

Salt bridge

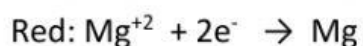
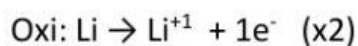
The glass tube connecting the two half cells is known as a salt bridge.

It is filled with an electrolyte and there is cotton wool on each end- so that the electrolyte solution stays in the tube.



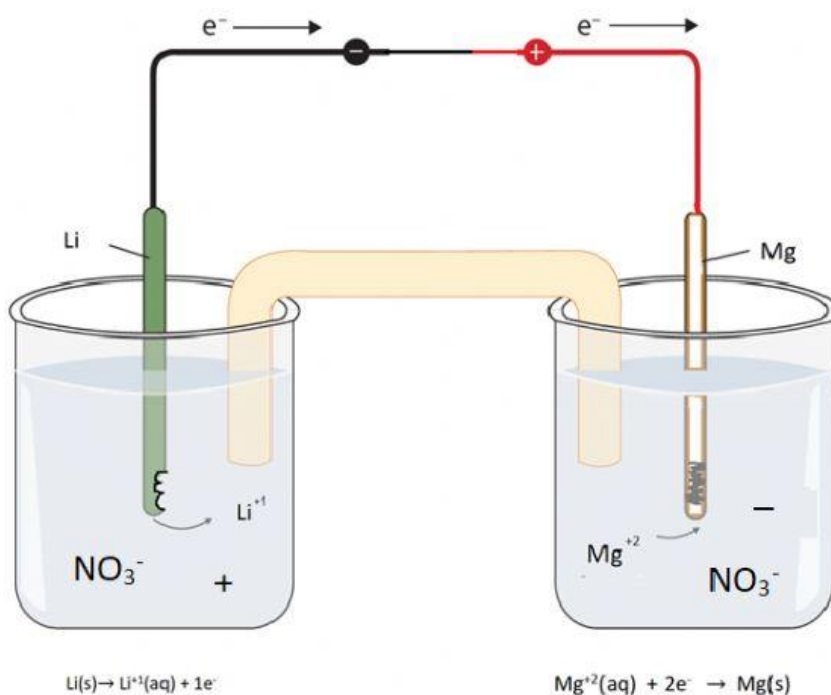
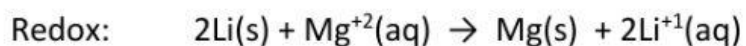
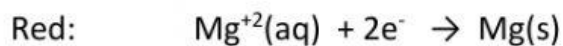
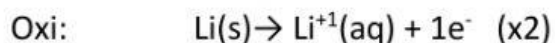
Let's examine this Li-Mg cell

Firstly- galvanic cells are spontaneous reactions and thus follow the **C-rule**



Now add the phases to the substances

If the element is neutral then it is in solid phase(s), and if it is charged, it is in aqueous solution (aq).



Li half-cell

Not let's examine what is actually happening in the cell

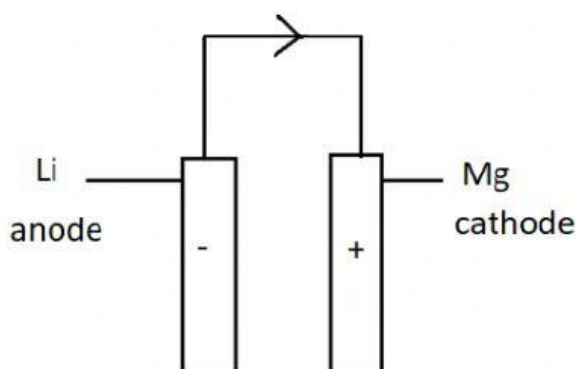
The Li(s) electrode is breaking up and becoming Li^{+1} ions that go into the solution and the electrons that it is giving off moves through the wire to the Mg half-cell. The Li electrode is corroding.

- The Li **solution** ends up becoming more and more **positively charged** (from the Li^{+1} ions that are constantly being given off into the solutions)
- AND the Li electrode becomes more negatively charged (from the electrons being given off every time a Li atom breaks up and collecting on the electrode.)

The Mg half-cell

The $\text{Mg}^{+2}(\text{aq})$ ions in the solution are attracted to the Mg electrode and gain electrons from the Mg electrode. The Mg^{+2} ions and the electrons join up and form Mg(s) ions, that plate onto the Mg electrode.

