NAME ............................................ Chemistry Class .......................

Student Number ……….

Redox Answers

Oxidation Is Loss of electrons

OIL RIG

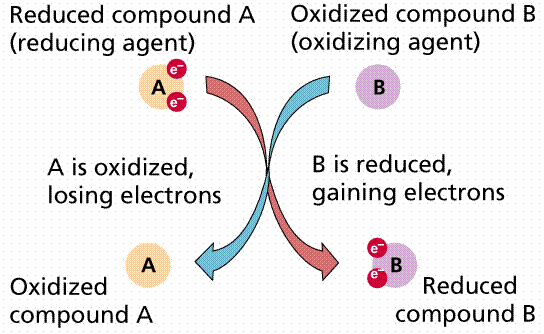
Reduction Is Gain of electrons

**Topic 3: Redox 1**

1. know what is meant by the term ‘oxidation number’
2. be able to calculate the oxidation number of elements in compounds and ions

*The use of oxidation numbers in peroxides and metal hydrides is expected.*

1. understand oxidation and reduction in terms of electron transfer and changes in oxidation number, applied to reactions of *s-* and *p-* block elements
2. understand oxidation and reduction in terms of electron loss or electron gain
3. know that oxidising agents gain electrons
4. know that reducing agents lose electrons
5. understand that a disproportionation reaction involves an element in a single species being simultaneously oxidised and reduced
6. know that oxidation number is a useful concept in terms of the classification of reactions as redox and as disproportionation
7. be able to indicate the oxidation number of an element in a compound or ion, using a Roman numeral
8. be able to write formulae given oxidation numbers
9. understand that metals, in general, form positive ions by loss of electrons with an increase in oxidation number
10. understand that non-metals, in general, form negative ions by gain of electrons with a decrease in oxidation number
11. be able to write ionic half - equations and use them to construct full ionic equations



**Introduction to Redox**

**New ReferenceRefs:-**

Facer AS Chemistry Chapter 9, p181-194

Fullick p71 – 73, p168 - 170

**Department Website** / Resources

See Factsheets:-

|  |  |
| --- | --- |
| 11 | Oxidation and Reduction 1 |
| 78 | Recognising, Constructing and Interpreting Redox Reactions |
| 88 | Disproportionation |
| 104 | A few H’s will produce any half equation |

**Websites:**-

Explaining the basics of redox

<http://www.chemistry.co.nz/redox_new.htm>

Step by step- how to write balanced redox equations

<http://www.chemguide.co.uk/inorganic/redox/equations.html>

Rules and Exceptions to oxidation numbers

<http://www.chemguide.co.uk/inorganic/redox/oxidnstates.html>

Oxidation Numbers- practice assigning an oxidation number to an element in an ion or molecule

<http://www.science.uwaterloo.ca/~cchieh/cact/c123/oxidstat.html>

**Contents**

Oxidation and reduction as electron transfer

Assigning oxidation numbers

Oxidation numbers in terms of number changes

Classification of reactions as redox and disproportionation

Half and full redox equations

Redox titrations

**Note: Worksheet – KEYNOTE – Rules for assigning oxidation numbers - HGB**

**Oxidation and reduction**

**New exercise**Originally oxidation was defined as gain of oxygen or loss of hydrogen. Reduction was defined as loss of oxygen or gain of hydrogen.

e.g. in the reaction

CuO + H2 🡪 H2O + Cu

Which substance is being oxidised? H Explain why gained an O atom

Which substance is being reduced? CuO Explain why lost an O atom

Later, it was realised that similar reactions did not involve oxygen or hydrogen

e.g. CuCl2 + Mg 🡪 MgCl2 + Cu

Using the same logic as before:

Which substance is being oxidised? Mg Reduced? Cu in CuCl2

In both the above examples a Cu2+ ion is changing into a Cu atom by **gaining** 2 **electrons.** This is the new definition of reduction.

OXIDATION is when a species LOSES one or more ELECTRONS

REDUCTION is when a species GAINS one or more ELECTRONS

Remember: **‘OIL RIG’**

Oxidation and reduction involve the transfer of electrons. It is impossible to have oxidation without reduction or vice versa. Therefore we refer to **redox reactions** as reactions which involve **transfer of electrons.**

**Redox reactions and half equations**

From your work on Group II you will remember the reaction between magnesium and oxygen (optional demonstration).

Write a balanced equation for this reaction

Mg(s) + ½O2(g) 🡪 MgO(s)

This is an example of a **redox reaction.**

What is happening to the each magnesium atom (in terms of electrons)? Lost 2e-

What is happening to each of the 2 oxygen atoms (in terms of electrons)? Gained 2e-

Write the oxidation half equation showing the electrons lost

OIL Mg(s) 🡪 Mg2+(s) + 2e-

Write the reduction equation showing the electrons gained

RIG ½O2(g) + 2e- 🡪 O2-(s)

We call the reduction and oxidation equations half equations, because they are only half of a complete redox reaction.

Now consider the reaction of magnesium with chlorine;-

Write a balanced equation for this reaction

Mg(s) + Cl2(g) 🡪 MgCl2(s)

What is happening to the magnesium atoms? They are forming Mg2+ ions they are being oxidised

What is happening to each chlorine atom? They are forming Cl- ions

This is also a redox reaction.

