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Chemistry

In society

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Objectives

1. REDOX
2. CHROMATOGRAPHY
3. VOLUMETRIC ANALYSIS

REDOX


REDOX REACTION

REDUCTION-OXIDATION REACTIONS

Redox is a portmanteau of **REDUCTION** & **OXIDATION**


ELECTRON TRANSFER

REDOX

REDUCTION

A **reduction** reaction is a **GAIN** of electrons. If a substance is being reduced it is gaining electrons

REDOX

OXIDATION

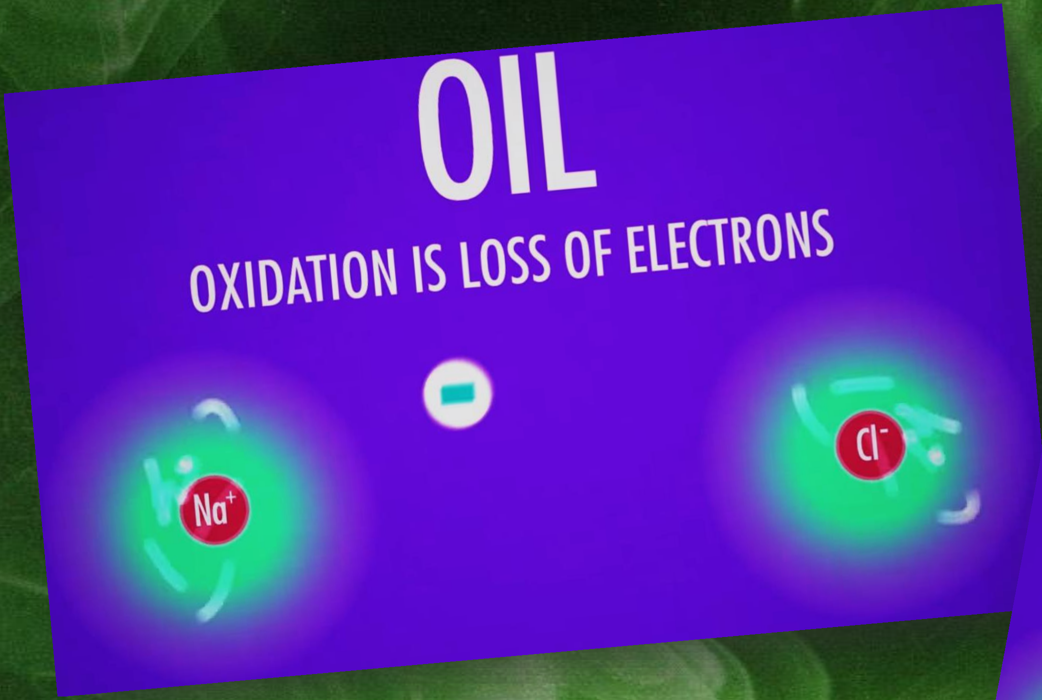
An **oxidation** reaction is a **LOSS** of electrons. If a substance is being oxidised it is losing electrons.

REDOX

OILRIG

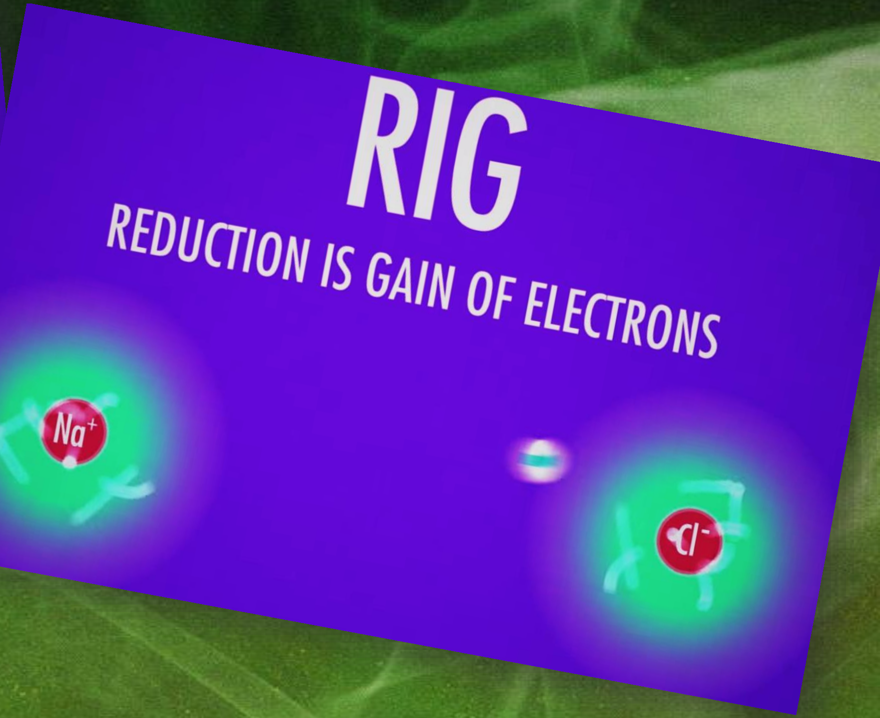
Oxidation
Is
Loss;
Reduction
Is
Gain

REDOX



OIL

OXIDATION IS LOSS OF ELECTRONS



RIG

REDUCTION IS GAIN OF ELECTRONS

REDOX

OXIDATION STATE

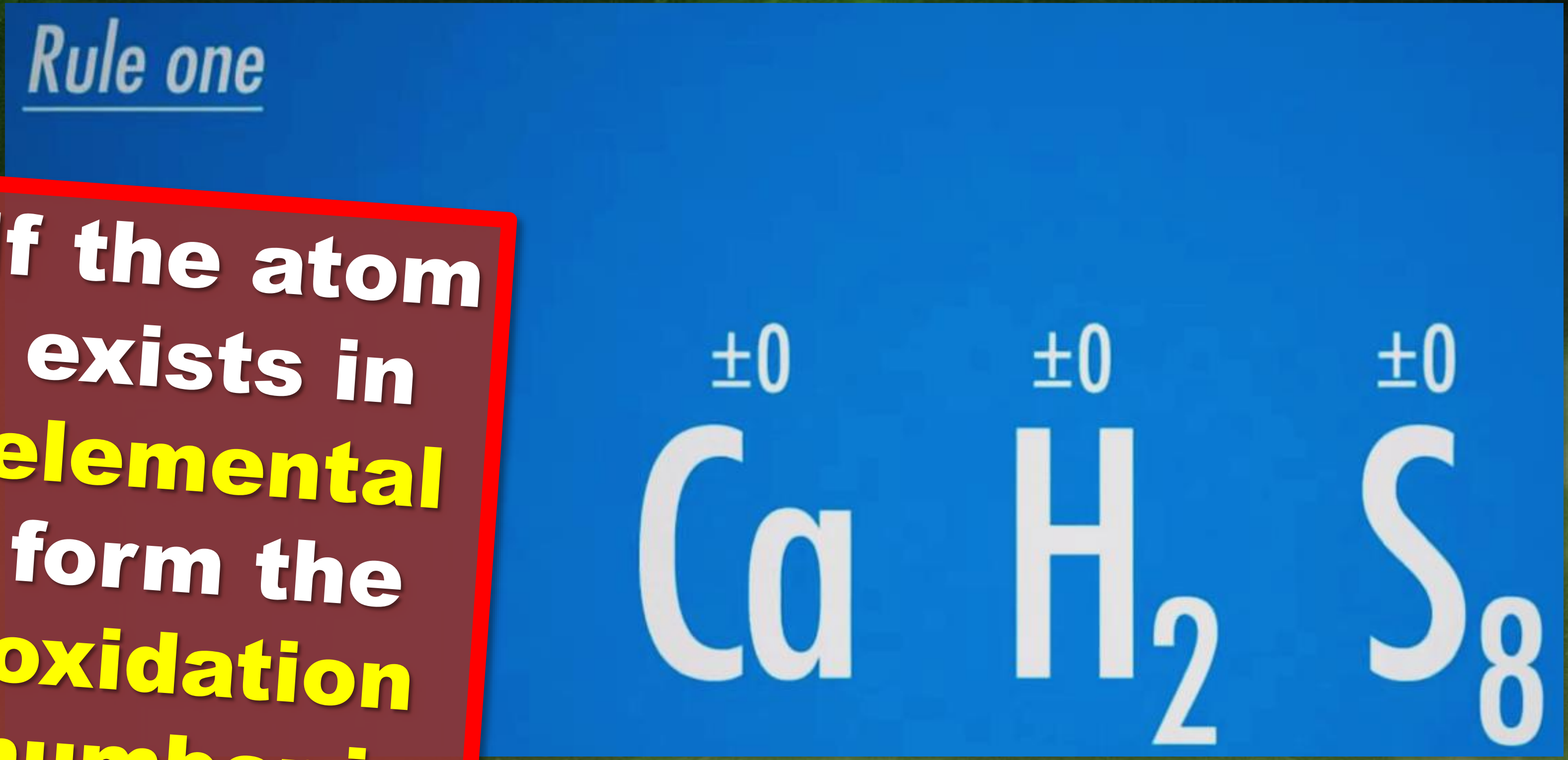
When bonds break who gets the electrons?
The atoms' oxidation state / oxidation number is the number of electrons assigned to them.

REDOX

OXIDATION STATE

How do we figure out the oxidation numbers for the elements?

