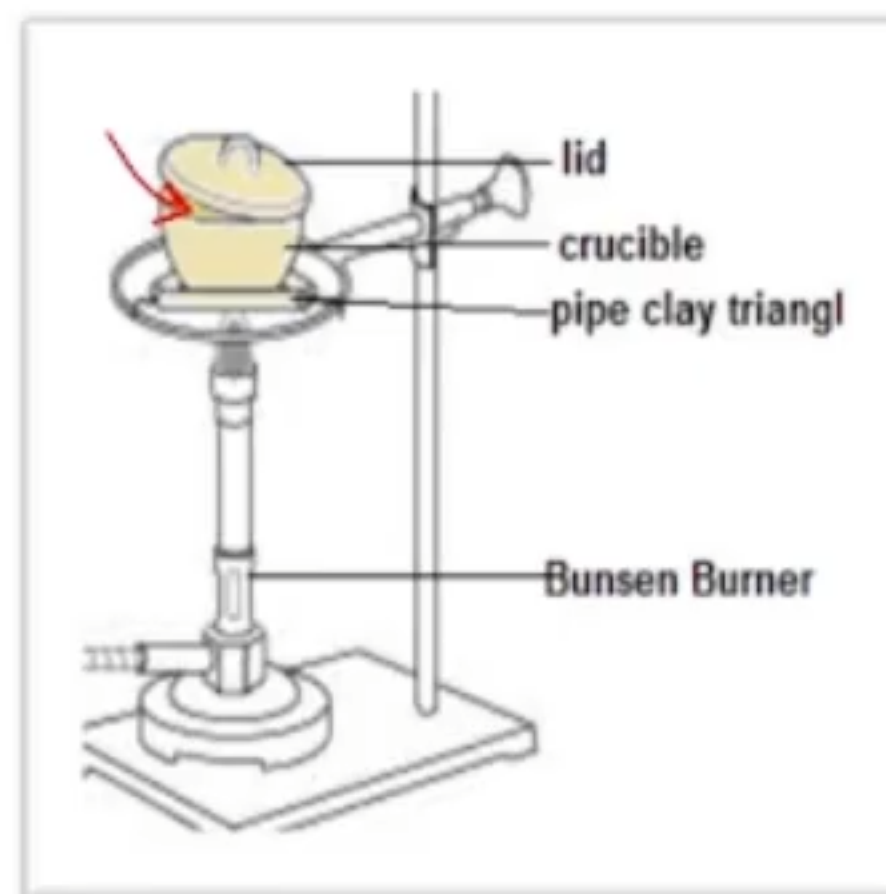
$Mg:O$

## In this experiment:

- Heat a known mass of Mg in a slightly open crucible.
- Mg will react with both  $O_2$  and  $N_2$  gases.



Magnesium nitride  $\rightarrow$  side product



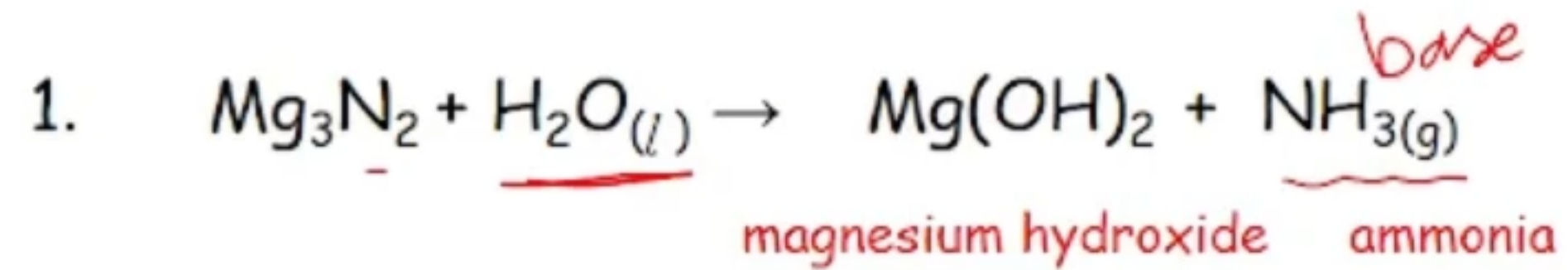
## Experiment 2

# Empirical Formula of a Compound

- The gray powder (ash) contains magnesium oxide MgO and magnesium nitride Mg<sub>3</sub>N<sub>2</sub>
- The amount of oxide MgO in the product is higher than nitride Mg<sub>3</sub>N<sub>2</sub>. why? (in air amount  $N_2 > O_2$ )

In this experiment:

- The nitride  $\text{Mg}_3\text{N}_2$  in the product can be converted to hydroxide by adding  $\text{H}_2\text{O}$ ; then to oxide via heating



- The released ammonia can be detected via smelling or via a wet litmus paper.

red  $\rightarrow$  blue

**Remove the burner  
and cool the crucible**

- Since the product at this stage is composed of only Mg and O, the mass of O can be found

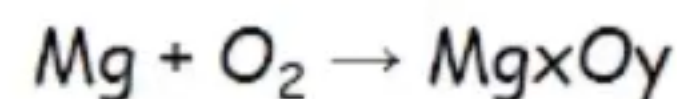
$$\text{Mass O} = \text{mass Mg}_x\text{O}_y - \text{mass Mg}$$

- From the masses of O and Mg, you can determine the empirical formula of this oxide.

## How can you determine an empirical formula?

1) Determine the masses of two atoms ( Mg and O )

To find oxygen mass:



$$\text{Mass O} = \text{mass Mg}_x\text{O}_y - \text{mass Mg}$$

2) Determine the number of moles by dividing the grams by the atomic mass.

$$\frac{\text{mass Mg}}{\text{Molar mass (24.31)}} : \frac{\text{mass O}}{\text{Molar mass (16.0)}} \quad \text{O}_2$$

3) Divide the number of moles of each element by the smallest number of moles.

4) Convert numbers to integer numbers by

- rounding to the nearest whole number

or

- if values are fractions of 0.25 then multiply by 4

- if values are fractions of 0.50 then multiply by 2

- if values are fractions of 0.333 then multiply by 3

5) Find molar ratio of each element (this will give the empirical formula)



**Experiment (2)**  
**Empirical Formula of a Compound**

**PostLab**

$\frac{Mg}{O}$

1. What is the effect of  $Mg_3N_2$  if it is not decomposed on the reported (Mg to oxygen) mole ratio?

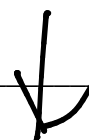


2. Describe the effect of each of the following factors, whether increases, decreases, or has no effect on the reported value of (Mg to O) mole ratios:

- a. If carbon is deposited on the crucible's surface (because of improper heating) and the crucible with contents is weighed without removing the carbon residue by further heating.



- b. If the magnesium oxide ash is not dried completely.



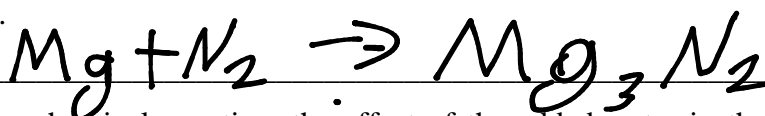
- c. If rapid oxidation of magnesium occurs by too much air, and some Mg is lost.



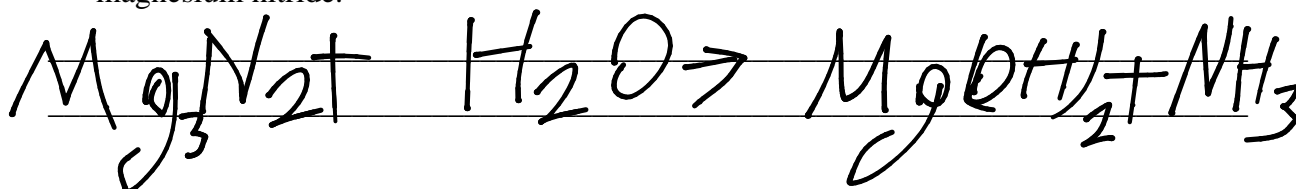
- d. If air is not sufficient to react with all the magnesium.



3. Explain how magnesium nitride is formed during heating? Write a chemical equation for its formation.



4. Explain, by a chemical equation, the effect of the added water in the decomposition of magnesium nitride.



**Example 1:**

Nicotine is a compound containing C, H, and N. A 2.50 g sample of the compound is burned and produces 6.78 g of CO<sub>2</sub>, 1.94 g of H<sub>2</sub>O, and 0.43 g of N<sub>2</sub>. what is the empirical formula of nicotine?

First we calculate the number of moles of each element present.

$$\text{mole C} = \text{mole CO}_2 = \frac{6.78\text{g}}{44.0\text{g/mol}} = 0.154 \text{ mole}$$

$$\text{mole H} = 2 \times \text{mole H}_2\text{O} = 2 \times \frac{1.94\text{g}}{18.0\text{g/mol}} = 0.216 \text{ mole}$$

$$\text{mole N} = 2 \times \text{mole N}_2 = 2 \times \frac{0.43\text{g}}{28.0\text{g/mol}} = 0.031 \text{ mole}$$

ratio of moles = ratio of atoms

The empirical formula is C<sub>0.154</sub> : H<sub>0.216</sub> : N<sub>0.031</sub>

Dividing by 0.031, (the smallest number) this will give the smallest whole-number ratio:

C<sub>5</sub>H<sub>7</sub>N = the empirical formula

In the chemical synthesis method, a known mass of a metal is combined with a non metal. The product is weighed and the mass of nonmetal is determined. From the mass of the metal and nonmetal, the number of moles is calculated and then the empirical formula is determined.

**Example 2:**

When 0.288 g of P is burned, 0.660 g of white phosphorus oxide is obtained, determine the empirical formula of this oxide?

$$\text{mass of oxygen} = 0.660 - 0.188 = 0.372 \text{ g}$$

$$\text{mole P} = \frac{0.288\text{g}}{31.0\text{g/mol}} = 0.0093 \text{ mole}$$

$$\text{mole O} = \frac{0.372\text{g}}{16.0\text{g/mol}} = 0.0233 \text{ mole}$$

$$\text{mole P: mole O} = 0.0093: 0.0233 = 2 : 5$$

Empirical formula is P<sub>2</sub>O<sub>5</sub>

In today's experiment you will prepare magnesium oxide by burning a known mass of magnesium in air (O<sub>2</sub>). From the mass of product, and magnesium, the empirical formula of magnesium oxide is determined.

1. How many grams of magnesium combine with 1.50 g of chloride ions in  $\text{MgCl}_2$ ?

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2. If 11.80 g of iron reacts with 5.06 g of oxygen. Determine the empirical formula of the resulting oxide?

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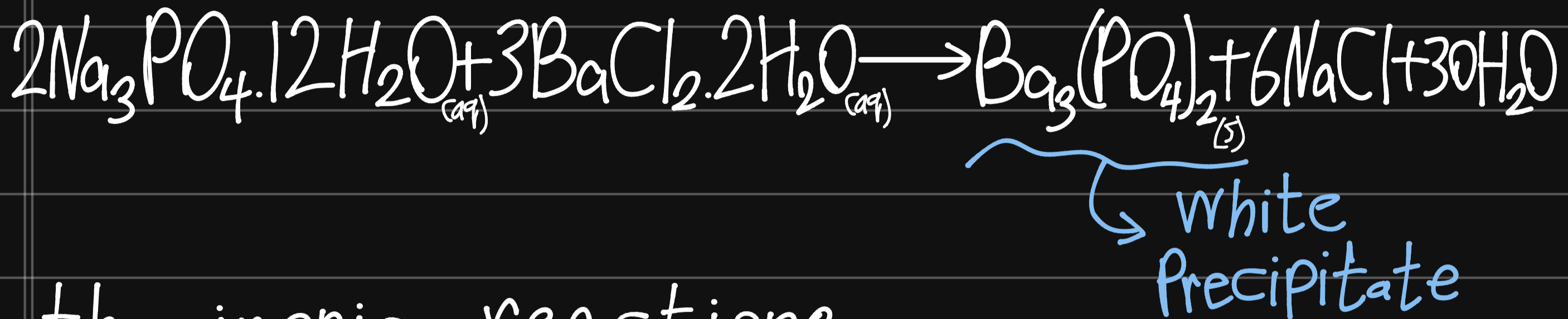
# Fill in the blanks



Solid magnesium metal is heated and reacts with both oxygen and nitrogen in the air to produce ----- and magnesium nitride. Upon addition of water, ----- is released and magnesium hydroxide is formed. A final heating step will produce magnesium oxide and ----- will be released

# Ex. 3<sup>o</sup> Limiting reactants:-

\* Precipitation reactions:-



\* the ionic reactions:-



Aim:- ① LR / Excess  $\begin{cases} \rightarrow \text{reacted} \\ \rightarrow \text{unreacted} \end{cases}$

② mass of  $\text{Ba}_3(\text{PO}_4)_2 \equiv$  Actual yield.

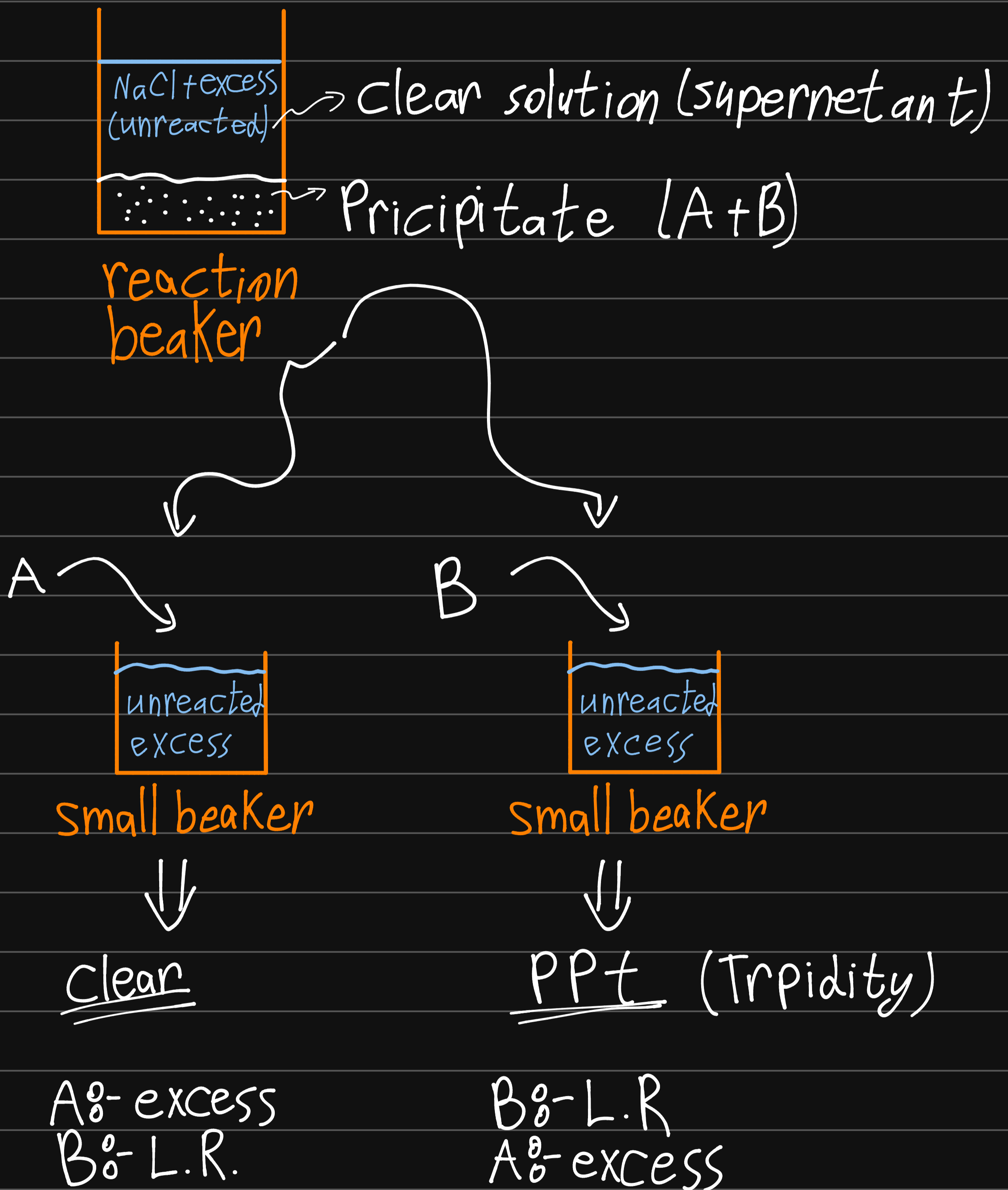
after filtration

③ %A, %B in the mixture

$$\textcircled{4} \% \text{ yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100\%$$

\* Heating  $\rightarrow$  Coagulation  $\rightarrow$  to increase particle size

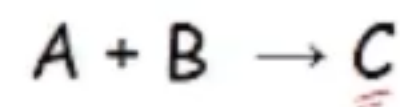
# \* finding the L.R & excess



# Introduction



- In a chemical reaction, the reaction will continue towards equilibrium or until one of the reactants is completely consumed.

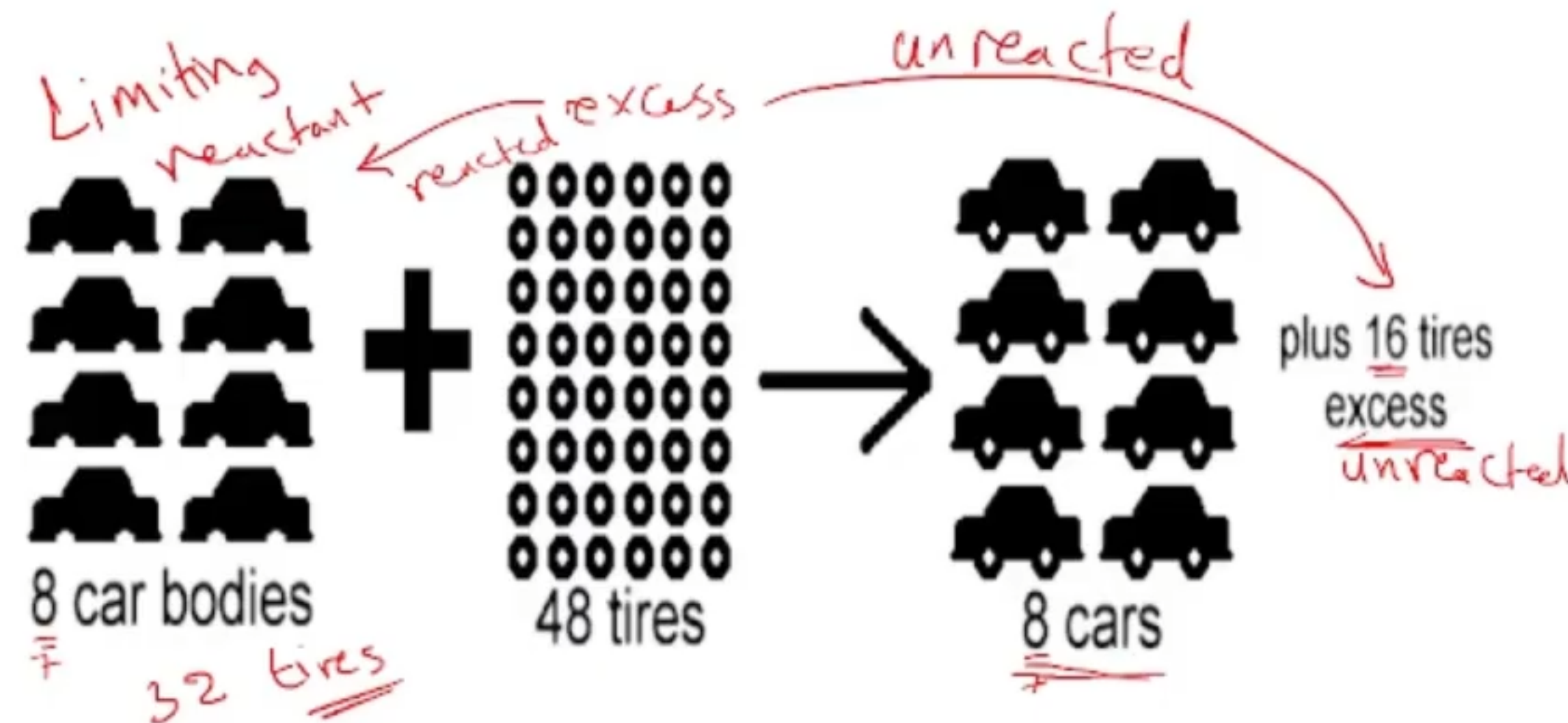


## Limiting reactant

- It is the reactant that is consumed firstly in the chemical reaction.
- Determines when the reaction stops.
- Determines the amount of product formed.

