

Standard Reduction Potential

The potential generated by a **reduction half-reaction** at an electrode measured relative to the **standard hydrogen electrode** under standard condition (temperature of 25 °C , ion concentration of 1 M and partial pressure of gas at 1 atm).


- E°_{red} measured the ease of reduction under standard conditions.

Example :

Reduction



E°_{red} $E^{\circ}_{\text{Cu}^{2+}/\text{Cu}}$



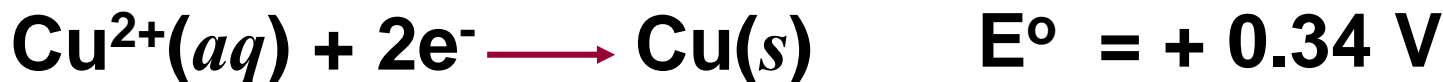
$E^{\circ} = + 0.34 \text{ V}$

Standard electrode potential is the potential difference produced when a half cell of an element is connected to **Standard Hydrogen Electrode** (S.H.E) under **standard conditions** (i.e. a temperature of 25 °C, ion concentration of 1M, partial pressure of gas of 1 atm)

- By convention, all **Standard Electrode Potentials** are written as **Standard Reduction Potential**, E°_{red}
- **SHE** is used as a reference electrode and is arbitrarily assigned a **standard reduction potential** of 0 volt.

- The **more positive** E°_{red} (or less negative), the easier the species undergo reduction.

Example :



Cu²⁺ is easier to be reduced than Zn²⁺.

Example :



Zn²⁺ is easier to be reduced than Na⁺.

- To obtain the **standard oxidation potentials**, the reduction half-reactions are reversed and the sign of E° changes.

Example :

Reduction

E°_{red} $E^\circ_{\text{Cu}^{2+}/\text{Cu}}$
↓



Oxidation

E°_{ox} $E^\circ_{\text{Cu}/\text{Cu}^{2+}}$
↓



- The **more -ve** the E°_{red} value, the greater tendency for the **reverse reaction (oxidation)** to occur **or**
- The **more +ve** E°_{ox} (or less -ve), the **easier** the species **undergo oxidation**.



Cu^{2+} is easier to be **reduced** than Zn^{2+} .

Zn is easier to be **oxidised** than Cu .

- Changing the stoichiometric coefficients of a half-cell reaction **does not change the value of E° .**

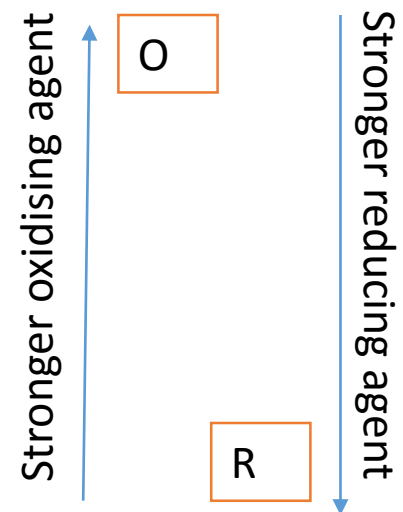
Example :



