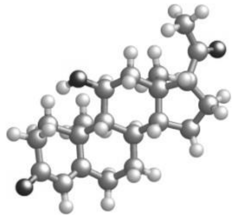

*Chem!stry*

**Name:** …………………… (    )

**Class:** ………………

**Date:** …… / …… / ……

**Oxidation and Reduction – Redox Reactions – Macroconcept: Change**

Activity 1:

**Definition Number One (Easy):**

**Oxidation** is the addition of …………………… to a substance or the removal of …………………… from a substance.

**Reduction** is the addition of …………………… to a substance or the removal of …………………… from a substance.

Activity 2:

• Use a coloured pencil to circle the elements that have been *oxidised* and a different coloured pencil to circle the elements that have been *reduced* in each of the following reactions:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

$$C_6H_6(l) + 3H_2(g) \rightarrow C_6H_{12}(l)$$

Key: ☐ (colour) = *oxidised*.
☐ (colour) = *reduced*.

Activity 3:

• Write your own balanced chemical equations below and ask another student in your class to identify the elements that have been *oxidised* and the elements that have been *reduced* in each case:

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Activity 4:

## Definition Number Two (Moderate):

**Oxidation** is the …………………… of …………………… from a substance.

**Reduction** is the …………………… of …………………… to a substance.

There is a simple mnemonic to help you remember this: ***OIL RIG***

***O***…………………… ***I***… ***L***…………………… (of ……………………).

***R***…………………… ***I***… ***G***…………………… (of ……………………).

For example, the balanced chemical equation for the reaction between sodium and chlorine is:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

During this reaction each sodium atom transfers its single valence electron into the valence shell of a chlorine atom. The sodium atoms have ***lost*** electrons and have therefore been ***oxidised*** to form positive sodium ions:

$$2Na(s) \rightarrow 2Na^{+}(s) + 2e^{-} \text{ (ionic half-equation for sodium)}$$

*Note:* an alternate way of writing the above equation is $2Na(s) \rightarrow 2Na^{+}(s) + 2e^{-}$

The chlorine atoms have ***gained*** electrons and have therefore been ***reduced*** to form negative chloride ions:

$$Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(g) \quad \text{(ionic half-equation for chlorine)}$$

*Note:* a reaction in which one element is reduced while another element is oxidised is called a ***redox*** reaction.

Activity 5:

• Break each of the following balanced chemical equations down into its component ionic half-equations and hence state which elements have been oxidised and which elements have been reduced:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

………………………………………………………………………………………………

………………………………………………………………………………………………

$$CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$$

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Activity 6:

• Write your own balanced chemical equations in the space provided below. Ask another student in your class to break each balanced chemical equation down into its component ionic half-equations and hence identify the elements that have been *oxidised* and the elements that have been *reduced*:

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## Definition Number Three (Advanced):

• Definition of oxidation state: Oxidation state is the hypothetical charge that an atom would have if all bonds to atoms of different elements were completely ionic, with no covalent character (note: this is never true for real bonds). Oxidation states maybe positive, negative, or zero.

Activity 7:

**Oxidation** is an …………………… in oxidation number or oxidation state.
**Reduction** is a …………………… in oxidation number or oxidation state.

The basic rules for the calculation of an oxidation state are as follows:

• The oxidation state of a pure element is always 0.

*e.g.* *For metallic iron (Fe) the oxidation state of the iron = 0.*
*For nitrogen gas ($N_2$) the oxidation state of the nitrogen = 0.*

• The oxidation state of fluorine in a compound is always –1.

*e.g.* *In sodium fluoride (NaF) the oxidation state of the sodium is +1 and the oxidation state of the fluorine is –1.*

• The oxidation state of oxygen in a compound is nearly always –2 (unless combined with fluorine). In peroxides (e.g. hydrogen peroxide, $H_2O_2$) it is –1.

