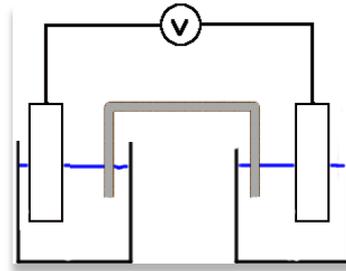


Standard Electrode Potentials

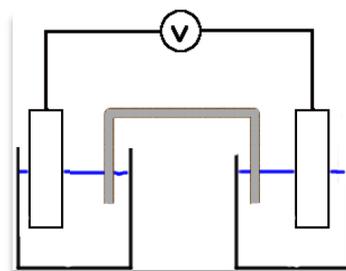
chemistry and conventions



1. Introduction
 - a. REDOX, half-equations and voltage
 - b. Why do we want to measure electrode potentials?
2. Describing the Electrochemical Cell
 - a. Technical terms
 - i. Electrode
 - ii. Half-cell
 - iii. Salt bridge
 - iv. Electromotive force (e.m.f.), potential difference, voltage
 - v. Electrode potential
3. Chemical Causes and Conventions
 - a. Cause of the electromotive force (e.m.f.)
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Standard Electrode Potentials

chemistry and conventions



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REDOX, half-equations and voltage

Oxidation Is Loss Reduction Is Gain

REDOX

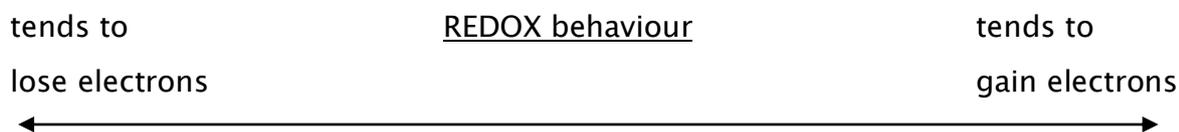
half-equations

Na	\rightarrow	$\text{Na}^+ + \text{e}^-$
$\frac{1}{2} \text{Cl}_2 + \text{e}^-$	\rightarrow	Cl^-
$\text{Na} + \frac{1}{2} \text{Cl}_2 + \text{e}^-$		
		$\rightarrow \text{Na}^+ + \text{e}^- + \text{Cl}^-$

voltage

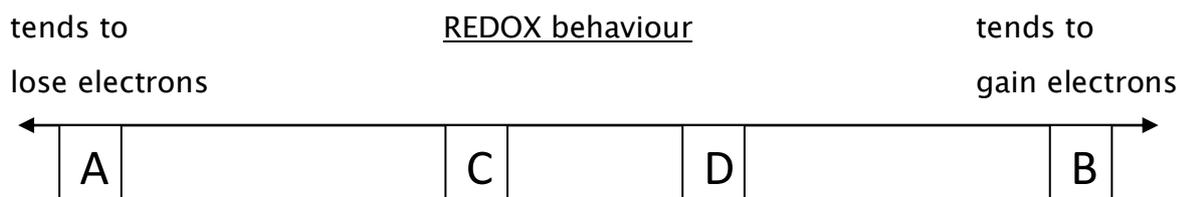
A diagram showing a battery labeled "1.5V" with a red '+' terminal and a black '-' terminal. A voltmeter (circle with 'V') is connected in a circuit with the battery. The voltmeter shows a reading of "+1.5V".

Why do we want to measure electrode potentials?



- quantitative reactivity series
- use to predict reactions
- make batteries

Where do you think the following would be on this reactivity series?



mild oxidising agent, such as iodine

reactive metal, such as potassium

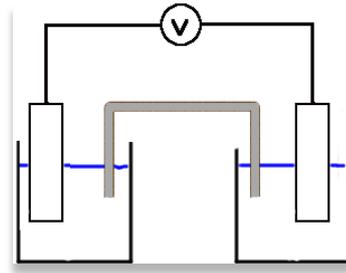
transition metal, such as iron

reactive non-metal, such as fluorine

answers (starting from top): D, A, C, B

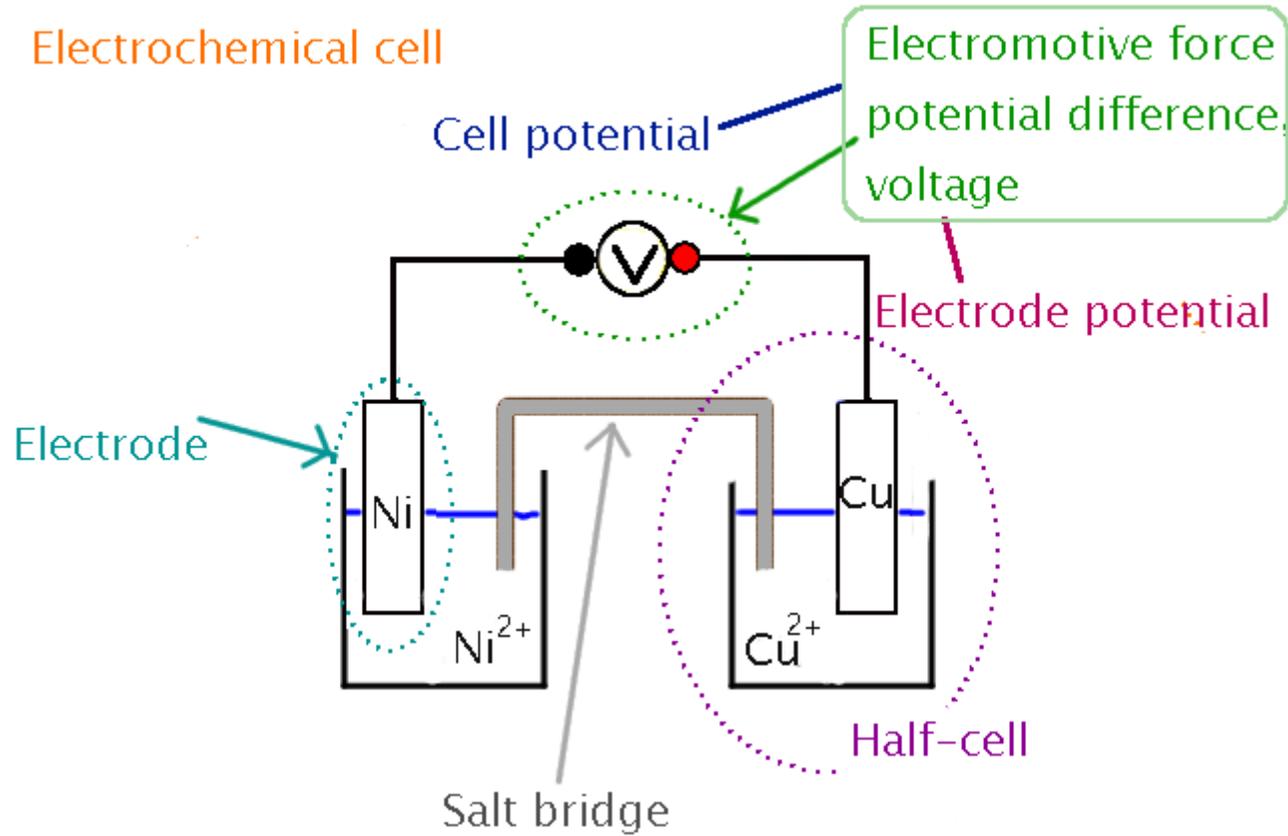
Standard Electrode Potentials

chemistry and conventions



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What technical terms do we need to know?



Match the term to its description
→

Salt Bridge

Electrode potential

Half-cell

Cell potential

Electromotive force (e.m.f.)

