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VCE Chemistry ¾
Spontaneous Redox Reactions [1.7]
Workbook

Outline:

Learning Objectives:

- ☐ **CH34 [1.7.1]** – Apply the ECS to predict spontaneous reactions.
- ☐ **CH34 [1.7.2]** – Identify differences between direct & indirect redox reactions, & features of ECS.
- ☐ **CH34 [1.7.3]** – Find the strongest oxidants/reductants by constructing your own ECS.

Section A: Predicting Spontaneous Redox Reactions

Sub-Section: Recap

Active Recall: What is the difference between reduction and oxidation?

Reduction Reaction	Oxidation Reaction
Involves the **[gain]/[loss]** of electrons.	Involves the **[gain]/[loss]** of electrons.

Active Recall: What is the chemical symbol for each of the following metals?

- Lead: _____
- Tin: _____
- Silver: _____
- Iron: _____
- Sodium: _____
- Cobalt: _____
- Magnesium: _____
- Manganese: _____

Question 1 Walkthrough.

Thiosulphate ions can react and turn into sulphate ions in **basic** conditions.

a. Write the balanced equation for this reaction.

_____

_____

_____

b. The type of reaction which occurs is a **[reduction] / [oxidation]** reaction.

Question 2

For the following redox reactions, assume they occur in **acidic** conditions.

a. Balance the following half-equations and indicate if they are reduction or oxidation.

i. Zinc metal ($Zn(s)$) turning into zinc ions ($Zn^{2+}(aq)$).

_____

Type of Reaction: **[reduction]/[oxidation]** reaction.

ii. Permanganate $\left(MnO_4^-(aq)\right)$ turning into manganese ions ($Mn^{2+}(aq)$).

_____

_____

Type of Reaction: **[reduction]/[oxidation]** reaction.

b. Given the above half-equations, write the overall reaction which takes place.

_____

_____

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Sub-Section: Net Ionic Equations

Exploration: Net Ionic Equations

- Consider the following reaction:

$$\mathbf{Ni(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Ni(s)}$$

- Rewrite all species as their cation and anion separately (*Label Below)*

$$\underline{\qquad\qquad\qquad\qquad} \mathbf{Zn(s)} \rightarrow \underline{\qquad\qquad\qquad\qquad\qquad\qquad} \mathbf{Ni(s)}$$

- **What is the unchanged species?** _____.
- **Unchanged Species Alternative Name:** _____.
- When writing **net ionic equations** or fully balanced equations, the **spectator ions can be omitted.**
- **Net Ionic Equation:**

Definition: Net Ionic Equations

- **Definition**: A balanced full equation with spectator ions omitted.
- **Spectator Ion**: Compound which is present but does not participate in the reaction.
- **Polyatomic Ions Data Book**: Page 6-7.

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Let's have a look at a question together!

Question 3 Walkthrough.

a. Given the following equation, balance the equation.

$$Cr_2(SO_4)_3(aq) + (NH_4)_2CO_3(aq) \rightarrow Cr_2(CO_3)_3(s) + (NH_4)_2SO_4(aq)$$

_____

_____

b. Identify any spectator ions and rewrite the net ionic equation.

_____

_____

Question 4 Walkthrough.

Write $(NH_4)_2Cr_2O_7(aq)$ in its ionised form.

_____

TIP: To find spectator ions, look for anything that remains aqueous (aq) in state before and after! If something changes state, it is not a spectator ion!

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Your turn!

Question 5

Write $Al_2(CO_3)_3(aq)$ in its ionised form.

_____

Question 6

Write $BaS_2O_3(aq)$ in its ionised form.

_____

Question 7

For each of the following, balance the equation and identify any spectator ions before writing the net ionic equation.

a. $3NiCl_2(aq) + 2Li_3PO_4(aq) \rightarrow Ni_3(PO_4)_2(s) + 6LiCl(aq)$

i. Spectator Ions: _____.

ii. Net Ionic Equation: _____.

b. $2AgNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + Ag_2CO_3(s)$

i. Spectator Ions: _____.

ii. Net Ionic Equation: _____.

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Question 8 Additional Question.

For the following, name all reactants/products, balance the equation, and identify any spectator ions before writing the net ionic equation.

$$2H_2O(l) + Ni(s) \rightarrow Ni(OH)_2(aq) + H_2(g)$$

i. Spectator Ions: _____.

ii. Net Ionic Equation: _____.

Space for Personal Notes

Sub-Section: The Electrochemical Series

Exploration: Oxidant and Reductant Strength

- Consider reactions involving sodium solid $(Na(s))$ and sodium ions $(Na^+(aq))$.

$Na^+(aq) + e^- \rightarrow Na(s)$	$Na(s) \rightarrow Na^+(aq) + e^-$
[reduction] / [oxidation] Reaction	[reduction] / [oxidation] Reaction
[oxidant] / [reductant]	[oxidant] / [reductant]

- Electron configurations:

Na^+	Na
Electron configuration: _____	Electron configuration: _____
[stable] / [unstable] Form	**[stable] / [unstable] Form**

- Likelihood of occurring naturally:

$Na^+(aq) + e^- \rightarrow Na(s)$	$Na(s) \rightarrow Na^+(aq) + e^-$
[more] / [less] Likely to occur naturally.	[more] / [less] Likely to occur naturally.

- Substances can be ranked based on their oxidising/reducing _____.
  - Sodium ion (Na^+): _____.
  - Sodium metal (Na): _____.
- Using this, it can be predicted whether a reaction will happen **spontaneously**.

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Discussion: How do we know that the oxidising/reducing strength of all chemicals relative to each other?

NOTE: The electrochemical series can be found on **page 2 of the databook.**

ALSO NOTE: The electrochemical series ranks substances in terms of increasing oxidant/reductant strength.