Write the reduction half equation Cl2(g) + 2e- 🡪 2Cl-(s) remember RIG

Write the oxidation half equation Mg(s) 🡪 Mg2+(aq) + 2e-

You might also have come across displacement reactions at GCSE, these are another example of redox reactions.

* Consider zinc reacting with copper(II) sulfate

Write the ionic equation for zinc and CuSO4 reacting to form copper and ZnSO4. The sulfate ions are spectator ions, i.e. they do not change and therefore do not appear in the half equations.

Full equation

Zn(s) + Cu2+(aq) 🡪 Zn2+(aq) + Cu(s)

Reduction half equation Cu2+(aq) + 2e- 🡪 Cu(s)

Oxidation half equation Zn(s) 🡪 Zn2+(aq) + 2e-

* Consider this reaction

2AgNO3 + Cu **🡪** 2Ag + Cu(NO3)2

Is this a redox reaction? yes

Reduction half equation (again ignore the spectator ions) Ag+(aq) + e- 🡪 Ag(s)

Oxidation half equation Cu(s) 🡪 Cu2+(aq) + 2e-

Redox is important in biology and physics as well as in chemistry. For example redox reactions underpin respiration in biology and electric cells in physics. In chemistry it explains rusting and why some materials will react with each other and others not. You will study redox further in the U6 where we will consider cells and batteries.

**Demonstration**

**Optional demo** – Blue Bottle (open evening demonstration).

Place 100cm3 of a solution of glucose in sodium hydroxide in a stoppered bottle.

Add 5 cm3 of methylene blue dye and wait. What do you observe?

Blue methylene blue formed a colourless solution

Glucose is a **reducing agent** (reducing sugar). The methylene blue dye has been reduced

Shake the bottle. What do you observe?

Solution turns blue

The methylene blue dye has now been **oxidised**. What do you think the **oxidising agent** is?

Oxygen in the air in the flask

Leave the bottle to stand again. What happens? Turned colourless, blue coloured at the meniscus

Eventually the colour will no longer change. Why not? All O2 in the flask air used up.**Oxidising agents and reducing agents**

An **oxidising agent brings about an oxidation**. It, itself, is reduced.

A **reducing agent brings about a reduction.** It, itself is oxidised.

New exercise

|  |  |  |
| --- | --- | --- |
| **Equation** | **Oxidising agent** | **Reducing agent** |
| CuO + Mg 🡪 MgO + Cu | Cu2+ | Mg |
| Zn + Cl2 🡪 ZnCl2 | Cl2 | Zn |
| CuSO4 + Fe 🡪 FeSO4 + Cu | Cu2+ | Fe |

**Covalent redox reactions**

So far we have only looked at ionic compounds where each of the species forms ions. However we can also look at redox reactions with covalent compounds:-

E.g. H2(g) + ½ O2(g) **→** H2O(l)

What has been oxidised? H2

What has been reduced? O2

Covalent bonds have been formed, so the electrons have not been fully transferred. In order to determine if a substance has been oxidised or reduced we **assign oxidation numbers** to each species in the reaction, even if the bonding is covalent.

**Oxidation Numbers**

Oxidation Number is the charge the species in a molecule or ion would have if the bonding was 100% ionic.

You have already used oxidation numbers when you have looked at copper(II) ion, iron(II) or iron(III). The roman numerals represent the oxidation number of the principal atom.

So for example

Cu2+ has an oxidation number of +2

I- has an oxidation number of -1

O2- has an oxidation number of -2

**New exercise**Note: we write the oxidation numbers as a sign followed by the number (charge is number then sign).

Assign oxidation numbers for

Mg in MgCl2 ………+2………………

Na in NaCl…………+1………………

Ca in CaO………+2………………

K in KI…………+1…………………

Fe in FeO…………+2……………

Cu in CuCl2………+2………………

It is fairly straightforward to assign oxidation numbers for ionic compounds, however for covalent compounds and ions containing more than one atom there are rules for assigning oxidation numbers which help us work out the oxidation numbers.

**LEARN THESE Rules for working out oxidation numbers:**

**1)** Elements in their **uncombined** state have an oxidation number of **zero**.

**2)** The **sum** of all the oxidation numbers **in an uncharged molecule is zero**.

**3)** The **sum** of all the oxidation numbers **in an** **ion** is the **charge** on the ion.

**4)** The oxidation number of **hydrogen is +1** except in metal hydrides when it is -1.

**5)** Being the most electronegative element, **fluorine** (unlike the other halogens) **always** has the oxidation number **-1**, and can bring out the highest oxidation number in any element it reacts with.

**6)** The oxidation number of **oxygen is -2** apart from in compounds with fluorine or in peroxides (-1).

**7)** The oxidation number of **chlorine is -1** apart from in compounds with fluorine or oxygen

**8)**.The oxidation number of all the **group one** and **two** metals is **+1** and **+2** respectively.

**9)** In any combination of two elements, the **more electronegative** one has the **negative** oxidation number.



e.g. H2SO4 KMnO4 Na2Cr2O7 KClO4



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New exercise ..... ..... ...... .....