Standard Reduction Potentials

Potential (V)	Reduction Half-Reaction
+2.87	Strongest oxidising agent $F_2(g) + 2 e^- \longrightarrow 2 F^-(aq)$
+1.51	$MnO_4^-(aq) + 8 H^+(aq) + 5 e^- \longrightarrow Mn^{2+}(aq) + 4 H_2O(l)$
+1.36	$Cl_2(g) + 2 e^- \longrightarrow 2 Cl^-(aq)$
+1.33	$Cr_2O_7^{2-}(aq) + 14 H^+(aq) + 6 e^- \longrightarrow 2 Cr^{3+}(aq) + 7 H_2O(l)$
+1.23	$O_2(g) + 4 H^+(aq) + 4 e^- \longrightarrow 2 H_2O(l)$
+1.06	$Br_2(l) + 2 e^- \longrightarrow 2 Br^-(aq)$
+0.96	$NO_3^-(aq) + 4 H^+(aq) + 3 e^- \longrightarrow NO(g) + 2 H_2O(l)$
+0.80	$Ag^+(aq) + e^- \longrightarrow Ag(s)$
+0.77	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$
+0.68	$O_2(g) + 2 H^+(aq) + 2 e^- \longrightarrow H_2O_2(aq)$
+0.59	$MnO_4^-(aq) + 2 H_2O(l) + 3 e^- \longrightarrow MnO_2(s) + 4 OH^-(aq)$
+0.54	$I_2(s) + 2 e^- \longrightarrow 2 I^-(aq)$
+0.40	$O_2(g) + 2 H_2O(l) + 4 e^- \longrightarrow 4 OH^-(aq)$
+0.34	$Cu^{2+}(aq) + 2 e^- \longrightarrow Cu(s)$
0 [defined]	$2 H^+(aq) + 2 e^- \longrightarrow H_2(g)$
-0.28	$Ni^{2+}(aq) + 2 e^- \longrightarrow Ni(s)$
-0.44	$Fe^{2+}(aq) + 2 e^- \longrightarrow Fe(s)$
-0.76	$Zn^{2+}(aq) + 2 e^- \longrightarrow Zn(s)$
-0.83	$2 H_2O(l) + 2 e^- \longrightarrow H_2(g) + 2 OH^-(aq)$
-1.66	$Al^{3+}(aq) + 3 e^- \longrightarrow Al(s)$
-2.71	$Na^+(aq) + e^- \longrightarrow Na(s)$
-3.05	$Li^+(aq) + e^- \longrightarrow Li(s)$ Strongest reducing agent

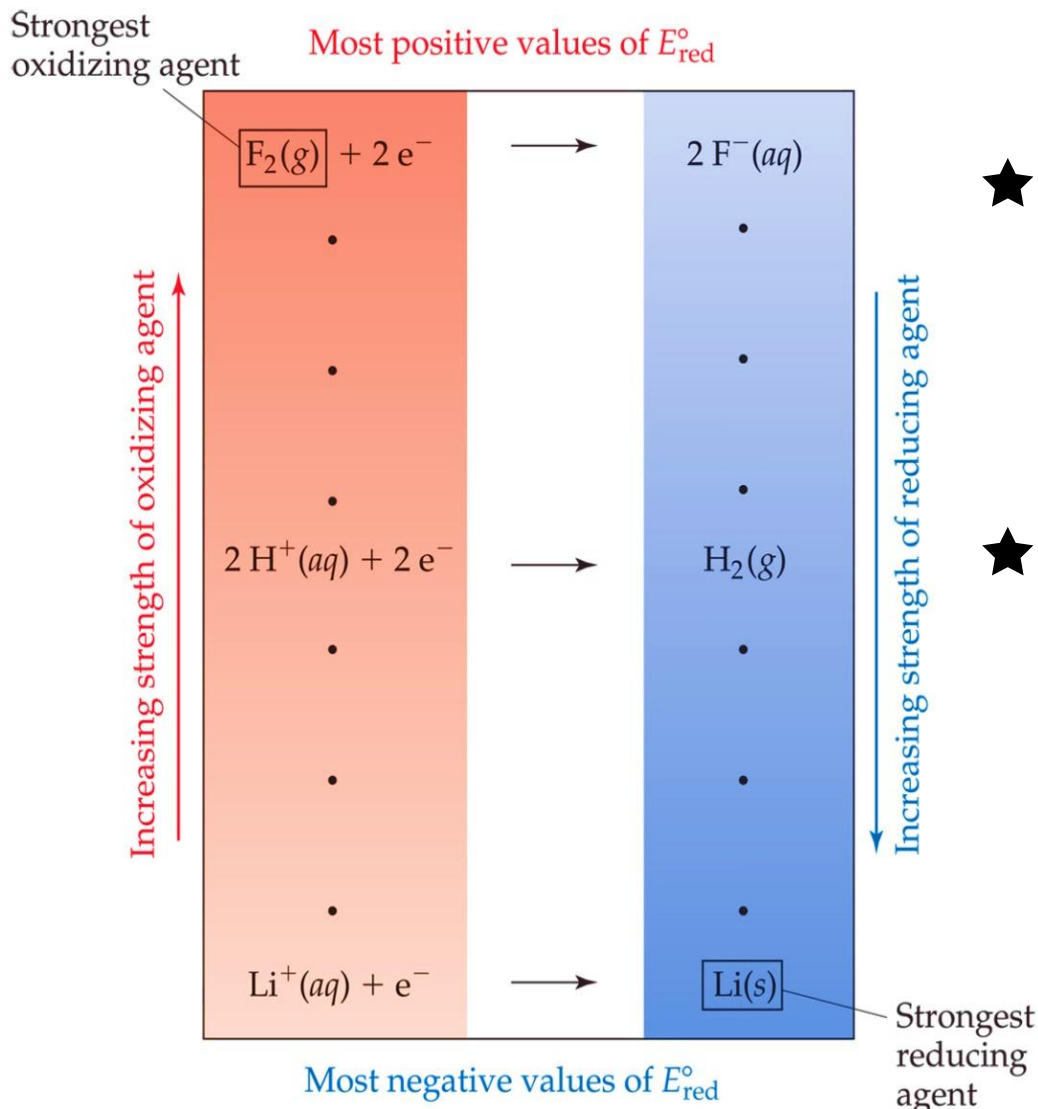
Standard Reduction Potentials for many electrodes have been measured and tabulated.