REDOX

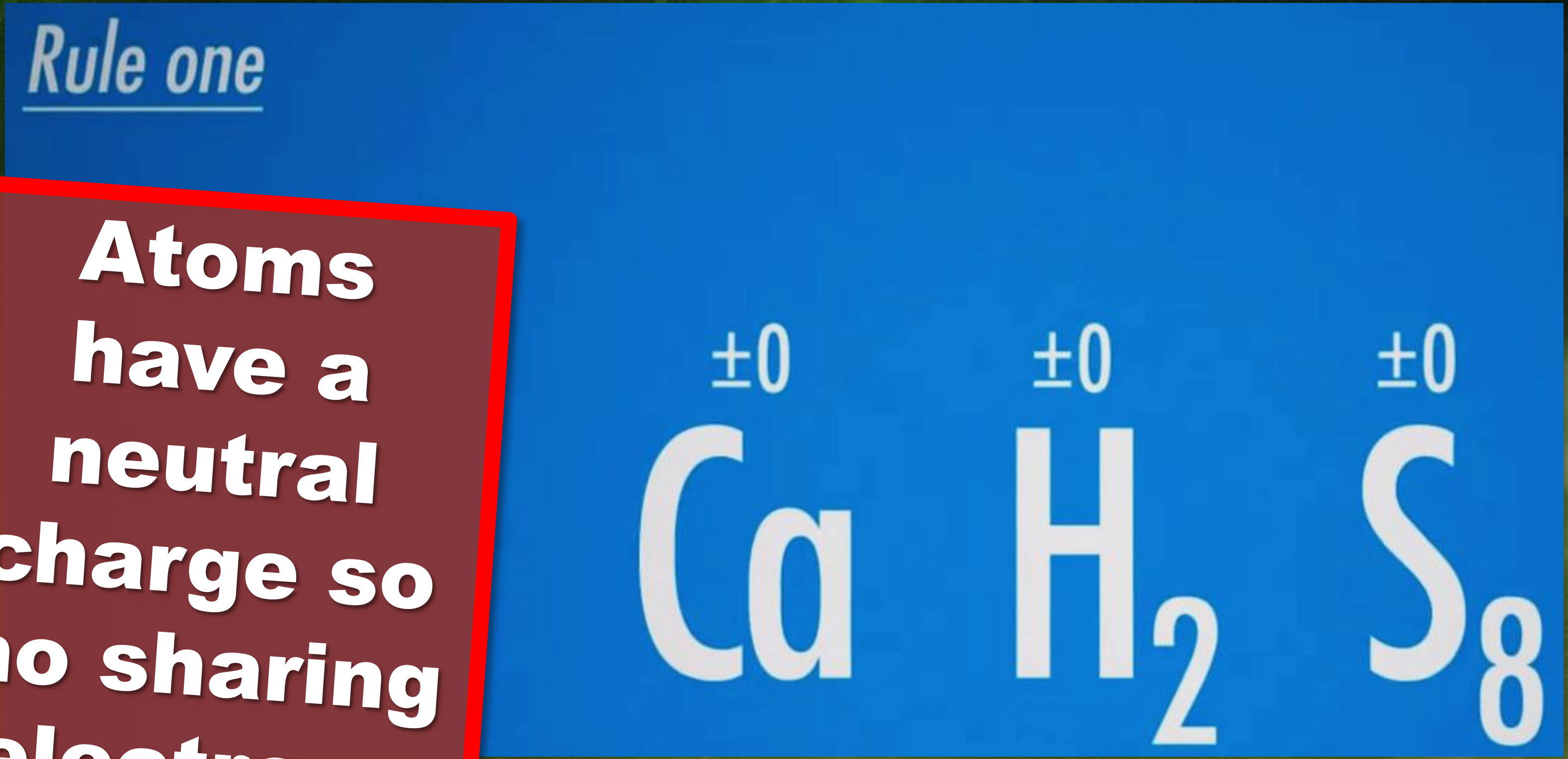


Rule one

$\overset{\pm 0}{\mathrm{Ca}}$ $\overset{\pm 0}{\mathrm{H}_2}$ $\overset{\pm 0}{\mathrm{S}_8}$

If the atom exists in elemental form the oxidation number is zero!

REDOX

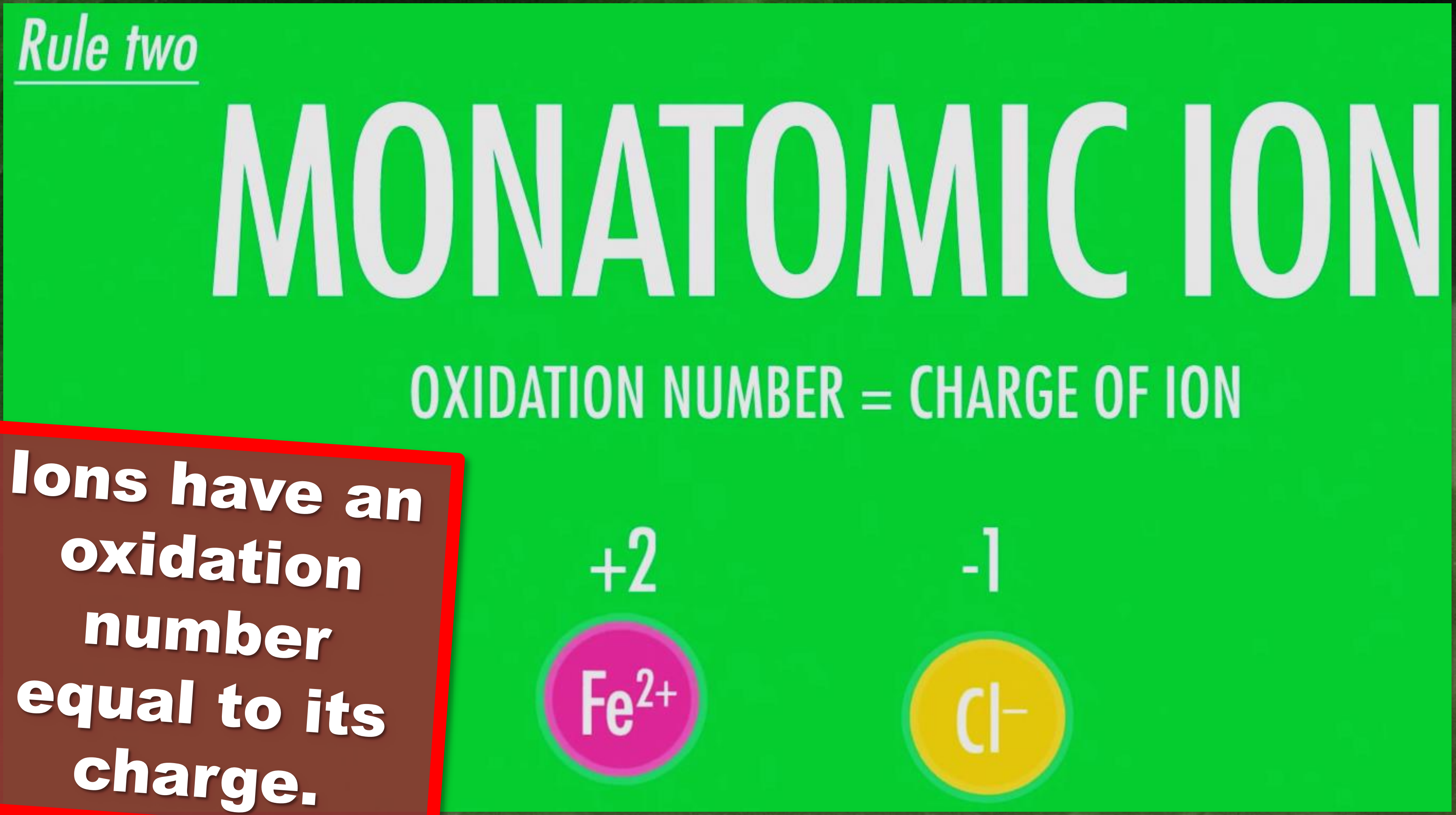


Rule one

$\overset{\pm 0}{Ca}\quad \overset{\pm 0}{H_2}\quad \overset{\pm 0}{S_8}$

Atoms have a neutral charge so no sharing electrons with itself.

REDOX



Rule two

MONATOMIC ION

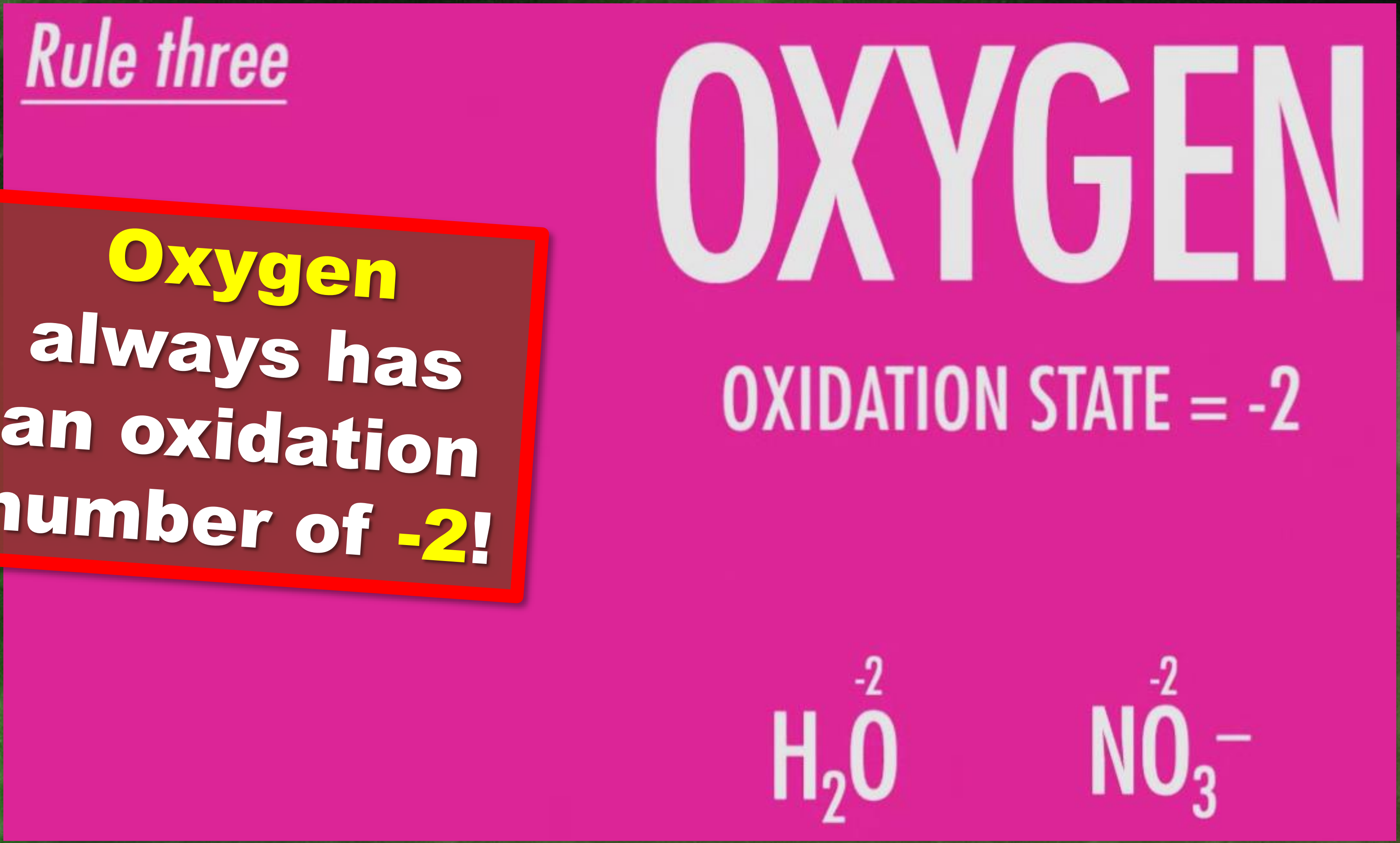
OXIDATION NUMBER = CHARGE OF ION

+2 Fe^{2+}

-1 Cl^-

Ions have an oxidation number equal to its charge.

