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مانيوال :
اساسيات كيمياء عملية
Chemistry Lab

اللجنة الأكاديمية لقسم الهندسة الصناعية

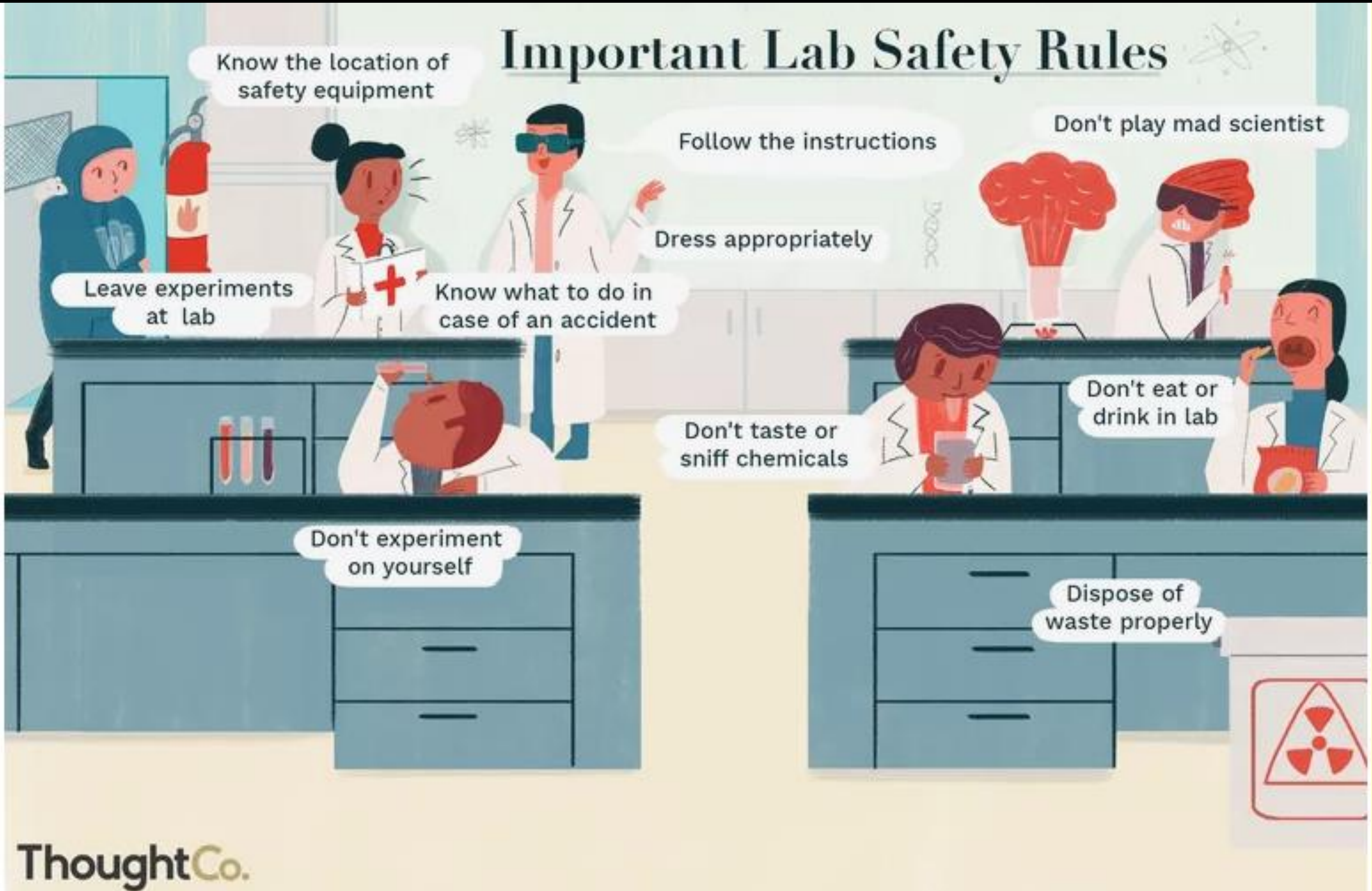
2023

1. SAFETY RULES

Are the basic rules that provide behavior, hygiene, and safety information to avoid accidents in the laboratory

Why? **to protect yourself and others from injuries**

1. SAFETY RULES



1. SAFETY RULES

10 most important lab safety rules

1. Follow the instructions
2. Know the Location of Safety Equipment
3. Dress for the Lab
4. Don't Eat or Drink in the Laboratory
5. Don't Taste or Sniff Chemicals
6. Don't Play Mad Scientist in the Laboratory
7. Dispose of Lab Waste Properly
8. Know What to Do With Lab Accidents
9. Don’t Leave your Experiment unattended
10. Don't Experiment on Yourself or other

Think before you act

When in doubt. **ASK**

1. SAFETY RULES

- **In our lab:**
- Lab coat and eye goggles must be worn during lab sessions.

1. SAFETY RULES

- **In our lab:**
- Food, drinks, smoking, and gum are not allowed in the labs at any time.

1. SAFETY RULES

- **In our lab:**

- Treat every chemical as if it is hazardous!

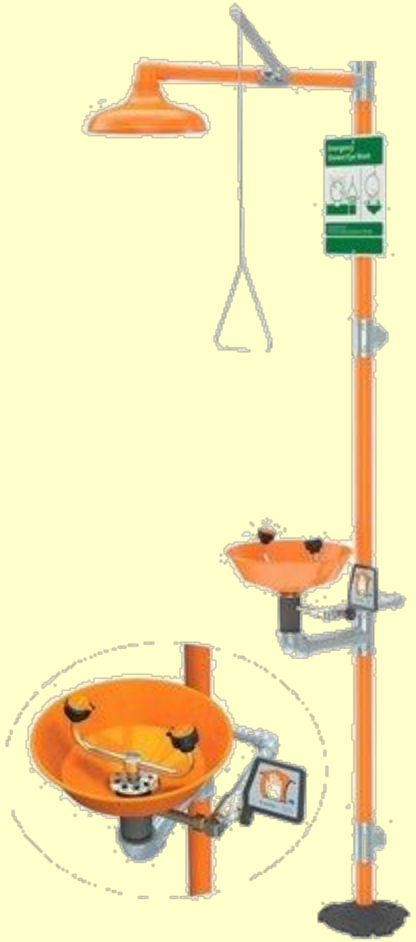
1. SAFETY RULES

- **In our lab:**
- Know the locations of:
  - *Fire alarms,*
  - *Fire extinguishers,*
  - *Chemical fume hoods,*

1. SAFETY RULES

- **In our lab:**

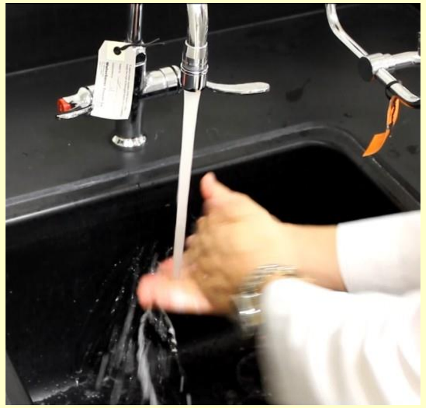
- Know the locations of:
- *Safety showers,*
- *Emergency eye washes.*



1. SAFETY RULES

- **In our lab:**

- ❑ Notify your instructor immediately of any injury, spill, fire, or explosion:
- ❑ Whenever your skin (hands, face, ...) comes into contact with chemicals, wash it quickly and thoroughly with warm water and soap.



- ❑ Notify your lab instructor if there is a mercury spill due to a broken mercury thermometer.

1. SAFETY RULES

- **In our lab:**

- Do not taste or smell anything.

- Never remove shared chemicals from their original locations.

- Label all containers to avoid errors. Read labels carefully.

1. SAFETY RULES

- **In our lab:**

- Dispose of all waste chemicals in an appropriate manner:

- *Never put solids in the sink.*

Experiment 2
The Empirical Formula of an Oxide

Empirical Formula of an Oxide

- **Objectives:**

Determine the empirical formula for magnesium oxide experimentally

What is the ratio of Mg to O in the magnesium oxide

- **Method:**

Convert a sample of magnesium to magnesium oxide and measure mass changes

Empirical Formula of an Oxide

Introduction:

- Magnesium reacts violently with oxygen at high temperature, producing magnesium oxide

$$2\ Mg + O_2 \rightarrow 2\ MgO$$

Empirical formula of magnesium oxide

Empirical Formula of an Oxide

Introduction:

Empirical formula: is the <u>**simplest**</u> <u>**integer**</u> **ratio** of atoms present in a compound

Empirical formula of magnesium oxide can be determined theoretically from the oxidation numbers of magnesium and oxygen

	Mg	O
Oxidation number =	**+2**	**-2**

They are neutral, therefor the Mg to O ratio must be 1:1; $\mathbf{Mg_1O_1}$ (reported as MgO) ← **Theoretical empirical formula**

Empirical Formula of an Oxide

Introduction:

Empirical formula of magnesium oxide can be determined **experimentally** by determining how many moles of oxygen react with certain number of moles of magnesium

Main steps:

- **Start with known amount of Mg**
- **React Mg with O_2 by combusting Mg in air**
- **Obtain the mass of the oxide**

Empirical Formula of an Oxide

Introduction:

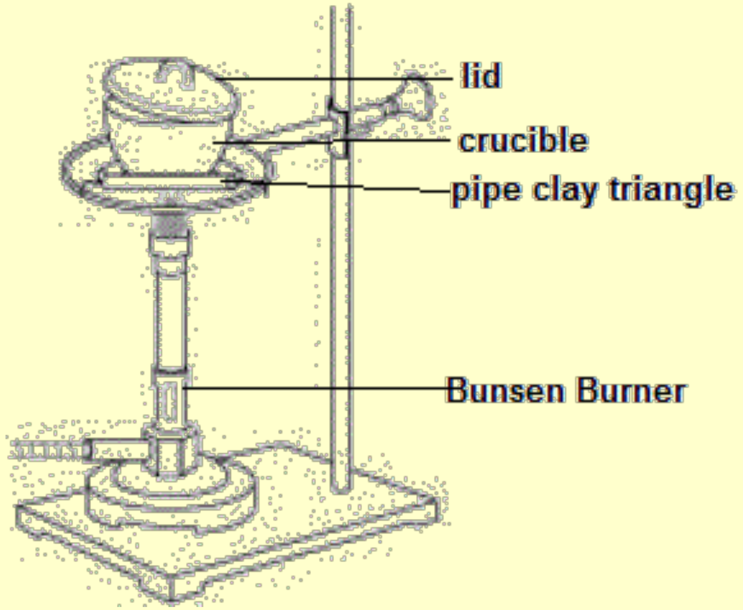
Air contains O_2 and N_2
so
The reaction products of combusting Mg in air:

- Magnesium oxide
- Magnesium nitride (Mg_3N_2)

Empirical Formula of an Oxide

Experimental:

- Same as last experiment we start with clean empty crucible (mass = m_1)
- Place 0.20 g of Mg inside the crucible and obtain their mass (mass = m_2)
- Partially cover the crucible (why?) then heat on soot-less fire until all content become white with now glow



The crucible content now are magnesium oxide and nitride

Empirical Formula of an Oxide

Experimental:

The next steps are to convert the nitride to oxide:

- **Cool the crucible and then add 10 drops of distilled water on the crucible contents. The reaction:**

$$Mg_3N_2 + 6\,H_2O \rightarrow 3\,Mg(OH)_2 + 2\,NH_3$$

(produced ammonia will convert litmus paper to blue)

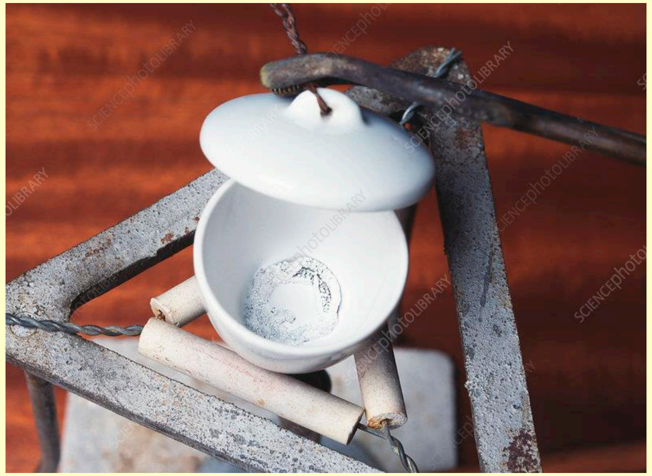
- **Heat the crucible on the flame for 5 minutes. The reaction:**

$$Mg(OH)_2 + heat \rightarrow MgO + H_2O$$

Now all Mg we started with is converted to oxide

After this step, obtain the mass of the crucible and its content (mass = m_3)