## Excess

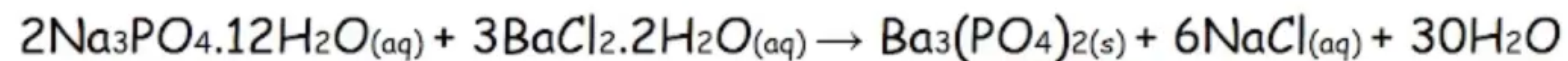
Is the reactant that present at the end of the reaction.



- Theoretical yield:** the amount of product that is expected to be produced when the limiting reagent is completely consumed. (can be calculated based on the limiting reactant)
- Actual yield:** is the actual amount of product that is produced in a chemical reaction.

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100\%$$

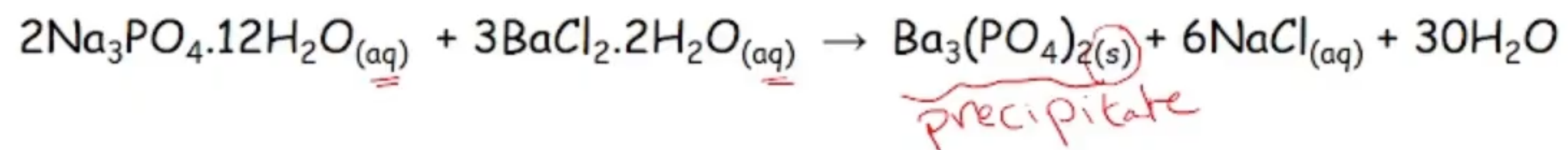
## Objective:



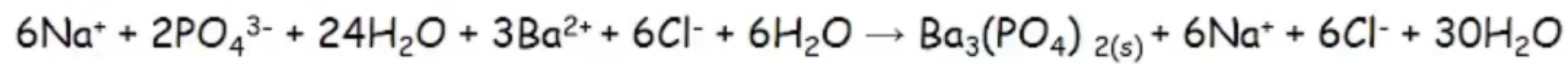
*white precipitate*

1. To determine the limiting reactant in a reaction of sodium phosphate  $\text{Na}_3\text{PO}_4$  and barium chloride  $\text{BaCl}_2$  solution.
2. To determine the actual yield (mass  $\text{Ba}_3(\text{PO}_4)_2$ )
3. To determine the percent composition of a salt mixture (%  $\text{BaCl}_2$  and %  $\text{Na}_3\text{PO}_4$ )

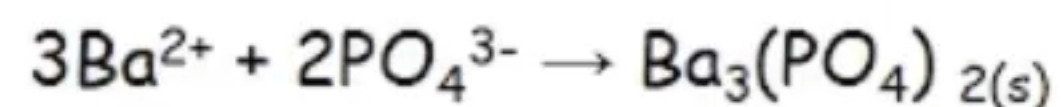
In this experiment, you will carry out a precipitation reaction between aqueous sodium phosphate and aqueous barium chloride.



and the ionic equation is:



The net ionic equation of the reaction is:

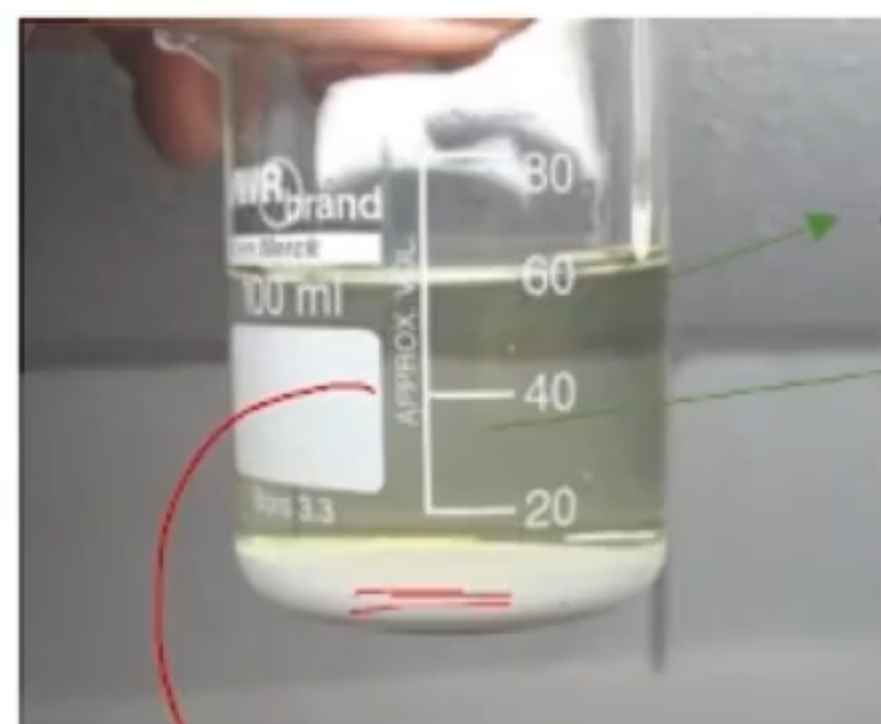


## Determination of the limiting reactant

The limiting reactant is determined using the solutions in beakers 1 & 2 as follows:

- **Test for excess  $\text{PO}_4^{3-}$ :** add 2 drops of 0.5 M  $\text{BaCl}_2$  to the clear solution in beaker 1. If a precipitate is formed, then  $\text{PO}_4^{3-}$  is in excess and  $\text{Ba}^{2+}$  is the limiting reactant.
- **Test for excess  $\text{Ba}^{2+}$ :** add 2 drops of 0.5 M  $\text{Na}_3\text{PO}_4$  to the clear solution in beaker 2. If a ppt. is formed, then  $\text{Ba}^{2+}$  is in excess and  $\text{PO}_4^{3-}$  is the limiting reactant.

**Clear solution = supernatant:** the clear liquid found above a precipitate (filtrate)



$\text{Na}^+, \text{Cl}^-$ , excess  
unreacted  
Supernatant  
Precipitate

$\text{BaCl}_2$   
excess  
↓  
turbidity  
ppt

$\text{Na}_3\text{PO}_4$   
excess  
↓  
clear  
excess:  $\text{Na}_3\text{PO}_4$   
Limiting reactant =  $\text{BaCl}_2$



## Post Lab

1. What is the effect of heating the solution on the particle size of  $\text{Ba}_3(\text{PO}_4)_2$  precipitate?

increase (Coagulation)

2. Describe the effect of each of the following factors (whether increase, decrease, or has no effect) on the actual yield of  $\text{Ba}_3(\text{PO}_4)_2$

a. Using a coarse filter paper instead of one with fine porosity.



b. Insufficient washing of the precipitate.



c. If the precipitate is washed with an acidic solution.



d. If the precipitate was not dried completely.

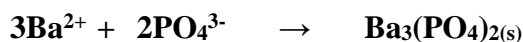


### Example 1:

A mixture containing 40.00 g of  $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$  and 30.00 g of  $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$  is dissolved in water. A precipitate of  $\text{Ba}_3(\text{PO}_4)_2$  weighing 22.65 g is produced. Calculate the % yield of  $\text{Ba}_3(\text{PO}_4)_2(\text{s})$ .

(FM) of  $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O} = 380.2 \text{ g/mole}$ ; (FM) of  $\text{BaCl}_2 \cdot 2\text{H}_2\text{O} = 244.2 \text{ g/mole}$ .

The net ionic equation is:



mole  $\text{PO}_4^{3-} = \text{mole } \text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O} = 40.00\text{g} / 380.2 \text{ g mole}^{-1} = 0.105 \text{ mole}$

mole  $\text{Ba}^{2+} = \text{mole } \text{BaCl}_2 \cdot 2\text{H}_2\text{O} = 30.00 \text{ g} / 244.2 \text{ g mole}^{-1} = 0.123 \text{ mole}$

0.105 mole  $\text{PO}_4^{3-}$  requires  $0.105 \times 3/2 = 0.158 \text{ mole } \text{Ba}^{2+}$ . Since we have only 0.123 mole of  $\text{Ba}^{2+}$ , so the **limiting reactant** is  $\text{Ba}^{2+}$ , and the **reactant in excess** is  $\text{PO}_4^{3-}$ .

The **theoretical yield** of product is calculated from the limiting reactant as follows:

$$\text{mole } \text{Ba}_3(\text{PO}_4)_2 = \text{moles } \text{Ba}^{2+} \times \frac{1 \text{ mole } \text{Ba}_3(\text{PO}_4)_2}{3 \text{ mole } \text{Ba}^{2+}}$$

$$= 0.123 \times \frac{1}{3}$$

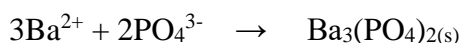
$$= 0.041 \text{ mole}$$

mass of  $\text{Ba}_3(\text{PO}_4)_2 = 0.041 \text{ mole} \times 602.2 \text{ g mole}^{-1} = 24.69 \text{ g}$  (this is the theoretical yield)

$$\% \text{ yield} = \frac{22.65 \text{ g}}{24.69 \text{ g}} \times 100 = 91.7 \%$$

### Example 2:

10.0 g of a unknown mixture containing  $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$  and  $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$  is dissolved in distilled water. The mass of  $\text{Ba}_3(\text{PO}_4)_2$  precipitated is 3.50 g. Calculate the % of each salt present in the mixture if  $\text{BaCl}_2$  is the limiting reactant?



$$\text{mole } \text{Ba}_3(\text{PO}_4)_2 = \frac{3.50 \text{ g}}{602.2 \text{ g/mol}} = 5.81 \times 10^{-3} \text{ mole}$$

$$\text{mole } \text{Ba}^{2+} = \text{mole } \text{BaCl}_2 \cdot 2\text{H}_2\text{O} = 3 \times (\text{mole } \text{Ba}_3(\text{PO}_4)_2)$$

$$\text{mole } \text{BaCl}_2 \cdot 2\text{H}_2\text{O} = 5.81 \times 10^{-3} \times 3 = 1.74 \times 10^{-2} \text{ mole}$$

$$\text{mass of } \text{BaCl}_2 \cdot 2\text{H}_2\text{O} = 1.74 \times 10^{-2} \text{ mole} \times 244.2 \text{ g mole}^{-1} = 4.25 \text{ g}$$

$$\text{mass of } \text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O} = 10.0 - 4.25 = 5.75 \text{ g}$$

$$\% \text{ BaCl}_2 \cdot 2\text{H}_2\text{O} = \frac{4.25}{10.0} \times 100 = 42.5\%$$

$$\% \text{ Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O} = \frac{5.75}{10.0} \times 100 = 57.5\%$$

1. A 25.00 g sample of  $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$  reacts with excess  $\text{BaCl}_2 \cdot \text{H}_2\text{O}$ . If the mass of  $\text{Ba}_3(\text{PO}_4)_2$  obtained is 17.56 g. Calculate the % yield of  $\text{Ba}_3(\text{PO}_4)_2$ ?



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## TEST YOURSELF

Decide if the following statements are TRUE or FALSE

- If addition of  $\text{Na}_3\text{PO}_4$  aqueous solution to the filtrate yield a precipitate,  $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$  is the limiting reactant.
- An underestimate of the limiting reactant is obtained when a portion of the precipitate mass is lost.

# Ex. 4<sup>o</sup> Physical properties

اللو 2

① solubility<sup>o</sup>-

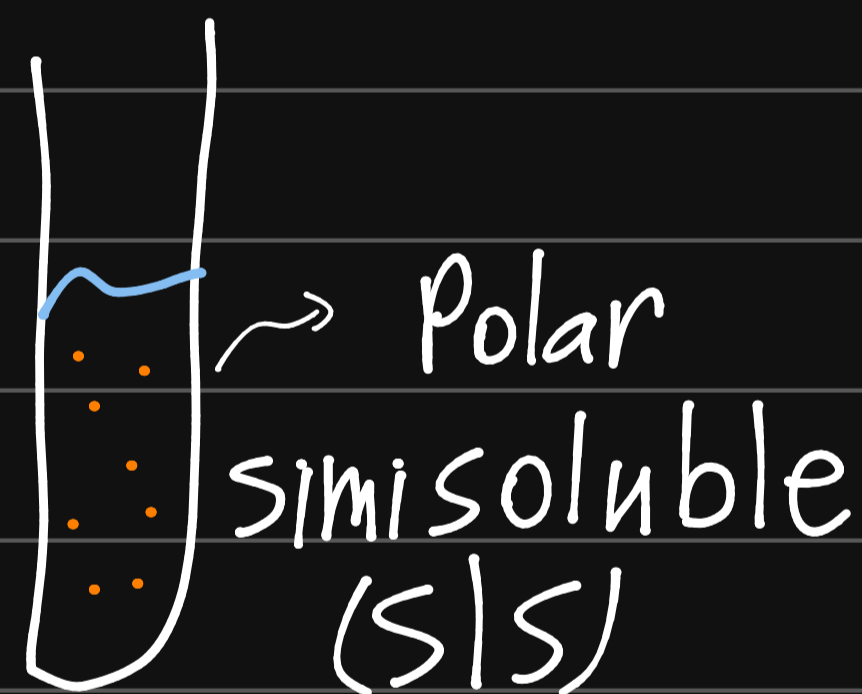
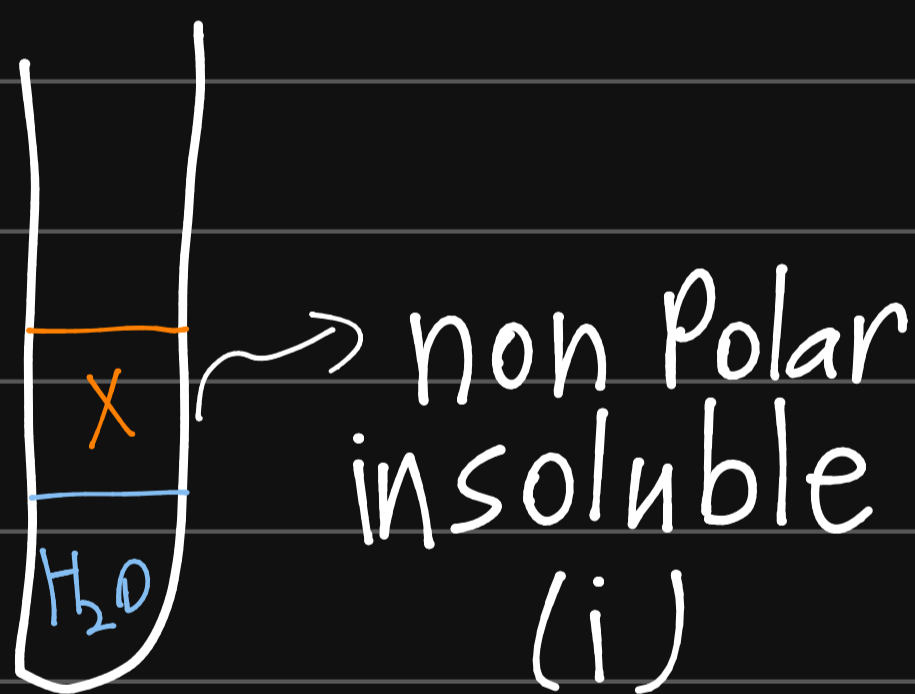
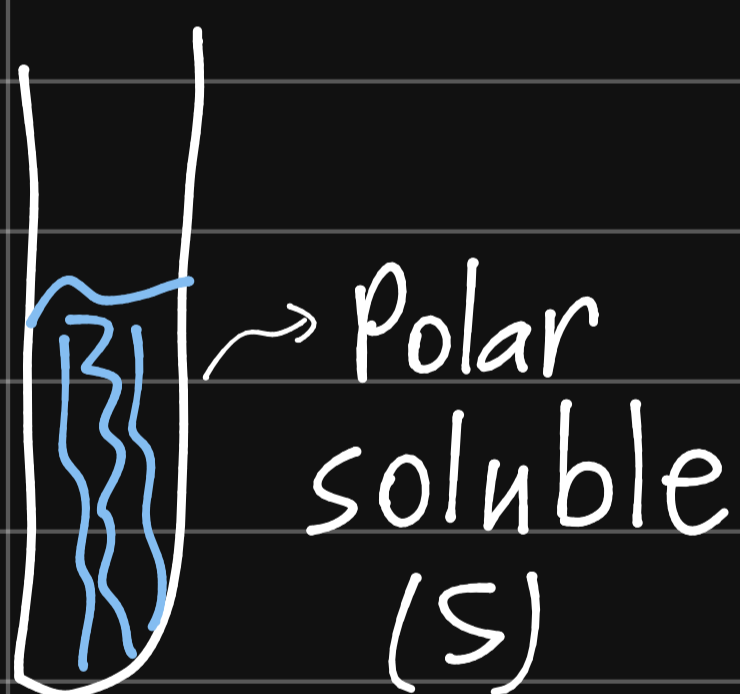
"Like dissolve like"

\* Polarity<sup>o</sup>- ① Polar ② non-Polar

Water ( $H_2O$ )  $\rightarrow$  Polar

- solubility of X in water<sup>o</sup>-

① X<sup>o</sup> Polar      ② X<sup>o</sup> non Polar      ③ X<sup>o</sup> semi-soluble



② density =  $\frac{\text{mass}}{\text{volume}}$

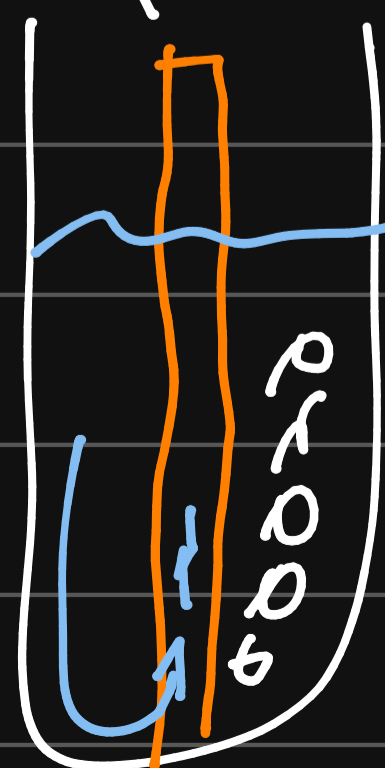
③ Boiling Point<sup>o</sup>-

(a)  $V.P > \text{Atm. P} \rightsquigarrow T > B.P$

(b)  $V.P = \text{Atm. P} \rightsquigarrow T = B.P$

(c)  $V.P < \text{Atm. P} \rightsquigarrow T < B.P$

capillary tube



## Objective:

- 1) To identify a compound based on its physical properties.
- 2) To learn how to properly assemble a boiling point apparatus.

The purpose of this experiment to identify a chemical compound from the data collected on its physical properties:

1. Solubility
2. Density
3. boiling point

**A physical property** is a characteristic of a substance that can be observed or measured without changing the identity of the substance.

