

Standard Electrode Potentials (E°)

**we will be concerned with standard
reduction potentials**

**for half reactions written as reductions
(electrons as reactants)**

all substances in their standard state

Standard Electrode Potentials (E°)

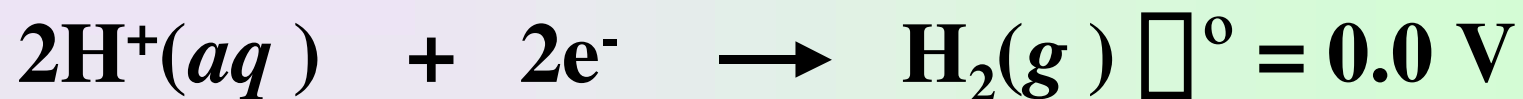
just as the over all cell reaction can be thought of as the sum of two half-cell reactions

the measured emf of the cell can be treated as the sum of the electrical potentials.

It is impossible however to measure the potential of a single electrode.

The standard hydrogen electrode*

reference all reduction half-reactions to:



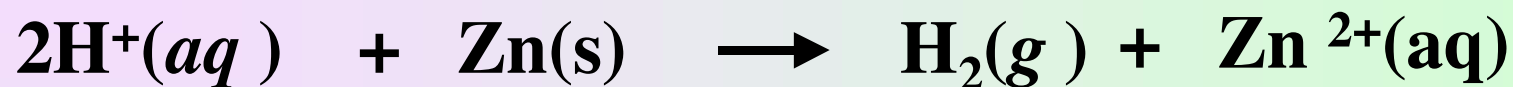
is assigned the arbitrary potential value
of 0 volts

* H_2 pressure = 1atm; $[\text{HCl}] = 1\text{M}$

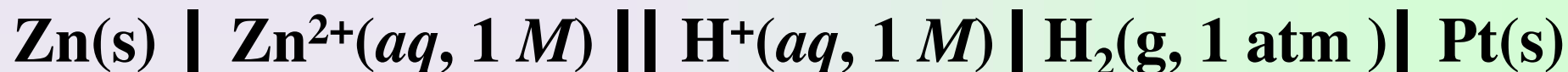
Standard Electrode Potentials (E°)

$$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$$

The reaction



represented by the cell diagram:

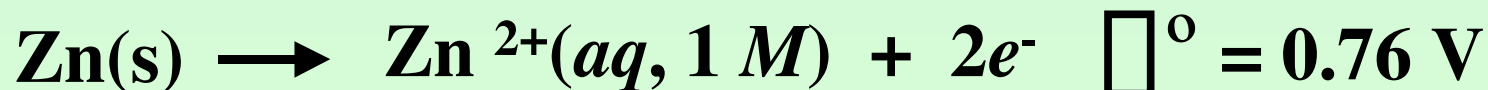


has a standard cell potential $E^\circ = 0.76 \text{ V}$

by convention:



therefore:



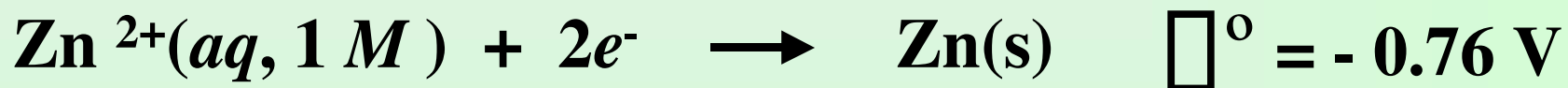
**0.76 V is the standard
oxidation potential for
Zn/Zn²⁺**

the custom is to tabulate E° values for half-reactions written as reductions

the standard oxidation potential for the half-reaction:

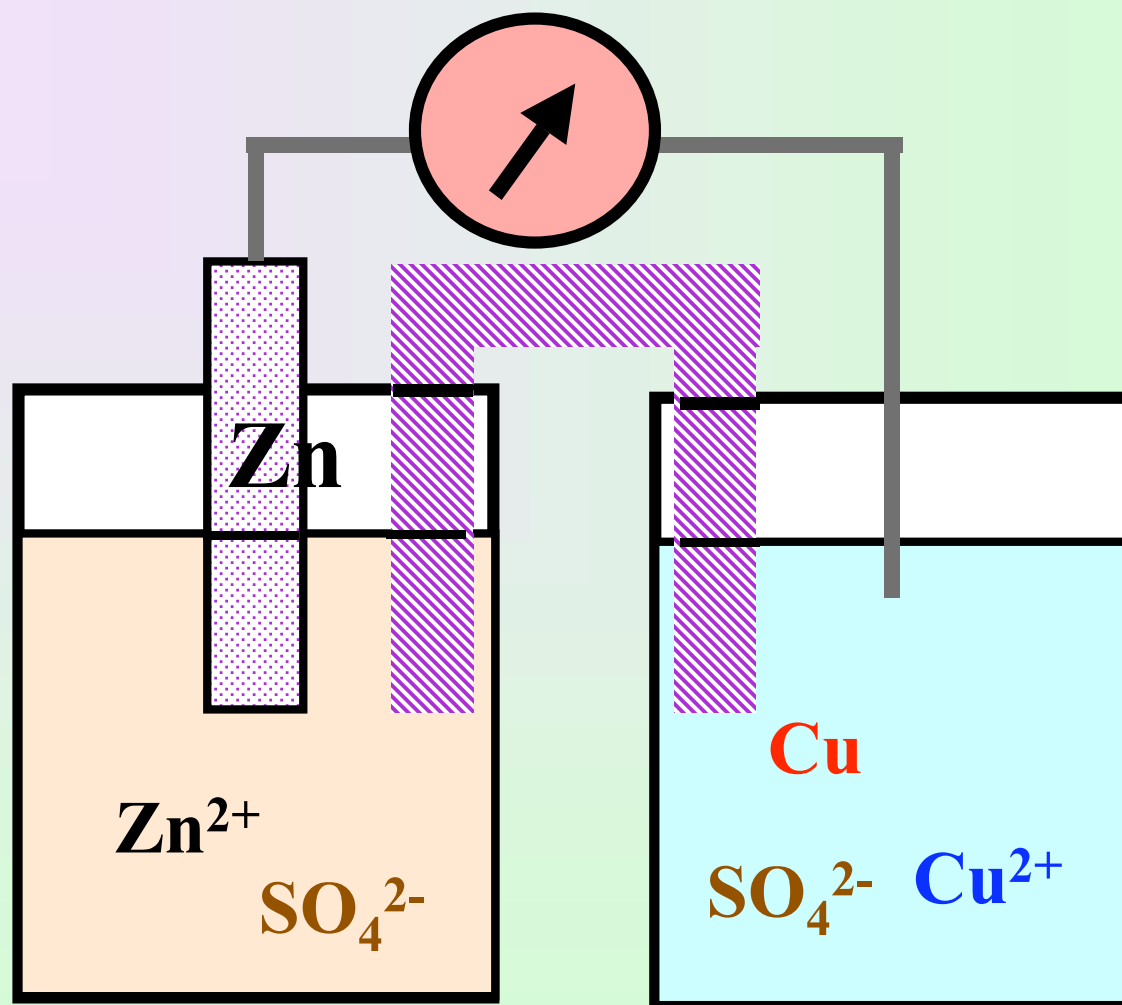


the standard reduction potential for the half-reaction:



Cell potential = 1.10 V

**Recall
the
galvanic
cell**

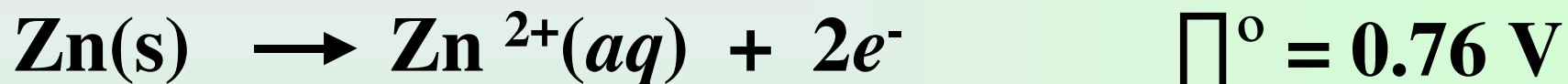


when all species are in their standard state

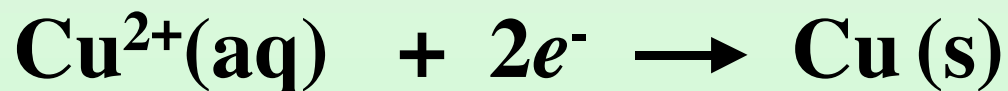
use the measured standard cell potential and the standard reduction potential of Zn^{2+} to calculate the standard reduction potential of Cu^{2+}



$$\square^{\circ} = 1.10 \text{ V}$$



Therefore:



$$\square^{\circ} = 1.10 \text{ V} - 0.76 \text{ V} = +0.34 \text{ V}$$

Standard Reduction Potentials

E° values refer to standard reduction potentials

the more positive the value of E° , the better the oxidizing agent

Standard Reduction Potentials

Better reducing agent

\square°



Better oxidizing agent

An Analogy



$K > 1$ when

Stronger acid \longrightarrow **Weaker acid**

as measured by K_a

An Analogy

oxidizing agent + reducing agent \rightleftharpoons oxidized material + reduced material

K > 1 when

Stronger oxidant \longrightarrow Weaker oxidant

as measured by E°

Example

Can $\text{Zn}^{2+}(\text{aq})$ oxidize Sn under standard state conditions?