Empirical Formula of an Oxide



Calculations:

- Mass of Magnesium = $m_2 - m_1$
- Mass of Oxygen = $m_3 - m_2$
- Moles of Mg = $\frac{mass\ of\ Mg}{24.3}$
- Moles of O = $\frac{mass\ of\ O}{16.0}$
- Formula : $Mg_{(moles\ of\ Mg)}O_{(moles\ of\ O)}$
- Empirical formula: divide by the least number of moles

Asample of a report sheet

Mass of empty crucible (m_1)	**54.78 ± 0.01**	g
Mass of crucible and Mg (m_2)	**54.97 ± 0.01**	g
Mass of Mg	**54.97 – 54.78 = 0.19**	g
Moles of Mg (n_1)	**0.19 ÷ 24.3 = 0.0078**	mol
Final mass of crucible and Mg-oxide (m_3)	**55.11 ± 0.01**	g
Mass of Mg-oxide produced	**55.11 – 54.78 = 0.33**	g
Mass of oxygen gained	**0.33 – 0.19 = 0.14**	g
Moles of oxygen atoms (n_2)	**0.14 ÷ 16.0 = 0.0088**	mol
Formula of magnesium oxide ($Mg_{n1}O_{n2}$)	$\mathbf{Mg_{0.0078}O_{0.0088}}$	
Empirical formula of magnesium oxide $\frac{0.0088}{0.0078}$	$\mathbf{Mg_1O_{1.1}}$ → **MgO**	
Mass percent of Mg in the oxide (x_1) (experimentally) (from you data)	**0.19 ÷ (0.019+0.14)×100% = 58%**	
Mass percent of Mg in the oxide (x_2) (calculated for MgO) (from the molar masses)	**24.3 ÷ (24.3+16.0)×100% = 60%**	

Limiting Reactant

Limiting Reactant

Objectives:

- Determine the limiting reactant in a mixture of $BaCl_2$ and Na_3PO_4
- Calculate the original masses of reactants and their mass percentages

Limiting Reactant

Introduction:

For reactions with two (or more reactants) the reaction will proceed until one reactant is totally consumed

This reactant is called the limiting reactant

Amount of this reactant determines:

- Amount of product(s) produced
- Amount of the excess reactant consumed

Limiting Reactant

Introduction:

In this experiment, you will carry out an aqueous reaction between sodium phosphate and barium chloride:

$$2\ Na_3PO_{4(aq)} + 3\ BaCl_{2(aq)} \rightarrow Ba_3(PO_4)_{2(s)} + 6\ NaCl_{(aq)}$$

the net ionic equation for the reaction will be:

$$2\ PO_4{}^{3-}{}_{(aq)} + 3\ Ba^{2+}{}_{(aq)} \rightarrow Ba_3(PO_4)_{2(s)}$$

You have to know the difference between actual yield and theoretical yield

Actual yield is: the amount of a product we obtain from the experiment,

(experimental yield)

Theoretical yield is: the value calculated from the equation of the chemical reaction while considering stoichiometry (true yield)

Note : actual yield may be lower or higher, depending on your work and considering the errors may be encountered during the work

Limiting Reactant

Experimental:

You will be experimenting on a salt mixture that contains:

$Na_3PO_4.12H_2O$ molar mass=380.2 g/mol
$BaCl_2.2H_2O$ molar mass=244.2 g/mol

You have to find:
a) Which is the limiting reactant
b) Mass for each reactant
c) Mass % for each reactant

Limiting Reactant

Experimental:

Dissolve accurately weighed ~0.70 g of the mixture.
The reaction is instant and precipitate will appear

In order to separate the solid from the liquid:

- Coagulate the precipitate
- Filtrate using filter paper

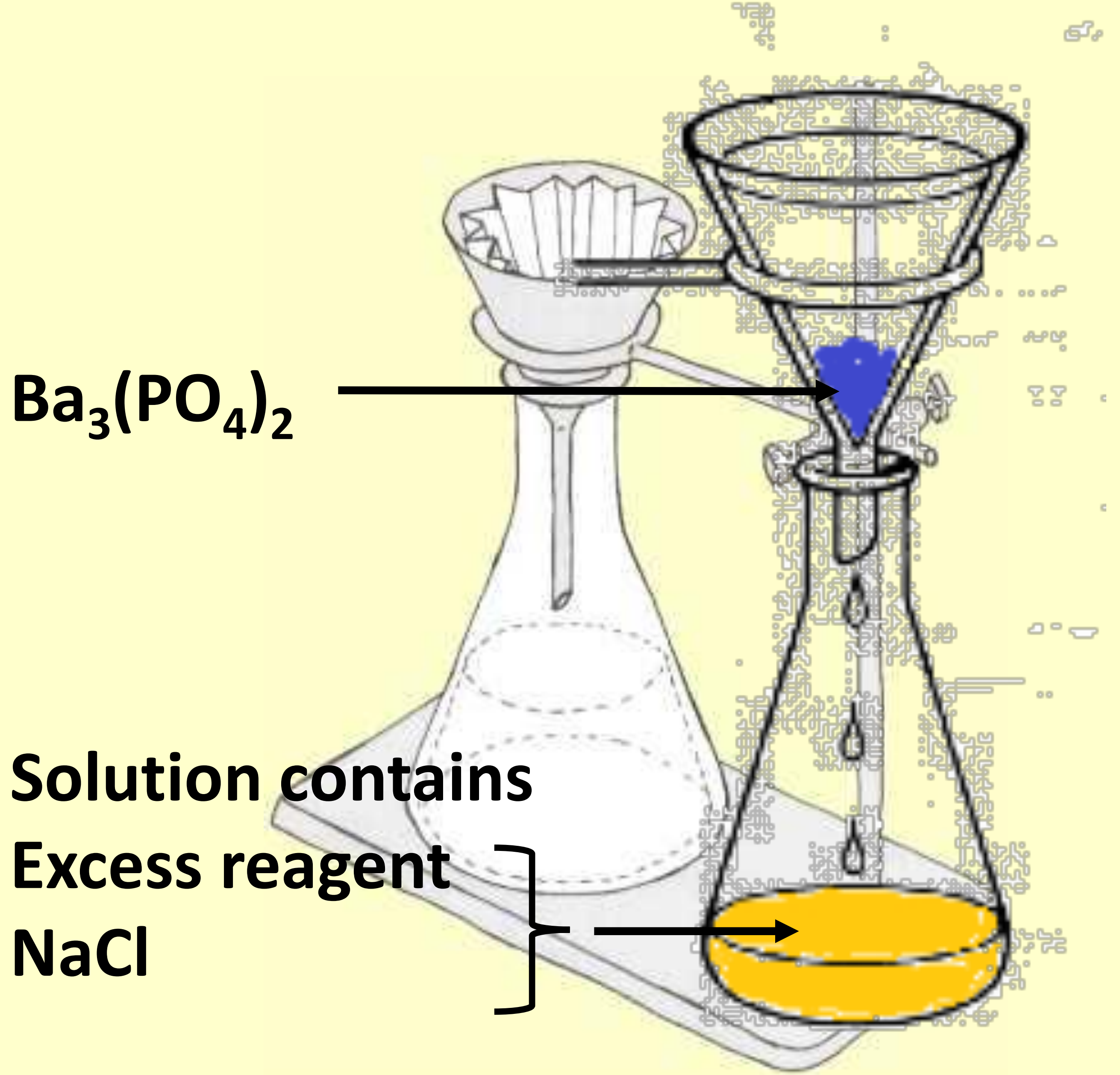
Heating the solution mix to 80-90°C for 20 minutes

Limiting Reactant

Experimental:

Filtration:

- Assemble the filtration setup
- Pour the solution mix into the filter paper
- Collect all solid particles
- Wash the solid with 5 mL of hot water

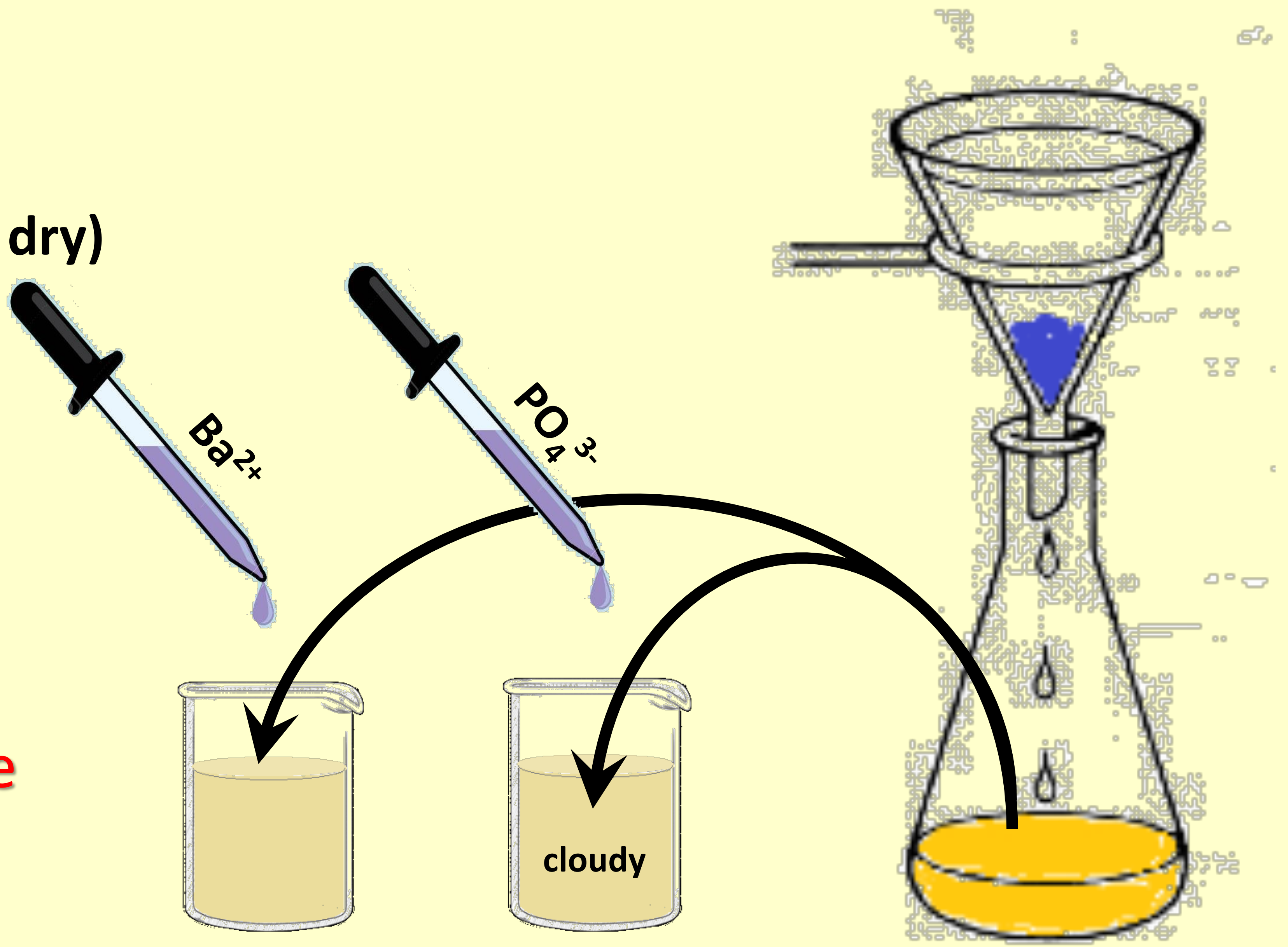


Limiting Reactant

Experimental:

Place the solid in an oven (to dry) and then obtain its mass

Transfer some of the filtrate into two small beakers



In this case Na_3PO_4 is the **limiting reactant**

Limiting Reactant

Calculations:

Results you collected:

Mass of salt mixture

Mass of $Ba_3(PO_4)_2$

The limiting reactant : Na_3PO_4

Limiting Reactant

If the limiting reactant is $Na_3PO_4.12H_2O$:		
Mass of $Ba_3(PO_4)_2$ precipitated	0.21±0.01	g
Number of moles of $Ba_3(PO_4)_2$ precipitated	0.21/601.9=0.00035	mol
Number of mole of $Na_3PO_4.12\ H_2O$ reacted	0.00035×2=0.00070	mol
Number of moles of $BaCl_2.2H_2O$ reacted	0.00035×3=0.0010	mol
Mass of $Na_3PO_4.12H_2O$ reacted	0.00070×380.2=0.27	g
Mass of $BaCl_2.2H_2O$ reacted	0.0010×244.2=0.24	g
Mass of excess $BaCl_2.2H_2O$	0.70-(0.27+0.24)=0.20	g
Mass percentage of $Na_3PO_4.12H_2O$	(0.27/0.70)×100%=39	%

Experiment (4)

Identification of a Compound: Physical Properties

What do we mean by the physical property?

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

For example:

- Silver is a shiny metal that conducts electricity very well.
- Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily.

** Physical properties of matter include :

-Color

-Hardness

-Solubility

-electrical conductivity

-density

-melting point

-boiling point

Aim of experiment:

- Identification of a chemical compound using it’s physical properties

1- Solubility

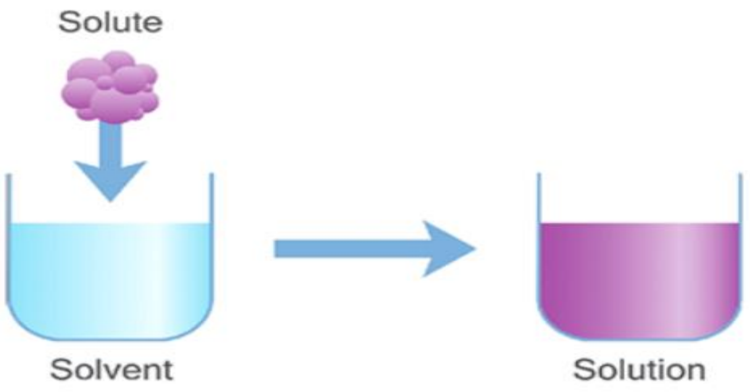
2- Density

3- Boiling point

1. Solubility

- Solvent + Solute = Solution

- The solubility of a substance in a particular solvent at a particular temperature is the maximum quantity of the substance dissolved in a fixed quantity of the solvent to form a saturated solution at that temperature. It is determined by the nature of the solute, solvent, temperature and pressure



Solubility of Liquids in Liquid

- liquid-liquid systems are classified according to the solubility to:

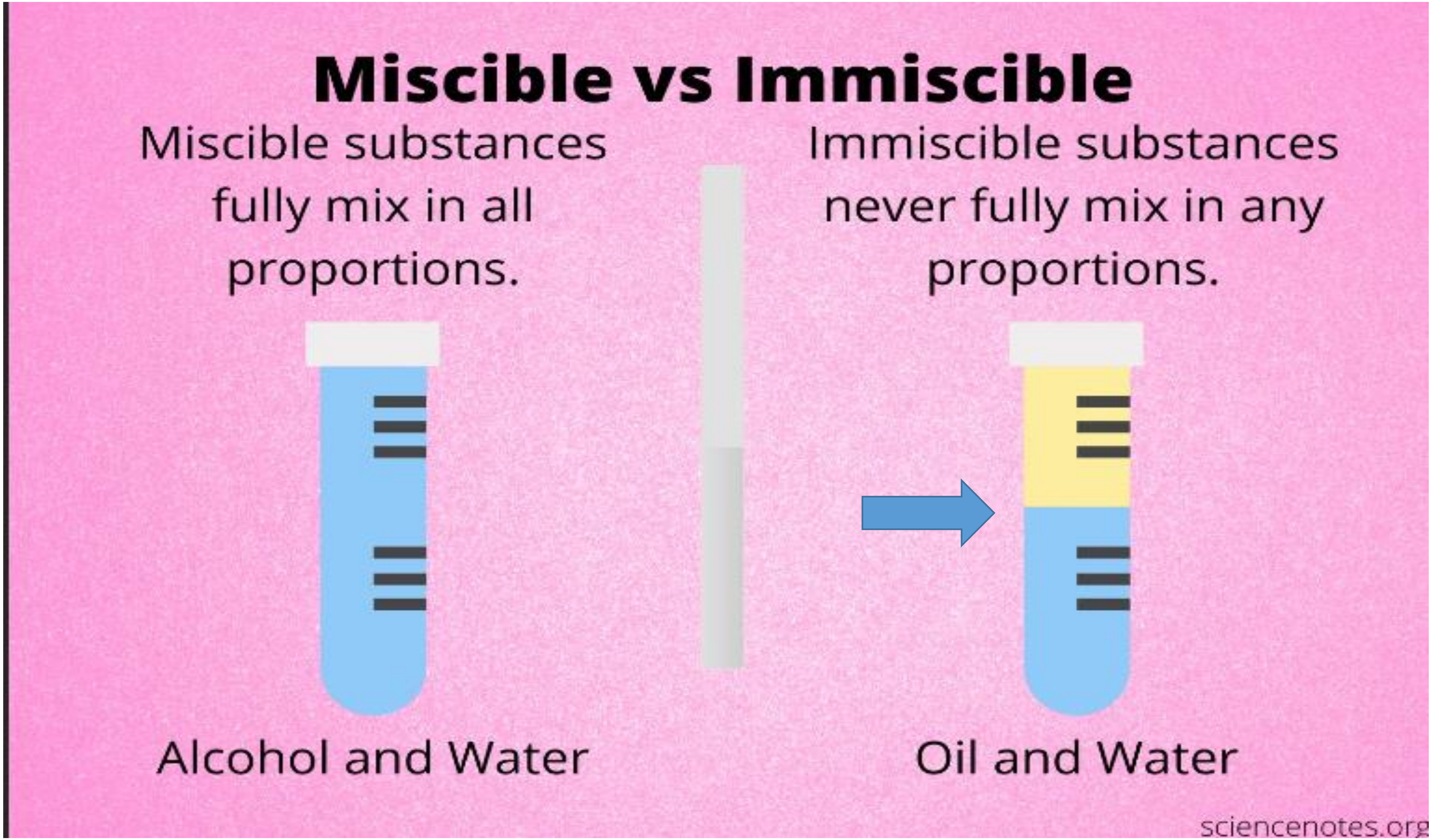
1- Completely miscible:
polar with polar such as water-alcohol

Non-polar with non-polar such as CCl_4 and Benzene

completely miscible = one layer observed = clear solution

2- Immiscible:
These liquids do not mix in any proportion, such as water and benzene

Immiscible = not mixed = two layers = turbidity = cloudiness


Miscible vs Immiscible
Miscible substances fully mix in all proportions.
Immiscible substances never fully mix in any proportions.
Alcohol and Water
Oil and Water
sciencenotes.org

2. Density

** **Density** is mass of a unit volume of a material substance.

d = m/v

d is density
m is mass
v is volume

** **Density** is commonly expressed in units of grams per cubic centimetre

$g/cm^3 = g/mL$

Different liquids ➔ Different densities
For example

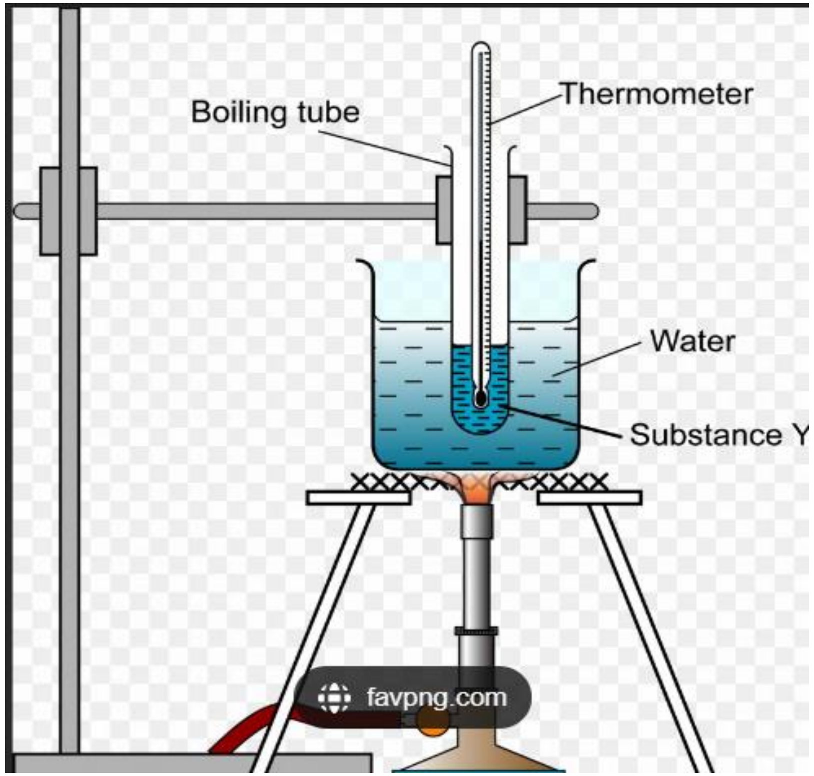
Density of Various Liquids

Liquid	Density (g/cm^3)	Temperature (°C)
acetone	0.792	20
alcohol (ethyl)	0.791	20
alcohol (methyl)	0.810	0
gasoline	0.66-0.69	
corn syrup	1.38	20
castor oil	0.969	15
olive oil	0.918	15
linseed oil	0.942	15
vegetable oil	0.91-0.93	20
turpentine	0.87	
water	1.00	4

3- Boiling point

- The **boiling point** of a substance is the temperature at which the vapor pressure of a liquid equals the external pressure surrounding the liquid

(at this point the liquid changes into a vapor)



What will you do in the experimental part during measuring the boiling point

NOTE

Boiling point is directly proportional to external pressure

Boiling point is inversely proportional to vapor pressure

- At the beginning of heating, vapor pressure of the liquid < Atmospheric pressure

(**temperature will not be recorded, since it will be lower than the boiling point**)

- As the temperature increases, the vapor pressure of the liquid increases until

Vapor pressure of the liquid = Atmospheric pressure

(temperature recorded = boiling point)

(We cannot detect it)

- The temperature will complete increasing more and more

vapor pressure of the liquid > Atmospheric pressure

Liquid starts escaping from the tube

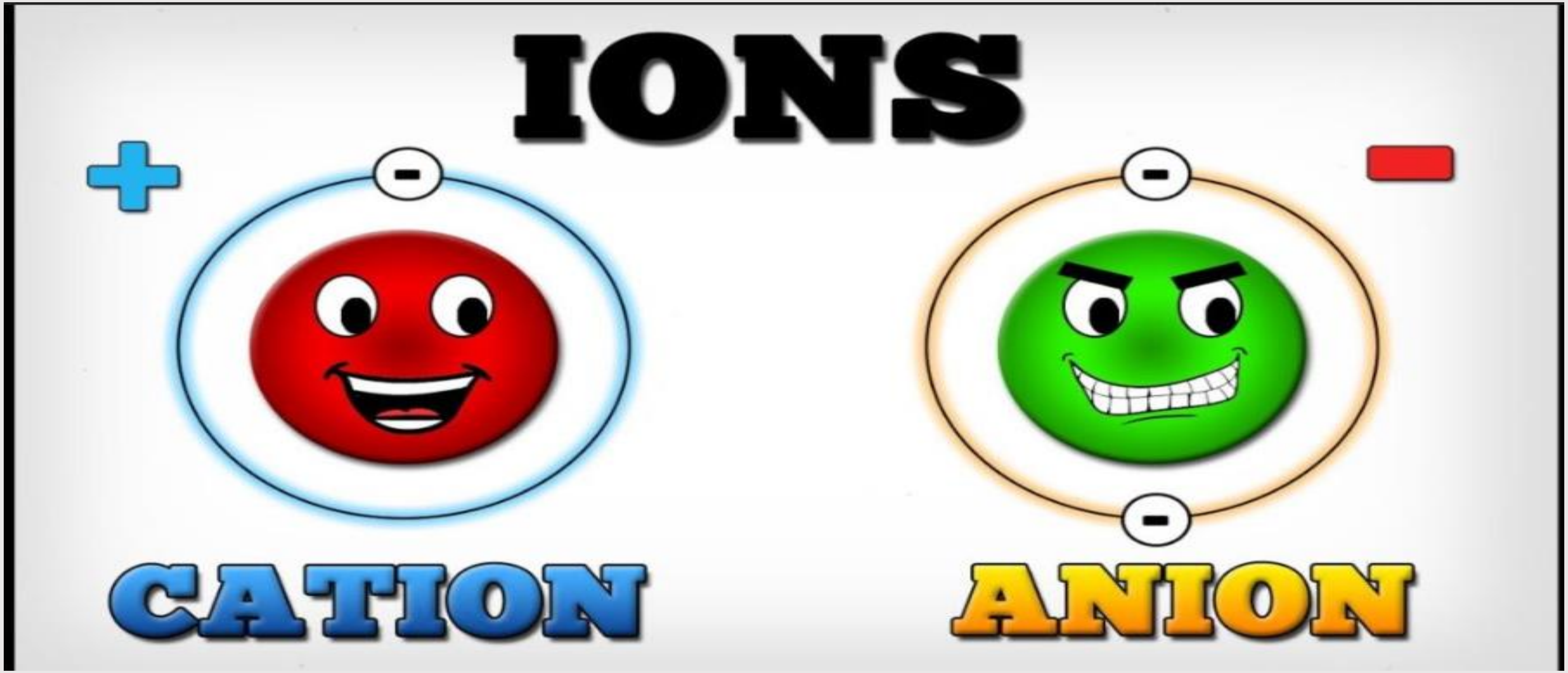
(temperature recorded > boiling point)

(We will not record it)

- Turn off the burner, the temperature decreases and thus the vapor pressure decreases

(As the liquid starts entering the tube, record the temperature and this will be the nearest value to the true boiling point)

Cations and Anions



What are cations and anions?