*e.g.* *In carbon dioxide ($CO_2$) the oxidation state of the carbon is +4 and the oxidation state of each oxygen is –2.*

• The oxidation state of chlorine in a compound is nearly always –1 (unless combined with fluorine or oxygen).

*e.g.* *In magnesium chloride ($MgCl_2$) the oxidation state of the magnesium is +2 and the oxidation state of each chlorine is –1.*

• The oxidation state of a Group 1 element in a compound is always +1.

*e.g.* *In potassium bromide (KBr) the oxidation state of the potassium is +1 and the oxidation state of the bromine is –1.*

• The oxidation state of a Group 2 element in a compound is always +2.

*e.g.* *In calcium oxide (CaO) the oxidation state of the calcium is +2 and the oxidation state of the oxygen is –2.*

• The oxidation state of a Group 13 element in a compound is always +3.

*e.g.* *In aluminium chloride ($AlCl_3$) the oxidation state of the aluminium is +3 and the oxidation state of each chlorine is –1.*

• The oxidation state of hydrogen in a compound is nearly always +1 (unless combined with a reactive metal, in which case it is –1).

*e.g.* *In water ($H_2O$) the oxidation state of each hydrogen is +1 and the oxidation state of the oxygen is –2. In sodium hydride (NaH) the oxidation state of the sodium is +1 and the oxidation state of the hydrogen is –1.*

• The *sum* of the oxidation states of the elements in a compound is equal to the *charge* carried by the compound.

*e.g.* *In the nitrate ion ($NO_3^-$) the overall charge on the ion is –1, so the sum of the oxidation state must equal –1. Because the oxidation state of oxygen in a compound is known to be –2, the oxidation state of nitrogen can be calculated:*

$$N + (3 \times O) = -1$$

$$N + (3 \times -2) = -1$$

$$N + (-6) = -1$$

$$N = -1 + 6$$

$$N = +5$$

Activity 8:

*e.g.* *In the sulfate ion ($SO_4^{2-}$) the overall charge on the ion is –2, so the sum of the oxidation state must equal –2. Because the oxidation state of oxygen in a compound is known to be –2, the oxidation state of sulfur can be calculated:*

$$S + (4 \times O) = -2$$
$$S + (4 \times -2) = -2$$
$$S + (-8) = -2$$
$$S = -2 + 8$$
$$S = +6$$

• Calculate the oxidation state of each of the elements in bold italics:

| | |
|---|---|
| $\boldsymbol{O}_2$ | $\boldsymbol{S}O_2$ |
| $\boldsymbol{Zn}O$ | $\boldsymbol{S}O_3$ |
| $\boldsymbol{N}O_2$ | $H_2\boldsymbol{S}$ |
| $\boldsymbol{C}O_3^{2-}$ | $\boldsymbol{Cr}O_4^{2-}$ |
| $\boldsymbol{Cu}^{2+}$ | $\boldsymbol{S}O_3^{2-}$ |
| $\boldsymbol{V}_2O_5$ | $\boldsymbol{I}O_4^{-}$ |
| $\boldsymbol{Cr}_2O_7^{2-}$ | $\boldsymbol{P}O_4^{3-}$ |
| $\boldsymbol{Fe}Cl_3$ | $H\boldsymbol{N}O_3$ |

## Activity 9:

• Calculate the oxidation state of each of the elements in bold italics at the start and at the end of the reaction and hence state whether the element has been oxidised or reduced during the reaction:

**(a)** $\boldsymbol{Fe}_2O_3(s) + 3\boldsymbol{C}O(g) \rightarrow 2\boldsymbol{Fe}(s) + 3\boldsymbol{C}O_2(g)$

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**(b)** $\boldsymbol{Mg}(s) + \boldsymbol{H}_2SO_4(aq) \rightarrow \boldsymbol{Mg}SO_4(aq) + \boldsymbol{H}_2(g)$

…………………………………………………………………………………………………

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**(c)** $\boldsymbol{Zn}(s) + \boldsymbol{Cu}SO_4(aq) \rightarrow \boldsymbol{Zn}SO_4(aq) + \boldsymbol{Cu}(s)$

…………………………………………………………………………………………………

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**(d)** $\boldsymbol{Mn}O_2(s) + 4HCl(aq) \rightarrow \boldsymbol{Mn}Cl_2(aq) + 2H_2O(l) + Cl_2(g)$

…………………………………………………………………………………………………

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**(e)** $\boldsymbol{Cr}_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2\boldsymbol{Cr}^{3+}(aq) + 7H_2O(l)$

…………………………………………………………………………………………………

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**(f)** $2\boldsymbol{S}_2O_3^{2-}(aq) + I_2(aq) \rightarrow \boldsymbol{S}_4O_6^{2-}(aq) + 2I^-(aq)$

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**(g)** $5\boldsymbol{Fe}^{2+}(aq) + \boldsymbol{Mn}O_4^-(aq) + 8H^+(aq) \rightarrow 5\boldsymbol{Fe}^{3+}(aq) + \boldsymbol{Mn}^{2+}(aq) + 4H_2O(l)$

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**(h)** $Na_2\boldsymbol{S}_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + \boldsymbol{S}O_2(g) + \boldsymbol{S}(s) + H_2O(l)$

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## Oxidising and Reducing Agents:

Potassium manganate(VII) $KMnO_4$ can act as an oxidising agent under acidic conditions. The ionic half-equation for the manganate(VII) ion behaving as an oxidising agent is:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l) \qquad \text{Equation 1.}$$

*Note*: the potassium ion $K^+$ is a spectator ion. It is not actually involved in the chemical reaction and is therefore omitted from the balanced chemical equation.

Activity 10:

• By looking at Equation 1, how can you tell that the manganate(VII) ion is behaving as an oxidising agent?

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Activity 11:

• Calculate the change in oxidation state of the manganese during this chemical reaction and hence state whether the oxidising agent has itself been oxidised or reduced:

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Iron(II) ions $Fe^{2+}$ can be oxidised to iron(III) ions $Fe^{3+}$ by acidified potassium manganate(VII). The half-equation for the oxidation of the iron(II) to iron(III) is:

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^- \qquad \text{Equation 2.}$$

This reaction only supplies one electron to the manganate(VII) ion, and so 5 moles of iron(II) ions are required to react with 1 mole of manganate(VII) ions:

$$5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + 5e^- \qquad \text{Equation 3.}$$

Combining together ionic half-equations (1) and (3) gives:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq) + 5e^-$$

The $5e^-$ on either side of the equation cancel out to give the overall balanced chemical equation for the oxidation of iron(II) to iron(III) by acidified manganate(VII):

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

Another oxidising agent is acidified potassium dichromate(VI) $K_2Cr_2O_7$:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

Again, the potassium ions are spectator ions that are not directly involved in the chemical reaction and so are omitted for simplicity.

Activity 12:

• Construct the overall balanced chemical equation for the oxidation of iron(II) to iron(III) by acidified dichromate(VI):

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The *Halogens* tend to behave as *oxidising agents*. Fluorine is an extremely powerful oxidising agent and the oxidising power of chlorine is comparable with that of oxygen. The oxidising power of the Halogens *decreases* down Group 17. The *Alkali Metals* tend to behave as *reducing agents*. The reducing power of the Alkali Metals *increases* down Group 1.

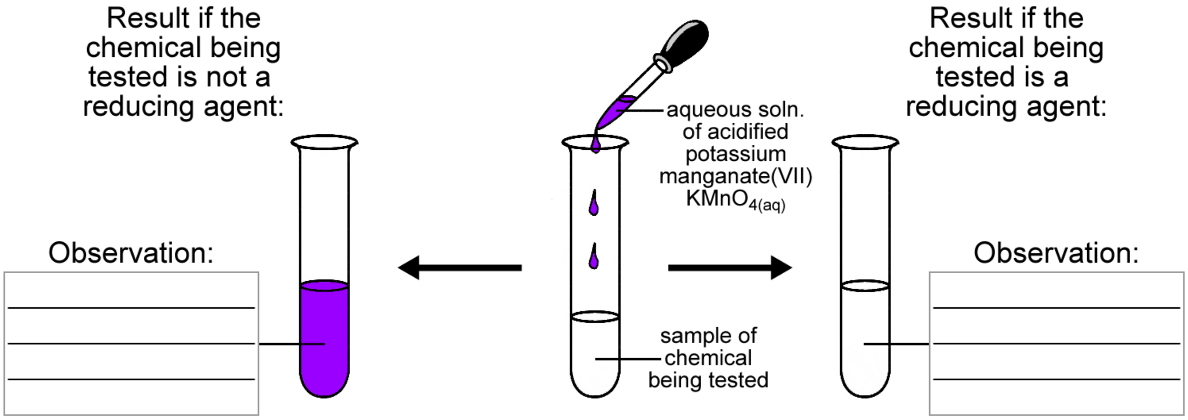
← Increasing Reducing Power : Increasing Oxidising Power →

**Cs Rb K Na Li : I Br C*l* F**

# Test for Reducing Agents and Oxidising Agents

## Activity 13:

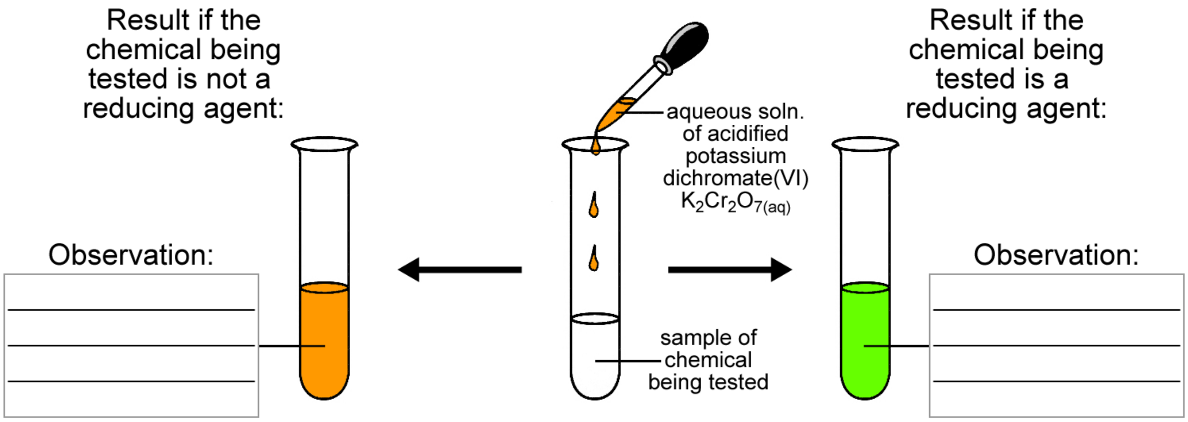
Test for reducing agents using acidified potassium manganate(VII):




Manganese is reduced from +7 to +2: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

## Activity 14:

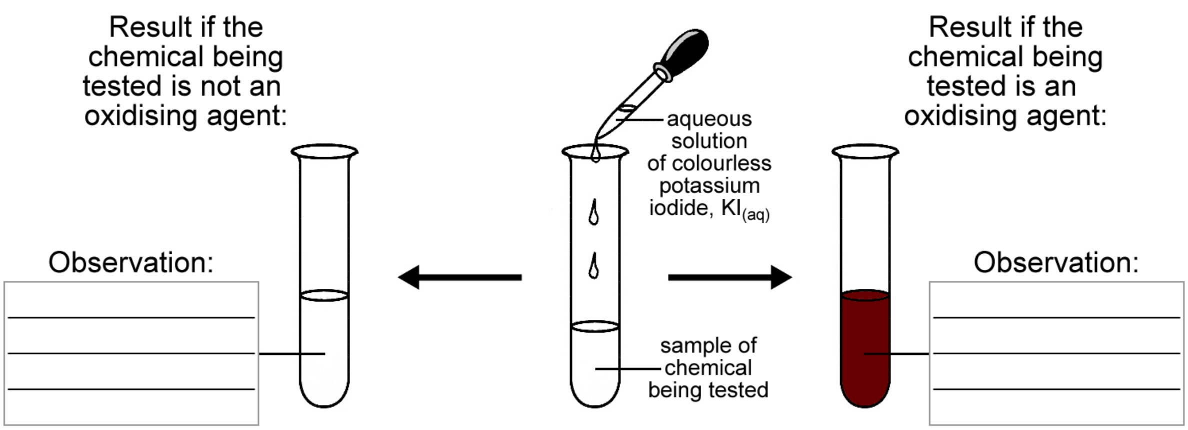
Test for reducing agents using acidified potassium dichromate(VI):




Chromium is reduced from +6 to +3: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

## Activity 15:

Test for oxidising agents using potassium iodide:




Iodine is oxidised from –1 to 0: $2I^-(aq) \rightarrow I_2(aq) + 2e^-$

**Disproportionation Reactions (Enrichment):**

Consider the decomposition of hydrogen peroxide into water and oxygen:

$$\text{hydrogen peroxide} \rightarrow \text{water} + \text{oxygen}$$

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

Activity 16:

• Calculate the change in oxidation state of the *oxygen* during this chemical reaction. Has the oxygen been oxidised or reduced during this reaction? What is unusual about this reaction?

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Consider the reaction between copper oxide and hot concentrated sulfuric acid:

$$\text{copper oxide} + \text{sulfuric acid} \rightarrow \text{copper} + \text{copper sulfate} + \text{water}$$

$$Cu_2O(s) + H_2SO_4(aq) \rightarrow Cu(s) + CuSO_4(aq) + H_2O(l)$$

Activity 17:

• Calculate the change in oxidation state of the *copper* during this chemical reaction. Has the copper been oxidised or reduced during this reaction? What is unusual about this reaction?

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Activity 18:

• The two chemical reactions given above are both examples of *disproportionation reactions*. Using what you have learned from the two chemical reactions given above, state what is meant by the term *disproportionation reaction*:

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……………………………………………………………………………………………

## An Application of Redox – Vitamin C as an Anti-oxidant (Enrichment):

Most animals can produce their own vitamin C, but humans and their close relatives the primates (apes, gorillas and chimpanzees) have lost this ability. Because of this, vitamin C is an essential nutrient in our diets.

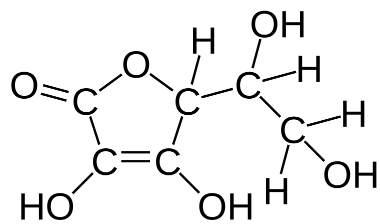
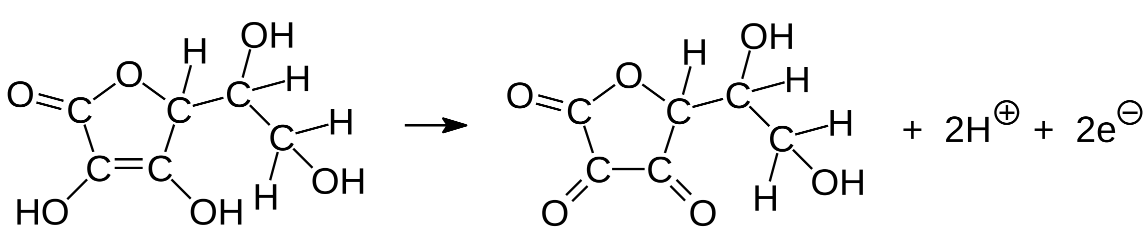


**Figure 1.** The structural formula of vitamin C.

Why is vitamin C an essential component of our diets? What is its function within our bodies? Vitamin C is vital to the production of *collagen*. Collagen is an essential component of the connective tissue within our bodies which gives our bodies form and supports our internal organs. Vitamin C also prevents and cures the disease scurvy, and can be beneficial in the treatment of iron deficiency, anaemia.

