

REDOX REACTIONS

Lecture 12



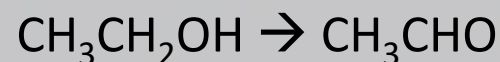
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RED-OX REACTIONS

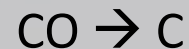
OXIDATION:

substance gains oxygen
substance loses hydrogen
substance loses electron(s)



REDUCTION:

substance loses oxygen
substance gains hydrogen
substance gains electron(s)



REDOX reactions always occur simultaneously

gunpowder: KNO_3 , C, S

NO_3^- reduced to NO and NO_2

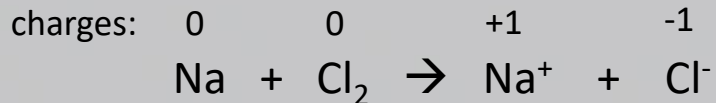
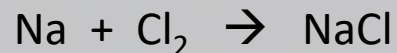
C oxidised to CO and CO_2

S oxidised to SO_2 and SO_3



REDOX REACTIONS OF IONIC COMPOUNDS

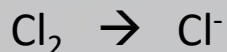
1. Recognise the RED and OX processes:



charge on Na increased

Na loses electron

OXIDATION

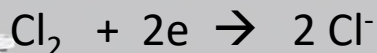
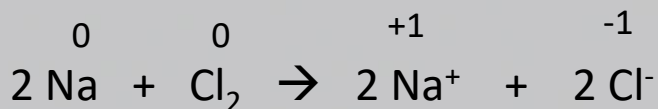


charge on Cl decreased

Cl gains electron

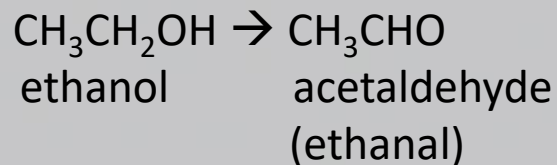
REDUCTION

2. Balance the atoms and charges: conservation of MASS and conservation of CHARGES



REDOX REACTIONS OF ORGANIC COMPOUNDS

For the covalent species:



OXIDATION NUMBER → 'FAKE CHARGE'

a **number** assigned to a compound or element to show how many electrons were lost or gained

The oxidation state of atoms in a pure element is 0

In an uncharged compound, the sum of all the oxidation states is zero

Average oxidation number for an atom

1. Each type of atom in a species gets an oxidation number
2. Total oxidation number = charge on the species
3. If still not able to assign numbers, follow the rules:



ASSIGNING OXIDATION NUMBERS

1. Metal cations: oxidation number = charge on the cation

group 1 metals	+1
group 2 metals	+2
group 3 metals	+3 (Al, Ga)
etc	

2. F -1 (always)

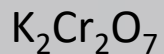
3. H +1 (except metal hydrides, -1)

4. O -2 (except O_2^{2-} , O_2^-)

5. If you still cannot assign all numbers, then the best practice is to assign an appropriate negative charge to the unassigned atom with the highest electronegativity



EXAMPLES



Overall charge on this molecule is 0

K is +1, O is -2

$$(2 \times 1) + (2 \times \text{Cr}) + (7 \times -2) = 0$$

$$\text{Cr} = +6$$



Overall charge on this ion is -2

O is -2

$$(1 \times \text{S}) + (4 \times -2) = -2$$

$$\text{S} = +6$$

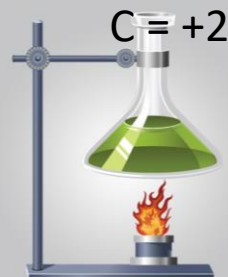
CN^- , cyanide

Overall charge on this ion is -1

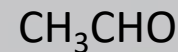
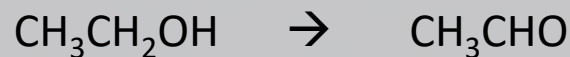
N is closer to F and would make -3 ion