- Cations and anions are both ions.
  The difference between a cation and an anion is the net electrical charge of the ion.

- Ions are atoms or molecules which have gained or lost one or more valence electrons, giving the ion a net positive or negative charge.

- Positive charge ------- Cation
- Negative charge ------ Anion

Aim of experiment

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

One of the most common and important tasks for a chemist is to identify unknown compounds. In this experiment you will use simple chemical tests to identify an unknown salt containing one cation and one anion. We have limited the possible ions in your unknown to the following:

Cations	Anions
potassium, K^+	Sulfate, SO_4^{2-}
calcium, Ca^{2+}	Hydrogen carbonate, HCO_3^-
iron(III), Fe^{3+}	Chloride, Cl^-
ammonium, NH_4^+	Bromide, Br^-

How can you discover the presence of an ion in the salt?

Ion	Reagent used	Observation
Ca^{2+}	$Na_2C_2O_4$ (sodium oxalate)	White precipitate (CaC_2O_4)
Fe^{3+}	KSCN (pottassium thiocyanate)	Blood red solution ($Fe(SCN)_3$)
NH_4^+	NaOH (sodium hydroxide)	Evolution of NH_3 (basic gas) Changing the red litmus paper to blue
SO_4^{2-}	$BaCl_2$ (barium chloride)	White precipitate ($BaSO_4$)
HCO_3^-	HCl (hydrochloric acid)	Bubbles of CO_2 gas
Br^-	$AgNO_3$ (silver nitrate)	Yellow precipitate (AgBr)
Cl^-	$AgNO_3$ (silver nitrate)	White precipitate (AgCl)

How to find the chemical formula of a salt consisting of a cation and an anion ?

- Suppose the salt consists of A^{x+} and B^{y-} , then the formula will be $(A)_y(B)_x$ >> >> >> A_yB_x

Examples:

Cation	Anion	Formula of salt
Fe^{3+}	SO_4^{2-}	$Fe_2(SO_4)_3$ Iron(III)sulfate
Ca^{2+}	Cl^-	$CaCl_2$ Calcium chloride
Ca^{2+}	Br^-	$CaBr_2$ Calcium bromide
NH_4^+	HCO_3^-	NH_4HCO_3 Ammonium bicarbonate
Fe^{3+}	Cl^-	$FeCl_3$ Iron(III)chloride

Molar Mass of Volatile Liquid

Molar Mass of Volatile Liquid

Objectives:

Determine the molar mass of a volatile liquid using Dumas method

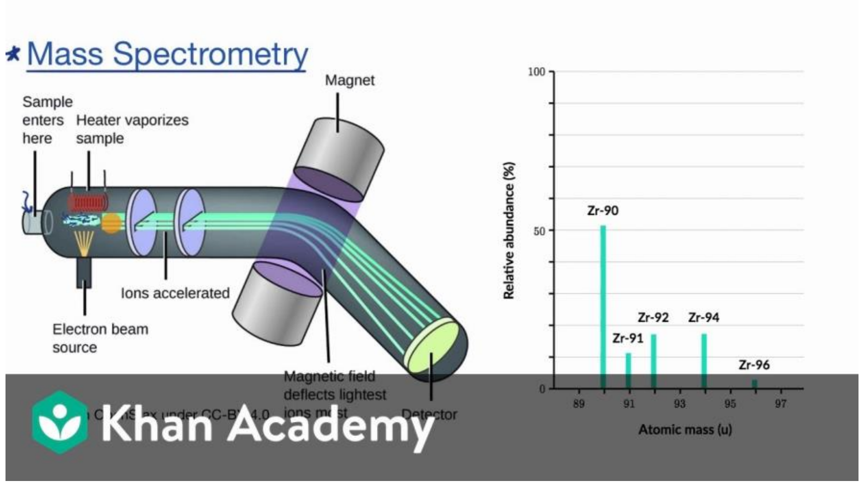
Molar Mass of Volatile Liquid

Introduction:

- Molar mass of many compounds can be determined using modern techniques as the mass spectrometer

- Earlier , very simple methods where used for that purpose

- Among these methods is Dumas method that is applied to volatile liquids



Molar Mass of Volatile Liquid

Introduction:

- **Volatile liquid** is a liquid that is easily evaporate. In other words, it has a low boiling point (below the boiling point of water)

- **Dumas method** is applicable for volatile liquids, because we want to evaporate the sample and deal with it in the gas phase

Conditions:
So our sample must be liquid at room temperature and boils at temperature lower that the boiling point of water

Molar Mass of Volatile Liquid

Introduction:

When the sample is evaporated, we assume that the vapor behaved ideally:

Ideal gas law:

$$PV = nRT$$

$$PV = \frac{mass}{molar\ mass} RT$$

$$molar\ mass = \frac{mass \times R \times T}{P \times V}$$

To determine the molar mass we need to:

	Unit
Determine the mass of the vapor	(g)
Measure its temperature	(K)
Measure its pressure	(atm)
Measure the volume it occupies	(L)

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

Important Conversions

$$T = 0^{o}C + 273 = 273\ K$$

1 atm = 760 mmHg

1 atm = 760 torr

1 atm = 1.01325×10^5 Pa

1 atm = 101.325 kPa.

1 liter = 1000 milliliters

During the experiment

- Waiting the water to boil and all the volatile liquid evaporate.
- Boiling stones will be added to avoid bumping
- The hole will be made to equalize internal pressure with the external pressure and let the excess vapor molecules to escape
- Leaving the flask to cool and the condensation takes place.
- Computing the mass of condensed vapor
- Finding the molar mass according to the ideal gas law

Sample of report sheet

Molar Mass of Volatile Liquid

Boiling point of water	96.0	ºC
Atmospheric pressure	685	mmHg
Mass of empty flask	86.05	g
Mass of flask and condensed vapor	86.50	g
Mass of condensed vapor	86.50-86.05=0.45	g
Volume of flask	220	mL
Boiling point of water	96.0+273=369	K
Atmospheric pressure	685/760=0.901	atm
Volume of flask	220/1000=0.220	L
Gas constant (R)	0.0821	L.atm /mol.K
Molar mass of unknown	68	g/mol

$$\frac{0.45 \times 0.0821 \times 369}{0.901 \times 0.220}$$

Molar volume of hydrogen gas

- Acids react with most metals to form a salt and hydrogen gas
- Metals that are more active than acids can undergo a **single displacement reaction**.

For example, zinc metal reacts with hydrochloric acid, producing zinc chloride and hydrogen gas.

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

- Molar volume is defined as the *volume of one mole of gas* , depending on pressure and temperature
- **Common symbols** V_m, $\tilde{V}$
- Molar volume of any ideal gas (at STP conditions) is 22.4 liters
- What are **STP** conditions ?
  (Standard Temperature and Pressure)
  0 °C (273.15 K) and 1 atm (101325 Pa)

- For ideal gases

$$PV = nRT$$

- To get the molar volume of H_2 gas produced in our experiment, we have to apply **Boyle's Law**
- **What is Boyle's Law?**
- Boyle's law is a gas law which states that the pressure exerted by a gas (of a given mass, kept at a constant temperature) is inversely proportional to the volume occupied by it. In other words, the pressure and volume of a gas are inversely proportional to each other as long as the temperature and the quantity of gas are kept constant.

- If temperature is constant, then:

↓

$$P_1V_1 = P_2V_2$$

P = Pressure of the gas
V = Volume of the gas

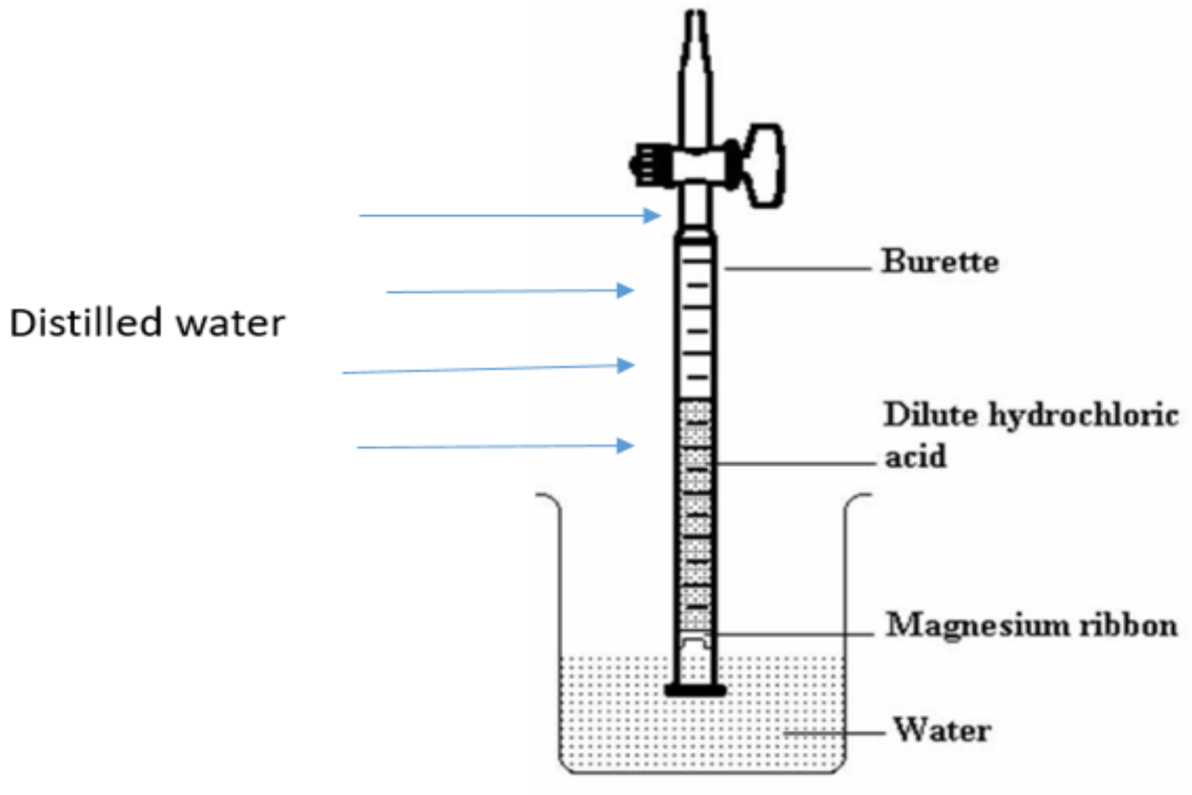
Temperature must be constant

- If temperature is variable, then:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P = Pressure of the gas
T = Temperature of the gas
V = Volume of the gas

Experimental part:



The magnesium metal is attached to the stopper of the burette. The aqueous hydrochloric acid is in the burette. The inverted burette is placed inside a water bath. When the metal comes in contact with the acid, the reaction produces hydrogen gas. The hydrogen gas fills the eudiometer. The reaction continues until all the magnesium metal (which is the limiting reagent in this experiment) is completely consumed.

As shown in Figure , the $H_{2(g)}$ that is formed is combined with water vapor. The water vapor is a result of the vapor pressure of water found in the aqueous medium. Therefore, the two gases: $H_{2(g)}$ and $H_2O_{(g)}$ are both found in the burette

Assuming the surface of liquid in the burette is level with the water outside, the total pressure of the gases inside the burette is the same as the atmospheric pressure. A barometer that is found in the laboratory is used to determine the atmospheric pressure.

