

Name: _____

Date: _____

Period: _____

SL IB Chemistry 9.1_ Redox - Guided Notes

Slide 3: Instead of looking for where electrons have been gained or lost, we can see if something has been oxidised or reduced by looking at its oxidation state.

Every element within a compound has an oxidation state.

If the oxidation state increases → oxidation!

If the oxidation state decreases → reduction!

NB: sometimes you will hear the term “oxidation number” used interchangeably with oxidation state, however IUPAC states that oxidation number refers to Roman numerals used with transition metals.

- What are oxidation states?

Slide 4: Rules for assigning oxidation states

- 1. The oxidation state of any _____ element is 0.
- 2. The _____ state of a monatomic ion equals the charge on the ion.
- 3. The more-_____ element in a binary compound is assigned the number equal to the charge it would have if it were an ion.
- 4. The _____ state of fluorine in a compound is always -1.
- 5. Oxygen has an _____ state of -2 unless it is combined with F (when it is +2), or it is in a peroxide (such as H_2O_2 or Na_2O_2), when it is -1.
- 6. The _____ state of hydrogen in most of its compounds is +1 unless it is combined with a metal, in which case it is -1.
- 7. In compounds, the elements of groups 1 and 2 as well as aluminum have oxidation states of +1, +2, and +3 _____.
- 8. The sum of the _____ states of all atoms in a neutral compound is 0.
- 9. The sum of the oxidation states of all atoms in a _____ ion equals the charge of the ion.

Slide 5: Important ones to know

Slide 6: PbCl₄

OS of Cl = -1

$$4 \times (-1) = -4$$

Overall OS = zero

OS of Pb = +4

- **Examples**

Slide 7: CO₃²⁻

OS of O = -2

$$3 \times (-2) = -6$$

Overall OS = -2

OS of C = +4

- **Examples**

- SbCl_3
- TiCl_4
- BrF_5
- lead(IV) oxide
- tin(II) chloride
- _____(III) chloride
- _____(IV) chloride
- bromine(V) fluoride
- Oxidation numbers of _____ elements are indicated by Roman numerals.

Slide 12: Redox reactions

- In any reaction where one species is oxidised (loses _____), another must be reduced (gains electrons). Both reduction and oxidation take place and so they are called redox reactions.
- _____ in terms of electrons:
- _____ = loss of electrons
- _____ = gain of electrons
- _____ agent = species that removes electrons from another
- reducing agent = species that gives _____ to another

Slide 13: Reminder

If the oxidation state increases → oxidation!

If the oxidation state decreases → reduction!

- What are oxidation states?

LEO
 (Lose Electrons = Oxidized)
The Lion, Says...
GER
 (Gain Electrons = Reduced)

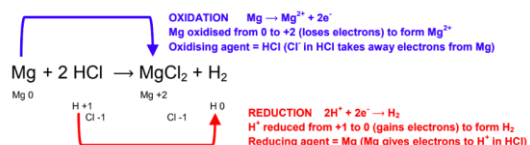


Slide 14: Redox reactions

- Note that the _____ agent removes electrons from a

species and so is itself reduced as it gains electrons. Similarly, a reducing agent is itself oxidised.

- _____ **agents** get reduced
- _____ **agents** get oxidized



Slide 15: The difference between redox and non-redox reactions is that only redox reactions involve change in oxidation state. Identify the reactions as redox or non-redox:

- _____ **States in Redox Reactions**

Slide 16: Check your work

Slide 17: In a disproportionation reaction, a species is simultaneously oxidised and reduced.



Show that these are all disproportionation reactions.

- _____ **Reactions**

Slide 18: Copy and complete the table

- https://docs.google.com/document/d/1hy3LmQAsDJVu4pPQ_rGBwll6fVPIId7F-mn9_AYoorbk/edit?usp=sharing

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Slide 19: Copy and complete the table

Slide 20: Check your work

Slide 21: Check your work

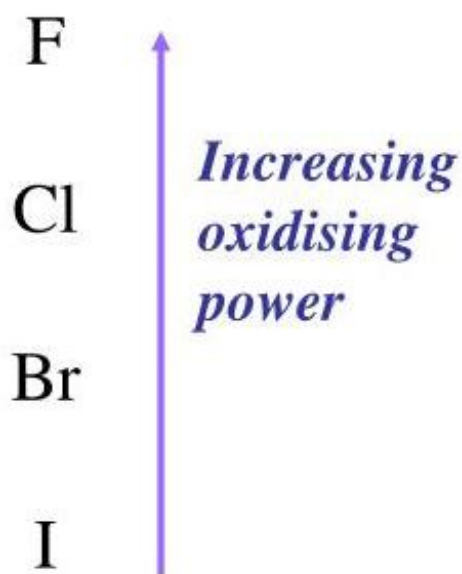
Slide 22: Activity series

- The activity series ranks metals _____ to the ease with which they undergo **oxidation**.
- _____ 25 of data books
- More _____ **metals** are stronger **reducing agents**