e.m.f. caused by joining two half-cells together

prevents build up of charge in the solution

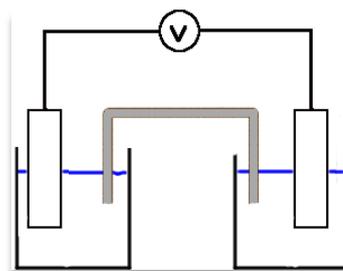
half of an electrochemical cell

potential energy due to build up of charge

e.m.f. caused by a half-cell relative to the hydrogen half-cell

Standard Electrode Potentials

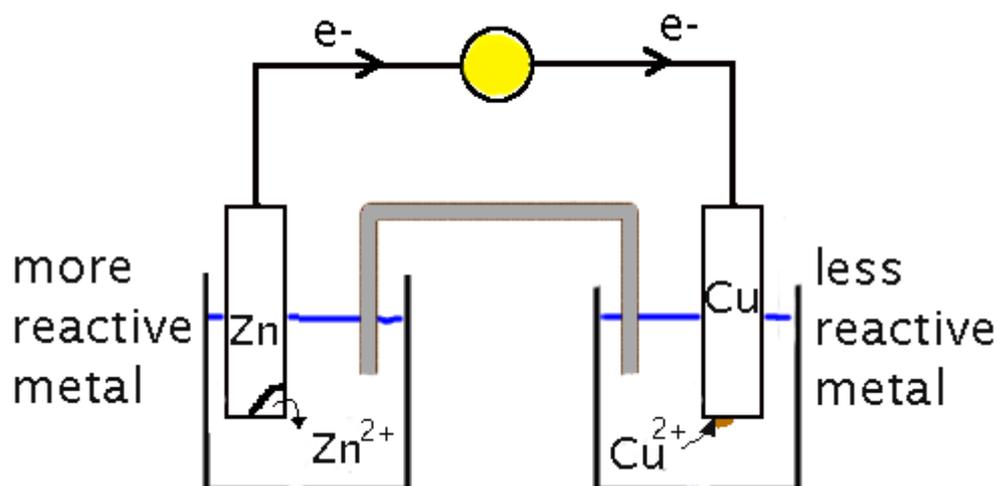
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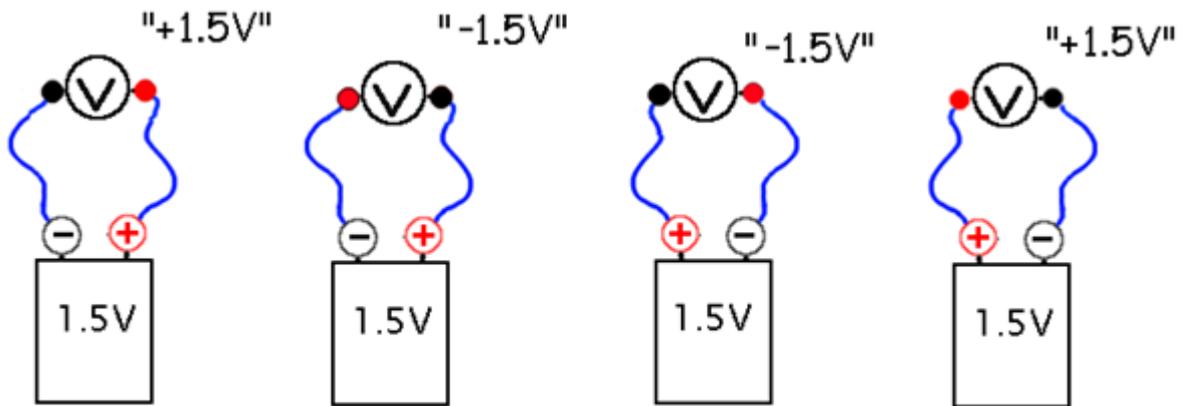
Cause of the electromotive force (e.m.f.)

Why does linking two half-cells cause a voltage?