- The Mg solution ends up becoming more and more negatively charged (from the positive Mg^{+2} ions leaving it and plating onto the electrode and thus leaving negative spectator ions NO_3^- behind)
- AND the Mg electrode becomes more positively charged (from the electrons being taken away from the Mg electrode and joining up with the Mg^{+2} ions)



AN OX - Anode is where
oxidation takes place

RED CAT - REDuction takes
place at the CATode

What role does the salt bridge play

To complete the circuit and

To maintain electrical neutrality in the cell

It contains positive and negative ions.

If they ask you when you can use in the salt bridge- KNO_3 is always a good option.

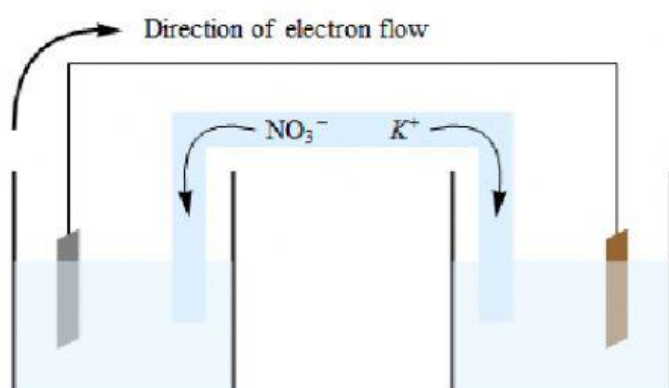
The reason is because you want to use something in the salt bridge that can't react with anything in the half-cells and form a precipitate (a solid) and clog the salt bridge.

You can learn all the rules again from grade 10- but it really isn't necessary. KNO_3 is always a good choice

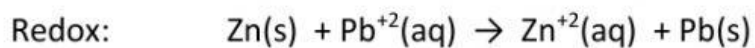
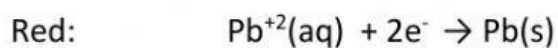
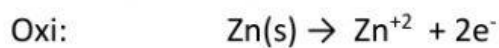
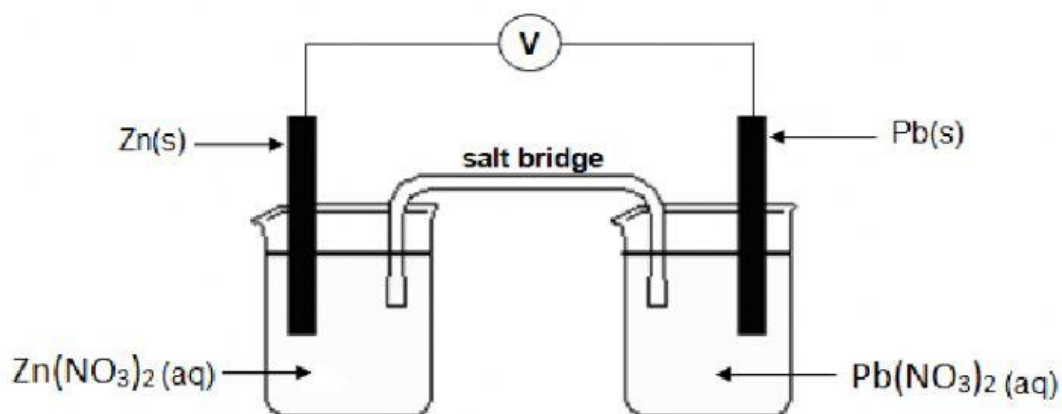
1. All common compounds of **Group I** and ammonium ions are soluble.
2. All **nitrates**, acetates, and chlorates are soluble.
3. All binary compounds of the halogens (other than F) with metals are soluble, except those of Ag, Hg(I), and Pb. Pb halides are soluble in hot water.)
4. All sulfates are soluble, except those of barium, strontium, calcium, lead, silver, and mercury (I). The latter three are slightly soluble.
5. Except for rule 1, carbonates, hydroxides, oxides, silicates, and phosphates are insoluble.
6. Sulfides are insoluble except for calcium, barium, strontium, magnesium, sodium, potassium, and ammonium.

Without the salt bridge, the **solution** in the anode compartment would become positively charged and the **solution** in the cathode compartment would become negatively charged, because of the charge imbalance, the electrode reaction would quickly come to a halt.