The more physical properties include:

Color, odor, density, solubility, melting point, boiling point, and whether the substance is a (solid, liquid, or gas).

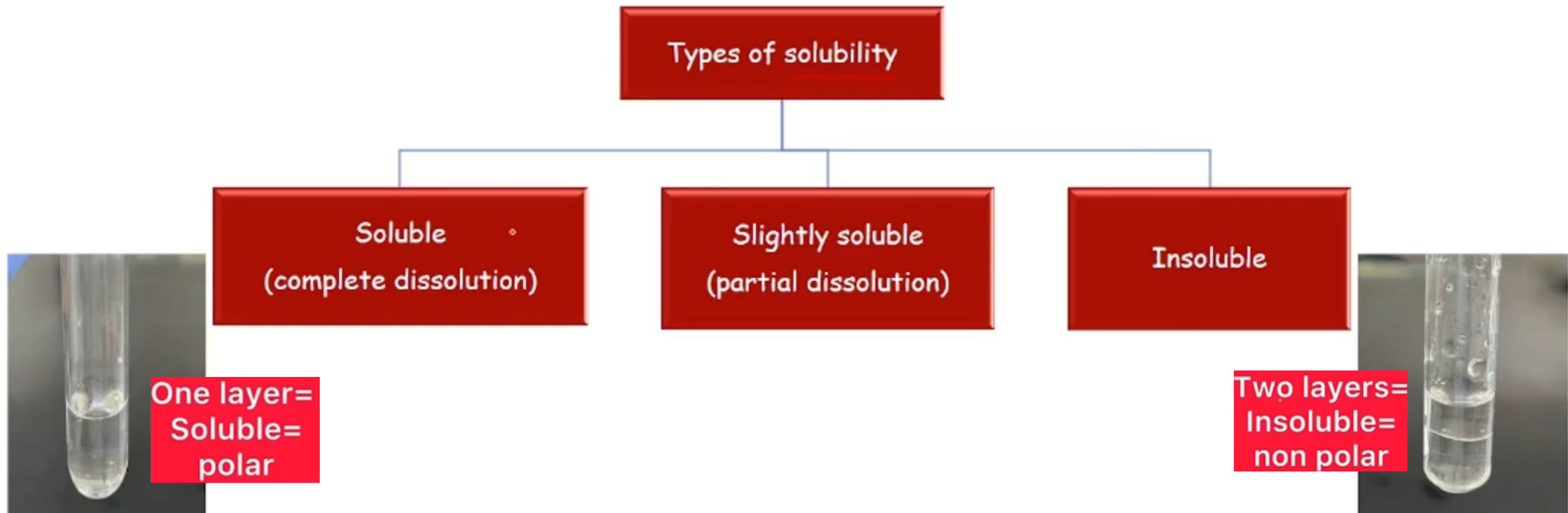


## 1) Solubility:

The maximum mass (usually in grams) of the substance that dissolves in a fixed mass (usually 100 g) of solvent at a given temperature.

A chemical has different solubility in different solvents due to the differences in the molecular composition of the solute from that of the solvent according to the solubility rule: like dissolve like

**Like dissolve like:** polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.



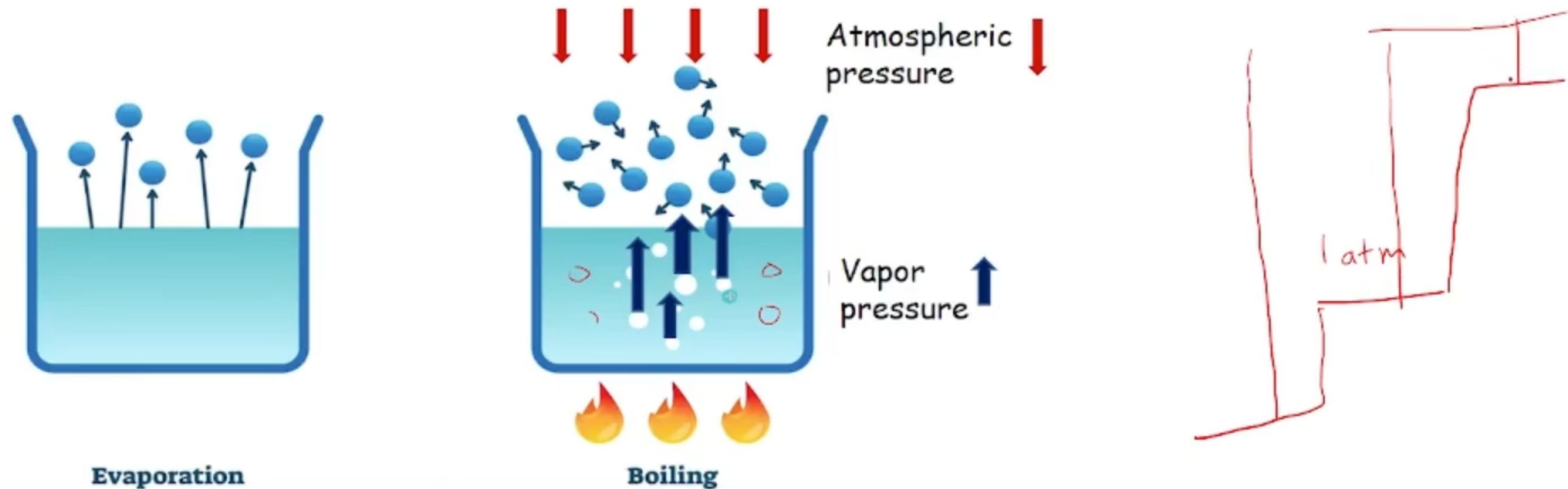
## 2) Density:

$$\frac{\text{mass} \uparrow}{\text{Volume} \downarrow}$$

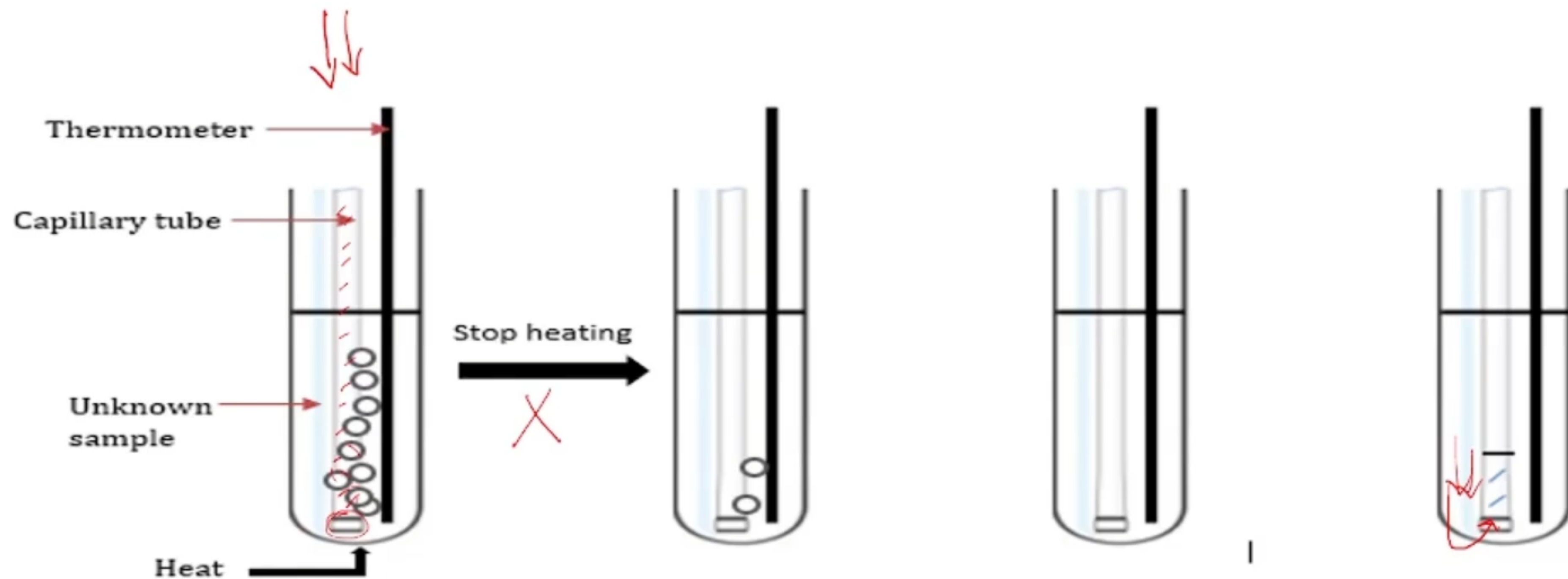
- The mass per unit volume.
- Substances with a large density have a large mass in a small volume.

### 3) Boiling point:

- The temperature at which the vapor pressure of the liquid equals of the atmospheric pressure above the liquid.
- Vapor pressure: is the pressure exerted by the vapor of a liquid when its equilibrium with its liquid phase.
- As the temperature of a liquid increases, the vapor pressure of the liquid increases until it equals the atmospheric. Bubbles of vapor begin to form throughout the liquid, and the liquid begins to boil.



- Boiling point  $\propto$  intermolecular force; (The greater the magnitude of the intermolecular forces, the higher will be boiling point)
- Normal boiling point ( if atmospheric pressure = 1 atm)



Rapid and continuous stream of bubbles escapes the capillary tube .

Vapor pressure > atm. pressure

$$T > B.P.$$

The stream of bubbles slows as the bath cools.

When the bubbles cease to escape and before the liquid re-enters the capillary tube.

Atm. Pressure = vapor pressure

At this point record the temperature (boiling point of the liquid)

$$T = B.P$$

The liquid enters the capillary tube.

Atm. Pressure > vapor pressure

$$T < B.P$$

*Table (1) physical properties of some common laboratory chemicals*

*Symbols used: i = insoluble, sls = slightly soluble, s = soluble*

Compound	Density(g/ml)	Boiling Point(C)	Solubility		
			H <sub>2</sub> O	C <sub>6</sub> H <sub>12</sub>	C <sub>2</sub> H <sub>5</sub> OH
Acetone	0.79	56	s	s	s
2-butanone	0.805	80	s	s	s
Cyclohexane	0.79	80.74	i	-	s
Cyclohexene	0.81	83	i	s	s
Ethanol	0.79	79	sls	s	-
Ethylacetate	0.90	77	s	s	s
Heptane	0.684	98	i	s	s
n-hexane	0.66	68	i	s	s
1-hexene	0.67	63	i	s	s
Isopropanol	0.79	83	s	s	s
Methanol	0.79	65	s	s	s
n-propanol	0.805	97	s	s	s
Water	1.00	100	-	i	s

60655565506

## Experiment (4)

### Identification of a Compound: Physical Properties

#### Post Lab

1. How does atmospheric pressure affect the boiling point of a liquid? Ask your instructor.

$$T_{b.p} \propto atm.P$$

2. If several drops of liquid unknown (Part B) cling to the pipette wall after delivery will the density of the unknown be reported too high or too low? Explain

$$\downarrow \quad \sim \quad d = \frac{m}{V}$$

3. (A). In part C.2 if the boiling point is recorded when bubbles are rapidly escaping the capillary tube, will it be recorded too high or too low? Explain.

$$\sim \quad V.P > atm.P$$

- (B). If the boiling point is recorded after the liquid enters the capillary tube (after the heat is removed), will it be recorded too high or too low? Explain

$$\sim \quad V.P < atm.P$$

1. Define:-

A. Solubility:

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B. Boiling point:

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2. A student's liquid unknown boils at approximately 69°C, is insoluble in water but soluble in cyclohexane. Its density is 0.65g / ml. Which chemical in table (1) is the unknown?

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3. What physical property, measurable in this experiment, distinguishes *cyclohexane* from *cyclohexene*?

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4. Using the apparatus described in this experiment, when should the boiling point of a liquid be recorded?

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If an unknown X is mixed with a small amount of cyclohexane, and gives two layers; then X is :

1. Polar
2. Non polar



# Ex. 50 Test of Cations & Anions

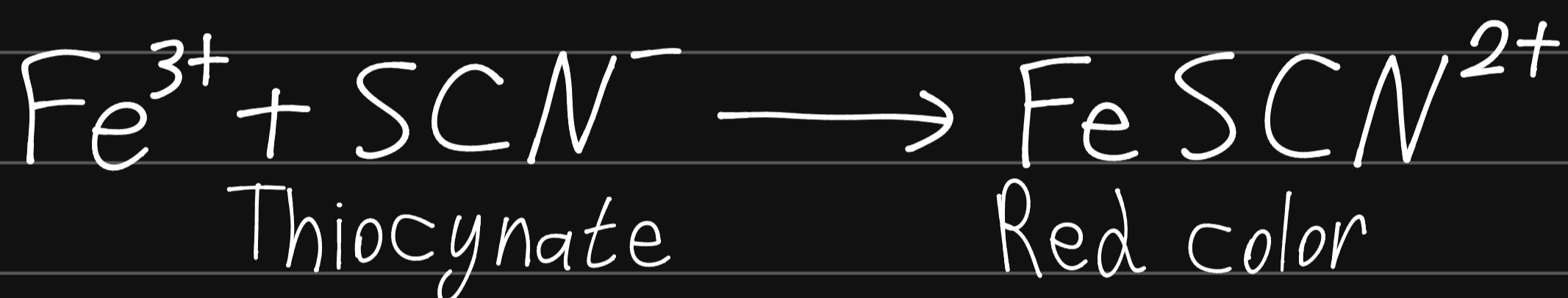
\* فيديو م. سامر

- Cations

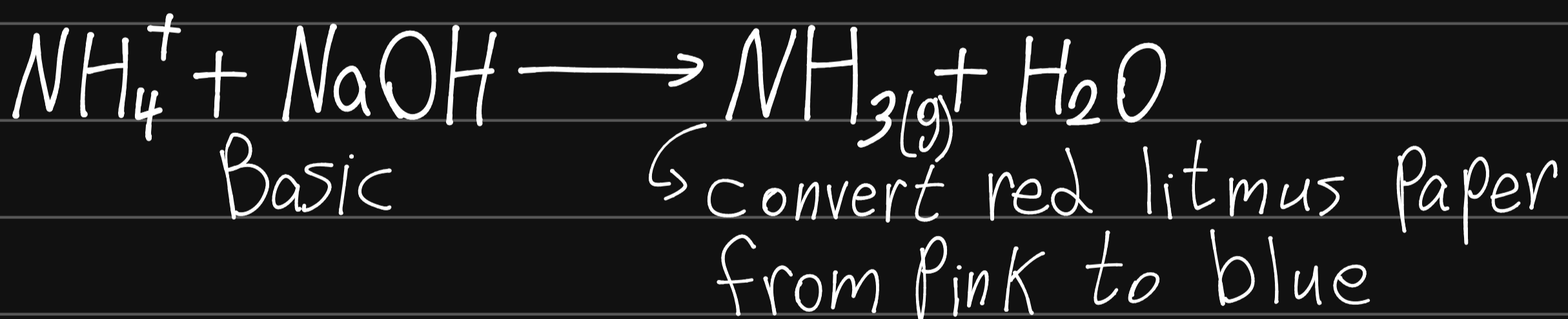
\*  $\text{Ca}^{2+}$ -test (Calcium test)



\*  $\text{Fe}^{3+}$ -test (Iron III)

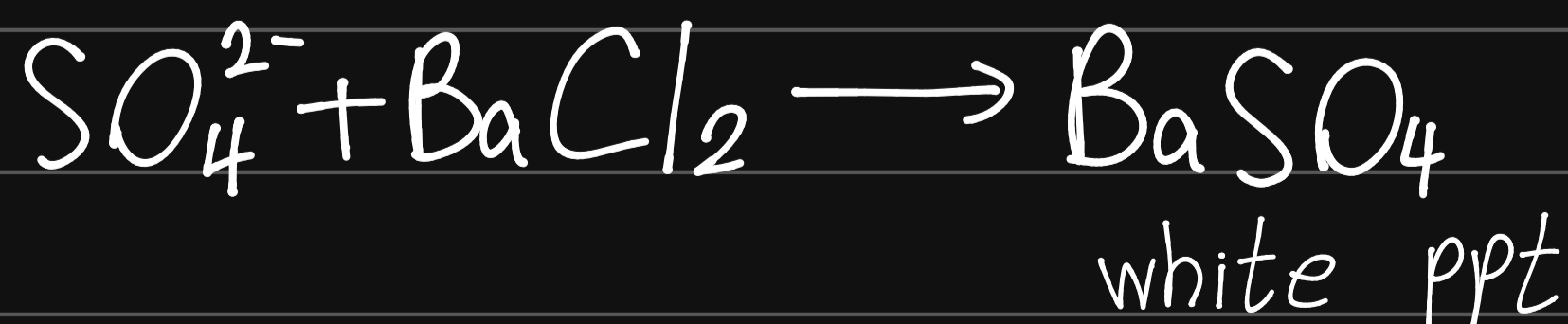


\*  $\text{NH}_4^+$ -test (Ammonium test)



- Anions

\*  $\text{SO}_4^{2-}$ -test (sulfate)



\*  $\text{Cl}^-$  &  $\text{Br}^-$  :



\*  $\text{HCO}_3^-$  - test : (Bicarbonate)



## *Objective:*

- To observe reactions of several common cations and anions.
- To use chemical tests to identify an unknown salt.

We have limited the possible ions in your unknown salt to the following:

Cations	Anions
calcium, $\text{Ca}^{2+}$ iron(III), $\text{Fe}^{3+}$ ammonium, $\text{NH}_4^+$	Sulfate, $\text{SO}_4^{2-}$ Chloride, $\text{Cl}^-$ Bromide, $\text{Br}^-$ Hydrogen carbonate, $\text{HCO}_3^-$

## Introduction:

The common procedure for testing any unknown sample is to make its solution and test this solution with various reagents for the ions present in it.