Assign oxidation numbers to:-

|  |  |  |  |
| --- | --- | --- | --- |
| Ba in BaCl2 | +2 | P in PO43- | +5 |
| Li in Li2O | +1 | Cr in Cr2O72- | +6 |
| P in P2O5 | +5 | O in H2O2 | -1 |
| P in P2O3 | +3 | Mn in MnO2 | +4 |
| C in CO | +2 | Xe in XeF4 | +4 |
| I in I2 | 0 | N in NH3 | -3 |
| C in CCl4 | +4 | N in N2O4 | +4 |
| I in I- | -1 | H in LiH | -1 |
| Cr in CrO4-2 | +6 | S in Na2S4O6 | +2.5 |

What is the problem with the oxidation number of sulfur in Na2S4O6?

S has an oxidation number of +2.5, cannot have ½ electron! But this method works for electron counting purposes

This is due to the sulphur atoms not all being equivalent in the bonding arrangement. In this situation it is acceptable to quote non whole number oxidation numbers.

**Redox Reactions**

We can then classify reactions as redox or not by looking to see if there is a change in the oxidation numbers of any of the species. If there is a **change in the oxidation numbers** then a redox reaction has occurred.

If the **oxidation number** becomes **more positive (increases)**, the element has been **oxidised**

If the **oxidation number** becomes **less positive (decreases)**, the element has been **reduced**

e.g.

**1)** 2Na + 2HCl **🡪**  2NaCl + H2

To see if this is a redox reaction we need to look at the oxidation numbers in the reactants and products.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  | Reactants | | | Products | | |
| Element | Na | H | Cl | Na | H | Cl |
| Oxidation no. | 0 | +1 | -1 | +1 | 0 | -1 |

We can see from the table above that Na has changed from 0 to +1, it has lost an electron and oxidation number has become more positive therefore it has been **oxidised**.

H has changed from +1 to 0, it has gained an electron and oxidation number has become less positive therefore it has been **reduced**.

Therefore this is a **redox reaction**

New exercise**2)** CuCl2  + Zn **🡪**  ZnCl2  + Cu

Assign oxidation numbers to the atoms in the table below.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  | Reactants | | | Products | | |
| Element | Cu | Cl | Zn | Cu | Cl | Zn |
| Oxidation no. | +2 | -1 | 0 | 0 | -1 | +2 |

Which element has been oxidised? zinc

Explain zinc oxidation number has increased from 0 to +2, it has lost two electrons

Which element has been reduced? Copper (II) ions

Explain Cu2+ ions have gained two electrons to form Cu(0), the oxidation number has decreased from +2 to 0

Which is the spectator ion? Cl-

**Half Equations**

We can write redox reactions as full equations where all the ions are shown, or we can write them as ionic equations, leaving out the spectator ions. We can also write half equations for either reduction or oxidation and then combine them to make total redox equations

Class video

See Video Redox 3245 18.20-22.30

**Writing Half equations**

Let us look at some previous examples:-

**Example 1**

Full equation 2Na + 2HCl 🡪 2NaCl + H2

0 +1 -1 +1 -1 0

Write the oxidation numbers below each atom / ion.

The Cl- ion (spectator ion) does not change. i.e. it starts as -1 and ends as -1, therefore we can write the above equation as:-

Full ionic equation 2Na + 2H+ 🡪 2Na+ + H2

We can also split the reaction up into:-

Note the sodium half equation is for **ONE** Na atom.

oxidation half equation:- Na 🡪 Na+ + e-

reduction half equation 2H+ + 2e- 🡪 H2

**Example 2**

Write out the full, ionic and 2 half equations for this reaction of zinc with copper(II) chloride forming zinc chloride and copper. It will help if you write the oxidation numbers under each atom/ion in the full equation.

Full equationCuCl2 + Zn **🡪** ZnCl2 + Cu

Oxidation numbers: +2 -1 0 +2 -1 0

Spectator ion(s): 2Cl-

Ionic equation (leaving out spectator ions)

Zn(s) + Cu2+ 🡪 Zn2+(aq) +Cu(s)

Oxidation half equation

Zn(s) 🡪 Zn2+(aq) + 2e-

Reduction half equation

Cu2+ + 2e- 🡪 Cu(s)

WorksheetPossible homework at this stage:-

QuestionsWorksheet [Oxidation and reduction practice](file:///F:\2009%20Packs\redox%20worksheets\Oxidation%20and%20Reduction%20Practice.doc)

Facer page 194 questions 1, 2, 5, 6, and 7.**How to write and use half equations to write an overall redox ionic equation**

1. Identify the atoms that change using oxidation numbers, write out their oxidation numbers above or below each atom.
2. Decide which substance has been oxidised and which reduced – Remember OIL RIG
3. Start to write an **oxidation half equation** by adding the correct number of electrons to account for the change of oxidation number
4. Balance hydrogen atoms with H+ and oxygen atoms with H2O
5. Check charges and atoms are equal either side.

This gives the **balanced** **half equation** for oxidation.

1. Repeat steps 3 to 5 to write a balanced half equation for **reduction**.
2. Multiply up one or both ½ equations so that the electrons are equal.
3. Add equations, cancelling electrons and any H+ / H2O which appear on both sides of the combined equation.

New exerciseThis results in a **Total redox equation**

**Examples**

**1)** **The reaction of chlorine with bromide ions** in seawater is used in the commercial production of bromine. Write the 2 half equations and the overall redox equation:-

Oxidation: bromide is oxidised to bromine

Oxidation numbers: -1 0

**Oxidation half equation**

(balanced with electrons) ……………………………………...........................…………...……

Do the atoms balance? 2Br- 🡪 Br2 + 2e-

Do the atoms balance? Yes

Therefore the chlorine must have been reduced to chloride ions.

Oxidation nos: 0 -1

**Reduction half equation**

(balanced with electrons) Cl2 + 2e- 🡪 2Cl-

As the number of electrons are equal in both we just add the two half equations together to give:-

**Overall redox equation**

Cl2 + 2Br- 🡪 Br2 + 2Cl-

**2)** **The reaction of iron with chlorine to give iron(III) chloride**

Write the 2 half equations and the overall redox equation. (Remember to balance electrons)

**Oxidation half equation**

Fe 🡪 Fe3+ + 3e- X 2

**Reduction half equation**

Cl2 + 2e- 🡪 2Cl- X 3

**Total redox equation** (multiply each half equation above so electrons balance then combine them)

2Fe + 3Cl2 + 6e- 🡪 2Fe3+ + 6e- + 6Cl-

2Fe + 3Cl2 🡪 2Fe3+ + 6Cl-

*The next 3 examples involve ‘oxo’ ions so are more complicated as the oxygen must be balanced*

**3)** **MnO4- oxidises FeCl2 in an acid solution to produce FeCl3 and Mn2+**

**Oxidation ½ equation** (remember to add the correct number of electrons)

+2 +3

Fe2+ 🡪 Fe3+ + e- X 5

**Reduction ½ equation**

* Write down reactant and product ions and work out the oxidation numbers of the Mn
* Add the correct number of electrons to account for the change in oxidation number.
* Balance the oxygen atoms in MnO4- by adding H2O’s to the right hand side of the equation
* Balance the hydrogen atoms by adding H+ ions to the left hand side of the equation.

+7 +2

MnO4- + 8H+ + 5e- 🡪 Mn2+ + 4H2O

**Total redox equation** (multiply oxidation half equation to make sure the electrons balance)

5Fe2+ + MnO4- + 8H+ + 5e- 🡪 5Fe3+ + 5e- + Mn2+ + 4H2O

Cancelling electrons

5Fe2+ + MnO4- + 8H+ 🡪 5Fe3+ + Mn2+ + 4H2O

*Using a similar method, complete the following:-*

**4)** **Cr2O72- oxidises I- in an acid solution to produce I2 and Cr3+**

**Reduction ½ equation**

+6 +3

Cr2O72- + 14H+ + 6e- 🡪 2Cr3+ + 7H2O

**Oxidation ½ equation**

-1 0

2I- 🡪 I2 + 2e-. X 3

**Total redox reaction**

Cr2O72- + 14H+ + ~~6e~~~~-~~ + 6I- 🡪 2Cr3+ + 7H2O + 3I2 + ~~6e~~~~-~~

Cancelling electrons

Cr2O72- + 14H+ + 6I- 🡪 2Cr3+ + 7H2O + 3I2

**5)** **MnO4- in an acid solution oxidises nitrate(III) to nitrate(V) and is reduced to Mn2+.**

**Reduction ½ equation**

+7 +2

MnO4- + 8H+ + 5e- 🡪 Mn2+ + 4H2O X 2

**Oxidation ½ equation**

+3 +5

NO2- + H2O 🡪 NO3- + 2H+ + 2e- X 5

**Total redox reaction**

6 3

2MnO4- + 16H+ + 10e- + 5NO2- + 5H2O 🡪 2Mn2+ + 8H2O + 5NO3- + 10H+ + 10e-

Cancelling electrons, water and protons

2MnO4- + 6H+ + 5NO2- 🡪 2Mn2+ + 3H2O + 5NO3-

**More practice.** Balance the half equations for each of these reactions (you will need to assign oxidation numbers and use the method explained on the previous pages)

**New exercise**

**Reduction half equations**

-1 -2

1) H2O2 + 2H+ + 2e- 🡪 2H2O

+7 +2

2) MnO4- + 8H+ + 5e- 🡪 Mn2+ + 4H2O

+6 +3

3) Cr2O72- + 14H+ + 6e- 🡪 2Cr3+ + 7H2O

+5 0

4) IO3- + 6H+ + 5e- 🡪 ½I2 + 3H2O

0 -1

5) Br2 + 2e- 🡪 2Br-

0 -1

6) I2 + 2e- 🡪 2I-

**Oxidation half equations**

+2 +3

1) Fe2+ 🡪 Fe3+ + e-

+2 +2.5

2) 2S2O32- 🡪 S4O62- + 2e-

+4 +6

3) SO32- + H2O 🡪 SO42- + 2H+ + 2e-

-1 0

4) 2I- 🡪 I2 + 2e-

+3 +5

5) NO2- + H2O 🡪 NO3- + 2H+ + 2e-

-1 0

6) H2O2 🡪 O2 + 2H+ + 2e-

**Full redox reactions.**  Combine the relevant half equations above to give full ionic redox equations. Remember to multiply up as necessary so electrons cancel.

1) MnO4- + 5Fe2+ + 8H+ 🡪 Mn2+ + 5Fe3+ + 4H2O

2) Cr2O72- + 3SO32- + 8H+ 🡪 2Cr3+ + 3SO42- + 4H2O

3) 2IO3- + 10I- + 12H+ 🡪 6I2 + 6H2O

4) Br2 + 2S2O32- 🡪 S4O62- + 2Br-

5) H2O2 + 2Fe2+ + 2H+ 🡪 2Fe3+ + 2H2O

6) 2MnO4- + 5H2O2 6H+ 🡪 2Mn2+ + 5O2 + 8H2O

WorksheetSee Worksheets [Oxidation states and redox equations](file:///F:\2009%20Packs\redox%20worksheets\redox%20reactions%20and%20ox%20nos.doc) OR [oxidation states 2](file:///F:\2009%20Packs\redox%20worksheets\Oxidation%20states.doc)

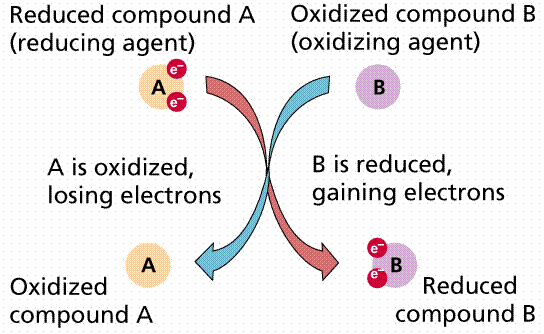
and [Redox or not?](file:///F:\2009%20Packs\redox%20worksheets\Redox%20reactions%20or%20not.doc)

**Identifying oxidising and reducing agents in complicated redox reactions**

An oxidising agent oxidises another atom, **it accepts electrons** so the oxidising agent is reduced.

A reducing agent reduces another atom, **it donates electrons**  so the reducing agent is oxidised

Oxidising and reducing agents can be atoms, molecules or ions. However, when identifying **what is being oxidised or reduced**, state the **element which is changing oxidation number.**



For example CuSO4 + Zn **🡪** Cu + ZnSO4

If we split this up into the 2 half equations we get

Zn **🡪** Zn2+ + 2e- Oxidation reaction

Cu2+ + 2e- **🡪** Cu Reduction reaction

The **Zn** has been oxidised so it is the **reducing agent** as it has caused the copper to be reduced from Cu2+ to Cu (oxidation number has changed from +2 to 0).

New exerciseThe **Cu2+** has been reduced so it is the **oxidising agent** as it has caused the zinc to be oxidised from Zn to Zn2+ (oxidation number has changed from 0 to +2).

Then identify the reducing and oxidising agents in this table

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Equation  (write oxidation numbers above atoms) | Element oxidised | Element reduced | Oxidising agent | Reducing agent |
| 1 | 0 +1 +2 0  Mg + 2 HCl 🡪 MgCl2 + H2 | Mg | H | HCl | Mg |
| 2 | +4 0 +6 -2  2 SO2 + O2 🡪 2 SO3 | S in SO2 | O2 | O2 | SO2 |
| 3 | 0 -2 -1 0  2Cl2 + 2 H2O 🡪 4 HCl + O2 | O in H2O | Cl2 | Cl2 | H2O |
| 4 | +7 +2 +2 +3  MnO4- + H+ + 5Fe2+ 🡪 Mn2+ + 4 H2O + 5Fe3+ | Fe2+ | Mn in MnO4- | MnO4- | Fe2+ |
| 5 | -3 +6 +3 0  (NH4)2Cr2O7 🡪 Cr2O3 + N2 + 4 H2O | N in NH4+ | Cr in Cr2O72- | Cr2O72- | NH4+ |
| 6 | -1 -2 0  2H2O2 🡪 2H2O + O2 | O in H2O2 | O in H2O2 | H2O2 | H2O2 |

What do you notice about the oxidising and reducing agents in equation 6? They are both the same

This is an example of a disproportionation reaction**Disproportionation**

Disproportionation is the simultaneous oxidation and reduction of atoms of the same element

For example:-

2 H2O2 🡪 2 H2O + O2

Just as before, work out:-

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Reactants | | Products | | |
| Element | H | O | H | O in O2 | O in H2O |
| Oxidation number | +1 | -1 | +1 | 0 | -2 |

The H atom does not change i.e. it is a spectator ‘ion’, but the oxygen goes from -1 to -2 **and** from -1 to 0(i.e. the element). The oxygen atom has been simultaneously oxidised and reduced. Therefore this is a disproportionation reaction.

Example 2)

**Cl2 + 2NaOH 🡪 NaCl + NaOCl + H2O**

Write out the oxidation numbers above or below each atom.

Is this a disproportionation reaction? Yes

Explain your answer using oxidation numbers Chlorine (0) has been oxidised to OCl- (+1) losing an electron. Chlorine (0) has been reduced to Cl- (-1) gaining an electron. Simultaneous oxidation and reduction of atoms of a species is disproportionation.