Slide 23: Activity of Halogens

- Group 17 on the periodic table ranks halogens _____ to the ease with which they undergo **reduction**.
- More reactive _____ are stronger **oxidizing agents**

This is because the *electronegativity* increases as we go up the group, which means that the elements gain electrons more easily.



For example: $\text{Cl}_2(\text{aq}) + 2\text{Br}^-(\text{aq}) \rightarrow \text{Br}_2(\text{aq}) + 2\text{Cl}^-(\text{aq})$

Slide 24: Practice

- **PROFICIENT**

Slide 25: Redox half equations

- Can you recall the meaning of these _____:
- Oxidation
- Reduction
- Redox
- Oxidising agent
- Reducing agent
- Reductant
- Oxidant
- Half equation

Slide 26: Half Equations

- It can be _____ to see what is being oxidised and what is being reduced when you look at a full chemical equation:
- We use half _____ to show the transfer of electrons.

Slide 27: Half Equations

- 2. Write the elements _____ as equations.
- $\text{H}_2 \rightarrow 2\text{H}^+$
- $\text{F}_2 \rightarrow 2\text{F}^-$
- 3. Work out the overall charge of each side of the _____.
- Write the _____ states for the elements.

- 0
- 0
- +1
- -1
- 0
- 0
- +2
- -2
- 4. Balance the charges with _____ (always add to the more positive side of the equation).

Slide 28: Half Equations

- $\text{H}_2 \rightarrow 2\text{H}^+$
- $\text{F}_2 \rightarrow 2\text{F}^-$
- 0
- 0
- +2
- -2
- 4. Balance the charges with electrons (always add to the more positive side of the equation).
- Remember, it doesn't need to be 0, both sides just need to be the same. Adding electrons will reduce the charge.
- $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
- $2\text{e}^- + \text{F}_2 \rightarrow 2\text{F}^-$
- **Which equation shows oxidation and which show reduction?**

Slide 29: You Try

- $\text{Na} \rightarrow \text{Na}^+$
- $\text{Pb}^{4+} \rightarrow \text{Pb}^{2+}$
- $\text{H}_2 \rightarrow \text{H}^+$
- $\text{Br}^- \rightarrow \text{Br}_2$

Slide 30: Check your work



Slide 31: https://docs.google.com/document/d/1h5-vh_NNXnr8zxIdEUVhzQowpcxLAGGSdooVCwnqlwQ/edit?usp=sharing

- This document outlines how to balance redox equations in acidic and basic solutions.

Slide 32: You Try

- $\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}$
- $\text{SO}_4^{2-} \rightarrow \text{S}$

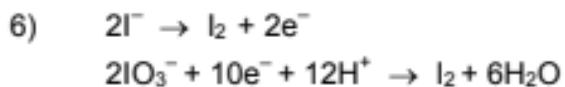
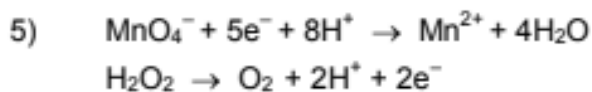
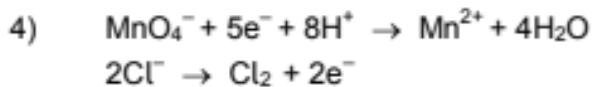
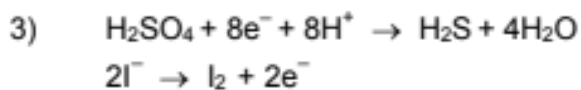
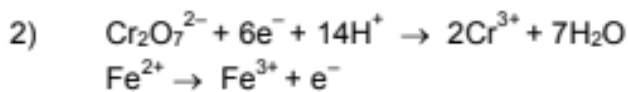
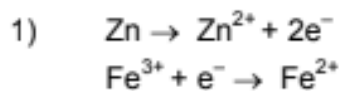
Slide 33: Combining Half Equations

- If you add the 2 half _____ for a redox reaction, you will return to the original *net ionic* equation.
- $\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$
- $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
- $\text{F}_2 + \text{H}_2 \rightarrow 2\text{F}^- + 2\text{H}^+$

- $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
- $2\text{e}^- + \text{F}_2 \rightarrow 2\text{F}^-$

Slide 34: Try for yourself

Combine these half equations to produce redox reactions.



Slide 35: Check your work

Slide 36: Redox titrations

Slide 37: What's the difference?

- Rather than an acid/base we have an oxidising agent/substance to be oxidised.
- So you need to find out the amount of substance that has been reduced or oxidised.
- Why can't we use the basic $M_1V_1 = M_2V_2$ formula?
- What else do we need to think about?

Slide 38: Redox Titrations

- Indicators are sometimes not needed in redox titrations because some redox changes are _____ by a color change (e.g. iodine in water, methylene blue and glucose).
- Redox _____ are used in:
- Food and _____ industry (e.g. wines analyzed for presence of SO_2 , analysis of vitamin C content in foods, etc.)
- _____ industry (e.g. purity and content analysis)
- Water and _____ analysis (e.g. determination of dissolved oxygen)

Slide 39: Steps to finding the mass/conc

- **Write out the two half ionic _____.**
- **Combine for the full ionic _____.**
- **Calculate the moles of the substance you have the values for (e.g. in the burette) by using the volume of the solution needed to reach _____.**
- _____ the moles of the reacting substance (e.g. in the flask).
- **Work out the moles in the _____ solution** (may need to multiply up for 1 dm^3 or full volume of starting solution).
- **Work out the mass or _____ of the original sample.**
- **Titre:** The volume of solution added from the burette to reach _____

Slide 40: Equations for typically used reagents

- manganate (VII) MnO_4^- (or _____) turns to Mn^{2+} and is an oxidising agent.
- $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
- _____ ($\text{C}_2\text{O}_4^{2-}$) becomes carbon dioxide (reducing agent)
- $\text{C}_2\text{O}_4^{2-}(\text{aq}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{e}^-$
- Iron loses _____ (reducing agent). Although can be both ways round.
- $2\text{Fe}^{2+}(\text{s}) \rightarrow 2\text{Fe}^{3+}(\text{s}) + 2\text{e}^-$

- _____(VI) ions (oxidising agent)
- $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$

Slide 41: Example-Analysis of iron with manganate (VII)

- **24.30cm³ of 0.010 mol dm⁻³ aqueous potassium manganate (VII) reacted with 25.0cm³ acidified FeSO₄(aq) to form Mn²⁺. This was taken from a solution originally made in a 250cm³ flask. What was the original mass of Fe²⁺ used?**
- $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
- $5\text{Fe}^{2+} \rightarrow 5\text{Fe}^{3+} + 5\text{e}^-$ (make both e- the same)
- $\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$
- $\text{MnO}_4^- : 5\text{Fe}^{2+}$ is a 1:5 ratio
- Moles $\text{MnO}_4^- = c \times v$
- $= 0.010 \times 0.0243 = 0.000243 \text{ mol}$
- Moles of Fe^{2+} in 25cm³ flask = 1:5 ratio
- $= 0.000243 \text{ mol} \times 5 = 0.001215 \text{ mol}$
- Moles of Fe^{2+} in 250cm³ = $0.001215 \times 10 = 0.01215 \text{ mol}$
- Mass = $M_r \times \text{mol} = 56 \times 0.01215 = \mathbf{0.68\text{g}}$

Slide 42: You Try-Analysis of Iron with manganate (VII)

- **All the iron in a 2.000 g tablet was dissolved in an acidic solution and converted to Fe²⁺. This was then titrated with KMnO₄. The titration required 27.50 cm³ of 0.100 mol dm⁻³ KMnO₄. Calculate the total mass of iron in the tablet and its percentage by mass. Describe what would be observed during the reaction, and how the equivalence can be detected.**
- Hint: During this reaction, Fe²⁺ is oxidized to Fe³⁺, and MnO₄⁻ is reduced to Mn²⁺

Slide 43: Example-Iodine-thiosulfate reaction

- **Several reactions use an oxidizing agent to react with excess iodide ions to form iodine.**
- $2\text{I}^-(\text{aq}) + \text{oxidizing agent} \rightarrow \text{I}_2(\text{aq}) + \text{reduced product}$
- Iodine is then titrated with sodium thiosulfate (Na₂S₂O₃), using starch as an indicator:
- **Oxidation: $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^-$**
- **Reduction: $\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$**

- **Overall equation:**
- $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow 2\text{I}^-(\text{aq}) + \text{S}_4\text{O}_6^{2-}(\text{aq})$
- Forms deep blue color in presence of starch
- Color disappears at equivalence point

Slide 44: You Try-Iodine-Thiosulfate Titration

- **Household bleach is an oxidizing agent that contains sodium hypochlorite, NaOCl, as the active ingredient. It reacts with iodide ions in acidic solution as follows:**
- A 10.00 cm³ sample of bleach was reacted with a solution of excess iodide ions, and the liberated iodine was then titrated with Na₂S₂O₃. The titration required 38.65 cm³ of 0.0200 mol dm⁻³ Na₂S₂O₃. Determine the concentration of OCl⁻ in the bleach.
- $\text{OCl}^-(\text{aq}) + 2\text{I}^-(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{I}_2(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$

Slide 45: Measures the dissolved oxygen content in water.

Dissolved oxygen is one of most important indicators of water quality (more pollution = less oxygen)

- Winkler Method
- Biological Oxygen Demand (BOD): amount of oxygen used to decompose the organic matter in a sample of water over a specified time period (usually 5 days) at a given _____ (high BOD = lower oxygen levels)

BOD Level (in ppm)	Water Quality
1 - 2	Very Good There will not be much organic waste present in the water supply.
3 - 5	Fair: Moderately Clean
6 - 9	Poor: Somewhat Polluted Usually indicates organic matter is present and bacteria are decomposing this waste.
10 or greater	Very Poor: Very Polluted Contains organic waste.

Slide 46: How are algal blooms related to increases in BOD?

- Turn and Talk

Slide 47: Redox titrations used to measure the dissolved oxygen in water and calculate the BOD.

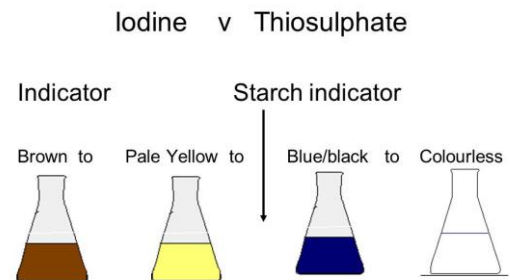
- Winkler Method
- The _____ O_2 (g) in the water is 'fixed' by the addition of a manganese (II) salt such as $MnSO_4$. O_2 reacts in basic solution to oxidize Mn (II) to higher oxidation states.
- _____ iodide ions, I^- , are added to the solution and oxidized by the Mn(IV) to I_2 .
- The iodine produced is titrated with sodium _____.
- $2Mn^{2+}(aq) + O_2(g) + 4OH^-(aq) \rightarrow 2MnO_2(s) + 2H_2O(l)$
- $MnO_2(s) + 2I^-(aq) + 4H^+(aq) \rightarrow Mn^{2+}(aq) + I_2(aq) + 2H_2O(l)$
- $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$
- **+2**
- **+4**
- **+4**
- **+2**

Slide 48: The Winkler method

- Under alkaline conditions, Mn^{2+} ions are rapidly oxidized to Mn^{3+} by dissolved oxygen, producing a pale brown _____ oh $Mn(OH)_3$.
- $4Mn(OH)_2(aq) + O_2(g) + 2H_2O(l) \rightarrow 4Mn(OH)_3(s)$
- A sample of river/stream water is shaken with excess alkaline Mn^{2+} ions and the _____ is reacted with an excess of KI, which oxidises to I_2 .
- $2KI(aq) + 2Mn(OH)_3(s) \rightarrow I_2(aq) + 2Mn(OH)_2(aq) + 2KOH(aq)$

Slide 49: The Winkler method

- The amount of I_2 is then determined by titration with sodium thiosulfate of known _____. Starch is used to show the end point.
- $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$



Slide 50: A 500 cm³ sample of water was collected and tested for dissolved oxygen by the addition of MnSO₄ in basic solution, followed by the addition of acidified KI. It was found that 12.50 cm³ of 0.0500 mol dm⁻³ Na₂S₂O₃ (aq) was required to react with the iodine produced. Calculate the dissolved oxygen content of the water in g dm⁻³, using the equations given on the previous slide.