REDOX



Rule three

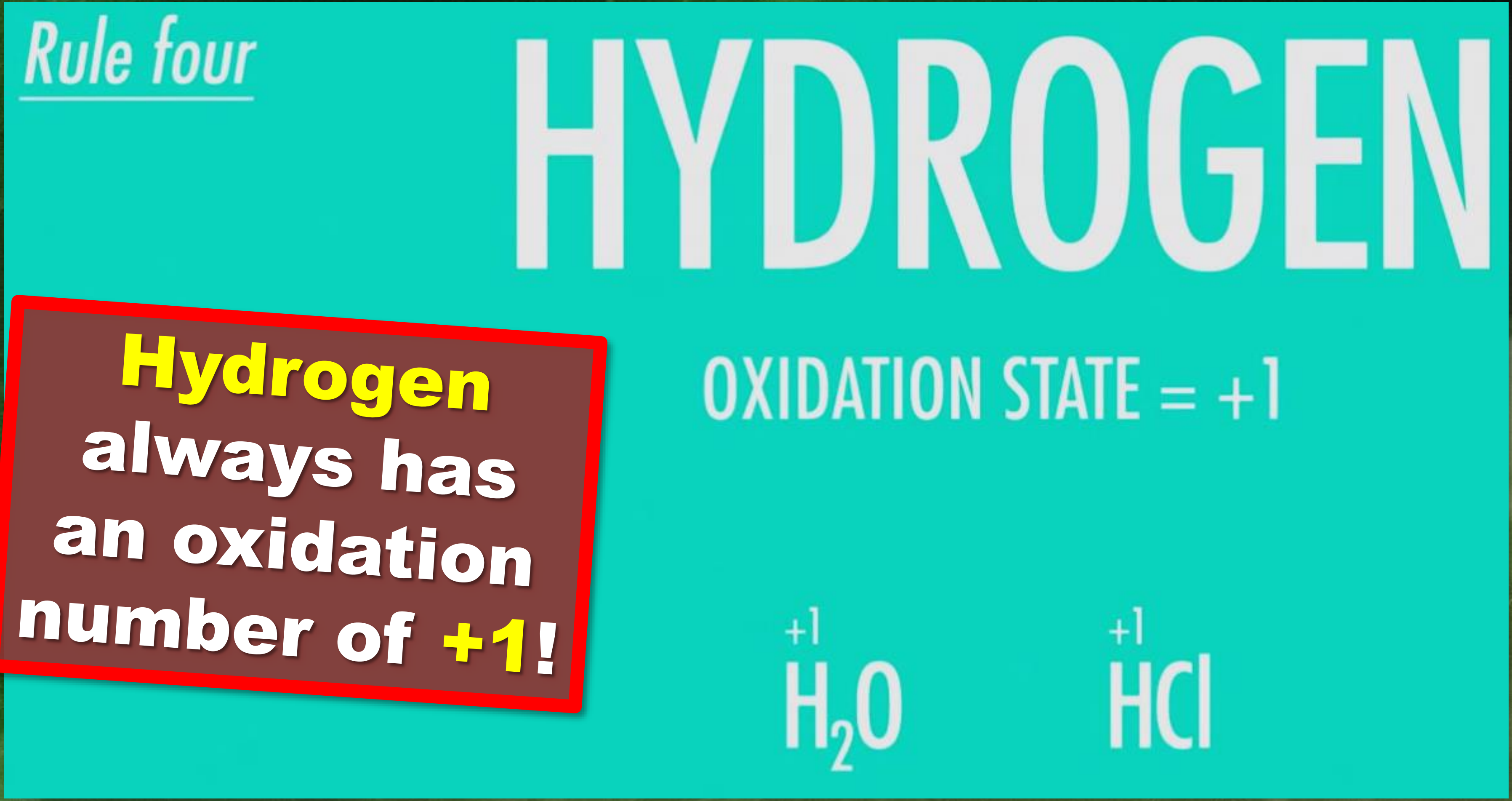
OXYGEN

OXIDATION STATE = -2

$H_2\overset{-2}{O}$ $N\overset{-2}{O}_3^{-}$

Oxygen always has an oxidation number of -2!

REDOX



Rule four

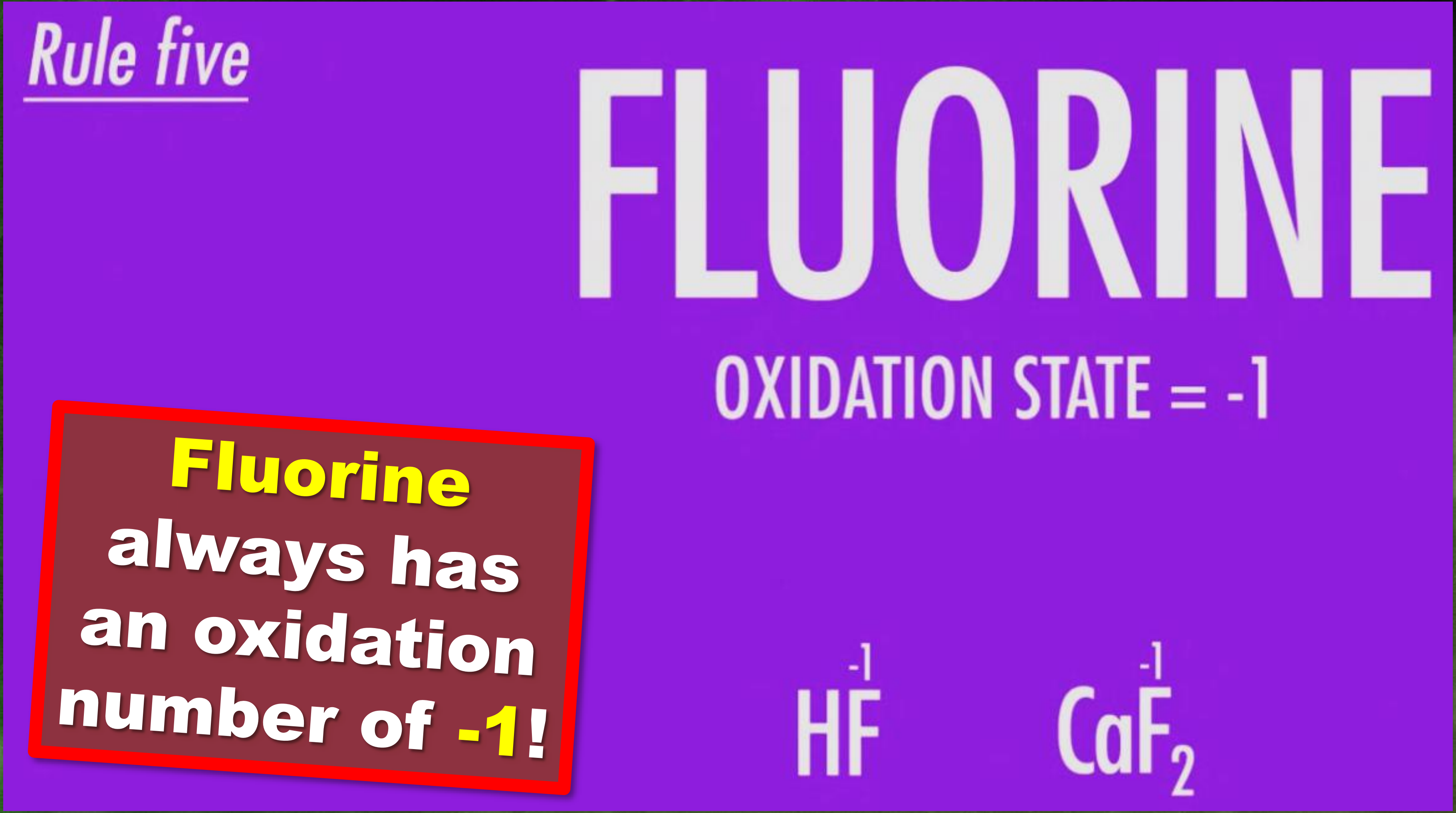
HYDROGEN

OXIDATION STATE = +1

$\overset{+1}{H}_2O$ $\overset{+1}{H}Cl$

Hydrogen always has an oxidation number of **+1**!

REDOX



Rule five

FLUORINE

OXIDATION STATE = -1

$H\overset{-1}{F}$ $Ca\overset{-1}{F}_2$

Fluorine always has an oxidation number of -1!

REDOX

RULE SIX

Other atoms get the charge they prefer, as long as the sum of oxidation numbers for all atoms = the total charge on the particle.

REDOX

RULE SIX

Other atoms get the charge they prefer, as long as the sum of oxidation numbers for all atoms = the total charge on the particle.

Fe_2O_3

?

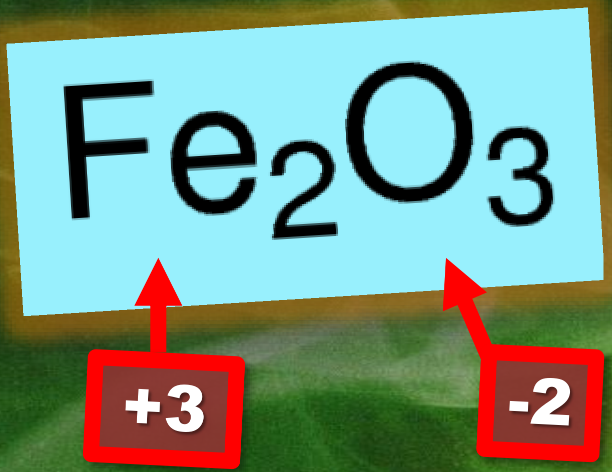
-2

ELEMENT	“CHARGE”
OXYGEN	(3 X -2)
IRON	?
	0

REDOX

RULE SIX

Other atoms get the charge they prefer, as long as the sum of oxidation numbers for all atoms = the total charge on the particle.



Fe_2O_3

+3

-2

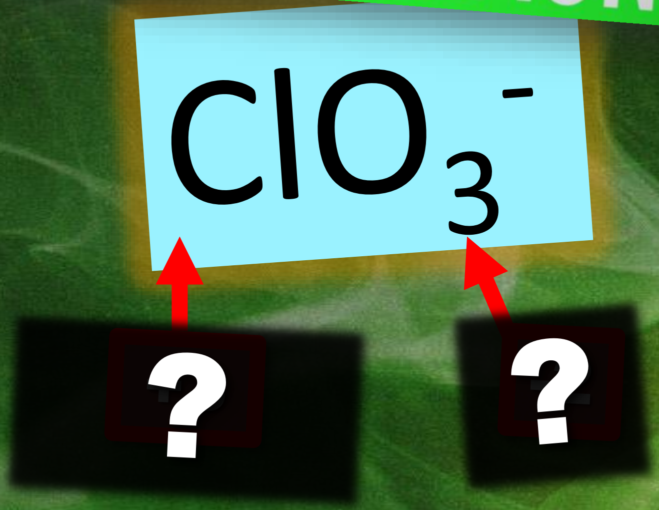
ELEMENT	“CHARGE”
OXYGEN	(3 x -2) = -6
IRON	(2 x +3) = 6
	0

REDOX

RULE SIX

Other atoms get the charge they prefer, as long as the sum of oxidation numbers for all atoms = the total charge on the particle.

POLYATOMIC ION


ClO_3^-

?

?

ELEMENT	"CHARGE"
Chlorine	?
Oxygen	?
	?

REDOX

RULE SIX

Other atoms get the charge they prefer, as long as the sum of oxidation numbers for all atoms = the total charge on the particle.