Therefore, according to **Dalton's law**:

$$P_{atm} = P_{H_2} + P_{H_2O}$$

Sample of data and calculations

In a certain experiment, 0.0369 grams of magnesium was reacted with excess HCl. The reaction resulted in 38.2 mL of $H_{2(g)}$ at 22°C. The barometric pressure was recorded as 749.2 mm of Hg. Calculate the molar volume of the gas at STP.

Mass of magnesium metal = 0.0369 grams

Atomic mass of magnesium metal = 24 grams/mol

Moles of magnesium metal = 0.00154 moles

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

$$\text{Moles of } H_{2(g)} = \text{Moles of Mg} \times \frac{1 \text{ mole } H_2}{1 \text{ mole Mg}} = 0.00154 \text{ moles}$$

Temperature of $H_{2(g)}$ = 22°C = 295 K = $\mathbf{T_1}$

P_{H2O} = 19.8 mm of Hg

Barometric pressure = 749.2 mm of Hg

P_{H2} = 749.2 – 19.8 = 729.4 mm of Hg = $\mathbf{P_1}$

Volume of $H_{2(g)}$ = 38.2 mL = 0.0382 L = $\mathbf{V_1}$

At STP:

Pressure = 760 mm of Hg = $\mathbf{P_2}$

Temperature = 273 K = $\mathbf{T_2}$

$$V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2} = \frac{729.4 mmHg \times 0.0382 L \times 273 K}{295 K \times 760 mmHg} = 0.0339 L$$

Moles of H_2 = 0.00154 moles

Therefore molar volume = $\frac{0.0339 L}{0.00154 moles} = 22.0 \frac{L}{mol}$

Colligative Properties

Molar mass determination

What is colligative property?

- is a property of a solution that is <u>dependent</u> on the ratio between the total number of solute particles (in the solution) to the total number of solvent particles.
- Colligative properties are <u>not dependent</u> on the chemical nature of the solution's components. Thus, colligative properties can be linked to several quantities that express the concentration of a solution, such as molarity, normality, and molality.

The four colligative properties that can be exhibited by a solution are given below:

- Boiling point elevation
- Freezing point depression
- Relative lowering of vapour pressure
- Osmotic pressure

In our experiment, we will study
Freezing point depression

What do we mean by freezing point ?

- It’s the temperature at which the vapour pressure of the substance in it’s liquid phase is equal to it’s vapour pressure in the solid phase.

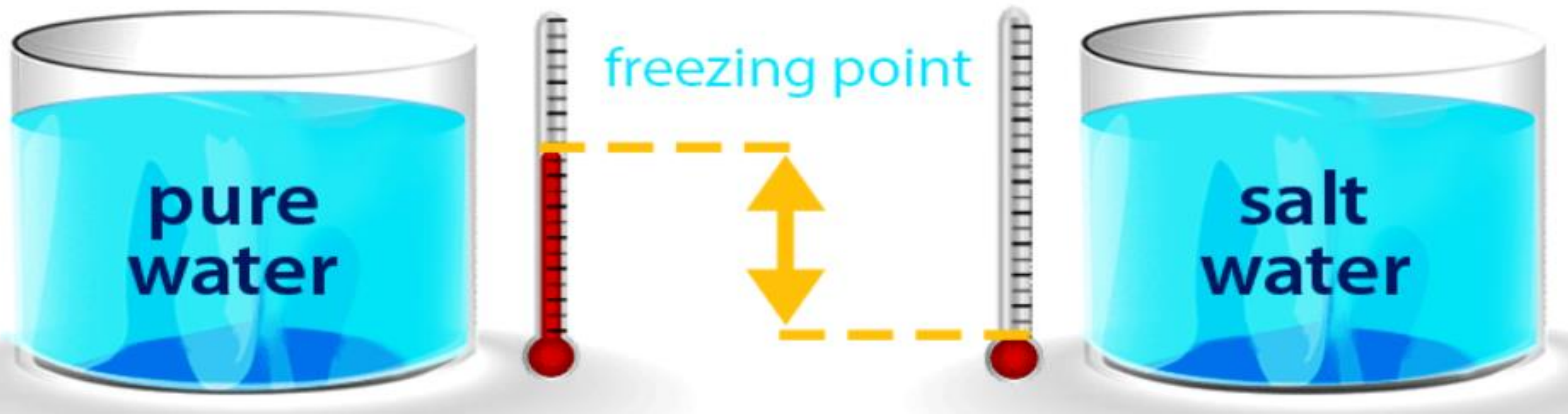
- when a non-volatile solid is added to the solvent, it’s vapour pressure decreases

- The difference between the freezing point of the pure solvent and it’s solution is called freezing point depression .
- **Freezing point depression** is a phenomenon that lowers the freezing point of a solvent when a solute is added to it.

An example from our life:

The freezing point of seawater is below 0^oC; it remains liquid at temperatures below the freezing point of pure water. This is caused by the salts that are dissolved in it.

The difference in temperature between a solution's freezing point and the freezing point of its pure solvent

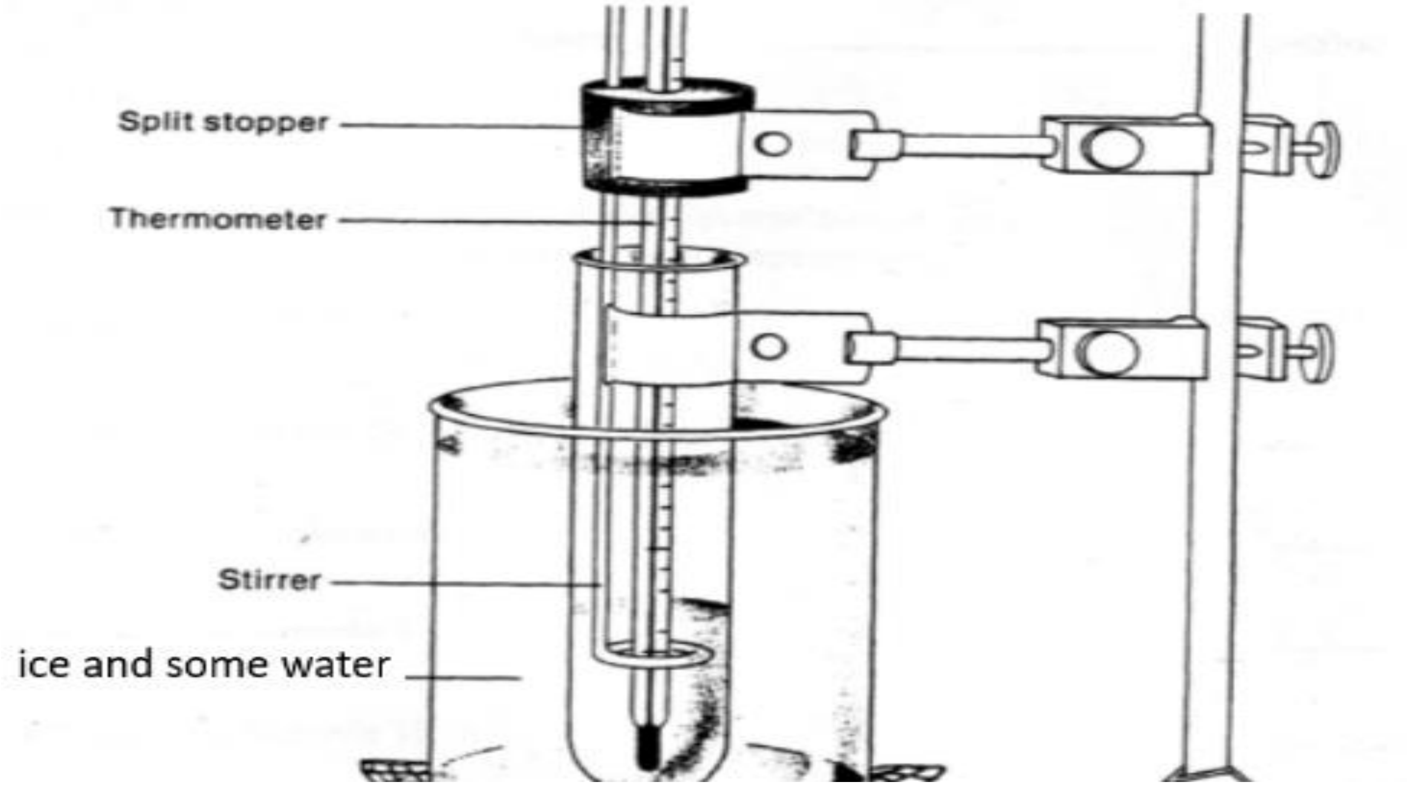


freezing point depression

How to use freezing point depression to determine molar mass of a solute?

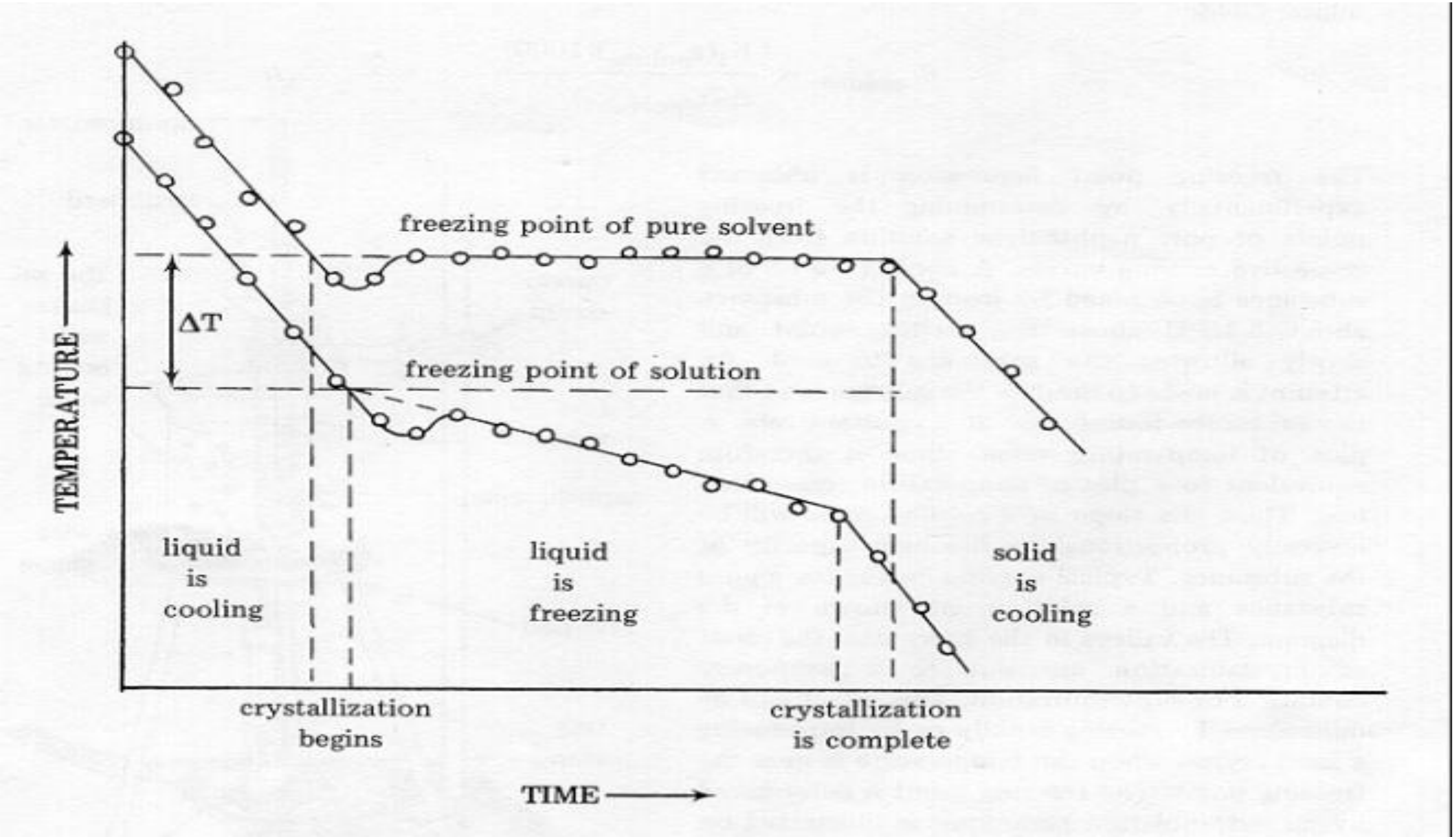
Refer to the link for procedure you will use

Freezing point depression apparatus



(Freezing point curves)

<u>Note that the long stability in temperature indicates the curve of the pure solvent</u>



Notes:

1) Dissociation of a solute in a solvent will lead to more particles in the solution, thus causing a larger depression

Example:

- Dissociation of salt molecules in water cause a larger decrease in the freezing point of the mixture,
- Electrolytes, when dissolved into water, dissociate into ions, generating more particles in solution (moles solute increase), which causes a larger freezing point depression than that of a non-electrolyte at the same molality
- Example: NaCl dissociates in water into Na^+ and Cl^- to cause increasing in moles of solute thus increasing depression in freezing point

Notes :

2) Association of a solute in a solvent will lead to less particles in the solution, thus causing a lower depression

Example:

- When dissolving of benzoic acid (**slightly polar since the bulk of it is nonpolar**) in benzene will lead to association of benzoic acid due to the intermolecular hydrogen bond formed between solute molecules. Thus, moles of solute in the solution decrease causing a lower depression

Calculations:

- **Freezing-Point Depression (ΔT_f):** The freezing point of the pure solvent (T°_f) minus the freezing point of the solution (T_f).

$$\Delta T_f = T^\circ_f - T_f$$

ΔT_f is proportional to concentration:

$$\Delta T_f = K_f\, m$$

K_f = molal freezing-point depression constant.