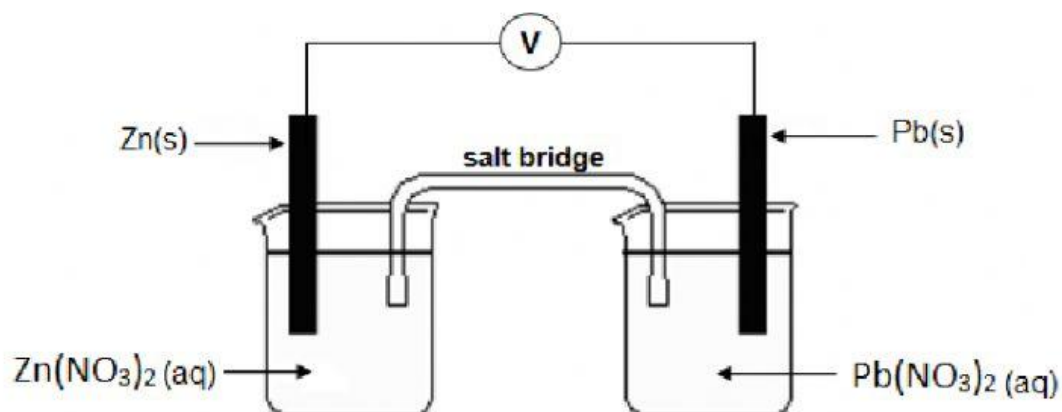
It helps to maintain the flow of electrons from the oxidation half-cell to a reduction half-cell, this completes the circuit



Example 2

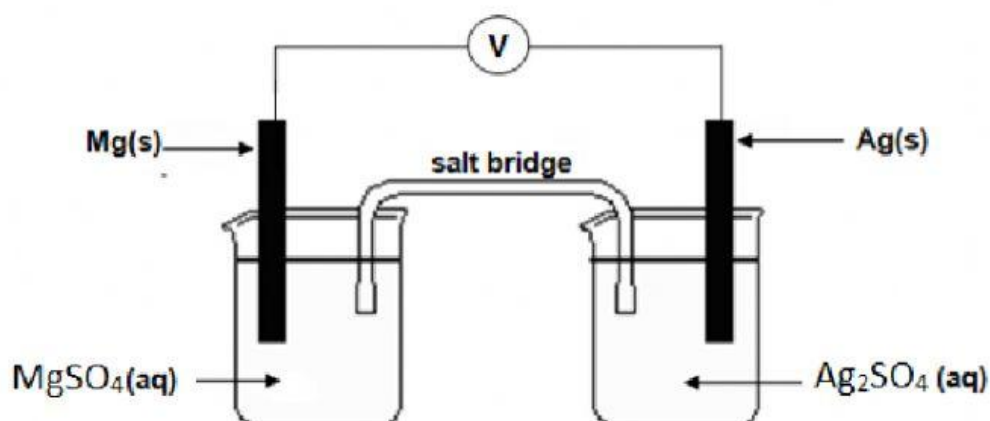


Fill in the missing words



Exercise:

1)



1.1) oxi: \rightarrow

1.2) red: \rightarrow

1.3) redox: \rightarrow

1.4) The anode:

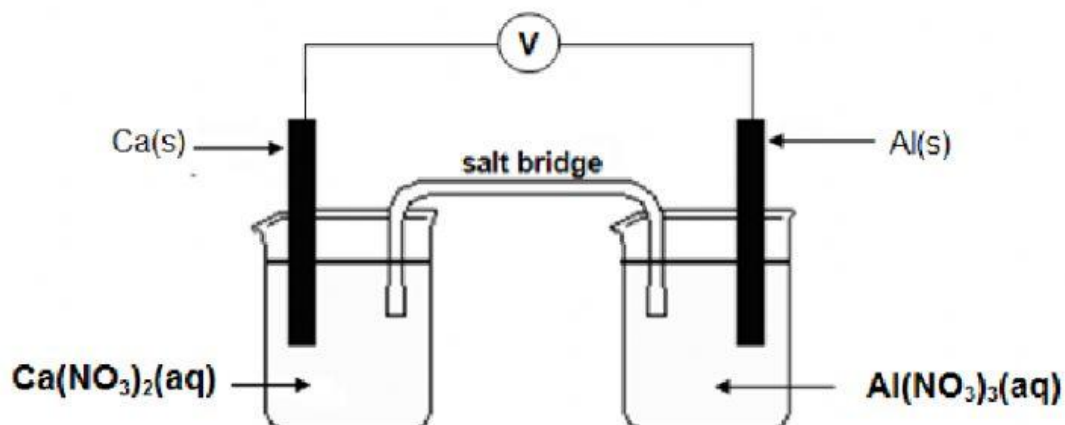
1.5) The cathode:

1.6) The Mg is charged

1.7) The Ag is charged

1.8) The spectator ion is: (write the name):

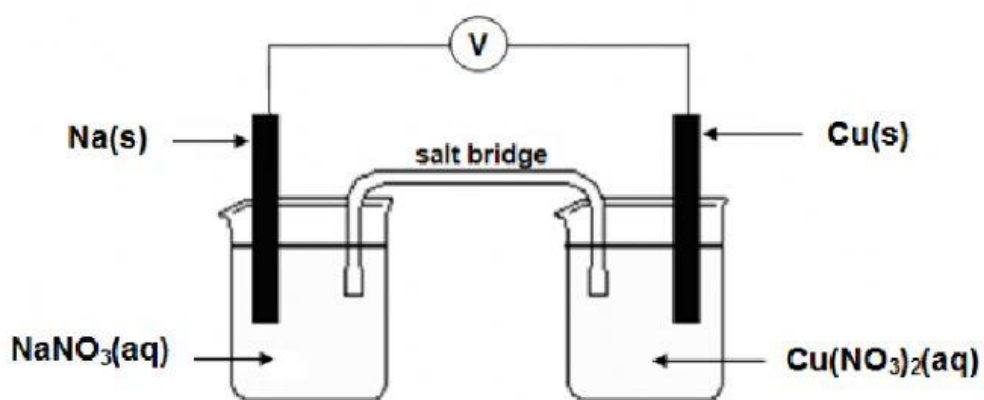
2)



2.1) oxi: \rightarrow

- 2.2) red: →
- 2.3) redox: →
- 2.4) The anode:
- 2.5) The cathode:
- 2.6) The Al is charged
- 2.7) The Ca is charged
- 2.8) The spectator ion is: (write the name):

3)



- 3.1) oxi: →
- 3.2) red: →
- 3.3) redox: →
- 3.4) The anode:
- 3.5) The cathode:
- 3.6) The Na is charged
- 3.7) The Cu is charged
- 3.8) The spectator ion is: (write the name):

Now let's start calculating the voltage that these cells would provide:

$$E_{\text{cell}}^{\ominus} = E_{\text{cathode}}^{\ominus} - E_{\text{anode}}^{\ominus}$$

or/of

$$E_{\text{cell}}^{\ominus} = E_{\text{reduction}}^{\ominus} - E_{\text{oxidation}}^{\ominus}$$

or/of

$$E_{\text{cell}}^{\ominus} = E_{\text{oxidising agent}}^{\ominus} - E_{\text{reducing agent}}^{\ominus}$$

$E_{\text{cell}}^{\ominus}$ is the voltage/potential difference the cell will create.

Eg) When Li and Mg reacts

$$E_{\text{cell}}^{\ominus} = E^{\ominus}_{\text{reduction}} - E^{\ominus}_{\text{oxidation}}$$

$$= -2,36 - (-3,05)$$

$$= 0,69 \text{ V}$$

Half-reactions		$E^{\ominus} \text{ (V)}$
$\text{Li}^+ + \text{e}^-$	$\rightleftharpoons \text{Li}$	-3,05
$\text{K}^+ + \text{e}^-$	$\rightleftharpoons \text{K}$	-2,93
$\text{Cs}^+ + \text{e}^-$	$\rightleftharpoons \text{Cs}$	-2,92
$\text{Ba}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Ba}$	-2,90
$\text{Sr}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Sr}$	-2,89
$\text{Ca}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Ca}$	-2,87
$\text{Na}^+ + \text{e}^-$	$\rightleftharpoons \text{Na}$	-2,71
$\text{Mg}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Mg}$	-2,36

Eg 2) Zn and Pb reacting

$\text{Zn}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Zn}$	-0,76
$\text{Cr}^{3+} + 3\text{e}^-$	$\rightleftharpoons \text{Cr}$	-0,74
$\text{Fe}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Fe}$	-0,44
$\text{Cr}^{3+} + \text{e}^-$	$\rightleftharpoons \text{Cr}^{2+}$	-0,41
$\text{Cd}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Cd}$	-0,40
$\text{Co}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Co}$	-0,28
$\text{Ni}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Ni}$	-0,27
$\text{Sn}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Sn}$	-0,14
$\text{Pb}^{2+} + 2\text{e}^-$	$\rightleftharpoons \text{Pb}$	-0,13

$$E_{\text{cell}}^{\ominus} = E^{\ominus}_{\text{reduction}} - E^{\ominus}_{\text{oxidation}}$$