THE USES OF STANDARD REDUCTION POTENTIAL (S.R.P) VALUES

- 1. To show the relative oxidising and reducing power of different species.**
- 2. To calculate standard cell potential (standard e.m.f.)**
- 3. To predict spontaneous redox reaction**

1. To show the relative oxidising and reducing power of different species.



★ More positive E°

- More easily electron is gained
- More easily reduced
- Stronger oxidizing agent

★ More negative E°

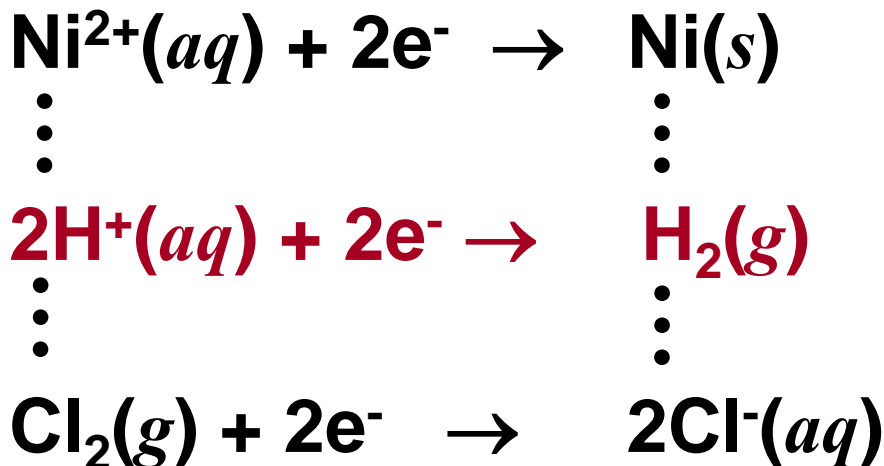
- More easily electron is lost
- More easily oxidised
- Stronger reducing agent

Oxidising agent

Reducing agent

E°_{red} (V)