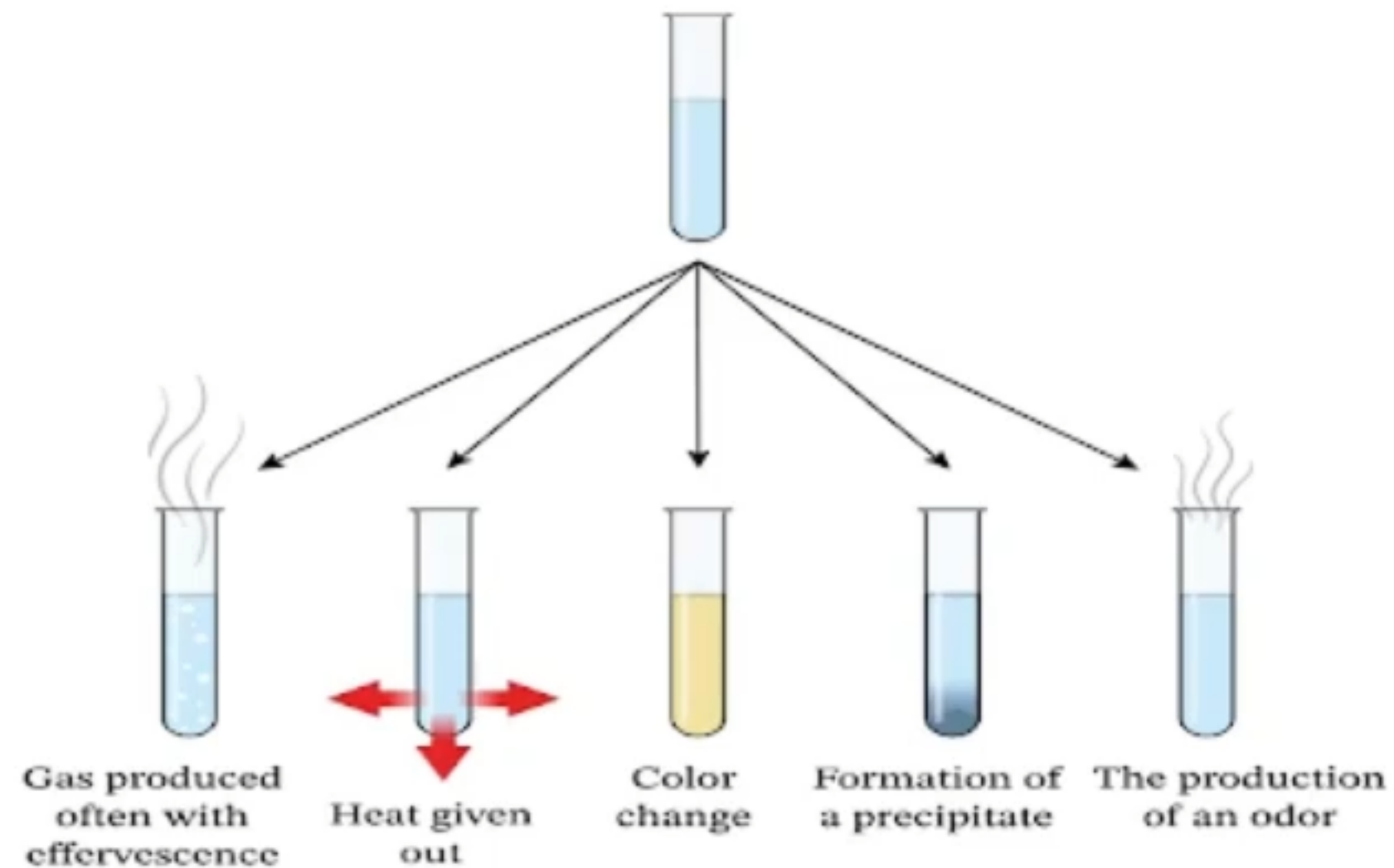
- So the analysis is carried out through the reactions between ions and reagents:

- Such reactions involve the following observation:

(a) Formation of a precipitate

(b) Change in colour

(c) Evolution of gas, etc.



- The procedure used in these tests is an example of **qualitative** analysis. You do not need to measure out volumes exactly using a measuring cylinder.

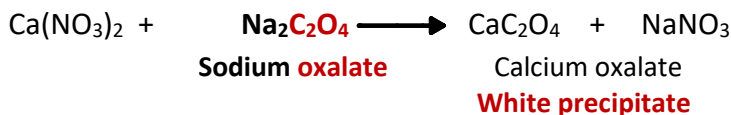
- Use **clean** test tubes and flasks. Dirty glassware will give confusing or misleading results. Before use, wash the test tubes several times with tap water and finally twice with distilled water from your wash bottle.

### Cations:

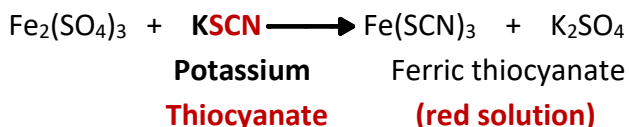
1.  $\text{Ca}(\text{NO}_3)_2 + \text{Na}_2\text{C}_2\text{O}_4 \rightarrow$  White precipitate
2.  $\text{Fe}_2(\text{SO}_4)_3 + \text{KSCN} \rightarrow$  red solution
3.  $\text{NH}_4\text{NO}_3 + \text{NaOH} \rightarrow$  red litmus paper blue

## Cations

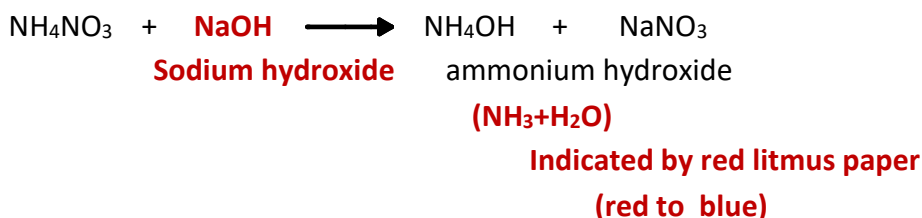
### 1. Calcium, $\text{Ca}^{2+}$



### 2. Ferric, Iron (III) $\text{Fe}^{3+}$

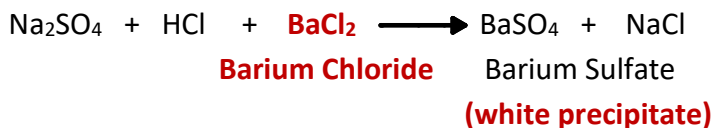


### 3. Ammonium, $\text{NH}_4^+$

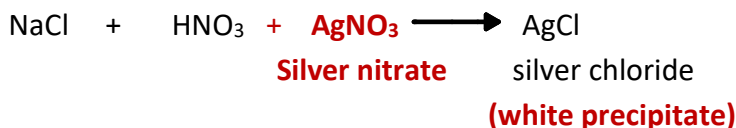


## Anions

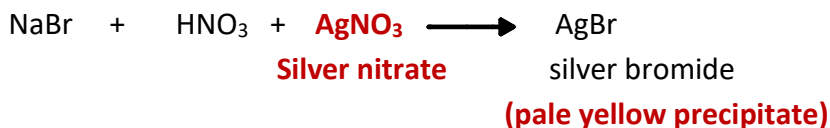
### 1. Sulfate, $\text{SO}_4^{2-}$



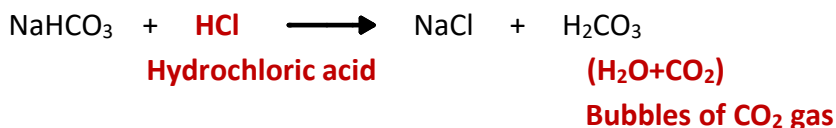
### 2. Chloride, $\text{Cl}^-$



### 3. Bromide, $\text{Br}^-$



### 4. Hydrogen carbonate, $\text{HCO}_3^-$



### Anions:

1.  $\text{Na}_2\text{SO}_4 + \text{HCl} + \text{BaCl}_2 \rightarrow$  white precipitate
2.  $\text{NaCl} + \text{HNO}_3 + \text{AgNO}_3 \rightarrow$  white precipitate
3.  $\text{NaBr} + \text{HNO}_3 + \text{AgNO}_3 \rightarrow$  pale yellow precipitate
4.  $\text{NaHCO}_3 + \text{HCl} \rightarrow$  Bubbles of  $\text{CO}_2$  gas

# Question

When a solution of an unknown salt X is treated with thiocyanate reagent, a blood-red solution is formed. When another portion of the same salt solution is treated with  $\text{BaCl}_2$ , a white precipitate appears. What is the probable chemical formula of the unknown salt X ?

Fill in the blanks:



ion	Reagent	Observation
Calcium	$\text{Na}_2\text{C}_2\text{O}_4$	-----
Chloride	-----	White precipitate
-----	$\text{SCN}^-$	Red solution
$\text{Br}^-$	Silver nitrate	-----
Hydrogen carbonate	-----	Bubbles of $\text{CO}_2$
-----	$\text{BaCl}_2$	White precipitate
Ammonium	Base	-----

# Ex 6° Molar mass of vol. Liq°-

\* فيديو م. ساصر

- Purpose° - Det of molar mass of a Vol. Liq by applying of Ideal gas law.

$$PV = nRT \Rightarrow PV = \frac{m}{M_r} RT$$

$$M_r = \frac{mRT}{pV} \approx \frac{g \cdot L \cdot atm \cdot mol^{-1} \cdot K^{-1} \cdot K}{atm \cdot L} = g/mol$$
$$\rightarrow M_r = \frac{dRT}{p}$$

\* Where°-

$M_r$ ° Molar mass (g/mol)

$m$ ° mass of vapor (g)

$$R = 0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$$

$T$ ° Temp (K) (of boiling water)

$$T(K) = t(C^\circ) + 273$$

$P$ ° Pressure (atm)

$$1 \text{ atm} = 760 \text{ torr} = 760 \text{ mmHg} = 101325 \text{ Pa}$$

V% Volume (L)

$$1 \text{ L} = 1000 \text{ cm}^3 = 1000 \text{ mL} = 1 \text{ dm}^3$$

d<sup>o</sup> density (g/L)

$$d = m/v$$

## Experiment (6)

### Molar Mass of a Volatile Liquid

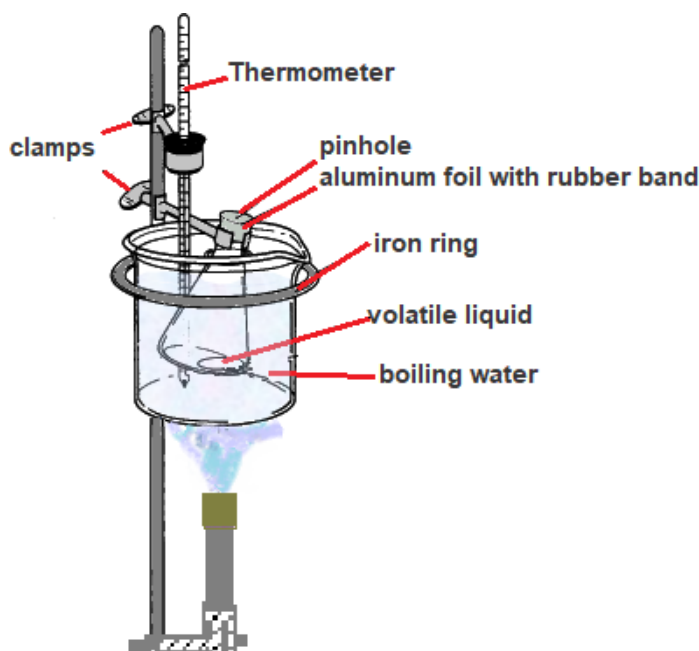
**Aim:** to determine the molar mass of an unknown volatile liquid using Dumas method and ideal gas law (boiling point of volatile liquid  $<100\text{ }^{\circ}\text{C}$ )

**Avogadro's law:** equal volumes of any two gases at the same temperature and pressure contain the same number of molecules.

**Ideal gas law:**  $PV = nRT$  can be applied on gases or (liquids in the gaseous state). Liquids must be appreciably volatile (high vapor pressure).

In this experiment the molar mass of a volatile liquid is determined as this procedure (**Dumas method**)

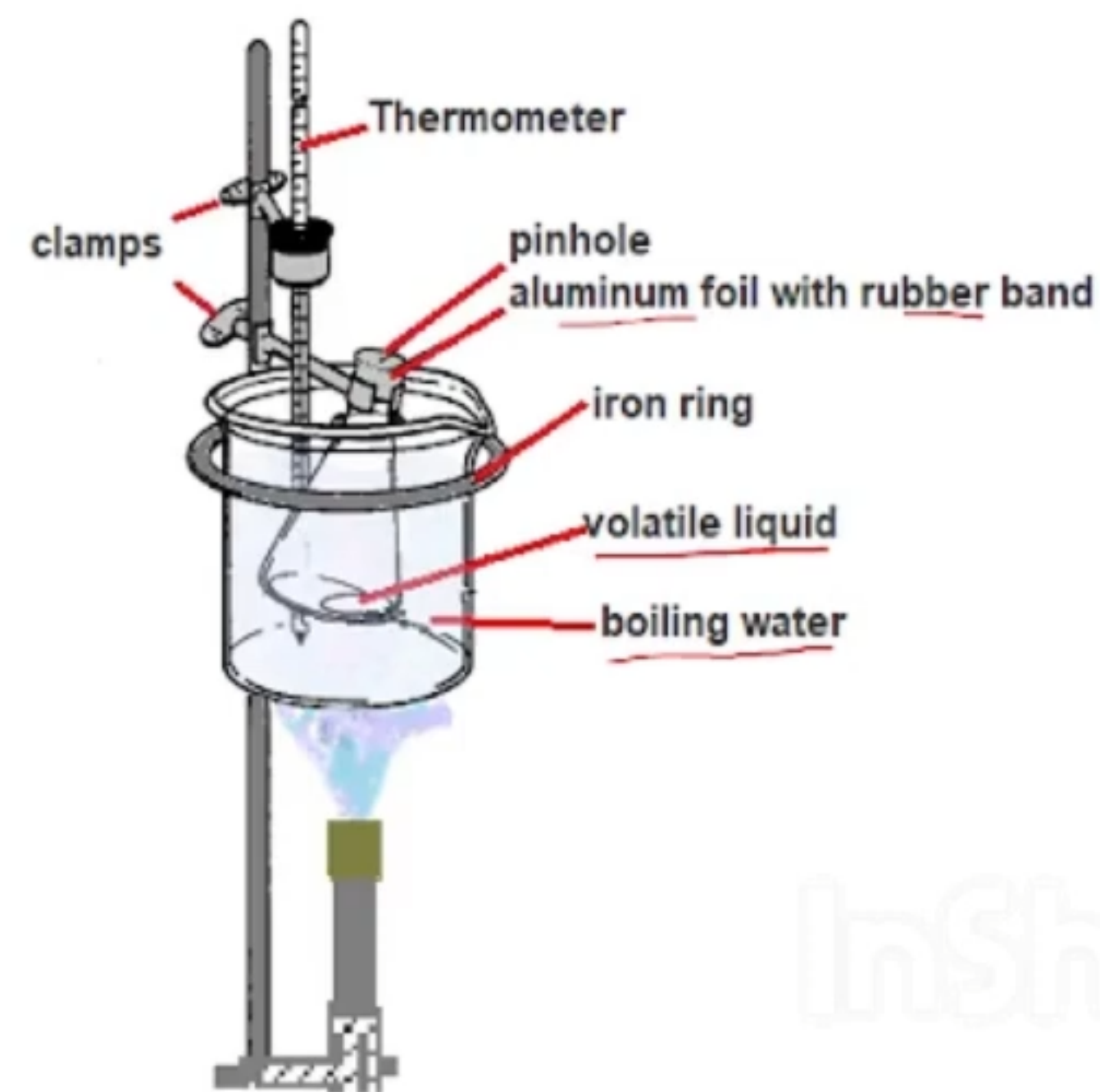
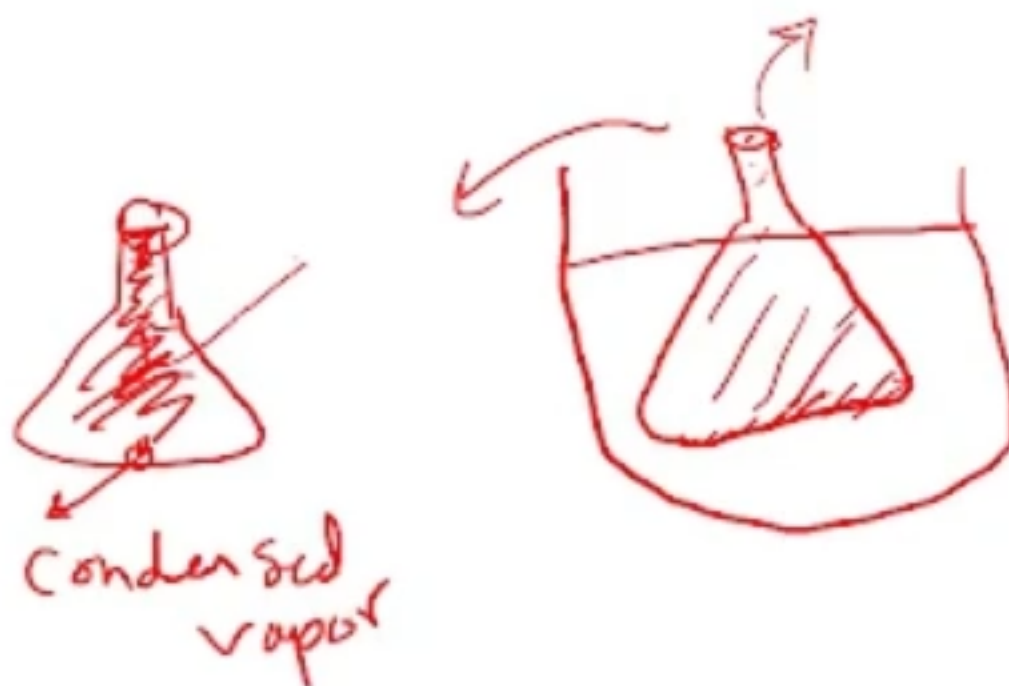
- 1) A volatile liquid is heated to a known temperature (above its boiling point) and allowed to escape from a container through tiny hole.
- 2) Once the liquid has vaporized, the container is cooled to room temperature.
- 3) The flask with contents is weighed again and the mass of the condensed vapor is calculated.
- 4) Volume of the container can be measured = volume of the gas
- 5) The high temperature can be measured (boiling point of water)
- 6) Pressure of the gas inside the container = room pressure (because the system is open to the atmosphere through the tiny hole.)



## The dumas method include the following steps:

- The volatile liquid is heated to known temperature (above the boiling point and allow the excess vapor to escape from tiny hole in container).
- Once the liquid has vaporize, the container is cooled to room temperature.
- Then the vapor remain in container at low temperature condenses to liquid and its mass is measured.
- The volume of container can be measured by measured the volume of water occupied by this container.
- The pressure of vapor gas is equal to atmospheric pressure (because the system is open to the atmosphere through the tiny hole).
- The high temperature can be measured (boiling point of water).
- The molar mass then can be calculated using the following law:

$$PV = \frac{\text{mass}}{M} RT$$



The molar mass is determined using the **ideal gas law**:

$$PV = nRT$$

The number of moles, n, is given by  $n = (m/M.w.)$

The ideal gas equation becomes:

$$PV = \frac{m}{MW} RT$$

$$M.w. = \frac{mRT}{PV}$$

Where: P = atmospheric pressure (atm), 1 atm = 760 mmHg = 760 torr = 101325 Pa

V = volume of vapor (L), 1 L = 1000 mL = 1 dm<sup>3</sup> = 1000 cm<sup>3</sup>

m = mass of condensed vapor (g)

M.w. = molar mass of volatile liquid (g/mol)

R = gas constant = 0.0821 atm L K<sup>-1</sup> mol<sup>-1</sup>

T = temperature of boiling water (K), K = °C + 273

Where  $\frac{m}{V} = d$  ( **d = density (g/L)** )

$$M.w. = \frac{dRT}{P}$$

$$d = \frac{MwP}{RT}$$

$$PV = nRT$$

Where:

P: atmospheric pressure (atm)

V: volume of vapor (L)

n: mole of vapor (mole)

T: temperature in kelvin

R: gas constant (0.08205 L atm/mol K)

## Experiment (6)

### Molar Mass of a Volatile Liquid

#### Prelab

- 1) For which of the following compounds can we determine its molar mass using the method described in this experiment? Give reasons. Benzene (b.p.78°C), Glycerol (b.p.180°C).

Benzene ,

b.p. benzene < b.p. water (100 °C)

- 2) A cylinder contains compressed hydrogen gas and the mass of the hydrogen is 20.0 g. What mass of oxygen would be contained in an identical cylinder at the same temperature and pressure?

$$PV = n_{H_2}RT$$

$$PV = n_{O_2}RT \quad (T, P, V, R) \text{ the same}$$

$$n_{O_2} = n_{H_2}$$

$$\text{mass } O_2 / 32 = 20 / 2$$

$$\text{mass } O_2 = 320 \text{ g}$$

## Experiment (6)

### Molar Mass of a Volatile Liquid

#### PostLab

1. A flask weighs 40.1305g when clean, dry, evacuated, 138.2410g when filled with water (density = 0.9970 g/ml) and 40.2487g when filled with a gaseous substance at 470.4 torr and 96°C. What is the molar mass (g/mol) of the gas? [R = 0.082 L.atm/mol. K]

Molar mass = 58.8 g/mol

$$MM = \frac{mRT}{PV}$$



Which of the following errors will increase the calculated molar mass of the volatile liquid?

