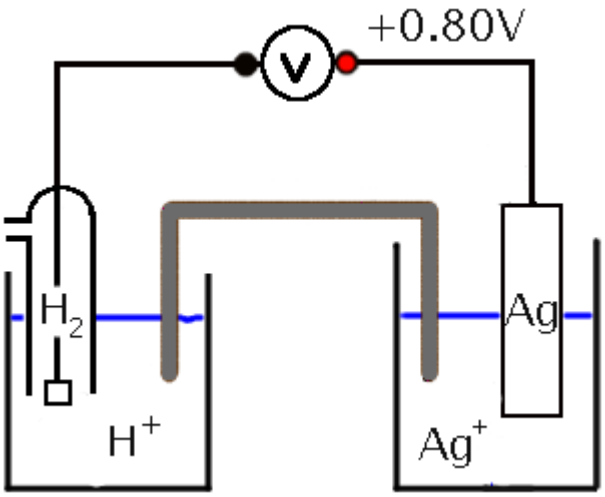
The Electrochemical Series

using standard electrode potentials

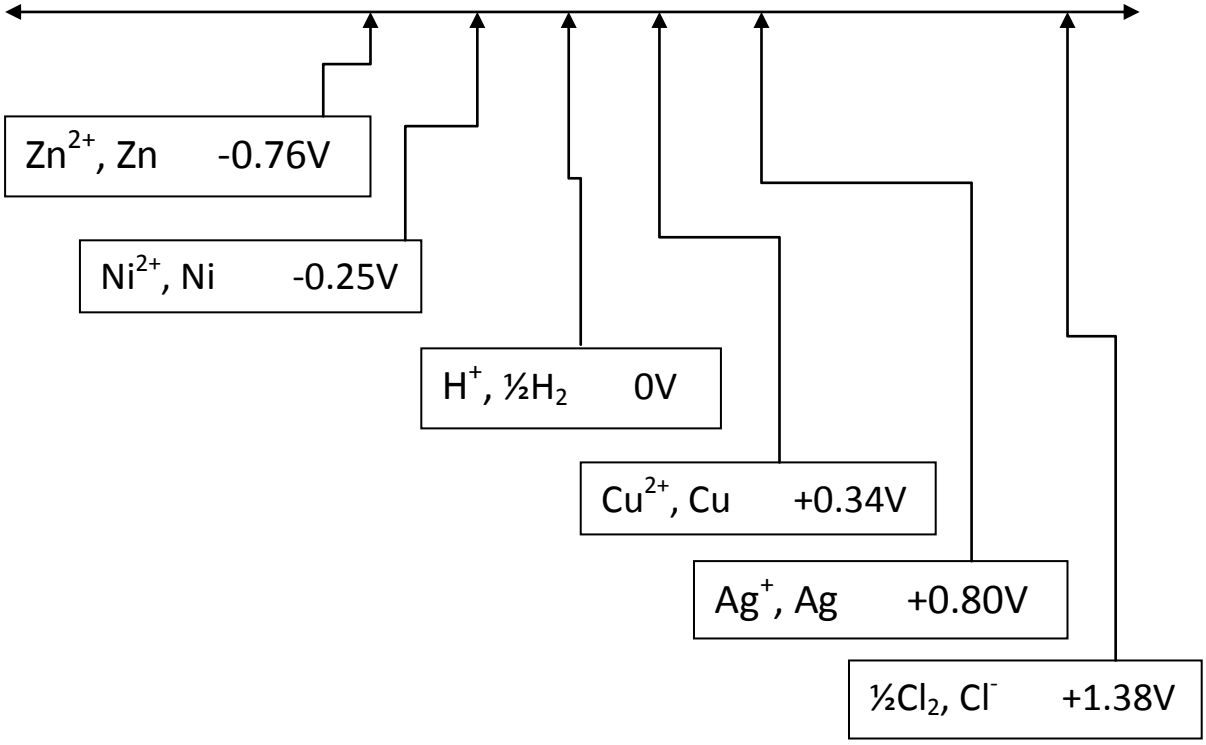
1. The Electrochemical Series
    a. Right-hand electrode relative to hydrogen
    b. REDOX behaviour
2. Using Standard Electrode Potentials
    a. Proving $E^{o}_{cell} = E^{o}_{(right\ hand\ electrode)} - E^{o}_{(left\ hand\ electrode)}$
    b. Identifying reaction for which E^{o}_{cell} is +ve

Right-hand electrode relative to hydrogen

Standard conditions (1atm, 1moldm^{-3}, standard states, 298K)



tends to lose electrons — REDOX behaviour — tends to gain electrons



More commonly written in a table of **reduction** half-equations

		Electrode				E^o / V	
-ve	reverse (←)	Zn^{2+}	+ $2e^-$	⇌	Zn	−0.76	*tends to lose electrons*
		Ni^{2+}	+ $2e^-$	⇌	Ni	−0.25	
		H^+	+ e^-	⇌	½ H_2	0	
		Cu^{2+}	+ $2e^-$	⇌	Cu	+0.34	
		Ag^+	+ e^-	⇌	Ag	+0.80	
		½ Br_2	+ e^-	⇌	Br^-	+1.09	*tends to gain electrons*
+ve	forward (→)	½ Cl_2	+ e^-	⇌	Cl^-	+1.38	

SIGN of E^o indicates bias of equilibrium

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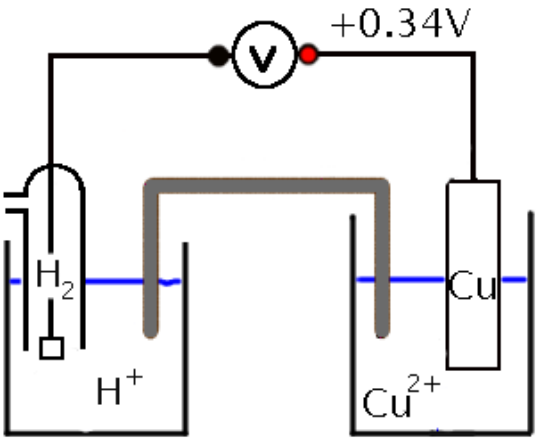
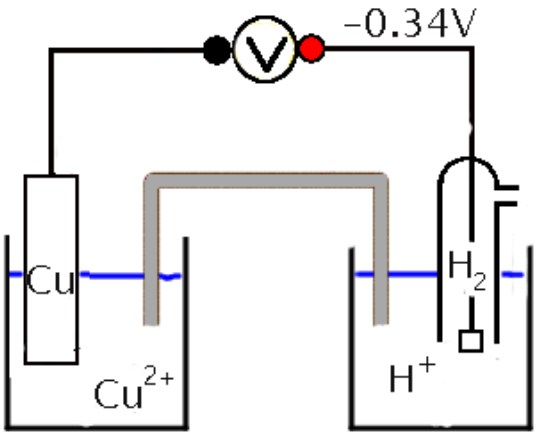
Proving $E^o_{cell} = E^o_{(right\ hand\ electrode)} - E^o_{(left\ hand\ electrode)}$

$Cu, Cu^{2+} = +0.34V$

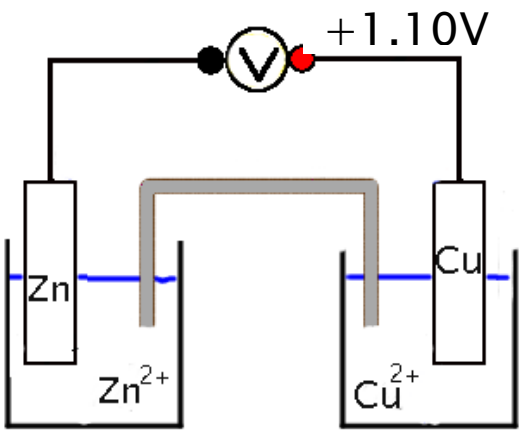


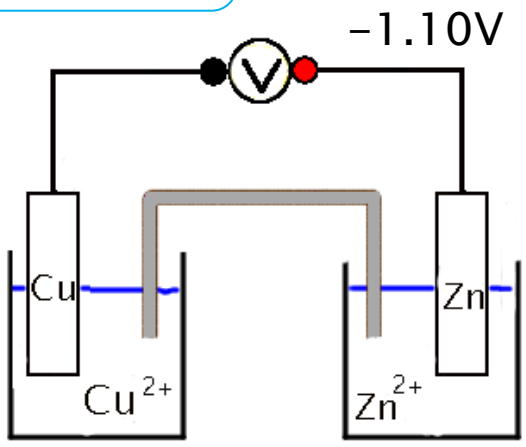


∴ if an electrode is on the L.H.S., its e.m.f. is reversed

$Cu, Cu^{2+} = +0.34V$

$Zn, Zn^{2+} = -0.76V$





e.m.f. = +0.34 + 0.76 = +1.10V

e.m.f. = −0.76 − 0.34 = −1.10V

∴ $E^o_{cell} = E^o_{(right\ hand\ electrode)} - E^o_{(left\ hand\ electrode)}$

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Identifying reaction for which E^{o}_{cell} is +ve *Anticlockwise rule*