Increasing strength of oxidising agent



-0.25

⋮

0.00

⋮

+1.36

More +ve



Increasing strength of oxidising agent



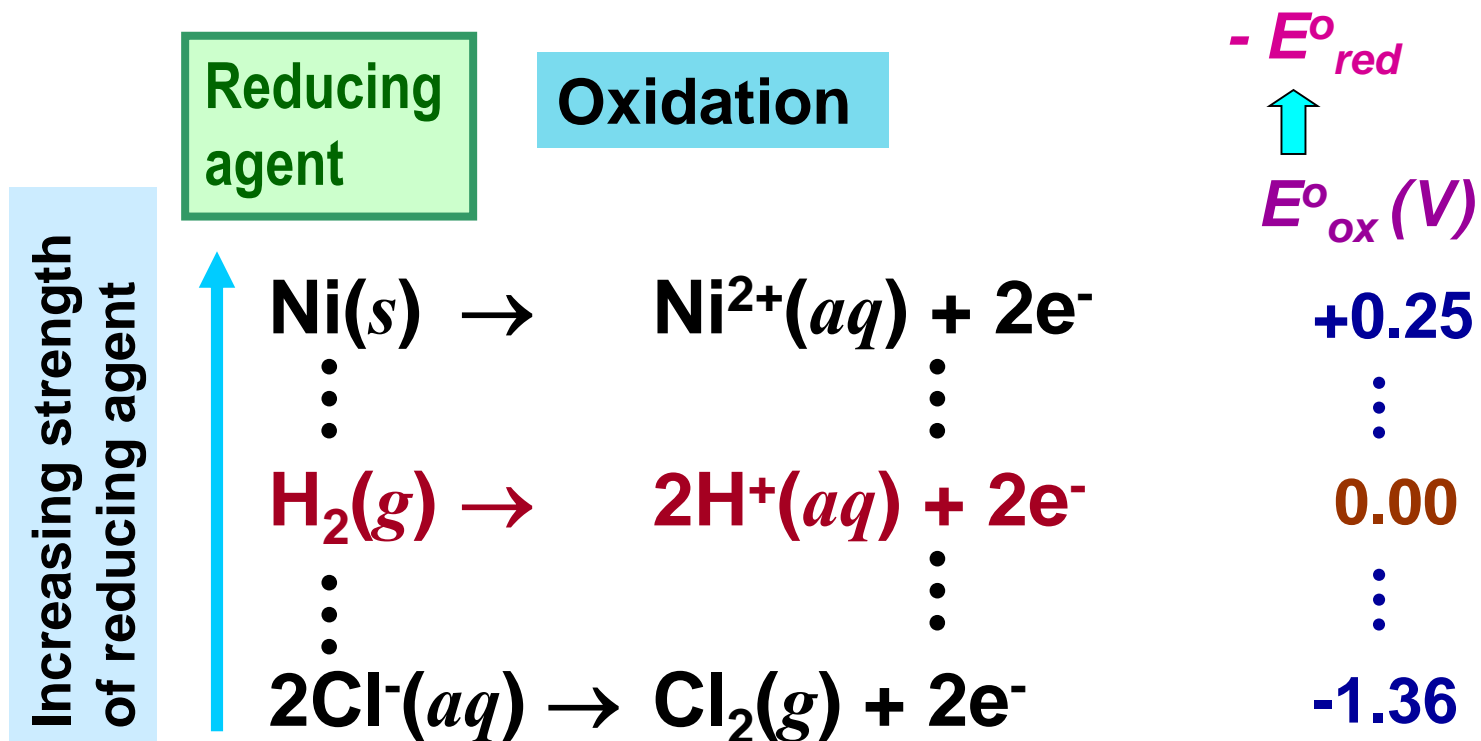
Increasing strength of reducing agent

The more positive E°_{red} , the stronger the oxidising agent.

The more negative E°_{red} , the stronger the reducing agent

Note !

For a reducing agent to supply electrons, the oxidation reaction must occur in the reverse direction.

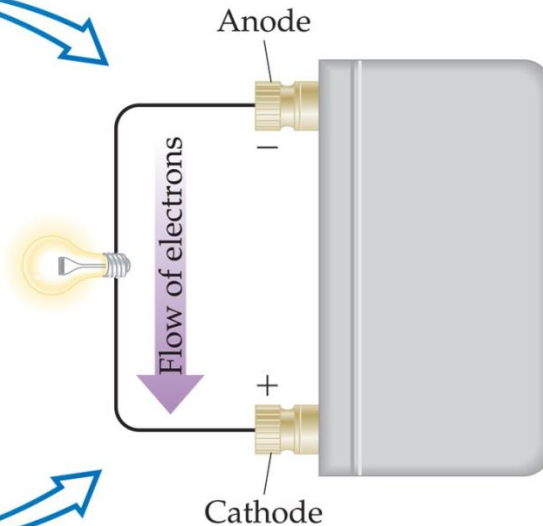


2. To calculate standard cell potential (standard e.m.f)

Electromotive Force (emf)