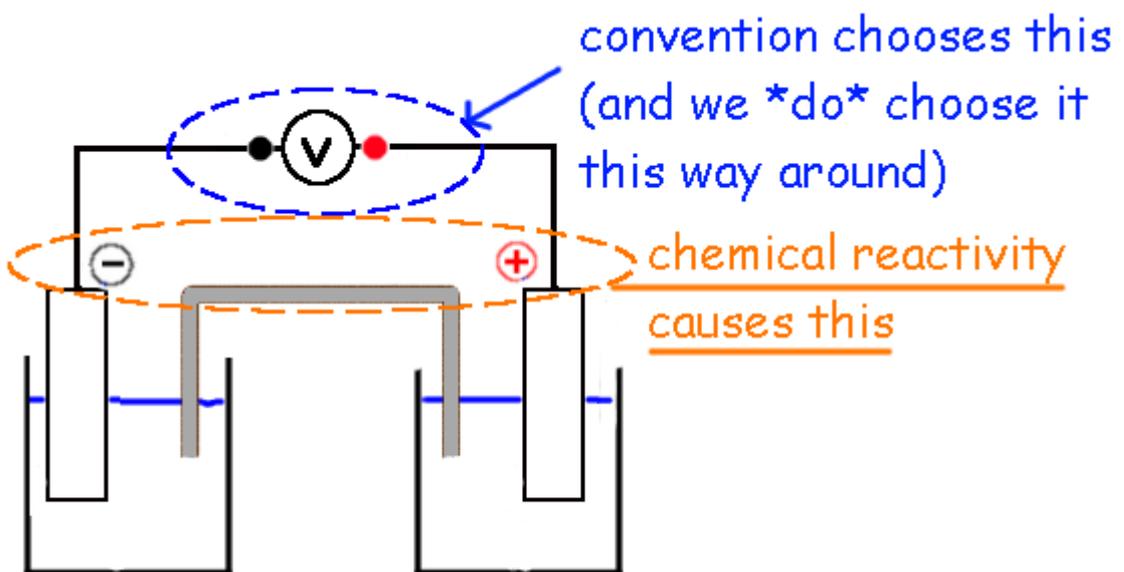
Convention of the voltmeter

Why does it matter which way around I plug in the voltmeter?