**Potassium manganate(VII) and glycerol** (Classic Chemistry Demonstrations 29)

Manganese can exist in a variety of oxidation states each oxidation state has a characteristic colour associated with it (also shown on your coloured Periodic Table):-

|  |  |  |
| --- | --- | --- |
| Oxidation state | Species | Colour |
| +7 | MnO4- | purple |
| +6 | MnO42- | green |
| +4 | MnO2 | Demonstration**Demonstration**brown |
| +2 | Mn2+ | very pale pink |

We can follow a reaction using the associated colour changes

Your teacher will demonstrate adding glycerol to KMnO4. Remember potassium ions produce a purple flame colour.

What colour is the manganate compound at the start of the experiment? purple

What colour are the manganese compounds when the residue is dissolved in water at the end of the experiment? green

WorksheetQuestionsWhat has happened to the oxidation number of the manganese? Gone down = reduction

What type of reaction is this? redox

**More practice:-** Facer page 194 questions 3, 4, 8, 9, 10, 11 and 12

Worksheets ‘Oxidising and Reducing Agents’ and ‘More work on Redox’

**Redox reactions** (Micro-scale chemistry 6 p26-27)

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In these experiments you will observe and interpret redox reactions

**1) Copper(II) ions and halide ions**

**Method**

* Cover the table below with clear plastic sheet
* Put 1 drop of copper(II) sulfate solution in each box
* Add 1 drop of chloride to the 1st box, bromide to the 2nd, and iodide to the 3rd
* Add one drop of starch to each of the reactant mixtures
* Carefully slide the sheet away and note your observations in the table

Note starch turns blue black in the presence of iodine

|  |  |  |  |
| --- | --- | --- | --- |
|  | Chloride ions Cl-(aq) | Bromide ions Br-(aq) | Iodide ions I-(aq) |
| Copper(II) sulphate | solution remained pale blue | solution remained pale blue | solution turned brown. When starch added it turned blue/black |
| Initial colour of Cu2+  pale blue |

Which of the halides reacted? iodide

What happened when the starch was added? Turned blue/black so I2 present

Write the half equations for

* the reaction of the iodide ions 2I- 🡪 I2 + 2e-
* the reaction of copper ions 2Cu2+ + 2e- 🡪 2Cu+ divide by 2:-

Cu2+ + e- 🡪 Cu+

**2) Iron Chemistry** (Micro scale chemistry 12 p35)

You will compare the chemistry of the two main oxidation states of iron

**Method**

* Cover the page with a clear plastic sheet
* Put 1 drop of iron(II) solution in each box in the 2nd row
* Put 1 drop of iron(III) solution in each box of the 3rd row
* Add 2 drops of NaOH to the first column and observe over 10 min
* Add 1 drop of potassium thiocyanate solution to the second column
* Add 1 drop of iodide and after 1 min 1 drop of starch to the third column
* Add 1 drop of KMnO4 to the fourth column and observe over 10 min
* Add 1 drop silver nitrate to the fifth column and observe closely
* Carefully slide the sheet away and note your observations in the table. **Give the formula of any precipitate formed.**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Solutions of  (ions) | Hydroxide  OH- | Thiocyanate  SCN- | Iodide  I- | Manganate  MnO4- | Silver  Ag+ |
| Iron(II) ions  Fe2+  Colour:  .................. | pale green ppt.  starts to turn brown with time  Fe(OH)2(s) | No change  (turned slightly pink with time) | No change | pale pink/slight brown ppt  REDOX | sparkly particles appeared  Ag(s)  REDOX |
| Iron(III) ions  Fe3+  Colour:  .................. | Red/brown ppt.  Fe(OH)3(s) | Blood red solution  Fe(SCN)3(s) | orange solution I2  turned black with starch  REDOX | No change | No change |

**Reactions of iron(II)**

Which ions reacted with iron(II)? OH- MnO4- and Ag+

Using your understanding of redox explain these reactions.

**Identification:-**

Fe2+(aq) + 2OH-(aq) 🡪 Fe(OH)2(s) pale green ppt indicates Fe2+ (oxidised by air to Fe3+)

**Redox**

Fe2+(pale green) is oxidised by MnO4-(purple) to form Fe3+(yellow) and Mn2+(v pale pink)

Fe2+(pale green) is oxidised by Ag+(colourless) to form Fe3+(yellow) and Ag(s) (sparkly bits!)

**Reactions of Iron (III)**

Which ions reacted with iron(III)? OH- SCN- and I-

Using your understanding of redox explain these reactions.

**Identification:-**

Fe3+(aq) + 3OH-(aq) 🡪 Fe(OH)3(s) red/brown ppt indicates Fe3+

Fe3+(aq) + 3SCN-(aq) 🡪 Fe(SCN)3(s) blood red solution indicates Fe3+

**Redox**

Fe3+(yellow) is reduced by I-(colourless) to form Fe2+(pale green) and I2(brown)

2Fe3+(aq) + 2I- 🡪 2Fe2+(aq) + I2(aq)

**Revision Page**

**Multiple choice questions.**

**1.** What is the oxidation number of chlorine in the ClO3– ion?

**A** –1

**B** +4

**C** +5

**D** +6

(Total 1 mark)

**2.** Which of these reactions is **not** a redox reaction?

**A** Mg(NO3)2(s) → MgO(s) + 2NO2(g) + ½O2(g)

**B** HCl(aq) + NaOH(aq) → NaCl(aq) + H2O(l)

**C** Fe(s) + CuSO4(aq) → FeSO4(aq) + Cu(s)

**D** Cl2(aq) + 2Br–(aq) → 2Cl–(aq) + Br2(aq)

(Total 1 mark)

**3.** What is the oxidation number of **sulfur** in sodium tetrathionate, Na2S4O6?

**A** –½

**B** +1½

**C** +2½

**D** +5

**5.** Chemical reactions may involve

**A** oxidation

**B** reduction

**C** no change in oxidation number

**D** disproportionation

Which of the terms above best describes what happens to the **chlorine** in the following reactions?

(a) Cl2(g) + H2O(l) → HCl(aq) + HOCl(aq)

**A**

**B**

**C**

**D**

(1)

(b) Cl2(g) + 2Na(s) → 2NaCl(s)

**A**

**B**

**C**

**D**

(1)

(c) NaCl(s) + H2SO4(l) → HCl(g) + NaHSO4(s)

**A**

**B**

**C**

**D**

(1)

(Total 3 marks)

**Structured Questions**

**1.** Bromine is extracted from seawater using chlorine.

(a) (i) Write the equation for the reaction of chlorine with sodium bromide solution.  
Do **not** include state symbols.

Cl2 + 2NaBr → Br2 + 2NaCl

(1)

(ii) The seawater is acidified before the reaction with chlorine to prevent the bromine produced reacting with the water.

Br2 + H2O  HBr + HOBr

Name the type of reaction taking place between bromine and water. Explain your answer in terms of the changes in oxidation number of bromine.

**Type of reaction** Disproportionation **(1)**

**Explanation** Bromine is simultaneously oxidised from 0 to +1 and reduced from 0 to –1 **(1)**

(iii) Bromine vapour reacts with sulphur dioxide and water as follows.

Br2 + SO2 + 2H2O → 2HBr + H2SO4

State the oxidation number of sulphur in

SO2 +4

H2SO4 .+6

(2)

(iv) Use the data from (iii) to show that bromine is acting as an oxidising agent.

(1)

The oxidation number of S is increasing (so bromine is acting as an oxidising agent)  
Or

Oxidation number of Br is decreasing so it must be acting as an oxidising agent

(b) The ionic half-equation for the reduction of iodate(V) ions, IO3–, to iodine in acid solution is

2IO3– + 12H+ + 10e– → I2 + 6H2O

(i) Write the ionic half-equation for the oxidation of SO2 in water to SO42– and H+ ions. Do **not** include state symbols.

(1)

SO2 + 2H2O → SO42– + 4H+ +2e(–)

(ii) Combine the reduction reaction of iodate(V) ions, IO3–, with the oxidation reaction of SO2 to give the full ionic equation for the reaction of IO3– with SO2.  
Do **not** include state symbols

**Correct balanced equation  
2 IO3– + 5 SO2 + 4H2O → I2 + 5 SO42– + 8H+** **(2)**

(2)

(Total 9 marks)

**2.** (a) Explain the term **reducing agent** in terms of oxidation number change.

Substance that can lower/reduce the oxidation number (of an element in another substance)

(1)

(b) Write ionic half-equations (do **not** include state symbols) to show:

(i) chlorate(I) ions, ClO–, in **acidic** solution, being reduced to chlorine molecules and water.

2ClO– + 4H+ + 2e(–) → Cl2 + 2H2O

(1)

(ii) chloride ions being oxidised to chlorine molecules.

2Cl– → Cl2 + 2e(–)

(1)

(c) Combine the two equations in (b) to show the effect of adding an acid to a mixture of chlorate(I) ions and chloride ions.

ClO– + Cl– + 2H+ → Cl2 + H2O

(1)

(d) Potassium chlorate, KClO3, decomposes on heating to give potassium chloride, KCl, and oxygen, O2.

(i) Write the equation for this reaction. State symbols are **not** required.

2KClO­ → 2KCl + 3O2

(1)

(ii) Show, by the use of **oxidation numbers**, why this is a redox reaction.

Oxidation numbers all correct **(1)** Cl O  
Start +5 –2  
End –1 0

Chlorine reduced as oxidation number decreases/ changes  
from +5 to –1 **(1)**

Oxygen oxidised as oxidation number increases/changes from  
–2 to 0 **(1)**

(3)

(Total 8 marks)

**3.** This question is about the manufacture of bromine from bromide ions found in seawater.

(a) In the first step, chlorine gas is bubbled into acidified seawater. This converts the bromide ions to bromine. The low pH prevents hydrolysis of the liberated bromine.

(i) Complete and balance the equation for the hydrolysis of bromine with water which is a disproportionation reaction.

Br2(aq) + H2O(l)  2H+(aq) + Br–(aq) + BrO–(aq)

(2)

(iii) Explain, using oxidation numbers, why this reaction is known as disproportionation.

(2)

bromine has been both oxidised and reduced **(1)**  
must mention bromine (Br/Br2 for first mark)  
from 0 to +1 and –1 **(1)**

(iv) Write the ionic equation, including state symbols, for the reaction of chlorine gas with bromide ions.

Cl2(g) + 2Br–(aq) → 2Cl–(aq) + Br2(aq)

(2)

(b) In the second step, air is blown through the reaction mixture to remove the bromine as a vapour which is then mixed with sulphur dioxide gas and water vapour.