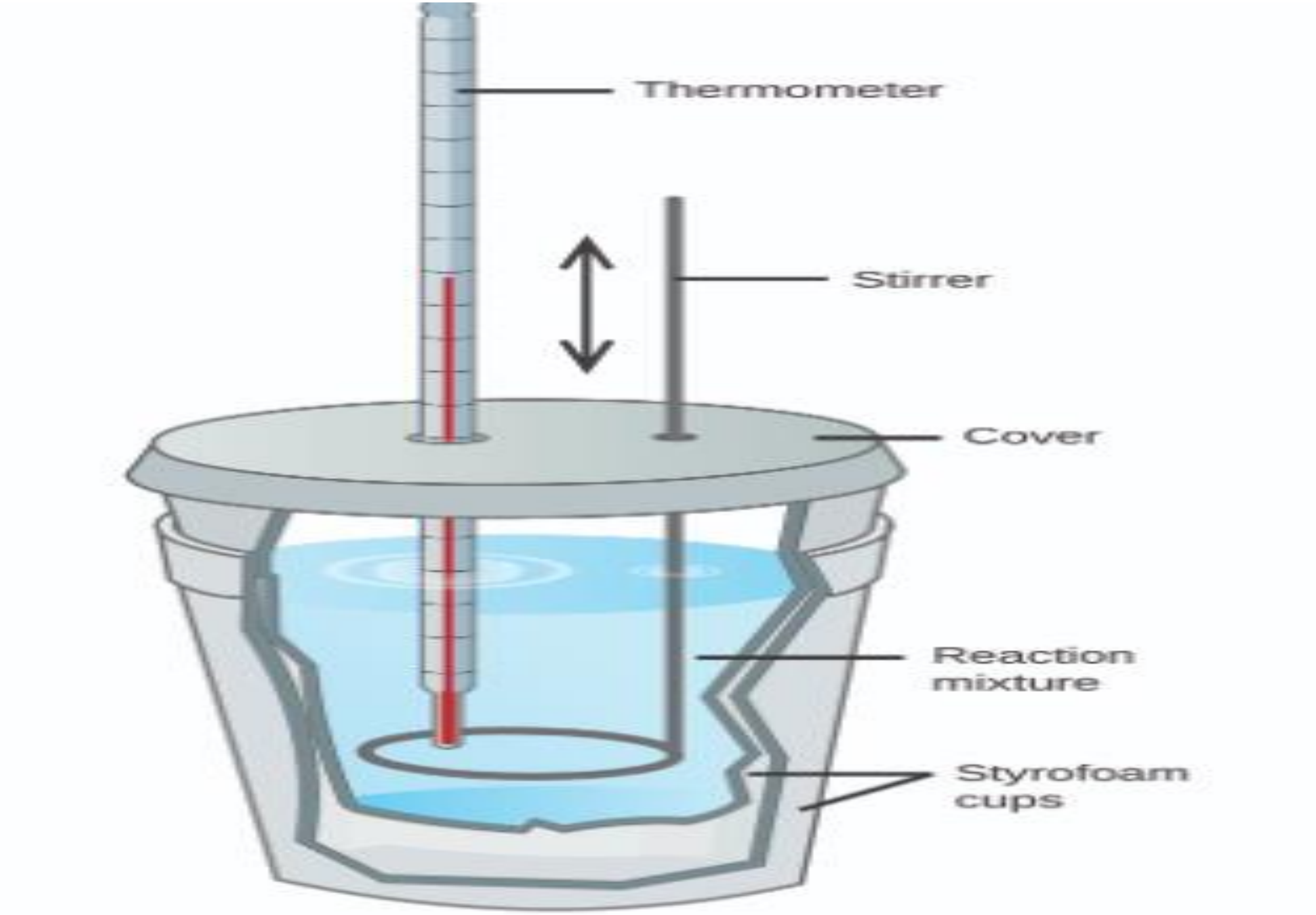
$$\text{molality} = \frac{T_f(\textit{solvent}) - T_f(\textit{mixture})}{k_f(\textit{solvent})} \qquad (1)$$

$$n_{\text{unknown added}} = \text{molality}_{\text{unknown}}(\text{mass}_{\text{solvent (in kg)}}) \qquad (2)$$

$$\text{M.W.} = \frac{\text{total g of unknown}}{n_{\text{unknown}}} \qquad (3)$$

$$\text{Molality(m)} = \frac{\textit{moles of solute}}{\textit{mass(in Kg) of Solvent}}$$

Calorimetry



Introduction:

Heat lost equals heat gained in an isolated system

- This law is known as the law of energy conservation
- It will be the basis of our experiment today

Introduction:

- Chemical and physical processes are accompanied by changes in energy.
- Mostly, this energy is observed as heat flow **into** or **out** of the system under study.
- **First law** of thermodynamics states that:

Heat lost equals heat gained in an isolated system

Objectives:

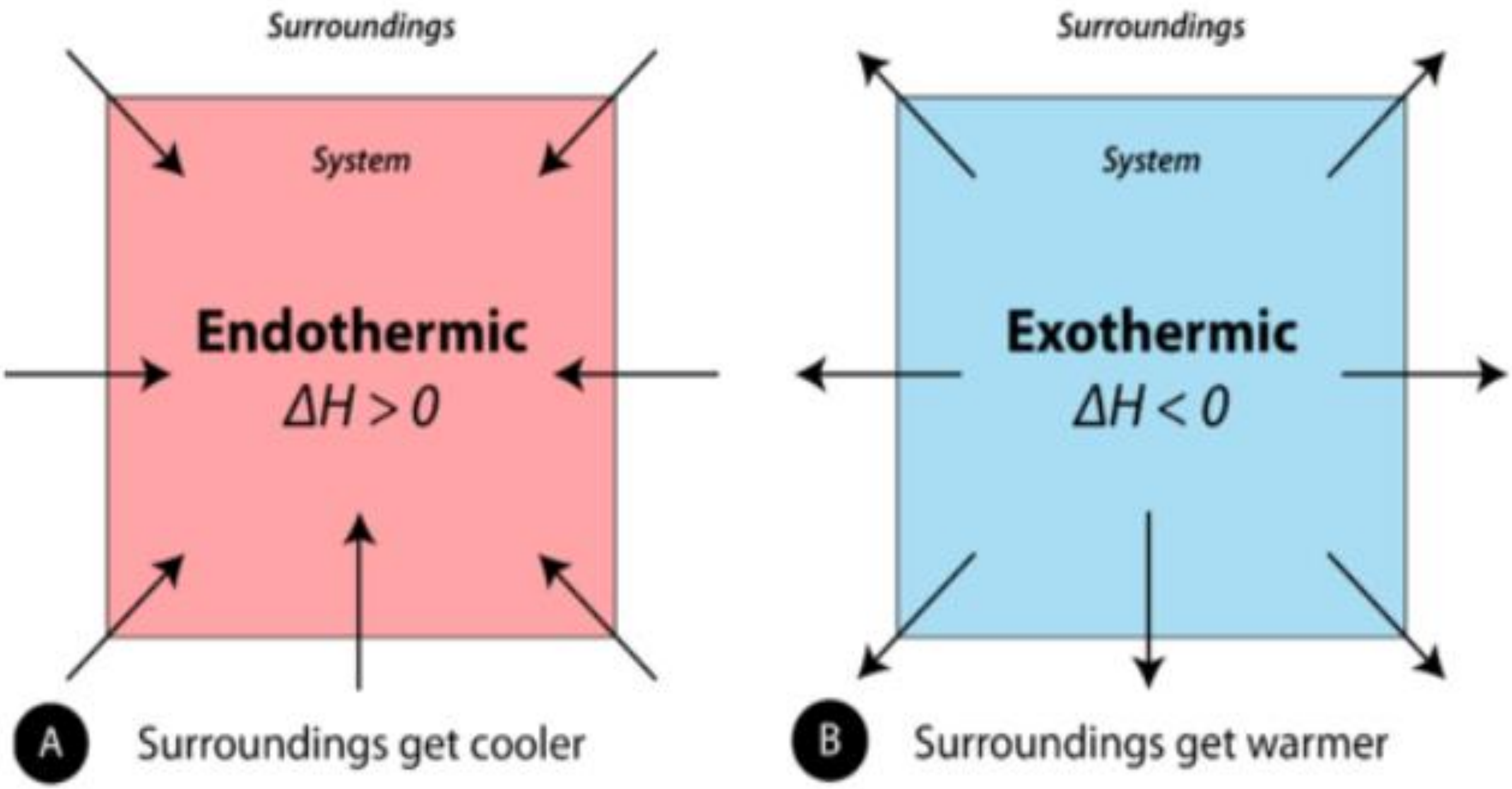
- Measure the enthalpy change when solution of sodium hydroxide (NaOH) is prepared in a coffee-cup calorimeter **(Heat of solution)**

- Measure the enthalpy change when sodium hydroxide (NaOH) and hydrochloric acid (HCl) are mixed in the above coffee-cup calorimeter **(Heat of reaction)**

Reactions are classified into:

1- **Endothermic reactions** require energy, so energy is a reactant. Heat flows from the surroundings to the system (reaction mixture) and the enthalpy of the system increases (ΔH is positive).

2- **Exothermic reactions** evolve energy, so energy is a product. Heat flows from the system (reaction mixture) to surroundings and the enthalpy of the system decreases (ΔH is negative).


Surroundings
System
Endothermic
ΔH > 0
Surroundings
System
Exothermic
ΔH < 0
A
Surroundings get cooler
B
Surroundings get warmer

The **Heat of solution** is defined as the difference in the enthalpy related to the dissolving substance in a solvent at constant pressure which is leading in infinite dilution. The unit of solution enthalpy is KJ/mol. The enthalpy change is observed when the solute is dissolved in the solvent

The **Heat or enthalpy of neutralization** is the change in enthalpy occurring when an aqueous acid reacts with an aqueous base to form salt and one mole of water under standard conditions

The **Heat of Reaction** (also known and Enthalpy of Reaction) is the change in the enthalpy of a chemical reaction that occurs at a constant pressure.

Experimental:

- Our calorimeter is simply a polystyrene coffee cup

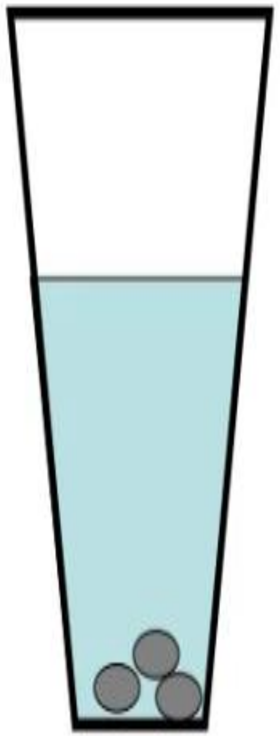
- Polystyrene is heat insulator which (does not allow heat to be transferred through it) to insure **insulated system conditions**

Experimental:

Heat of solution

1. $NaOH(s) + H_2O \rightarrow NaOH(aq) + H_2O \quad \Delta H_1$

- **Weigh** 50.0 ml of cold water inside the calorimeter, **record its mass**.
- **Pour** the cold water into the calorimeter and **record its temperature**.
- **Weigh** a piece of NaOH
- Dissolve NaOH in the cold water
- **While dissolving, Record the highest temp. of the solution obtained**.