The term *anti-oxidant* refers to vitamin C's ability to protect the fat-soluble vitamins A and E, as well as fatty acids, from oxidation. If a mixture of vitamins A, E and C are exposed to an oxidising agent, then the vitamin C will be oxidised first, thus saving the vitamins A and E from oxidation.

During its oxidation, 1 mole of vitamin C loses 2 moles of electrons as shown in the balanced chemical equation below. *Note:* the reaction also produces 2 moles of hydrogen ions which is why vitamin C is often referred to as *ascorbic acid*:



Activity 19:

- Write a balanced chemical equation to describe the oxidation of vitamin C:

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## An Application of Redox – How Batteries Work (Enrichment):

When a strip of zinc metal is immersed in a beaker of aqueous copper(II) sulfate, the following redox reaction takes place:

zinc + copper(II) sulfate → zinc sulfate + copper

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Activity 20:

• Write ionic half–equations to describe what happens to **(a)** the zinc and **(b)** the copper when zinc metal reacts with an aqueous solution of copper(II) sulfate:

**(a)** ………………………………………………………………………………………….

**(b)** ………………………………………………………………………………………….

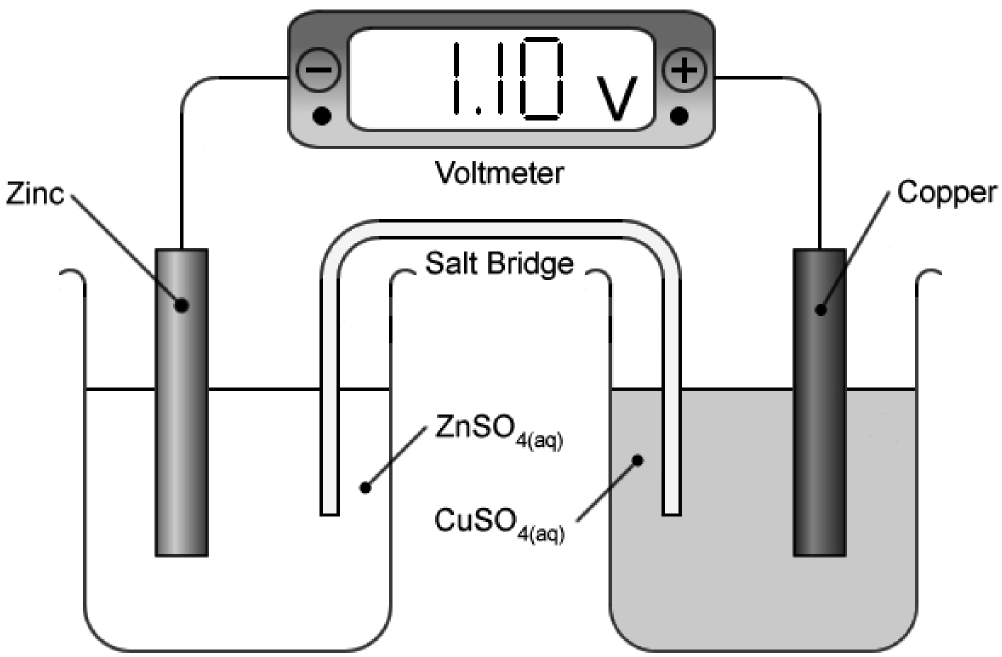


**Figure 2.** A simple electrochemical cell or battery.

Activity 21:

• Explain how the apparatus set up in **Figure 2.** generates a potential difference of 1.10 V.

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• Scan the QR code below for the answers to this assignment.

http://www.chemist.sg/redox/redox_worksheet_ans.pdf