

Redox Questions on oxidising and reducing agents

Table of standard electrode potentials for use in the questions below.

$\text{Br}_2(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	$E^\ominus = +1.09 \text{ V}$
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ce}^{3+}(\text{aq})$	$E^\ominus = +1.61 \text{ V}$
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Co}(\text{s})$	$E^\ominus = -0.28 \text{ V}$
$\text{Co}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Co}^{2+}(\text{aq})$	$E^\ominus = +1.82 \text{ V}$
$\text{Cr}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	$E^\ominus = -0.74 \text{ V}$
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cr}^{2+}(\text{aq})$	$E^\ominus = -0.41 \text{ V}$
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	$E^\ominus = +1.36 \text{ V}$
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	$E^\ominus = +0.34 \text{ V}$
$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}^+(\text{aq})$	$E^\ominus = +0.15 \text{ V}$
$\text{Cu}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	$E^\ominus = +0.52 \text{ V}$
$\text{Eu}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Eu}^{2+}(\text{aq})$	$E^\ominus = -0.43 \text{ V}$
$\text{F}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{F}^-(\text{aq})$	$E^\ominus = +2.87 \text{ V}$
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	$E^\ominus = -0.45 \text{ V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	$E^\ominus = +0.77 \text{ V}$
$\text{Ho}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ho}(\text{s})$	$E^\ominus = -2.32 \text{ V}$
$\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$	$E^\ominus = +1.46 \text{ V}$
$\text{I}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$	$E^\ominus = +0.54 \text{ V}$
$2\text{IO}_3^-(\text{aq}) + 12\text{H}^+(\text{aq}) + 10\text{e}^- \rightleftharpoons \text{I}_2(\text{s}) + 6\text{H}_2\text{O}(\text{l})$	$E^\ominus = +1.19 \text{ V}$
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	$E^\ominus = +1.51 \text{ V}$
$\text{Na}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Na}(\text{s})$	$E^\ominus = -2.71 \text{ V}$
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	$E^\ominus = +0.80 \text{ V}$
$\text{Pu}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Pu}^{3+}(\text{aq})$	$E^\ominus = +0.97 \text{ V}$
$\text{Pu}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Pu}(\text{s})$	$E^\ominus = -2.03 \text{ V}$
$\text{U}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{U}^{3+}(\text{aq})$	$E^\ominus = -0.61 \text{ V}$
$\text{U}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{U}(\text{s})$	$E^\ominus = -1.79 \text{ V}$
$\text{UO}_2^{2+}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{U}^{4+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$	$E^\ominus = +0.33 \text{ V}$

Redox Questions on oxidising and reducing agents

1 Which is the strongest oxidising agent in this table?F₂..... Not Na⁺

2 Which is the strongest reducing agent in this table?Na.....

3 State whether each of the following is TRUE or FALSE

(a) Ce⁴⁺ is a stronger oxidising agent than Br₂

TRUE or FALSE Ce⁴⁺ has a more positive E° value

(b) Co is a stronger reducing agent than Cr

TRUE or **FALSE** Looking at the oxidation reaction (reverse reaction) – the potential is more positive for Cr

(c) Cr²⁺ is a stronger reducing agent than Co²⁺

TRUE or FALSE Looking at the oxidation reaction M²⁺ → M³⁺ + e⁻

(d) Pu will reduce Cr²⁺ to Cr

TRUE or FALSE

(e) MnO₄⁻/H⁺ will oxidise Pu³⁺ to Pu⁴⁺

TRUE or FALSE Pu has a more negative E° value, therefore is a stronger reducing agent than Cr

(f) Ce⁴⁺ will oxidise F⁻ to F₂

TRUE or **FALSE**

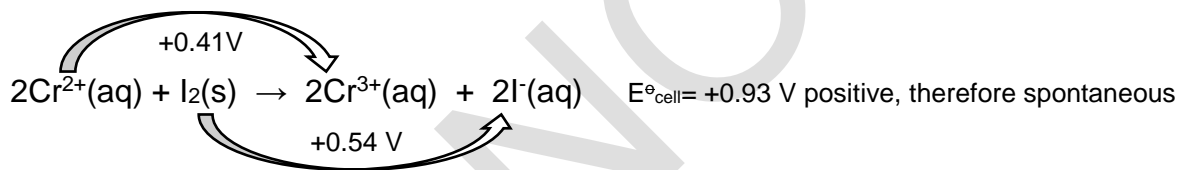
(g) Br₂ is a stronger oxidising agent than Pu⁴⁺