Dissolving solid sodium hydroxide in water

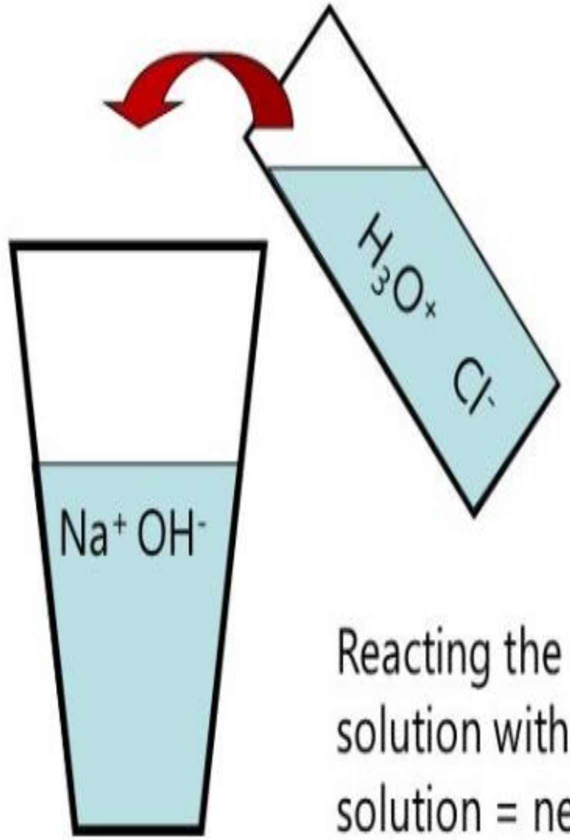
This process produces sodium and hydroxide ions ie. NaOH (aq) solution.

Experimental:

Heat of reaction

2. NaOH (aq) +HCl (aq) → NaCl (aq) + H_2O ΔH_2

- Weigh a piece of NaOH (similar as in part 1) solution inside the calorimeter
- Put in the calorimeter
- Weigh 50.0 mL of HCl solution. **record its temperature**.
- Pour the acid solution into the calorimeter, stir with the thermometer and monitor the change in temp. **Record the highest temp of the mixture**.



Reacting the sodium hydroxide solution with a hydrochloric acid solution = neutralization

Results:

Enthalpy of a reaction (ΔH_{soln}) and (ΔH_{rxn}) as KJ/mol

For part 1:

Heat of solution = mass of solution * S * ΔT

Heat calculated previously is divided by number of moles of NaOH

NOTE: temperature increases, thus (ΔH_{soln}) is negative

Exothermic

Results:

For part 2 :

Heat of reaction = mass of solution * S * ΔT

Heat calculated previously is divided by number of moles of NaOH

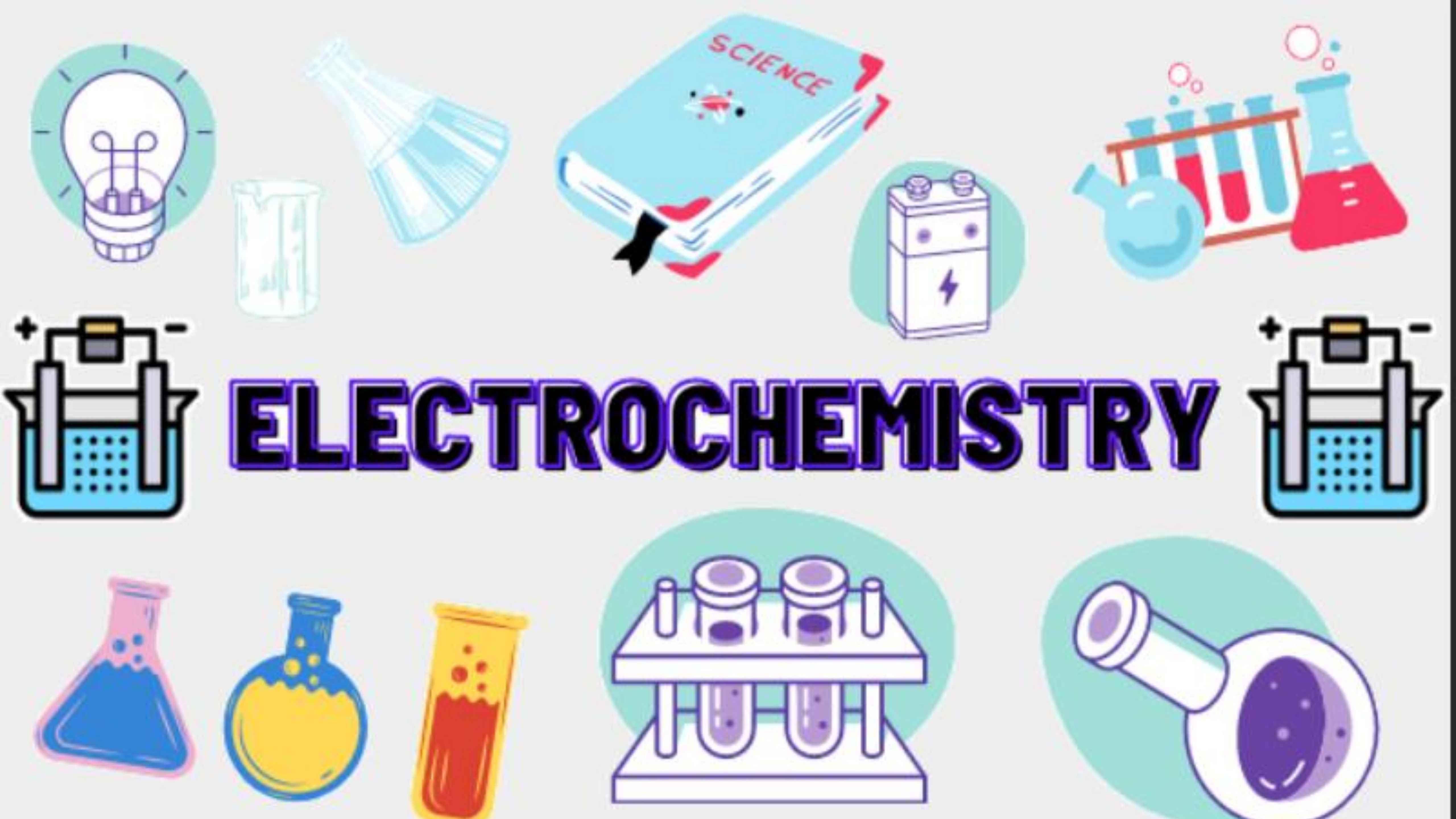
NOTE: temperature increases, thus (ΔH_{rxn}) is negative

Exothermic reaction

$$\Delta H_{rxn} = \Delta H_{soln} + \Delta H_{neutralization}$$

Thus,

$$\Delta H_{neutralization} = \Delta H_{rxn} - \Delta H_{soln}$$


SCIENCE
ELECTROCHEMISTRY

What is electrochemistry?

- A science that deals with the study of the relationship between electrical energy and chemical changes.
- Chemical reactions that involve the input or generation of electric currents are called electrochemical reactions.

Such reactions are broadly classified into two categories:

1- Production of chemical change by electrical energy ➡ electrolysis

2- Conversion of chemical energy into electrical energy ➡ redox reactions.

Electrochemical cells

They are the equipments in which the conversion of chemical energy into electrical energy takes place.

Types of electrochemical cells:

1- Galvanic cells

2- Electrolytic cells

In our experiment we will deal with galvanic cells in which redox reactions take place

Galvanic cells

- Conversion of chemical energy into electrical energy
- Electricity can be obtained with the help of a redox reaction.
- Redox reactions: reactions involve oxidation and reduction reactions

  Oxidation reaction: loss of electrons (increasing in oxidation number)

  Reduction reaction: gain of electrons (decreasing in oxidation number)
- The oxidation and reduction take place in two separate compartments.
- Each compartment consists of an electrolyte solution and a metallic conductor, which acts as an **electrode**
- Each compartment is called **half cell**
- Oxidation half is called an **anode**
- Reduction half is called **cathode**

Components of galvanic cell:

- 1- **Anode**
- 2- **Cathode**
- 3- **Electrolyte solutions** in each compartment
- 4- **Salt bridge**: an inverted U-tube filled with a concentrated solution of inert electrolytes. It is used to:

  1-maintain the charge balance

  2-complete the circuit by allowing the flow of ions through it
- 5- **External Circuit:** Conducts the flow of electrons between electrodes in an external circuit. (Note : electrons flow from anode to cathode)
- 6- **Voltmeter** : to measure the potential of the cell

For a galvanic cell consisting of Zinc and Copper

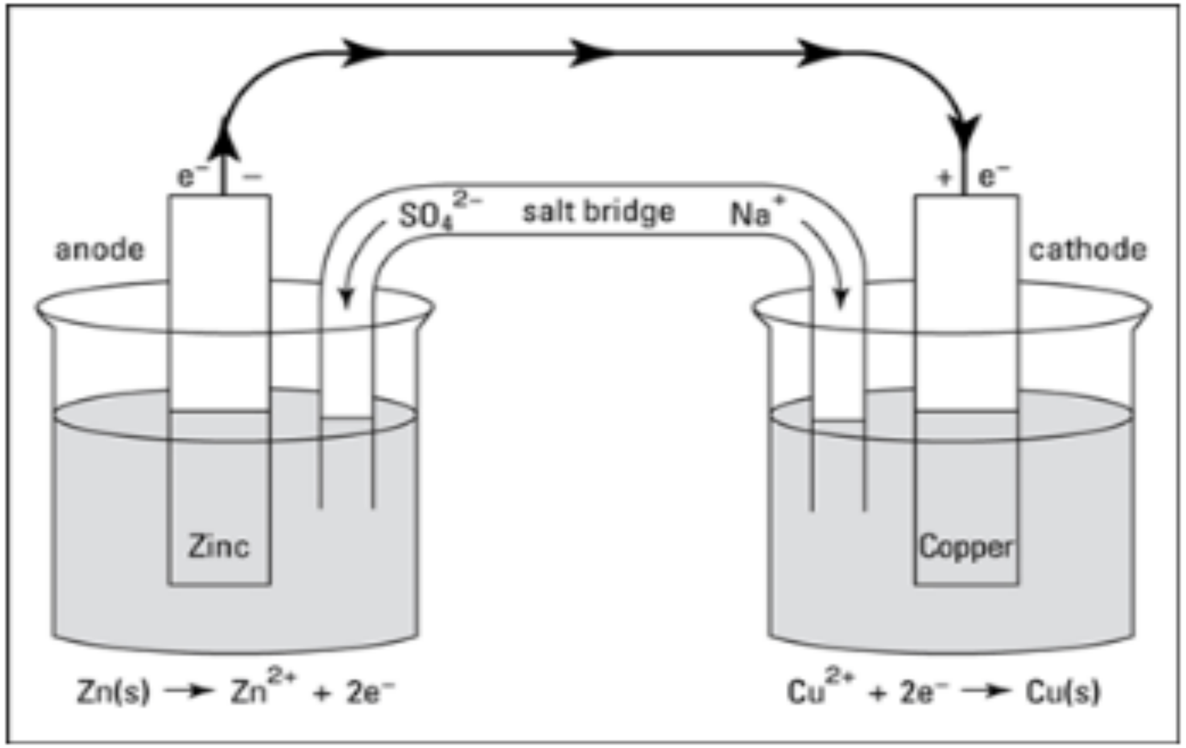
$Zn(s) + Cu^{+2}(aq) \rightarrow Zn^{+2} + Cu(s)$

At anode (oxidation half),

$Zn(s) \rightarrow Zn^{+2} + 2e^{-}$

At cathode(reduction half),

$Cu^{+2}(aq) + 2e^{-} \rightarrow Cu(s)$



Electrode potential

- When two-electrodes are dipped in their respective ions, a tendency for one of the electrodes (anode) to undergo oxidation, whereas the ion of the other electrode (cathode) has the tendency to gain an electron.
- This tendency of losing of electrons (oxidation) or gaining of electrons (reduction) is called **electrode potential**.

- **Standard electrode potential (E^0)**: the electrode potential of an electrode relative to a standard hydrogen electrode under standard conditions.

The standard conditions taken are as follows:

- 1 molar concentration of each ion in the solution.
- A temperature of 298 K.
- 1 atm pressure.