ClO_3^-

+5

-2

ELEMENT	“CHARGE”
Chlorine	+5
Oxygen	(3 x -2) = -6
	-1

REDOX

$$\overset{\pm0}{N_2} + 3\,\overset{\pm0}{H_2} \rightarrow 2\,\overset{-3}{N}\overset{+1}{H_3}$$

Able to work out which elements are being oxidised and which are being reduced through oxidation states.

REDOX

$$\overset{\pm 0}{N_2} + 3\,\overset{\pm 0}{H_2} \rightarrow 2\,\overset{-3}{N}\overset{+1}{H_3}$$

Nitrogen & Hydrogen reactants are elemental so oxidation number = 0

REDOX

$$\overset{\pm 0}{N_2} + 3\,\overset{\pm 0}{H_2} \rightarrow 2\,\overset{-3}{N}\overset{+1}{H_3}$$

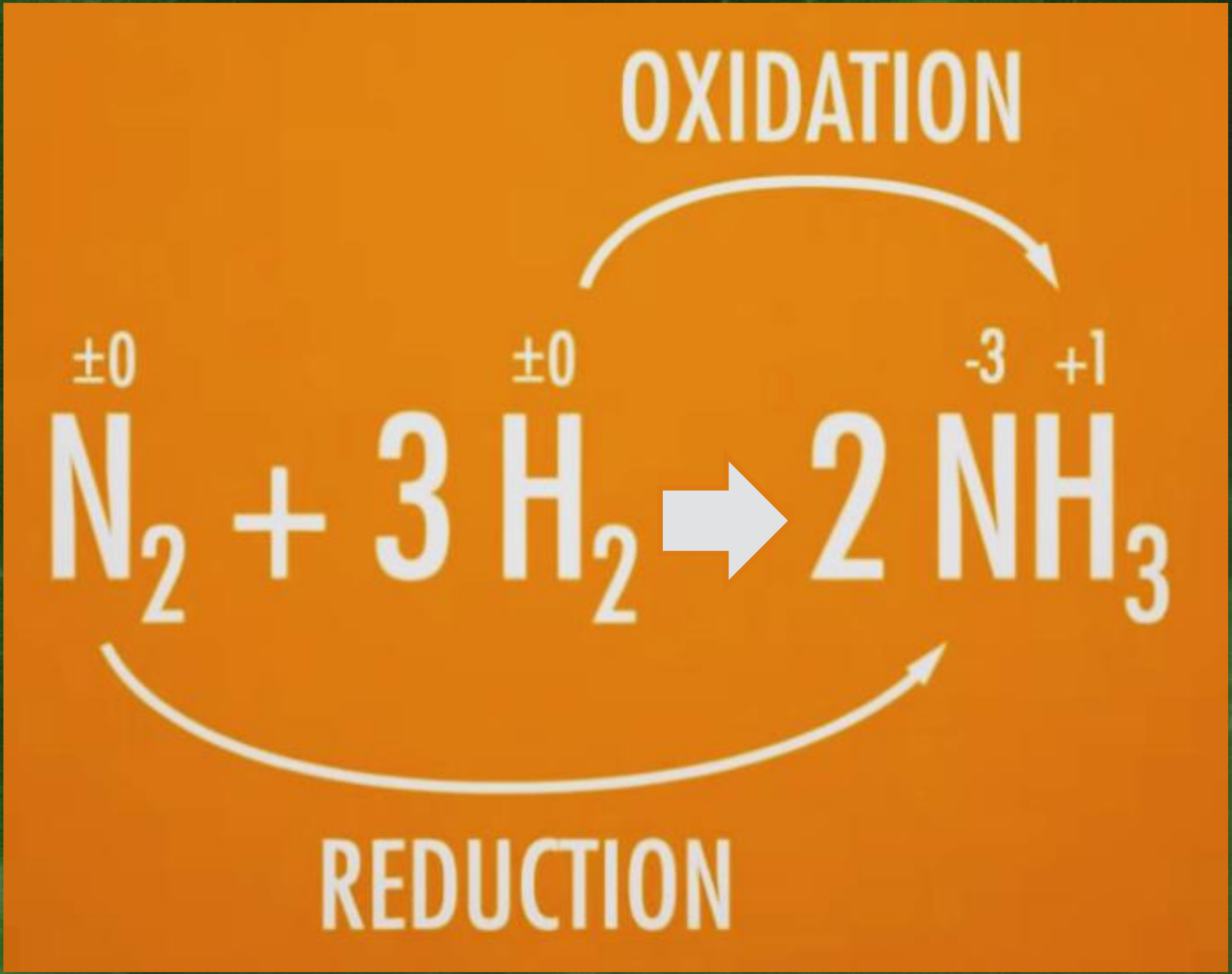
Hydrogen always has an oxidation number = +1

REDOX

$$\overset{\pm 0}{N_2} + 3\,\overset{\pm 0}{H_2} \rightarrow 2\,\overset{-3}{N}\overset{+1}{H_3}$$

Nitrogen will have -3 to balance charge on ammonia.

REDOX



$$\overset{\pm 0}{N}_2 + 3\,\overset{\pm 0}{H}_2 \rightarrow 2\,\overset{-3}{N}\overset{+1}{H}_3$$

Hydrogen
is oxidised
(loss of e^-)

Nitrogen is
reduced
(gain of e^-)

REDOX

$3H_2 \rightarrow 2NH_3 + 3e^-$

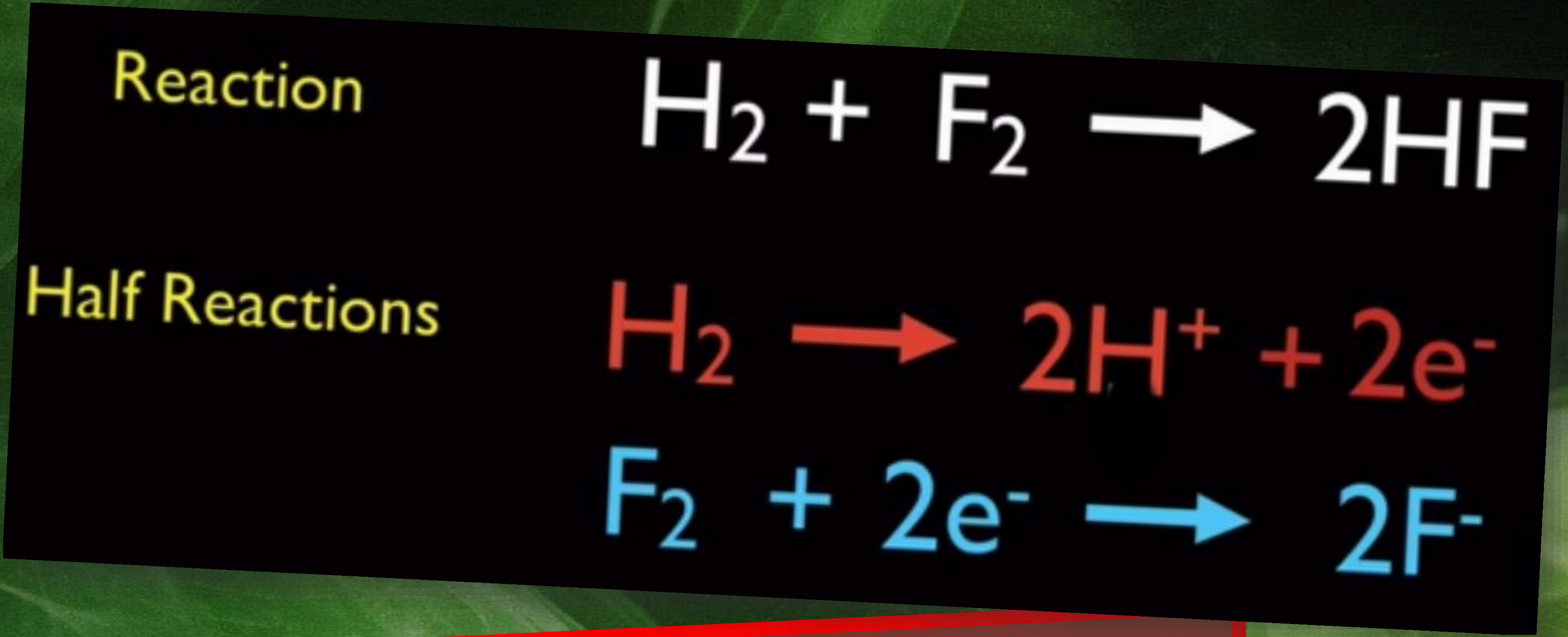
$N_2 + 3e^- \rightarrow 2NH_3$

$N_2 + 3H_2 \rightarrow 2NH_3$

Hydrogen is **oxidised** (loss of e^-)

Nitrogen is **reduced** (gain of e^-)

REDOX



Reaction $H_2 + F_2 \rightarrow 2HF$

Half Reactions $H_2 \rightarrow 2H^+ + 2e^-$

$F_2 + 2e^- \rightarrow 2F^-$

For some reactions to fully understand what is being oxidised and reduced we need to write half-reactions.

REDOX

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + 2H_2(g)$$

For **half-reactions** first work out the oxidation numbers.

REDOX

$$\overset{0}{\mathrm{Mg}}(s) + 2\overset{+1}{\mathrm{H}}\overset{-1}{\mathrm{Cl}}(aq) \longrightarrow \overset{+2}{\mathrm{Mg}}\overset{-1}{\mathrm{Cl}_2}(aq) + 2\overset{0}{\mathrm{H}_2}(g)$$

Mg and H_2 are elemental so oxidation number = 0

REDOX

$$\overset{0}{\mathrm{Mg}}(s) + 2\overset{+1}{\mathrm{H}}\overset{-1}{\mathrm{Cl}}(aq) \longrightarrow \overset{+2}{\mathrm{Mg}}\overset{-1}{\mathrm{Cl}_2}(aq) + 2\overset{0}{\mathrm{H}_2}(g)$$

H_2 always has an oxidation number = +1

REDOX

$$\overset{0}{\mathrm{Mg}}(s) + 2\overset{+1}{\mathrm{H}}\overset{-1}{\mathrm{Cl}}(aq) \longrightarrow \overset{+2}{\mathrm{Mg}}\overset{-1}{\mathrm{Cl}_2}(aq) + 2\overset{0}{\mathrm{H}_2}(g)$$

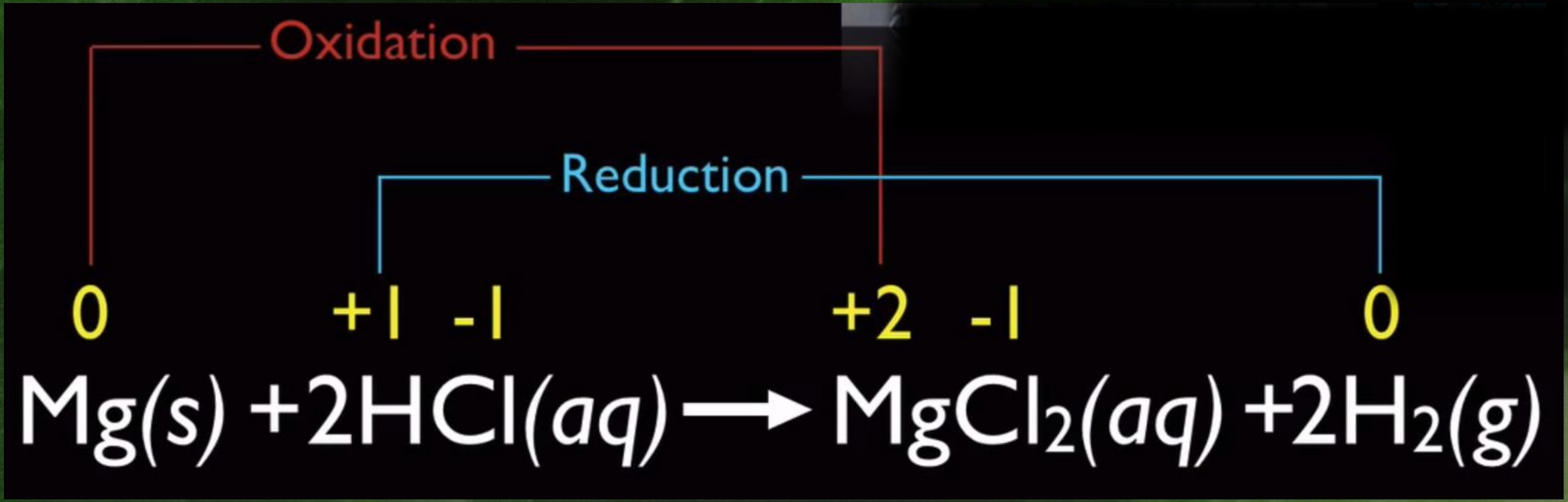
Cl will therefore be -1 in order to balance the charge.