1. The atmospheric pressure is assumed to be 747 torr instead of 755 torr which is recorded.
2. The volume is erroneously recorded 270 mL, while the correct volume should be 255 mL.

# Ex. 7% Det of Molar Volume of $H_2$ gas%

\* فيديو م. س. م.

\* Molar Volume% -  $\bar{V}$ % - Volume occupy per mol

- One mole of any ideal gas at STP will occupy volume = 22.4 L

$$(\bar{V})_{STP} = 22.4 \text{ L/mol}$$

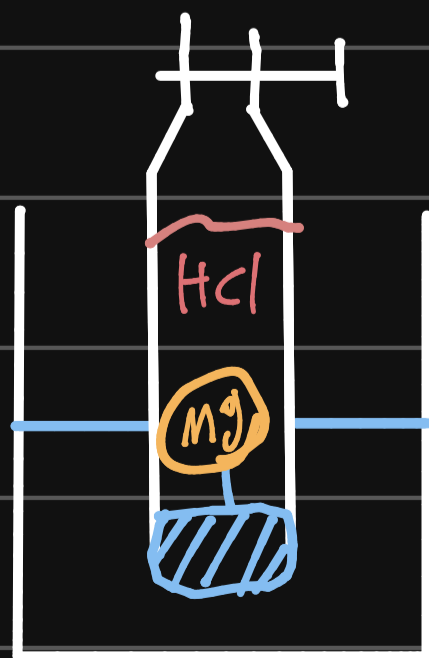
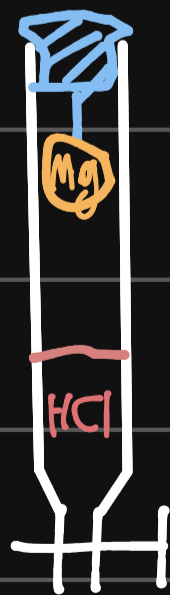
STP% - Standard Temp. Pressure

$$STP\% \quad t = 0^\circ\text{C}, \quad P = 1 \text{ atm}$$

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

$\sim 760 \text{ torr}$   
 $\sim 273$

$$\bar{V} = \frac{V}{\text{moles } H_2} \sim \text{moles } H_2 = \text{moles } Mg$$
$$= \frac{m \text{ Mg}}{M_r \text{ Mg}} \sim 24.31 \text{ g/mol}$$



$$\sim P_{tot} = P_1 + P_{H_2O}$$

$$P_1 = P_{tot} - P_{H_2O}$$

## Experiment (7)

### Molar Volume of Hydrogen Gas

#### Objective

Determine the molar volume of hydrogen gas and to compare it with the “22.4 L /mole” theoretical STP value.

Molar volume,  $\bar{V}$ : volume occupy per mol  
unit of molar volume is (L/mol)

One mole of any ideal gas at STP will occupy volume = 22.4 L

#### STP: Standard Temperature Pressure

Temperature: 0 °C (273 K), pressure : 1 atm (760 torr)

Combined gas law:  $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

Annotations: 760 torr (above  $P_2$ ), 273 K (next to  $T_2$ ), and a question mark (?) (above  $V_2$ ). Red arrows point from these annotations to their respective variables in the equation.

Classroom condition      STP condition

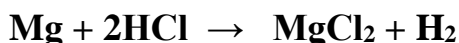
$P_1$ ,  $V_1$ ,  $T_1$  will be found experimentally,

Where  $P_1$  is hydrogen gas pressure,  $T_1$  is room temperature,  $V_1$  is the volume of gas produced.

$$\bar{V} = \frac{V_2}{\text{mol H}_2}$$

Hydrogen gas can be produced from the reaction of active metals with acids.

*In this experiment:*  $H_2$  will be generated from reacting magnesium metal (Mg) with HCl.



$H_2$  Will be collected in an inverted burette by water displacement.

*In our experiment we use small piece of Mg and excess of HCl, so Mg is the limiting reactant*

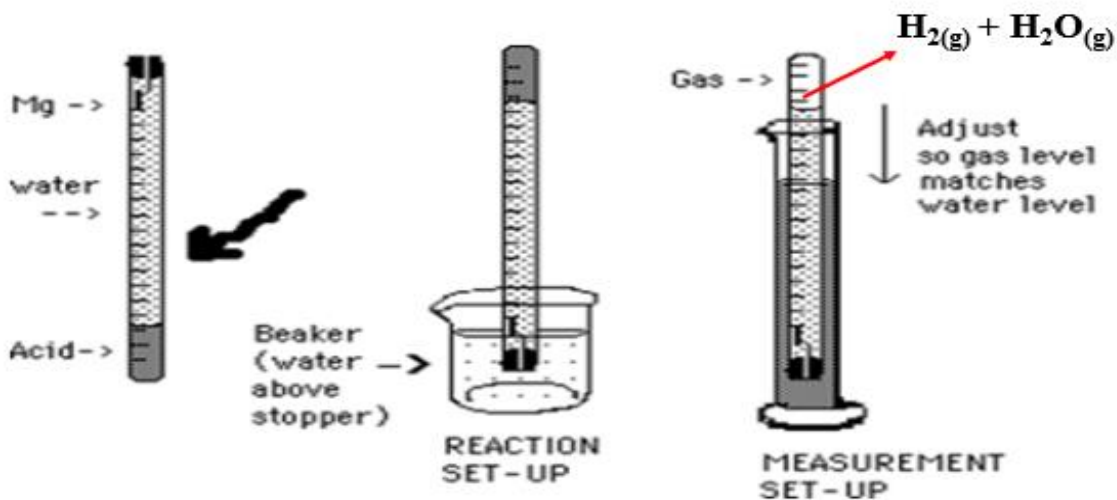
$$\text{Thus, mol H}_2 = \text{mol Mg, mol Mg} = \frac{\text{mass Mg}}{\text{M.w Mg}}$$

Hydrogen gas  $H_2$  will be collected over water, so that the pressure in the burette will be

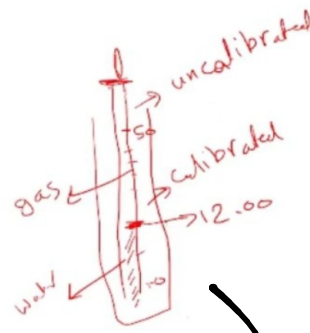
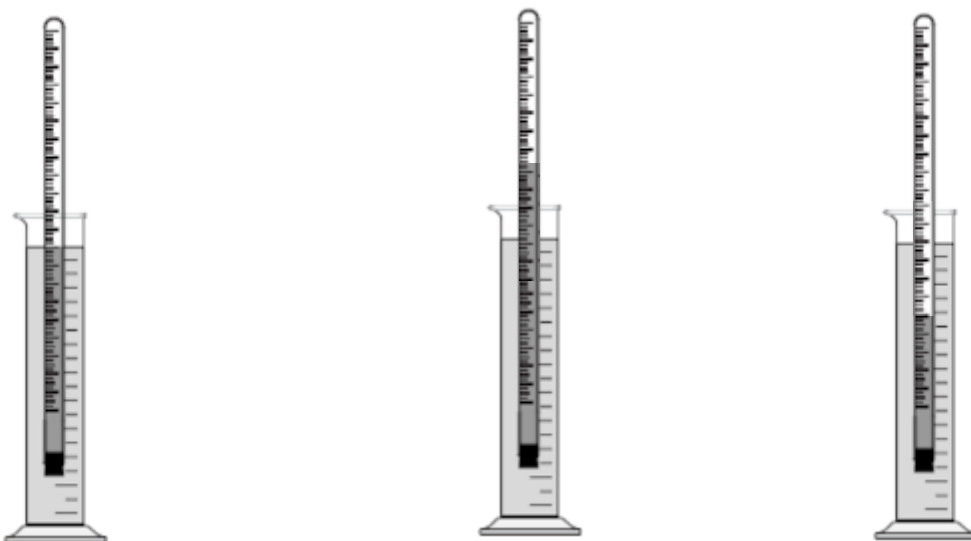
$$P_{\text{total}} = P_{H_2} + P_{H_2O}, \quad P_{\text{total}} = P_{\text{atm}}$$

$$P_{H_2} = P_{\text{atm}} - P_{H_2O}$$

\* $P_{H_2O}$  depends on the temperature of water



When the internal gas pressure ( $P_{\text{total}}$ ) matches the room atmospheric pressure ( $P_{\text{atm}}$ ) when the burette is adjusted into the large cylinder until the two liquid levels are equal.



$P_{\text{atm}} = P_{\text{inside burette}} (P_{\text{total}})$

$P_{\text{atm}} > P_{\text{inside burette}} (P_{\text{total}})$

$P_{\text{atm}} < P_{\text{inside burette}} (P_{\text{total}})$

*3.60 mL*                      *50 - 12*

$$\text{Volume of gas} = (\text{Volume of burette from 50 mL to sealed end}) + (50.0 \text{ mL} - \text{burette reading}).$$

## Vapor Pressure of Water at Various Temperatures

Temperature °C	Pressure torr	Temperature °C	Pressure torr
16	13.6	21	18.6
17	14.5	22	19.8
18	15.5	23	21.0
19	16.5	24	22.4
20	17.5	25	23.8

## PreLab

A student at The Hashemite University wants to determine experimentally the volume occupied by one mole of  $\text{H}_2$  gas at STP. She reacts 0.1471 g of Zn with excess  $\text{HCl}(aq)$  and collects 56.09 ml of gas over water at 22 °C and 757.8 torr. The vapor pressure of water at 22 °C is 19.8 torr.

1. Use the data given above to calculate:

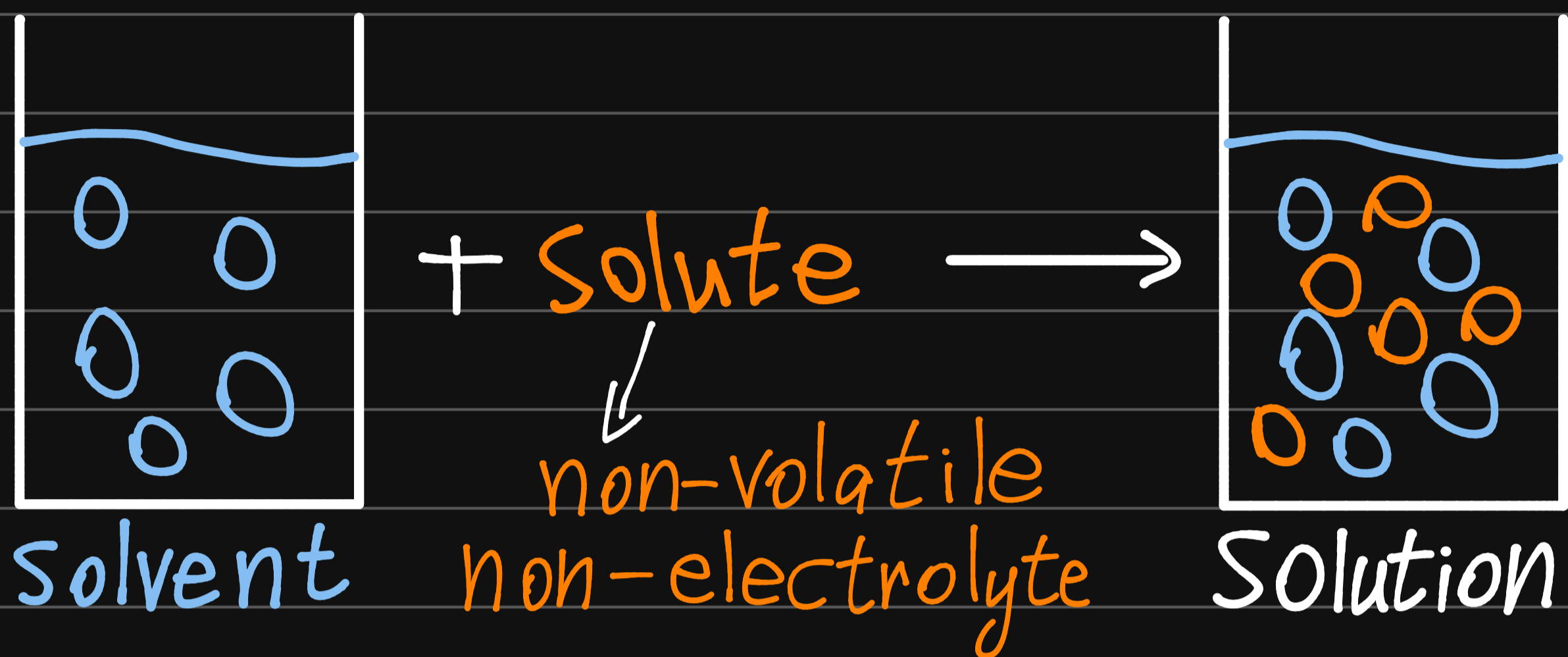
- (i) The volume occupied by one mole of dry  $\text{H}_2$  at 22 °C and 760 torr.
- (ii) The volume occupied by one mole of dry  $\text{H}_2$  at STP.

Note: do NOT simply use the ideal gas law equation!

2. Name the gas laws which you used in your calculations.

# Ex 8<sup>o</sup> Colligative properties:-

\* اللوز



\* the effect of adding <sup>non-volatile</sup><sub>non-electrolyte</sub> Solute:-

- Boiling point elevation
- freezing point depression
- Vapor pressure lowering

\* all are proportional with the amount of solute.

Boiling point elevation

freezing point depression

Vapor pressure lowering

}  $\propto$  amount of solute

$\Delta T_f \propto m$  → molality.  
 $T_{\text{solvent}} - T_{\text{solution}}$

$$\Delta T_f = K_f \times m \rightarrow \frac{C^\circ \text{ Kg}}{\text{mol}}$$

$$\Delta T_f = K_f \frac{\text{mol solute}}{\text{mass solvent (Kg)}}$$

$$\Delta T_f = K_f \frac{\text{mass solute (g)}}{M_{r \text{ solute}} \times \text{mass solvent (Kg)}}$$

$$M_{r \text{ (g/mol)}} = \frac{K_f \left( \frac{C^\circ \text{ Kg}}{\text{mol}} \right) \times \text{mass solute (g)}}{\Delta T_f (C^\circ) \times \text{mass solvent (Kg)}}$$

\* فيديو م. سافر

- When a non vol. solute is dissolved in certain amount of solvent there will be f.p.d., B.P. elevation and V.P. inc<sup>↑</sup> of solvent

~ f.p.d<sup>o</sup> - freezing Point depression  $\equiv \Delta T_f$

$\Delta T_f \propto$  amount of solute in solvent (quantitative)

$$\Delta T_f \propto m \Rightarrow \Delta T_f = K_f m$$

$$m \text{ molality} = \frac{\text{moles solute}}{\text{mass solvent (Kg)}}$$

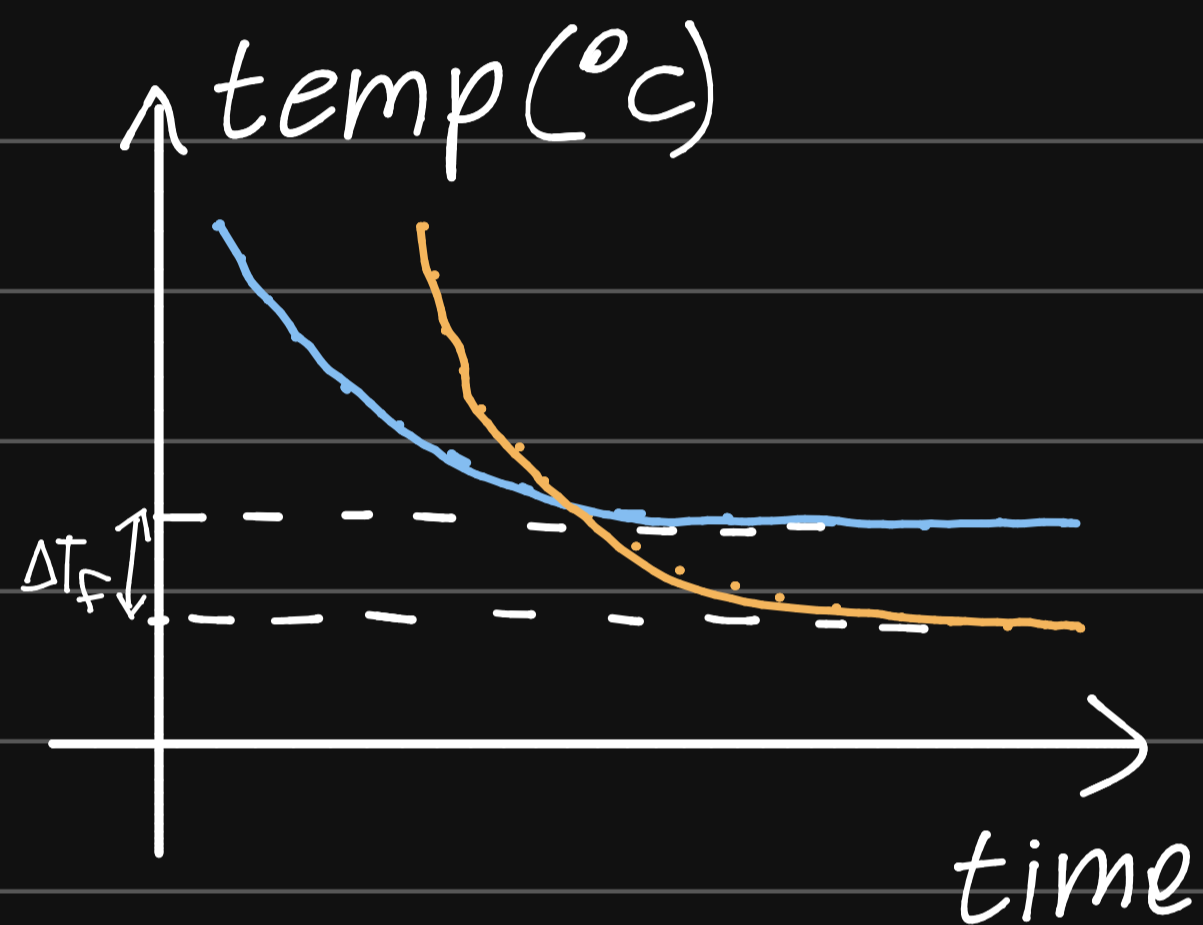
- f.p.d means the f.p of solvent will be lower than if it were pure
- f.p° - Temp at which solid & liq states coexists in eq. at atmospheric pressure

$$\Delta T_f = K_f \times m$$

$$\Delta T_f = K_f \frac{\text{mol solute}}{\text{mass solvent (Kg)}}$$

$$\Delta T_f = K_f \frac{\text{mass solute (g)}}{M_{r \text{ solute}} \times \text{mass solvent (Kg)}}$$

$$M_r \left( \frac{\text{g}}{\text{mol}} \right) = \frac{K_f \left( \frac{\text{C}^\circ \text{Kg}}{\text{mol}} \right) \times \text{mass solute (g)}}{\Delta T_f (\text{C}^\circ) \times \text{mass solvent (Kg)}}$$



$K_f$ ° f.p.d const for solvent

## Experiment (8)

### Colligative Properties: Molar Mass Determination

#### Objective:

To determine the molar mass of a non-volatile, non-electrolyte by observing the difference between the freezing points of a solvent and a solution.

#### Introduction:

The addition of a non-volatile solute to a solvent produces characteristic changes in a solvent's physical properties. These changes are: **vapor pressure lowering**, **boiling point elevation** and **freezing point depression**. For example when antifreeze (ethylene glycol) is added to an automobile's cooling system it will prevent freeze-up in the winter and boiling in the summer because the antifreeze-water solution has a lower freezing point and a higher boiling point than pure water.

Vapor pressure lowering, boiling point elevation, and freezing point depression are called **colligative properties** of solutions. These properties are governed by the number, rather than the type, of solute particles dissolved in the solvent.

The change in colligative properties of a solution is directly proportional to the amount of solute dissolved in the solvent. The freezing point depression,  $\Delta T_f$ , and the boiling point elevation,  $\Delta T_b$ , are proportional to the molality,  $m$ , of the solute in the solution:

$$\Delta T_f = K_f m \quad (1)$$

$$\Delta T_b = K_b m \quad (2)$$

Where

$K_f$  and  $K_b$  are the molar freezing point and boiling point constants for the solvent.

$$m = \frac{\text{mol solute}}{\text{Kg solvent}} = \frac{(\text{g / mol. wt.}) \text{ solute}}{\text{Kg solvent}} \quad (3)$$

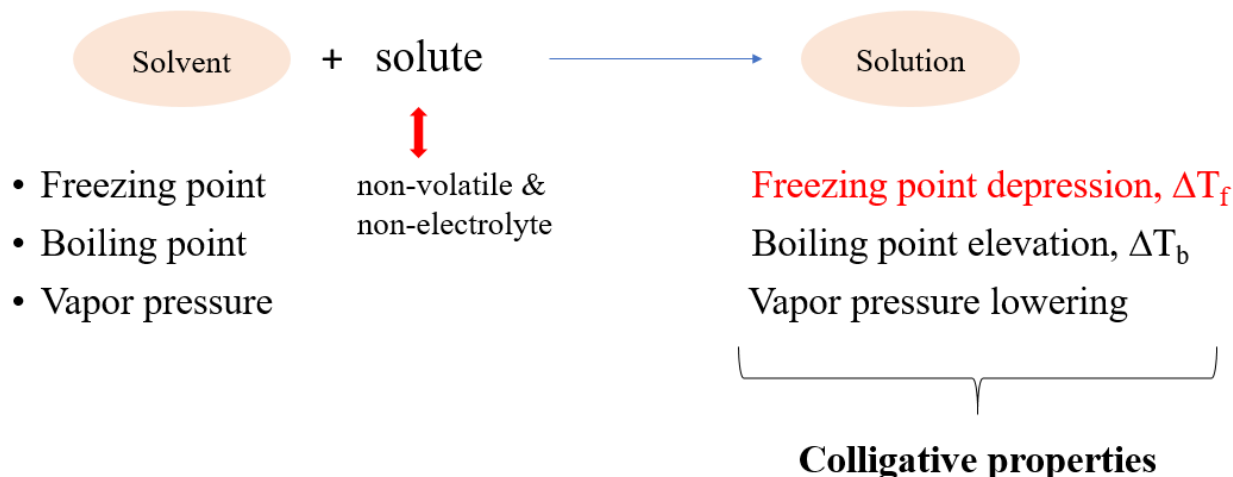
In this experiment you will determine the molar mass of a non-volatile, non-electrolyte (solute) in cyclohexane by measuring the freezing points for the pure cyclohexane and the solution, as well as the mass of solute and solvent. From  $K_f$  and equations (1) and (3), you will calculate the formula weight of the solute.

Cyclohexane's freezing point and that for the solutions are obtained from a cooling curve; a graph representing a decreasing temperature as a function of time (Figure 1). Cyclohexane's cooling curve reaches a plateau at its freezing point; extrapolation of the plateau to the temperature axis determines its freezing point. The solution's cooling curve does not reach a plateau, but continues to decrease slowly as the cyclohexane freezes. Its freezing point is determined at the intersection of two lines drawn tangent to the curves above and below the freezing point (See Figure 1).

## Experiment (8)

### Colligative Properties

When a non volatile solute is dissolved in certain amount of solvent, some of physical properties of solvent are changed as follows:



Colligative properties: depend on amount of solute (not on the type) “quantitative”.

As amount of solute was dissolved in solvent increase  $\longrightarrow$  freezing point decrease  $\longrightarrow$  freezing point depression increase

### **Objective:**

Determine the molar mass of a non-volatile solute by observing the difference between the freezing points of a solvent and a solution.

Freezing point of a solution is lower than the freezing point of pure solvent.

*In our experiment:* the molar mass of non-volatile, non-electrolyte solute will be determined by measuring the freezing point of pure solvent (cyclohexane) and freezing point of solution (solute and solvent).

The freezing point depression is proportional to the molality,  $m$ , of the solute in the solution:

$$\Delta T_f \propto m$$

$$\Delta T_f = K_f m,$$

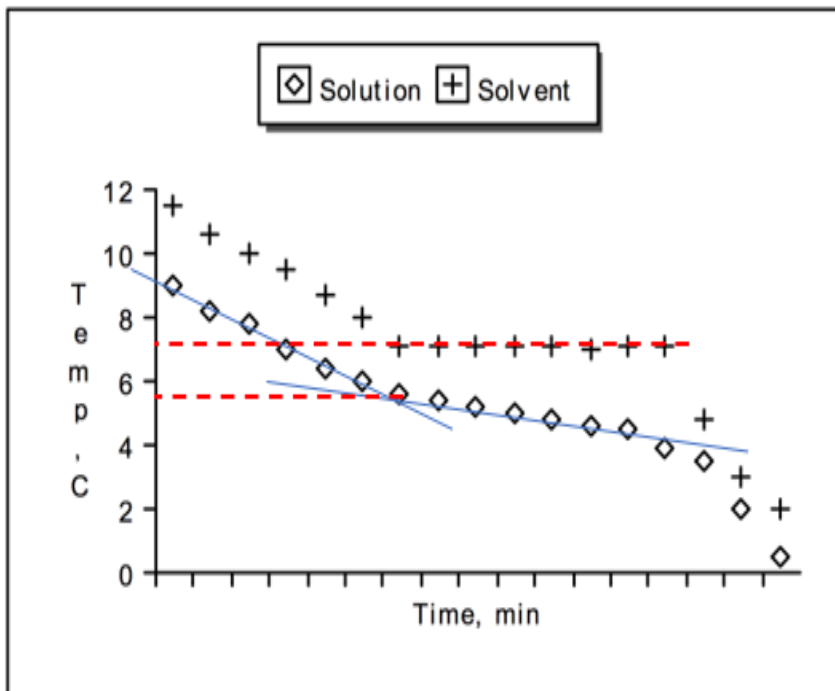
$K_f$  : molar freezing point constant for the solvent ( $^{\circ}\text{C}\cdot\text{Kg/mol}$ ) or ( $^{\circ}\text{C}/m$ )

$$\Delta T_f: T_{f,\text{solvent}} - T_{f,\text{solution}} (^{\circ}\text{C})$$

$$m: \text{molality or molal concentration (mol/Kg)} = \frac{\text{mol solute}}{\text{mass solvent (kg)}} = \frac{\left(\frac{\text{mass}}{M_w}\right)\text{solute}}{\text{mass solvent (kg)}}$$

$$\Delta T_f = K_f \frac{\left(\frac{\text{mass}}{M_w}\right)\text{solute}}{\text{mass solvent (kg)}}$$

$$M.w. = \frac{K_f (^{\circ}\text{C}\cdot\frac{\text{Kg}}{\text{mol}}) \cdot \text{mass solute (g)}}{\Delta T_f (^{\circ}\text{C}) \cdot \text{mass solvent (kg)}}$$



**Freezing point of solution lower than freezing point of pure solvent**

**Cooling curves for solvent and solution**

## Experiment (8)

### Colligative Properties: Molar Mass Determination

#### PostLab

$$\text{Mol. wt. solute} = \frac{k_f \text{ mass solute}}{\Delta T_f \text{ mass solvent (kg)}}$$

1. If the solution's freezing point is erroneously read 0.2°C lower than it should be, will the unknown's calculated molar mass be too high or too low? Explain!

\_\_\_\_\_  $\Delta T_f = T_f \text{ solvent} - T_f \text{ solution}$ ,  $T_f \text{ solution decrease} \rightarrow \Delta T_f \text{ increase}$ , \_\_\_\_\_  
\_\_\_\_\_ If  $\Delta T_f$  increase, the molar mass decrease \_\_\_\_\_

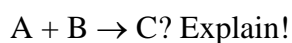
2. How will the freezing point change of cyclohexane be affected (compared with the freezing point change by a non-volatile, non-associating and non-dissociating solute) by:

- (a) A non-volatile solute that dissociates? Explain.

**Dissociate**

\_\_\_\_\_  $AB \rightarrow A + B$ , dissociation, moles inc,  $\Delta T_f$  inc, molar mass dec \_\_\_\_\_  
\_\_\_\_\_ 1 mol                  2 mol \_\_\_\_\_

- (b) Two solutes that react according to the equation,



**Associate**

\_\_\_\_\_  $2 \text{ mol} \rightarrow 1 \text{ mol}$ , moles dec,  $\Delta T_f$  dec, molar mass inc \_\_\_\_\_

3. If some solute adheres to the test tube's wall in Part B.1, is the freezing point change greater or less than it should be? Explain!

\_\_\_\_\_ Mass solute dec, mol solute dec,  $\Delta T_f$  dec \_\_\_\_\_

1. Students prepared two cyclohexane solutions having the same mass of solute. However Student 1 used 13 g of cyclohexane, Student 2 used 15 g. Which student will observe the larger freezing point change? Explain.

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2. A 0.597 g sample of a non-electrolyte dissolves in 20.0g of cyclohexane. The freezing point depression is 3.62 °C. What is the molar mass of the non-electrolyte? ( $K_f$  for Cyclohexane is 20.0 °C Kg/mol).

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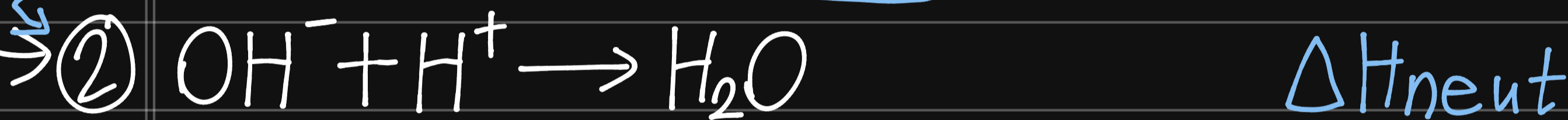
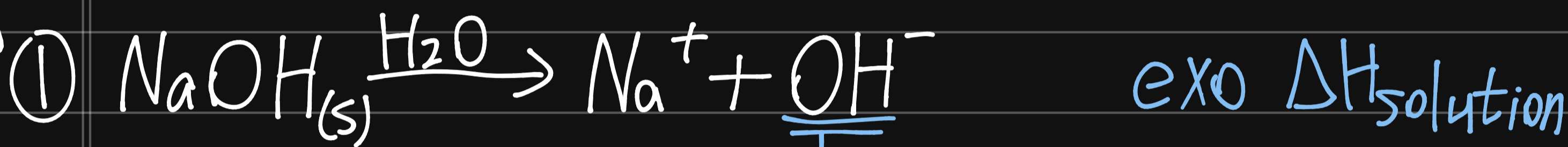
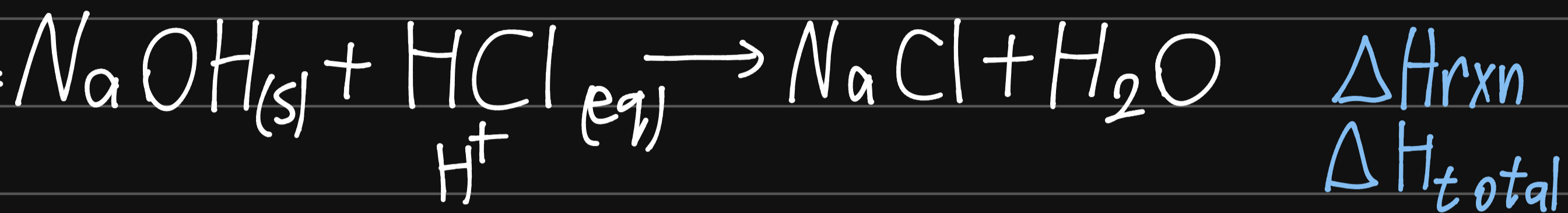
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# Ex 9% - Calorimetry

\* heat of neutralization % -

\* الحرارة



$$\Delta H_{\text{rxn}} = \Delta H_{\text{solution}} + \Delta H_{\text{neut}}$$

$$\Delta H_{\text{neut}} = \Delta H_{\text{solution}} - \Delta H_{\text{rxn}}$$

$\downarrow \times 10^3 / \text{mol}$

$$\text{J} \rightarrow \text{kJ/mol} \sim 10^3 \text{ J/mol H}_2\text{O}$$

$\uparrow \times \text{mol} \times 10^3$

mol L.R (NaOH)

$$\Delta H = m_{\text{solution}} * S * \Delta T \rightarrow S_{\text{water}} = 4.184 \text{ J/g} \cdot \text{C}^\circ$$

$$\Delta H = (m_{\text{NaOH}} + m_{\text{H}_2\text{O}/\text{HCl}}) * S * (T_{\text{solution}} - T_{\text{solvent}})$$

J      g      J/g.C<sup>o</sup>      C<sup>o</sup>

\* فيديو م. س. ص

- Heat of soln° amount of heat required or released when a certain amount of solute is dissolved in certain amount of solvent

- Heat of sol  $\equiv \Delta H_{solu}$

$$\Delta H(J) = \text{mass solu (g)} * \text{s.p. heat } \left(\frac{J}{g \cdot ^\circ C}\right) * |\Delta t (^{\circ}C)|$$

$\Delta H$  can have +ve or -ve sign

the sign is added and it depends on  $\Delta t$

- If  $t_f > t_i \Rightarrow \Delta t \text{ } ^\circ +ve$ ,  $\Delta H \text{ } ^\circ -ve$  (exothermic)  
(sign is add to  $\Delta H$ )

- If  $t_f < t_i \Rightarrow \Delta t \text{ } ^\circ -ve$ ,  $\Delta H \text{ } ^\circ +ve$  (endothermic)  
(sign is add to  $\Delta H$ )

Ⓘ  $H_2O + NaOH \text{ solu}$   
 $\Delta H_{solu} = m_{solu} * S.P * |\Delta t|$

Ⓜ  $HCl + NaOH \text{ } \Delta H_{solu} + \Delta H_{nat}$   
 $\Delta H_{rxn} = \Delta H_{solu} + \Delta H_{nat}$

$\Delta H_{\text{nut}} \equiv$  Heat of Neutralization<sup>o</sup>-heat that is released from rxn of an acid with a base and can't be experimentally determined.

$$\Delta H_{\text{nut}} = \Delta H_{\text{rxn}} - \Delta H_{\text{solu}}$$

(HCl+NaOH) (H<sub>2</sub>O+NaOH)

- s.p. heat<sup>o</sup>- amount of heat required to raise the temp. of 1g sample by 1°C
- $\Delta H^{\circ}$  quantitative property

$$\Delta H \text{ Per mole NaOH} = \frac{\Delta H \text{ (J)}}{\text{mole NaOH (mol)}} = \text{---} \frac{\text{J}}{\text{mol}}$$

$$\Delta H \text{ Per mass NaOH} = \frac{\Delta H \text{ (J)}}{\text{mass NaOH (mol)}} = \text{---} \frac{\text{J}}{\text{g}}$$

# *Objective:*

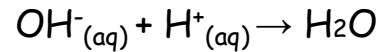
To measure the heat of the reaction:

- Heat of solution
- Heat of neutralization

# Heat of neutralization

**Heat of neutralization ( $\Delta H_{\text{neu}}$ ):** the heat **evolved**, at constant pressure, when acids are neutralized with bases.

Or is the heat released when 1 mole of water is formed from the neutralization between one mole of hydrogen ions,  $\text{H}^+$  from an acid and one mole of hydroxide ions,  $\text{OH}^-$  from a base.



\*Heat of neutralization is always exothermic;  $\Delta H_{\text{neu}}$  **-ve**

Heat of neutralization can not be determined experimentally.

It can be determined from the difference between the heat of reaction obtained by dissolving solid NaOH in aqueous HCl solution (**part B**) and the heat of solution obtained by dissolving solid NaOH in water (**part A**).

$$\Delta H_{\text{neu}} = \Delta H_2 - \Delta H_1$$

$$\Delta H = m \times s \times \Delta T, \Delta H \text{ in J}$$

**m** : mass of solution (mass of solvent + mass of solute) (g)

**s**: specific heat : the quantity of heat required to raise (increase) the temperature of one gram of a substance by one Celsius degree ( $J/g.^{\circ}C$ );  $s = 4.184 J/g.^{\circ}C$  for water

**$\Delta T$** : temperature change ( $^{\circ}C$ )

Standard unit of neutralization reaction = KJ/mol

$$\Delta H \text{ in KJ/mol} = \Delta H/1000 \div n(H_2O_{(l)})$$

\* If  $T_2 > T_1$ , the reaction is exothermic (the heat is released)

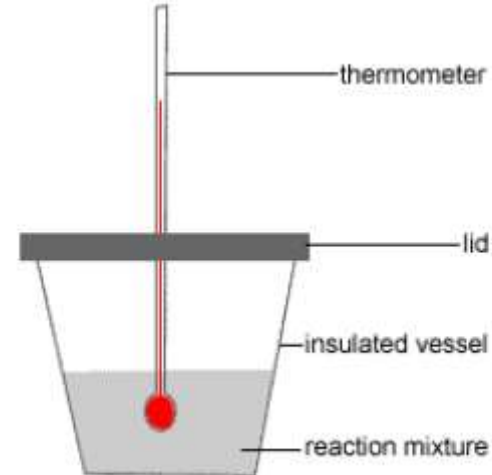
**-ve sign should be added to  $\Delta H$**

\* If  $T_1 > T_2$ , the reaction is endothermic (the heat is absorbed)

**$\Delta H$  is positive**

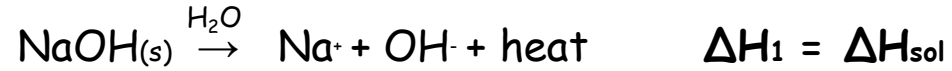
# calorimeter

- A calorimeter is a device used to measure heat changes during a chemical reaction.
- In this experiment, a coffee cup calorimeter made of polystyrene foam (Styrofoam) is used, which provide good insulation. This insulation minimizes heat loss, allowing accurate measurement of energy changes.



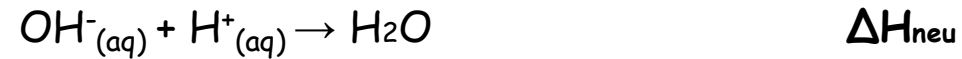
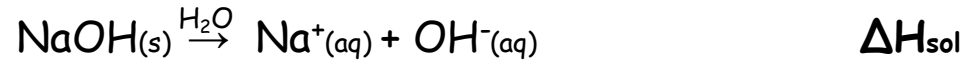
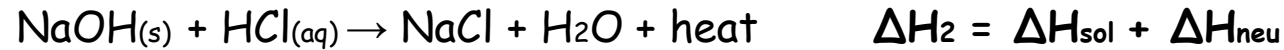
## Part A

**Heat of solution ( $\Delta H_{\text{sol}}$ ):** amount of heat **released** or **absorbed** when solute being dissolved in solvent.



## Part B

**Heat of reaction; combined heat of solution and heat of neutralization:**



$$\Delta H_{\text{neu}} = \Delta H_2 - \Delta H_1$$

### Example 1

A 2.00 g sample of solid CsOH is dissolved in 200.0 mL of water in a calorimeter. The temperature of the water was raised from 22.3 to 23.4 °C. Calculate the heat of solution in kJ/mol. (assume the specific heat of the solution to be 4.184 J/g °C and the density of the solution to be 1.00 g/mL).