Standard electrode potential (E^0):

$F_2(g) + 2e^-$	→	$2F^-(aq)$	+2.87
$Au^+(aq) + e^-$	→	$Au(s)$	+1.68
$Cl_2(g) + 2e^-$	→	$2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^-$	→	$2H_2O(1)$	+1.23
$Ag^+(aq) + e^-$	→	$Ag(s)$	+0.80
$Fe^{3+}(aq) + e^-$	→	$Fe^{2+}(aq)$	+0.77
$I_2(s) + 2e^-$	→	$2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^-$	→	$4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^-$	→	$Cu(s)$	+0.34
$2H^+(aq) + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^-$	→	$Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^-$	→	$Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^-$	→	$Ni(s)$	−0.23
$Co^{2+}(aq) + 2e^-$	→	$Co(s)$	−0.28
$Fe^{2+}(aq) + 2e^-$	→	$Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^-$	→	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	→	$H_2(g) + 2OH^-(aq)$	−0.83
$Al^{3+}(aq) + 3e^-$	→	$Al(s)$	−1.67
$Mg^{2+}(aq) + 2e^-$	→	$Mg(s)$	−2.34
$Na^+(aq) + e^-$	→	$Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^-$	→	$Ca(s)$	−2.87
$K^+(aq) + e^-$	→	$K(s)$	−2.93

- As **E^0** is higher, the ability for reduction increases

For example:

A cell consists of Cu electrode (**E^0** = +0.34 V**)** and Ag electrode (**E^0** = +0.80 V), then :

Anode is Cu (oxidation takes place) **E^0** for the oxidation reaction is -0.34 V

Cathode is Ag (reduction takes place) **E^0** for the oxidation reaction is +0.80 V

$$\text{oxidation: } Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

$$\text{reduction: } 2 \times (Ag^+(aq) + e^- \longrightarrow Ag(s)) \quad \text{or} \quad 2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$$

$$\text{overall: } 2Ag^+(aq) + Cu(s) \longrightarrow 2Ag(s) + Cu^{2+}(aq)$$

NOTE : E^0 is qualitative property (although the half cell of Ag is multiplied by (2), E^0 is not multiplied

E^0 for the cell = **E^0** cathode - **E^0** anode = 0.80 -0.34 0.46 V

OR

E^0 for the cell = **E^0** (for oxidation reaction) + **E^0** (for reduction reaction) =-0.34 + 0.80 = 0.46 V

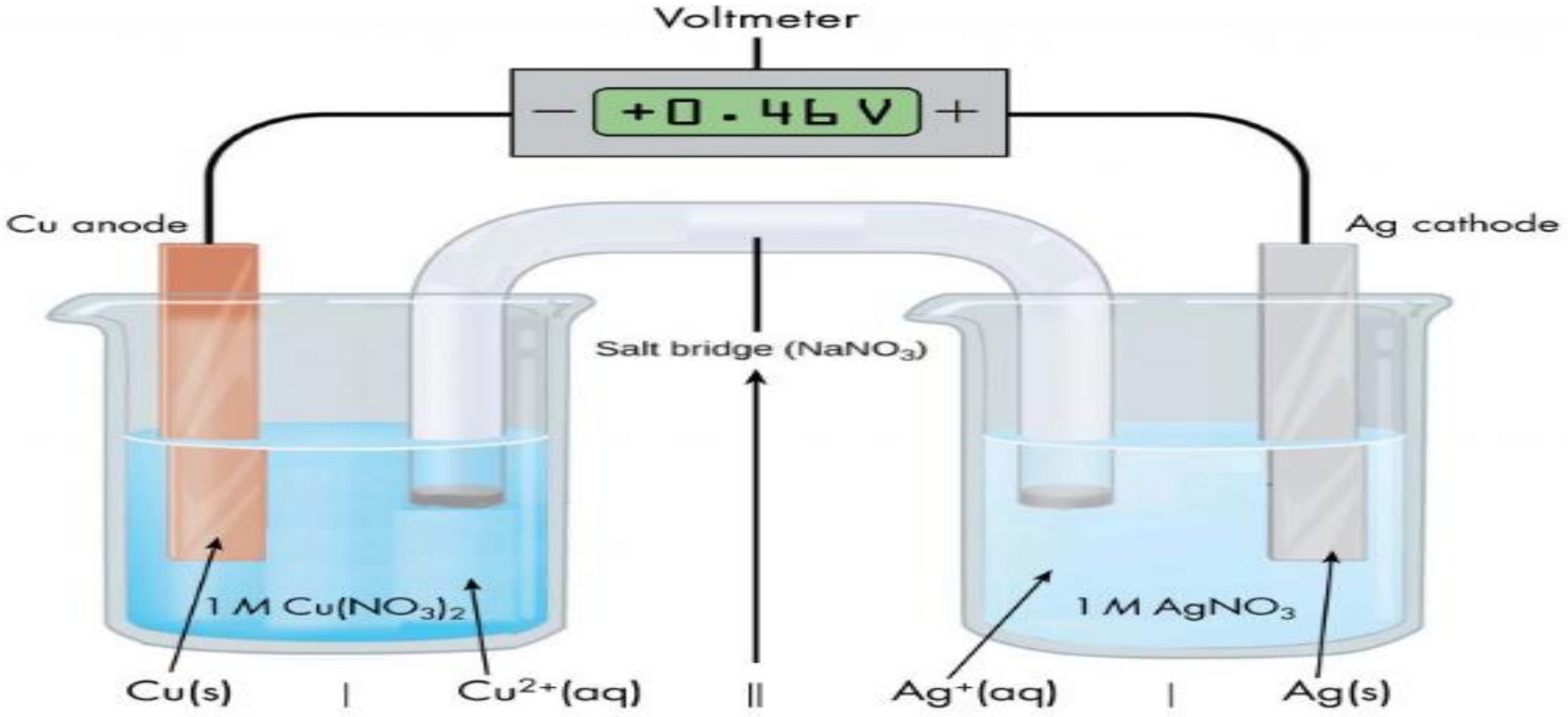
We can represent the cell as follows:

$$\mathrm{Cu}(s)\,|\,\mathrm{Cu}^{2+}(aq,\ 1\ M)\,\|\,\mathrm{Ag}^{+}(aq,\ 1\ M)\,|\,\mathrm{Ag}(s)$$

Oxidation

salt bridge

Reduction



$E^o_{cell} = E_{Cathode} - E_{anode}$

$E_{Cathode}$= standard reduction potential of the cathode.

E_{anode}= standard reduction potential of the anode.

If E^o_{cell} is positive, then the reaction is feasible.

If E^o_{cell} is negative, then the reaction is not feasible.

$Ni^{+2}(aq) + Cu(s) \rightarrow Ni(s) + Cu^{+2}(aq)$

$E^o_{cell} = 0.25 - 0.34$

$= -0.11v$

$Cu(s) + 2H^+(aq) \rightarrow Cu^{+2} + H_2(g)$

$E^o_{cell} = 0 - 0.34 = -0.34\ v$

To calculate the potential of the cell (E) at other conditions, use Nernst equation:

$$E = E^0 - \frac{0.0591}{n}\log(Q)$$

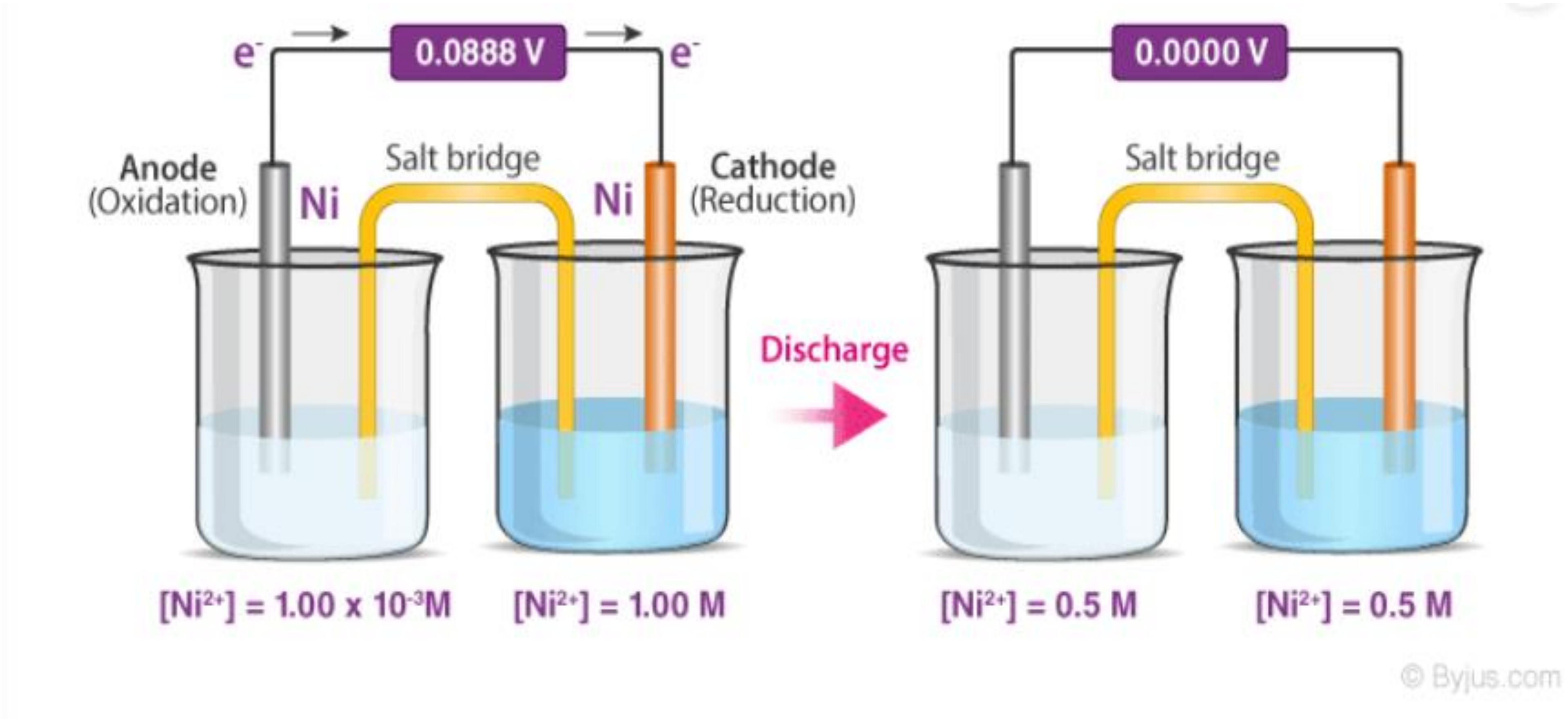
$$Q = \frac{[\text{products}]}{[\text{reactants}]}$$

number of electrons

Concentration cell

- Concentration cells : electrochemical cells that consist of two half-cells wherein the electrodes are the same, but they vary in concentration.
- To reach equilibrium, the more concentrated half cell is diluted and the half cell of lower concentration has its concentration increased via the transfer of electrons between these two half cells.
- Therefore, as the cell moves towards chemical equilibrium,a potential difference is created.

E^0 for the concentration cell = 0

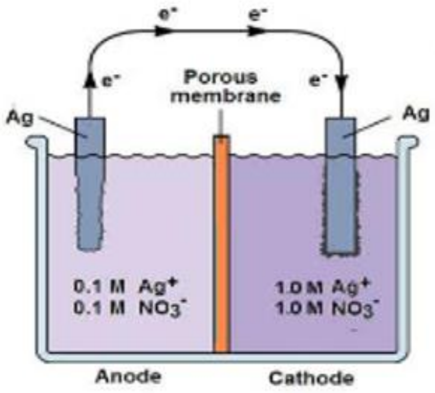


Nernst equation for concentration cell

$$E_{cell} = -\left(\frac{0.0591\ V}{n}\right)\log Q$$

$$Q = \frac{\text{diluted}}{\text{concentrated}}$$

Concentration and the Nernst Equation



Anode: $Ag \rightarrow Ag^+ + e^-$ $E_{1/2} = -0.80\ V$

Cathode: $Ag^+ + e^- \rightarrow Ag$ $E_{1/2} = 0.80\ V$

$$Q = \frac{[Ag^+]_{anode}}{[Ag^+]_{cathode}} = \frac{0.1}{1} = 0.1$$

$$E_{cell} = E^{\circ}_{cell} - (0.0591/n)\log(Q)$$

0 V 1

$$E_{cell} = -(0.0591)\log(0.1) = 0.0591\ V$$