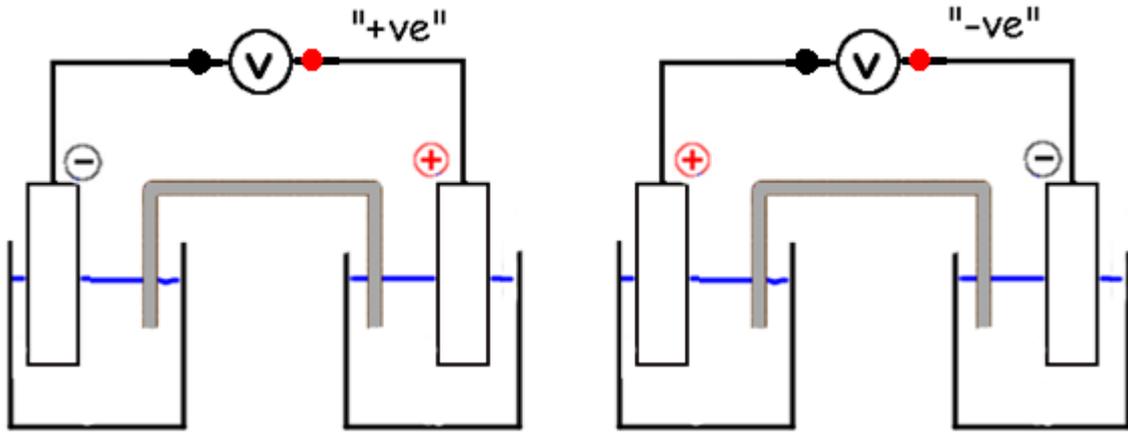


Convention of the voltmeter

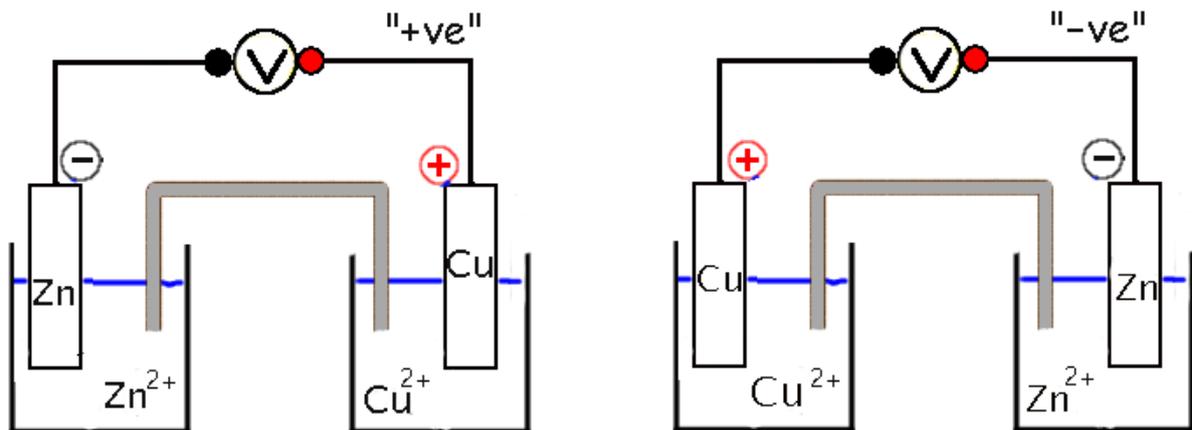
Why does it matter which way around I plug in the voltmeter?

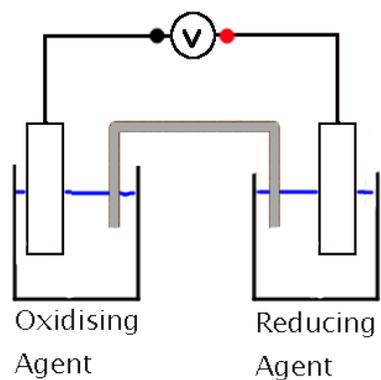
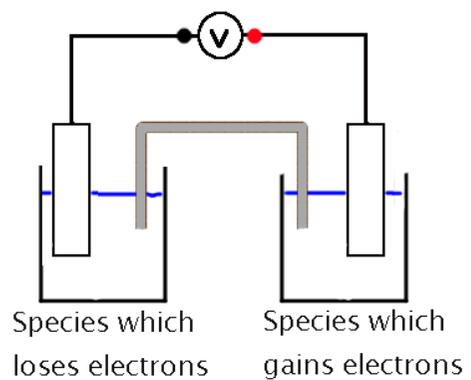


What does the sign of the electrochemical cell tell me about the REDOX behaviour of the half-cells?



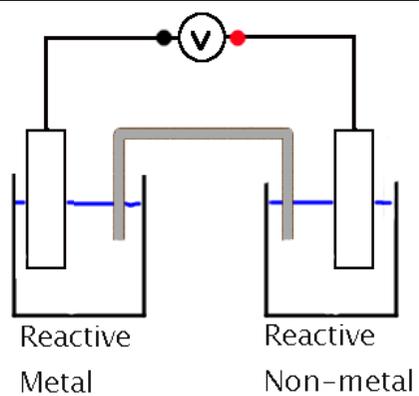
What does the sign of the electrochemical cell tell me about the REDOX behaviour of the half-cells?





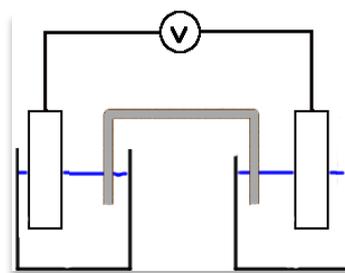
Would you expect a +ve or -ve voltage for the following electrochemical cells?

answers (clockwise starting from top left):
+ve, -ve, +ve



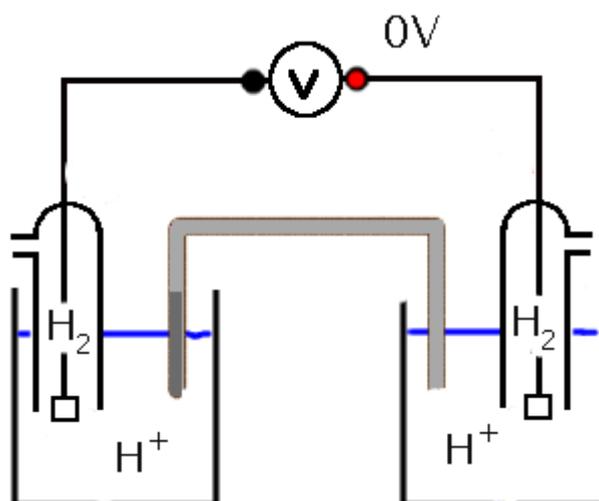
Standard Electrode Potentials

chemistry and conventions

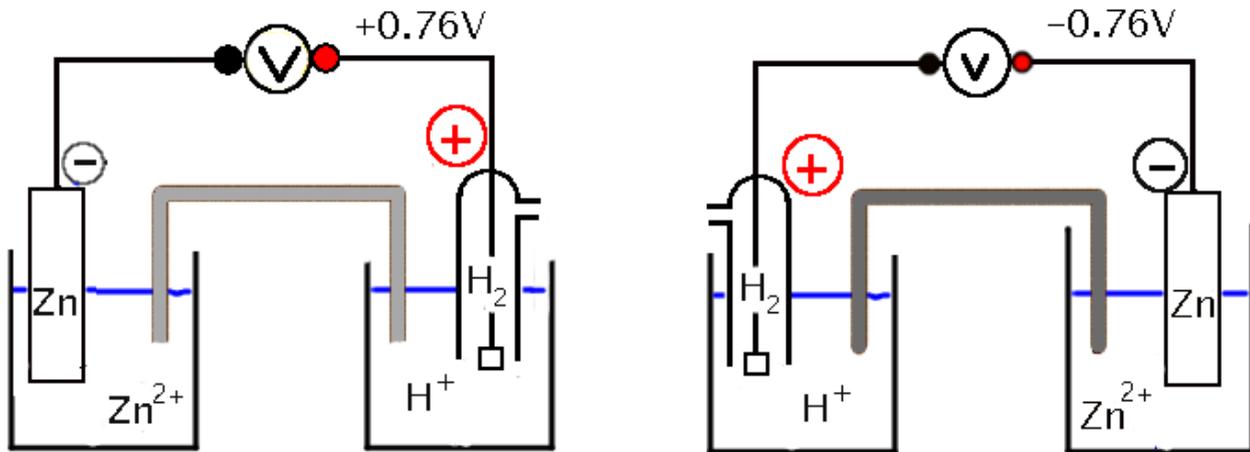


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If hydrogen is used as the reference electrode, what e.m.f. should it have?



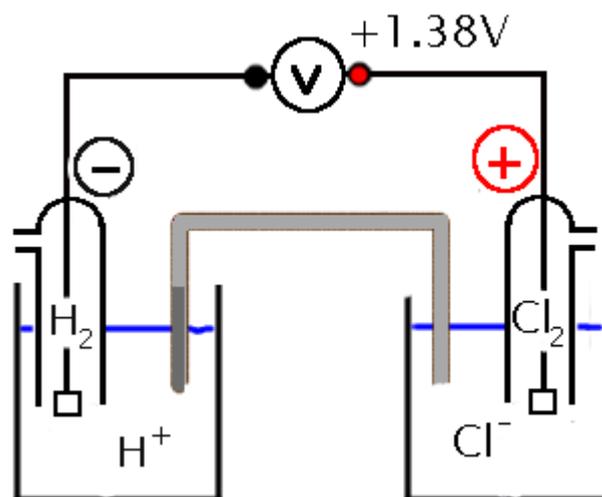
Why does it matter which side we put the reference electrode on?



CONVENTION: reference electrode on the **left hand side** (L.H.S.)

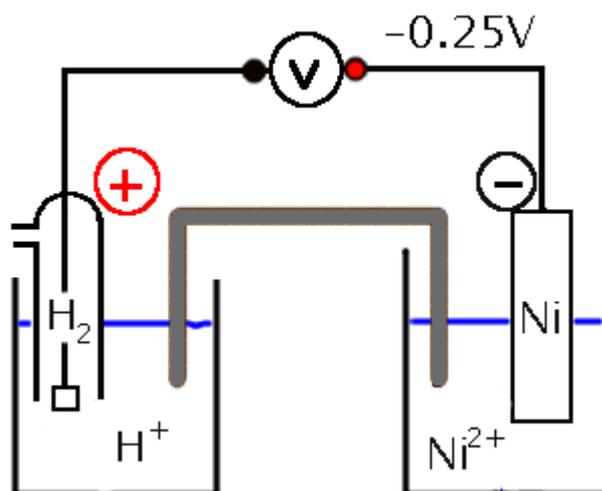
What things do we need to do consistently in order to build up a table of standard electrode potentials?

1. Left hand half-cell into negative terminal of voltmeter
2. Left hand half-cell to be H_2 , H^+
3. Standard conditions (1atm, 1mol dm^{-3} , standard states, 298K)



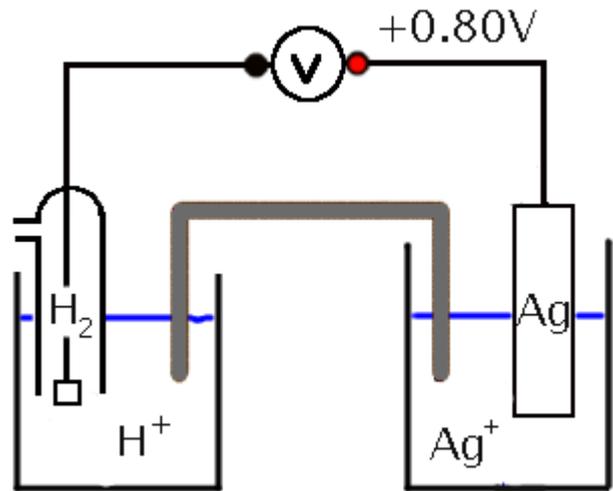
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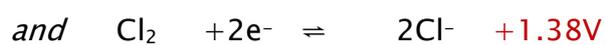
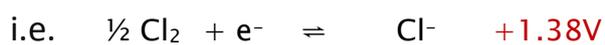


More commonly written in a table of **reduction** half-equations

SIGN of E° indicates bias of equilibrium

	Electrode	E° / V	
-ve reverse (\leftarrow)	$\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}$	-0.76	↑ tends to lose electrons
	$\text{Ni}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ni}$	-0.25	
	$\text{H}^+ + \text{e}^- \rightleftharpoons \frac{1}{2} \text{H}_2$	0	
	$\text{Cu}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cu}$	+0.34	↓ tends to gain electrons
	$\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}$	+0.80	
+ve forward (\rightarrow)	$\frac{1}{2} \text{Br}_2 + \text{e}^- \rightleftharpoons \text{Br}^-$	+1.09	
	$\frac{1}{2} \text{Cl}_2 + \text{e}^- \rightleftharpoons \text{Cl}^-$	+1.38	

note: the number of electrons in the half-equation does not affect E°



Glossary

Electrochemical cell	Two half-cells connected by a salt bridge
Electrochemical series	Shows the redox behaviour of half-reactions relative to the hydrogen electrode
Electrode	The terminal through which electric current passes between metallic and nonmetallic parts of an electric circuit
Electrode potential <i>for a cell</i> (cell potential)	The e.m.f. caused by joining two half-cells together
Electrode potential <i>for a half-cell</i>	The e.m.f. caused by a half-cell relative to the hydrogen half-cell, using a right-hand electrode system.
Electromotive force (e.m.f.) potential difference, voltage	I do not distinguish between these three terms. They all describe the extent to which there is a build up of opposite charges on either side of the voltmeter.
Half-cell	Electrode dipped into a solution of ions
Half-equation	Ionic equation showing the gain or loss of electrons of a species. In the electrochemical series half-equations are shown as <i>reductions</i> , i.e. with the electrons on the left hand side of the equation.
Hydrogen electrode	half-cell based upon the half equation $\text{H}^+ + \text{e}^- \rightleftharpoons \frac{1}{2} \text{H}_2$, and used as a reference in the electrochemical series
Salt bridge	Either a glass tube or filter paper saturated with K^+ and NO_3^- ions. Used to link two half-cells. Ions from the salt bridge can flow into the half-cells to prevent a build up of charge within the solutions as REDOX reactions occur.

References

Andrew Hunt, *Complete A-Z Chemistry Handbook* (2003). Hodder Arnold.

Nuffield Advanced Science, *Book of Data* (1984). Longman.