$$(1 \times \text{C}) + (1 \times -3) = -1$$

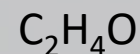
$$\text{C} = +2$$



REDOX REACTIONS OF ORGANIC COMPOUNDS



For easier work rewrite: $\text{C}_2\text{H}_6\text{O}$



Overall charge on this molecule is 0

H is +1, O is -2

$$(2 \times \text{C}) + (6 \times 1) + (1 \times -2) = 0$$

$$\text{C} = -2$$

Overall charge on this molecule is 0

H is +1, O is -2

$$(2 \times \text{C}) + (4 \times 1) + (1 \times -2) = 0$$

$$\text{C} = -1$$

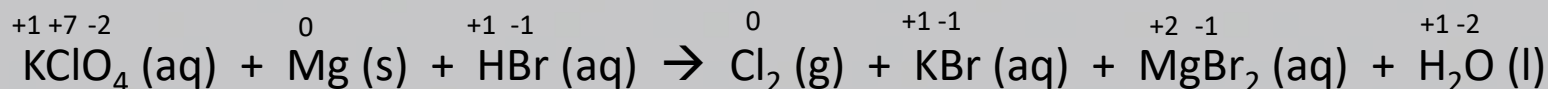
carbon lost electron

OXIDATION



BALANCING REDOX EQUATIONS

1. Assign oxidation numbers:

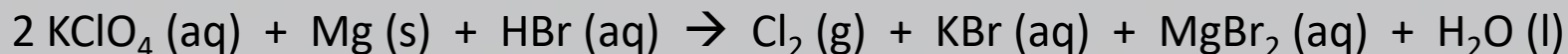


2. Identify atoms that are reduced and oxidised:

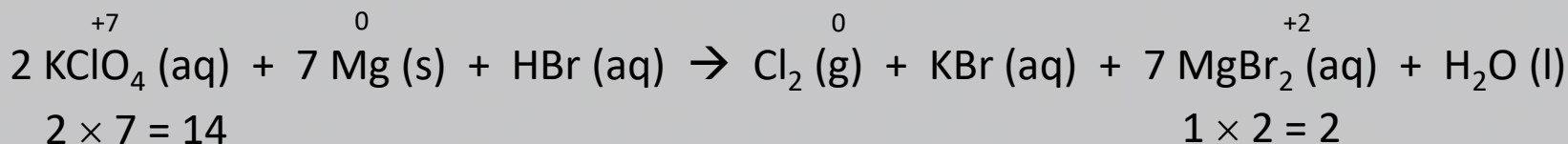
Cl	+7	→ 0	Reduced
Mg	0	→ +2	Oxidized

3. Balance electron transfer

a) balance 'redox-ed' atoms using correct coefficients on whole species

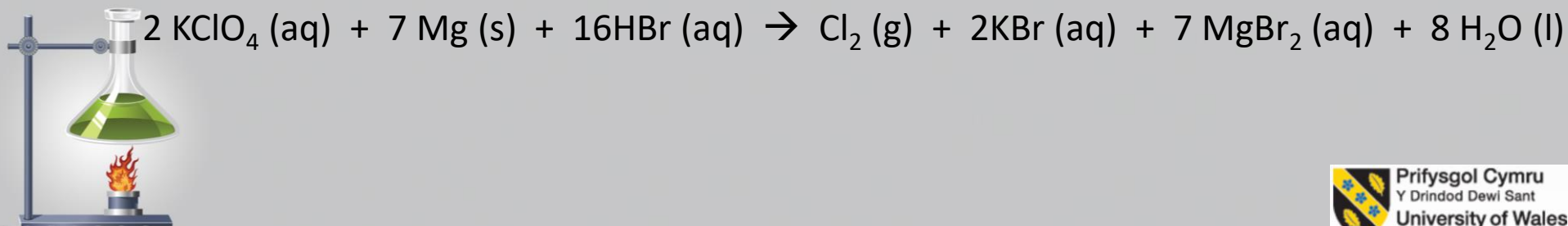


b) balance electrons using correct coefficients on whole species



BALANCED 'IMPORTANT ATOMS' (Mg and Cl) AND ELECTRONS

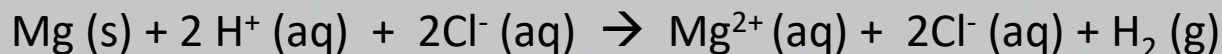
4. Balance remaining atoms



HALF EQUATIONS



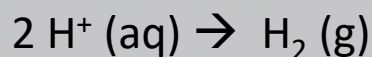
1. Write the equation in ionic form



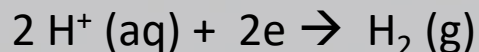
2. Remove spectator ions (Cl^-)