- Example: Winkler Method
- Moles S₂O₃²⁻ = c x v
- = 0.0500 x 0.01250 = 0.000625 mol
- Moles of S₂O₃²⁻: I₂ = 2:1 ratio
- mol I₂ = 0.000625 mol x 0.5 = 0.0003125 mol
- Moles of I₂: MnO₂ = 1:1, Moles of MnO₂: O₂ = 2:1
- mol O₂ = 0.0003125 mol x 0.5 = 0.00015625 mol
- Mass = M_r x mol = 32.00 x 0.00015625 = **5.080 x 10⁻³ g in 500 cm³**
- **Dissolved oxygen: 0.0102 g dm⁻³**

Slide 51: The Winkler method was used to measure the concentration of dissolved oxygen in a sample of water. Manganese(II) sulfate, sulfuric acid and potassium iodide were added to 50.0 cm³ of the water. The iodine that was formed was titrated against a sodium thiosulfate solution with a concentration of 2.00 x 10⁻³ mol dm⁻³. It was found that 17.40 cm³ of sodium thiosulfate was required for the titration.

Calculate the concentration of dissolved oxygen in the original sample.

- You Try: Winkler Method

Slide 52: Moles S₂O₃²⁻ = c x v

$$= 0.00200 \times 0.01740 = 0.0000348 \text{ mol}$$

Moles of I₂: I₂ = 2:1 ratio

$$\text{mol I}_2 = 0.000348 \text{ mol} \times 0.5 = 0.000174 \text{ mol}$$

Moles of I₂: MnO₂ = 1:1, Moles of MnO₂: O₂ = 2:1

$$\text{mol O}_2 = 0.000174 \text{ mol} \times 0.5 = 8.70 \times 10^{-6} \text{ mol}$$

$$\text{Dissolved oxygen: } 8.70 \times 10^{-6} \text{ mol} / 0.050 \text{ dm}^{-3} = 1.74 \times 10^{-4} \text{ mol dm}^{-3}$$

- Answer: _____ Method
- $2\text{Mn}^{2+}(\text{aq}) + \text{O}_2(\text{g}) + 4\text{OH}^{-}(\text{aq}) \rightarrow 2\text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- $\text{MnO}_2(\text{s}) + 2\text{I}^{-}(\text{aq}) + 4\text{H}^{+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{I}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
- $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow 2\text{I}^{-}(\text{aq}) + \text{S}_4\text{O}_6^{2-}(\text{aq})$

Slide 53: Electrochemical Cells

- Practical _____ of Redox
- **Voltaic (_____)** cells
- _____ cells (HL)

Slide 54: Voltaic Cells

- _____ called galvanic cells.
- Can be made from two _____ metals **dipped in solutions of their own ions** and **connected by a wire**.
- Always two _____ – oxidation and reduction

Red Cat

Reduction is gain at the **Cathode**



An Ox

Oxidation is **Loss** at the **Anode**



Slide 55: Voltaic Cells

- Redox reactions are organized so that energy released is available as _____ energy.
- Oxidation and reduction reactions are separated into **half-cells (aka _____)** and electrons flow through an external circuit.
- _____ **flow from the most reactive (anode) to least (cathode).**
- A voltmeter shows the voltage between the two half cells, **e. m. f** (_____ **force**)

Slide 56: Voltaic Cells

- Zinc is more reactive than copper, so **loses _____ more easily** than copper to form Zn^{2+} (**most negative electrode = as undergoes oxidation**)
- _____ released into the circuit and same no. of electrons are taken to reduce the Cu^{2+} to form copper atoms (**most positive electrode = as undergoes reduction**)

Slide 57: Salt bridge

- **A salt bridge (_____ KNO_3) allows free moving ions to conduct the charge.**
- The salt bridge provides an **ionic connection** between the two ionic solutions. The ions are free to move in the bridge so charge is transferred through the bridge solution and this **keeps each half of the cell _____ neutral.**

Slide 58: Salt bridge

- The salt used should be _____ with the electrodes and electrode solutions. E.g. potassium chloride would not be suitable for copper systems as chloride ions can form complexes with copper ions

Slide 59: Electrode Potentials

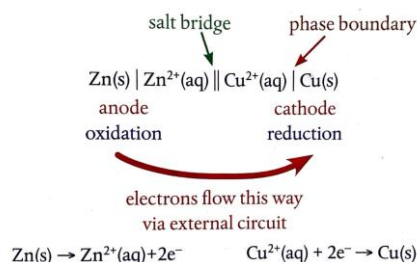
- $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
- $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
- Anode
- Oxidation
- Cathode
- Reduction
- -
- +

Slide 60: You Try

- Draw a voltaic cell with one half-cell consisting of Mg and a solution of Mg^{2+} ions and the other consisting of Zn and a solution of Zn^{2+} ions. Label the electrodes with name and charge, the direction of electron and ion movement, and write equations for the reactions occurring at each electrode.