High potential energy



Low potential energy

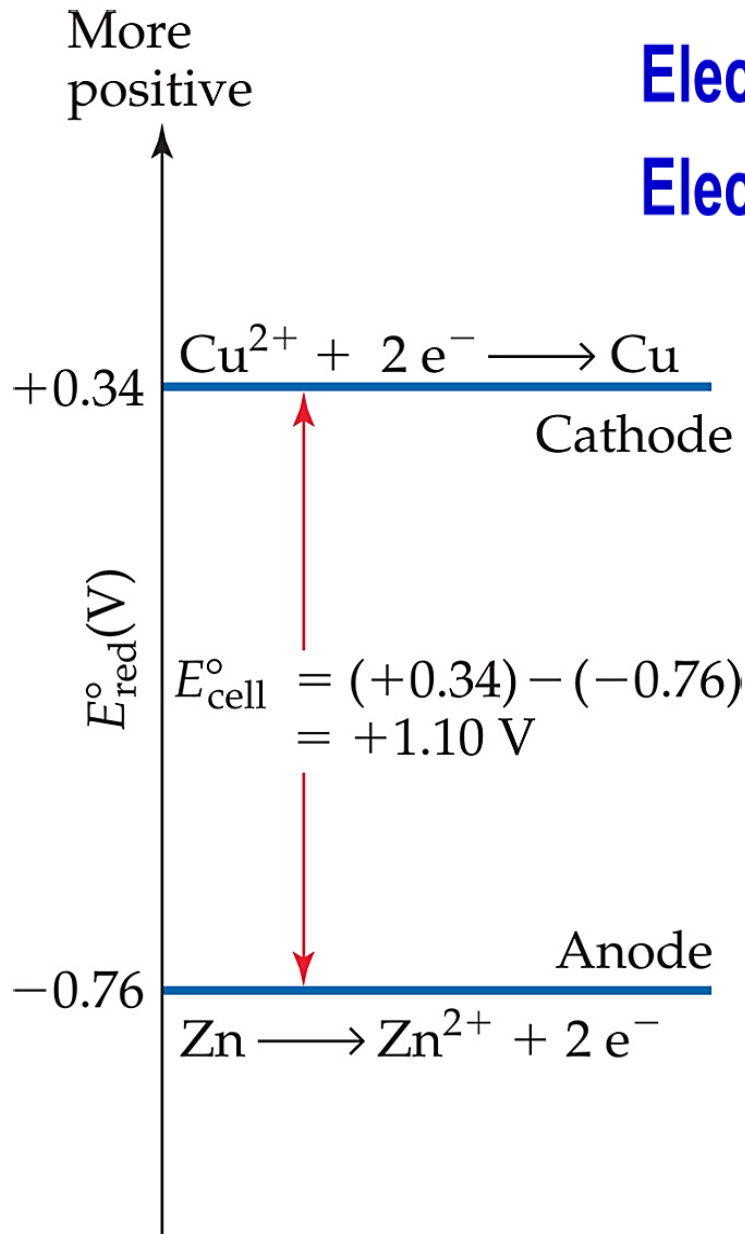
- Water only spontaneously flows one way in a waterfall.
- Likewise, electrons only spontaneously flow one way in a redox reaction—from higher to lower potential energy.

ELECTROMOTIVE FORCE (EMF)

- The potential difference between the anode and cathode in a cell is called the **electromotive force (emf)**.
- It is also called the **cell potential**, and is designated E_{cell} .
- **Cell potential at standard conditions** is known as **Standard Cell Potential** and is calculated using formula:

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

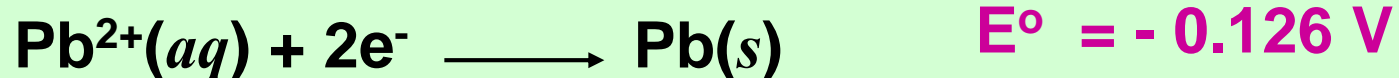
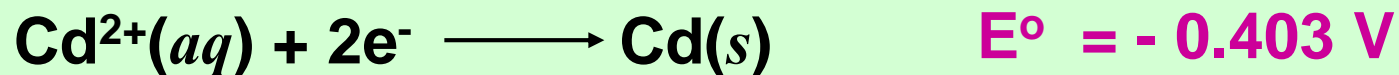
Electrode with more positive E° is cathode.
Electrode with more negative E° is anode.



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

The greater the difference in E° between the two electrodes or half-cells, the greater the voltage of the cell.