positive electrode goes in the **forward** direction

this is **reduction** and goes on the **right** of the cell

e.g. nickel and copper

$\leftarrow$

$Ni^{2+} + 2e^- \rightleftharpoons Ni \quad -0.25$

$Cu^{2+} + 2e^- \rightleftharpoons Cu \quad +0.34$

$\rightarrow$

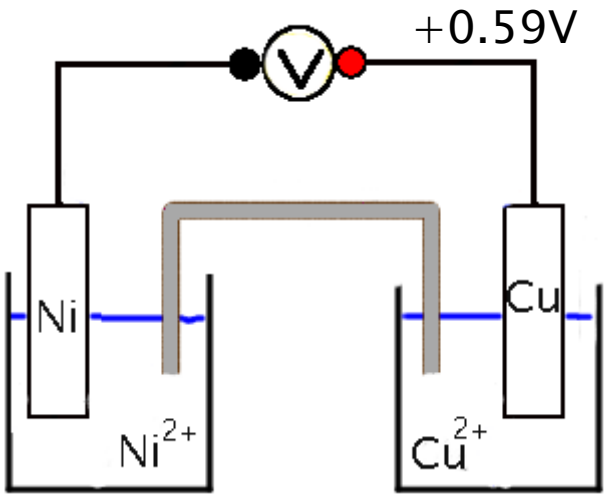
cell e.m.f. = R.H.S. – L.H.S.

= +0.34 + 0.25

= +0.59V

Electrode				E^{o}/ V
Zn^{2+}	+ $2e^-$	$\rightleftharpoons$	Zn	−0.76
Ni^{2+}	+ $2e^-$	$\rightleftharpoons$	Ni	−0.25
H^+	+ e^-	$\rightleftharpoons$	½ H_2	0
Cu^{2+}	+ $2e^-$	$\rightleftharpoons$	Cu	+0.34
Ag^+	+ e^-	$\rightleftharpoons$	Ag	+0.80
½ Br_2	+ e^-	$\rightleftharpoons$	Br^-	+1.09
½ Cl_2	+ e^-	$\rightleftharpoons$	Cl^-	+1.38

+ve value indicates that the feasible reaction is $Cu^{2+} \rightarrow Cu$ and $Ni \rightarrow Ni^{2+}$



Identifying reaction for which E^{o}_{cell} is +ve *Anticlockwise rule*

positive electrode goes in the **forward** direction

this is **reduction** and goes on the **right** of the cell

e.g. chlorine and bromine

	←			
½ Br_2	+ e^-	⇌	Br^-	+1.09
½ Cl_2	+ e^-	⇌	Cl^-	+1.38
	→			

Electrode				E^o/ V
½ Br_2	+ e^-	⇌	Br^-	+1.09
½ Cl_2	+ e^-	⇌	Cl^-	+1.38

cell e.m.f. = R.H.S. – L.H.S.

= +1.38 – 1.09

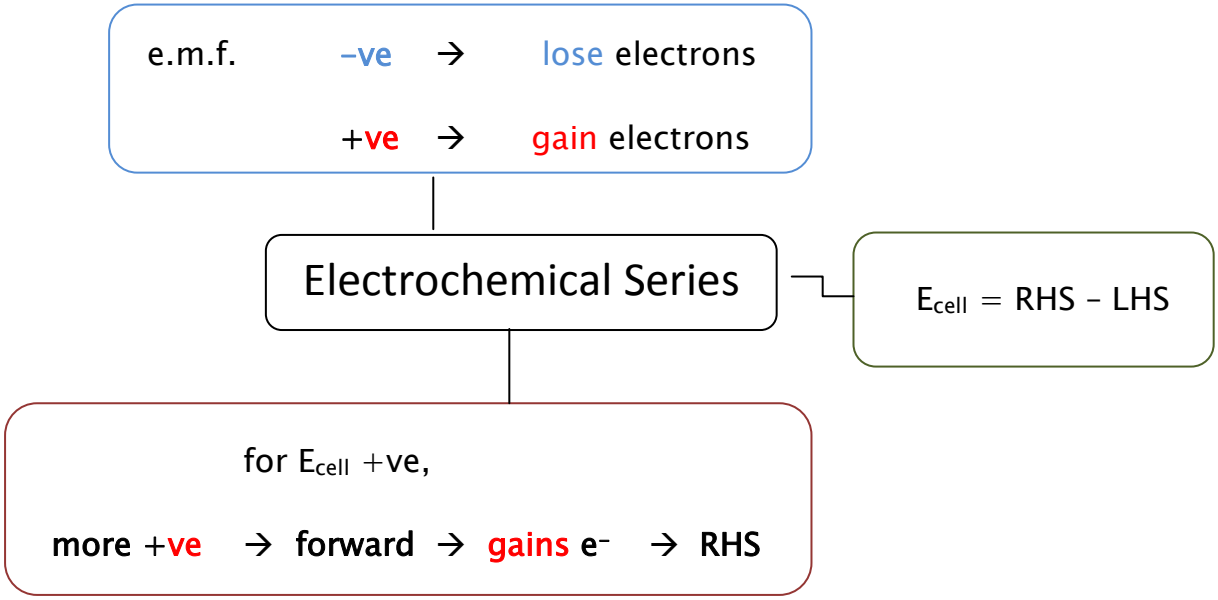
= +0.29V

The feasible, spontaneous reactions are

½ Cl_2 + e^- → Cl^-

Br^- → ½ Br_2 + e^-

½ Cl_2 + Br^- → Cl^- + ½ Br_2



References

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