TRUE or FALSE

4 Deduce whether each of the following reactions will be spontaneous or not

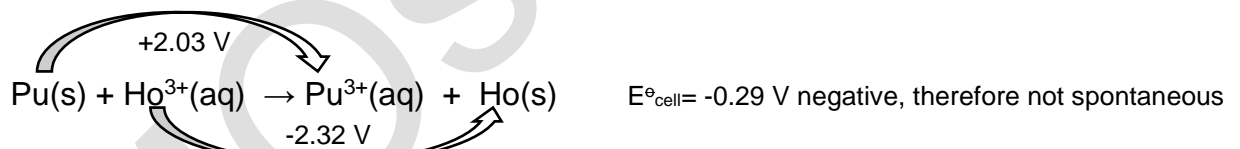
(a) 2Cr²⁺(aq) + I₂(s) → 2Cr³⁺(aq) + 2I⁻(aq)

SPONTANEOUS or NOT SPONTANEOUS



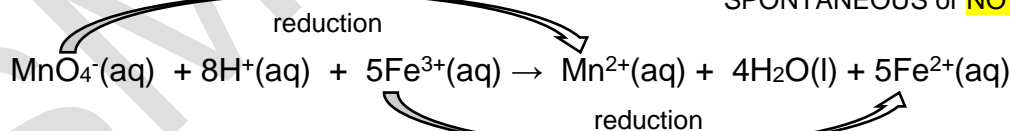
(b) Pu(s) + Ho³⁺(aq) → Pu³⁺(aq) + Ho(s)

SPONTANEOUS or **NOT SPONTANEOUS**



(c) MnO₄⁻(aq) + 8H⁺(aq) + 5Fe³⁺(aq) → Mn²⁺(aq) + 4H₂O(l) + 5Fe²⁺(aq)

SPONTANEOUS or **NOT SPONTANEOUS**



Both half-reactions involve reduction, therefore this cannot be spontaneous.

Value between -1.51 V and -1.36 V for the oxidation reaction (Pb²⁺ to PbO₂)

5 (a) Identify a species that can be oxidised by MnO₄⁻/H⁺ but not by Cr₂O₇²⁻/H⁺.....Pb²⁺.....

(b) Identify a species that will oxidise Fe²⁺ to Fe³⁺ but not Pu³⁺ to Pu⁴⁺.....Ag⁺.....

E° value between +0.77 V and +0.97 V. Must be Ag⁺, not Ag – Ag⁺ is the oxidising agent

Redox Questions on oxidising and reducing agents

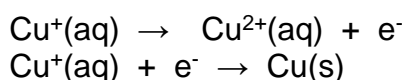
6 (a) By considering the following electrode potentials



deduce what will happen to Cu^{+} ions in solution – write an equation.

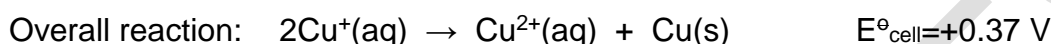
Cu^{+} ions will be oxidised to Cu^{2+} and reduced to Cu

The same species is oxidised and reduced – a *disproportionation* reaction



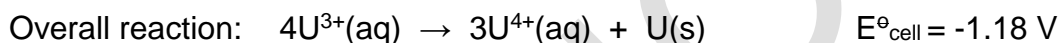
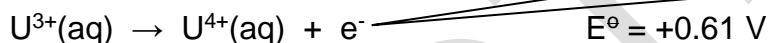
$$\begin{aligned} E^{\circ}_{\text{ox}} &= -0.15 \text{ V} \\ E^{\circ} &= +0.52 \text{ V} \end{aligned}$$

Potential for oxidation



Positive value, therefore spontaneous.

(b) Deduce whether $\text{U}^{3+}(\text{aq})$ do a similar reaction to the reaction in (a)



Need to multiply this half-equation by 3 to balance the electrons when writing the overall redox equation

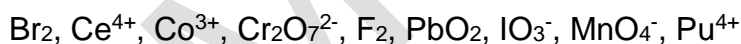
U^{3+} will not undergo a similar reaction (will not disproportionate) because the cell potential is negative – the reaction shown is not spontaneous.

7 Deduce an equation for a spontaneous reaction involving Ag .

The only possibility using the half equations given in the table is the oxidation of Ag to Ag^{+} :



Anything with a standard electrode potential more positive than 0.80 V will oxidise Ag to Ag^{+} , so there are several possibilities:



Balanced equations:

