Chem Factsheet

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Applying electrode potential data

Reminders from Chem FactSheet 199:

1. Each **half equation** has an associated **standard electrode potential** e.g. $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$; $E^\ominus = +0.34V$
2. Each electrode potential is measured relative to a **standard hydrogen electrode**. By convention, this is made to correspond to the zero point of a scale of measurement of redox powers by **assigning it an $E^\ominus$ value of 0.00 volts**.
3. All $E^\ominus$ measurements are made at **1.00mol dm^{-3}, 298K and 100kPa**.
4. A **more negative** $E^\ominus$ value indicates a **more reducing** system. i.e. one that is more likely to **give electrons**, causing **reduction**, but become **oxidised itself**.
5. A **more positive** $E^\ominus$ value indicates a **more oxidising** system. i.e. one that is more likely to **take electrons**, causing **oxidation**, but become **reduced itself**.
6. Arranged in order of $E^\ominus$ values, a list of half equations with corresponding $E^\ominus$ values is called the **electrochemical series**.
7. Usually, they are arranged such that the **more negative $E^\ominus$ values appear at the top** of the list with the **oxidising agents (oxidants) on the left** and the **reducing agents (reductants) on the right**.

OXIDANTS		REDUCTANTS	$E^\ominus$ /V
1 $Zn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Zn(s)$	- 0.76
2 $V^{3+}(aq) + e^-$	$\rightleftharpoons$	$V^{2+}(aq)$	- 0.26
3 $2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2(g)$	0.00
4 $SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2SO_3(aq) + H_2O(l)$	+ 0.17
5 $Cu^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cu(s)$	+ 0.34
6 $VO^{2+}(aq) + 2H^+(aq) + e^-$	$\rightleftharpoons$	$V^{3+}(aq) + H_2O(l)$	+ 0.34
7 $Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$	+ 0.77
8 $VO_2^+(aq) + 2H^+(aq) + e^-$	$\rightleftharpoons$	$VO^{2+}(aq) + H_2O(l)$	+ 1.00
9 $MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\rightleftharpoons$	$Mn^{2+}(aq) + 4H_2O(l)$	+ 1.51

Given such data, you may be asked to carry out any of the following tasks. In the self-assessed questions, S1 etc, delete one or more of the alternatives or insert words or phrases to create correct statements. Check your answers at the end of this FactSheet.

1. PLACE SPECIES IN ORDER OF REDUCING or OXIDISING POWER

S1 The **more reducing** a system is, then the more **NEGATIVE / POSITIVE** its associated $E^\ominus$ value is.

S2 The **more oxidising** a system is, then the more **NEGATIVE / POSITIVE** its associated $E^\ominus$ value is.

S3 Hence, in Table 1,

(a) the **most oxidising** species is ____________________________

(b) the **least oxidising** species is ____________________________

(c) the **most reducing** species is ____________________________

(d) the **least reducing** species is ____________________________

S4 Arrange acidified $VO_2^+(aq)$, $V^{3+}(aq)$, acidified $SO_4^{2-}(aq)$ and $Zn^{2+}(aq)$ in order of **increasing oxidising power**.

____

S5 Arrange acidified $V^{3+}(aq)$, $Fe^{2+}(aq)$, $Zn(s)$ and $Mn^{2+}(aq)$ in order of **increasing reducing power**.

____

2. DECIDE WHICH SPECIES ARE CAPABLE OF OXIDISING OR REDUCING ANOTHER SPECIES

Remember a redox reaction **MUST** occur between a reducing agent and an oxidising agent so that one can give electrons and the other can accept them.

S6 Reducing agent "A" *should* reduce oxidising agent "B" provided A's $E^\ominus$ value is more **NEGATIVE / POSITIVE** than B's

S7 Alternatively, oxidising agent "B" *should* oxidise reducing agent "A" provided B's $E^\ominus$ value is more **NEGATIVE / POSITIVE** than A's.

Note the use of the word "should"! Predictions based on $E^\ominus$ values give the **energetic feasibility** under **standard conditions**. However, a predicted "positive" reaction may not actually occur because:
(a) it is **kinetically controlled** by a high activation energy or
(b) it is carried out under **non-standard conditions**.

Using the data from Table

S8 Zinc **WILL / WILL NOT** *reduce* iron(III) because its $E^\ominus$ value is **MORE / LESS** negative than that of iron(III).

S9 Iron(II) **WILL / WILL NOT** *reduce* vanadium(III) because its $E^\ominus$ value is **MORE / LESS** negative than that of vanadium(III).

S10 Iron(III) **WILL / WILL NOT** *oxidise* $VO^{2+}(aq)$ because its $E^\ominus$ value is **MORE / LESS** positive than that of $VO^{2+}(aq)$.

S11 Acidified $VO_2^+(aq)$ **WILL / WILL NOT** *oxidise* $V^{2+}(aq)$ because its $E^\ominus$ value is **MORE / LESS** positive than that of $VO^{2+}(aq)$.

3. DECIDE THE REACTION PRODUCTS OF A REDOX REACTION

Remembering that the **reducing agent is oxidised** during the course of a redox reaction, and the **oxidising agent is reduced**, then:

(a) the half equation with the **more positive $E^\ominus$ value** (i.e. more oxidising) will proceed from **left to right** (i.e. reduction). and

(b) the half equation with the more negative $E^\ominus$ value (i.e. more reducing) will proceed from right to left (i.e. oxidation).

S12 Using S8, zinc will *reduce* iron(III) because its $E^\ominus$ value is more negative than that of iron(III). Hence the zinc will change to __________ and the iron(III) will change to ___________.

S13 Using S11, acidified $VO_2^+(aq)$ will *oxidise* $V^{2+}(aq)$ because its $E^Ø$ value is more positive than that of $VO^{2+}(aq)$. Hence $VO_2^+(aq)$ will change to __________ and the $V^{2+}(aq)$ will change to ___________.

Be careful to look for **multiple** oxidations or reductions! This can occur when a an element has **several different oxidation states**. Vanadium is a good example because it can be V(+2), (+3), (+4) or (+5). Hence, you might have predicted a reduction of V(+5) to (+4) but you should check whether further reduction to (+3) and/or (+2) is possible for the reducing agent being used.

e.g. **Oxidising** $V^{2+}(aq)$ using $Fe^{3+}(aq)$ will firstly produce $V^{3+}(aq)$ since +0.77V for $Fe^{3+}(aq)$ is more positive than -0.26V for $V^{2+}(aq)$.

Also, provided there is sufficient oxidising agent, the V^{3+}(aq) product will then be further oxidised to VO^{2+}(aq) since +0.77V for Fe^{3+}(aq) is also more positive than +0.34V for V^{3+}(aq). However, the VO^{2+}(aq) will **NOT** then be oxidised to VO_2^+(aq) since +0.77V for Fe^{3+}(aq) is NOT more positive than +1.00V for VO^{2+}(aq).

S14 **Reducing** VO_2^+(aq) using H_2SO_3(aq) will firstly produce ________ since +0.17V for H_2SO_3(aq) is more negative than +1.00V for VO_2^+(aq).
Also, provided there is sufficient reducing agent, this product will then be further reduced to ________ since +0.17V for H_2SO_3(aq) is also more negative than +0.34V for VO^{2+}(aq).
However, this product will **NOT** then be reduced to ________ since +0.17V for H_2SO_3(aq) is NOT more negative than -0.26V for V^{3+}(aq).

4.`CALCULATE $E^\ominus_{cell}$ FOR A GIVEN REDOX REACTION

> For a general redox reaction,
> OXIDANT + REDUCTANT → PRODUCTS.
> $E^\ominus_{cell}$ for the reaction is given by $\mathbf{E^\ominus_{OXIDANT} - E^\ominus_{REDUCTANT}}$. You may see this presented as: $\mathbf{E^\ominus_{cell} = E^\ominus_{REDUCED\ SUBSTANCE} - E^\ominus_{OXIDISED\ SUBSTANCE}}$ since the oxidant is reduced and the reductant is oxidised.

e.g. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Zn is the reductant since it gives electrons to the Cu^{2+}.

Cu^{2+} is the oxidant since it takes electrons from the Zn.

$Zn^{2+}(aq) + 2e- \rightleftharpoons Zn(s)$ - 0.76V
$Cu^{2+}(aq) + 2e- \rightleftharpoons Cu(s)$ + 0.34V
Hence, $E^\ominus_{cell}$ = +0.34 – (-0.76) = **+1.10V**.

S15 $Zn^{2+}(aq) + 2Fe^{2+}(aq) \rightarrow Zn(s) + 2Fe^{3+}(aq)$

_______ is the reductant since it gives electrons to the ________.
_______ is the oxidant since it takes electrons from the _______.

$Zn^{2+}(aq) + 2e- \rightleftharpoons Zn(s)$ - 0.76V
$Fe^{3+}(aq) + e- \rightleftharpoons Fe^{2+}(aq)$ + 0.77V
Hence, $E^\ominus_{cell}$ = ________________ = ____________.

5. USE $E^\ominus_{cell}$ TO DECIDE THE FEASIBILITY OF A REDOX REACTION

> A redox reaction is energetically feasible under standard conditions if, and only if, $\mathbf{E^\ominus_{cell} > 0}$

Hence, using the example from S15.

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

IS energetically feasible since its $\mathbf{E^\ominus_{cell}}$ **is positive** (+1.10V).

Similarly, using the example from S15,

$Zn^{2+}(aq) + 2Fe^{2+}(aq) \rightarrow Zn(s) + 2Fe^{3+}(aq)$

IS NOT energetically feasible since its $\mathbf{E^\ominus_{cell}}$ **is negative** (-1.53V).

S16Use $E^\ominus_{cell}$ to decide the energetic feasibility of each of the following reactions under standard conditions:

(a) $Br_2(aq) + 2Cl^-(aq) \rightarrow 2Br^-(aq) + Cl_2(aq)$
(b) $Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

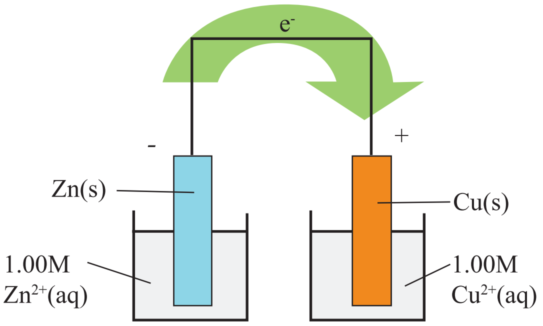
Data: $Cl_2 + 2e- \rightleftharpoons 2Cl^-$ +1.36V
$Br_2 + 2e- \rightleftharpoons 2Br^-$ +1.07V
$I_2 + 2e- \rightleftharpoons 2I^-$ +0.54V

For (a) Oxidant = **BROMINE / CHLORIDE** ;
Reductant = **BROMINE / CHLORIDE**.
Hence, $E^\ominus_{cell}$ = _________ - __________ = _________.
Hence reaction (a) **IS / IS NOT** energetically feasible under standard conditions.

For (b) Oxidant = **CHLORINE / IODIDE** ;
Reductant = **CHLORINE / IODIDE**.
Hence, $E^\ominus_{cell}$ = _________ - __________ = _________.
Hence reaction (b) **IS / IS NOT** energetically feasible under standard conditions.

6. DEDUCE A CELL REACTION

This is the **chemical reaction responsible for the current when the electrodes of an electrochemical cell are connected by a wire**.



Consider the following half reactions being used to construct a cell as shown.
$Cu^{2+}(aq) + 2e- \rightleftharpoons Cu(s)$; $E^\ominus$ = +0.34V
$Zn^{2+}(aq) + 2e- \rightleftharpoons Zn(s)$; $E^\ominus$ = -0.76V

Since the zinc electrode potential is more negative than copper, electrons are released at the zinc electrode and travel round the circuit to the positive copper electrode.

This means zinc atoms change to zinc ions at the negative electrode.
i.e. **oxidation occurs at the negative electrode** :
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e-$ ——(1)

and (b) copper ions change to copper atoms at the positive electrode
i.e. **reduction occurs at the positive electrode** :
$Cu^{2+}(aq) + 2e- \rightarrow Cu(s)$ ——(2)

> Oxidation occurs at the negative electrode but reduction occurs at the positive electrode. Hence the **half equation with the MORE NEGATIVE electrode potential is REVERSED** during the cell reaction.

Combining these half equations in appropriate proportions to balance the electrons then gives the required cell reaction.

Hence equations (1) and (2) combine 1:1 to give:

$\mathbf{Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)}$

S17 Deduce the cell reaction for the cell constructed using the 2 half-reactions shown below.
$Ag^+(aq) + e- \rightleftharpoons Ag(s)$ +0.80V
$Fe^{3+}(aq) + e- \rightleftharpoons Fe^{2+}(aq)$ + 0.77V
These data show the **more negative** electrode is **Ag / Fe** electrode.
Hence, this half reaction is reversed giving ____________________
This combined with the other half reaction in the ratio __________to give the cell reaction:______________________________________

S18 Deduce the cell reaction for the cell constructed using the 2 half-reactions shown below.

$V^{3+}(aq) + e- \rightleftharpoons V^{2+}(aq)$ - 0.26 A
$VO^{2+}(aq) + 4H^+(aq) + 2e- \rightleftharpoons V^{3+}(aq) + 2H_2O(l)$ + 1.00 B

These data show the more negative electrode is electrode A / B.
Hence, this half reaction is reversed giving ___________________

This is combined with the other half reaction in the ratio __________ to give the cell reaction:

____

S19 Deduce the cell reaction for the cell constructed using the 2 half-reactions shown below.
$Fe^{3+}(aq) + e- \rightleftharpoons Fe^{2+}(aq)$ + 0.77V A
$MnO_4^-(aq) + 8H^+(aq) + 5e- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$ + 1.51V B

These data show the more negative electrode is electrode A / B.

Hence, this half reaction is reversed giving ___________________

This is combined with the other half reaction in the ratio __________ to give the cell reaction:

____

7. PREDICT & EXPLAIN DISPROPORTIONATION REACTIONS

Disproportionations are reactions involving the simultaneous oxidation and reduction of the same species.

e.g. $2V^{3+}(aq) + 2OH^- \rightarrow V^{2+} + VO^{2+} + H_2O$
This is a disproportionation because V^{3+} is simultaneously oxidised (+3 to +4) to VO^{2+} and reduced (+3 to +2) to V^{2+}.

For a disproportionation to occur, there are 2 requirements:
(a) an intermediate oxidation state must exist.
e.g. V(+3) between V(+2) and V(+4) in the above example.
(b) the electrode potential for the oxidation from intermediate oxidation state to the higher oxidation state must be more negative than the electrode potential for the reduction from intermediate oxidation state to the lower oxidation state.

e.g. $Cu^{2+}(aq) + e– \rightleftharpoons Cu^+(aq)$; $E^\ominus = + 0.15V$ A
$Cu^+(aq) + e– \rightleftharpoons Cu(s)$; $E^\ominus = + 0.52V$ B
Here the intermediate oxidation state is Cu^+.

The electrode potential for half reaction A is more negative than B. Hence **half reaction A will be reversed** giving
$Cu^+(aq) \rightarrow Cu^{2+}(aq) + e–$.

Combining this with half reaction B gives the equation for the disproportionation i.e. $2Cu^+(aq) \rightarrow Cu(s) + Cu^{2+}(aq)$.
Furthermore, $E^\ominus$cell = $E^\ominus$oxidant - $E^\ominus$reductant
= $E^\ominus$more +ve - $E^\ominus$more –ve
= (+0.52) - (+0.15) = + 0.37V.

This positive $E^\ominus$cell value confirms this disproportionation is energetically feasible under standard conditions.

S20 Use the following data to predict whether chlorine will or will not disproportionate in acid or alkaline solution.
$2ClO^-(aq) + 2H_2O(l) + 2e- \rightleftharpoons Cl_2(aq) + 4OH^-(aq)$; +0.95V A
$Cl_2(aq) + 2e- \rightleftharpoons 2Cl^-(aq)$; +1.36V B
$2ClO^-(aq) + 4H^+(aq) + 2e- \rightleftharpoons Cl_2(aq) + 2H_2O(l)$; +1.59V C

1. In acid conditions, half reactions B and C need to be used since H^+ ions are involved in C.
Of these, B / C is the more negative and should be reversed giving:

____

Combining this with the other half-equation gives:

____

This IS / IS NOT a disproportionation of chlorine.

2. In alkaline conditions, half reactions A and B need to be used since OH- ions are involved in A.
Of these, A / B is the more negative and should be reversed giving:

____

Combining this with the other half-equation gives:

____

This IS / IS NOT a disproportionation of chlorine.

8. DEDUCE THE EFFECTS OF CONCENTRATION CHANGES ON $E^\ominus$ AND $E^\ominus$cell VALUES

This can be done using Le Chatelier's Principle.
Consider a general half equation: Oxidant + ne- $\rightleftharpoons$ Reductant ---(A)

S21 If the oxidant concentration is increased, equilibrium (A) will shift to the LEFT / RIGHT. Hence MORE / LESS electrons will be taken from the measuring electrode. This will make the measured electrode potential MORE / LESS positive.

If the oxidant concentration is decreased, equilibrium (A) will shift to the LEFT / RIGHT. Hence MORE / LESS electrons will be taken from the measuring electrode. This will make the measured electrode potential MORE / LESS positive.

If the reductant concentration is increased, equilibrium (A) will shift to the LEFT / RIGHT. Hence MORE / LESS electrons will be given to the measuring electrode. This will make the measured electrode potential MORE / LESS negative.

If the reductant concentration is decreased, equilibrium (A) will shift to the LEFT / RIGHT. Hence MORE / LESS electrons will be given to the measuring electrode. This will make the measured electrode potential MORE / LESS negative.

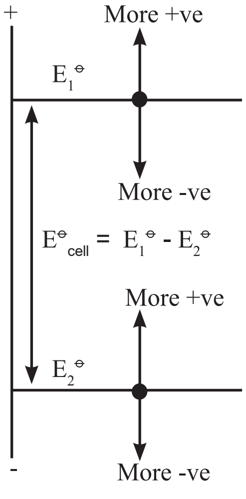
For example:
$MnO_4^-(aq) + 8H^+(aq) + 5e- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$; $E^\ominus = + 1.51V$

This electrode potential (+1.51V) can be made:
1. more positive (less negative) by increasing / decreasing $[MnO_4^-(aq)]$
2. more positive (less negative) by increasing / decreasing $[H^+(aq)]$
3. more positive (less negative) by increasing / decreasing $[Mn^{2+}(aq)]$
4. more negative (less positive) by increasing / decreasing $[MnO_4^- (aq)]$
5. more negative (less positive) by increasing / decreasing $[H^+(aq)]$
6. more negative (less positive) by increasing / decreasing $[Mn^{2+}(aq)]$.

But how will these effects change any $E^\ominus$cell value?

S22 Referring to your S21 answers and the diagram below:

(a) increasing [oxidant] for the E_1 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(b) decreasing [oxidant] for the E_1 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(c) increasing [reductant] for the E_1 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(d) decreasing [reductant] for the E_1 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(e) increasing [oxidant] for the E_2 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(f) decreasing [oxidant] for the E_2 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(g) increasing [reductant] for the E_2 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.
(h) decreasing [reductant] for the E_2 system will increase / decrease E_{cell} relative to $E^\ominus_{cell}$.



S23 Now apply this to :
$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Cu(s) + Zn^{2+}(aq)$; $E^{\ominus}_{cell} = +1.51V$.

In order to increase the value of $E^{\ominus}_{cell}$,
$[Cu^{2+}(aq)]$ should be INCREASED / DECREASED.

In order to decrease the value of $E^{\ominus}_{cell}$,
$[Cu^{2+}(aq)]$ should be INCREASED / DECREASED.

ANSWERS TO SELF-ASSESSED QUESTIONS

S1 more NEGATIVE -- .

S2 more POSITIVE -- .

S3 (a) Acidified MnO_4^- (b) Zn^{2+} (c) Zn (d) Mn^{2+}

S4 $Zn^{2+}(aq) < V^{3+}(aq) <$ acidified $SO_4^{2-}(aq) <$ acidified $VO_{-}^{+}(aq)$

S5 $Mn^{2+}(aq) < Fe^{2+}(aq) < V^{3+}(aq) < Zn(s)$

S6 A's $E^{\ominus}$ value is more NEGATIVE than B's

S7 B's $E^{\ominus}$ value is more POSITIVE than A's.

S8 WILL reduce iron(III) because its Eê value is MORE negative

S9 WILL NOT reduce vanadium(III) because its $E^{\ominus}$ value is LES negative --

S10 WILL NOT oxidise $VO^{2+}(aq)$ because its $E^{\ominus}$ value is LESS positive --

S11 WILL oxidise $V^{2+}(aq)$ because its $E^{\ominus}$ value is MORE positive

S12 zinc will change to $Zn^{2+}(aq)$ and the iron(III) will change to $Fe^{2+}(aq)$.

S13 $VO^{2+}(aq)$ will change to $VO^{2+}(aq)$ and the $V^{2+}(aq)$ will change to $V^{3+}(aq)$.

S14 $VO^{2+}(aq)$ -- $V^{3+}(aq)$ -- $V^{2+}(aq)$

S15 $Fe^{2+}(aq)$ is the reductant since it gives electrons to the $Zn^{2+}(aq)$.
$Zn^{2+}(aq)$ is the oxidant since it takes electrons from the $Fe^{2+}(aq)$.
$E^{\ominus}cell = -0.76 - (+0.77) = -1.53V$.

S16 For (a) Oxidant = BROMINE ; Reductant = CHLORIDE.
Hence, $E^{\ominus}cell = +1.07 - (+1.36) = -0.29V$.
Hence reaction (a) IS NOT energetically feasible under standard conditions.
For (b) Oxidant = CHLORINE ; Reductant = IODIDE.
Hence, $E^{\ominus}cell = +1.36 - (+0.54) = +0.82V$.
Hence reaction (b) IS energetically feasible under standard conditions.

S17 These data show the more negative electrode is Fe electrode.
half reaction is reversed giving $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e-$
combined with the other half reaction in the ratio 1:1 –
the cell reaction: $Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$

S18 These data show the more negative electrode is electrode A.
Hence, this half reaction is reversed giving $V^{2+}(aq) \rightarrow V^{3+}(aq) + e-$
combined with the other half reaction in the ratio 1:2 –
the cell reaction:
$VO^{2+}(aq) + 4H^{+}(aq) + 2V^{2+}(aq)\ V^{3+}(aq) + 2H_2O(l)$

S19 These data show the more negative electrode is electrode A.
Hence, this half reaction is reversed giving $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e-$
combined with the other half reaction in the ratio 1:5 –
$MnO_4^-(aq) + 8H^{+}(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$

S20 1. Of these, B is the more negative and should be reversed giving:
$2Cl^-(aq) \rightarrow Cl_2(aq) + 2e-$
$2Cl^-(aq) + 2ClO^-(aq) + 4H^{+}(aq) \rightarrow 2Cl_2(aq) + 2H_2O(l)$
This IS NOT a disproportionation of chlorine.
2. Of these, A is the more negative and should be reversed giving:
$Cl_2(aq) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 2H_2O(l) + 2e-$
$2Cl_2(aq) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 2Cl^-(aq) + 2H_2O(l)$
This IS a disproportionation of chlorine.

S21 oxidant concentration is increased -- RIGHT -- MORE -- MORE positive.
oxidant concentration is decreased -- LEFT -- LESS -- LESS positive.
reductant concentration is increased – LEFT – MORE -- MORE negative.
reductant concentration is DEcreased – RIGHT – LESS -- LESS negative.
1. more positive (less negative) by INCREASING $[MnO_4^-(aq)]$
2. more positive (less negative) by INCREASING $[H^{+}(aq)]$
3. more positive (less negative) by DECREASING $[Mn^{2+}(aq)]$
4. more negative (less positive) by DECREASING $[MnO_4^-(aq)]$
5. more negative (less positive) by DECREASING $[H^{+}(aq)]$
6. more negative (less positive) by INCREASING $[Mn^{2+}(aq)]$.

S22 Referring to your S21 answers and the diagram opposite:
(a) increasing [oxidant] for the E1 system will INCREASE E_{cell}
(b) decreasing [oxidant] for the E1 system will DECREASE E_{cell}
(c) increasing [reductant] for the E1 system will DECREASE E_{cell}
(d) decreasing [reductant] for the E1 system will INCREASE E_{cell}
(e) increasing [oxidant] for the E2 system will DECREASE E_{cell}
(f) decreasing [oxidant] for the E2 system will INCREASE E_{cell}
(g) increasing [reductant] for the E2 system will INCREASE E_{cell}
(h) decreasing [reductant] for the E2 system will DECREASE E_{cell}

S23 In order to increase the value of $E^{\ominus}cell$, $[Cu^{2+}(aq)]$ should be INCREASED.
In order to decrease the value of $E^{\ominus}cell$, $[Cu^{2+}(aq)]$ should be DECREASED.

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