- Temperature change,  $\Delta T = 23.4 - 22.3 = 1.1 \text{ }^\circ\text{C}$
- Total mass of solution, m, (200.0 mL water) (1.00 g/mL) + 2.00 g CsOH = 202.0 g
- Heat absorbed by the solution = heat evolved by CsOH =  
Mass of solution x specific heat of solution x  $\Delta T =$   
 $202.0 \text{ g} \times 4.184 \text{ J/g }^\circ\text{C} \times 1.1^\circ\text{C} = -930 \text{ J}$
- Moles of CsOH = mass/Molar mass =  $2.00 \text{ g} / 149.9 \text{ g/mol} = 0.01334 \text{ mol}$
- Heat evolved per mole of CsOH =  $930 \text{ J} / 0.01334 \text{ mol} = -7.0 \times 10^4 \text{ J/mol}$
- Heat of solution in KJ/mole =  $(-7.0 \times 10^4 \text{ J/mol}) \times (1 \text{ kJ}/10^3 \text{ J}) = -70 \text{ kJ/mol}$

Note: the minus sign shows that the reaction is exothermic; i.e heat is given off (evolved) by the system.

### Example 2

A 2.00 g sample of solid CsOH reacted with 200.0 mL of an aqueous solution of hydrogen chloride (HCl) in a calorimeter, the temperature of the solution increased from 22.3 to 24.3 °C. Calculate the heat of reaction in kJ/mol? (Assume the specific heat of the solution to be 4.184 J/g °C and the density of the solution to be 1.00 g/mL)

- Temperature change,  $\Delta T = 2.0 \text{ }^\circ\text{C}$
- Total mass of solution = 202.0g
- Heat absorbed by the solution =  $202.0 \times 4.184 \times 2.0$   
 $= 1.7 \times 10^3 \text{ J}$
- Moles of CsOH = 0.01334 mol
- Heat evolved per mole of CsOH =  $1.3 \times 10^5 \text{ J/mol}$
- Heat of reaction in kJ/mol =  **$-1.3 \times 10^2 \text{ kJ/mol}$**

Note: The difference in the heat of reaction (Example 2) and the heat of solution (Example 1) is the **heat of neutralization**. So the heat of neutralization of CsOH with HCl is 60 KJ/mol

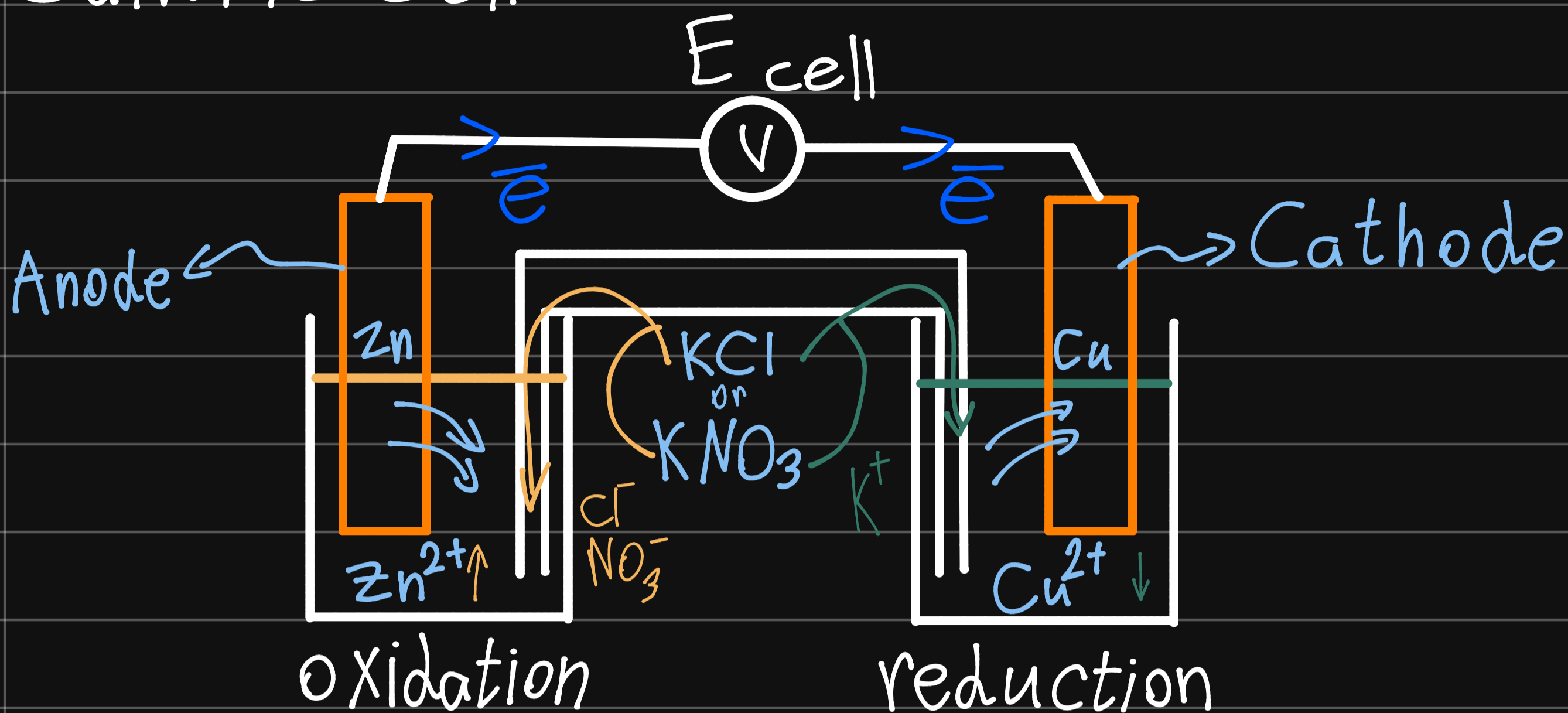
$$1.3 \times 10^2 - 70 = 60$$

$$\text{Heat of neutralization} = - 60 \text{ kJ/mol}$$

# Ex. 10° - Electrochemistry

\* اللوز

## \* Galvanic cell°-



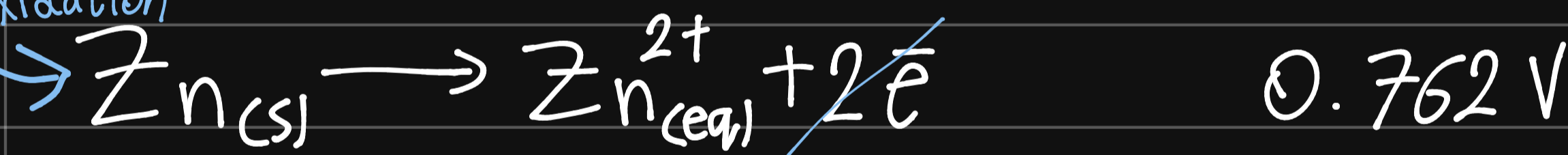
oxidation



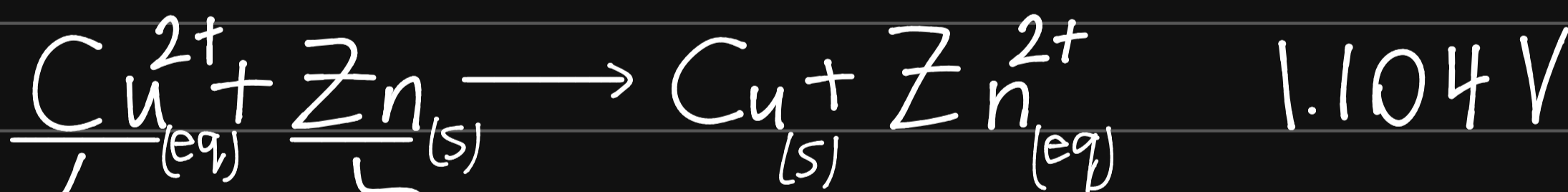
reduction



oxidation



net. equ



oxidizing agent

reducing agent

## \* Line cell°-

Oxidation || reduction



\*  $E_{cell}^{\circ}$

$E$  at  $25^{\circ}$  cell Potential (Voltage)

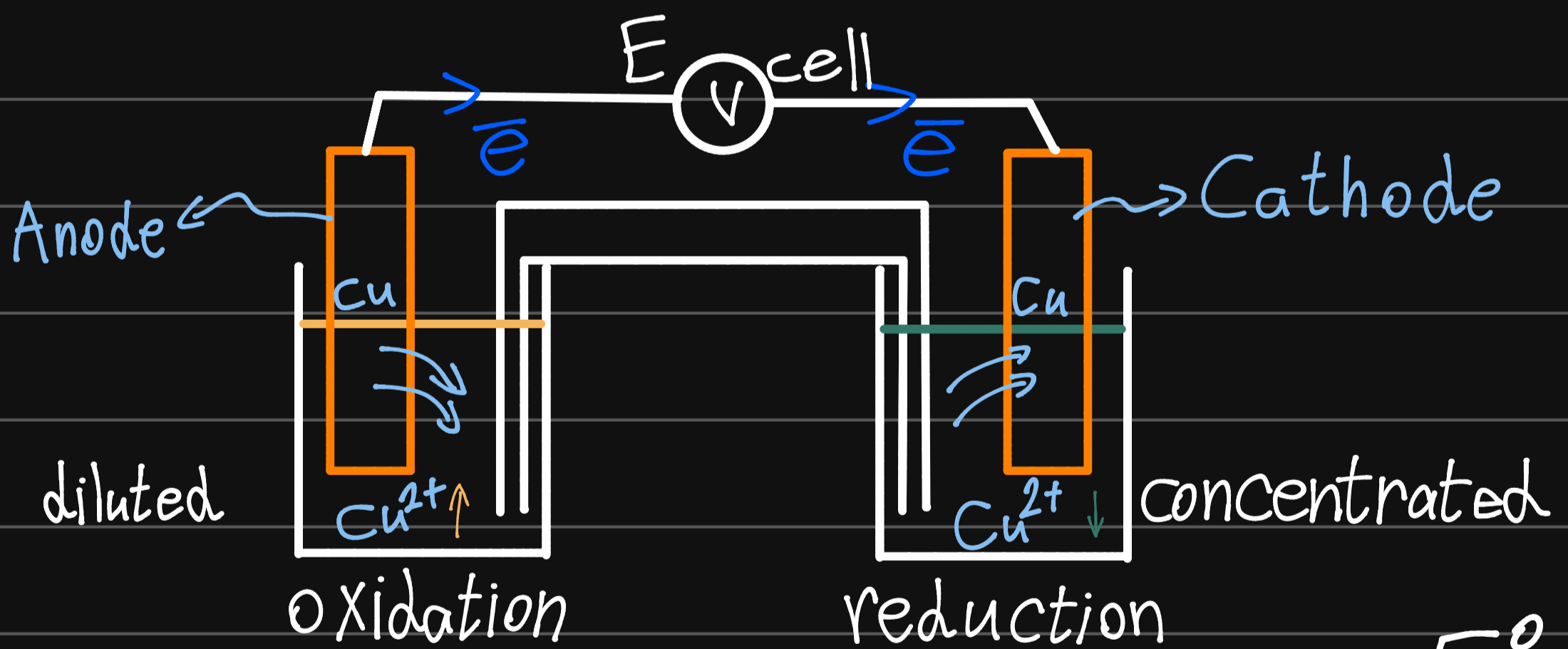
$$E = E^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Product}]}{[\text{Reactant}]}$$

عدد الإلكترونات  
الكلية

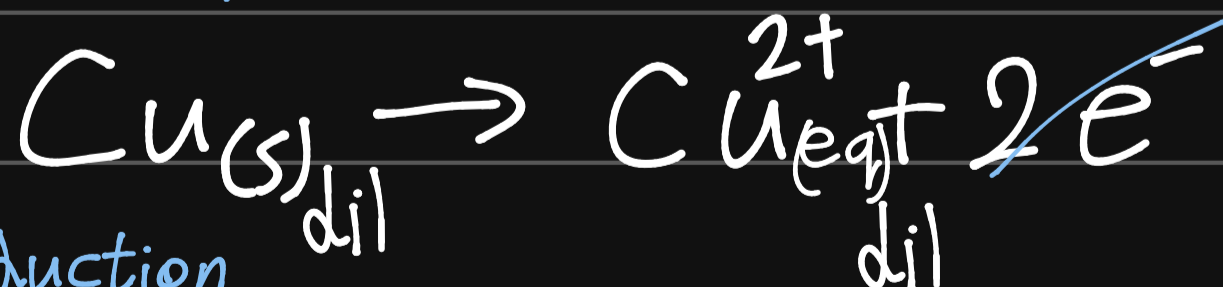
→ For the previous Zn||Cu Cell<sup>o</sup>

$$E = 1.104 - \frac{0.592}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

\* Concentration Cell<sup>o</sup>

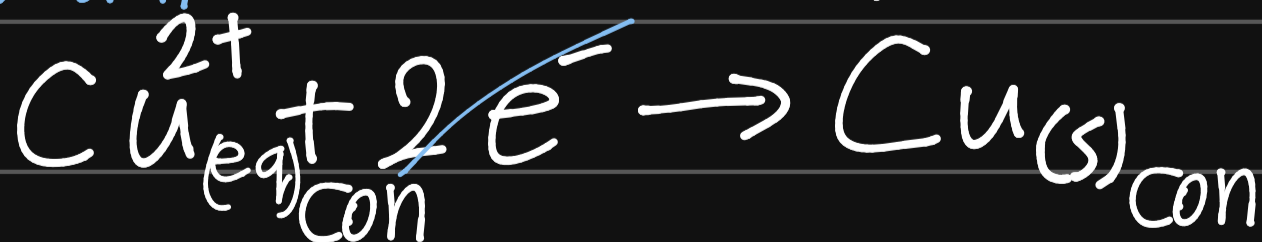


oxidation



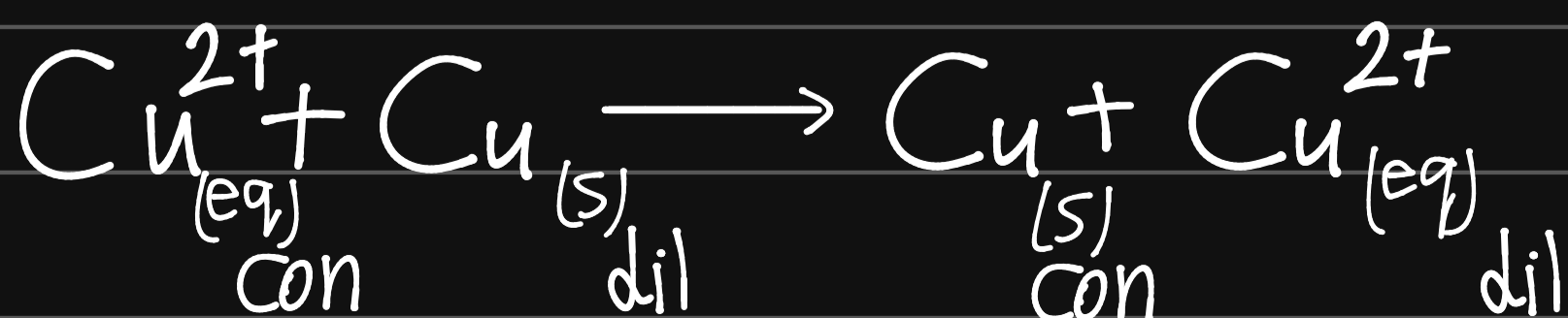
$E^{\circ}$   
-0.342 V

reduction



0.342 V

net. equ



0 V

$$E_{\text{cell}} = 0 - \frac{0.0592}{2} \log \frac{[\text{Cu}^{2+}_{\text{dil}}]}{[\text{Cu}^{2+}_{\text{con}}]}$$

at 25°C

\* فیدو م. س. م.

- In some of chemical rxns there is lose & gain of  $\bar{e}$ 's, this is called electrochemistry and can be shown in what is called Galvanic Cell.

\* Components of Galvanic cell.

① Cpds to be Oxidized & reduced

- Oxidation: loss of  $\bar{e}$ 's :  $M \rightarrow M^{n+} + n\bar{e}$

- Reduction: gain of  $\bar{e}$ 's :  $M + n\bar{e} \rightarrow M^{n-}$

- substance that is oxidized is called Reducing agent

- substance that is Reduced is called oxidizing agent

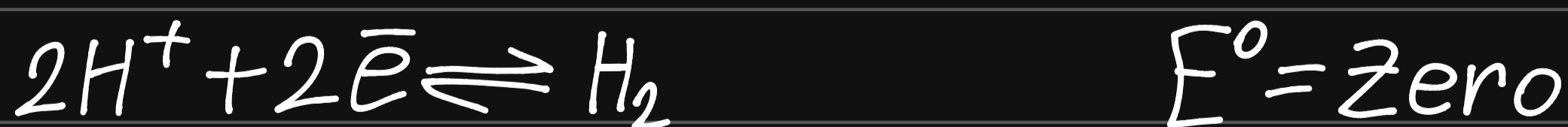
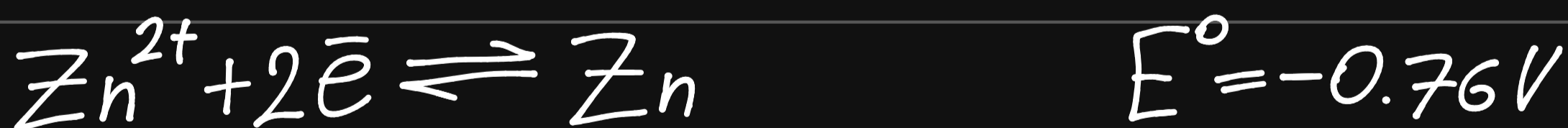
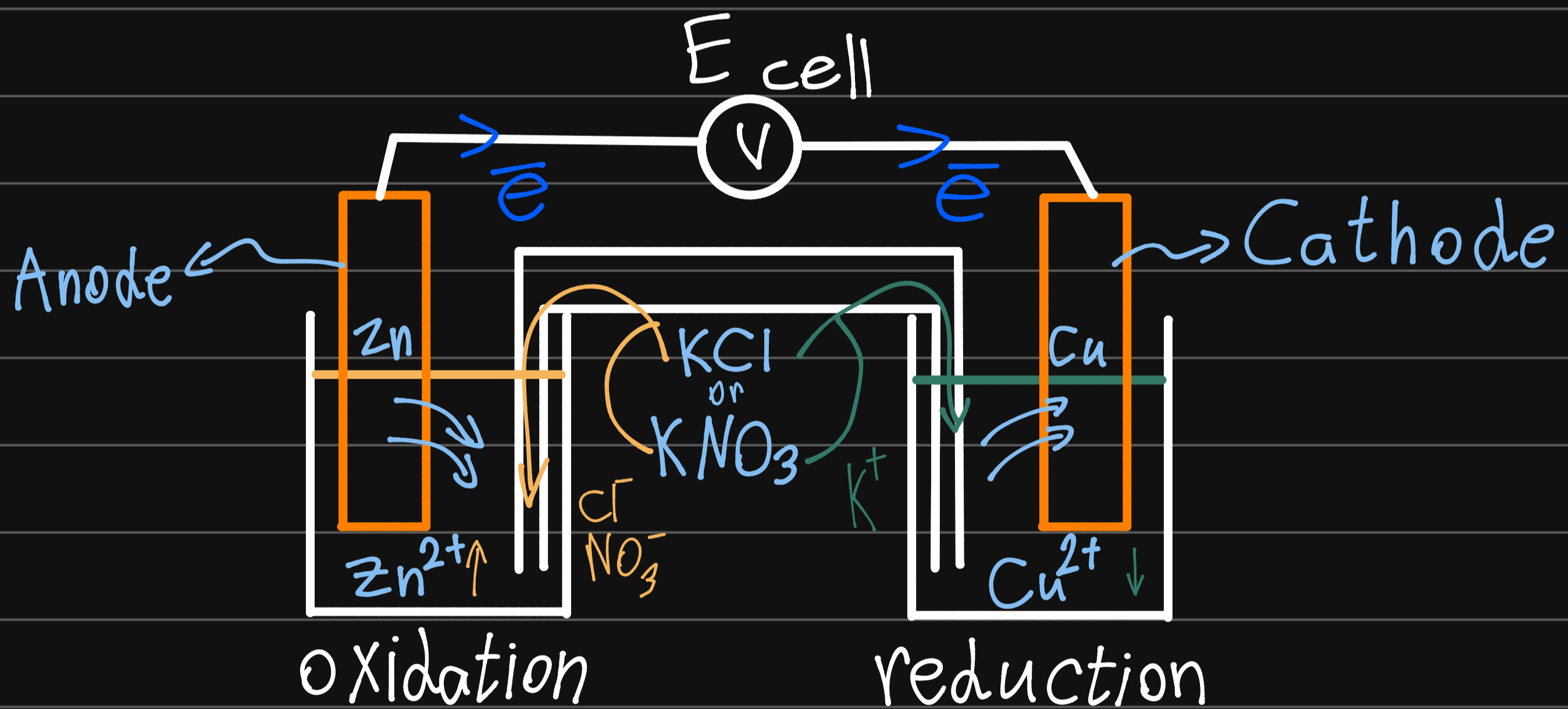
② Electrodes:-

- cathode:- Electrode of which reduction occurs.