REDOX

$$\overset{0}{Mg}(s) + 2\overset{+1}{H}\overset{-1}{Cl}(aq) \longrightarrow \overset{+2}{Mg}\overset{-1}{Cl_2}(aq) + 2\overset{0}{H_2}(g)$$

To balance the charge on the product side the (2x) -1 charge of the Cl will balance with the Mg; so Mg has oxidation number = +2

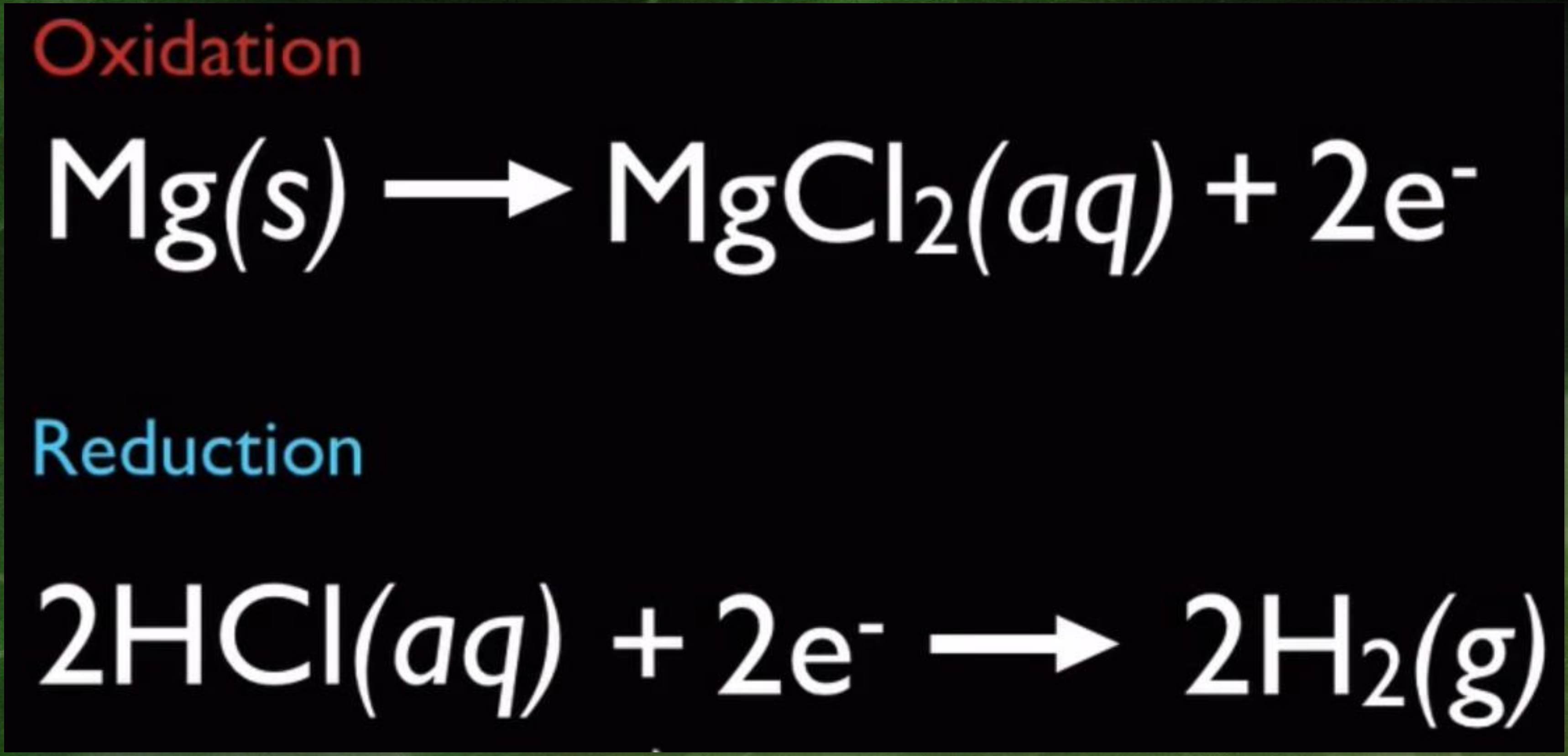
REDOX



$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + 2H_2(g)$$

The Mg is oxidised as it loses e^-
The Hydrogen is reduced as it gains e^-

REDOX



Oxidation

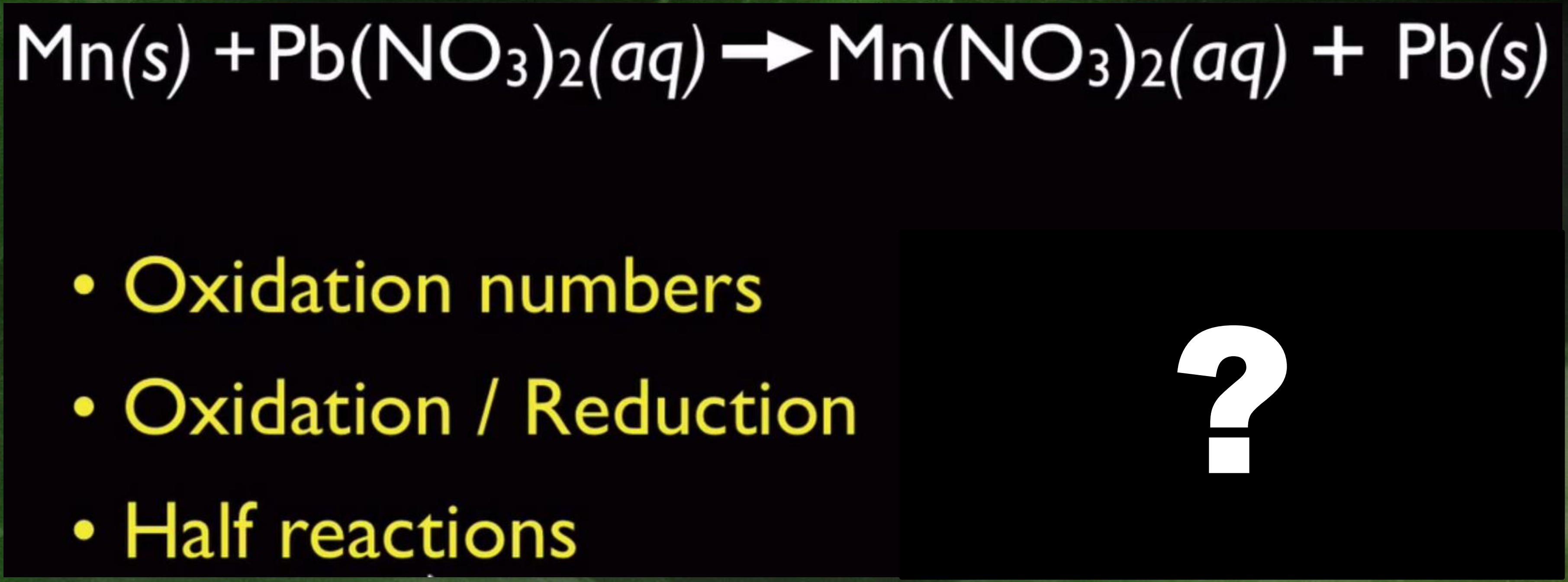
$$Mg(s) \longrightarrow MgCl_2(aq) + 2e^-$$

Reduction

$$2HCl(aq) + 2e^- \longrightarrow 2H_2(g)$$

Half-reactions include the electrons!

REDOX



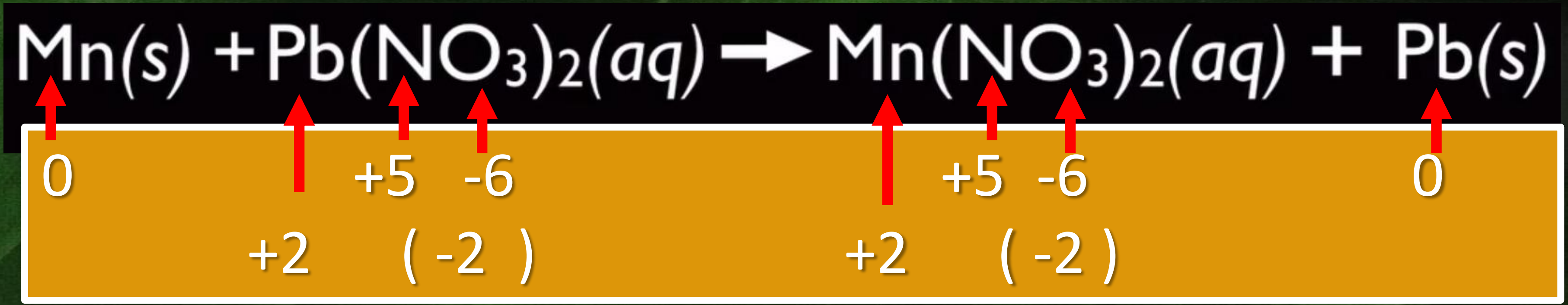
$Mn(s) + Pb(NO_3)_2(aq) \rightarrow Mn(NO_3)_2(aq) + Pb(s)$

- Oxidation numbers
- Oxidation / Reduction
- Half reactions

?

Work out the half-reactions!

REDOX



$Mn(s) + Pb(NO_3)_2(aq) \rightarrow Mn(NO_3)_2(aq) + Pb(s)$

0 +5 -6 +5 -6 0

+2 (-2) +2 (-2)

0
Mn → **+2**
$Mn(NO_3)_2$

+2
$Pb(NO_3)_2$ → **0**
Pb

REDOX

$$Mn(s) + Pb(NO_3)_2(aq) \rightarrow Mn(NO_3)_2(aq) + Pb(s)$$

Oxidation

$$Mn_{(s)} \rightarrow Mn(NO_3)_2\ _{(aq)} + 2e^-$$

(Mn is the Reducing agent)

Reduction

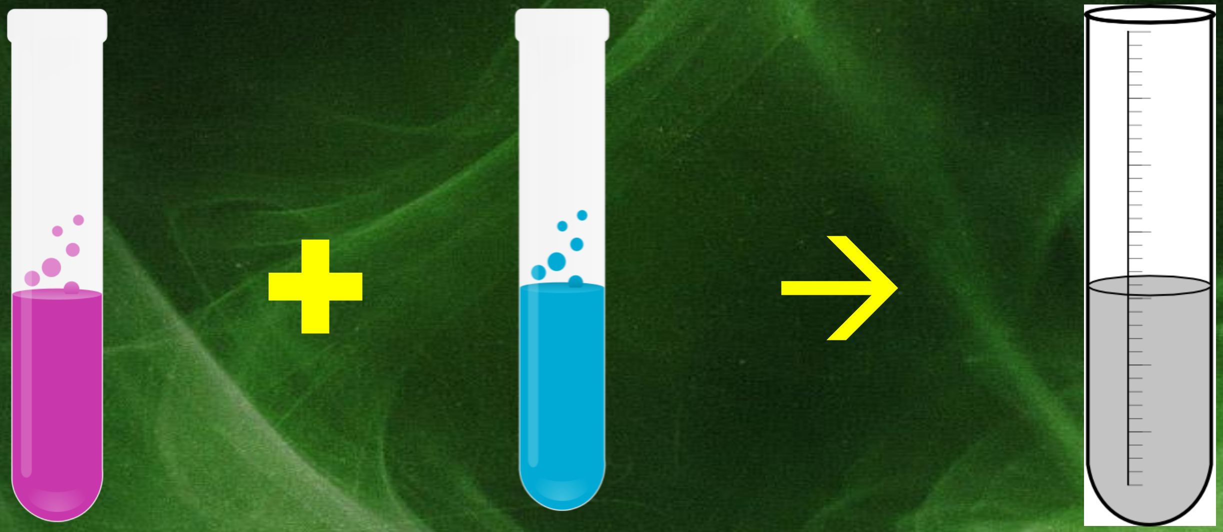
$$Pb(NO_3)_2\ _{(aq)} + 2e^- \rightarrow Pb\ _{(s)}$$

(Pb is the Oxidizing agent)

REDOX-REACTIONS

1. Balance all atoms on either side of equation except hydrogen & oxygen.
2. To balance oxygen add water to the oxygen deficient side.
3. To balance hydrogen add H^+ or to deficient side for acid conditions. For basic conditions add OH^- to BOTH sides, to neutralize H^+.
4. Balance the charge on either side by adding e^-. make sure charge is equal on both sides!

REDOX



Acidified MnO_4^-

Fe^{2+}

Colourless solution contains Mn^{2+} and Fe^{3+}

REDOX EQUATION?

REDOX

Fe^{2+} → $Fe^{3+} + e^-$

(Fe^{2+} is the Reducing agent)

MnO_4^- → Mn^{2+}

Purple → colourless

(MnO_4^- is the Oxidizing agent)

REDOX

Redox

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

$MnO_4^- \rightarrow Mn^{2+}$

1. Balance all atoms on either side of equation except hydrogen & oxygen.

REDOX

Redox

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

2. To balance oxygen add water to the oxygen deficient side.

REDOX

Redox

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$

3. To balance hydrogen add H^+ to deficient side.

REDOX

Redox

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

4. Balance the charge on either side by adding e^-. make sure charge is equal on both sides!

REDOX

Redox

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

To cancel out the electrons for the redox the iron equation must be multiplied by 5.