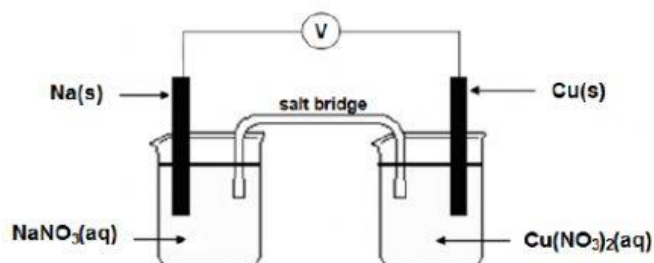
$$= -0,13 - (-0,76)$$

$$= 0,63 \text{ V}$$

The fact that the $E_{\text{cell}}^{\ominus}$ values for these turn out **positive** is another indication that the reaction is **spontaneous**.

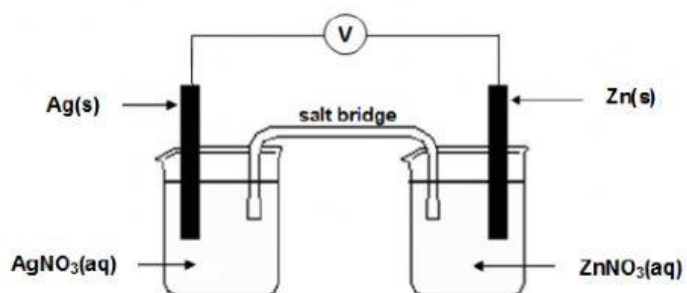
Exercise: (use a comma, not a dot for decimals)

1) Calculate the $E^{\ominus}_{\text{cell}}$ /potential difference values for the following cells



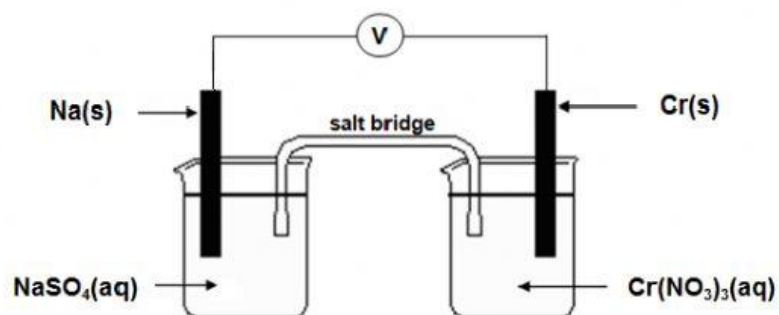
$E^{\ominus}_{\text{cell}} =$

2)



$E^{\ominus}_{\text{cell}} =$

3)



$E^{\ominus}_{\text{cell}} =$

Standard conditions

These potential difference values will only be true if the cell is functioning under standard conditions

There are:

At 25°C

Electrolyte concentration must be 1 mol.dm⁻³

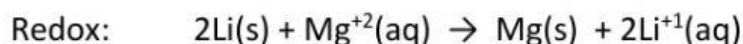
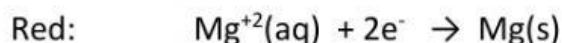
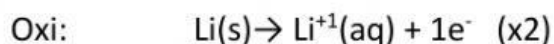
If there is a gas involved the pressure of the gas must be 1 atmosphere (101,325 kPa)

If these conditions are not met, then the calculated potential difference on the table will not be accurate.

Cell notation/Daniel cell notation

Cell notation is another way of writing a redox reaction

Eg 1)



Cell notation: $\text{Li(s)}/\text{Li}^{\text{+}}(\text{aq})(1 \text{ mol.dm}^{-3})// \text{Mg}^{\text{+2}}(\text{aq})(1 \text{ mol.dm}^{-3})/\text{Mg(s)}$

The /
represents a
phase
boundary
(separation
between the
different
phases)

Always write the one
that oxidises first.
(in the order it was
written in the half
reaction)

The //
represents the
salt bridge

Always write the one
that reduces second.
(in the order it was
written in the half
reaction)