Reduction:



gains electrons



reactants

Oxidation:



loses electrons



products

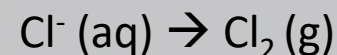
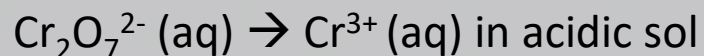


MORE COMPLEX HALF-EQUATIONS

	$\text{MnO}_4^- (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq})$ in acidic sol	$\text{MnO}_4^- (\text{aq}) \rightarrow \text{MnO}_2 (\text{s})$ in basic sol
Ox number	+7 +2	+7 +4
Red or Ox	Reduction, e - reactants	Reduction, e - reactants
No Imp Atoms	1 manganese	1 manganese
No electrons	$7 - 2 = 5$ electrons	$7 - 4 = 3$ electrons
Charge:	$\text{MnO}_4^- (\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+} (\text{aq})$ -6 +2	$\text{MnO}_4^- (\text{aq}) + 3\text{e}^- \rightarrow \text{MnO}_2 (\text{s})$ -4 0
# O atoms	$\text{MnO}_4^- (\text{aq}) + 8 \text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} (\text{aq})$ 4 0	$\text{MnO}_4^- (\text{aq}) + 3\text{e}^- \rightarrow \text{MnO}_2 (\text{s}) + 4 \text{OH}^-$ 4 6
Check H	$\text{MnO}_4^- (\text{aq}) + 8 \text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} (\text{aq}) + 4 \text{H}_2\text{O}$	$2 \text{H}_2\text{O} + \text{MnO}_4^- (\text{aq}) + 3\text{e}^- \rightarrow \text{MnO}_2 (\text{s}) + 4 \text{OH}^-$



MORE COMPLEX HALF-EQUATIONS



Ox number

+6

+3

-1

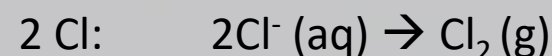
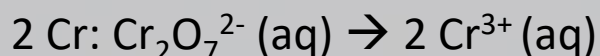
0

Red or Ox

Reduction, e - reactants

oxidation, e - products

No Imp Atoms

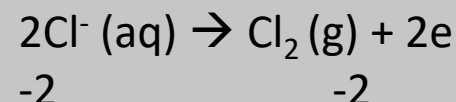
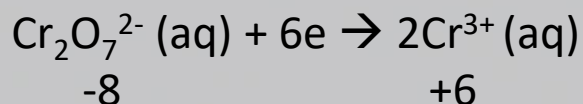


No electrons

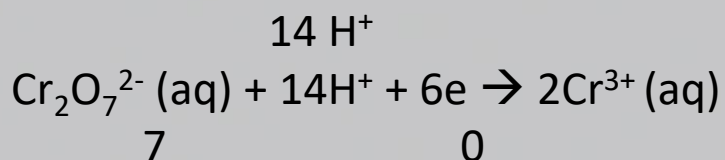
12-6=6 electrons

-2-0= 2 electrons

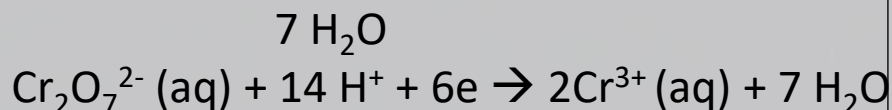
Charge:



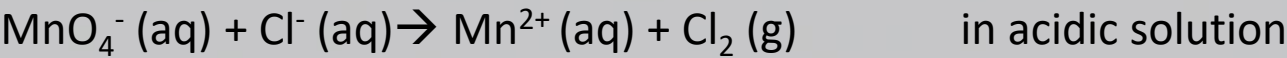
0 atoms



Check H



BALANCING EQUATIONS



Ox number +7 +2

Half-equation $\text{MnO}_4^- (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq})$ $\text{Cl}^- (\text{aq}) \rightarrow \text{Cl}_2 (\text{g})$

Red or Ox Reduction, e - reactants oxidation, e - products

No Imp Atoms 1 Mn 2 Cl $2\text{Cl}^- (\text{aq}) \rightarrow \text{Cl}_2 (\text{g})$

No electrons $7 - 2 = 5$ electrons $-2 - 0 = 2$ electrons



Combine with balanced e's $2\text{MnO}_4^- (\text{aq}) + 10\text{Cl}^- (\text{aq}) \rightarrow 2\text{Mn}^{2+} (\text{aq}) + 5\text{Cl}_2 (\text{g})$

Charge -12 +4
 16 H⁺