REDOX

Redox

$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

REDOX

Redox

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

CHARGES

1-	8+	10+	→	2+	15+	0
	17+		→		17+	

REDOX

For basic conditions…

$3H_2O + 5ClO_3^- + 3I_2 \rightarrow 5Cl^- + 6IO_3^- + 6H^+$

Add OH^- to both sides to neutralize H^+

REDOX

For basic conditions…

$3H_2O + ClO_3^- + I_2 \rightarrow Cl^- + 2IO_3^- + 6H^+$

$+ 6OH^-$

$+ 6OH^-$

REDOX

For basic conditions…

$$3H_2O + ClO_3^- + I_2 \rightarrow Cl^- + 2IO_3^- + \cancel{6H^+}$$

$+ 6OH^-$

$+ \cancel{6OH^-}$

$+ 6H_2O$

REDOX

For basic conditions…

$$\cancel{3H_2O} + ClO_3^- + I_2 \rightarrow Cl^- + 2IO_3^- + \cancel{6H^+}$$

$+ 6OH^-$

$+ \cancel{6OH^-}$

$+ \cancel{6}H_2O$

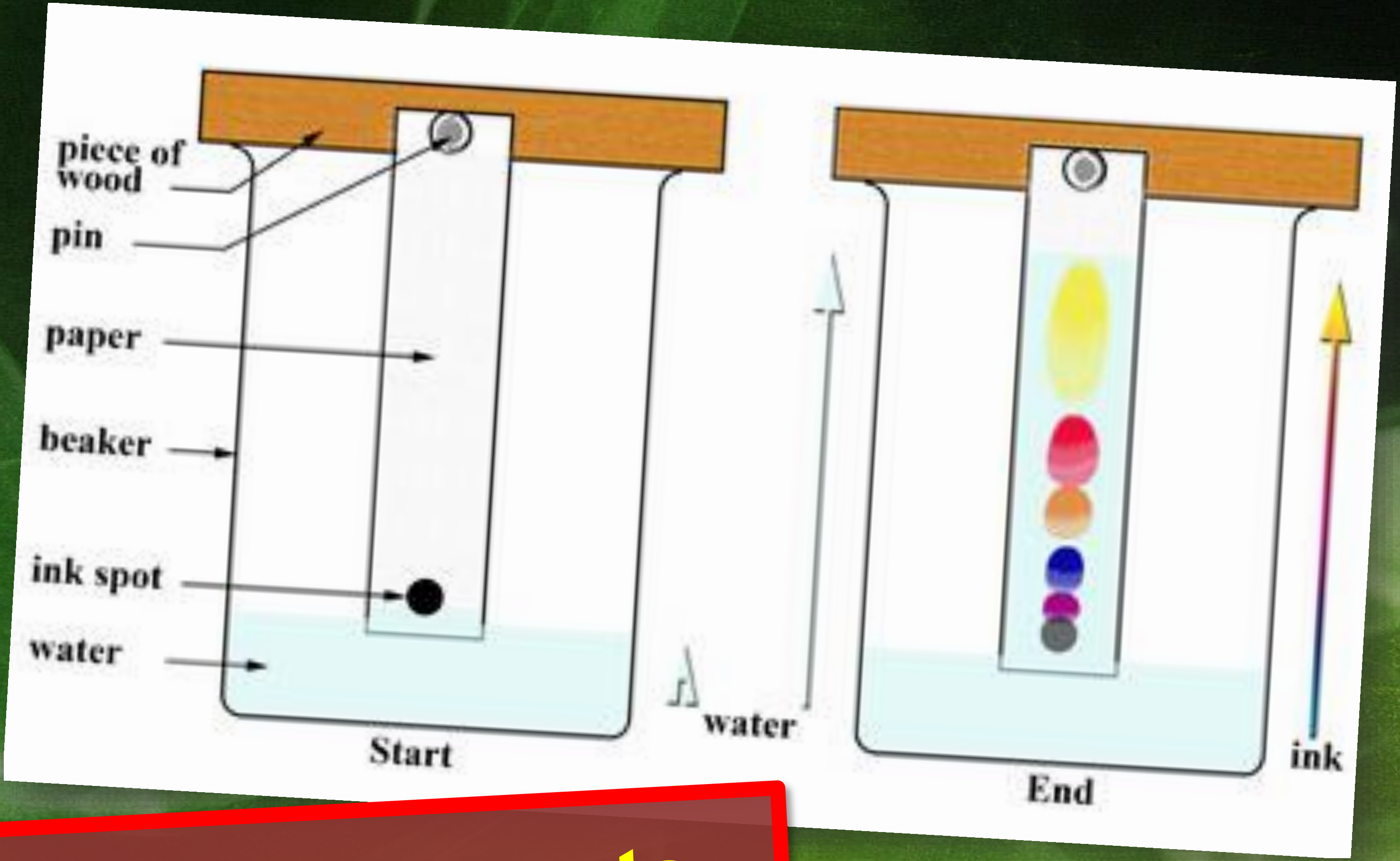
$+ 3H_2O$

REDOX

For basic conditions…

$$6OH^- + ClO_3^- + I_2 \rightarrow Cl^- + 2IO_3^- + 3H_2O$$

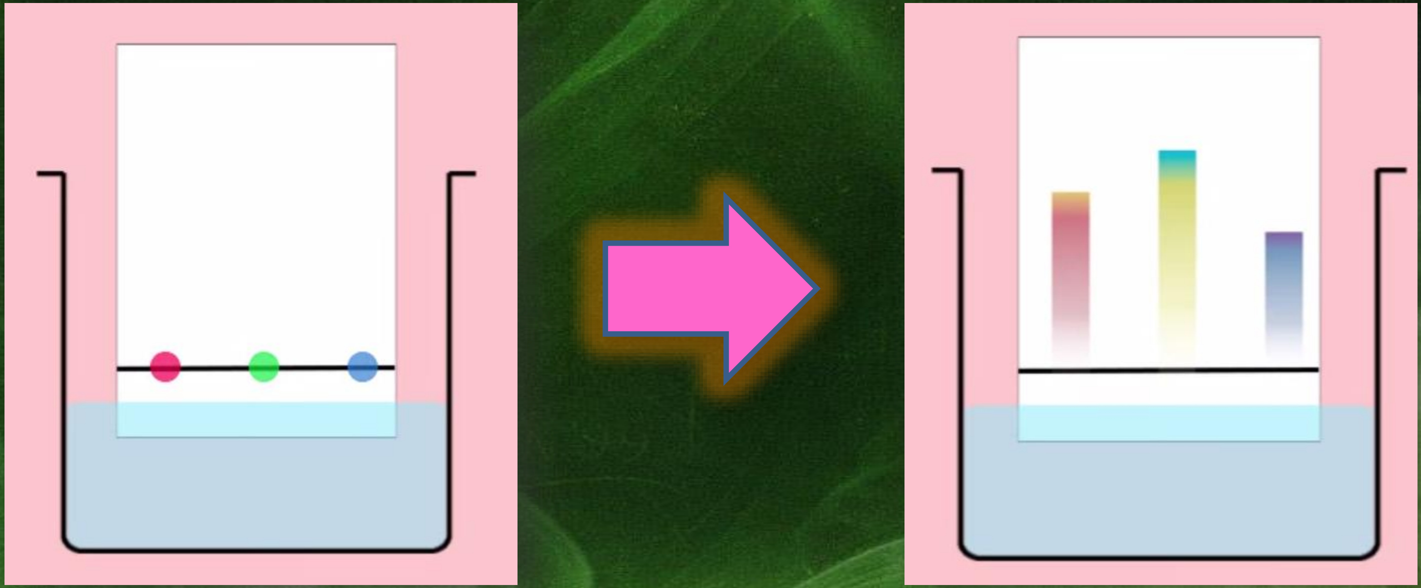
CHROMATOGRAPHY



Technique to separate components of a mixture

CHROMATOGRAPHY

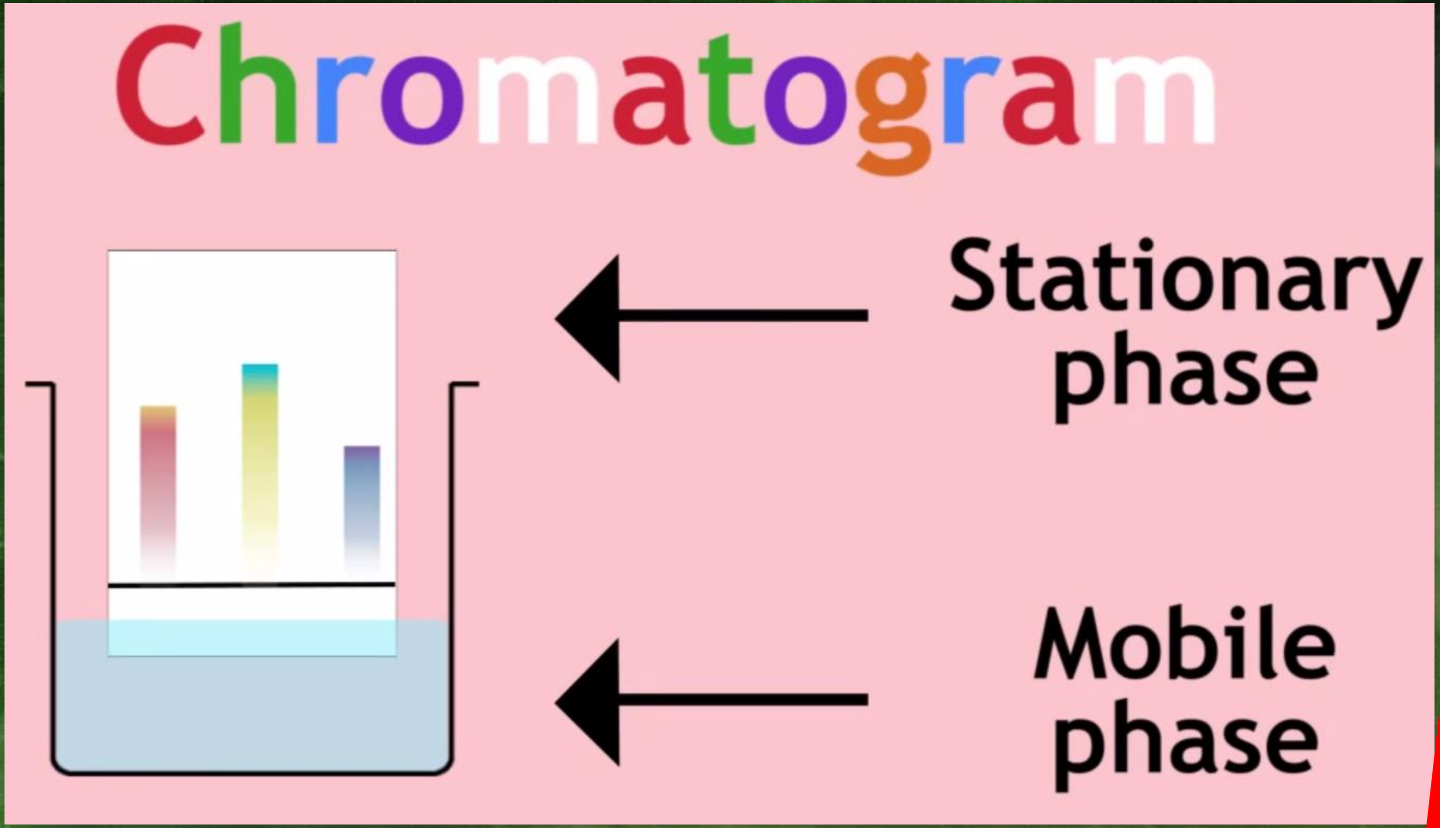
Paper chromatography



Reference spots are dipped in solvent and over time the solvent moves up paper and separating components based on solubility.

CHROMATOGRAPHY

Paper chromatography

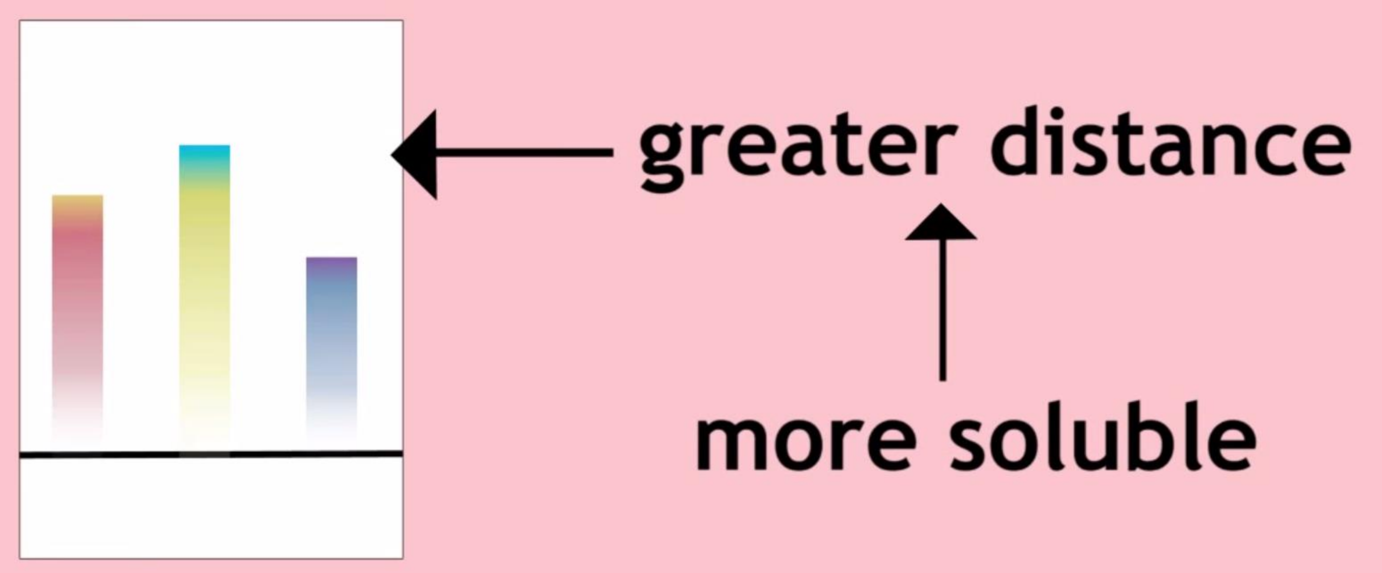


Paper is stationary phase: doesn’t move

Solvent is mobile phase: does move

CHROMATOGRAPHY

Paper chromatography



greater distance

more soluble

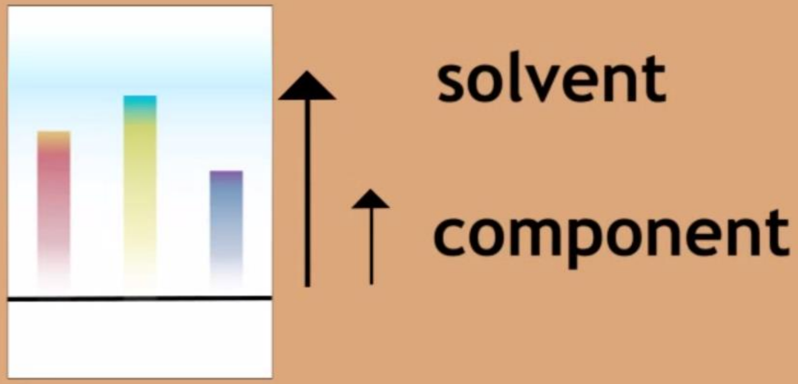
The more soluble the component the further it travels up the stationary phase.

CHROMATOGRAPHY

Paper chromatography

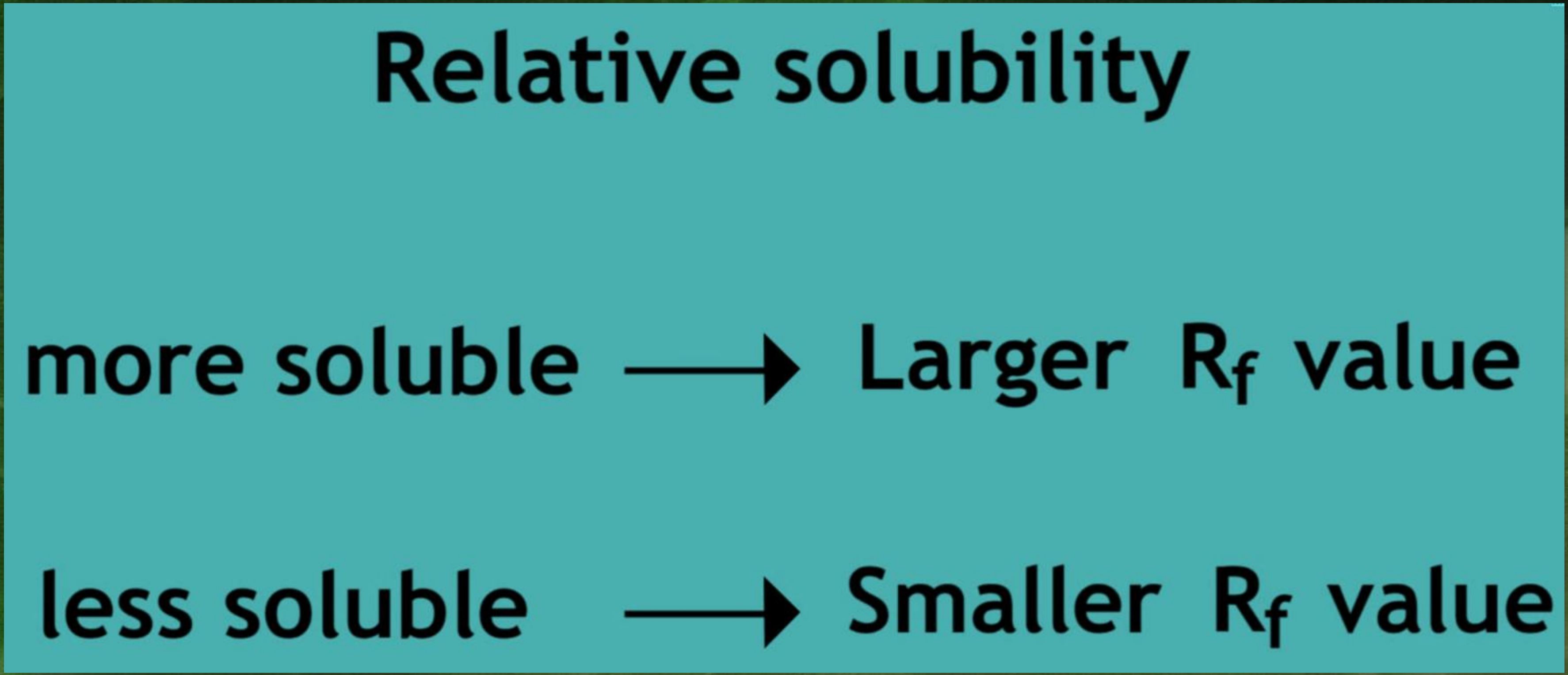
Retention factor (R_f)

$$R_f = \frac{\text{Distance travelled by a component}}{\text{Distance travelled by the solvent}}$$