Let's have a look at the Electrochemical Series on the next page first!

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Databook: The Electrochemical Series

Reaction	Standard electrode potential (E^0) in volts at 25 °C
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$PbO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2H_2O(l)$	+1.47
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.36
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

What do the double arrows mean?

Exploration: Writing Half-Equations

- Consider the copper half-equation:

$$\mathbf{Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)}$$

- **Forwards Reaction:**
  - Type of Reaction: [reduction] / [oxidation]
  - Forwards Reaction Written: _____
- **Reverse Reaction:**
  - Type of Reaction: [reduction] / [oxidation]
  - Reverse Reaction Written: _____

NOTE: When writing half-equations, use single arrows (→).

ALSO NOTE: The oxidation reaction has to be written _____ around!

Exploration: Standard Electrode Potential

- The **standard electrode potential** will be covered in depth later on and is a way to measure the **electromotive force (EMF).**
- Species with a standard electrode potential of zero (0.00): _____
- **Name:** _____
- The E^0 values of all other species are made relative to that.

The Electrochemical Series

Reduction Reaction	Oxidation Reaction
[forward] / [reverse] Reaction	[forward] / [reverse] Reaction

Oxidants	Reductants
Positioned on the [left] / [right] side	Positioned on the [left] / [right] side

Strongest Oxidants	Strongest Reductants
Positioned [top] / [bottom] - [left] / [right]	Positioned [top] / [bottom] - [left] / [right]

- Standard Electrode Potential Definition: Method to measure electromotive force (EMF).
- Standard Hydrogen Electrode (SHE): $\mathrm{H^+(aq)/H_2(g)}$ which has $E^0 = 0.00\ V$.

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Key Takeaways

- ☑ Net Ionic Equation Definition: A balanced full equation with spectator ions omitted.
- ☑ **Spectator Ion**: Compound which is present but does not participate in the reaction..

Reduction Reaction	Oxidation Reaction
[forward] / [reverse] Reaction	[forward] / [reverse] Reaction

Oxidants	Reductants
Positioned on the [left] / [right] Side	Positioned on the [left] / [right] Side

Strongest Oxidants	Strongest Reductants
Positioned [top] / [bottom] - [left] / [right]	Positioned [top] / [bottom] - [left] / [right]

- ☑ Standard Electrode Potential Definition: Method to measure electromotive force (EMF).
- ☑ Standard Hydrogen Electrode (SHE): $\mathrm{H^+(aq)/H_2(g)}$ which has $E^0 = 0.00\,V$.

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Section B: Spontaneous Reactions

Sub-Section: Introduction to Spontaneous Reactions

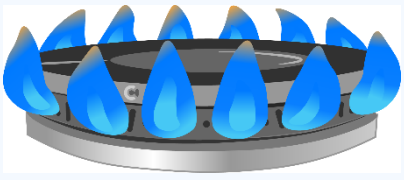
Context

- We use the electrochemical series to predict whether a _____reaction will occur or not.

What is a spontaneous reaction?

Spontaneous Reactions

- **Definition**: A **spontaneous reaction** is a reaction that happens _____ without being forced if the reactants are in contact with each other.
- **Example**:

Combustion	Acid-Base Reactions
Methane burning	Vinegar and baking soda
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$	$CH_3COOH + NaHCO_3 \rightarrow CH_3COONa + CO_2 + H_2O$
	

- **Example of Spontaneous Reaction:**
  - **Link (Video 1):** https://youtube.com/playlist?list=PL2hVhOVQpiYQ9qBzM2ob1xkc-eUSIb66a&si=d5RskLFh7K8gOIm5

Non-Spontaneous Reaction

- **Definition**: A non-spontaneous reaction only occurs when energy is inputted.
- For a non-spontaneous reaction, if **energy is not inputted**, **will a reaction occur?** [yes] / [no]
- **Example of non-spontaneous reaction:**
  - **Link (Video 2):** https://youtube.com/playlist?list=PL2hVhOVQpiYQ9qBzM2ob1xkc-eUSIb66a&si=d5RskLFh7K8gOIm5

NOTE: We'll cover these non-spontaneous reactions in depth in the next AOS in Electrolysis!

Active Recall: Where is the strongest oxidant and strongest reductant located on the electrochemical series?

Strongest Oxidants	Strongest Reductants
Positioned [top] / [bottom] - [left] / [right]	Positioned [top] / [bottom] - [left] / [right]

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Exploration: Spontaneous vs Non-Spontaneous Reaction

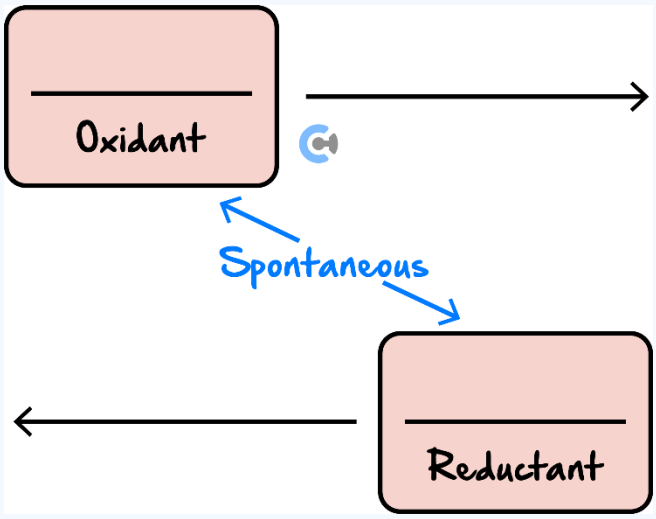
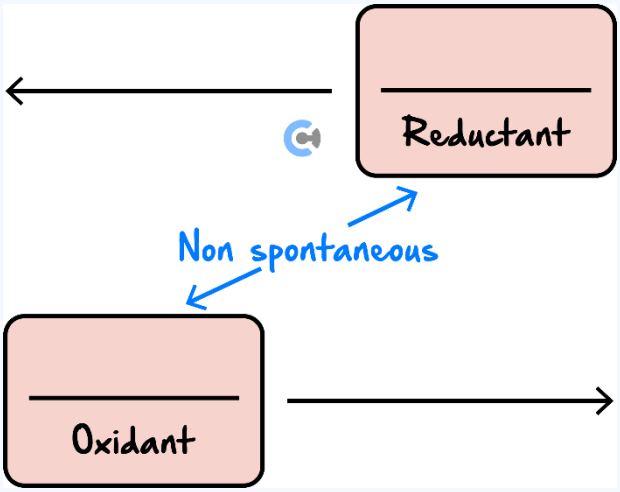
Spontaneous Reactions