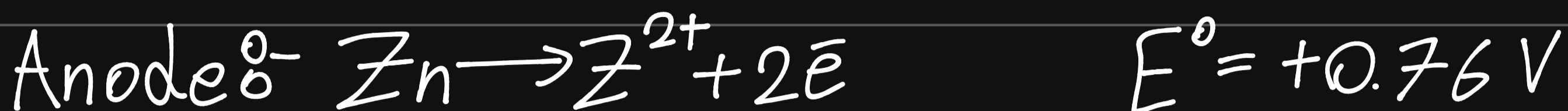
- Anode & Electrode of which Oxidation occurs.

\* In any galvanic cell  $e^-$ 's flow from anode to cathode.

③ Salt bridge: U-tube that is filled with salt saturated solu such as KCl or  $KNO_3$  used to compensate transfer of ions.

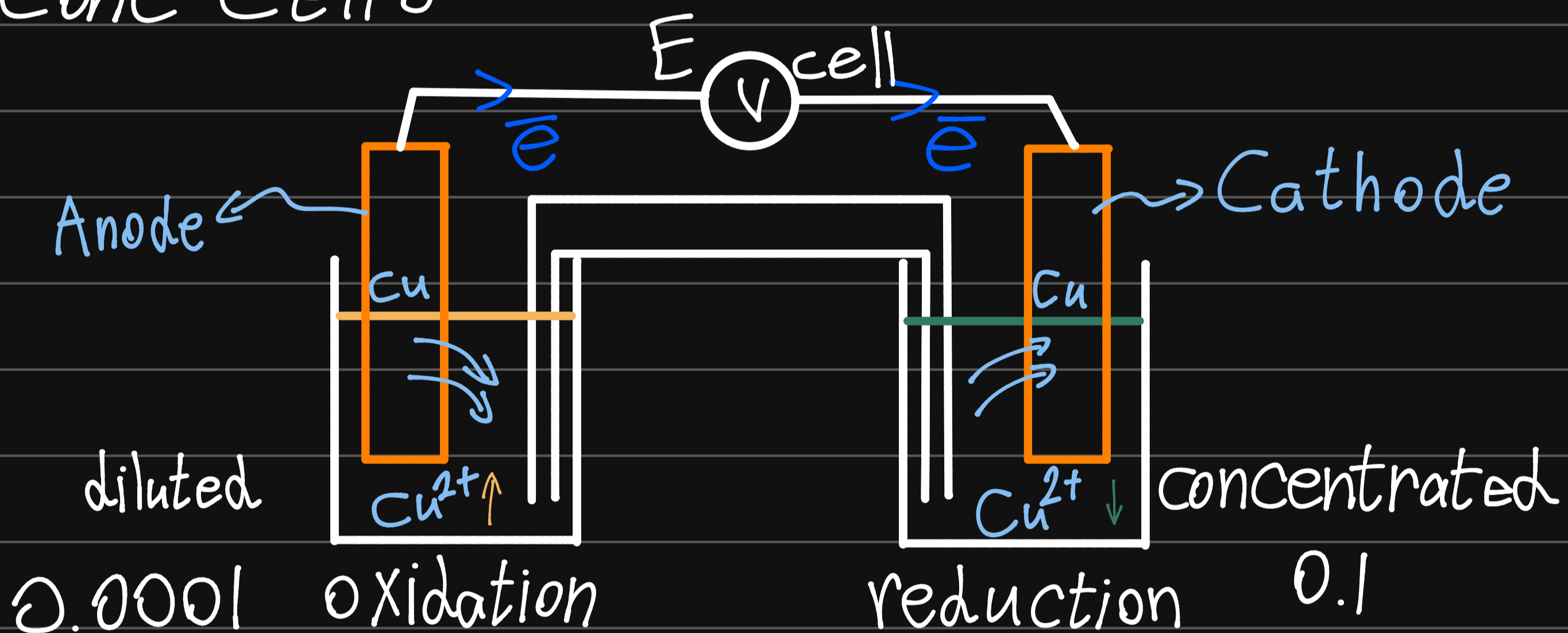


$E^\circ \rightarrow$  std conditions = 1 atm,  $25^\circ C$ .



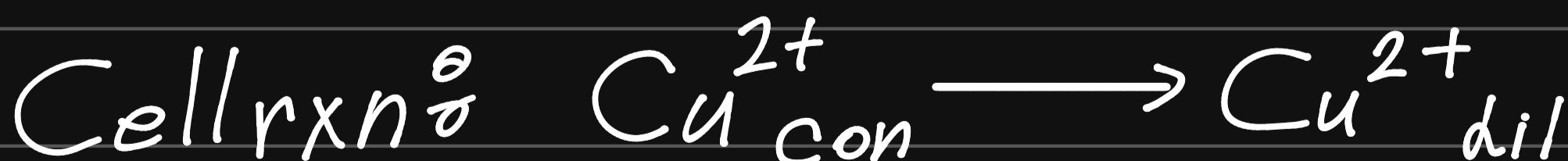
$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q \Rightarrow Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

\* Conc Cell<sup>o</sup>-



$$E^{\circ}_{\text{cell}} = \text{Zero}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q \Rightarrow Q = \frac{[\text{Cu}^{2+}]_{\text{dil}}}{[\text{Cu}^{2+}]_{\text{conk}}}$$



## Experiment (10)

### Electrochemistry

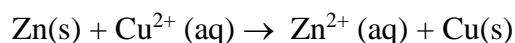
#### Objectives:

- To measure the cell potential for various electrochemical cells.
- To study the dependence of the cell potential concentration on the cell.

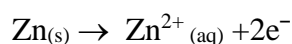
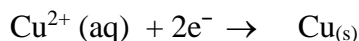
#### Introduction:

##### A. Electrochemical Cells:

An oxidation-reduction reaction (redox reaction) such as:



can be separated into two half-reactions, called half-cell reactions.



The first half-reaction is the reduction half-cell and the second half-reaction is the oxidation half-cell.

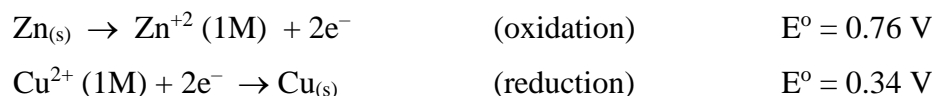
The above redox reaction can be carried out in a way such that the components of the oxidation and reduction half-cells are not mixed together, and the potential of the reaction can be measured, as shown in the Figure below.

Such a cell can be represented by:

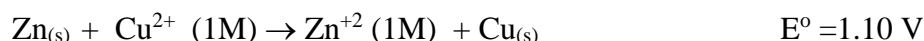


The measured potential for this cell is 1.10 V. This value is referred to as the standard cell potential,  $E^{\circ}$  which indicates the voltage observed when all species are in their standard states (1.0 M for species in solution, and 1.0 atm. pressure for gases). Most commonly  $E^{\circ}$  values are measured at 298 K.

This cell potential,  $E^\circ$ , can be calculated from the two half-cell potentials:



Net reaction:



When a cell reaction is written so that  $E^\circ$  is positive, the reaction will proceed spontaneously from left to right when all components are in their standard states.

## **B. Concentration Cells**

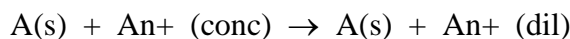
In this experiment you will observe the effect of varying the concentration of a species, participating in an electrochemical oxidation-reduction reaction, on the cell potential. Since the potential of an electrode depends on the concentration of the ions used in the electrode half-cell, a cell may be constructed from two half-cells differing only in the concentration of ions. The principle behind this is simple, dilution of a concentrated solution is spontaneous, so this process gives a cell with positive  $E$ . For the process:



a cell can be constructed such as:



which represents the reaction:



Since the two half-reactions are the same,  $E_o(\text{anode}) = E_o(\text{cath})$  so that  $E_o (\text{cell}) = 0$ .

Applying Nernst equation, at  $25^\circ\text{C}$ , we get:

$$E = -\frac{0.0592}{n} \log \frac{[\text{A}^{n+}]_{\text{dil}}}{[\text{A}^{n+}]_{\text{conc}}}$$

where  $n$  is the number of electrons transferred.

## Experiment (10) Galvanic Cell

**Galvanic cell:** Electrochemical cell in which spontaneous oxidation-reduction reactions produce electric current.

Components of Galvanic cell:

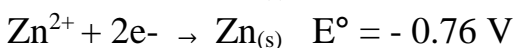
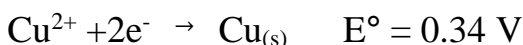
### 1. Substances that be oxidized and reduced.

The **reduction** occurs on the half cell that has **higher +ve value of  $E^\circ$**

Then the oxidation occurs on the other half cell.

Take the example as shown in the figure below:

From  $E^\circ$  values are listed in the table,  $E^\circ$  of  $\text{Cu}^{2+}$  (0.34 V) is more positive than  $\text{Zn}^{2+}$  (-0.76 V). So,  $\text{Cu}^{2+}$  will be reduced and Zn will be oxidized.



Then the reactions are written as:



$E^\circ$ : voltage (potential) of the half cell at standard condition

\*standard conditions: 25 °/298 K, 1 atm, 1 M solution

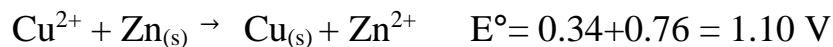
reduction reaction	$E^\circ$ (V)
$\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$	-1.662
$\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$	-0.762
$\text{Cr}^{3+} + 3e^- \rightarrow \text{Cr}$	-0.744
$\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}$	-0.447
$\text{Cd}^{2+} + 2e^- \rightarrow \text{Cd}$	-0.403
$\text{Ni}^{2+} + 2e^- \rightarrow \text{Ni}$	-0.257
$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}$	-0.126
$2\text{H}^+ + 2e^- \rightarrow \text{H}_2$	0.000
$\text{AgCl} + e^- \rightarrow \text{Ag} + \text{Cl}^-$	+0.222
$\text{Hg}_2\text{Cl}_2 + 2e^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	+0.268
$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$	+0.342
$\text{Cu}^+ + e^- \rightarrow \text{Cu}$	+0.521
$\text{Ag}^+ + e^- \rightarrow \text{Ag}$	+0.780
$\text{Au}^{3+} + 3e^- \rightarrow \text{Au}$	+1.498
$\text{Au}^+ + e^- \rightarrow \text{Au}$	+1.692

Increasing tendency to get reduced. More



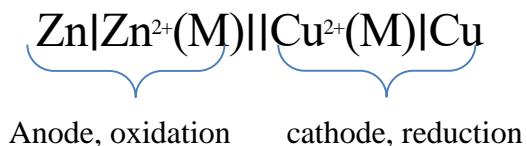
The net

equation is:



\*  $E^\circ$  is always positive

This galvanic cell can be represented by line cell:



Oxidation: loss of electrons

Reduction: gain of electrons

Reducing agent: substance that is oxidized.

oxidizing agent: substance that is reduced.

So,

Reducing agent: Zn

Oxidizing agent:  $\text{Cu}^{2+}$

## 2. Electrodes:

Cathode : electrode at which reduction occurs

Anode: electrode at which oxidation occurs

Electrons flow from anode to cathode

So,

Zn is anode

Cu is cathode

And electrons flows from Zn to Cu

## 3. Salt bridge:

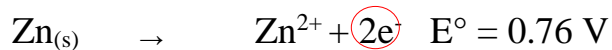
U-shaped glass tube filled with saturated electrolyte solution (e.g., KCl,  $\text{KNO}_3$ ) to compensate the transfer of ions (keep charges in each half cell balanced).

- $\text{Cl}^-$  or  $\text{NO}_3^-$  ions move toward the anode/oxidation side (to neutralize excess  $\text{Zn}^{2+}$ )
- $\text{K}^+$  ions move toward the cathode/ reduction side (to replace the  $\text{Cu}^{2+}$  ions that are reduced)

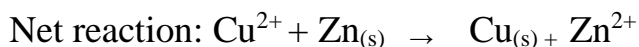
## Calculation of the cell potential (Nernst equation):

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q, \quad Q = \frac{[\text{product}]^x}{[\text{reactant}]^y}$$

X and y are stoichiometric coefficients (the numbers in front of each species) from the balanced redox reaction.



Note: the moles of electrons in the two half reactions must be equal



$$E^{\circ}_{\text{cell}} = 0.34 + 0.76 = 1.10 \text{ V}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log \frac{[\text{product}]^x}{[\text{reactant}]^y} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log \frac{[\text{oxidation}]^x}{[\text{reduction}]^y}$$

$$= 1.10 - \frac{0.0592}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}, \quad n = 2, \quad x \text{ and } y = 1$$

n = moles of electrons in balanced redox equation

for example: if the concentration of  $\text{ZnSO}_4$  is 0.02 M and concentration of  $\text{CuSO}_4$  is 0.4 M, calculate the cell potential

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log \frac{[\text{product}]}{[\text{reactant}]} = E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

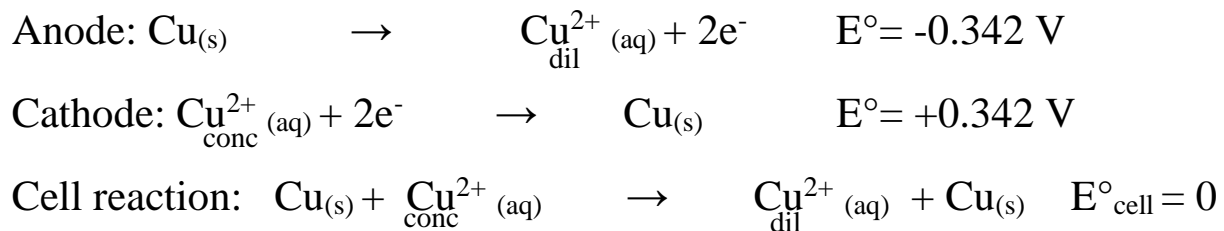
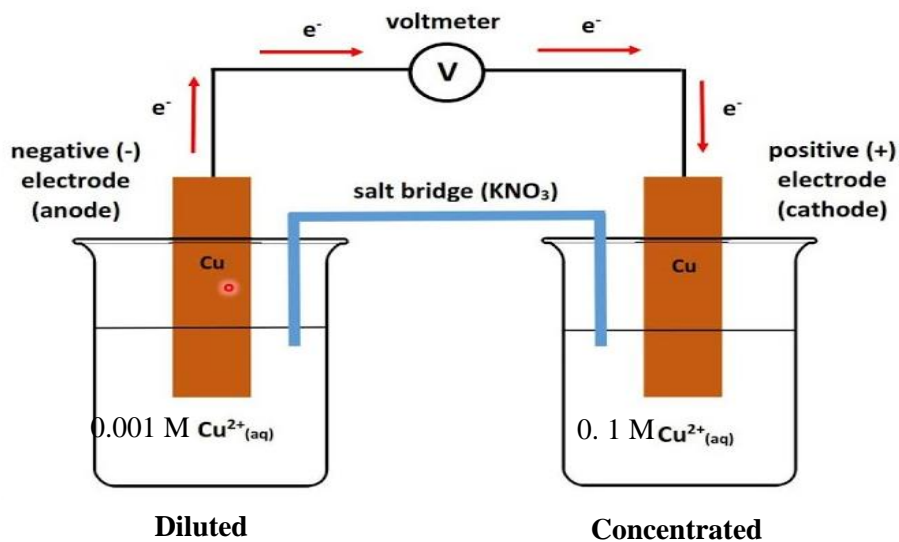
$$= 1.10 - 0.0296 \log \frac{0.02}{0.4}$$

$$= 1.14 \text{ V}$$

## Concentration Cell

**Concentration cell:** is an electrolytic cell that consists of two half-cells with the same electrodes, but different concentrations.

Electrons flow from the cell with the lower concentration (diluted) (Anode) to the cell with the higher concentration (concentrated) (cathode).



$$E_{\text{cell}} = \cancel{E^\circ_{\text{cell}}} - \frac{0.0592}{n} \log Q, \quad E_{\text{cell}} = -\frac{0.0592}{2} \log \frac{[\text{Cu}^{2+}]_{\text{dil}}}{[\text{Cu}^{2+}]_{\text{conc}}}$$

\*Always  $E^\circ_{\text{cell}}$  of concentration cell is 0

## B. Concentration cell (1) page 65

Volume of one drop = 0.05 mL

3 drops of  $\text{Cu}(\text{NO}_3)_2$ :  $3 \times 0.05 = 0.15$  mL (volume of concentrated solution)

Volume of solution = 50 mL

Volume of diluted solution =  $50 + 0.15 = 50.15$  mL

Concentration of  $\text{Cu}(\text{NO}_3)_2 = 0.1$  M

$$(\text{MV})_{\text{dil.}} = (\text{MV})_{\text{conc.}}$$

$$M_{\text{dil.}} \times 50.15 = 0.1 \times 0.15$$

$$M_{\text{dil.}} = (0.1 \times 0.15) / 50.15 =$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log \frac{[\text{Cu}^{2+}]_{\text{dil}}}{[\text{Cu}^{2+}]_{\text{conc}}}$$

## B. Concentration Cells

Set up a simple concentration cell by using a Cu/0.10 M Cu(NO<sub>3</sub>)<sub>2</sub> half-cell as that prepared in section A, and a second half-cell prepared as follows:

1. Take 30 mL (the same volume as before) of 0.1M KNO<sub>3</sub> solution and add to it exactly one drop of 0.10M Cu(NO<sub>3</sub>)<sub>2</sub> solution.
2. Clean a Cu strip with a sand paper and insert it into the beaker.
3. Connect the two half-cells by a salt bridge and measure the cell voltage.
4. ~~Now add 10 drops of 0.10M Na<sub>2</sub>S solution to the second, more dilute, half-cell. Since CuS is very insoluble the addition of Na<sub>2</sub>S has the effect of decreasing Cu<sup>2+</sup> concentration to an extremely small value.~~
5. Measure the cell voltage again.
6. ~~Label the electrode signs and processes (draw a diagram).~~
7. ~~Measure the temperature of the solution.~~

## B. Concentration Cells

### Calculations:

1. One drop has a volume of about 0.05 mL, thus the concentration of Cu<sup>2+</sup> in the dilute half-cell is about  $(0.05/30) \times 0.1M = 2 \times 10^{-4}M$ . Use the Nernst equation to calculate the value of E at 298 K and compare it with your observed value. Why might they differ?

E<sub>calculated</sub> (show calculations) =

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