CHROMATOGRAPHY

Paper chromatography

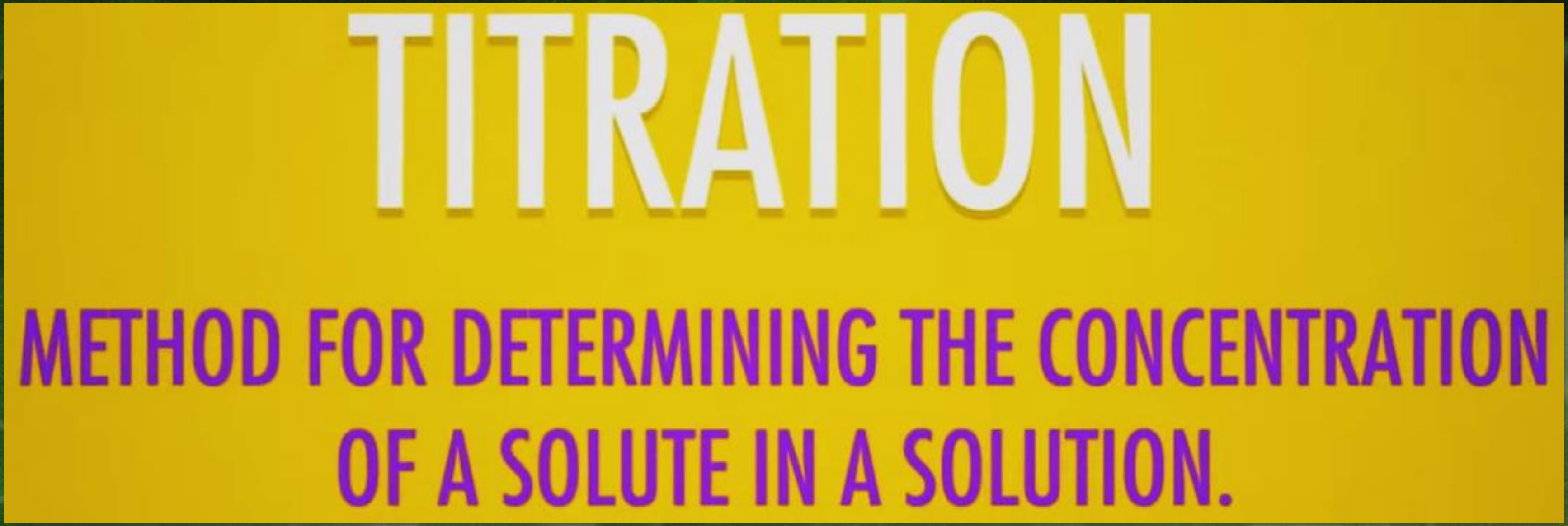


Relative solubility

more soluble → Larger R_f value

less soluble → Smaller R_f value

VOLUMETRIC ANALYSIS

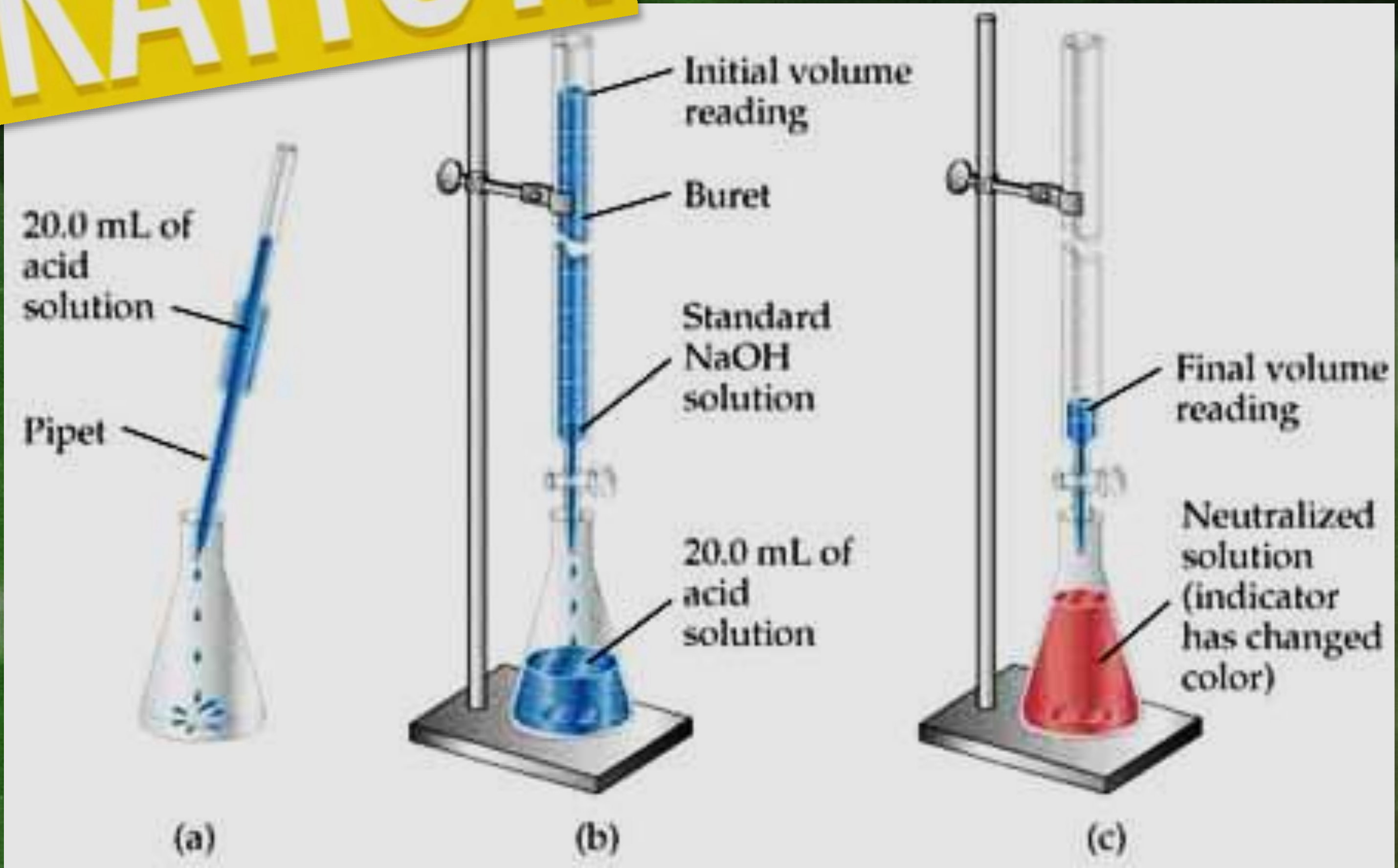


TITRATION

METHOD FOR DETERMINING THE CONCENTRATION OF A SOLUTE IN A SOLUTION.

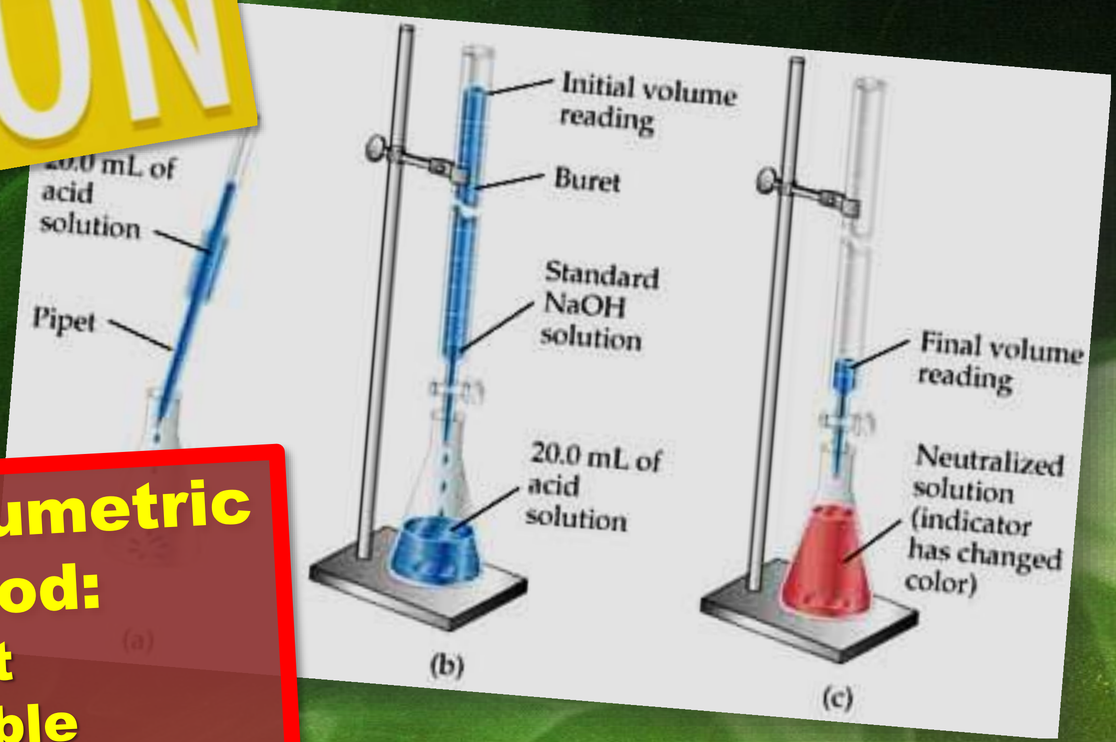
VOLUMETRIC ANALYSIS

TITRATION



VOLUMETRIC ANALYSIS

TITRATION



Advantages of volumetric titration method:

1. Easy to carry out
2. Versatile & reliable
3. Many reagents are cheap
4. Vast range of applications

VOLUMETRIC ANALYSIS

TITRATION

For a titration:

1. The reaction must proceed according to a definite chemical equation with **NO SIDE REACTIONS**.
2. Must be **Not reversible**.
3. There must be some method of detecting the equivalence point – **an indicator**.
4. The reaction should be **rapid**!