Reaction
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$
$PbO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2H_2O(l)$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$
$Na^+(aq) + e^- \rightleftharpoons Na(s)$
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$
$K^+(aq) + e^- \rightleftharpoons K(s)$
$Li^+(aq) + e^- \rightleftharpoons Li(s)$

Non-Spontaneous Reactions

Reaction
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$
$PbO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2H_2O(l)$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$
$Na^+(aq) + e^- \rightleftharpoons Na(s)$
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$
$K^+(aq) + e^- \rightleftharpoons K(s)$
$Li^+(aq) + e^- \rightleftharpoons Li(s)$

Spontaneous Redox Reaction

Spontaneous Redox Reaction	Non-Spontaneous Redox Reaction
	

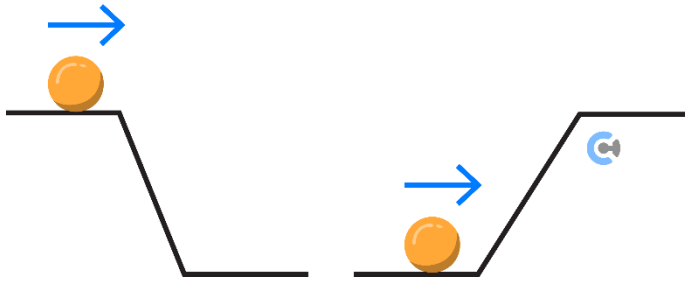
How can we remember this?

TIP: Spontaneous Reaction

$$S - Shape: for\ Spontaneous$$

Analogy: Rolling Ball

- Imagine a ball rolling to the right:



- How will the ball naturally roll? [uphill] / [downhill]

Sub-Section: Predicting Simple Spontaneous Reactions

How do we predict spontaneous reactions?

Exploration: Predicting Spontaneous Reactions I

- Consider a potential reaction which may occur between nickel sulphate $(NiSO_4(aq))$ and zinc metal $(Zn(s))$.
- All ions/species which are present: _____
- Where are these species on the electrochemical series? *(Label Below)*

Reaction	Standard Electrode Potential ($E°$) in volts at 25°C
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn\ (s)$	-0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni\ (s)$	-0.25
$Co^{2+}\ (aq) + 2e^- \rightleftharpoons Co(s)$	-0.28
$Fe^{2+}\ (aq) + 2e^- \rightleftharpoons Fe\ (s)$	-0.44
$Zn^{2+}\ (aq) + 2e^- \rightleftharpoons\ Zn\ (s)$	-0.76
$2H_2O\ (l) + 2e^- \rightleftharpoons H_2\ (g) + 2OH^-(aq)$	-0.83

- Does a reaction occur spontaneously? [yes] / [no]
- Half-Equations:
  - Reduction: _____.
  - Oxidation: _____.

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REMINDER

- The oxidation reaction will usually be _____ around!

Exploration: Predicting Spontaneous Reactions II

- Consider a potential reaction which may occur between zinc sulfate ($ZnSO_4(aq)$) and nickel metal ($Ni(s)$).
- All ions/species which are present: _____
- Where are these species on the electrochemical series? *(Label Below)*

Reaction	Standard Electrode Potential ($E°$) in volts at 25°C
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn\ (s)$	-0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni\ (s)$	-0.25
$Co^{2+}\ (aq) + 2e^- \rightleftharpoons Co(s)$	-0.28
$Fe^{2+}\ (aq) + 2e^- \rightleftharpoons Fe\ (s)$	-0.44
$Zn^{2+}\ (aq) + 2e^- \rightleftharpoons Zn\ (s)$	-0.76
$2H_2O\ (l) + 2e^- \rightleftharpoons H_2\ (g) + 2OH^-(aq)$	-0.83

- Does a reaction occur spontaneously? [yes] / [no]

TIP: Draw a vertical line to split oxidants and reductants apart!

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Predicting Spontaneous Reactions Steps

1. Split all species into cations and anions. Some cations/anions are inert.
2. Locate all species on the electrochemical series. Draw a vertical line to split oxidants and reductants apart.
3. Draw a mini-electrochemical series version.
4. Find the strongest oxidant (top-left) and strongest reductant (bottom-right).
5. Check for the downhill gradient.
6. Write out half-equations.

Let's look at some questions together!

Question 9 Walkthrough.

Nickel chloride is mixed with lead metal. Predict if the reaction is spontaneous or non-spontaneous. If spontaneous, write the half-equations out.

NOTE: Metal by itself → solid, Metal as part of an ionic compound → aqueous.

TIP: When figuring out what reaction will occur, list all the species/ions present and locate them on the electrochemical series. Draw out your own 'mini' electrochemical series.

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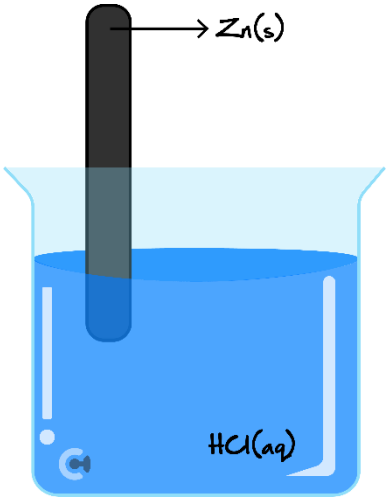
Question 10 Walkthrough.

Zinc metal sheet is dipped into a solution of hydrochloric acid $(HCl(aq))$.



a. Write the half-equation which occurs for the:

i. Reduction reaction.

_____

ii. Oxidation reaction.

_____

b. Write the balanced equation for the overall reaction.

_____

_____

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Your turn!

Question 11

A solution of nickel nitrate has zinc metal dipped inside of it.

a.

i. Write the reduction reaction which takes place.

_____

ii. Write the oxidation reaction which takes place.

_____

b. Write the full balanced ionic equation.

_____

_____

Question 12

A cobalt rod is dipped into a solution of lead sulphate.

a.

i. Write the reduction reaction which takes place.

_____

ii. Write the oxidation reaction which takes place.

_____

b. Write the full balanced ionic equation.

_____

_____

Question 13

For each of the following, circle to indicate whether a reaction will occur or not. If there is a reaction, write the relevant reduction and oxidation reactions.

a. A solid lead rod dipped into a solution containing silver nitrate.

Spontaneous Reaction: [occurs] / [does not occur]

b. A solution containing nickel nitrate is mixed with tin metal.

Spontaneous Reaction: [occurs] / [does not occur]

c. A solution containing hydrochloric acid (HCl) is mixed with a strip of iron metal.

Spontaneous Reaction: [occurs] / [does not occur]

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Question 14 Additional Question.

Which of the following metals could be used to make a container to store an aqueous tin (II) sulphate solution?

A. Zn

B. Pb

C. Ni

D. Fe

Question 15 Additional Question.

Use the electrochemical series to determine whether the following reaction is a spontaneous redox reaction, giving justification for your reasoning.

$$\mathbf{2Ag^{+}(aq) + 2Br^{-}(aq) \longrightarrow 2Ag(s) + Br_2(l)}$$

NOTE: Sometimes, an acid such as Hydrochloric acid (HCl), Nitric acid $(\mathrm{HNO_3})$ or Sulphuric acid $(\mathrm{H_2SO_4})$ is used. They all contain $\mathrm{H^+}$ ions, which are located at an E^0 value of $0\ V$.

ALSO NOTE: Generally, the metal solids are located on the right side of the electrochemical series, whereas metal ions are located on the left side of the electrochemical series.

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REMINDER: Don't forget when making an overall balanced equation from the two half-equations, be sure to **cancel out the electrons!**

Misconception

"Any of the following equations can be used if there is hydrogen ions (H^+) present."

Reaction	Standard electrode potential (E^0) in volts at 25°C
$H_2O_2(l) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00

TRUTH: Only one of them can be used!

Space for Personal Notes

Sub-Section: Unique Species on the Electrochemical Series

Context

- There are some weird species on the electrochemical series which we'll cover right now!

Exploration: Chemical Reaction

- Consider a solution which contains silver nitrate and tin (II) nitrate.
- Species Present: _____
- Where are they on the electrochemical series? *(Label Below)*

Reaction
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$

- **Reduction Reaction:** _____

- **Oxidation Reaction:** (Double Check: Are we missing any equations?)

_____

Misconception

"All metal ions only exist on the left side of the electrochemical series."

TRUTH:
"While most of the metal ions do exist on the left side of the electrochemical series, there are some exceptions to this rule!"

NOTE: The four exceptions to this rule which is on VCAA's electrochemical series are:

- _____
- _____
- _____
- _____

Exploration: Charges on Ions

- Some ions exist in multiple different forms such as $Fe^{3+}(aq)$ and $Fe^{2+}(aq)$.
- There are **two** ways to figure out the charge on the ion.
  - **Roman Numerals**
  - **Ionic Compound**
- **Roman Numerals:**

Iron (II) nitrate	Iron (III) nitrate

Ionic Compound:

What is the charge on a nitrate? _____

$Fe(NO_3)_2$	$Fe(NO_3)_3$

NOTE: The **number of anions** and their **charge** indicate the charge on the metal cation.

Try some questions

Question 16

Predict the reaction that will occur between $SnCl_4$ and tin metal. If a reaction is expected to occur, write the full balanced ionic equation. If not, justify why no reaction will occur.

a.

i. Write the reduction reaction which takes place.

_____

ii. Write the oxidation reaction which takes place.

_____

b. Write the overall reaction.

_____

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Question 17

Chlorine gas is pumped into a solution containing iron (II) nitrate. Write the full balanced equation which occurs.

_____

_____

_____

_____

Question 18

Hydrogen peroxide (H_2O_2), which is a chemical commonly used for cleaning, is placed in a beaker, and it is found to react with itself.

a. Write the half equation for the:

i. Reduction reaction.

_____

ii. Oxidation reaction.

_____

b. Write the overall reaction that takes place.

_____

_____

Question 19 Additional Question.

Fluoride acid, HF(aq), is poured onto a bar of manganese. Write the full balanced equation for the reaction which occurs.

Space for Personal Notes

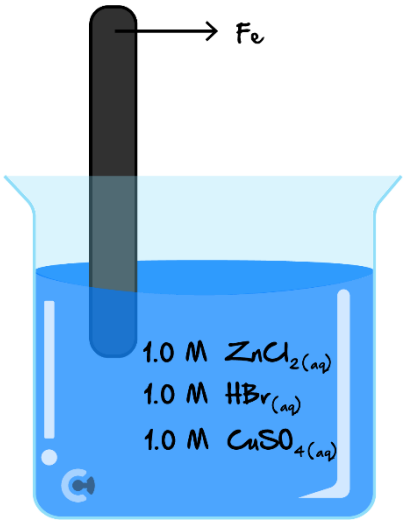
Sub-Section: Multiple Oxidants and Reductants

REMINDER: Don't forget that the strongest oxidant always reacts with the strongest reductant!

Let's have a look at a question together!

Question 20 Walkthrough.

An iron rod dipped into a 1.0 M solution containing zinc chloride, hydrogen bromide, and copper sulphate.



Write the full balanced half-equations that occur.

a. Oxidation reaction.

_____

_____

b. Reduction reaction.

_____

_____

NOTE: Sometimes, there are some species that do not react at all because there are even stronger oxidants/reductants!

Your turn!

Question 21

A mixture of lead (II) nitrate, zinc nitrate, and copper nitrate are mixed together. A nickel coin is then dipped into the mixture. Write the two half-equations which are expected to occur.

a. Reduction half-equation:

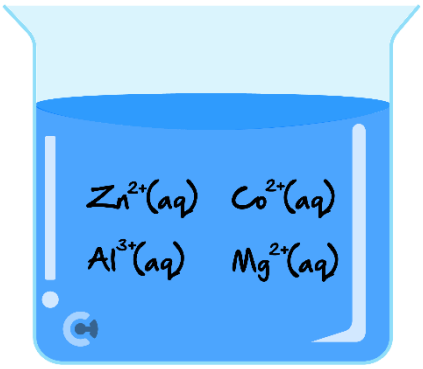
_____

b. Oxidation half-equation:

_____

Question 22

A beaker happens to contain zinc, cobalt, aluminium, and magnesium ions.



An iron bar is dipped into the beaker of water.

a. Will the iron bar reduce or oxidise? Write the half-reaction out.

_____

b. Write the other half-equation that occurs.

_____

Question 23

A room happens to have fluorine and chlorine gas mixed together.

a. Will a reaction take place? Justify your answer.

_____

_____

b. Hydrogen gas is now added to the mixture. Write the overall reaction that takes place.

_____

_____

_____

_____

NOTE: Whenever there are ion reactants and products that can be combined into a neutral compound, try to combine them! VCAA is slowly forcing students to write these out!

Space for Personal Notes

Question 24 Additional Question.

A mixture of iron (II) chloride and sodium chloride are mixed together with bromine liquid. Write the overall reaction that takes place.

_____

_____

_____

_____

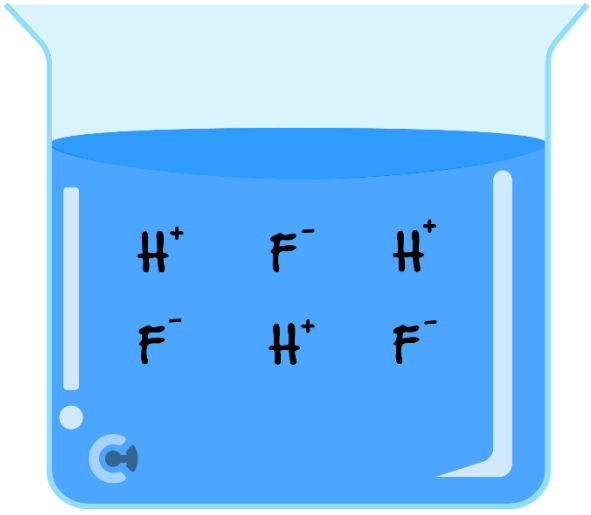
NOTE: While the products have cations and anions, they cannot be combined to form a neutral compound, and thus, they do not need to be combined here!

Exploration: Ion Reactants and Products

- Imagine the following reaction:

$$H_2(g) + F_2(g) \rightarrow 2H^+(aq) + 2F^-(aq)$$

- Is there a way to simplify this? What do the products look like in a chemical beaker?



$$H_2(g) + F_2(g) \rightarrow \_\_\_\_\_$$

Sub-Section: Energy

Energy Conversions During Spontaneous Direct Contact Reaction

- Are these reactions **endothermic** or **exothermic** process? [endothermic] / [exothermic]
- Type of Energy Released: _____
- Energy Conversion: _____

Strongest Oxidant	Strongest Reductant
[highest] / [lowest] E^0 Value	[highest] / [lowest] E^0 Value

NOTE: Indirect redox reactions are also known as _____ which will be covered in the next booklet!

Energy Conversions

Direct Contact Spontaneous Redox Reaction	Indirect Contact Spontaneous Redox Reaction

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Sub-Section: Rate of Reaction

Rate of Reaction

- **Definition:** How quickly the reaction occurs!

Exploration: Reactions with Acid (H^+)

- Consider some reactions with hydrochloric acid (H^+).
- **Which metals will react with H^+ spontaneously?** *(Label Y/N in First Row Below)*

Metals	Sn	Fe	Mg	Cu	Zn	Ca
Reaction? (Theoretically) (Y/N)						
Reaction? Experimentally (Y/N)						

- Watch the video below. Does a reaction occur for all of them? *(Label Y/N in Second Row Above)*
- **Link (Video 4):** Play the video at 2x speed.

https://youtube.com/playlist?list=PL2hVhOVQpiYQ9qBzM2ob1xkc-eUSIb66a&si=13HOf03f0Arw-Lri

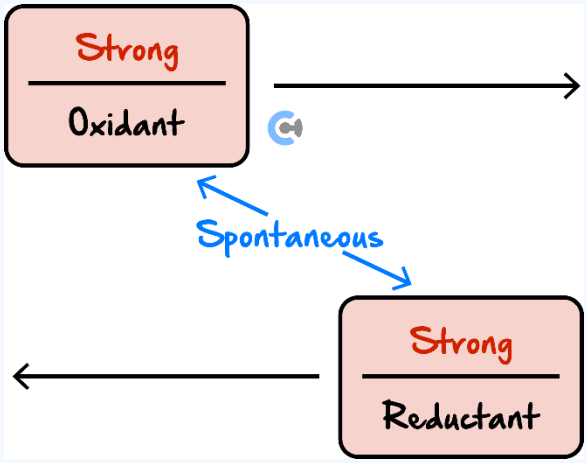
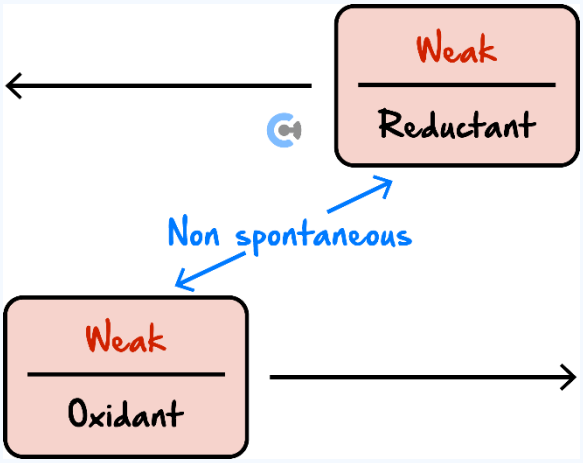
- Why does a reaction seem to not occur for some of the metals?
  - _____
  - _____

Rate of Reaction using Electrochemical Series

- Sometimes a reaction is predicted to occur but, in reality, does not, as the electrochemical series _____.
- Strength of oxidant/reductant [does] / [does not] affect rate of reaction!

NOTE: The rate of reaction will be properly covered in U3 AOS 2!

Key Takeaways

Spontaneous Redox Reaction	Non-Spontaneous Redox Reaction
	

- ☑ Steps to predicting spontaneous reaction:
  1. Split all species into cations and anions. Some cations/anions are inert.
  2. Locate all species on the electrochemical series. Draw a vertical line to split oxidants and reductants apart.
  3. Draw a mini-electrochemical series version.
  4. Find the strongest oxidant on the top left and the strongest reductant on the bottom right.
  5. Check for the downhill gradient.
  6. Write out half-equations.
- ☑ When multiple oxidants/reductants are present, the strongest oxidant reacts with the strongest reductant.
- ☑ The electrochemical series does not predict the rate of reaction.

- The four ions which appear on both sides of the electrochemical series:
  - $Fe^{3+}(aq)/Fe^{2+}(aq)$
  - $Sn^{4+}(aq)/Sn^{2+}(aq)$
  - $Pb^{2+}(aq)/Pb(s)$
  - $Mn^{2+}(aq)/Mn(s)$

Direct Contact Spontaneous Redox Reaction	Indirect Contact Spontaneous Redox Reaction
Chemical → thermal	Chemical → electrical

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Section C: Constructing an Electrochemical Series

Context

- Sometimes, **equations** and/or reactants that can undergo reduction/oxidation **will be provided** but **do not appear on the electrochemical series** in the databook.
- Sometimes, the order of the electrochemical series must be _____.

Active Recall: What is the E^0 value for each of the following?

- Electrochemical series ordered from [lowest → highest] / [highest → lowest] E^0 value.

Strongest Oxidant	Strongest Reductant
[highest] / [lowest] E^0 value	[highest] / [lowest] E^0 value

Let's look at a question together!

TIP: Write the conjugate redox pair for each of them.

Space for Personal Notes

Question 25 Walkthrough.

The following table lists a selection of standard redox potentials.

Half-Reaction	E^0/V
$AuCl_4^-(aq) + 3e^- \rightleftharpoons Au(s) + 4Cl^-(aq)$	$+0.99$
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	-0.40
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	$+1.51$
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightleftharpoons NO(g) + 2H_2O(l)$	$+0.96$

a. From the table above, the strongest oxidant and the strongest reductant are:

Strongest Oxidant	Strongest Reductant

b.

i. Which set of the following chemical species is likely to produce a spontaneous reaction when mixed?

A. $Mn^{2+}(aq)$ and $AuCl_4^-(aq)$

B. $Cd^{2+}(aq)$ and $NO(g)$

C. $NO(g)$ and $AuCl_4^-(aq)$

D. $NO_3^-(aq)$, $H^+(aq)$ and $Cd(s)$

ii. Explain your answer from **part b.i.**

_____

_____

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TIP: Write the conjugate redox pair for each of them.

Your turn!

Question 26

The following information is provided:

$$Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s) \qquad -2.90\ V$$

$$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s) \qquad -0.40\ V$$

$$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s) \qquad -1.03\ V$$

$$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s) \qquad -0.28\ V$$

State the weakest oxidant and weakest reductant.

Weakest Oxidant	Weakest Reductant

Space for Personal Notes

Question 27

Consider the following conjugate redox pairs and their E^0 values.

$\mathbf{Cl_2(g) \,/\, Cl^-(aq)}$	$\mathbf{+1.36\,V}$
$\mathbf{I_2(s) \,/\, I^-\,(aq)}$	$\mathbf{+0.54\,V}$
$\mathbf{Al^{3+}(aq) \,/\, Al(s)}$	$\mathbf{-1.67\,V}$
$\mathbf{MnO_4^-(aq) \,/\, Mn^{2+}(aq)}$	$\mathbf{+1.52\,V}$
$\mathbf{Pb^{2+}(aq) \,/\, Pb(s)}$	$\mathbf{-0.13\,V}$

a. Which species is:

i. The strongest oxidising agent?

ii. The strongest reducing agent?

iii. The weakest oxidising agent?

iv. The weakest reducing agent?

b. Explain whether a reaction occurs between solid iodine and manganese (II) ions.

_____

_____

c. Explain whether a reaction occurs between aluminium metal and lead (II) ions.

_____

_____

Question 28 Additional Question.

Some standard electrode potential (E^0) of four redox pairs are provided below. On the basis of these values choose the correct option.

$$Br_2/Br^- = +1.09\,V$$

$$Cu^{2+}/Cu = +0.34\,V$$

$$Ag^+/Ag = +0.80\,V$$

$$I_2/I^- = +0.54\,V$$

A. Br^- is the strongest oxidant.

B. Ag will reduce Br_2.

C. Cu^{2+} will oxidise I^-.

D. Cu is the strongest oxidant.

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Sub-Section: Deriving an Electrochemical Series

Context

- Sometimes, equations not in the electrochemical series are provided.
- For these, we must **construct our own electrochemical series**.

Creating Electrochemical Series Yourself Steps

1. Draw a vertical line to separate oxidants and reductants.
2. Using information, place oxidants/reductants on this mini electrochemical series.

Spontaneous Reactions	Non-Spontaneous Reactions
[positive] / [negative] gradient	[positive] / [negative] gradient

Strong Oxidant	Weak Oxidant
[top] / [bottom] – [left] / [right]	[top] / [bottom] – [left] / [right]

3. Write the conjugate version of the oxidant/reductant.
4. Repeat for each piece of information.

Space for Personal Notes

Let's look at a question together!

Question 29 Walkthrough.

There are three unknown substances, P, Q and R. The following half-equations are given, but their E^0 values are not given.

Reaction
$P^{2+}(aq) + 2e^- \rightleftharpoons P(s)$
$Q^{2+}(aq) + 2e^- \rightleftharpoons Q(s)$
$R^{2+}(aq) + 2e^- \rightleftharpoons R(s)$

It is known that when Q is mixed into a solution of R^{2+}, a reaction begins to occur.

It is also known that when P is mixed into a solution containing R^{2+}, no reaction occurs.

Rank the three metals in terms of their decreasing oxidant strength.

- Decreasing oxidant strength: _____.

TIP: Ions are usually on the left, and metal solids are usually on the right.

Question 30 Walkthrough.

Cathy is given five metals and 1 M solutions of nitrates of the metals.

The metals are labelled M, N, O, P and Q, and the solutions are labelled M^{2+}, N^{2+}, O^{2+}, P^{2+} and Q^{2+}.

The student carries out a number of experiments and the results obtained are listed below.

i. Metal M reacts with N^{2+} spontaneously.

ii. When the solution O^{2+} has metals M, N, Q or P dipped inside of it, it coats itself around the metal.

iii. When metal P is dipped in a solution with all solutions, a reaction occurs with all of them except with metal M^{2+}.

iv. When a solution of $Q(NO_3)_2$ has metal N placed within it, no reaction is observed to occur.

Rank each of the 5 metals in order of decreasing E^0 values.

_____

_____

_____

_____

_____

Active Recall: What are the steps to create your own electrochemical series?

1. Draw a _____ to separate oxidants and reductants.
2. Using information, place _____ on this mini electrochemical series.
3. Write the _____ version of the oxidant/reductant.
4. Repeat for each piece of information.

Space for Personal Notes

Your turn!

Question 31

Melissa is similarly given five metals and 1 M solutions of nitrates of the metals.

The metals are labelled A, B, C, D and E, and the solutions are labelled A^{2+}, B^{2+}, C^{2+}, D^{2+} and E^{2+}.

The student carries out a number of experiments and the results obtained are listed below.

i. Metal A remains unchanged in all solutions.

ii. Metal C becomes coated with another metal when placed in each of the solutions A^{2+}, B^{2+}, D^{2+} and E^{2+}.

iii. Metal D becomes coated with another metal when placed in each of the solutions A^{2+} and B^{2+}, but not when placed in the solution E^{2+}.

Rank each of the 5 metals in order of increasing E^0 values.

_____

_____

_____

NOTE: Greater E^0 values lie on the top of the electrochemical series!

ALSO NOTE: These questions are a little tougher and can come up on the exam sometimes, so just be aware of them!

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Question 32

A student conducted an experiment using metals P, Q, R and S and solutions of their ions: P^{2+}, Q^{+}, $R^{2+}$and S^{2+} to determine an electrochemical series. The results are shown in the table below.

Chemicals mixed	Observations
Metal S with $Q^{+}(aq)$ ions	Reaction
Metal R with $P^{2+}(aq)$ ions	Reaction
Metal S with $P^{2+}(aq)$ ions	No Reaction

List the four metal ions in terms of decreasing oxidant strength.

_____

Question 33

Consider the following information about the reaction of Ru^{2+} with various reagents.

$$Ru^{2+}(aq) + Fe^{2+}(aq) \rightarrow \text{no observed reaction.}$$

$$Ru^{2+}(aq) + Ni(s) \rightarrow Ru(s) + Ni^{2+}(aq)$$

$$Ru^{2+}(aq) + Ag(s) \rightarrow \text{no observed reaction.}$$

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

Where would the following reaction be placed in the electrochemical series if the above tests were carried out under standard conditions?

$$Ru^{2+}(aq) + 2e^{-} \rightarrow Ru(s)$$

A. Below $-0.23\ V$

B. Between $-0.44\ V$ and $-0.23\ V$

C. Between $0.77\ V$ and $0.34\ V$

D. Above $0.77\ V$

Question 34

When metal X is placed in a solution of Y^{2+} ions, Y and X^{2+} are formed. When both metals are placed in an acidified solution, no reaction occurs. The order in which the species X^{2+}, Y^{2+} and H^{+} decrease in oxidising strength is:

A. $Y^{2+} > X^{2+} > H^{+}$

B. $X^{2+} > Y^{2+} > H^{+}$

C. $H^{+} > Y^{2+} > X^{2+}$

D. $X^{2+} > H^{+} > Y^{2+}$

Question 35 Additional Question.

When a drop of 1.0 *M* $XCl_2(aq)$ is placed on a lead plate, and no reaction is observed. When a drop of 1.0 *M* $YCl_3(aq)$ is placed on a different part of the lead plate, a reaction is observed. The results of these experiments indicate that:

A. X is a stronger reductant than Pb and that Y^{3+} is a stronger oxidant than Pb^{2+} .

B. Pb is a stronger reductant than X and that Pb^{2+} is a stronger oxidant than Y^{3+}.

C. X is a stronger oxidant than Pb and that Y^{3+} is a stronger reductant than Pb^{2+}.

D. Pb is a stronger oxidant than X and that Pb^{2+} is a stronger reductant than Y^{3+}.

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Contour Check

Learning Objective: [1.7.1] Apply the ECS to predict spontaneous reactions.

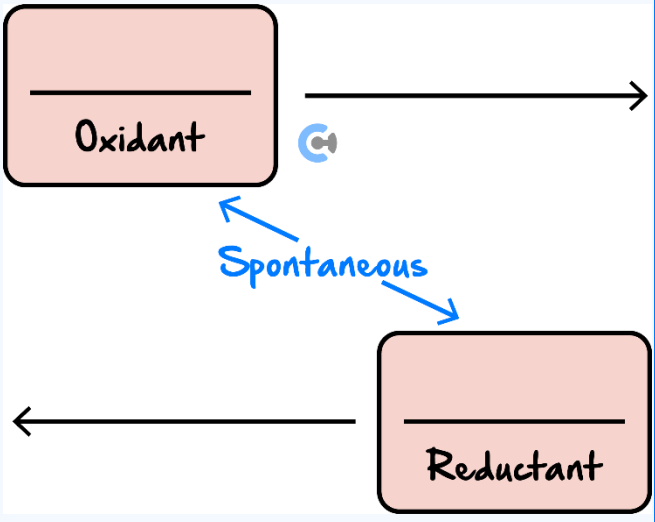
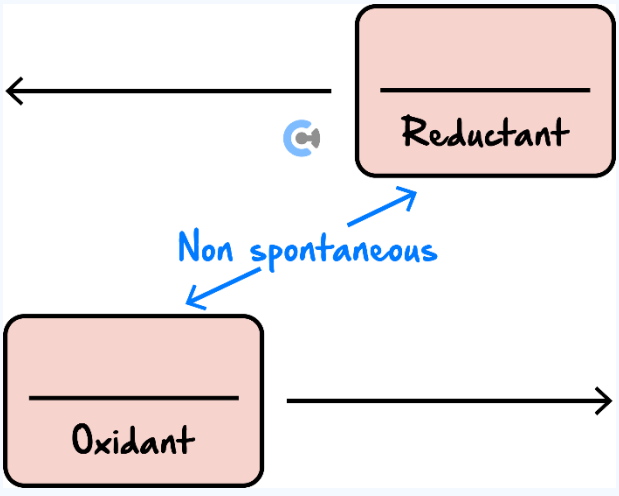
Key Takeaways

- Net Ionic Equation Definition: A balanced full equation with _____ omitted.
- **Spectator Ion**: Compound which is present but does not _____.

Reduction Reaction	Oxidation Reaction
[forward] / [reverse] reaction on ECS	[forward] / [reverse] reaction on ECS

Oxidants	Reductants
Positioned on the [left] / [right] side	Positioned on the [left] / [right] side

Strongest Oxidants	Strongest Reductants
Positioned [top] / [bottom] - [left] / [right]	Positioned [top] / [bottom] - [left] / [right]

Spontaneous Redox Reaction	Non-Spontaneous Redox Reaction
	

- Steps to predicting spontaneous reaction:
  1. Split all species into _____. Some cations/anions are _____.
  2. Locate all species on the _____. Draw a _____ to separate oxidants and reductants apart.
  3. Draw a _____.
  4. Find the strongest oxidant ([top] / [bottom] - [left] / [right]) and strongest reductant ([top] / [bottom] - [left] / [right])
  5. Check for _____.
  6. Write out half-equations.
- When multiple oxidants/reductants are present, the _____ oxidant reacts with _____ reductant.
- The four ions which appear on both sides of the electrochemical series:
  - _____
  - _____
  - _____
  - _____

Learning Objective: [1.7.2] Identify differences between direct & indirect redox reactions, & features of ECS

Key Takeaways

- Standard Electrode Potential Definition: Method to measure _____.
- Standard Hydrogen Electrode (SHE): $H^+(aq)/H_2(g)$ which has $E^0 =$ _____.
- The electrochemical series does not predict the _____.

Direct Contact Spontaneous Redox Reaction	Indirect Contact Spontaneous Redox Reaction

Learning Objective: [1.7.3] Find strongest oxidants/reductants by constructing your own ECS

Key Takeaways

- Electrochemical series ordered from [lowest → highest] / [highest → lowest] E^0 value.

Strongest Oxidant	Strongest Reductant
[highest] / [lowest] E^0 value	[highest] / [lowest] E^0 value

- Creating electrochemical series yourself steps:

1. Draw a _____ to separate oxidants and reductants.
2. Using information, place oxidants/reductants on this mini electrochemical series.

Spontaneous Reactions	Non-Spontaneous Reactions
[positive] / [negative] gradient	[positive] / [negative] gradient

3. Write the _____ version of the oxidant/reductant.
4. Repeat for each piece of information.

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