Slide 61: Cell Diagram Convention

- _____ **version of voltaic cell:**
- Single vertical line _____ phase boundary (e.g. between solid electrode and aqueous solution)
- Double vertical line _____ salt bridge
- Aqueous _____ of each electrode are placed next to the salt bridge
- Anode _____ put on the left and cathode on the right (electrons flow from left to right)
- _____ ions omitted
- If half-cells contain two ions, they are _____ by a comma



Slide 62: Pt(s) Fe²⁺(aq) Fe³⁺(aq) Ti³⁺(aq) Ti+(aq) Pt(s)

- Left hand side, negative _____, **oxidation (anode)**. Higher on activity series.
- Positive _____, right hand side, **reduction (cathode)**
- **Writing half _____ from cell diagrams**

Slide 63: You Try

- Draw a cell diagram to represent the voltaic cell drawn on slide 61.

Slide 64: Electrode Potentials

- Reactions that occur at each electrode in the cell are _____.
- $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$
- $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$
- Direction depends upon how easily electrons are lost by each metal (how easily it is **oxidised**) – measured using electrode _____.

Slide 65: Electrode potentials

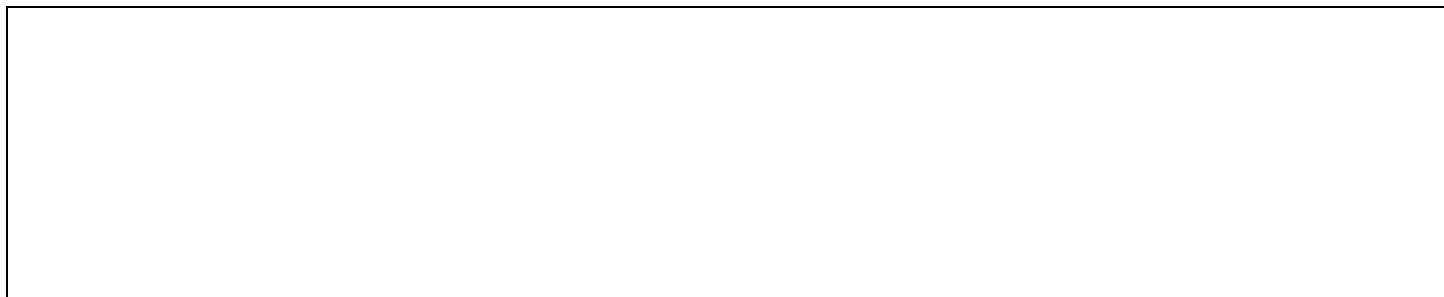
- Zinc half cell **more negative** _____ potential.
- Zinc is oxidised so reaction **goes** _____
- Copper is reduced so _____ goes forwards.
- The more reactive a metal, the more negative its _____ potential
- $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$
- $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$

Slide 66: Electrode potentials

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

Slide 67: Activity series

- The activity series (section 25 of data booklet) ranks metals in order of the ease with which they undergo oxidation, and enables us to predict whether redox reactions will occur.
- More reactive elements will displace less reactive elements from their compounds.
- For example:
- $\text{Fe}(\text{s}) + \text{CuSO}_{4(\text{aq})} \rightarrow \text{FeSO}_{4(\text{aq})} + \text{Cu}(\text{s})$
- **In cells, the most reactive metal will produce electrons (oxidation - anode) and the least reactive metal will take electrons (reduction - cathode)**



Slide 68: Different half-cells make voltaic cells with different voltages

Slide 69: Different half-cells make voltaic cells with different voltages

Slide 71: You Try

- Use the metal reactivity series in section 25 of the data booklet to predict which electrode will be the anode and which will be the cathode when the following half-cells are connected. Write half-equations for the reactions occurring at each electrode.
- Zn/Zn²⁺ and Fe/Fe²⁺
- Fe/Fe²⁺ and Mg/Mg²⁺
- Mg/Mg²⁺ and Cu/Cu²⁺