Practise in writing oxidation and reduction ½ equations.

1. Use **page 7 of the data book** to complete the ion-electron ½ equations below by putting the **electrons** on the correct side and state which are **oxidation** and which are **reduction.** The first one has been done for you.

a)	Br_2	+ $2e^-$	$\longrightarrow$	$2\,Br^-$	**reduction**
b)	Fe^{2+}		$\longrightarrow$	Fe^{3+}	_____
c)	Mg		$\longrightarrow$	Mg^{2+}	_____
d)	Cu^{2+}		$\longrightarrow$	Cu	_____
e)	$2I^-$		$\longrightarrow$	I_2	_____
f)	$2H^+$		$\longrightarrow$	H_2	_____

2. In an ion-electron equation, which side are the electrons on?

i) an **oxidation equation -** _____ hand side**.**

ii) a **reduction equation -** _____ hand side**.**

3. State which of the equations below is an **oxidation** and which shows **reduction.**

i) $SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2e^-$ _____

ii) $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4\,H_2O$ _____

Constructing REDOX equations

Redox equations show both the oxidation and reduction equations together.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \quad \text{oxidation}$$

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s) \quad \text{reduction}$$

When combining the above equations to form a REDOX equation then the **electrons can be cancelled out** as they will appear on both sides of a combined equation. So the combined **REDOX** equation is:-

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

The above equation is straightforward because the number of electrons in both ion-electron equations is the same. Some equations are not so easy.

Before you construct a redox equation you must **equalise electrons** and cancel i.e. make sure the **number of electrons are the same** in each of the **reduction and oxidation equations.** To do this you may have to multiply one or both of the ½ equations by a whole number.

e.g.1 zinc displacing silver(I) ions from solution to form silver and zinc(II) ions.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \quad \text{oxidation}$$

$$Ag^+(aq) + e^- \longrightarrow Ag(s) \quad \text{reduction} \quad \times 2$$

The reduction equation above only has 1 e^- so the whole equation has to be multiplied by 2 to achieve the same number of electrons as the oxidation ½ equation.

$$2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$$

Now we can cancel out the electrons and construct the **REDOX** equation.

$$Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s) \quad \text{REDOX}$$

e.g.2 aluminium displacing silver(I) ions from solution to form silver and aluminium ions.

$$Al(s) \longrightarrow Al^{3+}(aq) + 3e^- \quad \textbf{oxidation}$$

$$Ag^+(aq) + e^- \longrightarrow Ag(s) \quad \textbf{reduction} \quad X\ 3$$

REDOX equation:-

$$3Ag^+(aq) + Al(s) \longrightarrow 3Ag(s) + Al^{3+}(aq)$$

Now try the following examples on your own using page 7 in the data book:-

Remember to **equalise and cancel electrons** before writing the **REDOX** equation.

1) magnesium displacing silver(I) ions from solution to form silver and magnesium ions.

_____ ⟶ _____ + ____ **oxidation**

_____ + ____ ⟶ _____ **reduction**

REDOX equation:-

_____ + _____ ⟶ _____ + _____

2) aluminium displacing copper(II) ions from solution to form copper and aluminium ions.

_____ ⟶ _____ + ____ **oxidation**

_____ + ____ ⟶ _____ **reduction**

REDOX equation:-

_____ + _____ ⟶ _____ + _____

And now for one **without a metal.**

3) permanganate ions (**MnO_4^-**) reacting with bromide ions(**Br^-**) to form manganese(II) (**Mn^{2+}**) ions and bromine **(Br_2)**

_____ ⟶ _____ + ____ **oxidation**

_____ + ____ + _____ ⟶ _____ + _____ **reduction**

REDOX equation:

_____ + _____ + _____ ⟶ _____ + _____ + _____