The unbalanced equation for this reaction is Br2 + H2O + SO2 → H+ + Br− + SO42–

(i) Identify the elements which are oxidised and reduced and give their oxidation numbers.

Element **oxidised** Sulfur

Oxidation number initial +4 . final +6

Element **reduced** Bromine

Oxidation number initial 0 final -1

(2)

(ii) Using this information, or otherwise, balance the equation.

Br2 + 2H2O + SO2 → 4H+ + 2Br– + SO42–

(1)

(Total 9 marks)

**4.** Hydrochloric acid, formed when hydrogen chloride is dissolved in water, can be converted to chlorine using an aqueous solution of hydrogen peroxide:

2HCl(aq) + H2O2(aq)  Cl2(g) + 2H2O(l)

(i) Give the oxidation numbers of

chlorine in HCl -1. chlorine in Cl2 0

oxygen in H2O2 -1 oxygen in H2O -2 (2)

(ii) Name the reducing agent in this reaction.

hydrochloric acid/hydrogen chloride/chloride ion/ HCl/Cl–  
iii) Explain why the oxidation numbers you have given in (i) are consistent with the fact that **two** moles of hydrochloric acid react with **one** mole of hydrogen peroxide.

*N.B. Read whole thing through.  
Look for good use of chemical language eg use of molecule / atom*

O.N. of 2 Cls has increased by (1 × 2=) TWO  
O.N. of 2 Os has decreased by (1 × 2=) TWO

OR

In terms of electron exchange from one **oxidation state** to another.

OR

Can consider total oxidation numbers remaining constant eg “sum of O and Cl  
O.N. = –4 on left, sun of O and Cl O.N. = –4 on right too”. 1

(1)

(Total 4 marks)

**5.** (a) Deduce the oxidation number of iodine in the following species.

(i) I2O7 + 7

(1)

(ii) IO +7

(1)

(b) Iodine, I2, can be reduced to iodide ions, I–, by tin(II) ions, Sn2+, which are themselves oxidised to tin(IV) ions, Sn4+.

(i) Construct the oxidation and reduction half-equations for the above system.

Sn2+  Sn4+ + 2e(–) *OR* Sn2+ – 2e(–)  Sn4+**(1)**

I2 + 2e(–)  21– **(1)**

(2)

(ii) Use the above half-equations to construct the overall ionic equation for the reaction.

Sn2+ + I2  Sn4+ + 2I–

(1)

(Total 5 marks)

**6.** (a) Hydrogen sulphide is produced when concentrated sulphuric acid is added to solid sodium iodide, but sulphur dioxide is produced when concentrated sulphuric acid is added to solid sodium bromide.

(i) Complete the following table:

|  |  |  |
| --- | --- | --- |
| Compound | Formula | Oxidation number of sulphur in compound |
| Sulphuric acid | H2SO4 | +6 |
| Hydrogen sulphide | H2S | -2 |
| Sulphur dioxide | SO2 | +4 |

(3)

(ii) Use your answers to part (a)(i) to suggest which of the ions, iodide or bromide, has the greater reducing power.

Iodide has greater reducing power **(1)** *with some attempt at using answer from part (i)*

Reduces sulphur by more oxidation numbers / or correctly uses their numbers from part (i) / or an  
‘electron gain’ type argument **(1)**

(2)

(b) (i) Write an ionic half-equation to show the oxidation of chloride ions, Cl–, to chlorine, Cl2.

2Cl  Cl2 + 2e–

(1)

(ii) Write an ionic half-equation to show the reduction of chlorate(I) ions, OCl–, to chloride ions, in acidic conditions.

OCl– + 2H+ + 2e  Cl– + H2O

(2)

(iii) Bleach is a solution of chlorate(I) ions and chloride ions. Combine the two ionic half-equations above to produce an equation which shows the effect of adding acid to bleach.

OCl– + Cl- + 2H+ Cl2 + H2O

(1)

(Total 9 marks)

**7.** (a) Define the term **oxidation number**.

formal charge an atom would have in a compound if bonding was 100% ionic

(2)

(b) The equation below shows the disproportionation of chlorine.

Cl2(g) + H2O(l)  HClO(aq) + HCl(aq)

0 +1 -1.

(i) Underneath the chlorine-containing species write the oxidation number of chlorine in each case.

(1)

(ii) Use these oxidation numbers to explain the term **disproportionation**

One of the chlorines in each molecule (0 to +1) has lost  
an electron / been oxidised 1

The other chlorine in the chlorine molecule has  
gained an electron / been reduced to –1 1

a simple statement / definition of disproportionation max 1

.

(2)

(c) Explain why hydrogen chloride forms an acidic solution when dissolved in water.

reacts / changes / dissociates / ionises / HCl donates a proton to the water (1)

donating/releasing a H+ (aq) or H3O+ (aq) ions(1)

(Total 7 marks)

2015

NAME ...........................……... HOMEWORK DEADLINE .....................

Student Number ………… Chemistry Class ………

Student targets from **previous pack**

Redox

|  |  |
| --- | --- |
| **Task** | Mark |
| Notes | /10 |
| Revision Notes /Summary page | /10 |
| Exam questions | /57 |
| Overall Grade for this work | A B C D E U |

Student comments

Tutor comments

Tutor signature Date

Student targets for **next pack**