Determine the E°_{cell} for a voltaic cell which is based on the following two standard half-reactions :



Solution

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

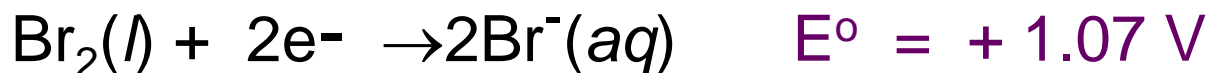
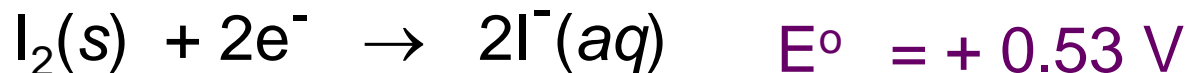
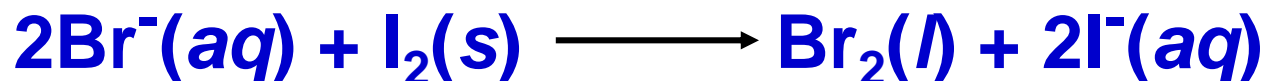
$$= E^\circ_{\text{Pb}^{2+}/\text{Pb}} - E^\circ_{\text{Cd}^{2+}/\text{Cd}}$$

$$= -0.126 - (-0.403)$$

$$= +0.277 \text{ V}$$

3. To Predict Spontaneous Redox Reaction

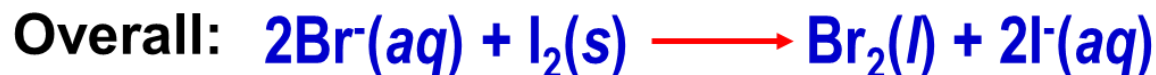
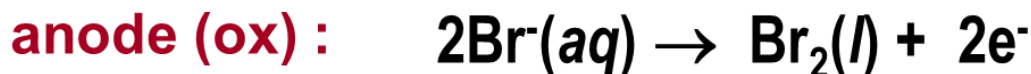
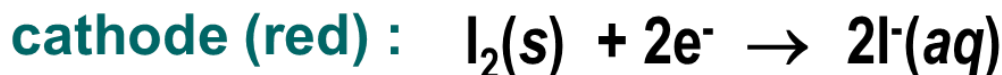
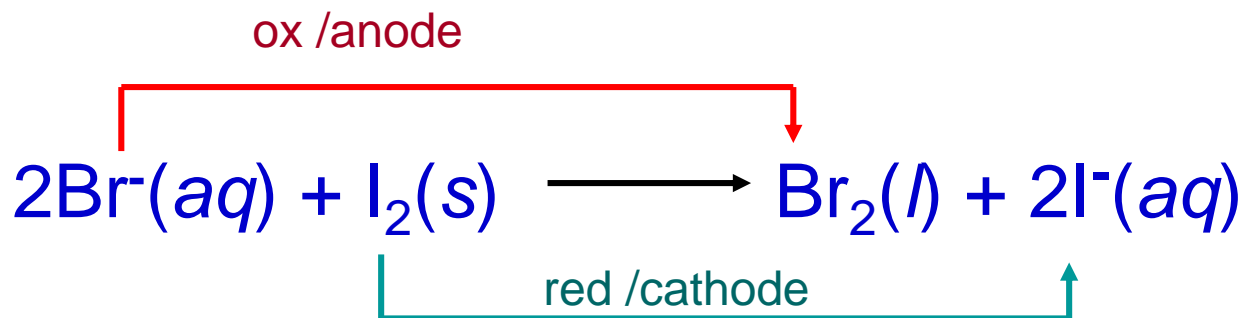
By calculating the E° , determine whether Br^- ion can reduce I_2 under standard conditions.



Note :

In predicting the spontaneity of a reaction, anode and cathode is based on equation given, not based on E° .

Solution :



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$= E^\circ_{\text{I}_2/\text{I}^-} - E^\circ_{\text{Br}_2/\text{Br}^-}$$

$$= 0.53 - 1.07$$

$$= -0.54 \text{ V} //$$

$$E^\circ_{\text{cell}} < 0$$

\therefore The Br⁻ ions **will not** reduce I₂ spontaneously.