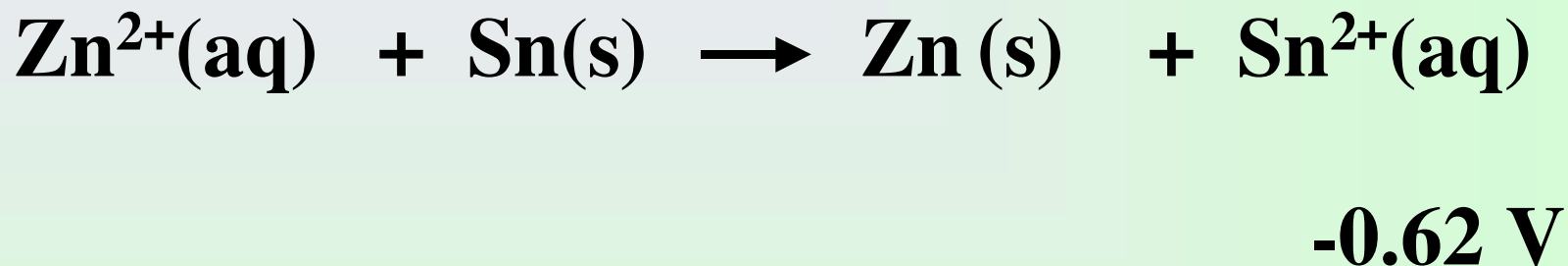
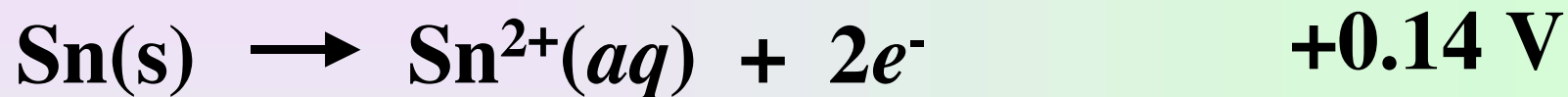
What are the half- reactions?



Zn^{2+} is a worse oxidizing agent than Sn^{2+} ;

Zn^{2+} will not oxidize Sn

Example: what is the cell potential? \square°



If the sign is negative; the reaction is not spontaneous in the direction written.

Standard Reduction Potentials

Better reducing agent

\square°

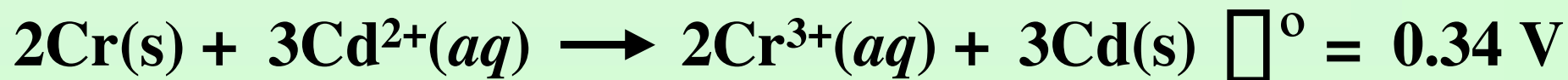
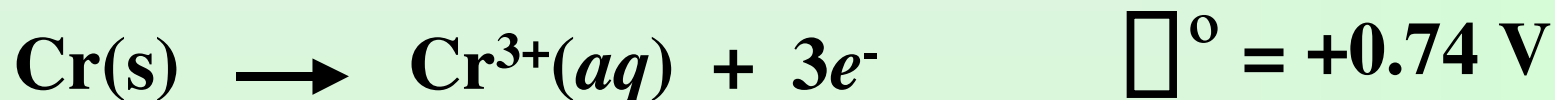


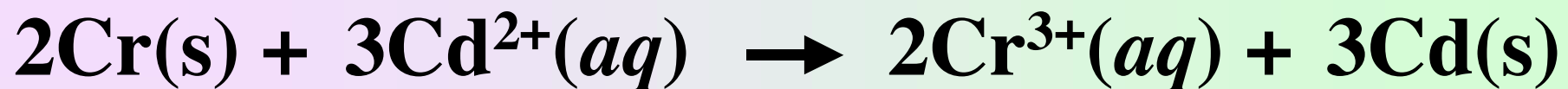
Better oxidizing agent

Example: What is the the standard emf of a galvanic cell made of a Cd electrode in a 1.0 M $\text{Cd}(\text{NO}_3)_2$ solution and a Cr electrode in a 1.0 M $\text{Cr}(\text{NO}_3)_3$ solution?



reverse the equation with more negative \square° and add





$$E^{\circ} = 0.34 \text{ V}$$

E° is an intensive property

so multiplying the reduction of Cd^{2+} by 3 and the oxidation of Cr by 2 has no effect on the value of the voltage

e^- 

Cr anode

Cd cathode

1M $\text{Cr}(\text{NO}_3)_3$

1M $\text{Cd}(\text{NO}_3)_2$

emf = 0.34 V

